International Junior Science Olympiad (IJSO)

TARGET IJSO

Contents

CHEMICAL REACTIONS & EQUATIONS

INTRODUCTION

Change is the law of nature. We observe various types of changes around us. Plants grow into trees, a child grows into an adult, fruits ripen, water evaporates, water freezes in refrigerator, mercury rises in a thermometer on a hot day, iron articles rust in moist air, milk changes into curd, oil burns in stoves, sometimes a glass tumbler breaks, etc. Scientists classify these changes as physical and chemical changes.

PHYSICAL CHANGE

A physical change is one in which the substance undergoing the change is not destroyed. That is, the substance does not lose its identity. **In a physical change, no new element or no new compound is formed**. When the source responsible for the physical change is removed, the substance regains its original state.

Examples of Physical Changes

(i) Evaporation of a liquid

(ii) Sublimation

(iii) Melting of a solid

(iv) Dissolution of sugar or salt in water

(v) Powdering of sugar

(vi) Mixing of iron particles and sand

CHEMICAL CHANGE

A chemical change is one in which the identity of the original substance changes and a new substance or new substances are formed. In a chemical change, the properties of the substances before and after the change are entirely different.

Examples of Chemical Changes

(i) Ripening of fruits (ii) Digestion of food

(iii) Souring of milk

(iv) Rusting of iron

(v) Burning of a candle

(vi) combustion of glucose in metabolic process

(vii) Preparation of water from hydrogen and oxygen

(viii) Photosynthesis by plants

CHEMICAL REACTION

Chemical reaction is a process in which some known substances are changed into new substance or new substances. The properties of new substances are different from the properties of the starting substances.

In the language of chemistry,

The starting substances are called reactants.

The new substances formed are called products.

A chemical reaction is represented by shorthand notation called chemical equation as

$$\textbf{Reactants} \longrightarrow \textbf{Products}$$

The arrow ($\longrightarrow$) indicates the direction of the reaction that reactants are changed into products.

Examples of Some Chemical Reactions

Hydrogen and oxygen combine so that water is formed under suitable experimental conditions.

Hydrogen +	Oxygen	$\longrightarrow$	Water	+ Heat
(Reactant)	(Reactant)		(Product)	

Methane gas burns in oxygen so that carbon dioxide and water are produced and heat energy is generated.

Methane + Oxygen $\longrightarrow$ Carbon dioxide + Water + Heat energy

CHARACTERISTICS OF CHEMICAL REACTIONS

A change is called a chemical reaction if it shows all or some of the following characteristics :

(i) Change of Physical State (ii) Evolution of gas (iii) Change in energy
(iv) Change in the colour (v) Formation of precipitate

Change of Physical State

In certain reactions, the physical state of products is different from that of the reactants. That is, there is a change of state during a chemical reaction.

When a mixture of hydrogen and oxygen is ignited with an electric spark at room temperature, liquid water is formed.

$$2H_2\,(g) + O_2\,(g) \xrightarrow[\text{room temperature}]{\text{electric spark}} 2H_2O\,(l)$$

Hydrogen gas Oxygen gas Water (liquid)

When ammonia (NH_3) gas is allowed to come in contact with hydrogen chloride (HCl) gas, solid ammonium chloride is obtained.

$$NH_3\,(g) + HCl\,(g) \longrightarrow NH_4Cl\,(s)$$

Ammonia Hydrogen chloride Ammonium chloride (white powder)

Evolution of Gas

Some chemical reactions take place with the evolution of a gas.

Some reactions in which a gas is evolved are described below :

Heating a mixture of potassium chlorate ($KClO_3$) and manganese dioxide (MnO_2) gives oxygen gas.

$$2KClO_3\,(s) \xrightarrow[MnO_2]{\text{heat}} 2KCl\,(s) + 3O_2(g)$$

Potassium chlorate Potassium chloride Oxygen gas

MnO_2 is used as a catalyst in this reaction.

Energy Changes

During a chemical change, energy is either evolved or absorbed. The energy evolved or absorbed may be in the form of heat, light, electricity, sound etc.

When coal (or carbon) is burnt, heat and light are produced.

$$C\,(s) + O_2\,(s) \longrightarrow CO_2\,(g) + \text{Heat} + \text{Light}$$

carbon/coal Oxygen (from the air) Carbon dioxide

When a small quantity of water is added to quicklime, a large amount of heat is evolved.

$$CaO\,(s) + H_2O\,(l) \longrightarrow Ca(OH)_2\,(s) + \text{Heat}$$

quicklime water slaked lime

This reaction takes place when lime is added to water for preparing the lime suspension which is used for whitewashing.

Limestone ($CaCO_3$) is burnt to obtain lime (or quicklime). In this reaction, heat is absorbed.

$CaCO_3$ (s) +	Heat	$\longrightarrow$	CaO (s)	+	CO_2 (g)
Limestone	Quicklime		Carbon dioxide (or lime)		

Change of Colour

There are some reactions in which there is a colour change.

For example :

When red lead oxide is heated, yellow lead monoxide is formed.

$2Pb_3O_4$ (s)	$\xrightarrow{\text{heat}}$	**6PbO(s)**	+	O_2 (g)
Red lead oxide		**Lead monoxide**		
(red colour)		**(yellow colour)**		

Formation of Precipitate

In certain reactions, when solution of two reagents are mixed, one of the products formed gets precipitated immediately. Colour of the precipitate depends upon the reagents used in the reaction.

When silver nitrate solution is mixed with a solution of sodium chloride, **white precipitate of silver chloride** is formed.

$AgNO_3$ (aq)	+	NaCl (aq)	$\longrightarrow$	**AgCl (s)**	+	$NaNO_3$ (aq)
Silver nitrate		sodium chloride		**silver chloride**		sodium nitrate
				(white precipitate)		

When potassium iodide solution is added to solution of lead nitrate, **yellow precipitate of lead iodide is formed.**

$Pb(NO_3)_2$ (aq)	+	2KI (aq)	$\longrightarrow$	**PbI_2 (s)**	+	$2KNO_3$ (aq)
lead nitrate		potassium iodide		**lead iodide**		potassium nitrate
				(yellow precipitate)		

CHEMICAL EQUATION

A chemical equation is defined as a shorthand notation of an actual chemical reaction in terms of the symbols and formulae along with the number of atoms and molecules of its reactants and products.

Methane + Oxygen $\longrightarrow$ Carbon dioxide + Water +Heat ...(i)

In the form of chemical equation, the reaction is represented as :

$CH_4(g) + 2O_2(g) \longrightarrow CO_2\,(g) + 2H_2O\,(\ell) + \text{Heat}$ ..(ii)

Equation (ii) represents a chemical reaction in terms of the formula of reactants and products. It also tells us about the number of atoms and molecules of various substances of the reaction.

INFORMATION CONVEYED BY CHEMICAL EQUATION

We get the following information from a balanced chemical equation :

(a) The formula, symbol, names and physical states of the reactants and products.

(b) The relative number of atoms and molecules of the reactants that take part in the reaction.

(c) The relative number of atoms and molecules of the products formed in the reaction.

ADVANTAGES OF THE USES OF A CHEMICAL EQUATION

(a) It saves time and space in writing.

(b) To prepare a known amount of the product we can calculate the amount of the reactant to be taken.

(c) The chemical equation can be understood by the chemists belonging to any country of the world irrespective of the language they speak. For example, the chemical equation $N_2 + 3H_2 \longrightarrow 2NH_3$, can be understood by chemist as : nitrogen and hydrogen combine in 1 : 3 mole ratio to give 2 mole ratio of ammonia.

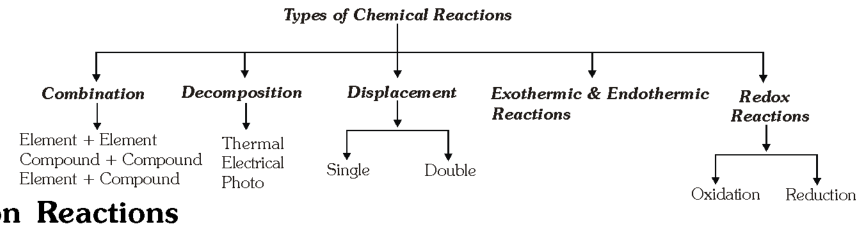
LIMITATIONS OF A CHEMICAL EQUATION

(a) A chemical equation does not tell us about the rate of the reaction and the time taken for the completion of the reaction.

(b) Some reactions may be even explosive. This is not revealed by a chemical equation.

(c) As such chemical equation does not tell about the actual quantity of the reactants consumed or products formed.

TYPES OF CHEMICAL REACTIONS



Combination Reactions

Those chemical reactions which involve the combination of two or more substances to form a single new substance are called combination reactions. Combination reactions may involve either.

(i) Combination of two elements or, (ii) Combination of an element and a compound or,

(iii) Combination of two compounds.

Let us now discuss all these types of combination reactions one by one.

(a) Combination reactions involving two elements

Some examples of combination reactions involving two elements are :

(i) Magnesium ribbon burns in oxygen with a brilliant flame and a white residue of magnesium oxide is formed.

$$\underset{\text{Magnesium}}{2\,Mg\,(s)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\substack{\text{Magnesium oxide} \\ \text{(White residue)}}}{2MgO(s)}$$

(ii) Carbon (charcoal) burns in air to form carbon dioxide

$$\underset{\text{Carbon (charcoal)}}{C\,(s)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2\,(g)}$$

In all the above reactions, two elements combine to form a single new compound and are, therefore, combination reactions. These reactions are also called **synthesis reactions**.

(b) Combination reactions involving an element and a compound

Some examples of combination reactions involving an element and a compound are :

(i) Carbon monoxide burns in oxygen to form carbon dioxide

$$\underset{\text{Carbon monoxide}}{2CO\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)} \longrightarrow \underset{\text{Carbon dioxide}}{2CO_2\,(g)}$$

(ii) Nitric oxide combines with oxygen at room temperature to form brown fumes of nitrogen dioxide (NO_2)

$$\underset{\text{Nitric oxide}}{2NO} + \underset{\text{Oxygen}}{O_2} \longrightarrow \underset{\text{Nitrogen dioxide}}{2NO_2}$$

(c) Combination reactions involving two compounds

Some examples of such reactions are :

(i) Quick lime (CaO) reacts with water to form calcium hydroxide (slaked lime)

$CaO\ (s) +$	$H_2O\ (l)$	$\longrightarrow$	$Ca(OH)_2\ (aq)$
Calcium oxide	Water		Calcium hydroxide
(Quick lime)			(Slaked lime solution) or (Lime water)

Decomposition Reactions

The conversion of a single compound into two or more simple substances is called decomposition. For example:

$$AB \longrightarrow A + B$$

Here AB is the original compound and A and B are simple substances. A and B may be elements or simpler compounds into which the original compound AB decomposes. We can see that decomposition is reverse of combination.

Types of Decomposition Reactions

Decomposition reactions are promoted by heat, electricity or light. Therefore, these reactions are classified as follows :

(a) Thermal decomposition (caused by heat)

(b) Electrolytic decomposition (caused by electricity)

(c) Photo decomposition (caused by light)

Thermal Decomposition

Heat is called thermal energy. On heating, some substances undergo decomposition and new substances are formed. This process is called thermal decomposition or thermal dissociation. For example :

Mercury (II) oxide decomposes on heating to form mercury and oxygen.

$2HgO(s)$	$\xrightarrow{\Delta}$	$2Hg\ (l)$	$+$	$O_2(g)$
Mercury(II)oxide		Mercury		Oxygen

Electrolytic Decomposition Reaction

The decomposition reactions caused by electric current is called electrical decomposition reaction or electrolysis. For example :

(a) On passing electricity through molten sodium chloride, it decomposes into sodium metal and chlorine gas.

$2NaCl\ (l)$	$\xrightarrow{\text{Electric current}}$	$2Na\ (s)$	$+$	$Cl_2\ (g)$
Sodium chloride		Sodium		Chlorine
(Molten)				

(b) On passing electricity through acidified water, it decomposes into hydrogen and oxygen.

$$2H_2O\ (l) \xrightarrow{\text{Electric current}} 2H_2(g) + O_2(g)$$

Photodecomposition Reactions

A decomposition reaction caused by light is called photodecomposition reaction.

For example:

In the presence of sunlight

$$2AgCl\ (s) \xrightarrow{\text{Sunlight}} 2Ag\ (s) + Cl_2(g)$$

Displacement Reactions

The chemical reactions in which an atom or a group of atoms in the molecule is replaced by another atom or a group of atoms are called displacement reactions. Let AL be any molecule and X be the displacing group, then the displacement reaction may be written as :

$$X + A - L \longrightarrow X - L + A$$

Here A is the leaving group.

These reactions are generally found to occur in the solution. The elements involved may be metals or non-metals, i.e. a more active metal may displace a less active metal or a more active non-metal may displace a less active non-metal from its compound.

Before we take up examples of displacement reactions, it is important to understand the relative reactivities of metals.

	Element	***Symbol***	
Metals More reactive than Hydrogen	Potassium	K	Most Reactive
	Sodium	Na	**Reactivity Decreases downward**
	Calcium	Ca	
	Barium	Ba	
	Magnesium	Mg	
	Aluminium	Al	
	Zinc	Zn	
	Iron	Fe	
	Nickel	Ni	
	Tin	Sn	
	Lead	Pb	
	Hydrogen	H	
Metals less reactive than Hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	Least Reactive
	Platinum	Pt	

Reaction 1.

When an iron nail is dipped in a copper sulphate solution, it gets coated with copper.

$Fe(s)$	+	$CuSO_4\ (aq)$	$\longrightarrow$	$FeSO_4\ (aq)$	+	$Cu\ (s)$
Iron		Copper sulphate		Iron sulphate		Copper

In this reaction, Fe has taken the place of Cu in the compound $CuSO_4$. In other words, we say that Fe has displaced Cu from the compound $CuSO_4$.

Conclusion

From this reaction we conclude that Fe (iron) is more reactive metal than Cu (Copper).

Displacement of a Less Active Non metal by a More Active Non metal

Reaction 1.

When chlorine gas is passed through sodium bromide solution, sodium chloride and bromine are formed.

$$\underset{\text{Sodium bromide}}{2NaBr(aq)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\ (aq)} + \underset{\text{Bromine}}{Br_2(g)}$$

Conclusion

In this reaction, chlorine has displaced bromine from NaBr. Therefore, chlorine is more reactive than bromine.

Reaction 2.

When chlorine is bubbled through sodium iodide solution, sodium chloride and iodine are formed

$$\underset{\text{Sodium iodide}}{2NaI(aq)} + \underset{\text{Chlorine}}{Cl_2\ (g)} \longrightarrow \underset{\text{Sodium chloride}}{2NaCl\ (aq)} + \underset{\text{Iodine}}{I_2\ (g)}$$

Conclusion

In this reaction Cl has displaced I from sodium iodide (NaI).

Therefore, Cl (chlorine) is more reactive than I (Iodine).

Double Displacement Reactions

The reactions given under displacement reactions involve displacement of an atom or group of atoms in a molecule. But there are some reactions in which two different atoms or groups of atoms are displaced by other atoms or groups of atoms. Such **reactions in which two compounds react by an exchange or displacement of ions to form new compounds are called double displacement reactions.** These reactions generally occur in solutions and in some cases, one of the products being insoluble, precipitates out as a solid and settles down.

Example 1.

If we add silver nitrate solution to sodium chloride solution, a white precipitate of silver chloride is formed along with the formation of sodium nitrate solution.

$$\underset{\text{Silver nitrate}}{AgNO_3\ (aq)} + \underset{\text{Sodium chloride}}{NaCl\ (aq)} \longrightarrow \underset{\substack{\textbf{Silver chloride}\\ \textbf{(White ppt.)}}}{\mathbf{AgCl\ (s)}} + \underset{\substack{\text{Sodium}\\ \text{Nitrate}}}{NaNO_3(aq)}$$

In this reaction, NO_3^- ions displace Cl^- ions and Cl^- ions displace NO_3^- ions. It may be noted that in this reaction, silver chloride is formed as an insoluble white solid known as white precipitate.

Example 2.

If ammonium hydroxide solution is added to aluminium chloride solution, a white precipitate of aluminium hydroxide along with ammonium chloride solution is obtained.

$$\underset{\substack{\text{Aluminium}\\ \text{chloride}}}{AlCl_3(aq)} + \underset{\substack{\text{Ammonium}\\ \text{hydroxide}}}{3NH_4OH(aq)} \longrightarrow \underset{\substack{\textbf{Aluminium}\\ \textbf{hydroxide}\\ \textbf{(White ppt.)}}}{\mathbf{Al(OH)_3\ (s)}} + \underset{\substack{\text{Ammonium}\\ \text{chloride}}}{3NH_4Cl(aq)}$$

In this double displacement reaction, Cl^- ions and OH^- ions have displaced each other to form insoluble aluminium hydroxide and ammonium chloride solution.

Double decomposition reactions can be further classified in two types :

(a) Precipitation reactions

(b) Neutralisation reactions

Let us discuss these double decomposition reactions individually.

(a) Precipitation reactions

Those reactions in which two clear and transparent solutions on mixing result in the formation of an insoluble product are known as precipitation reactions, and the insoluble product is known as precipitate.

Some examples of precipitation reactions are given here :

On mixing aqueous solution of sodium sulphide and lead acetate, a black precipitate of lead sulphide (PbS) is formed. The precipitation reaction may be represented as:

$Na_2S(aq)$	+	$(CH_3COO)_2Pb(aq)$	$\longrightarrow$	**PbS (s)**	+ $2CH_3COONa$ (aq)
Sodium sulphide		Lead acetate		**Lead sulphide (Black precipitate)**	Sodium acetate

On mixing the aqueous solutions of silver nitrate and sodium iodide, a yellow coloured precipitate of silver iodide will be obtained as shown above. In this reaction, iodide ions (I^-) have replaced nitrate ions (NO_3^-) and vice versa.

$AgNO_3(aq)$ +	$NaI(aq)$	$\longrightarrow$	**AgI (s)**	+ $NaNO_3$ (aq)
Silver nitrate	Sodium iodide		**Silver iodide (Yellow precipitate)**	Sodium nitrate

(b) Neutralization Reactions

When an aqueous solution of hydrochloric acid is mixed with an aqueous solution of sodium hydroxide in equal amounts, a reaction takes place to form sodium chloride and water

HCl (aq)	+	$NaOH(aq)$	$\longrightarrow$	$NaCl(aq)$	+ $H_2O(\ell)$
Hydrochloric acid		Sodium hydroxide		Sodium chloride	Water

Such a reaction is termed as a **neutralization reaction.** The hydrogen (H^+) ions which were responsible for the acidic properties of HCl have reacted with hydroxyl (OH^-) ions which were responsible for the basic properties of NaOH, there occurs a chemical change and appear in the form of crystalline sodium chloride on evaporation.

Such reactions in which an acid and a base react with each other to produce salt and water are known as neutralization reactions.

Exothermic and Endothermic Reactions

On the basis of energy changes chemical reactions are classified under two categories that are exothermic and endothermic reactions.

(a) Exothermic Reaction

The chemical reactions in which formation of products is accompanied by evolution of heat are known as exothermic reaction.

In such cases, the sign " + Heat" is written alongwith the products i.e.

Reaction $\longrightarrow$ Products + Heat

Examples :

(1) $C(s) + O_2(g) \longrightarrow CO_2(g) + \text{Heat}$ (2) $CH_4(g) + O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) + \text{Heat}$

(3) $2Mg(s) + O_2(g) \longrightarrow 2MgO(s) + \text{Heat}$ (4) $2H_2S(g) + 3O_2(g) \longrightarrow 2H_2O(l) + 2SO_2(g) + \text{Heat}$

(5) $CaO(s) + H_2O(l) \longrightarrow Ca(OH)_2(aq) + \text{Heat}$

(6) $C_6H_{12}O_6(aq) + 6O_2(g) \longrightarrow 6O_2(aq) + 6H_2O(l) + \text{Energy}$

(b) Endothermic Reaction

The chemical reactions in which formation of products is accompanied by the absorption of heat are known as endothermic reactions.

In such cases, the sign "+ Heat" is written alongwith the reactants or – 'Heat' with the products

Reactants $\longrightarrow$ Products - Heat

Example :

(1) $Ba(OH)_2(aq) + 2NH_4Cl(aq) \longrightarrow BaCl_2(aq) + 2NH_4OH(aq)$ - Heat

(2) $NH_4Cl(s) + H_2O(l) \longrightarrow NH_4OH(aq) + HCl(aq)$ - Heat

(3) $\underset{\text{Hot coke}}{C(s)} + \underset{\text{Steam}}{H_2O(g)} \longrightarrow \underbrace{CO(g) + H_2(g)}_{\text{Water gas}}$ - Heat

(4) $\underset{\text{Nitrogen}}{N_2(g)} + \underset{\text{Oxygen}}{O_2(g)} + \text{Heat} \longrightarrow \underset{\text{Nitric oxide}}{2NO(g)}$

This combination reaction is endothermic in nature.

(5) $C(s) + 2S(g) \longrightarrow CS_2(l)$ - Heat

Oxidation and Reduction reactions

In our daily life we come across processes like rusting of objects made of iron, fading of the colour of the clothes, burning of the combustible substances such as cooking gas, wood, coal, etc. All such processes fall in the category of specific type of chemical reactions called oxidation - reduction reactions or redox reactions.

Most of the elements are reactive and react with oxygen and hydrogen. Initially, on the basis of addition of oxygen and hydrogen, the chemical reactions were considered as oxidation and reduction reactions but afterwards, the definition was expanded, on the basis of addition or displacement of other elements except O_2 and H_2, which are as follows–

Oxidation

The oxidation of a substance takes place when :

(a) There is addition of oxygen to a substance.

The chemical reactions in which oxygen is added to a substance.

$C(s) + O_2(g) \longrightarrow CO_2(g)$

$2Mg(s) + O_2(g) \longrightarrow 2MgO(g)$

$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$

$2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$

(b) There is removal of hydrogen from a substance

The chemical reactions in which hydrogen is lost from a substance.

$H_2S(g) \longrightarrow H_2(g) + S(s)$

$2HI \longrightarrow H_2 + I_2$

Reduction

The reduction of a substance takes place when :

(a) There is addition of hydrogen to a substance.

Examples

$H_2 + S \longrightarrow H_2S$

$\mathbf{Cl_2 + H_2S \longrightarrow 2HCl + S}$

$H_2 + Cl_2 \longrightarrow 2HCl$

$2Na + H_2 \longrightarrow 2NaH$

(b) There is removal of oxygen from a substance

Examples :

$$ZnO + H_2 \longrightarrow Zn + H_2O$$

Here zinc oxide has been reduced to zinc by the removal of oxygen

$$ZnO + C \longrightarrow Zn + CO$$

Reduction

(Removal of oxygen)

$$CuO + H_2 \longrightarrow Cu + H_2O$$

Reduction

(Removal of oxygen)

So, in short, **reduction is a chemical reaction in which substance combine with hydrogen or an electropositive element or loses oxygen or an electronegative element.**

A substance that brings about oxidation that is addition of oxygen or electronegative element and removal of hydrogen or electropositive element is called oxidizing agent.

On the other hand, a substance that brings about reduction, that is removal of oxygen or electronegative element and addition of hydrogen or electropositive element is called reducing agent.

Consider the reaction

$$\boldsymbol{CuO + H_2 \longrightarrow Cu + H_2O}$$

In this reaction, hydrogen removes oxygen from copper oxide. Thus, CuO is reduced and hydrogen behaves as reducing agent.

Copper oxide gives oxygen to hydrogen and hydrogen is oxidised to water by CuO. Therefore, CuO is acting as oxidizing agent

(Removal of oxygen)

Reduction

$$CuO + H_2 \longrightarrow Cu + H_2O$$

(Addition of oxygen)

Oxidation

CuO makes oxidation to occur $\longrightarrow$ Oxidizing agent

H_2 makes reduction to occur $\longrightarrow$ Reducing agent

Note : (i) The substance to which oxygen is added or substance from which hydrogen is removed is said to be oxidized

(ii) The substance from which oxygen is removed substance to which hydrogen is added is said to be reduced

(iii) **The substance which gets oxidized acts as on reducing agent**

(iv) **The substances which gets reduced acts as oxidizing agent**

REDOX REACTIONS

Generally, oxidation and reduction reaction takes place simultaneously because in a chemical reaction, one atom, ion or molecule of a substance loses electron while another atom, ion or molecule gains electron.

In this way oxidation of one substance while reduction of another substance takes place. Thus such reactions are known as Redox reactions.

Example

(1) (oxidation)

$$2Mg + O_2 \longrightarrow 2MgO \quad (Mg^{2+}\ O^{-2})$$

(Reduction)

(oxidation)

(2) $Fe + S \xrightarrow{\Delta} FeS \quad (Fe^{2+}\ S^{-2})$

(Reduction)

EFFECTS OF OXIDATION REACTION IN EVERYDAY LIFE

There are a number of oxidation reactions taking place around us which affect our everyday life. Two of these are briefly described below :

Corrosion

The process of slowly eating up of the metals due to attack of atmospheric gases such as oxygen, carbon dioxide, hydrogen sulphide, water vapour etc. on the surface of the metals so as to convert the metal into oxide, carbonate, sulphide etc. is known as corrosion.

The most common example of corrosion is rusting i.e. corrosion of iron. When an iron article remains exposed to moist air for a long time, its surface is covered with a brown, flaky (non-sticky) substance called rust. Rust is mainly hydrated ferric oxide ($Fe_2O_3 .xH_2O$). It is formed due to attack of oxygen gas and water vapour present in the air on the surface of iron.

$$\underset{\text{Iron}}{2Fe(s)} + \underset{\text{(From air)}}{3/2O_2(g)} + \underset{\text{Moisture}}{xH_2O(l)} \longrightarrow \underset{\substack{\text{Hydrated ferric oxide}\\\text{(Rust)}}}{Fe_2O_3.xH_2O(s)}$$

Similarly , copper objects lose their lusture or shine after some time. The surface of these objects acquire a green coating of basic carbonate, $CuCO_3.Cu(OH)_2$ when exposed to air. This is due to attack of O_2, CO_2 and water vapour present in the air on the surface of copper.

$$\underset{\text{Copper}}{2Cu(s)} + \underbrace{CO_2(g) + O_2(g) + H_2O(l)}_{\text{From moist air}} \longrightarrow \underset{\substack{\text{Basic copper carbonate}\\\text{(Green)}}}{CuCO_3.Cu(OH)_2}$$

Likewise, the surface of silver metal gets tarnished (i.e. loses lustre and becomes dull) on exposure to air. This is due to the formation of a coating of black silver sulphide (Ag_2S) on its surface by the action of H_2S present in the air.

$$\underset{\text{Silver}}{2Ag(s)} + \underset{\substack{\text{Hydrogen sulphide}\\\text{(From air)}}}{H_2S(g)} \longrightarrow \underset{\substack{\text{Silver sulphide}\\\text{(Black)}}}{Ag_2S(s)} + H_2(g)$$

Rusting is a serious problem because it weakens the structure of bridges, iron railings, automobile parts etc. Every year a large amount of money is spent to replace rusted iron and steel structure. The reason is that the reddish brown crust of rust does not stick to the surface. It peels off (or falls down) exposing fresh surface for rusting. Thus, corrosion of iron is a continuous process which ultimately eats up the whole iron object.

Methods to prevent rusting; Rusting can be prevented if iron objects are not allowed to come in contact with the damp air. Some common methods generally used are given below.

(i) By painting the iron articles such as window grills, iron gates, steel furniture, railway coaches bodies of cars, buses etc.

(ii) By greasing and oiling the iron articles such as mechanical tools, machine parts etc.

(iii) By galvanisation, i.e. coating the surface of iron objects with a thin layer of zinc.

Rancidity

Oxidation also has damaging effect on foods containing fats and oils. When the food materials prepared in fats and oils are kept for a long time, they start giving unpleasant smell and taste. The fat and oil containing food materials which give unpleasant smell and taste are said to have become rancid (sour or stale). This happens as follows :

When the fats and oils present in food materials get oxidised by the oxygen (of air), their oxidation products have unpleasant smell and taste. Due to this, the smell and taste of food materials containing fats and oils change and become very unpleasant (or obnoxious). **The condition produced by aerial oxidation of fats and oils in foods marked by unpleasant smell and taste is called rancidity.** Rancidity spoils the food materials prepared in fats and oils which have been kept for a considerable time and makes them unfit for eating. The characteristics of a rancid food are that it gives out unpleasant smell and also has an unpleasant taste. Rancidity is called 'vikritgandhita' in Hindi.

The development of rancidity of food can be prevented or retarded (slowed down) in the following ways

1. Rancidity can be prevented by adding anti-oxidants to foods containing fats and oils.

Anti-oxidant is a substance (or chemical) which prevents oxidation. Anti-oxidants are actually reducing agents. When anti-oxidants are added to foods, then the fats and oils present in them do not get oxidised easily and hence do not turn rancid. So the foods remain good to eat for a much longer time. The two common anti-oxidants used in foods to prevent the development of rancidity are BHA (Butylated Hydroxy-Anisole) and BHT (Butylated Hydroxy-Toluene).

2. Rancidity can be prevented by packaging fat and oil containing foods in nitrogen gas.

3. Rancidity can be retarded by keeping food in a refrigerator.

4. Rancidity can be retarded by storing food in air-tight containers.

5. Rancidity can be retarded by storing foods away from light.

EXERCISE

CHEMICAL REACTIONS & EQUATIONS

INTRODUCTION

1. Which is not a physical change ?
(1) Digestion of food (2) Formation of dew
(3) Sublimation of camphor (4) Ringing of an electric bell

2. Which is not a chemical change ?
(1) Rusting of iron (2) Evaporation of water
(3) Clotting of blood (4) Burning of wood

3. Which of the following is not a physical change?
(1) Boiling of water to give water vapour (2) Melting of ice to give water
(3) Dissolution of salt in water (4) Combustion of liquefied petroleum Gas (LPG)

4. Which of the following is not the characteristic of a chemical change?
(1) It always involves absorption or liberation of energy
(2) No net energy change is involved
(3) Chemically new substance is formed
(4) It is irreversible

5. Which among the following is not a chemical change?
(1) Melting of ice (2) Carbon cycle
(3) Dehydration of substances (4) Fermentation of substances

6. An example of a chemical change is _____.
(1) formation of clouds (2) glowing of an electric light
(3) dropping sodium into water (4) dissolving of salt in water

7. Which of the following is a physical change?
(1) Solubility in water (2) Combustibility
(3) Aerial oxidation (4) Reaction of sodium metal with water

8. Which among the following is not a physical change?
(1) Melting of solids to liquids (2) Vaporisation of liquids to gases
(3) Liquefaction of gases to liquids (4) Decay of matter

9. Which among the following is not a chemical change?
(1) Melting of ice (2) Carbon cycle
(3) Dehydration of substances (4) Fermentation of substances

10. An example of a chemical change is _____.
(1) formation of clouds (2) glowing of an electric light
(3) dropping sodium into water (4) dissolving of salt in water

11. Which of the following is a physical change?
(1) Solubility in water (2) Combustibility
(3) Aerial oxidation (4) Reaction with water

12. Which of the following is not a physical change?
(1) Boiling of water to give water vapour (2) Melting of ice to give water
(3) Dissolution of salt in water (4) Combustion of Liquefied Petroleum Gas (LPG)

13. Which of the following process is reversible?

(1) Melting of ice at 100° C
(2) Evaporation of water at room temperature
(3) Mixing of carbon dioxide and oxygen
(4) None of the above

14. During an irreversible chemical reaction the

(1) concentration of the reactants and products remain same
(2) concentration of reactants decreases and for products increases
(3) concentration of reactants increases and for products decreases
(4) None of the above.

15. What is the value of x in given equation?

$$yAl + xH^+ \rightarrow yl^{3+} + zH_2$$

(1) 2
(2) 4
(3) 6
(4) 8

CHEMICAL REACTIONS AND EQUATIONS

16. When lead nitrate reacts with potassium iodide, yellow precipitate of :

(1) PbI_2 is formed
(2) KNO_3 is formed
(3) $Pb(NO_3)_2$ is formed
(4) $PbIO_3$ is formed

17. Formation of carbon disulphide from carbon and sulphur takes place by :

(1) Absorption of heat
(2) Evolution of heat
(3) No change in heat content
(4) None of the above

18. In which of the following equations, the abbreviations represent the correct states of the reactants and products involved at room temperature?

(1) $2H_2(\ell) + O_2(\ell) \rightarrow 2H_2O(g)$
(2) $2H_2(g) + O_2(\ell) \rightarrow 2H_2O(\ell)$
(3) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$
(4) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

19. A chemical equation is balanced in accordance with the law of -

(1) Conservation of mass (2) Multiple proportion (3) Constant proportion (4) Reciprocal proportion

20. $xAgI + Na_2S \rightarrow yAg_2S + zNaI$

x, y and z respectively are

(1) 2, 1, 2
(2) 1, 1, 1
(3) 1, 1, 2
(4) 1, 2, 2

21. Which of the following reaction is balanced?

(1) $Ba_3N_2 + 6H_2O \rightarrow 3Ba(OH)_2 + 2NH_3$
(2) $3Hg(OH)_2 + 2H_3PO_4 \rightarrow Hg_3(PO_4)_2 + 6H_2O$
(3) Both (1) and (2)
(4) $As + 6NaOH \rightarrow 2Na_3AsO_3 + 3H_2$

22. $2Fe + xHC_2H_3O_2 \rightarrow 2Fe(C_2H_3O_2)_3 + 3H_2$ the value of x will be

(1) 7
(2) 6
(3) 12
(4) 2

23. Which of the following equations is most informative:

(1) $N_2 + H_2 \longrightarrow NH_3$
(2) $N_2 + 3H_2 \longrightarrow 2NH_3$
(3) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g) + 92\ kJ$
(4) $2N_2 + H_2 \longrightarrow NH_3$

24. Which of the following equations is representing combination of two elements?

(1) $CaO + CO_2 \rightarrow CaCO_3$
(2) $4Na + O_2 \rightarrow 2Na_2O$
(3) $SO_2 + 1/2O_2 \rightarrow SO_3$
(4) $2Na + 2H_2O \rightarrow 2NaOH + H_2$

25. In one molecule of ammonium sulphide there are ___________.

(1) 2 atoms of N, 8 atoms of H, and 1 atom of S
(2) 1 atom of N, 4 atoms of H, and 1 atom of S
(3) 1 atom of N, 4 atoms of H, and 2 atoms of S
(4) 2 atoms of N, 8 atoms of H, and 2 atoms of S

26. The correctly balanced equation for $FeS + O_2 \rightarrow Fe_2O_3 + SO_2$ is ________.

(1) $2FeS + O_2 \rightarrow Fe_2O_3 + 4SO_2$ (2) $2FeS + 3O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

(3) $4FeS + 4O_2 \rightarrow 2Fe_2O_3 + 2SO_2$ (4) $4FeS + 7O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

27. Choose the wrong statement :

(1) Chemical equation helps us to know the ratio of the moles of the reactants and products

(2) The relative number of atoms and molecules can be known by using chemical equation

(3) A chemical equation tells us about the rate of reaction and the time taken for completion

(4) The representation of actual chemical reaction becomes easy by using equation

28. The sign used to indicate a reversible reaction is _______.

(1) $\rightarrow$ (2) $\rightleftharpoons$ (3) $\leftarrow$ (4) $=$

29. The correctly balanced equation for $FeS + O_2 \rightarrow Fe_2O_3 + SO_2$ is ________.

(1) $2FeS + O_2 \rightarrow Fe_2O_3 + 4SO_2$ (2) $2FeS + 3O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

(3) $4FeS + 4O_2 \rightarrow 2Fe_2O_3 + 2SO_2$ (4) $4FeS + 7O_2 \rightarrow 2Fe_2O_3 + 4SO_2$

30. In which of the following chemical equations, the abbreviations represent the correct states of the reactants and products involved at room temperature?

(1) $2H_2(l) + O_2(l) \rightarrow 2H_2O(g)$ (2) $2H_2(g) + O_2(l) \rightarrow 2H_2O(l)$

(3) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ (4) $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

31. In the reaction given below a, b, c and d respectively are

$$aFe_2O_3 + bH_2 \rightarrow cFe + dH_2O$$

(1) 1, 1, 2, 3 (2) 1, 1, 1, 1 (3) 1, 3, 2, 3 (4) 1, 2, 2, 3

32. Which of the following is the coefficient for O_2 when the equation is balanced ?

$$C_4H_{10} + O_2(g) \rightarrow CO_2(g) + H_2O(g)$$

(1) 4 (2) 8 (3) 10 (4) 13

33. What is the sum of the coefficients of the following equation when it is balanced?

$$FeS(s) + HCl(g) \rightarrow FeCl_2(s) + H_2S(g)$$

(1) 5 (2) 4 (3) 3 (4) 2

34. Find x in the equation $3Fe + xH_2O \rightarrow Fe_3O_4 + 4H_2$.

(1) 1 (2) 2 (3) 3 (4) 4

35. What is the number of nitrogen atoms on the reactant side of the equation $H_2(g) + N_2(g) \rightarrow NH_3(g)$ after balancing?

(1) 1 (2) 2 (3) 3 (4) 4

36. Which of the following is not a clue that a chemical reaction has taken place ?

(1) A solid is formed when two clear solutions are mixed.

(2) A clear solution is added to a red solution and the result is a blue solution.

(3) A solid is added to water, and bubbles form.

(4) A pure solid is heated and turns into a pure liquid.

37. What does it mean for a chemical equation to be balanced ?

(1) The sums of the coefficients on each side of the equation are equal.

(2) The same number of each type of atom appears on both sides of the equation.

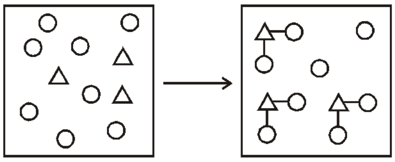
(3) The formulae of the reactants and the products are the same.

(4) The sums of the subscripts on each side of the equation are equal.

38. Which of the following physical states is not used for H_2O in a chemical equation ?

(1) (s) (2) (l) (3) (aq) (4) (g)

39. The reaction of an element X (4) with element Y (O) is represented in the following diagram. Which of the equations best describes this reaction ?



(1) $3X + 8Y \rightarrow X_3Y_8$ (2) $3X + 6Y \rightarrow X_3Y_6$

(3) $X + 2Y \rightarrow XY_2$ (4) $3X + 8Y \rightarrow 3XY_2 + 2Y$

40. In the reaction $FeSO_4 + x \rightarrow Na_2SO_4 + Fe(OH)_2$, x is

(1) Na_2SO_4 (2) H_2SO_4 (3) NaOH (4) None of these

41. Sum of coefficients of products in the following balanced chemical equations :

$Pb_3O_4 + HCl \rightarrow PbCl_2 + H_2O + Cl_2$

(1) 5 (2) 17 (3) 9 (4) 8

42. In the following equation :

$Na_2CO_3 + xHCl \rightarrow 2NaCl + CO_2 + H_2O$, the value of x is

(1) 1 (2) 2 (3) 3 (4) 4

43. Carbon dioxide gas is soluble in water. What happens when the temperature is increased?

(1) solubility of carbon dioxide decreases

(2) solubility of carbon dioxide increases

(3) temperature rise has no effect on solubility

(4) carbon dioxide decomposes into carbon and oxygen and hydrogen

44. The reaction $C + O_2 \longrightarrow CO_2$ + Heat is a –

(1) Combination reaction (2) Oxidation reaction (3) Exothermic reaction (4) All of the above

45. Give the number of molecules of carbon-dioxide formed upon complete oxidation of glucose.

(1) 1 (2) 3 (3) 6 (4) 4

46. Which of the following equations is not correctly balanced ?

(1) $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ (2) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

(3) $A1 + 2H^+ \rightarrow A1^{3+} + H_2$ (4) $Br_2 + 2I^- \rightarrow 2Br^- + I_2$

47. Formation of carbon disulphide from carbon and sulphur takes place by

(1) absorption of heat (2) evolution of heat

(3) no change in heat content (4) None of these

48. In the equation

$Cu + xHNO_3 \rightarrow Cu(NO_3)_2 + yNO_2 + 2H_2O$

The values of x and y are

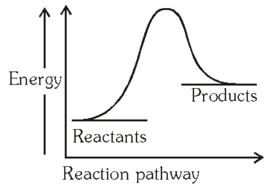
(1) 3 and 5 (2) 8 and 6 (3) 4 and 2 (4) 7 and 1

49. Which of the following chemical equations is an unbalanced one?

(1) $2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$ (2) $2C_4H_{10} + 12O_2 \rightarrow 8CO_2 + 10H_2O$

(3) $2Al + 6H_2O \rightarrow 2Al(OH)_3 + 3H_2$ (4) $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

50. The following reaction coordinate diagram represents



(1) an exothermic reaction.
(2) an endothermic reaction.
(3) a reaction that is neither endothermic nor exothermic.
(4) a reaction in which a catalyst is used.

51. When carbon dioxide gas is passed through lime water,
(1) calcium hydroxide is formed. (2) white precipitate of CaO is formed.
(3) white precipitate of $CaCO_3$ is formed. (4) colour of lime water disappears.

52. X is a metal which can replace Y and Z both from their salt solution. Y can replace Z but not W. W can replace both Y and Z but not X. What is the correct reactivity order of X, Y, Z and W?
(1) W > X > Y > Z (2) X > Y > Z > W (3) X > W > Y > Z (4) X > W > Z > Y

53. A dilute solution of sodium carbonate was added to two test tubes-one containing dil. HCl (1) and the other containing dilute NaOH (2). The correct observation was
(1) a brown coloured gas liberated in test tube A.
(2) a brown coloured gas liberated in test tube B.
(3) a colourless gas liberated in test tube A.
(4) a colourless gas liberated in test tube B.

54. The solution of a chemical compound x reacts with $AgNO_3$ to form a white precipitate of y. x and y can be–
(1) NaCl, AgCl (2) NaBr, AgBr (3) NaI, AgI (4) CH_3Cl, AgCl

55. Dilute hydrochloric acid is added to granulated zinc taken in a test tube. The following observations are recorded. Point out the correct observation.
(1) The surface of metal becomes shining
(2) The reaction mixture turns milky
(3) Odour of a pungent smelling gas is recorded
(4) A colourless and odourless gas is evolved

TYPES OF CHEMICAL REACTIONS

56. On heating crystals of ferrous sulphate products obtained are :
(1) Ferric oxide, Sulphur dioxide, Sulphur trioxide
(2) Ferric oxide, Ferrous sulphide, Oxygen
(3) Ferrous sulphide, Sulphur dioxide, Oxygen
(4) Ferric oxide, Sulphur trioxide, Oxygen

57. A small amount of quick lime is added to water and calcium hydroxide was obtained. What kind of chemical reaction takes place between them?
(1) Decomposition reaction (2) Combination between two elements
(3) Combination between two compounds (4) None of these

58. Which of the following reactions is/are a double displacement reaction (s)?
(i) $AgNO_3 + NaBr \rightarrow NaNO_3 + AgBr$ (ii) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$
(iii) $As_2O_3 + 3H_2S \rightarrow As_2S_3 + 3H_2O$ (iv) $NaOH + HCl \rightarrow NaCl + H_2O$
(1) (i) and (ii) only (2) Only (iii) (3) Only (iv) (4) All

59. The reaction: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$, is
(i) Neutralization reaction (ii) Double displacement reaction
(iii) Decompositon reaction (iv) Addition reaction
(1) (i) and (iv) only (2) (i) and (ii) only (3) (i) and (iii) only (4) (ii) and (iv) only

60. On keeping iron nail in copper sulphate solution for two hours, the colour of solution changes.

(1) from colourless to blue
(2) from light green to blue
(3) from blue to light green
(4) from light green to colourless

61. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$

The above reaction is an example of

(1) combination reaction
(2) double displacement reaction
(3) decomposition reaction
(4) displacement reaction

62. What happens when dilute hydrochloric acid is added to iron filling? Tick the correct answer:

(1) Hydrogen gas and iron chloride are produced
(2) Chlorine gas and iron chloride are produced
(3) No reaction takes place
(4) Iron salt and water are produced

63. Which of the following is a combustion reaction?

(1) Boiling of water (2) Melting of wax (3) Burning of petrol (4) None of these

64. The reaction $C + O_2 \longrightarrow CO_2$ + Heat is an example of :

(1) Combination reaction
(2) Oxidation reaction
(3) Exothermic reaction
(4) All of the above

65. Conversion of $CaCO_3$ into CaO as per following reaction is an example of –

$$CaCO_3 \longrightarrow CaO + CO_2$$

(1) Decomposition reaction
(2) Reduction reaction
(3) Oxidation reaction
(4) None of these

66. When one mole of lead nitrate is heated then:

(1) 2 moles of NO_2 gas are produced
(2) 1 moles of O_2 gas is produced
(3) both (1) and (2)
(4) 4 moles of NO_2 gas are produced

67. Copper (II) carbonate on thermal decomposition produces________ and __________ and the colour changes from _______ to _______ :

(1) CuO, CO_2, black, blue
(2) CO_2, Cu_2O, green, black
(3) Cu_2O, CO_2, black, blue
(4) CO_2, CuO, green, black

68. Which of the following is a decomposition reaction?

(1) $NaOH + HCl \longrightarrow NaCl + H_2O$
(2) $NH_4CNO \longrightarrow H_2NCONH_2$
(3) $2KClO_3 \longrightarrow 2KCl + 3O_2$
(4) $H_2 + I_2 \longrightarrow 2HI$

69. Which of the following reaction is not a precipitation reaction :

(1) $AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$
(2) $Pb(NO_3)_2 + 2KI \longrightarrow PbI_2 + 2KNO_3$
(3) $FeCl_3 + 3NH_4OH \longrightarrow Fe(OH)_3 + 3\,NH_4Cl$
(4) $2KI + Cl_2 \longrightarrow 2KCl + I_2$

70. $Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$

This reaction is an example of –

(1) Combination reaction
(2) Double displacement reaction
(3) Decomposition reaction
(4) Displacement reaction

71. Which of the following reaction is double displacement reaction?

(1) $FeCl_3 + 3NaOH \longrightarrow Fe(OH)_3 + 3NaCl$

(2) $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$

(3) $2CO + O_2 \longrightarrow 2CO_2$

(4) $N_2 + O_2 \longrightarrow 2NO$

72. Which of the following is a displacement reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$
(2) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$
(3) $Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$
(4) $NaOH + HCl \longrightarrow NaCl + H_2O$

73. $Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$. This reaction is an example of :

(1) Combination reaction
(2) Double displacement reaction
(3) Decomposition reaction
(4) Displacement reaction

74. What happens when dilute hydrochloric acid is added to iron fillings?

(1) Hydrogen gas and iron chloride are produced

(2) Chlorine gas and iron hydroxide are produced

(3) No reaction takes place

(4) Iron salt and water are produced.

75. When iron nails are added to an aqueous solution of copper sulphate, a chemical change occurs, which of the following is not true about this reaction?

(1) Blue colour of the solution fades.
(2) Iron nails become brownish in colour.
(3) It is a displacement reaction.
(4) Iron nails dissolve completely.

76. $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$ is an example of ________.

(1) neutralization reaction
(2) redox reaction
(3) double displacement reaction
(4) decomposition reaction

77. In the reaction : $BaCl_2 + ZnSO_4 \rightarrow ZnCl_2 + BaSO_4$, the white precipitate seen is due to ______.

(1) $ZnCl_2$
(2) $BaSO_4$
(3) $BaCl_2$
(4) $ZnSO_4$

78. Which of the following represent a double displacement reaction?

(1) $2H_2 + O_2 \longrightarrow 2H_2O$
(2) $2Mg + O_2 \longrightarrow 2MgO$
(3) $AgNO_3 + NaCl \longrightarrow AgCl \downarrow + NaNO_3$
(4) $H_2 + Cl_2 \longrightarrow 2HCl$

79. Which of the following reaction is double displacement reaction ?

(1) $FeCl_3 + 3NaOH \longrightarrow Fe(OH)_3 + 3NaCl$

(2) $Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$

(3) $2CO + O_2 \longrightarrow 2CO_2$

(4) $N_2 + O_2 \longrightarrow 2NO$

80. Which among the following is(are) double displacement reaction(s)?

(i) $Pb + CuCl_2 \rightarrow PbCl_2 + Cu$
(ii) $Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$
(iii) $C + O_2 \rightarrow CO_2$
(iv) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

(1) (i) and (iv)
(2) (ii) only
(3) (i) and (ii)
(4) (iii) and (iv)

81. Reaction between barium chloride and aluminium sulphate is an example of

(1) Neutralisation reaction
(2) Displacement reaction
(3) Precipitation reaction
(4) Combination reaction.

82. What are the products obtained when potassium sulphate reacts with barium iodide in an aqueous medium?

(1) KI and $BaSO_4$
(2) KI, Ba and SO_2
(3) K, I_2 and $BaSO_4$
(4) K, Ba, I_2 and SO_2

83. $AlCl_3 + 3NaOH \rightarrow Al(OH)_3 + 3NaCl$

Aluminum chloride + Sodium hydroxide → Aluminum hydroxide + Sodium chloride

The above reaction is an example of

(1) decomposition reaction (2) double displacement reaction

(3) synthesis reaction (4) single displacement reaction

84. Which one is the correct statement about the reaction?

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

(1) Iron is more reactive than copper

(2) Iron is less reactive than copper

(3) In this reaction Cu^{2+} is oxidized and Fe is reduced

(4) Copper is more reactive than iron

85. Following is the reactivity series in decreasing order of their reactivity -

Magnesium > Zinc > Iron > Lead > Copper > Silver > Gold

Which one of the following metals can displace copper from copper sulphate solution?

(1) Zn (2) Ag (3)Au (4) None

86. The reaction $Pb(NO_3)_2 + 2NaI \rightarrow 2NaNO_3 + PbI_2 \downarrow$ is classified as

(1) combination reaction (2) decomposition reaction

(3) single displacement reaction (4) double displacement reaction

87. The reaction between an acid and a base is ______ reaction.

(1) chlorination (2) hydrolysis (3) neutralisation (4) hydrogenation

88. The most reactive halogen is

(1) fluorine (2) iodine (3) chlorine (4) bromine

89. In a displacement reaction, chlorine can displace

(1) iodine (2) bromine (3) both (1) & (2) (4) neither (1) nor (2)

90. Which of the following pairs will give displacement reactions ?

(1) $ZnSO_4$ solution and Aluminium metal (2) $MgCl_2$ solution and Aluminium metal

(3) $FeSO_4$ solution and Silver metal (4) $NaNO_3$ solution and Copper metal

91. What happens when zinc granules reacts with hydrochloric acid?

(1) Precipitation (2) Neutralisation (3) Gas-formation (4) Oxidation

92. The reaction $H_2 + Cl_2 \longrightarrow 2HCl$ is a –

(1) Decomposition reaction (2) Combination reaction

(3) Double displacement reaction (4) Displacement reaction

93. Breaking of lead bromide into lead and bromine is an example of __________.

(1) decomposition reaction (2) synthesis reaction

(3) displacement reaction (4) neutralisation reaction

94. Which of the following is not a decomposition reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$ (2) $2KClO_3 \longrightarrow 2KCl + 3O_2$

(3) Digestion of food in the body (4) $H_2 + Cl_2 \longrightarrow 2HCl$

95. Which of the following is a decomposition reaction?

(1) $NaOH + HCl \longrightarrow NaCl + H_2O$ (2) $NH_4CNO \longrightarrow H_2NCONH_2$

(3) $2KClO_3 \longrightarrow 2KCl + 3O_2$ (4) $H_2 + I_2 \longrightarrow 2HI$

96. Conversion of $CaCO_3$ into CaO as per following reaction is an example of –

$$CaCO_3 \longrightarrow CaO + CO_2$$

(1) Decomposition reaction
(2) Reduction reaction
(3) Oxidation reaction
(4) None of these

97. Which of the following are exothermic processes?

(i) Reaction of water with quick lime
(ii) Dilution of an acid
(iii) Evaporation of water
(iv) Sublimation of camphor (crystals)

(1) (i) and (ii) (2) (ii) and (iii) (3) (i) and (iv) (4) (iii) and (iv)

98. Three beakers labelled as A, B and C each containing 25 mL of water were taken. A small amount of NaOH, anhydrous $CuSO_4$ and NaCl were added to the beakers A, B and C respectively. It was observed that there was an increase in the temperature of the solutions contained in beakers A and B, whereas in case of beaker C, the temperature of the solution falls. Which one of the following statement(s) is(are) correct?

(i) In beakers A and B, exothermic process has occurred.
(ii) In beakers A and B, endothermic process has occurred.
(iii) In beaker C exothermic process has occurred.
(iv) In beaker C endothermic process has occurred.

(1) (i) only (2) (ii) only (3) (i) and (iv) (4) (ii) and (iii)

99. Which among the following statement(s) is(are) true? Exposure of silver chloride to sunlight for a long duration turns grey due to

(i) the formation of silver by decomposition of silver chloride
(ii) sublimation of silver chloride
(iii) decomposition of chlorine gas from silver chloride
(iv) oxidation of silver chloride

(1) (i) only (2) (i) and (iii) (3) (ii) and (iii) (4) (iv) only

100. Electrolysis of water is a decomposition reaction. The mole ratio of hydrogen and oxygen gases liberated during electrolysis of water is

(1) 1 : 1 (2) 2 : 1 (3) 4 : 1 (4) 1 : 2

101. Which of the following is an endothermic process?

(1) Addition of quick lime to water
(2) Burning of LPG
(3) Decomposition of limestone
(4) Dilution of acid

102. The following reaction is used for the preparation of oxygen gas in the laboratory

$$2KClO_3\ (s) \xrightarrow[\text{Catalyst}]{\text{Heat}} 2KCl\ (s) + 3O_2\ (g)$$

Which of the following statement(s) is(are) correct about the reaction?

(1) It is a decomposition reaction and endothermic in nature
(2) It is a combination reaction
(3) It is a decomposition reaction and accompanied by release of heat
(4) It is a photochemical decomposition reaction and exothermic in nature

103. Which of the following are combination reactions?

(i) $2KClO_3 \xrightarrow{\text{Heat}} 2KCl + 3O_2$
(ii) $MgO + H_2O \rightarrow Mg(OH)_2$
(iii) $4Al + 3O_2 \rightarrow 2Al_2O_3$
(iv) $Zn + FeSO_4 \rightarrow ZnSO_4 + Fe$

(1) (i) and (iii) (2) (iii) and (iv) (3) (ii) and (iv) (4) (ii) and (iii)

104. Which of the following chemical reactions is balanced as well as undergo combustion ?

(1) $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
(2) $6CO_2 + 12H_2O \rightarrow C_6H_{12}O_6 + 6O_2 + 6H_2O$
(3) $C_3H_6 + 4O_2 \rightarrow 3CO_2 + 3H_2O$
(4) $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$

105. Which of the following statement is incorrect?

(1) In an exothermic reaction heat is evolved.
(2) An endothermic reaction must absorb energy before it can occur.
(3) For an endothermic reaction, the value of enthalpy is in positive.
(4) After an exothermic reaction, there is no change in temperature of the reaction mixture.

106. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ is an example of

(1) Displacement (2) Neutralization (3) Redox (4) Both (1) and (3)

107. Which of the following is a combustion reaction?

(1) Boiling of water (2) Melting of wax (3) Burning of petrol (4) None of these

108. The reaction, $H_2 + Cl_2 \longrightarrow 2HCl$ is

(1) decomposition reaction
(2) displacement reaction
(3) a combination reaction
(4) an isomerisation reaction

109. Decomposition of water is

(1) electrolytic (2) thermal (3) photolytic (4) all of these

110. Which of the following is a decomposition reaction?

(1) $NaOH + HCl \longrightarrow NaCl + H_2O$
(2) $NH_4CNO \longrightarrow H_2NCONH_2$
(3) $2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$
(4) $H_2 + I_2 \longrightarrow 2HI$

111. Heating limestone produces

(1) quick lime (2) carbon dioxide (3) both (1) and (2) (4) none of these

112. When copper powder is heated, it gets coated with

(1) black copper oxide
(2) yellow copper oxide
(3) red copper oxide
(4) none of these

113. Which of the following equations is representing combination of two compounds?

(1) $CaO + CO_2 \longrightarrow CaCO_3$
(2) $CO + \frac{1}{2}O_2 \longrightarrow CO_2$
(3) $SO_2 + \frac{1}{2}O_2 \longrightarrow SO_3$
(4) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

114. What happens when dilute hydrochloric acid is added to iron filings?

(1) Hydrogen gas and iron chloride are produced.
(2) Chlorine gas and iron hydroxide are produced.
(3) No reaction takes place.
(4) Iron salt and water are produced.

115. The reaction, $Fe_2O_3(s) + 2Al(s) \longrightarrow Al_2O_3 + 2Fe(s)$ is an example of

(1) combination reaction
(2) double displacement
(3) decomposition reaction
(4) single displacement reaction

116. Which of the following is a displacement reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$
(2) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$
(3) $Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$
(4) $NaOH + HCl \longrightarrow NaCl + H_2O$

117. What happens when copper rod is dipped in iron sulphate solution?

(1) Copper displaces iron.
(2) Blue coloured copper sulphate solution is obtained.
(3) No reaction takes place.
(4) Reaction is exothermic.

118. A student added dilute HCl to a test tube containing zinc granules and made following observations:

I. The zinc surface became dull and black.

II. A gas is evolved which burnt with a pop sound.

III. The solution remained colourless.

Correct observations are

(1) I & II (2) I & III (3) II & III (4) I, II & III

119. $AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$

Above reaction is

(1) precipitation reaction (2) double displacement reaction

(3) combination reaction (4) both (1) and (2)

120. $Zn + \text{(dil.)}\ H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$

Above reaction is

(1) decomposition reaction (2) single displacement reaction

(3) oxidation reaction (4) neutralisation reaction

121. $CuO + H_2 \rightarrow H_2O + Cu$, reaction is an example of

(1) redox reaction (2) decomposition reaction (3) neutralisation (4) analysis reaction

122. When lead nitrate reacts with potassium iodide, yellow precipitate of

(1) PbI_2 is formed. (2) KNO_3 is formed

(3) $Pb(NO_3)_2$ is formed. (4) $PbIO_3$ is formed.

123. ________is used in black and white photography.

(1) $AgCO_3$ (2) AgCl (3) $AgNO_3$ (4) AgS

124. 'M' is a metal combines with oxygen $4M + 3O_2 \rightarrow 2M_2O_3$. The reaction respresents________.

(1) combination reaction as well as reduction reaction

(2) decomposition as well as oxidation reaction

(3) oxidation reaction as well as displacement reaction

(4) combination reaction as well as oxidation reaction

125. _________, is not a decomposition reaction.

(1) Calcium carbonate is heated (2) Ammonium chloride is heated

(3) Potassium nitrate is heated (4) Magnesium ribbon is heated

126. (a) $AgNO_3(aq) + Cu(s) \longrightarrow CuNO_3(aq) + Ag(s)$

(b) $CuSO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + Cu(s)$

(c) $FeSO_4(aq) + Zn(s) \longrightarrow ZnSO_4(aq) + Fe(s)$

The most reactive among these metals is :

(1) Zn (2) Ag (3) Cu (4) Fe

127. The basis of black and white photography is -

(1) decomposition of lead salts (2) combination of lead salts

(3) decomposition of silver salts (4) combination of silver salts

128. When ferrous hydroxide reacts with hydrochloric acid, _______ and H_2O are produced.

(1) $FeCl_3$ (2) $FeCl_2$ (3) $FeCl_4$ (4) FeCl

129. Lead nitrate on heating gives

(1) lead oxide (2) nitrogen dioxide (3) oxygen (4) all of these

130. Barium chloride on reacting with ammonium sulphate forms barium sulphate and ammonium chloride. Which of the following correctly represents the type of the reaction involved?

(i) Displacement reaction
(ii) Precipitation reaction
(iii) Combination reaction
(iv) Double displacement reaction

(1) (i) only (2) (ii) only (3) (iv) only (4) (ii) and (iv)

131. __________is an exothermic reaction.

(1) Potassium nitrate dissolves in water
(2) Washing soda added to water
(3) Ammonium chloride added to water
(4) Sugar added to water

132. On thermal decomposition of $FeSO_4$, the gases evolved are

(1) SO_2 (2) SO_3 (3) O_2 (4) Both (1) and (2)

133. When ferrous sulphide is treated with dil. sulphuric acid, the products are :

(1) $FeSO_4$ and H_2S (2) FeS and H_2SO_4 (3) Fe_2SO_4 and H_2S (4) $Fe_2(SO_4)_3$ & H_2S

134. Three beakers labelled as A, B and C each containing 25 mL of water were taken. A small amount of NaOH, anhydrous $CuSO_4$ and NaCl were added to the beakers A, B and C respectively. It was observed after some time that there was an increase in the temperature of the solutions contained in beakers A and B, whereas in case of beaker C, the temperature of the solution falls. Which one of the following statement(s) is (are) correct?

(i) In beakers A and B, exothermic process has occurred.
(ii) In beakers A and B, endothermic process has occurred.
(iii) In beaker C exothermic process has occurred.
(iv) In beaker C endothermic process has occurred.

(1) (i) only (2) (ii) only (3) (i) and (iv) (4) (ii) and (iii)

135.

Column - I	**Column- II**
(i) Absorption of energy	(p) Physical change
(ii) Curdling of milk	(q) Exothermic reaction
(iii) Glowing of an electric bulb	(r) Chemical change
(iv) Releasing of energy	(s) Endothermic reaction

(1) (i)-(s), (ii)-(p), (iii)-(r), (iv)-(q)
(2) (i)-(s), (ii)-(r), (iii)-(p), (iv)-(q)
(3) (i)-(p), (ii)-(q), (iii)-(r), (iv)-(s)
(4) (i)-(s), (ii)-(r), (iii)-(q), (iv)-(p)

REDOX REACTION AND ITS APPLICATIONS

136. Which statement is incorrect for the reaction $CuO + H_2 \longrightarrow Cu + H_2O$:

(1) CuO is reduced
(2) H_2 is oxidized
(3) CuO is reduced and H_2 is oxidized
(4) Both CuO and H_2 are oxidized

137. In the reaction $2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$ which one is oxidized?

(1) Al (2) Fe (3) Fe_2O_3 (4) None

138. The process of reduction involves

(1) The gain of oxygen atoms
(2) The addition of hydrogen
(3) The loss of electrons
(4) Neither gain nor loss of electrons

139. Fatty foods become rancid because of which one of the following :

(1) Oxidation (2) Reduction (3) Hydrogenation (4) Corrosion

140. Rusting of iron is

(1) an oxidation process (2) A reduction process
(3) A redox process (4) A displacement process

141. Rancidity of fats and oils is caused by

(1) Aerial oxidation (2) Aerial reduction
(3) The presence of impurities (4) A displacement process

142. Which of the following element does not get corroded

(1) Au (2) Pt (3) Both (1) and (2) (4) Copper

143. Which of the following can prevent rancidity?

(1) Argon (2) Sulphur (3) Oxygen (4) Iodine

144. Which of the following is not a redox reaction:

(1) $SO_2 + 2H_2S \longrightarrow 3S + 2H_2O$ (2) $SO_2 + Cl_2 + 2H_2O \longrightarrow 2HCl + H_2SO_4$
(3) $2Mg + SO_2 \longrightarrow 2MgO + S$ (4) None of these

145. Which of the following is a redox reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$ (2) $H_2 + Cl_2 \longrightarrow 2HCl$
(3) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$ (4) $NaOH + HCl \longrightarrow NaCl + H_2O$

146. Which statement is correct about the following reaction? $ZnO + CO \longrightarrow Zn + CO_2$

(1) ZnO is being oxidized (2) CO is being reduced
(3) CO_2 is being oxidized (4) ZnO is being reduced

147. Which of the following statement is incorrect?

(1) In oxidation, oxygen is added to a substance.
(2) In reduction, hydrogen is added to a substance.
(3) Oxidizing agent is oxidized.
(4) Reducing agent is oxidized.

148. Rusting of iron is not observed with :

(1) ordinary water (2) damp air (3) boiled water (4) sea water

149. In the equation $PbO_2 + 4HCl \rightarrow PbCl_2 + 2H_2O + Cl_2$, the substance undergoing oxidation is_______.

(1) lead dioxide (2) hydrochloric acid (3) hydrogen (4) lead chloride

150. Which of the following statement is incorrect?

(1) In oxidation, oxygen is added to a substance.
(2) In reduction, hydrogen is added to a substance.
(3) Oxidizing agent is oxidized.
(4) Reducing agent is oxidized.

151. Which of the following is a redox reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$ (2) $H_2 + Cl_2 \longrightarrow 2HCl$
(3) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$ (4) $NaOH + HCl \longrightarrow NaCl + H_2O$

152. Which statement is correct about the following reaction? $ZnO + CO \longrightarrow Zn + CO_2$

(1) ZnO is being oxidized (2) CO is being reduced
(3) CO_2 is being oxidized (4) ZnO is being reduced

153. In reaction $SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$, the reducing agent is –

(1) SO_2 (2) H_2S (3) H_2O (4) S

154. The formula for rust is ________.

(1) CuO (2) $Fe_2O_3.xH_2O$ (3) Al_2O_3 (4) AgS

155. The main cause of rancidity in foods is

(1) Bacteria (2) Proteins

(3) Antioxidants (4) Oxidation of the fatty acid molecule

156. Which of the following gases can be used for storage of fresh sample of an oil for a long time?

(1) Carbon dioxide or oxygen (2) Nitrogen or oxygen

(3) Carbon dioxide or helium (4) Helium or nitrogen

157. Which of the following statements about the given reaction are correct?

$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$

(i) Iron metal is getting oxidised (ii) Water is getting reduced

(iii) Water is acting as reducing agent (iv) Water is acting as oxidising agent

(1) (i), (ii) and (iii) (2) (iii) and (iv) (3) (i), (ii) and (iv) (4) (ii) and (iv)

158. In the reaction, $2H_2S + SO_2 \rightarrow 3S + 2H_2O$

(1) H_2S has been oxidized (2) SO_2 has been oxidized

(3) H_2S is the oxidizing agent (4) SO_2 is the reducing agent

159. Which of the following does not corrode when exposed to the atmosphere ?

(1) Iron (2) Copper (3) Gold (4) Silver

160. Green coating on copper in rainy season is due to the formation of

(1) $CuCO_3$ (2) $Cu(OH)_2$ (3) $CuCO_3.Cu(OH)_2$ (4) CuS

161. Which of the following reaction is not a redox reaction as well as displacement reaction?

(1) $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$ (2) $ZnO + C \longrightarrow Zn + CO$

(3) $2Al + 6HCl \longrightarrow 2AlCl_3 + 3H_2$ (4) $H_2S + Cl_2 \longrightarrow 2HCl + S$

162. Assertion : In the reaction $Zn(s) + 2H^+ (aq) \rightarrow Zn^{2+} (aq) + H_2(g)$. Zn acts as an oxidising agent and H^+ acts as a reducing agent.

Reason : An oxidising agent accepts electrons while a reducing agent loses electrons.

(1) If both assertion and reason are CORRECT and reason is the CORRECT explanation of the assertion.

(2) If both assertion and reason are CORRECT, but reason is NOT THE CORRECT explanation of the assertion.

(3) If assertion is CORRECT, but reason is INCORRECT.

(4) If assertion is INCORRECT, but reason is CORRECT.

(5) If both assertion and reason are INCORRECT.

163. Which of the following reaction indicates the oxidising behavior of H_2SO_4 ?

(1) $2PCl_5 + H_2SO_4 \rightarrow 2POCl_3 + 2HCl + SO_2Cl_2$

(2) $NaCl + H_2SO_4 \rightarrow NaHSO_4 + HCl$

(3) $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

(4) $2HI + H_2SO_4 \rightarrow I_2 + SO_2 + H_2O$

164. When iron is added to $CuSO_4$ solution, copper is precipitated. It is due to

(1) Oxidation of Cu^{2+} (2) Reduction of Cu^{2+}

(3) Hydrolysis of $CuSO_4$ (4) Ionisation of $CuSO_4$

165. Which of the following characterizes the oxidation-reduction relationship?

(1) Element losing electrons is losing oxygen, element gaining electrons is gaining oxygen.

(2) Element gaining electrons is oxidized, element losing electrons is reduced.

(3) Element gaining electrons is losing hydrogen, element losing electrons is gaining hydrogen.

(4) Element losing electrons is oxidized, element gaining electrons is reduced.

166. When a chemical species loses one or more electrons, it is said to have been

(1) oxidised (2) reduced (3) remain unchanged (4) hydrolysed

167. Which of the following is a redox reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$ (2) $Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

(3) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$ (4) $NaOH + HCl \longrightarrow NaCl + H_2O$

168. Which of the following statements is incorrect?

(1) In oxidation, oxygen is added to a substance.

(2) In reduction, hydrogen is added to a substance.

(3) Oxidizing agent is oxidized.

(4) Reducing agent is oxidized.

169. A substance which oxidises itself and reduces other is known as

(1) oxidising agent (2) reducing agent (3) both (1) and (2) (4) none of these

170. In the reaction $PbO + C \longrightarrow Pb + CO$

(1) PbO is oxidised. (2) C acts as oxidising agent.

(3) C acts as reducing agent. (4) This reaction does not represent redox reaction.

171. Which of the following does not corrode when exposed to the atmosphere?

(1) Iron (2) Copper (3) Gold (4) Silver

172. Antioxidants are

(1) hydrating agents (2) dehydrating agents (3) oxidizing agents (4) reducing agents

173. In order to prevent the spoilage of potato chips, they are packed in plastic bags in an atmosphere of ______.

(1) chlorine (2) hydrogen (3) nitrogen (4) oxygen

174. Which of the following is a redox reaction?

(1) $CaCO_3 \longrightarrow CaO + CO_2$ (2) $H_2 + Cl_2 \longrightarrow 2HCl$

(3) $CaO + 2HCl \longrightarrow CaCl_2 + H_2O$ (4) $NaOH + HCl \longrightarrow NaCl + H_2O$

175. Which of the following gases can be used for storage of fresh sample of an oil for a long time?

(1) Carbon dioxide or oxygen (2) Nitrogen or helium

(3) Helium or oxygen (4) Nitrogen or oxygen

176. Which of the statements about the reaction below are incorrect?

$$2PbO(s) + C(s) \longrightarrow 2Pb(s) + CO_2(g)$$

(i) Lead is getting reduced (ii) Carbon dioxide is getting oxidised.

(iii) Carbon is getting oxidised. (iv) Lead oxide is getting reduced.

(1) (i) and (ii) (2) (i) and (iii) (3) (i), (ii) and (iii) (4) All of these

177. When the gases sulphur dioxide and hydrogen sulphide mix in the presence of water, the reaction is $SO_2 + 2H_2S \rightarrow 2H_2O + 3S$. Here, hydrogen sulphide is acting as

(1) an oxidising agent (2) a reducing agent (3) a dehydrating agent (4) a catalyst

178. In the reaction $2MnO_2 + 4Al \rightarrow 3Mn + 2Al_2O_3$ the oxidising agent is

(1) MnO_2 (2) Al (3) Al_2O_3 (4) Mn

179. In the equation, $NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$ nitric acid is acting as

(1) an oxidising agent (2) an acid (3) a reducing agent (4) a dehydrating agent

180. $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$. This is

(1) oxidation (2) reduction (3) redox reaction (4) none of these

181. Consider the reaction

$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

With reference to the above, which one of the following is the correct statement?

(1) Zn is reduced to Zn^{2+} (2) Zn is oxidised to Zn^{2+}.

(3) Zn^{2+} is oxidised to Zn (4) Cu^{2+} is oxidised to Cu.

182. In the reaction $PCl_3 + Cl_2 \rightarrow PCl_5$

(1) PCl_3 is acting as reductant.

(2) Cl_2 is acting as reductant.

(3) Both PCl_3 and Cl_2 are acting as reductant.

(4) Both PCl_3 and Cl_2 are acting as oxidant.

183. Check whether the reaction $H_2S + SO_2 \rightarrow H_2O + S$ is a redox reaction or not? If yes, what will be the oxidant?

(1) Yes, SO_2 (2) Yes, S (3) Yes, H_2S (4) No, this is not redox

184. Ozone is ________to oxygen in sunlight.

(1) reduced (2) oxidised (3) corroded (4) synthesised

185. Chemically rust is

(1) hydrated ferrous oxide (2) only ferric oxide

(3) hydrated ferric oxide (4) none of these

ANSWER KEY

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	1	2	4	2	1	3	1	4	1	3	1	4	4	2	3	1	1	3	1	1
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	3	2	3	2	1	4	3	2	4	3	3	4	1	4	2	4	2	3	4	3
Que.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	4	2	1	4	3	3	3	2	2	2	3	3	3	1	4	1	3	4	2	3
Que.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	4	1	3	4	1	1	4	3	4	4	1	3	4	1	4	3	2	3	1	2
Que.	81	82	83	84	85	86	87	88	89	90	91	92	93	94	95	96	97	98	99	100
Ans.	3	1	2	1	1	4	3	1	3	1	3	2	1	4	3	3	1	3	1	2
Que.	101	102	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118	119	120
Ans.	3	1	4	1	4	4	3	3	1	3	3	1	1	1	4	2	3	4	4	2
Que.	121	122	123	124	125	126	127	128	129	130	131	132	133	134	135	136	137	138	139	140
Ans.	1	1	2	4	4	1	3	2	4	4	2	4	1	3	2	3	1	2	1	3
Que.	141	142	143	144	145	146	147	148	149	150	151	152	153	154	155	156	157	158	159	160
Ans.	1	3	1	4	2	4	3	3	2	3	2	4	2	2	4	4	3	1	3	3
Que.	161	162	163	164	165	166	167	168	169	170	171	172	173	174	175	176	177	178	179	180
Ans.	1	4	4	2	4	1	2	3	2	3	3	4	3	2	2	1	2	1	2	2
Que.	181	182	183	184	185															
Ans.	2	1	1	1	3															

GENERAL CONCEPTS OF CHEMISTRY

LANGUAGE OF CHEMISTRY

Following terms are commonly used in chemistry.

1. Atom

Atom is the basic unit of matter that consists of a dense, central nucleus surrounded by a cloud of negatively charged electrons. The atomic nucleus contains a mixture of positively charged protons and electrically neutral neutrons (except in the case of hydrogen-1, which is the only stable nuclide with no neutrons). The electrons of an atom are bound to the nucleus by the electrostatic force.

2. Molecule

- A molecule is defined as an electrically neutral group of at least two atoms in a definite arrangement held together by very strong (covalent) chemical bonds.
- The smallest particle of an element, which may or may not have independent existence is called an atom, while the smallest particle of a substance which is capable of independent existence is called a molecule.
- Molecules are formed by combination of atoms by chemical bonding with each other.

Atom + Atom $\rightarrow$ Molecule

3. Atomicity

The number of atoms, which constitute one molecule of an element / a compound is called its atomicity.

- If the molecule is made up of atoms of the same element it is said to be homo-atomic. Examples: O_2, N_2
- If the molecule is made up of atoms of the different elements it is said to be hetero-atomic. Examples: CO_2, H_2O, SO_3, HCl

Classification of molecules:

(a) Monoatomic molecules:- Atomicity of these type of molecules is one.

Example: Helium, Neon, Argon, Krypton, Xenon

(b) Diatomic molecules: Two atoms are present in diatomic molecules. These two atoms are may be the combination of same **(homoatomic)** or different atoms **(heteroatomic).**

Example: For same combination of atom: homoatomic molecules.

H_2, Cl_2, O_2

Example: For different combination of atom: heteroatomic molecules.

HCl, CO, NaF

(c) Triatomic molecules: Three atoms are present in the triatomic molecules. Three atoms may be of same or different type.

Example: For same combination of atom: homoatomic molecules

O_3

Example: For different combination of atom: heteroatomic molecules

CO_2

(d) Tetratomic molecules: Four atoms are present in tetratomic molecules.

Example: For same combination of atom: homoatomic molecules

P_4

Examples: For different combination of atom: heteroatomic molecules PCl_3

NH_3

(e) Pentatomic molecules: Five atoms are present in pentatomic molecules.

Example: For pentatomic molecules:

CH_4, HNO_3, CCl_4

(f) Polyatomic molecules:

Example: S_8, C_{60}, C_2H_5OH

4. Element

It is defined as a substance that cannot be further reduced to simpler substances by ordinary processes. Elements are made up of particles/atoms of only one kind.

There are 118 elements known. Out of these 92 of them occur in nature and rest are synthetically made.

For example: Hydrogen and oxygen

5. Compound

It is a pure substance that can be decomposed into simpler substances by some suitable chemical techniques. A compound is formed by combination of two or more elements in a definite proportion. For example, water is a compound of hydrogen and oxygen elements present in the ratio of 1:8.

6. Mixture

A mixture is a material containing two or more elements or compounds that are in close contact and are mixed in any proportion. The components of a mixture can be separated by physical means.

For example, air, gun powder, etc.

Properties of a mixture :

(a) A mixture may be homogenous or heterogeneous. A homogenous mixture has a uniform composition throughout its mass. For example, sugar or salt dissolved in water, alcohol in water, etc. A heterogeneous mixture does not have a uniform composition throughout its mass. There are visible sharp boundaries.

For example: Oil and water, salt and sand, etc.

(b) The constituents of a mixture can be separated by physical means like filtration, evaporation, sublimation and magnetic separation.

(c) In the preparation of a mixture, energy is neither evolved nor absorbed.

(d) A mixture has no definite melting and boiling point.

(e) The constituents of a mixture retain their original set of properties. For example, sulphur dissolves in carbon disulphide and a magnet attracts iron filings.

7. Symbol

The symbol of an element is a short way of representing an element.

- Hydrogen was named by Lavoisier using the Greek words 'Hydro' ie., water and 'Genes' meaning forming.
- Chlorine has its origin from Greek word - 'Chloros' - meaning greenish yellow, 'Bromos' means stench in Greek and 'Iodes' in Greek means violet.
- The word nickel has originated from a German word meaning 'Saton' or 'Old nick'. Cobalt comes from a German word - 'Globlin' or 'Evil Spirit'.
- Argon from Greek meaning 'Inactive'.
- Gold is an Anglo-Saxon word - 'Aurum'.

Symbols can be formed as follows:

8. Valency

- Valency is the combining capacity of an atom. It is equal to the number of electrons the atom loses, gains or shares when it combines with one or more atoms.
- Depending on whether the atom loses, gains or shares electrons, the valency may be positive or negative or neutral respectively.

Example: Valency of sodium is +1, because it can lose one electron. It is represented as Na^+. Valency of oxygen may be –2 if it accepts 2 electrons i.e., O^{2-}.

- If two or more atoms combine by sharing electrons i.e., not losing or gaining electrons but sharing, then the valency numbers are neither positive nor negative. In the case of carbon dioxide, carbon and oxygen atoms, do not, lose or gain electrons, and their valencies are 4 and 2 respectively.
- Certain atoms exhibit different kinds of valencies. This is because such elements can loose more than one electron from their outer most shell depending on the available conditions. When such atoms exhibit variable valency, the name of the atom with the lower valency ends as **–ous**, and the one with the higher valency, ends as **–ic**.

Lower Valency		Higher Valency	
Cuprous	Cu^{+}	Cupric	Cu^{++}
Ferrous	Fe^{++}	Ferric	Fe^{+++}
Plumbous	Pb^{++}	Plumbic	Pb^{++++}
Mercurous	$Hg^{2+}{}_{2}$	Mercuric	Hg^{++}
Aurous	Au^{+}	Auric	Au^{+++}

- Atoms with variable valencies have their symbols same but are named differently.
- ***Atoms also combine in a group when the atoms of different elements combine in a group they are called radicals.*** They behave as a single unit and the valency of radicals is the overall net charge of the group. The radicals maintain their identity in the chemical changes but are incapable of independent existence.
- **Remember:**

  Radicals Formulae

  Hydroxide - OH^{-}

  Sulphate - SO_4^{2-}

  Nitrate - NO_3^{-}

List of Common Ions /Radicals with Positive Valency

Positive Valency 1	**Symbol**	**Positive Valency 3**	**Symbol**
1. Ammonium	NH_4^{+}	1. Aluminium	Al^{3+}
2. Hydrogen	H^{+}	2. Chromium	Cr^{3+}
3. Lithium	Li^{+}	3. Bismuth	Bi^{3+}
4. Sodium	Na^{+} .	4. Arsenic	As^{3+}
5. Potassium	K^{+}	5. Ferric [Iron (III)]	Fe^{3+}
6. Cuprous [Copper (I)]	Cu^{+}	6. Auric [Gold (III)]	Au^{3+}
7. Argentous [Silver (I)]	Ag^{+}		
8. Mercurous [Mercury (I)]	Hg^{+}		
9. Aurous [Gold (I)]	Au^{+}		

Positive Valency 2	**Symbol**	**Positive Valency 4**	**Symbol**
1.Magnesium	Mg^{2+}	1. Stannic [Tin (IV)]	Sn^{4+}
2.Calcium	Ca^{2+}	2. Plumbic (Lead (IV)]	Pb^{4+}
3. Zinc	Zn^{2+}	3. Platinic [Platinum (IV)]	Pt^{4+}
4. Barium	Ba^{2+}		
5. Nickel	Ni^{2+}		
6. Uranium	U^{2+}		
7. Cupric	[Copper (II)]		Cu^{2+}
8. Argentic [Silver (II)]	Ag^{2+}		
9. Mercuric [Mercury (II)]	Hg^{2+}		
10. Ferrous [Iron (II)]	Fe^{2+}		
11.Plumbous [Lead (II)]	Pb^{2+}		
12. Stannous [Tin (II)]	Sn^{2+}		
13. Platinous (Platinum (II)]	Pt^{2+}		

Radicals with Negative Valency

Negative Valency 1	Symbol	Negative Valency 2	Symbol
1. Fluoride	F^-	1. Sulphate	SO_4^{2-}
2. Chloride	Cl^-	2. Sulphite	SO_3^{2-}
3. Bromide	Br^-	3. Sulphide	S^{2-}
4. Iodide	I^-	4. Thiosulphate	$S_2O_3^{2-}$
5. Hypochlorite	ClO^-	5. Zincate	ZnO_2^{2-}
6. Chlorate	ClO_3^-	6. Plumbate	PbO_2^{2-}
7. Bicarbonate or hydrogen carbonate	HCO_3^-	7. Oxide	O^{2-}
8. Bisulphite or hydrogen sulphite	HSO_3^-	8. Peroxide	O_2^{2-}
9. Bisulphide or hydrogen sulphide	HS^-	9. Manganate	MnO_4^{2-}
10. Bisulphate or hydrogen sulphate	HSO_4^-	10. Dichromate	$Cr_2O_7^{2-}$
11. Hydride	H^-	11. Carbonate	CO_3^{2-}
12. Hydroxide	OH^-	12. Silicate	SiO_3^{2-}
13. Aluminate	AlO_2^-	13. Stannate	SnO_3^{2-}
14. Permanganate	MnO_4^-	14. Oxalate	$(COO)_2^{2-}$
15. Cyanide	CN^-		
16. Nitrite	NO_2^-		
17. Nitrate	NO_3^-		
18. Acetate	CH_3COO^-		

Negative Valency 3	Symbol	Negative Valency 4	Symbol
1. Nitride	N^{3-}	1. Carbide	C^{4-}
2. Phosphide	P^{3-}		
3. Phosphite	PO_3^{3-}		
4. Phosphate	PO_4^{3-}		

9. Ions

- By the loss or gain of electrons a neutral atom is changed to an ion. Ions are charged atoms or a group of atoms. In other words, ions are particles formed by atoms by the donation or acceptance of electrons.
- When an atom lose electrons it forms positive ion. Positively charged ions are called **cations**. On the other hand, when an atom gains an electron it acquires a negative charge. Such ions with a negative charge are known as **anions**.

An ion may be represented as follows:

Name of the element $\leftarrow [K]^{1+} \rightarrow$ Charge

Name of the element $\leftarrow [N]^{3-} \rightarrow$ Charge

Cation **Anion**

The charge on the cation indicates the number of electrons donated. In an anion it shows the number of electrons gained.

Differences between atoms and ions

Atom	Ion
An atom is a neutral particle [number of e^- = number of p^+]	An ion is a charged particle [number of e^- # number of p^+]
An atom is unstable.	An ion is stable. (in terms of octet)
An atom is incapable of independent existence.	An ion can exist independently.
An atom has an incomplete outermost shell.	An ion has a complete outermost shell.

The physical and chemical properties of an atom and ion of an element are quite different, as studies have proved.

10. Oxidation Number

In ionic substances, the redox reactions can be explained on the basis of electron transfer. However, the redox reactions of covalent compounds cannot be explained in terms of electron transfer. The term oxidation number explains the phenomenon of oxidation-reduction in covalent and ionic substances.

Oxidation number is defined as the charge, which an atom appears to have when all other atoms are removed from it as ions. The oxidation number is a fictitious charge in case of covalent species. It can have positive, zero or negative values depending upon their state of combination.

11. Names of compounds

(a) Binary compounds

Binary compounds are compounds containing only two elements. If one of the two is a metal (or ammonium), then the suffix - ide, is added to the non-metal atom.

Remember:

NaCl – Sodium chloride

CaO – Calcium oxide

AlN – Aluminium nitride

- If atoms of two elements form more than one type of compound, then, the prefix mono, di, tri, tetra, penta, etc. are used, to show the number of atoms in the compound.

Remember:

Dinitrogen monoxide – N_2O

Mononitrogen monoxide – NO

Dinitrogen trioxide – N_2O_3

Mononitrogen dioxide – NO_2

Dinitrogen pentoxide – N_2O_5

(b) Ternary compounds

Ternary compounds are compounds, which contain atoms of three elements. If one of them is a metal, it is used first and if oxygen is present along with another non-metal, then the suffix - ate is added to part of the name of the other non-metal.

Example: $KClO_3$ – Potassium chlorate.

- This suffix is used, if it is the only compound formed of the three elements. If there is one more compound formed by the same elements, but with a lesser number of oxygen, then instead of the suffix **ate**, suffix **ite** is used.

  Example: $KClO_2$ – Potassium chlorite.

- If a third compound is formed, with still lesser number of oxygen atoms, then, it is called - hypochlorite.

  Example:

KClO	–	Potassium hypochlorite.			
Chlorate	–	$ClO^-{}_3$	Chlorite	–	ClO_2^-
Nitrate	–	$NO^-{}_3$	Nitrite	–	NO_2^-
Sulphate	–	SO_4^{2-}	Sulphite	–	SO_3^{2-}
Phosphate	–	PO_4^{3-}	Phosphite	–	PO_3^{2-}
Carbonate	–	CO_3^{2-}	Silicate	–	SiO_3^{2-}
Oxalate	–	$(COO)_2^{2-}$	Acetate	–	CH_3COO^-
Aluminate	–	AlO_2^-	Zincate	–	ZnO_2^{2-}
Plumbate	–	PbO_2^{2-}	Chromate	–	CrO_4^{2-}
Manganate	–	MnO_4^{2-}			

12. Chemical equation

A chemical equation is a "balanced account of a chemical transaction." In any chemical transaction or reaction, the number of atoms of all the participating elements will remain proportionately constant before and after the reaction (The Law of Conservation of Mass).

- In a chemical equation, the formulae of the reactants and products are used. Reactants are substance(s) that undergo the chemical reaction. The products are the substances produced during the chemical reaction.
- The reactants and products are connected by an arrow (→). The arrow may be read as "to yield" or "to form" or "to give". The reactants are placed on the left side of the arrow and the products on the right side. The different reactants as well as products are connected by a plus sign (+).

13. S.I. UNITS (System International Units):

- **Seven Basic Units.** The seven basic physical quantities in the International System of Units, their symbols, the names of their units (called the base units) and the symbols of these units are given in Table.

TABLE: SEVEN BASIC PHYSICAL QUANTITIES AND THEIR S.I. UNITS

Physical	Quantity	Symbol	S.I.	Unit	Symbol
Length	ℓ	metre	m		
Mass	m	kilogram	kg		
Time	t	second	s		
Electric current	I	ampere	A		
Thermodynamic		temperature	T	kelvin	K
Luminous intensity	I_u	candela	cd		
Amount of the substance	n	mole mol			

- **Derived units.** The units of all other physical quantities are derived from those of above basic quantities. The units thus obtained are called the derived units. Some common physical quantities and their derived units are shown in Table.

TABLE : SOME COMMON PHYSICAL QUANTITIES AND THEIR DERIVED UNITS.

Physical Quantity	Description	Unit	Symbol
Area	Length square	square metre	m^2
Volume	Length cube	cubic metre	m^3
Density	Mass/unit volume	kilogram per cubic metre	$kg\ m^{-3}$
Velocity	Distance/unit time	metre per second	$m\ s^{-1}$
Acceleration	Speed change/unit time	metre per second square	$m\ s^{-2}$
Force	Mass × Acceleration	newton	$N = kg\ m\ s^{-2}$
Pressure	Force/unit area	pascal (newton per square metre)	$Pa = N\ m^{-2}$ $= kg\ m^{-1}\ s^{-2}$
Work, energy	Force × Distance	joule	$J = N\ m = kg\ m^2\ s^{-2}$
Frequency	Cycles/ second	hertz	$Hz = s^{-1}$
Electric charge	Current × Time	coulomb	$C = As$
Potential difference	—	volt	$V = kg\ m^2\ s^{-3}\ A^{-1}$ $= JA^{-1}\ s^{-1} = JC^{-1}$
Electric resistance	Reciprocal of resistance	ohm^{-1}	$\Omega^{-1} = A\ V^{-1}$
Electric conductance	Potential-difference/current	ohm	$\Omega = V\ A^{-1}$

- **Prefixes used with units.** The S.I. system recommends the multiples such as 10^3, 10^6, 10^9 etc. and fraction such as 10^{-3}, 10^{-6},10^{-9} etc., i.e. the powers are the multiples of 3. These are indicated by special prefixes. These along with some other fractions or multiples in common use, along with their prefixes are given below in Table and illustrated for length (m)

TABLE : SOME COMMONLY USED PREFIXES WITH THE BASE UNITS.

Prefix	Symbol	Multiplication Factor	Example
deci	d	10^{-1}	1 decimetre (dm) = 10^{-1} m
centi	c	10^{-2}	1 centimetre (cm) = 10^{-2} m
milli	m	10^{-3}	1 millimetre (mm) = 10^{-3} m
micro	μ	10^{-6}	1 micrometre (μm) = 10^{-6} m
nano	n	10^{-9}	1 nanometre (nm) = 10^{-9} m
pico	p	10^{-12}	1 picometre (pm) = 10^{-12} m
femto	f	10^{-15}	1 femtometre (fm) = 10^{-15} m
atto	a	10^{-18}	1 attometre (am) = 10^{-18} m
deka	da	10^1	1 decametre(dam) = 10^1 m
hecto	h	10^2	1 hectometre (hm) = 10^2 m
kilo	k	10^3	1 kilometre (km) = 10^3 m
mega	M	10^6	1 megamerte(Mm) = 10^6 m
giga	G	10^9	1 gigametre (Gm) = 10^9 m
tera	T	10^{12}	1 teramerte (Tm) = 10^{12} m
peta	P	10^{15}	1 petametre (Pm) = 10^{15} m
exa	E	10^{18}	1 exametre (Em) = 10^{18} m

As volume is very often expressed in litres, it is important to note that the equivalence in S.I. units for volume is as under: 1 litre (1L) = 1 dm^3 = 1000 cm^3 and 1 millilitre (1mL) = 1 cm^3 = 1cc

- **Some important unit conversions**

1. Length: 1 mile = 1760 yards

1 yard = 3 feet

1 foot = 12 inches

1 inch = 2.54 cm

1 Å = 10^{-10} m or 10^{-8} cm

2. Mass: 1 Ton = 1000 kg

1 Quintal = 100 kg

1 kg = 2.205 Pounds (lb)

1 kg = 1000 g

1 gram = 1000 milli gram

1 a.m.u. = 1.67×10^{-24} g

3. Volume: 1 litre = 1 dm^3 = 10^{-3} m^3 = 10^3 cm^3 = 10^3 mL = 10^3 cc

1 millilitre = 1 cm^3 = 10^{-6} m^3

= 1 cc

4. Energy: 1 calorie = 4.184 joules $\simeq$ 4.2 joules

1 joule = 10^7 ergs

1 litre atmosphere (L-atm) = 101.3 joule

1 electron volt (eV) = 1.602×10^{-19} joule

5. Pressure: 1 atmosphere (atm) = 760 torr

= 760 mm of Hg

= 76 cm of Hg

= 1.103×10^5 pascal (Pa)

= 1.103×10^5 N/m^2

- **Some important points to remember about S. I. units:**

1. The unit named after a scientist is started with a small letter and not with a capital letter e.g. unit of force is written as newton and not as Newton.

   Likewise unit of heat of work is written as joule and not as Joule.

2. Symbols of the units do not have a plural ending like 's'. For example we have 10 cm and not 10 cms.

3. Words and symbols should not be mixed e.g. we should write either joules per mole or J mol^{-1} and not joules mol^{-1}

4. Prefixes are used with the basic units e.g. kilometer means 1000 m (because meter is the basic unit).

   Exception. Though kilogram is the basic unit of mass, yet prefixes are used with gram because in kilogram, kilo is already a prefix.

5. A unit written with a prefix and a power is a power for the complete unit e.g. cm^3 means $(centimeter)^3$ and not centi $(meter)^3$.

LAWS OF CHEMICAL COMBINATION

(I) Law of mass conservation (or Law of indestructibility):

- It is given by **Lavoisier** and tested by **Landolt.**
- According to this law, the mass is not created or destroyed in a chemical reaction or physical reaction, but one form can change into another form.
- If the reactant is completely converted in products, then the sum of the mass of reactants is equal to the sum of the mass of products.

Total mass of reactants = Total mass of products.

If reactant are not completely consumed then the relationship will be:

Total mass of reactant = Total mass of product + Mass of unreacted reactant

Ex.

$$\underbrace{\underset{2}{H_2} + \underset{71}{Cl_2}}_{2 + 71 = 73\text{ g}} \longrightarrow \underbrace{\underset{2(1 + 35.5)}{2HCl}}_{73\text{ g}}$$

Special Point: The mass conservation law is not applicable in **nuclear reactions**.

(II) Law of definite proportion (or Law of constant composition):

- It is given by **Proust**.
- According to this law, a compound is obtained from different sources. But the ratio of each component (by weight) remain same. i.e. it does not depend on the method of its preperation or the source from which it has been obtained.
- For example, molecule of ammonia always has the formula NH_3. That is one molecule of ammonia always contains, one atom of nitrogen and three atoms of hydrogen or 17.0 g of NH_3 always contains 14.0 g of nitrogen and 3 g of hydrogen, no matter how it is obtained.

(III) Law of multiple proportion:

- It is given by **Dalton**.
- According to law of multiple proportion, if two elements combine to form more than one compound, then the different masses of one element which combine with a fixed mass of other element bear a simple ratio to one another.
- The following examples illustrate this law.

(i) Nitrogen and oxygen combine to form five oxides, which are : Nitrous oxide (N_2O), nitric oxide (NO), nitrogen trioxide (N_2O_3), nitrogen tetraoxide (N_2O_4) and nitrogen pentoxide (N_2O_5).

Weights of oxygen which combine with the fixed weight of nitrogen in these oxides are calculated as under:

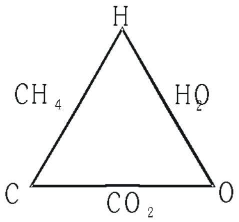
Oxide	Ratio of weights of nitrogen and oxygen in each compound
N_2O	28 : 16
NO	14 : 16
N_2O_3	28 : 48
N_2O_4	28 : 64
N_2O_5	28 : 80

Number of parts by weight of oxygen which combine with 14 parts by weight of nitrogen from the above are 8,16,24,32 and 40 respectively. Their ratio is 1 : 2 : 3 : 4 : 5, which is a simple ratio. Hence, the law is illustrated.

(IV) Law of reciprocal proportion (or Law of equivalent mass):

- It is given by **Richter**.
- The ratio of the weights of two elements A and B which combine separately with a fixed weight of the third element C is either the same or some simple multiple of the ratio of the weights in which A and B combine directly with each other.

- This law is illustrated with the help of the following examples :



(i) The elements C and O combine separately with the third element H to form

CH_4 and H_2O and they combine directly with each other to form CO_2 as shown in the adjoining figure. In CH_4, 12 parts by weight of carbon combine with 4 parts by weight of hydrogen. In H_2O, 2 parts by weight of hydrogen combine with 16 parts by weights of oxygen. Thus the weights of C and O which combine with fixed weight of hydrogen (say 4 parts of weight) are 12 and 32, i.e. they are in the ratio 12 : 32 or 3 : 8.

Now in CO_2, 12 parts by weight of carbon, combine directly with 32 parts by weight of oxygen, i.e. they combine directly in the ratio 12 : 32 or 3 : 8 which is the same as the first ratio.

(V) Law of gaseous volume:

- This law is given by **Gay Lussac**.
- According to this law, in the gaseous reaction, the reactants are always combined in a simple ratio by volume and form products, which is **simple ratio by volume** at same temperature and pressure.

Eg.1

$$\boxed{1 \qquad 1 \qquad 2} \longleftarrow \text{Stoichiometry}$$

$$H(g) + Cl(g) \longrightarrow 2\,HCl(g)$$

$$1 \text{ Volume} + 1 \text{ Volume} \qquad 2 \text{ Volume}$$

One volume of hydrogen combines with one volume of chlorine to produce 2 volumes of hydrogen chloride.

Simple ratio = 1 : 1 : 2.

Special Point: This law is used only for gaseous reaction. It relate volume to mole or molecules. But not relate with mass.

BERZELIUS & AVOGADRO'S HYPOTHESIS

(I) Berzelius hypothesis:

According to Berzelius, equal volume of all gases contain equal number of atoms at same temperature and pressure.

Ex.

$$\boxed{1 \qquad 1 \qquad 2} \longleftarrow \text{Stoichiometry}$$

$H_2(g)$	+	$Cl_2(g)$	$\longrightarrow$	$2HCl(g)$
1 Volume		1 Volume		2 Volume
N atoms		N atoms		2N atoms
$\boxed{\frac{1}{2}\text{atom}}$		$\frac{1}{2}$atom		1 atom

For the formation of one compound of hydrogen chloride, $\frac{1}{2}$ atom of chlorine and $\frac{1}{2}$ atom of hydrogen is required. These values are in against of Dalton's Atomic Theory.

- According to Dalton's theory, atom is not divisible. so Berzelius hypothesis was discarded.

(II) Avogadro's hypothesis:

According to this hypothesis, equal-volume of all gases contain equal number of molecules at same temperature and pressure.

Ex.

$$\boxed{1 \qquad 1 \qquad 2} \longleftarrow \text{Stoichiometry}$$

$H_2(g)$	+	$Cl_2(g)$	$\longrightarrow$	$2HCl(g)$
		1 Volume		1 Volume
		2 Volume		N molecules
		N molecules		2N molecules
$\boxed{\frac{1}{2}\text{molecule}}$ (1 atom)		$\frac{1}{2}$ molecule (1 atom)		1 molecule

It is correct as molecule is divisible.

- **Application of Avogadro's hypothesis:**

(I) Atomicity – Total number of atoms in a molecule of elementary substance is called as atomicity.

Ex.

Molecule	Atomicity
H_2	2
O_2	2
O_3	3
NH_3	4

(II) Relation between molecular weight and vapour density:

Vapour density (V.D.): Vapour density of a gas is the ratio of densities of gas & hydrogen at the same temperature & pressure.

$$\text{Vapour Density (V.D.)} = \frac{\text{Density of gas}}{\text{Density of hydrogen}} = \frac{d_{gas}}{d_{H_2}} \qquad \left\{ d = \frac{m(\text{mass})(g)}{V(\text{Volume})(mL)} \right.$$

$$\text{V.D.} = \frac{(m_{gas})\text{for certain V litre volume}}{(m_{H_2})\text{for certain V litre volume}}$$

- If N molecules are present in the given volume of a gas and hydrogen under similar condition of temperature and pressure.

$$\text{V.D.} = \frac{(m_{gas})\text{ of N molecules}}{(m_{H_2})\text{of N molecules}} = \frac{(m_{gas})\text{ of 1 molecule}}{(m_{H_2})\text{of 1 molecule}} = \frac{\text{Molecular mass of gas}}{2}$$

$\therefore$ Molecular mass of gas = 2 × V.D.

$$\boxed{\mathbf{M_w = 2V.D.}}$$

(III) Relation between molar mass (M_w) & volume:

$$\text{At STP. } M_{W.} = 2 \times \text{V.D.} = 2\times\frac{d_{gas}}{d_{H_2}} = 2 \times \frac{(m_{gas})\text{for certain V litre volume}}{(m_{H_2})\text{for certain V litre volume}}$$

or $$M_W = 2\times\frac{\text{mass of 1 litre gas}}{\text{mass of 1 litre } H_2} \qquad d_{H_2} = 0.000089\frac{g}{mL} = \frac{m}{V} = \frac{m}{1000mL}$$

or $$M_W = 2\times\frac{\text{Mass of 1 litre gas}}{0.089g} \qquad \Big| \quad V = 1 \text{ liter} = 1000 \text{ mL}$$

$$M_W\,(g) = 22.4 \times \text{mass of 1 litre gas} \qquad \text{then } m_{H_2} = 0.089g$$

$\boxed{\mathbf{M_W(g) = Mass\ of\ 22.4\ litre\ gas}}$ **or** $\boxed{\mathbf{M_w(g) \equiv 22.4\ litre\ (at\ STP)}}$

(IV) Gram Molecular Volume (GMV)

- At NTP, the volume of 1 mole of gaseous substance is 22.4 litre is called as gram molecular volume.
- At NTP, d_{H_2} = 0.000089 g/mL = mass/volume = mass/1000 mL

If volume = 1 litre = 1000 mL then mass = 0.089 g

$\because$ 0.089g H_2 occupies = 1 litre at STP

$\therefore$ 1 g H_2 occupies = $\frac{1\,\text{litre}}{0.089}$ at STP

$\therefore$ 2 g or 1 mol H_2 occupies = $\frac{1\,\text{litre}}{0.089} \times 2 = 22.4$ litre at STP

1 mole of any gaseous substance occupy 22.4 litre of volume at NTP or STP

$\boxed{\textbf{1 mol} \equiv \textbf{22.4 litre (at STP)}}$

MOLE CONCEPT

In SI Units we represent mole by the symbol 'mol'. It is defined as follows:

(i) A mole is the amount of a substance that contains as many entities (atoms, molecules or other particles) as there are atoms in exactly 0.012 kg (or 12g) of the carbon - 12 isotope.

- It may be emphasized that the mole of a substance always contains the same number of entities, no matter what the substance may be. In order to determine this number precisely, the mass of a carbon-12 atom was determined by a mass spectometer and found to be equal to 1.992648 × 10^{-23}g knowing that 1 mole of carbon weighs 12g, the number of atoms in it is equal to:

$$\frac{12\text{g/mol } C^{12}}{1.992648 \times 10^{-23}\text{g/ } C^{12}\text{ atom}} = 6.022137 \times 10^{23} \text{ atoms/mol}$$

(ii) In a simple way, we can say that mole has 6.022137 $\times 10^{23}$ entities (atom, molecules or ions etc.)

The number of entities in 1 mol is so important that it is given a separate name and symbol, known as '**Avogadro constant**' denoted by N_A in the honour of the nineteenth century Italian scientist, Amedeo Avogadro. Significant figures, in most of the calculations the value, 6.022 × 10^{23} entities/mol for N_A is generally used.

Note : 1 gram atom = mole atom 2 gram molecules = mole molecules

- **Different formula to get moles are following:**

(i) mole atoms = $\frac{\text{mass(g)}}{\text{atomic mass}}$ (ii) mole molecules = $\frac{\text{mass(g)}}{\text{molecular mass}}$

mole atoms = atomicity × mole molecules.

(iii) $n = \frac{V}{22.4}$ (V $\rightarrow$ Volume of gas in liter at NTP or STP)

(iv) mole particles $n = \frac{N}{N_A}$ (N $\rightarrow$ Number of particles)

mole atoms $= \frac{\text{number of atoms}}{N_A}$ and mole molecules $= \frac{\text{number of molecules}}{N_A}$

ATOMIC MASS, GRAM ATOMIC MASS, MOLECULAR MASS AND GRAM MOLECULAR MASS

(I) Atomic mass: It is defined as the number which indicates how many times the mass of one atom of an element is heavier in comparison to $\frac{1}{12}^{th}$ part of the mass of one atom of C-12.

- **Atomic mass unit (a.m.u.):** The quantity $\frac{1}{12}^{th}$ mass of an atom of C^{12} is known as atomic mass unit.

  Since mass of 1 atom of C - 12 = 1.9924×10^{-23} g

  $\therefore \frac{1}{12}^{th}$ part of the mass of 1 atom = $\frac{1.9924 \times 10^{-24} g}{12}$ = 1.67×10^{-24} g = 1 a.m.u. = $\frac{1}{6.023 \times 10^{23}}$

- It may be noted that the atomic masses as obtained above are the relative atomic masses and not the actual masses of the atoms. These masses on the atomic mass scale are expressed in terms of **atomic mass units** (abbreviated as amu). Today, 'amu' has been replaced by 'u' which is known as **unified mass**.

One atomic mass unit (amu) is equal to $\frac{1}{12}^{th}$ of the mass of an atom of carbon - 12 isotope. Thus the atomic mass of hydrogen is 1.008 amu while that of oxygen is 15.9994 amu (or taken as 16 amu).

(II) Gram atomic mass (or mass of 1 gram atom): When numerical value of atomic mass of an element is expressed in grams then the value becomes gram atomic mass.

gram atomic mass = mass of 1 **gram atom** = mass of 1 **mole atom**

= mass of N_A atoms = mass of 6.023×10^{23} atoms.

Ex. Gram atomic mass of oxygen = mass of 1 **g atom** of oxygen = mass of 1 **mol atom** of oxygen.

= mass of N_A atoms of oxygen. = $\left(\frac{16}{N_A} g\right) \times N_A$ = 16 g

Ex. Mass of one atom of oxygen = 16 amu or $16 \times 1.67 \times 10^{-24}$ g

Mass of N_A atoms of oxygen = $16 \times 1.67 \times 10^{-24} \times 6.023 \times 10^{-23}$ g = 16 g

(III) Molecular mass: The number which indicates how many times the mass of one molecule of a substance is heavier in comparison to $\frac{1}{12}^{th}$ part of the mass of an atom of C–12.

Ex. Molecular mass of oxygen (O_2) = 32 amu

Molecular mass of (O_3) = 48 amu

Molecular mass of HCl = 1 + 35.5 = 36.5 amu

Molecular mass of H_2SO_4 = 2 + 32 + 64 = 98 amu

(IV) Gram molecular mass (mass of 1 gram molecule): When numerical value of molecular mass of the substance is expressed in grams then the value becomes gram molecular mass.

Gram molecular mass = mass of 1 **gram molecule** = mass of 1 **mole molecule**

= mass of N_A molecules = mass of 6.023×10^{23} molecules

Ex. Gram molecular mass of H_2SO_4 = mass of 1 **gram molecule** of H_2SO_4

= mass of 1 **mole molecule** of H_2SO_4= mass of N_A molecules of H_2SO_4

= $\left(\frac{98}{N_A} g\right) \times N_A$ = 98 g

Ex. Molecular Mass of N_2 = 28 amu = $28 \times 1.67 \times 10^{-28}$ g

Mass of N_A molecules of N_2 = $28 \times 1.67 \times 10^{-24} \times 6.023 \times 10^{23}$ g = 28 g

Thus molecular mass can be defined as the absolute mass in grams of 6.023×10^{23} molecules of a substance.

(V) Actual Mass: The mass of one molecule of a substance in grams is known its actual mass.

For Ex. O_2 = 32 amu → relative mass

= $32 \times 1.67 \times 10^{-24}$ g → Actual mass

Actual mass of H_2O = 2 + 16 amu = $18 \times 1.67 \times 10^{-24}$g = 2.99×10^{-23} g

STOICHIOMETRY BASED CONCEPT (Problems based on chemical reactions)

One of the most important aspects of a chemical equation is that when it is written in the balanced form, it gives quantitative relationships between the various reactants and products in terms of moles, masses, molecules and volumes. This is called stoichiometry (Greek word, meaning 'to measure an element'). For example, a balanced chemical equation along with the quantitative information conveyed by it is given below:

$CaCO_3$	+	2HCl $\longrightarrow$	$CaCl_2$	+ H_2O	+ CO_2
1 Mole		2 Mole	1 Mole	1 Mole	1 Mole
40 + 12 + 3 × 16 = 100 g		2(1 + 35.5) = 73 g	40 + 2 × 35.5 = 111 g	2 × 1 + 16 = 18 g	12 + 2 × 16 = 44 g or 22.4 L at STP

Thus,

(i) 1 mole of calcium carbonate reacts with 2 moles of hydrochloric acid to give 1 mole of calcium chloride, 1 mole of water and 1 mole of carbon dioxide.

(ii) 100 g of calcium carbonate react with 73 g hydrochloric acid to give 111 g of calcium chloride, 18 g of water and 44 g (or 22.4 litres at STP) of carbon dioxide.

1	3		2 Stoichiometry
N_2 +	$3H_2$	→	$2NH_3$
1 mole +	3 mole	→	2 mole
22.4 litre +	3 × 22.4 litre	→	2 × 22.4 litre (at STP)
1 litre +	3 litre	→	2 litre
1000 mL +	3000 mL	→	2000 mL
1 mL +	3 mL	→	2 mL
28 gm +	6 gm	→	34 g (According to the law of conservation of mass)

*Gram can not be represented by stoichiometry

- The quantitative information conveyed by a chemical equation helps in a number of calculations. The problems involving these calculations may be classified into the following different types :-

(1) Mass - Mass relationships i.e. mass of one of the reactants or products is given and the mass of some other reactant or product is to be calculated.

(2) Mass - Volume relationships i.e. mass/volume of one of the reactants or products is given and the volume/mass of the other is to be calculated.

(3) Volume - Volume relationships i.e. volume of one of the reactants or the products is given and the volume of the other is to be calculated.

- The general method of calculations for all the problems of the above types consists of the following steps:-

(i) Write down the balanced chemical equation.

(ii) Write the relative number of moles or the relative masses (gram atomic or molecular masses) of the reactants and the products below their formula.

(iii) In case of a gaseous substance, write down 22.4 litres at STP below the formula in place of 1 mole

(iv) Apply unitary method to make the required calculations.

Quite often one of the reactants is present in larger amount than the other as required according to the balanced equation. The amount of the product formed then depends upon the reactant which has reacted completely. This reactant is called the limiting reactant. The excess of the other is left unreacted.

Combustion reaction: (Problem based on combustion reactions): For balancing the combustion reaction : First of all balance C atoms, Then balance H atom, Finally balance Oxygen atom.

- **For Example:** Combustion reaction of C_2H_6 : $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$ (skeleton equation)

balance C atoms $C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$

Now balance H atoms $C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$

Now balance oxygen atoms $C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$

Type I Involving Mass-Mass Relationship

Ex.1 How much iron can be theoretically obtained in the reduction of 1 kg of Fe_2O_3

Sol. Writing the balanced equation for the reaction.

$$\begin{array}{ccccccc} 1 & & 3 & & 2 & & 3 \\ Fe_2O_3 & + & 3H_2 & \longrightarrow & 2Fe & + & 3H_2O \end{array}$$

$$n = \frac{\text{Mass}}{M} = \frac{1000}{160}\text{mol}$$

The equation shows that 2 mol of iron are obtained from 1 mol of ferric oxide.

Hence, the obtained no. of moles of Fe = $\frac{2 \times 1000}{160}$ =12.5 mol = $\frac{\text{Mass}}{\text{Atomic Mass}} = \frac{\text{Mass}}{56}$

Mass of iron obtained = 12.5 × 56 g = 700 g

Ex.2 What amount of silver chloride is formed by the action of 5.850 g of sodium chloride on an excess of silver nitrate?

Sol. Writing the equation for the reaction

$$\begin{array}{ccccccc} 1 & & 1 & & 1 & & 1 \\ NaCl & + & AgNO_3 & \longrightarrow & AgCl & + & NaNO_3 \end{array}$$

$$n = \frac{\text{Mass}}{M} = \frac{5.85}{58.5} = 0.1\text{ mol}$$

1 mol of AgCl is obtained with 1 mol of NaCl

Hence, the number of moles of AgCl obtained with 0.1 mol of NaCl = 0.1 mol

$\because$ $n = \frac{\text{Mass}}{M}$ $\Rightarrow$ 0.1 mol $= \frac{\text{Mass}}{M} = \frac{\text{Mass}}{143.5}$ $\Rightarrow$ Mass = 0.1 × 143.5 g = 14.35 g.

Type II Mass - Volume Relationship

Ex.3 At 100°C for complete combustion of 3g ethane the required volume of O_2 & produced volume of CO_2 at STP will be.

Sol.

$$\begin{array}{ccccc} 2 & & 7 & 4 & 6 \\ 2C_2H_6 & + & 7O_{2(g)} & 4CO_{2(g)} & 6H_2O_{(g)} \end{array}$$

$n = \frac{\text{Mass}}{M} = \frac{3}{30} = \frac{1}{10} = 0.1$ mole

(a) Required moles of $O_2 = \frac{7}{2} \times 0.1$ = 0.35 mol

Volume of O_2 at STP = 0.35 × 22.4 = 7.84 litre

(b) Produced moles of $CO_2 = \frac{4}{2} \times 0.1 = 0.2$ mol

Volume of CO_2 at STP = 0.2 × 22.4 = 4.48 litre

Ex.4 In the following reaction, if 10 g of H_2, react with N_2. What will be volume of NH_3 at STP?

$N_2 + 3H_2 \longrightarrow 2NH_3$

Sol.

$$\begin{array}{ccccc} 1 & & 3 & & 2 \\ N_2 & + & 3H_2 & \longrightarrow & 2NH_3 \\ & & 10\text{ g} & & \end{array}$$

$$n = \frac{\text{Mass}}{M} = \frac{10}{2} = 5 \text{ mol.}$$

Produced moles of $NH_3 = \frac{2}{3} \times 5 = \frac{10}{3}$, Volume of NH_3 at STP $= \frac{10}{3} \times 22.4 = 74.67$ litre

Type III Volume-Volume Relationship

Ex.5 At 100°C for complete combustion of 1.12 litre of butane (C_4H_{10}), the produced volume of $H_2O(g)$ & $CO_2(g)$ at STP will be.

Sol.

$$\begin{array}{ccccccc} 1 & & 13/2 & & 4 & & 5 \\ C_4H_{10}(g) & + & \frac{13}{2}O_2(g) & \longrightarrow & 4CO_2(g) & + & 5H_2O(g) \\ 1.12\text{ litre} & & & & & & \end{array}$$

Volume of $H_2O(g)$ at STP $= 5 \times 1.12 = 5.6$ litre

Volume of $CO_2(g)$ at STP $= 4 \times 1.12 = 4.48$ litre

Ex.6 At 25°C for complete combustion of 5 mol propane (C_3H_8). What will be the required volume of O_2 at STP?

Sol. For C_3H_8, the combustion reaction is

$$\begin{array}{ccccccc} 1 & & 5 & & 3 & & 4 \\ C_3H_{8(g)} & + & 5O_{2(g} & \longrightarrow & 3CO_{2(g)} & + & 4H_2O_{(g)} \end{array} \qquad 5 \text{ mol}$$

Required moles of $O_2 = 5 \times 5 = 25 \text{ mol} = V/22.4$

Volume of O_2 gas at STP (V) $= 25 \times 22.4 = 560$ litre

Ex.7 3 litre of mixture of propane (C_3H_8) & butane (C_4H_{10}) on complete combustion gives 10 litre CO_2. Find the composition of mixture.

Sol. Let the volume of propane in the mixture = x litre,

∴ The volume of butane in the mixture = (3 – x) litre

Now let us calculate the volume of CO_2 evolved with the help of chemical equation.

Step I: Calculation of volume of CO_2 from x litre of propane.

$$\begin{array}{ccccccc} 1 & & 5 & & 3 & & 4 \\ C_3H_8 & + & 5O_2 & \longrightarrow & 3CO_2 & + & 4H_2O \\ x\text{ litre} & & & & 3x\text{ litre} & & \end{array}$$

Step II: Calculation of volume of CO_2 from (3 – x) litre of butane. The combustion equation for butane is:

$$\begin{array}{ccccccc} 1 & & 13/2 & & 4 & & 5 \\ C_4H_{10} & + & \frac{13}{2}O_2 & \longrightarrow & 4CO_2 & + & 5H_2O \\ (3-x)\text{ litre} & & & & 4(3-x)\text{ litre} & & \end{array}$$

Step III: Calculation of composition of the mixture.

Total volume of CO_2 formed in the step (I) and step (II) = [3x + 4(3 – x)] litre

But the volume of CO_2 actually formed = 10 litre

$3x + 4(3 - x) = 10$

or $3x + 12 - 4x = 10$ or $x = 2$ litre

∴ Volume of propane = x litre = 2 litre

∴ Volume of butane = (3 – x) litre = (3 – 2) = 1 litre

LIMITING REAGENT (L.R.) CONCEPT

Limiting Reagent (L.R.): The reactant which is completely consumed in a reaction is called as L.R.

Ex. 1	2		1	2	← Stoichiometry
A +	2B	⟶	C +	2D	
given 3 mol		9mol			
3 – 3 = 0 mol		9 – 6 = 3 mol		3 mol	6 mol

L.R. = A

Formula for checking L.R. = $\dfrac{\text{given value (moles, volume, or molecules)}}{\text{Stoichiometry Co-efficient}}$

Least value indicate the L.R.

Ex. A B

$3/1 = 3$ $9/2 = 4.5$

3 < 4.5 So A is L.R.

Identification : More than 1 initial quantities of reactants are given

Example Based on L.R.

Ex.1 $A + 5B \longrightarrow C + 3D$ In this reaction which is L.R.

Given 10 mol 10 mol

Sol. For A For B

$\frac{10}{1} = 10$ $\frac{10}{5} = 2$

2 < 10 So B is L.R.

Ex.2 Zinc and hydrochloric acid react according to the reaction

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

If 0.30 mol Zn are added to hydrochloric acid containing 0.52 mol of HCl, how many moles of H_2 are produced?

Sol. The reaction is

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Thus, 1 mol of zinc reacts with 2 moles of HCl.

$\therefore$ 0.30 mol of zinc will react with HCl = 2×0.30 = 0.60 mol

But we have only 0.52 mol of HCl. Hence, zinc cannot react completely and hence is not a limiting reactant.

Again, 2 moles of HCl react with zinc = 1 mol

$\therefore$ 0.52 mol of HCl will react with zinc

$= \frac{1}{2} \times 0.52 = 0.26$ mol

As we have 0.30 mol of zinc, therefore, HCl will react completely, i,e., HCl is the limiting reactant.

2 moles of HCl produce H_2 = 1 mol

$\therefore$ 0.52 mol of HCl will produce H_2

$= \frac{1}{2} \times 0.52 = 0.26$ mol

Ex.3 At NTP, In a container 100 mL N_2 and 100 mL of H_2 are mixed together. Then find out the produced volume of NH_3.

Sol. Balanced equation will be $N_2 + 3H_2 \longrightarrow 2NH_3$.

Given 100mL 100mL

For determination of Limiting reagent. Now divided the given quantities by stoichiometry coefficients

$$\frac{100}{1} = 100 \qquad \frac{100}{3} = 33.3 \text{ (Limiting reagent)}$$

In this reaction H_2 is limiting reagent so reaction will proceed according to H_2.

According to stoichiometry from 3 mL of H_2 produced volume of NH_3 = 2 mL

That is from 100 mL of H_2 produced volume of $NH_3 = \frac{2}{3} \times 100 = 66.6$ mL

SOLUTION

- When two or more chemically non-reacting substances are mixed and form homogeneous mixture it is called solution.
- When the solution is composed of only two chemical substances, it is termed a binary solution, similarly, it is called ternary and quaternary if it is composed of three and four components respectively.

  Solution = solute + solvent
- **Solute:-** Generally the component present in lesser amount than other component in solution is called solute.
- **Solvent:-** Generally, the component present in greater amount than any or all other components is called the solvent.
- Physical state of solvent and solution is same.
- Example:- In a syrup (liquid solution) containing 60% sugar (a solid) and 40% water (a liquid - same aggregation as solution), water is termed as the solvent and sugar as solute.
- **Dilute Solution**

  A solution in which relatively a small amount of solute is dissolved in large amount of solvent is called a dilute solution.
- **Concentrated solution**

  A solution in which relatively a large amount of the solute is present is called a concentrated solution.
- **Saturated solution**

  The maximum amount of solute in grams, that can be dissolved in 100 gm of a solvent at a particular temperature is called solubility of the solute and such a solution in which no more solute can be dissolved at a particular temperature is called saturated solution.
- **Supersaturated solution**

  A solution containing more amount of solute than that required for saturation of a given amount of solvent at a particular temperature, is called a supersaturated solution.

  It is an unstable system.
- **Types of solution**

  There may be the following seven types of binary solutions –

S.No.	Solute	Solvent	Example
1.	Gas	Gas	Air
2.	Gas	liquid	Oxygen in water
3.	Gas	solid	Hydrogen in Pd/pt.
4.	Liquid	liquid	Alcohol in water, benzene in toluene
5.	Liquid	solid	Mercury in zinc amalgam
6.	Solid	liquid	Sugar in water
7.	solid	solid	Alloys

Special point:- The solution of liquid in gas or solid in gas is not possible because the constituents can not form a homogeneous mixture.

Properties of a solution:

(i) A solution consists of a single phase i.e. it is a monophasic system.

(ii) A solution is uniform throughout, so it has uniform properties such as density, refractive index etc.

(iii) Size of solute particles in a solution is of the order $10^{-7} - 10^{-8}$ cm.

(iv) The components of a solution can not be easily separated by physical methods.

(v) The properties of a solution are the properties of its components. i.e. the components do not lose their properties when they form a solution.

(vi) The composition of a solution is not definite but can vary within certain limits.

(vii) Certain properties of solution such as density, viscosity, surface tension, boiling point, freezing point etc. vary with the composition of the solution.

21. EXPRESSING CONCENTRATION OF A SOLUTION:

(i) Molarity (M):- The number of moles of solute dissolved in one litre solution is called its molarity.

$$\text{Molarity} = \frac{\text{Number of moles of solute}}{\text{volume of solution in litres}} = \frac{n}{V}$$

$$\text{weight of solute in gram} = \frac{\text{Molarity} \times \text{Volume of solution in mL} \times \text{molecular weight}}{1000}$$

$$\text{Molarity} = \frac{\text{Number of moles of solute} \times 1000}{\text{Volume of solution in mL}} = \frac{\text{weight of solute in grams} \times 1000}{\text{Molecular weight} \times \text{Volume of solution in mL}}$$

$$= \frac{\text{strength of solution in gram / litre}}{\text{Molecular weight of solute.}} = \text{Normality} \times \frac{\text{Equivalent weight of solute}}{\text{Molecular weight of solute}}$$

(ii) Normality (N):- The number of equivalents or gram equivalents of solute dissolved in one litre of the solution is known as normality (N) of the solution.

$$\text{Normality(N)} = \frac{\text{Number of gram equivalents of solute}}{\text{volume of solution in litre}}$$

$$= \frac{\text{weight of solute in gram}}{\text{equivalent weight} \times \text{volume of solution (litre)}}$$

$$= \frac{\text{strength of solution in gram / litre}}{\text{Equivalent weight of solute}}$$

Equivalent weight of a substance is that weight which reacts with or displaces one gram of hydrogen, 8 grams of oxygen or 35.5 grams of chlorine.

(iii) Molality (m):- The number of moles or gram molecules of solute dissolve in 1000 gram of the solvent is called molality of the solution.

$$\text{Molality of a solution} = \frac{\text{Number of moles of solute}}{\text{Amount of solvent in kg.}}$$

$$= \frac{\text{Number of moles of solute} \times 1000}{\text{Amount of solvent in grams.}}$$

It is independent of temperature.

(iv) Formality (F):- It is the number of formula mass in grams present per litre of solution.

$$\text{formality} = \frac{\text{weight of solute in gram}}{\text{formula mass of solute in grams} \times \text{volume of solution in litre}}$$

- In case formula mass is equal to molecular mass, formality is equal to molarity.
- Depends on temperature.

(v). Concentration or strength of a solution (S):- The numbers of grams of solute dissolved in one litre solution is known as its strength in grams per litre.

- Strength in grams per litre

$$S = \frac{\text{weight of solute in grams} \times 1000}{\text{volume of solution in mL}}$$

S = Molarity of solution × molecular weight of solute.

S = Normality of solution × equivalent weight of solute.

(vi) Percentage by weight:- The number of grams of solute is dissolved in one gram of solution is called weight fraction of the solute. Thus,

$$\text{weight fraction} = \frac{w}{w+W}$$

Where 'w' grams of solute is dissolved in W grams of solvent.

$$\text{weight percent} = \frac{\text{weight of solute in grams} \times 100}{\text{weight of solution in grams.}} = \frac{w \times 100}{w+W}$$

(vii). Percent by volume (Volume fraction):- This method is used for solutions of liquid in a liquid. The volume of liquid (solute) in ml dissolved in one ml. of solution is called volume fraction.

$$\text{Volume fraction} = \frac{\text{Volume of liquid solute in mL}}{\text{volume of solution in mL}}$$

$$\text{Volume percent} = \frac{\text{Volume of solute} \times 100}{\text{Volume of solution}}$$

(viii) Parts per million (ppm): This method is used for expressing the concentration of very dilute solutions such as hardness of water, air pollution etc.

$$\text{ppm of substance} = \frac{\text{Mass of solute in gms} \times 10^6}{\text{Mass of solution}} = \frac{\text{Volume of solute} \times 10^6}{\text{volume of solution}}$$

(ix) Mole fraction:- The ratio of the number of moles of one component to the total number of all the components present in the solution, is called the mole fraction of that component.

Mole fraction of solute X_A is given by

$$X_A = \frac{n_A}{n_A + n_B}$$

Mole fraction of solvent X_B is given by $X_B = \dfrac{n_B}{n_A + n_B}$

where n_A is moles of solute A and n_B is moles of solvent B $\boxed{X_A + X_B = 1}$

(x) Relation between molarity and normality:-

S = Molarity × molecular weight of solute

and S = Normality × equivalent weight of solute.

So we can write

Molarity × Molecular weight of solute = Normality × equivalent weight of solute.

$$\text{Normality} = \frac{\text{molarity} \times \text{molecular weight of solute}}{\text{equivalent weight of solute}} = \frac{\text{molarity} \times \text{molecular weight of solute}}{\text{(molecular weight of solute/valency factor)}}$$

Normality = molarity × valency factor

$N = M \times n$; $N \geq M$

MOLECULAR FORMULA & EMPIRICAL FORMULA

(I) Molecular formula: The molecular formula of a compound represent the **actual number of atoms** present in 1 molecule of the compound i.e. it shows the real formula of its 1 molecule.

(II) Empirical Formula: The empirical formula of a compound express the **simplest whole number ratio of atoms** of various elements present in 1 molecule of the compound.

Ex.

Molecular Formula $\rightarrow$	H_2O_2	CH_4	C_2H_6	$C_2H_4O_2$
	2 : 2	1 : 4	2 : 6	2 : 4 : 2
	1 : 1		1 : 3	1 : 2 : 1
Empirical Formula $\rightarrow$	$\boxed{HO}$	$\boxed{CH_4}$	$\boxed{CH_3}$	$\boxed{CH_2O}$

Relationship b/w Empirical & Molecular Formula:

Molecular Formula = n × Empirical Formula

[Where n = natural no. (1, 2, 3,.........)]

or $n = \dfrac{\text{Molecular Formula}}{\text{Empirical Formula}}$ or $n = \dfrac{\text{Molecular mass}}{\text{Empirical mass}}$

Determination of empirical formula–

Following steps are involved to determine the empirical formula of the compounds–

(i) First of all find the % by weight of each element present in 1 molecule of the compound

(ii) The % by weight of each element is divided by its atomic weight. It gives atomic ratio of elements present in the compounds.

(iii) Atomic ratio of each element is divided by the minimum value of atomic ratio as to get simplest ratio of atoms.

(iv) If the value of simplest atomic ratio is fractional then raise the value to the nearest whole number.or Multiply with suitable coefficient to convert it into nearest whole number.

(v) Write the Empirical formula as we get the simplest ratio of atoms.

Example based on molecular and empirical formula

Ex.1 In a compound x is 75.8% and y is 24.2% by weight present. If atomic weight of x and y are 24 and 16 respectively. Then calculate the empirical formula of the compound.

Sol.

Elements	%	Atomic weight	$\dfrac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
x	75.8%	24	$\dfrac{75.8}{24} = 3.1$	$\dfrac{3.1}{1.5} = 2$	2
y	24.2%	16	$\dfrac{24.2}{16} = 1.5$	$\dfrac{1.5}{1.5} = 1$	1

Empirical formula = x_2y

Ex.2 In a compound Carbon = 52.2%, Hydrogen = 13%, Oxygen = 34.8% are present and molecular mass of the compound is 92. Calculate molecular formula of the compound ?

Sol.

Elements	%	Atomic weight	$\dfrac{\%}{\text{Atomic weight}}$	Simplest ratio	Ratio
C	52.2	12	$\dfrac{52.2}{12} = 4.35 = 4.4$	$\dfrac{4.4}{2.2} = 2$	2
H	13	1	$\dfrac{13}{1} = 13$	$\dfrac{13}{2.2} = 5.9$	6
O	24.8	16	$\dfrac{34.8}{16} = 2.2$	$\dfrac{2.2}{2.2} = 1$	1

Empirical formula = C_2H_6O Empirical mass = 12 × 2 + 16 + 6 = 46

$n = \dfrac{\text{Molecular mass}}{\text{Empirical mass}} = \dfrac{92}{46} = 2$ molecular formula = $2 \times (C_2H_6O) = C_4H_{12}O_2$

EQUIVALENT WEIGHT

The equivalent weight of a substance is the number of parts by weight of the substance that combine with or displace directly or indirectly 1.008 parts by weight of hydrogen or 8 parts by weight of oxygen or 35.5 parts by weight of chlorine or 108 parts by weight of Ag.

Calculation of equivalent weight:

(i) Equivalent weight $= \dfrac{\text{Atomic weight}}{\text{Valency factor}}$

(ii) Equivalent weight of ions $= \dfrac{\text{formula weight of ion}}{\text{Valency}}$

(iii) Equivalent weight of ionic compound = equivalent weight of cation + equivalent weight of anion

Ex. Equivalent weight of H_2SO_4 = Equivalent weight of H^+ + Equivalent weight of Anion(SO_4^{-2})

$= 1 + 48 = 49$

(iv) Equivalent weight of acid / base $= \dfrac{\text{Molecular weight}}{\text{Basicity/Acidicity}}$

(v) Equivalent weight of salt $= \dfrac{\text{Molecular weight}}{\text{Total charge on cation or anion}}$

Ex. Na_2SO_4 (salt) i.e. $2Na^+$ & SO_4^{-2}

Total charge on cation or anion is 2

molecular weight of Na_2SO_4 is $= (2 \times 23 + 32 + 16 \times 4) = 142$

Equivalent weight of $Na_2SO_4 = \dfrac{142}{2} = 71$

(vi) Equivalent weight of an oxidizing or reducing agent $= \dfrac{\text{Molecular weight of the substance}}{\text{Number of electrons gain/lost by one molecule}}$

Methods for determination of molecular weight:

(i) Molecular weight $= 2 \times$ vapour density

(ii) Diffusion method (only for gases):- According to Graham's diffusion law

$$\text{rate of diffusion of a gas} \propto \frac{1}{\sqrt{\text{Molecular weight or density}}} \quad \therefore \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

Miscellaneous Examples

Question 1: 25 mL of $\dfrac{N}{10}$ caustic soda solution exactly neutralizes 20 mL of an acid solution containing 7.875 g of acid per litre. Calculate the equivalent weight of the acid.

Ans. At the end point ,

Acid Base

$N_1V_1 = N_2V_2$ Where N_1 = Normality of acid = ?

N_2 = Normality of base $= \dfrac{N}{10}$

V_1 = Volume of acid = 20 mL

V_2 = Volume of base = 25 mL

Thus substituting these values in the above equation,

$N_1 \times 20 \text{ mL} = \dfrac{N}{10} \times 25$ $\qquad N_1 = \dfrac{N}{10} \times \dfrac{25}{20} = \dfrac{N}{8} = 0.125\text{ N}$

Calculation of equivalent weight of acid -

Strength = Normality $\times$ Equivalent weight Equivalent weight $= \dfrac{\text{Strength}}{\text{Normality}} = \dfrac{7.875}{0.125} = 63$

Question 2: 150 mL of $\frac{N}{10}$ HCl is required to react completely with 1 gram of a sample of limestone. Calculate the percentage purity of limestone.

Ans. $CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2$

100 gram   2 gram-equivalents

50 gram   1 gram-equivalents

$\therefore$ Equivalent weight of $CaCO_3$ = 50

150 mL of $\frac{N}{10}$ HCl $\equiv$ 150 mL of $\frac{N}{10}$ $CaCO_3$

$$\therefore \text{Strength} = \frac{\text{Normality} \times \text{Equivalent weight} \times \text{Volume in cc}}{1000} = \frac{1}{10} \times 50 \times \frac{150}{1000} = 0.75$$

1 gram of limestone contains 0.75 gram of $CaCO_3$ in it. Thus the sample is **75% pure**.

Question 3: 5 g of a mixture of NaCl and Na_2CO_3 were dissolved in water and volume made upto 250 mL. 15 mL of this solution required 50 mL of $\frac{N}{10}$ HCl for complete neutralisation. Calculate the percentage composition of original mixture.

Ans. In the mixture only Na_2CO_3 reacts with HCl. Thus,

50 mL of $\frac{N}{10}$ HCl $\equiv$ 50 mL of $\frac{N}{10}$ Na_2CO_3

50 mL of $\frac{N}{10}$ HCl $\equiv$ 15 mL of $\frac{N}{3}$ Na_2CO_3

$$\text{Strength of } Na_2CO_3 = \frac{\text{Normality} \times \text{equivalent weight} \times \text{volume}}{1000} = \frac{1}{3} \times \frac{53 \times 250}{1000} = 4.416 \text{ g}$$

% of Na_2CO_3 in the mixture = $100 \times \frac{4.416}{5}$ = **88.32 %**

% of NaCl in the mixture = 100 − 88.32 = **11.68 %**

Question 4: 25 mL of sodium carbonate solution are mixed with 8 mL of 0.75N HCl and for complete neutralisation it further required 15 mL of 0.8 N H_2SO_4. Find the strength of the given alkali solution in terms of normality and in gL^{-1}.

Ans. 8 mL of 0.75N HCl $\equiv$ 6 mL of 1 N HCl

15 mL of 0.8 N H_2SO_4 $\equiv$ 12 mL of 1 N H_2SO_4

Total volume of acid used = 6 + 12 = 18 mL of 1 N acid

$N_1V_1 = N_2V_2$

$N_1 \times 25 = 1N \times 18 \qquad N_1 = \frac{1N \times 18}{25} = 0.72$

Strength of sodium carbonate solution = normality $\times$ equivalent weight = **0.72 × 53 = 38.16 g L^{-1}**

Question 5: What volume of 6M HCl and 2 M HCl should be mixed to get one litre of 3M HCl?

Ans. Suppose the volume of 6M HCl required to obtain 1 L of 3 M HCl = xL

$\therefore$ Volume of 2M HCl required = (1 − x) L

Applying the molarity equation

$M_1V_1 + M_2V_2 = M_3V_3$

6M HCl + 2 M HCl = 3 M HCl

$6x + 2(1 - x) = 3 \times 1 \quad 6x + 2 - 2x = 3 \quad 4x = 1$

x = 0.25 L

Hence, volume of 6 M HCl required = 0.25 L, Volume of 2 M HCl required = 0.75 L

Question 6: To neutralize 200mL of $Ba(OH)_2$ solution, we have to add 100 mL of 0.1 M H_2SO_4 solution. What is the normality of Ba (OH_2)?

(1) 0.05 (2) 0.01 (3) 1.0 (4) 0.1

Ans. (0.1 M H_2SO_4 = 0.2 NH_2SO_4)

$[H_2SO_4]\ N_1V_1 = N_2V_2\ [Ba\ (OH)_2]$

$0.2 \times 100 = N_2 \times 200$

$0.1 = N_2$

Question 7: The concentration of "10 volume H_2O_2" is -

(1) 30 % (2) 3% (3) 1% (4) 10 %

Ans. 10 volume H_2O_2 means -

1 c.c. of the solution gives 10 c.c. of oxygen i.e. 100 c.c. of the solution gives 1000 c.c. of oxygen.

$$H_2O_2 \rightarrow H_2O + \frac{1}{2} O_2$$

34 g 11,200c.c oxygen

$\therefore$ 11, 200 c.c. of oxygen is obtained from = 34 g H_2O_2

$\therefore$ 1000 c.c. of oxygen is obtained from $= \frac{34}{11200} \times 1000$ g H_2O_2

$\therefore$ Concentration is 3% = 3 g H_2O_2

Question 8: Calculate the volume of N/100 oxalic acid solution obtained by dissolving 126g of oxalic acid.

(1) 100 litres (2) 300 litres (3) 200 litres (4) 250 litres

Ans. Equivalent weight of oxalic acid = 63 g

0.63 g oxalic acid is needed to prepare

N/100 solution of 1 litre

$\therefore$ 126 g oxalic acid is needed to prepare

$\frac{N}{100}$ solution of volume $\frac{1}{0.63} \times 126 = 200$ litres.

Question 9: 100 cc of 1.2 N acid should be diluted to .....in order to prepare its one normal solution

(1) 120 c.c. (2) 200 c.c. (3) 240 c.c. (4) 360 c.c.

Ans. $N_1V_1 = N_2V_2$

$1.2N \times 100 = 1N \times V_2$

$V_2 = 120$ c.c

Question 10: What is the volume of water that should be added to 150 mL of $\frac{N}{2}$ oxalic acid to prepare a solution of $\frac{N}{10}$ oxalic acid?

(1) 750 c.c. (2) 400 c.c. (3) 800 c.c. (4) 600 c.c.

Ans. $N_1V_1 = N_2V_2$

$$\frac{N}{2} \times 150 = \frac{N}{10} \times V_2$$

$750 = V_2$

Hence the volume of water added would be $= 750 - 150 = 600$ mL

Question 11: A sample of calcium carbonate is 80% pure. 25 g of this sample is treated with excess of HCl. How much volume of CO_2 will be obtained at N.T.P.?

(1) 4.48 litre (2) 5.6 litre (3) 11.2 litre (4) 2.24 litre

Ans. 25 g of sample contains = $\frac{80}{100} \times 25g$ of $CaCO_3$ = 20g

The eq^n. of the reaction is :

$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

100 g 22.4 litre

Hence the volume of CO_2 at N.T.P. would be $= \frac{22.4}{100} \times 20 = 4.48$ litre

Question 12: When excess of $CaCO_3$ is treated with 100 mL of HCl solution, the CO_2 gas obtained was found to be 1.12 litre (at N.T.P.). What is the normality of HCl?

(1) 0.2 N (2) 1N (3) 0.1N (4) 2N

Ans. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

2 mole 1 mole

2 equivalent 1 equivalent

2 equivalent 22.4 litre

1 equivalent 11.2 litre

0.1 equivalent 1.12 litre

Now using the formula

Normality = $\frac{w}{E} \times \frac{1000}{V(ml.)}$ = $0.1 \times \frac{1000}{100}$ = 1

Hence normality = 1 N

CHEMICAL ANALYSIS

It is of two types

(i) Qualitative analysis: Qualitative analysis is one in which we determine the kind of elements present in given sample of compound.

(ii) Quantitative analysis: Quantitative analysis is one in which the amount or concentration of a particular species in a sample is determined accurately. This is done in two ways-

(a) Gravimetric analysis **(b)** Volumetric analysis

- The Gravimetric analysis involves the measurement of the weight of the substance present in a solution, which is done by precipitation of the substance from the solution.
- **Volumetric analysis** involves the measurement of the volume of a known solution required to bring about the completion of the reaction with a measured volume of the unknown solution.

  The process of addition of the known solution from the burette to the measured volume of solution of the substance until the reaction between the two is just complete is called as **titration**.
- The titration mainly involves two solutions.

1. **Unknown solution** - The solution whose concentration is unknown and which is to be determined is called as unknown solution. This is called as **titrate**.

2. **Standard solution** - The solution in which an accurate amount of the reagent (titrant) has been dissolved in the known volume of the solution is called as **standard solution**. Thus titrant is of two types-

(a) Primary standard - The compounds which exist in solid state, which neither are deliquescent nor efflorescent in nature, they can be accurately weighed and their concentrations are exactly known, are called as primary standards. Important primary standards are crystalline oxalic acid, potassium dichromate, silver nitrate, copper sulphate, ferrous ammonium sulphate, sodium carbonate, NaCl, KCl and sodium thiosulphate etc.

(b) Secondary standard - These substances can not be weighed accurately. The solutions of these substances are to be titrated before use to determine their exact concentrations. The examples of secondary standards are NaOH, KOH hydrochloric acid, sulphuric acid, potassium permanganate, Iodine etc.

- **Law of equivalence** -

This law is applied to all the titrations, According to this law, the chemical substances react in the ratio of their chemical equivalent weights.

i.e, $\frac{\text{Mass of substance A}}{\text{Mass of substance B}} = \frac{\text{Chemical equivalent weight of A}}{\text{Chemical equivalent weight of B}}$

or $\frac{\text{Mass of substance A}}{\text{Chemical equivalent weight of A}} = \frac{\text{Mass of substance B}}{\text{Chemical equivalent weight of B}}$

or Gram equivalents of A = Gram equivalent of B

or Milligram equivalents of A = Milli gram equivalents of B

- **Equivalence point or End point** - The point at which the amounts of two reactants are just equivalent is called as end point of the titration.

- **Indicator**

The substance which helps in the detection of the end point or equivalence point of the titration is called as indicator. Those substances which undergo detectable change at the end point are used as indicators.

ACID-BASE TITRATIONS

Titration is a procedure for determining the concentration of a solution using another solution of known concentration, called standard solution. We will consider three types of reactions: (1) titrations involving a strong acid and a strong base. (2) titrations involving a weak acid and a strong base, and (3) titrations involving a strong acid and a weak base. Titrations involving a weak acid and a weak base are complicated by the hydrolysis of both the cation and the anion of the salt formed. For this reason, they are almost never carried out.

(i) STRONG ACID-STRONG BASE TITRATIONS

The reaction between HCl, a strong acid, and NaOH, a strong base, can be represented by

$$NaOH(aq) + HCl(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

or, in terms of the net ionic equation.

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Consider the addition of a 0.10 M NaOH solution (from a buret) to an Erlenmeyer flask containing 25 mL of 0.10 M HCI. Figure-1 shows the pH profile of the titration (also known as the titration curve). Before the addition of NaOH, the pH of the acid is given by –log (0.10). or 1.00. When NaOH is added, the pH of the solution in-creases slowly at first. Near the equivalence point the pH begins to rise steeply, and at the equivalence point (that is, the point at which equimolar amounts of acid and base have reacted) the curve rises almost vertically. In a strong acid-strong base titration, both the hydrogen ion and hydroxide ion concentrations are very small at the equivalence point (approximately 1×10^{-7} M); consequently, the addition of a single drop of the base can cause a large increase in $[OH^-]$ and in the pH of the solution. Beyond the equivalence point, the pH again increases slowly with the addition of NaOH. It is possible to calculate the pH of a solution at every stage of titration. Here are three sample calculations.

1. After the addition of10 mL of 0.10 M NaOH to 25 mL of 0.10 M HCI. The total volume of the solution is 35 mL. The number of moles of NaOH in 10 mL is

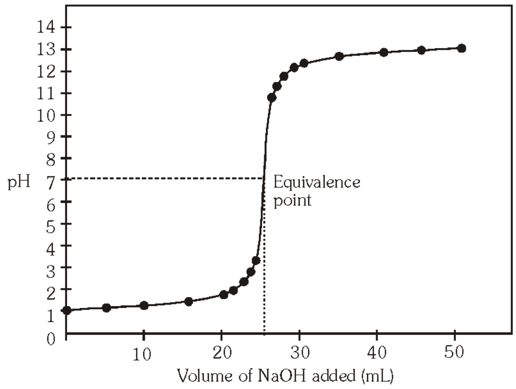


Figure 1: pH profile of a strong acid-strong base titration A 0.10 M NaOH solution is added from a buret to 25 ml of a 0.10 M HCl solution in on Erlenmeyer flask (see Figure). This curve is sometimes referred to as a titration curve.

$$10 \text{ ml} \times \frac{0.10 \text{mol NaOH}}{1\text{L NaOH}} \times \frac{1\text{L}}{1000\text{mL}} = 1.0 \times 10^{-3} \text{ mol}$$

The number of moles of HCl originally present in 25 mL of solution is

$$25 \text{ mL} \times \frac{0.10 \text{mol HCl}}{1\text{L HCl}} \times \frac{1\text{L}}{1000\text{mL}} = 2.5 \times 10^{-3} \text{ mol}$$

Thus, the amount of HCI left after partial neutralization is $(2.5 \times 10^{-3}) - (1.0 \times 10^{-3})$, or 1.5×10^{-3} mol. Next, the concentration of H^+ ions in 35 mL of solulion is found as follows:

$$\frac{1.5 \times 10^{-3} \text{mol HCl}}{35 \text{mL}} \times \frac{1000\text{mL}}{1\text{L}} = 0.043 \text{ mol HCl/L.} = 0.043 \text{ M HCl}$$

Thus $[H^+] = 0.043$ M, and the pH of the solution is pH = –log 0.043 = 1.37

2. *After the addition of 25 mL of 0.10 M NaOH to 25 mL of 0.10 M HCl.* This is a simple calculation because it involves a complete neutralization reaction and the salt (NaCl) does not undergo hydrolysis. At the equivalence point. $[H^+] = [OH^-]$ and the pH of the solution is 7.00.

3. *After the addition of 35 mL of 0.10 M NaOH to 25 mL of 0.10 M HCl.* The total volume of the solution is now 60 mL. The number of moles of NaOH added is

$$35 \text{ mL} \times \frac{0.10 \text{ mol NaOH}}{1\text{L NaOH}} \times \frac{1000\text{mL}}{1\text{L}} = 3.5 \times 10^{-3} \text{ mol}$$

The number of moles of HCl in 25 mL solution is 2.5×10^{-3}. After complete neutralization of HCl. the amount of NaOH left is $(3.5 \times 10^{-3}) - (2.5 \times 10^{-3})$ or 1.0×10^{-3} mol.

The concentration of NaOH in 60 mL of solution is

$$\frac{1.0 \times 10^{-3} \text{mol NaOH}}{60 \text{mL}} \quad \frac{1000\text{mL}}{1\text{L}} = 0.017 \text{ mol NaOH/L} = 0.017 \text{ M NaOH}$$

Thus $[OH^-] = 0.017$ M and pOH = –log 0.017 = 1.77. Hence, the pH of the solution is

pH = 14.00 – pOH

= 14.00 - 1.77 = 12.23

(ii) WEAK ACID-STRONG BASE TITRATIONS

Consider the neutralization reaction between acetic acid (a weak acid) and sodium hydroxide (a strong base):

$$CH_3COOH(aq) + NaOH(aq) \longrightarrow CH_3COONa(aq) + H_2O(l)$$

This equation can be simplified to

$$CH_3COOH(aq) + OH^-(aq) \longrightarrow CH_3COO^- (aq) + H_2O(l)$$

The acetate ion undergoes hydrolysis as follows:

$$CH_3COO^- (aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^- (aq)$$

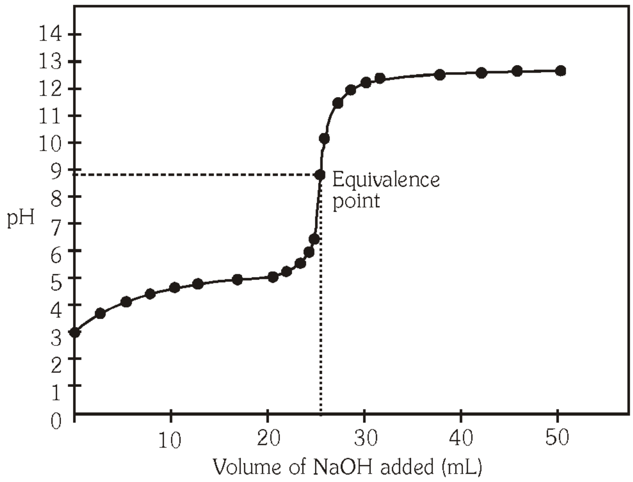
Therefore, at the equivalence point, when we only have sodium acetate present, the pH will greater than 7 as a result of the excess OH^- ions formed (Figure-2) that this situation is analogous to the hydrolysis of sodium acetate (CH_3COONa).

The following graph deals with the titration of a weak acid with a strong base

Figure 2 : pH profile of a weak acid-strong base titration. A 0.10 M NaOH solution is added from a buret to 25 ml of a 0.10 M CH_3COOH solution in an Erlenmeyer flask. Due to the hydrolysis of the salt formed, the pH at the equivalence point is

greater than 7.

(iii) STRONG ACID-WEAK BASE TITRATIONS



Consider the titration of HCl. a strong acid, with NH_3 a weak base:

$$HCl(aq) + NH_3(aq) \longrightarrow NH_4Cl(aq) \text{ or simply}$$

$$H^+(aq) + NH_3(aq) \longrightarrow NH_4^+ (aq)$$

The pH at the equivalence point is less than 7 due to the hydrolysis of the NH_4^+ ion:

$$NH_4^+ (aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+ (aq) \text{ or simply}$$

$$NH_4^+ (aq) + \rightleftharpoons NH_3(aq) + H^+ (aq)$$

Because of the volatility of an aqueous ammonia solution, it is more convenient to add hydrochloric acid from a buret to the ammonia solution. Figure 3 (a) shows the titration curve for this experiment and Figure 3 (b) shows the titration curve for the case in which a weak base is added from a buret to HCl.

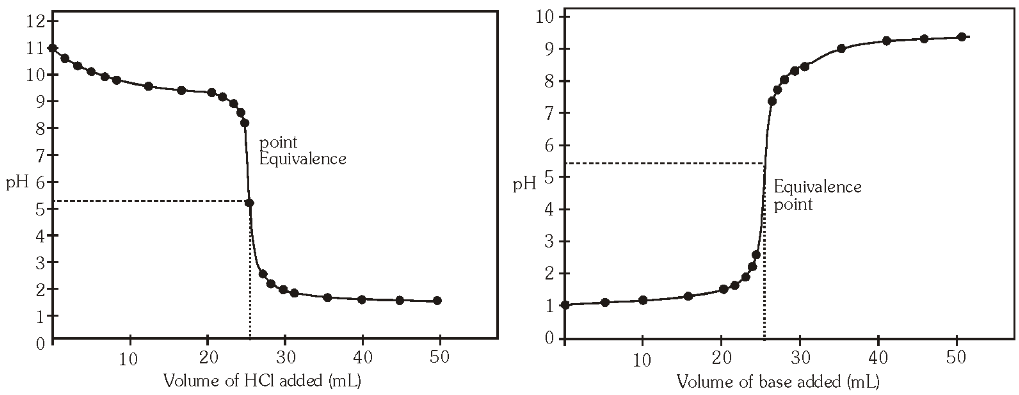


Figure 3 : pH profiles of a strong acid-weak base titration. 3 (a) A 0.10 M HCl solution is added from a buret to 25 ml of a 0.10 M NH_3 solution in an Erlenmeyer flask, 3 (b) A 0.10 M weak base solution is added from a buret to 25 mL of a 0.10 M HCl solution in an Erlenmeyer flask. As a result of salt hydrolysis, the pH at the equivalence point in both cases is lower than 7.

GENERAL CONCEPT OF CHEMISTRY

EXERCISE-I

INTRODUCTION

1. Which property of an element is not variable?
(1) Valency (2) Atomic weight (3) Equivalent weight (4) None

2. Which of the following is not a divalent ion ?
(1) Sulphate (2) Nitrate
(3 Hydrogen phosphate (4) Carbonate

3. The significant figures in 5.23×10^5 are –
(1) 8 (2) 3 (3) 4 (4) Infinite

4. The charge on cation M is +2 and anion A is –3. The compound has the formula :
(1) MA_2 (2) M_3A_2 (3) M_2A_3 (4) M_2A

5. Chemical formula of ferric oxide is
(1) FeO (2) Fe_2O_3 (3) Fe_3O_4 (4) none of these

LAWS OF CHEMCIAL COMBINATION

6. The law of multiple proportion was proposed by :
(1) Lavoisier (2) Dalton (3) Proust (4) Gay Lussac

7. In the reaction $N_2 + 3H_2 \longrightarrow 2\,NH_3$, ratio by volume of N_2, H_2 and NH_3 is 1 : 3 : 2. This illustrates law of -
(1) Definite proportion (2) Multiple proportion
(3) Reciprocal proportion (4) Gaseous volumes

8. Different proportions of oxygen in the various oxides of nitrogen prove the law of -
(1) Equivalent proportion (2) Multiple proportion
(3) Constant proportion (4) Conservation of matter

9. Law of definite proportions when expressed in terms of volumes becomes -
(1) Dalton's law (2) Berzelius hypothesis (3) Gay-Lussac's Law (4) Avogadro's law

10. The law of conservation of mass holds good for all of the following except -
(1) All chemical reactions (2) Nuclear reactions
(3) Endothermic reactions (4) Exothermic reactions

11. If law of conservation of mass was to hold true, then 20.8 g of $BaCl_2$ on reaction with 9.8 g of H_2SO_4 will produce 7.3 g of HCl and $BaSO_4$ equal to :-
(1) 11.65 g (2) 23.3 g (3) 25.5 g (4) 30.6 g

12. A chemical equation is balanced according to the law of –
(1) Multiple proportions (2) Constant proportions
(3) Reciprocal proportions (4) Conservation of mass

13. Among the following pairs of compounds, the one that illustrates the law of multiple proportions is
(1) NH_3 and NCl_3 (2) H_2O and D_2O (3) CuO and Cu_2O (4) CS_2 and $FeSO_4$

14. Law of reciprocal proportion can be verified by following compounds -
(1) CO_2, CH_4, H_2O (2) N_2O, N_2O_3, N_2O_5
(3) NaCl, Na_2CO_3, NaOH (4) H_2O, HCl, NaCl

15. Oxygen combines with two isotopes of carbon ^{12}C and ^{14}C to form two sample of carbon dioxide. The data illustrates -
(1) Law of conservation of mass (2) Law of multiple proportions
(3) Law of reciprocal proportions (4) None of these

MOLE CONCEPT

16. From the following the number of atoms is greatest in -
(1) 4 g hydrogen (2) 71 g chlorine (3) 48 g magnesium (4) 127 g iodine

17. One mole of CO_2 contains
(1) 6.02×10^{23} atoms of C (2) 6.02×10^{23} atoms of O
(3) 18.1×10^{23} molecules of CO_2 (4) 3 grams atoms of CO_2

18. Which has maximum molecules?
(1) 7 g N_2 (2) 2 g H_2 (3) 16 g NO_2 (4) 16 g O_2

19. The maximum number of molecules is present in:-
(1) 5L of N_2 gas at STP (2) 0·5g of H_2 gas
(3) 10g of O_2 gas (4) 15L of H_2 gas at STP

20. The actual weight of a molecule of water is -
(1) 18 g (2) 2.99×10^{-23} g
(3) both (1) & (2) are correct (4) None of these

21. The weight of substance that displace 22.4 litre air at NTP is :
(1) Molecular weight (2) Atomic weight (3) Equivalent weight (4) All

22. The volume of CO_2 at STP obtained by heating 1g of $CaCO_3$ will be
(1) 1 litre (2) 22.4 litre (3) 0.224 litre (4) 11.2 litre

23. 9 g of Al will react, with
(1) 6 g O_2 (2) 8 g O_2 (3) 9 g O_2 (4) 4 g O_2

24. Volume of CO_2 obtained at STP by the complete decomposition of 9.85 g. $BaCO_3$ is – (At. wt of Ba = 137)
(1) 2.24 lit (2) 1.12 lit (3) 0.84 lit (4) 0.56 lit

25. How many moles of potassium chlorate to be heated to produce 11.2 litre oxygen at NTP?
(1) $\frac{1}{2}$ mol (2) $\frac{1}{3}$ mol (3) $\frac{1}{4}$ mol (4) $\frac{2}{3}$ mol

26. 26 cc of CO_2 are passed over red hot coke. The volume of CO evolved is :–
(1) 15 cc (2) 10 cc (3) 32 cc (4) 52 cc

27. H_2 evolved at STP on complete reaction of 27 g of Aluminium with excess of aqueous NaOH would be
(1) 22.4 litres (2) 44.8 litres (3) 67.2 litres (4) 33.6 litres

28. Which contains least number of molecules ?
(1) 1 g CO_2 (2) 1 g N_2 (3) 1 g O_2 (4) 1 g H_2

29. The number of atoms in n mole of gas can be given by :
(1) $n \times N_A \times$ atomicity (2) $\frac{n \times N_A}{\text{Atomicity}}$ (3) $\frac{N_A \times \text{Atomicity}}{n}$ (4) None

30. Total number of protons, electrons and neutrons in 12g of ${}^{12}_{6}C$ is :-
(1) 1.8 (2) 12.044×10^{23} (3) 1.084×10^{25} (4) 10.84×10^{23}

31. The number of g molecules of oxygen in 6.02×10^{23} CO molecules is:-
(1) 1 gram molecule (2) 0.5 gram molecule (3) 5 gram molecule (4) 10 gram molecule

32. The number of moles of carbon dioxide which contain 8 g of oxygen is –
(1) 0.5 (2) 0.20 (3) 0.40 (4) 0.25

33. The number of gram molecules of oxygen in 6.02×10^{24} CO molecules is –
(1) 10 gram molecules (2) 5 gram molecules (3) 1 gram molecules (4) 0.5 gram molecules

34. The number of molecules in 16 g of methane is
(1) 3.0×10^{23} (2) 6.02×10^{23} (3) $\frac{16}{6.02} \times 10^{23}$ (4) $\frac{16}{3.0} \times 10^{23}$

35. The molecular weight of hydrogen peroxide is 34. What is the unit of gm molecular weight?
(1) g (2) mol (3) g mol^{-1} (4) mol g^{-1}

36. The number of moles of sodium oxide in 620 g of it is
(1) 1 mol (2) 10 moles (3) 18 moles (4) 100 moles

37. The number of molecules of SO_2 present in 64g of SO_2 is
(1) 6.0×10^{23} (2) 3×10^{23} (3) 12×10^{23} (4) 3×10^{10}

38. The number of atoms present in 16 g of oxygen is
(1) $6.02 \times 10^{11.5}$ (2) 3.01×10^{23} (3) $3.01 \times 10^{11.5}$ (4) 6.02×10^{23}

39. Number of oxygen atoms in 8 g of ozone is -
(1) 6.02×10^{23} (2) $\frac{6.02 \times 10^{23}}{2}$ (3) $\frac{6.02 \times 10^{23}}{3}$ (4) $\frac{6.02 \times 10^{23}}{6}$

40. The number of molecule in 4.25 g of NH_3 is -
(1) 1.505×10^{23} (2) 3.01×10^{23} (3) 6.02×10^{23} (4) None of these

41. Which of the following contains greatest number of oxygen atoms?
(1) 1 g of O (2) 1 g of O_2
(3) 1 g of O_3 (4) all have the same number of atoms

42. How many moles of water are present in 180 g of water?
(1) 1 mole (2) 18 mole (3) 10 mole (4) 100 mole

43. How many H atoms are in 3.42 g of $C_{12}H_{22}O_{11}$?
(1) 6.0×10^{23} (2) 1.3×10^{23} (3) 3.8×10^{22} (4) 6.0×10^{21}

44. The volume occupied by 16 g of oxygen under NTP conditions is :
(1) 1.12 dm^3 (2) 5.60 dm^3 (3) 11.20 dm^3 (4) 22.4 dm^3

45. The percentage value of nitrogen in urea is about
(1) 46 (2) 85 (3) 18 (4) 28

EMPIRICAL FORMULA AND MOLECULER FORMULA

46. The empirical formula of a compound is CH_2. One mole of this compound has a mass of 42 g. Its molecular formula is :
(1) CH_2 (2) C_2H_2 (3) C_3H_6 (4) C_3H_8

47. The empirical formula of a compound is CH_2O and its molecular weight is 120. The molecular formula of the compound is -
(1) $C_2H_4O_2$ (2) $C_3H_6O_3$ (3) $C_4H_8O_4$ (4) CH_2O

48. A hydrocarbon contains 75% of carbon. Then its molecular formula is -
(1) CH_4 (2) C_2H_4 (3) C_2H_6 (4) C_2H_2

49. An oxide of sulphur contains 50% of sulphur in it. Its empirical formula is -
(1) SO_2 (2) SO_3 (3) SO (4) S_2O

50. Empirical formula of glucose is -

(1) $C_6H_{12}O_6$ (2) $C_3H_6O_3$ (3) $C_2H_4O_2$ (4) CH_2O

51. A compound is found to contain 80% of carbon and 20% of hydrogen, then the molecular formula of the compound is-

(1) C_6H_6 (2) C_2H_5OH (3) C_2H_6 (4) C_2H_4

52. The simplest formula of a compound containing 50% of element X (atomic weight = 10) and 50% of element Y (atomic weight = 20) is:-

(1) XY (2) X_2Y (3) XY_2 (4) X_3Y

53. Which of the following compounds has same empirical formula as that of glucose?

(1) CH_3CHO (2) CH_3COOH (3) CH_3OH (4) C_2H_6

54. A compound of X and Y has equal mass of them. If their atomic weights are 30 and 20 respectively. Molecular formula of that compound (its mol. wt. is 120) could be -

(1) X_2Y_2 (2) X_3Y_3 (3) X_2Y_3 (4) X_3Y_2

55. Which compound has the empirical formula with the greatest mass ?

(1) C_2H_6 (2) C_4H_{10} (3) C_5H_{10} (4) C_6H_6

MOLARITY

56. The number of moles of OH^- in 0.3 litre of 0.005M $Ba(OH)_2$ is -

(1) 0.075 (2) 0.005 (3) 0.045 (4) 0.003

57. The molarity of 0.006 mole of NaCl in 100ml solution is -

(1) 0.6 (2) 0.06 (3) 0.006 (4) 0.066

58. What will be the molarity of a solution containing 5g of sodium hydroxide in 250ml solution -

(1) 0.5 (2) 1.0 (3) 2.0 (4) 0.1

59. Molarity is expressed as

(1) Gram/litre (2) Moles/litre (3) Litre/mole (4) Moles/1000gms

60. 171g of cane sugar ($C_{12}H_{22}O_{11}$) is dissolved in 1 litre of water. The molarity of the solution is

(1) 2.0 M (2) 1.0 M (3) 0.5 M (4) 0.25 M

61. Which statements is true for solution of 0.020M H_2SO_4?

(1) 2 litre of the solution contains 0.020 mole SO_4^{2-}

(2) 2 litre of the solution contains 0.080 mole SO_4^{2-}

(3) 2 litre of the solution contains 0.010 mole SO_4^{2-}

(4) 2 litre of the solution contains .04 mole of SO_4^{2-}

62. 2.0 molar solution is obtained, when 0.5 mole solute is dissolved in -

(1) 250 ml solvent (2) 250 g solvent (3) 250 ml solution (4) 1000 ml solvent

63. How many gram of HCl will be present in 150ml of its 0.52 M solution?

(1) 2.84 gm (2) 5.70 gm (3) 8.50 gm (4) 3.65 gm

64. The number of moles present in 2 litre of 0.5M NaOH is -

(1) 0.5 (2) 0.1 (3) 1 (4) 2

65. Molar solution means 1 mole of solute present in-

(1) 1000 g of solvent (2) 1 litre of solvent (3) 1 litre of solution (4) 1000 g of solution

66. What volume of 0.8 M solution contains 0.1 mole of the solute?
(1) 100 ml (2) 125 ml (3) 500 ml (4) 62.5 ml

67. 20 g of hydrogen is present in 5 litre vessel. The molar concentration of hydrogen is -
(1) 4 (2) 1 (3) 3 (4) 2

68. Molecular weight of urea is 60. A solution contain 6 gram of urea in one litre is-
(1) 1 molar (2) 1.5 molar (3) 0.1 molar (4) 0.01 molar

69. 8g NaOH is dissolved in one litre of solution, its molarity is
(1) 0.8 M (2) 0.4 M (3) 0.2 M (4) 0.1 M

70. For preparing 0.1M solution of H_2SO_4 in one litre, amount of H_2SO_4 needed is
(1) 0.98g (2) 4.9g (3) 49.0 g (4) 9.8 g

MOLALITY

71. The number of moles of solute per kg of solvent is called its -
(1) Molarity (2) Normality (3) Molar fraction (4) Molality

72. 7.45 g of potassium chloride is dissolved in 100 g of water. What will be the molality of the solution?
(1) 1m (2) 2m (3) 1.59m (4) 1.93m

73. The molality of a solution prepared by dissolving 5.64 g of glucose in 60g of water is :
(1) 5.22 m (2) 0.522 m (3) 10.44 m (4) 25.6 m

74. Calculate the molality of solute in aqueous solution containing 6 g of urea per 500 g of water (Mol. mass of urea = 60).
(1) 1.32m (2) 4.02 m (3) 0.2m (4) 0.22m

75. 2.82g of glucose is dissolved in 30g of water. The molality of solution is :
(1) 0.52m (2) 0.99m (3) 0.63m (4) 0.66m

MOLE FRACTION

76. If 5.85 gms of NaCl are dissolved in 90gms of water, the mole fraction of NaCl is -
(1) 0.1 (2) 0.2 (3) 0.3 (4) 0.019

77. A mixture has 18g water and 414g ethanol. The mole fraction of water in mixture is (assume ideal behaviour of the mixture)
(1) 0.1 (2) 0.4 (3) 0.7 (4) 0.9

78. 36 g water and 828 g ethyl alcohol form an ideal solution. The mole fraction of water in it, is -
(1) 1.0 (2) 0.7 (3) 0.4 (4) 0.1

79. The sum of the mole fraction of the components of a solution -
(1) 0 (2) 1 (3) 2 (4) 4

80. What is mole fraction of a solute in its one molal aqueous solution?
(1) 1 (2) 1.8 (3) 18 (4) 0.018

EQUIVALENT WEIGHT

81. Molecular weight of tribasic acid is W. Its equivalent weight will be :

(1) $\frac{W}{2}$ (2) $\frac{W}{3}$ (3) W (4) 3W

82. If 1.2 g of a metal displaces 1.12 litre of hydrogen at NTP, equivalent mass of the metal would be –

(1) 1.2×11.2 (2) 12 (3) 24 (4) $1.2 + 11.2$

83. 2.8 g of iron displaces 3.2 g of copper from a solution of copper sulphate solution. If the equivalent mass of iron is 28, then equivalent mass of copper will be –

(1) 16 (2) 32 (3) 48 (4) 64

84. Sulphur forms two chlorides S_2Cl_2 and SCl_2. The equivalent mass of sulphur in SCl_2 is 16. The equivalent weight of sulphur in S_2Cl_2 is -

(1) 8 (2) 16 (3) 32 (4) 64

85. 1.0 g of a metal combines with 8.89 g of Bromine. Equivalent weight of metal is nearly-

(1) 8 (2) 9 (3) 10 (4) 7

86. If equivalent weight of S in SO_2 is 8 then equivalent weight of S in SO_3 is -

(1) $\frac{8\times 2}{3}$ (2) $\frac{8\times 3}{2}$ (3) $8 \times 2 \times 3$ (4) $\frac{2\times 3}{8}$

87. The oxide of a metal has 32% oxygen. It's equivalent weight would be:-

(1) 34 (2) 32 (3) 17 (4) 16

88. A, E, M and n are the atomic weight, equivalent weight, molecular weight and valency of an element. The correct relation is

(1) $A = E \times n$ (2) $A = \frac{M}{E}$ (3) $A = \frac{M}{n}$ (4) $M = A \times n$

89. The equivalent mass of Fe in FeO is

(1) 56 (2) 28 (3) 36 (4) 18.66

90. The equivalent weihgt of a certain trivalent element is 20. Molecular weight of its oxide is

(1) 168 (2) 68 (3) 152 (4) 56

LAW OF GRAM EQUIVALENCE

91. The normality of 0.3M phosphorus acid (H_3PO_3) is-

(1) 0.1 (2) 0.9 (3) 0.3 (4) 0.6

92. The normality of 2.3M H_2SO_4 solution is -

(1) 2.3 N (2) 4.6 N (3) 0.46 N (4) 0.23 N

93. The amount of $K_2Cr_2O_7$ (eq. wt. 49.04) required to prepare 100ml of its 0.05N solution -

(1) 2.4924 g (2) 0.4904 g (3) 1.4712 g (4) 0.2452 g

94. Which of the following solutions has the highest normality?

(1) 8 g of KOH/litre (2) N phosphoric acid

(3) 8 g of NaOH/100 ml (4) 0.5M H_2SO_4

95. Normality (N) of a solution is equal to -

(1) $\frac{\text{No. of moles of solute}}{\text{Volume of solution in litre}}$ (2) $\frac{\text{No. of gram equivalent of solute}}{\text{Volume of solution in litre}}$

(3) $\frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$ (4) None of these

96. The weight of pure NaOH required to prepare 250cm^3 of 0.1N solution is -

(1) 4g (2) 1g (3) 2g (4) 10g

97. Molarity of 0.2N H_2SO_4 is -

(1) 0.2 (2) 0.4 (3) 0.6 (4) 0.1

98. The molarity of 0.2 N Na_2CO_3 solution will be :

(1) 0.5M (2) 0.2M (3) 0.1M (4) 0.4 M

99. The volume of water to be added to 100cm^3 og 0.5 M H_2SO_4 to get decinomal concentration is

(1) 400cm^3 (2) 500cm^3 (3) 450cm^3 (4) 100cm^3

100. The normality of H_2SO_4 having 0.05 equivalent in 2 litre is

(1) 0.1 N (2) 0.025 N (3) 0.5 N (4) 0. 25 N

TITRATION

101. NaClO solution reacts with H_2SO_3 as,$NaClO + H_2SO_3 \rightarrow NaCl + H_2SO_4$. A solution of NaClO used in the above reaction contained 15g of NaClO per litre. The normality of the solution would be -

(1) 0.8 (2) 0.4 (3) 0.2 (4) 0.33

102. NH_2SO_4 was diluted from 1 litre to 10 litrs. Normality of the solution obtained is -

(1) 10 N (2) 5 N (3) 1 N (4) 0.5 N

103. Calculate the volume of N/100 oxalic acid solution obtained by dissolving 126g of oxalic acid [$(COOH)_2.2H_2O$]

(1) 100 litres (2) 300 litres (3) 200 litres (4) 250 litres

104. 100 cc of 1.2 N acid should be diluted to .....in order to prepare its one normal solution

(1) 120 c.c. (2) 200 c.c. (3) 240 c.c. (4) 360 c.c.

105. What is the volume of water that should be added to 150 mL of $\frac{N}{2}$ oxalic acid to prepare a solution of $\frac{N}{10}$ oxalic acid?

(1) 750 c.c (2) 400 c.c. (3) 800 c.c. (4) 600 c.c.

106. How much of NaOH is required to neutralise 1500 cm^3 of 0.1N HCl (At wt. of Na = 23)?

(1) 4g (2) 6g (3) 40 g (4) 60 g

107. 15 ml of N/10 NaOH solution completely neutralises 12 ml of H_2SO_4 solution. The normality of H_2SO_4 solution will be -

(1) N/5 (2) N/10 (3) N/8 (4) N

108. 8 ml of $\frac{N}{10}$ HCl are required to neutralise 20ml solution of Na_2CO_3 in water. Normality of Na_2CO_3 solution is -

(1) 0.40 N (2) 0.040 N (3) 4.0 N (4) 1.4 N

109. The volume of 0.05M H_2SO_4 required to neutralize 80 ml of 0.13N NaOH will be-

(1) 104 ml (2) 52 ml (3) 10.4 ml (4) 26 ml

110. To neutralize 200mL of $Ba(OH)_2$ solution, we have to add 100 mL of 0.1 M H_2SO_4 solution. What is the normality of Ba (OH_2)?

(1) 0.05 (2) 0.01 (3) 1.0 (4) 0.1

EXERCISE-II

INTRODUCTION (VAPOUR DENSITY)

1. The vapour density of a gas is 11.2 the volume occupied by 11.2 g of this gas at NTP is –
(1) 1 litre (2) 11.2 litre (3) 22.4 litre (4) 20 litre

2. At normal temperature and pressure 0.24 g volatile substance produces 43 mL vapour. Then the vapour density of substance will be– (Density of hydrogen = 0.000089g/mL)
(1) 95.39 (2) 5.993 (3) 95.93 (4) 62.5

3. Vapour density of a volatile substance w.r.t CH_4 is 4. Its molecular weight would be –
(1) 8 (2) 32 (3) 64 (4) 128

4. Density of a gas at STP is 1.43 g/L What is its molecular weight?
(1) 16 (2) 32 (3) 48 (4) 64

5. The atomic mass of an element is 27. If valency is 3, the vapour density of the volatile chloride will be:-
(1) 66.75 (2) 6.675 (3) 667.5 (4) 81

6. A gas is found to have the formula $(CO)_x$. It's VD is 70 the value of x must be:-
(1) 7 (2) 4 (3) 5 (4) 6

7. The equivalent weight of an element is 4. It's chloride has a V.D. 59.25. Then the valency of the element is –
(1) 4 (2) 3 (3) 2 (4) 1

8. The vapour density of a gas is 35.5. The volume occupied by 3.55g of the gas at N.T.P. is
(1) 1.12 litres (2) 11.2 litres (3) 22.4 litres (4) 44.8 litres

9. The vapour density of pure ozone would be :
(1) 48 (2) 32 (3) 24 (4) 16

10. The vapour density of a gas is 11.2. The volume occupied by 11.2 g of this gas at N.T.P. is
(1) 22.4 litres (2) 11.2 litres (3) 1 litre (4) 2.24 litres

LAWS OF CHEMCIAL COMBINATION

11. One part of an element A combines with two parts of another element B. Six parts of the element C combines with four parts of the element B. If A and C combine together the ratio of their weights will be governed by -
(1) Law of definite proportion (2) Law of multiple proportion
(3) Law of reciprocal proportion (4) Law of conservation of mass

12. Which of the following combination illustrate law of reciprocal proportion :–
(1) N_2O_3, N_2O_4, N_2O_5(2) PH_3, P_2O_5, P_2S_5 (3) CS_2, CO_2, SO_2 (4) NaCl, NaBr, NaI

13. Which of the following is the best example of law of conservation of mass?
(1) 12 g of carbon combines with 32 g of oxygen to form 44 g of CO_2
(2) When 12 g of carbon is heated in a vacuum there is no change in mass
(3) A sample of air increases in volume when heated at constant pressure but its mass remains unaltered
(4)The weight of a piece of platinum is the same before and after heating in air

14. The law of definite proportions is not applicable to nitrogen oxide because.
(1) Nitrogen atomic weight is not constant (2) Nitrogen molecular weight is variable
(3) Nitrogen equivalent weight is variable (4) Oxygen atomic weight is variable

15. Hydrogen combines with chlorine to form HCl. It also combines with sodium to form NaH. If sodium and chlorine also combine with each other, they will do so in the ratio of their masses as:-

(1) 23 : 35.5 (2) 35.5 : 23 (3) 1 : 1 (4) 23 : 1

MOLE CONCEPT

16. How much amount of zinc is required to react with dilute H_2SO_4 for obtaining 224 mL hydrogen at STP-

(1) 0.65 g (2) 6.5 g (3) 65 g (4) 0.065 g

17. 50 g $CaCO_3$ will react with ...... g of 20% HCl by weight.

(1) 36.5 g (2) 73 g (3) 109.5 g (4) 182.5 g

18. A sample of calcium carbonate is 80% pure. 25 g of this sample is treated with excess of HCl. How much volume of CO_2 will be obtained at N.T.P.?

(1) 4.48 litre (2) 5.6 litre (3) 11.2 litre (4) 2.24 litre

19. If 32 g of O_2 contains 6.022×10^{23} molecules at NTP. Then 32 g of S, under the same conditions, will contain:-

(1) 6.022×10^{23} S (2) 3.011×10^{21} S (3) 1×10^{23} S (4) 12.044×10^{23}S

20. If V ml of the vapours of substance at NTP weighs W gm. Then mol. wt. of substance is:-

(1) $(W/V) \times 22400$ (2) $\frac{V}{W} \times 22.4$ (3) $(W - V) \times 22400$ (4) $\frac{W \times 1}{V \times 22400}$

21. A person adds 1.71 gram of sugar ($C_{12}H_{22}O_{11}$) in order to sweeten his tea. The number of carbon atoms added are (mol. mass of sugar = 342)

(1) 3.6×10^{22} (2) 7.2×10^{21} (3) 0.05 (4) 6.6×10^{22}

22. The total number of valence electrons in 4.2 g of N^{-3} ion is -

(1) 2.1 N_A (2) 4.2 N_A (3) 1.6 N_A (4) 2.4 N_A

23. In Victor Mayer's method 0.2g of an organic substance displaced 56mL of air at S.T.P. The molecular Weight of compounds

(1) 56 (2) 112 (3) 80 (4) 28

24. Which of the following contain highest number of molecules?

(1) 2.8 g of CO (2) 3.2 g of CH_4 (3) 1.7 g of NH_3 (4) 3.2 g of SO_2

25. 1.2 g of Mg (At mass 24) will produce MgO equal to -

(1) 0.05 mol (2) 40 g (3) 40 mg (4) 4 g

26. The volume of gas at NTP produced by 100 g of CaC_2 with water:-

(1) 70 litre (2) 35 litre (3) 17.5 litre (4) 22.4 litre

27. How many grams of H_2SO_4 are needed to neutralize completely 10 litres of ammonia gas (at NTP)?

(1) 21.87 g (2) 49.0 g (3) 24.5 g 4) 98.0 g

28. 4.4 g of an unknown gas occupies 2.24 litres of volume at STP. The gas may be :-

(1) N_2O (2) CO (3) CO_2 (4) 1 & 3 both

29. What is the mass of a molecule of CH_4?

(1) 16 (2) 26.6×10^{22} (3) 2.66×10^{-23} (4) 16 NA

30. Given that one mole of N_2 at NTP occupies 22.4 litre the density of N_2 is -

(1) 1.25 g/L (2) 0.80 g/L (3) 2.5 g/L (4) 1.60 g/L

31. 5.6 lt. of oxygen at STP contains -

(1) 6.02×10^{23} atoms (2) 3.01×10^{23} atoms

(3) 1.505×10^{23} atoms (4) 0.7525×10^{23} atoms

32. Which of the following has the highest mass ?

(1) 1 g atom of C (2) ½mole of CH_4

(3) 10 ml of water (4) 3.011×10^{23} atoms of oxygen

33. How many g of H_2SO_4 is present in 0.25 g mole of H_2SO_4?

(1) 24.5 (2) 2.45 (3) 0.25 (4) 0.245

34. Which of the following contains largest number of atoms ?

(1) 4 g of NH_3 (2) 16 g of O_2 (3) 28 g of N_2 (4) 18 g of H_2O

35. 500 ml. of a gaseous hydrocarbon when burnt in excess of O_2 gave 2.5 lt. of CO_2 and 3.0 lt. of water vapours under same conditions. Molecular formula of the hydrocarbon is -

(1) C_4H_8 (2) C_4H_{10} (3) C_5H_{10} (4) C_5H_{12}

EMPIRICAL FORMULA AND MOLECULER FORMULA

36. The number of atoms of Cr and O are 4.8×10^{10} and 9.6×10^{10} respectively. Its empirical formula is

(1) Cr_2O_3 (2) CrO_2 (3) Cr_2O_4 (4) None

37. The empirical formula of an organic compound containing carbon and hydrogen is CH_2. The mass of one litre of this organic gas is exactly equal to that of one litre of N_2. Therefore, the molecular formula of the organic gas is –

(1) C_2H_4 (2) C_3H_6 (3) C_6H_{12} (4) C_4H_8

38. A 400mg iron capsule contains 100mg of ferrous fumarate, $(CHCOO)_2Fe$. The percentage of iron present in it is approximately

(1) 33% (2) 25% (3) 14% (4) 8%

39. Two elements X (Atomic weight = 75) and Y (Atomic weight = 16) combine to give a compound having 75.8% of X. The empirical formula of compound is

(1) XY (2) X_2Y (3) X_2Y_2 (4) X_2Y_3

40. An oxide of metal M has 40% by mass of oxygen. Metal M has atomic mass of 24. The empirical formula of the oxide

(1) M_2O (2) M_2O_3 (3) MO (4) M_3O_4

41. On analysis, a certain compound was found to contain iodine and oxygen in the weight ratio of 254:80. The formula of the compound is : (At mass I = 127, O = 16)

(1) IO (2) I_2O (3) I_5O_2 (4) I_2O_5

42. Percentage of C, H and N are given as follows C = 40%, H = 13.33%, N = 46.67% The empirical formula will be –

(1) CH_2N (2) C_2H_4N (3) CH_4N (4) CH_3N

43. An organic compound containing carbon, hydrogen and oxygen contains 52.20% carbon and 13.04% hydrogen. Vapour density of the compound is 23. Its molecular formula will be :

(1) C_2H_6O (2) C_3H_8O (3) C_4H_8O (4) $C_5H_{10}O$

44. A compound contains 38.8% C, 16.0% H and 45.2% N. The empirical formula of the compound would be

(1) CH_3NH_2 (2) CH_3CN (3) C_2H_5CN (4) $CH_2(NH)_2$

45. A compound (80 g) on analysis gave C = 24 g, H = 4 g, O = 32 g. Its empirical formula is

(1) $C_2H_2O_2$ (2) C_2H_2O (3) CH_2O_2 (4) CH_2O

MOLARITY

46. What is the concentration of nitrate ions if equal volumes of 0.1 M $AgNO_3$ and 0.1 M NaCl are mixed together

(1) 0.1 M (2) 0.2 M (3) 0.05 M (4) 0.25 M

47. The density of a solution containing 13% by mass of sulphuric acid is 1.09g/mL. The molarity of the solution is :

(1) 1.445 M (2) 1.5 M (3) 2.89 M (4) None of these

48. The density of a solution containing 40% by mass of HCl is 1.2 g/mL. The molarity of the solution is :

(1) 1.45 (2) 0.78 (3) 13.89 (4) 13.15

49. 500ml of a glucose solution contains 6.02×10^{22} molecules. The concentration of solution is

(1) 0.1 M (2) 1.0 M (3) 0.2 M (4) 2.0 M

50. If 500ml of 1M solution of glucose is mixed with 500ml of 1M solution, final molarity of solution will be

(1) 1M (2) 0.5 M (3) 2M (4) 1.5M

MOLALITY

51. The molality of 90% weight/volume H_2SO_4 solution is [density = 1.8gm/ml]

(1) 1.8 (2) 48.4 (3) 9.18 (4) 91.8

52. The molality of a 13% solution (by weight) of sulphuric acid is : (density = 1.020 g/cm^3)

(1) 1.524m (2) 4.524m (3) 8.092m (4) 0.146m

53. The molality of K_2CO_3 solution formed by dissolving 2.5 g of in one litre of solution is : (density = 0.85g/ml)

(1) 0.431 m (2) 0.0213 m (3) 563.2 m (4) 85.8 m

54. The mass of urea (NH_2 $CONH_2$) required in making 2.5 kg of 0.25 molal aqueous solution :

(1) 37.50 g (2) 36.94 g (3) 25.50 g (4) 36.00 g

55. The density of a 2.03 M solution of acetic acid in water is 1.017 g/mL. The molality of the solution is :

(1) 2.267 m (2) 7.622 m (3) 6.722 m (4) 2.762m

MOLE FRACTION

56. The mole fraction of water in 20% aqueous solution of H_2O_2 -

(1) $\frac{77}{68}$ (2) $\frac{68}{77}$ (3) $\frac{20}{80}$ (4) $\frac{80}{20}$

57. If 5.85 g of NaCl are dissolved in 90 g of water, the mole fraction of solute is

(1) 0.0196 (2) 0.01 (3) 0.1 (4) 0.2

58. Calculate mole fraction of solute in an aqueous solution containing 6 g of urea per 500 g of water: (molar mass of urea = 60 g)

(1) 0.00456 (2) 0.00359 (3) 0.0573 (4) 0.3268

59. The mole fraction of sugar ($C_{12}H_{22}O_{11}$) and mass % of water in the solution prepared by mixing 50 g sugar in 100 g of water at 25°C is :

(1) 0.0256 , 66.67% (2) 0.146 , 44.36% (3) 0.0625 , 77.28% (4) 0.326 , 80.39%

60. Calculate the mole fraction of ethyl alcohol in a solution of total volume 95ml prepared by adding 50ml of ethyl alcohol (density = 0.789 mL^{-1}) to 50 mL water (density = 1.00 g mL^{-1})

(1) 0.236 (2) 9.027 (3) 1.0875 (4) 0.0093

EQUIVALENT WEIGHT

61. The formula of a metal oxide is Z_2O_3. If 6 mg of hydrogen is required for complete reduction of 0.1596 g metal oxide, then the atomic weight of metal is -

(1) 227.9 (2) 159.6 (3) 79.8 (4) 55.8

62. How many g of copper (at weight = 64) would be displaced from at the copper sulphate solution by adding 27g of aluminium (at weight = 27)?

(1) 32 (2) 64 (3) 96 (4) 160

63. Element 'A' reacts with oxygen to form a compound A_2O_3. If 0.359 gram of 'A' react to give 0.559 gram of the compound, atomic weight of 'A' in amu will be :–

(1) 51 amu (2) 43.08 amu (3) 49.7 amu (4) 47.9 amu

64. 1.6 g of Ca and 2.60 g of Zn when treated with an acid in excess separately, produced the same amount of hydrogen. If the equivalent weight of Zn is 32.6, what is the equivalent weight of Ca:-

(1) 10 (2) 20 (3) 40 (4) 5

65. The oxide of an element possess the molecular formula M_2O_3. If the equivalent mass of the metal is 9, the molecular mass of the oxide will be –

(1) 27 (2) 75 (3) 102 (4) 18

66. 1 g metal carbonate requires 200mL. of 0.1N HCl for complete neutralization. What is the equivalent weight of metal carbonate?

(1) 50 (2) 40 (3) 20 (4) 100

67. 20 g solute is present in 200 mL solvent. The density of solvent is 0.9 g/mL. What is the weight fraction and weight percent of the solute?

(1) 0.1 and 10 (2) 0.2 and 20 (3) 0.3 and 10 (4) 0.2 and 30

68. 0.45 gm, of acid (molecular wt. = 90) was exactly neutrilised by 20ml. of N NaOH. Basically of the acid is

(1) 1 (2) 2 (3) 3 (4) 4

69. 0.84 gm of a metal carbonate react's exactly with 40ml of N/2 H_2SO_4 . The equivalent weight of the metal carbonate is

(1) 84 (2) 64 (3) 42 (4) 32

70. A metal oxide is reduced by heating it in a steam of hydrogen. It is found that after complete reduction 3.15 gm of the oxide yields 1.05 gm of the metal. We may conclude that

(1) Atomic weight of the metal is 4 (2) Equivalent weight of the metal is 8

(3) Equivalent weight of the metal is 4 (4) Atomic weight of the metal is 8

LAW OF EQUIVALENCE

71. The volume of water that must be added to a mixture of 250 ml of 0.6 HCl and 750 ml of 0.2M HCl to obtain 0.25 M solution HCl is

(1) 750 ml (2) 100ml (3) 200ml (4) 300ml

72. What volume of 0.10 M H_2SO_4 must be added to 50 mL of a 0.10 M NaOH solution to make a solution in which the molarity of the H_2SO_4 is 0.050 M?

(1) 400mL (2) 50mL (3) 100mL (4) 150mL

73. What approximate volume of 0.40 M Ba $(OH)_2$ must be added to 50.0 mL of 0.30 M NaOH to get a solution in which the molarity of the OH^- ions is 0.50M?

(1) 33mL (2) 66mL (3) 133mL (4) 100mL

74. The volume water that must be added to a mixture of 250ml of 0.6M HCl and 750ml of 0.2M HCl to obtain 0.25M solution of HCl is

(1) 750ml (2) 100ml (3) 200ml (4) 300ml

75. Calculate the final molarity when 2.0 litre of 3.0 M sugar solution and 3.0 litre of 2.5 M sugar solution are mixed and then diluted to 10 litres with water.

(1) 56.2 M (2) 1.35 M (3) 0.78 M (4) 1.25 M

VOLUME STRENGTH OF H_2O_2 + PERCENTAGE OF FREE SO3

Labelled as 'volume H_2O_2, it means volume of O_2 (in litre) at 1 atm & 273 K that can be obtained from 1 litre of such a sample when it decomposes according to

$$H_2O_2 \rightarrow H_2O + \frac{1}{2}O_2$$

Volume Strength of H_2O_2 Solution = 11.2 x molarity

76. The % w/V of "10V" H_2O_2 solution is

(1) 10.2% (2) 3.03% (3) 22.4% (4) 5.04%

77. The volume of O_2 at STP that would be liberated by 20V H_2O_2 solution is:

(1) 56 mL (2) 112 mL

(3) 168 mL (4) 224 mL

Percentage labelling of Oleum: Labelled as '% oleum', it means maximum amount of H_2SO_4 that can be obtained from 100 gm of such oleum (mix of H_2SO_4 and SO_3) by adding sufficient water. For ex. 109% oleum sample means, with the addition of sufficient water to 100 gin oleum sample 109 gm H_2SO_4 is obtained.

% labelling of oleum sample = (100 + x)%

x = mass of H_2O required for the complete conversion of SO_3 in H_2SO_4

78. The mass of free SO_3 present in 100 gm, 109 % oleum sample is.

(1) 20g (2) 80g

(3) 40g (4) 60g

79. The % labelling of 100 g oleum sample if it contains 20 g SO_3

(1) 109% (2) 4.5%

(3) 100% (4) 104.5%

80. % labelling of oleum sample if it contains 10g SO_3 & 15g H_2SO_4:

(1) 27.25 % (2) 106%

(3) 109% (4) 118%

QUESTIONS ASKED IN PREVIOUS EXAMS

EXERCISE-III

1. A solution containing Na_2CO_3 + NaOH requires 300 ml of 0.1 N HCl using phenolphthalein indicator. Methyl orange is then added to the above titrated solution and a further 25 ml of 0.2 N HCl is required. What is the amount of NaOH present in the solution ?

(1) 0.8 g (2) 1.0 g (3) 1.5 g (4) 2.0 g

2. A certain quantity of gas occupied a volume of 0.1L, when collected over water at 10°C and a pressure 0.90 atm. The same gas occupied a volume of 0.080 L at STP in dry conditions. Calculate the aqueous tension at 10°C.

(1) 0.061 (2) 0.051 (3) 0.071 (4) 0.081

3. A chemist's report on a batch of pharmaceutical products, Aspirin ($C_9H_8O_4$) (250mg tablets) and Paracetamol ($C_8H_9NO_2$) (500mg tablets) indicated a '+0.5%' weight error in each tablet. Due to this error, the consumer gets extra 'x' molecules of aspirin per tablet and extra 'y' molecules of paracetamol per tablet. Choose the 'Best relation between x and y.

(1) x = y (2) x > y (3) y > x (4) x = 2y

4. Sodium reacts with excess oxygen to form sodium oxide. A student wants to prepare 1.24g sodium oxide. While doing the calculations, he uses atomic number of sodium instead of atomic mass. What is the approximate percentage error in the mass of sodium oxide obtained due to this mistake?

(1) 11% (2) 23% (3) 48% (4) 60%

5. Which of the following combinations of elements of given atomic numbers can lead to a compound with a chemical formula of XY_3 ?

(1) 2 and 6 (2) 5 and 15 (3) 3 and 18 (4) 13 and 17

6. Which list of formulas represents compounds only?

(1) CO_2,H_2O,NH_3 (2) H_2,N_2,O_2 (3) H_2,Ne, NaCl (4) MgO, NaCl, O_2

7. Which substance can not be broken down by a chemical reaction ?

(1) Ammonia (2) Argon (3) Methane (4) Water

8. At the same temperature and pressure which sample contains the same number of moles of particles as 1 litre of O_2(g) ?

(1) 1 L Ne (g) (2) 2L N_2 (g) (3) 0.5 L SO_2 (g) (4) 1 L H_2O (g)

9. Which sample at STP has the same number of molecules as 5 liters of NO_2(g) at STP ?

(1) 5 grams of H_2(g) (2) 5 litres of CH_4(g)

(3) 5 moles of O_2(g) (4) 5×10^{23} molecules of CO_2(g)

10. Calculate the mass of Litium that contains same number of atoms as present in 8g of Magnesium. Atomic masses of lithium and magnesium are 7 and 24 respectively.

(1) 8g (2) 3g (3) 7g (4) 2.3 g

11. Three containers A,B and C of equal volume contain N_2,NO_2 and CO_2 respectively at the same temperature and pressure, The ascending order of their masses is

(1) A, C, B (2) C, A, B (3) B, C, A (4) C, B, A

12. A certain sample of concentrated hydrochloric acid contains 50% HCl by mass and has density 1.20 gcm^{-3}. What is the molarity of this sample ?

(1) 16.4 M (2) 8.2 M (3) 32.8 M (4) 13.4 M

13. Given 1 dm^3 of 0.15 MHCl and 1 dm^3 of 0.40 MHCl. What is the maximum volume of 0.25 M HCl which can be made by directly mixing the two solutions without adding water.

(1) 1 dm^3 (2) 2 dm^3 (3) 0.667 dm^3 (4) 1.667 dm^3

14. A solution of H_2SO_4 is 80% by weight, having specific gravity 1.73 its normality is :

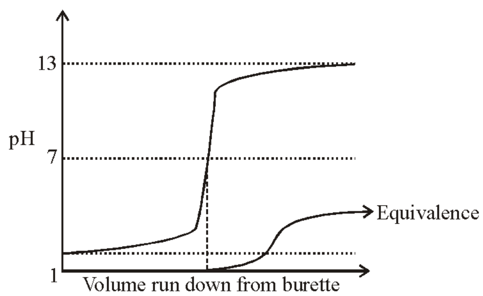
(1) 18.0 (2) 28.2 (3) 1.0 (4) 10.0

15. 6.02×10^{20} molecules of urea are present in 100 ml of its solution. The concentration of urea solution is

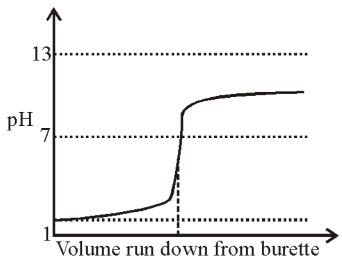
(1) 0.001 M (2) 0.01 M (3) 0.02 M (4) 0.1 M

16. An acid-base titration is a technique where a solution of known concentration of acid/base is used to determine the concentration of an unknown solution of acid/base. These titrations typically use a pH indicator solution to denote the end point of the reaction. A pH indicator is a compound added in small quantities to a solution to indicate the pH visually (generally by appearance/disappearance or change in colour). A typical procedure is as follows:

A certain volume 'V_1' of unknown concentration 'M_1' of HCl is taken in a conical flask, to which a few drops of phenolphthalein indicator solution is added. The solution remains colourless. From a burette (a graduated dropper) a solution of NaOH, whose concentration is known, 'M_2', is added dropwise into the conical flask until a pale pink colour is obtained and is termed as the end point. The amount of solution dispensed from the burette to obtain the end point is noted as 'V_2'. Phenolphthalein indicator changes its colour to pink only when the pH of the solution is above 9.5. Similarly, another indicator, methyl orange, is red in colour below pH 3.7 and yellow above. Give below is graph of pH of the solution in the conical flask and the reading of the burette in the course of the titration. The equivalent point is theoretically defined as the point in the graph where the number of moles HCl in the conical flask becomes equal to the number of moles of NaOH run down the burette. Note the difference between end point and equivalence point.



Which is the correct graph that represents the titration of NH_4OH (from burette) with HCl ?

(1)


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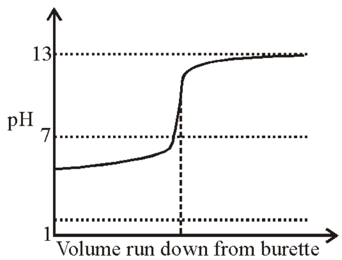

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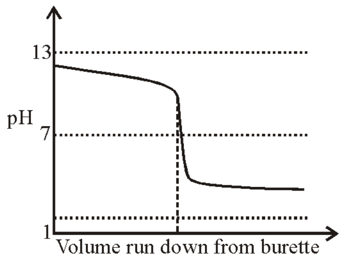

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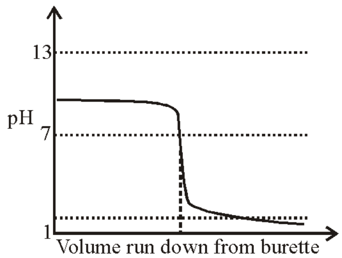


17. The pink colour of phenolphthalein in alkaline medium is due to -

(1) Negative ion (2) Positive ion (3) OH^- ions (4) Neutral form

18. For weak acid strong base titration, the indicator used is :-

(1) Potassium di-chromate (2) Methyl orange

(3) Litmus (4) Phenolphthalein

19. From the following in which titration methyl orange is a best indicator :-

(1) CH_3COOH + NaOH (2) $H_2C_2O_4$ + NaOH

(3) HCl + NaOH (4) $CH_3COOH + NH_4OH$

ANSWER KEY

Exercise-I

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	2	2	2	2	2	2	4	2	3	2	2	4	3	1	4	1	1	2	4	2
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	1	3	2	2	2	4	4	1	1	3	2	4	2	2	3	2	1	4	2	1
Que.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	4	3	2	3	1	3	3	1	1	4	3	2	2	3	2	4	2	1	2	3
Que.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	4	3	1	3	3	2	4	3	3	4	2	1	2	3	1	4	1	4	2	4
Que.	81	82	83	84	85	86	87	88	89	90	91	92	93	94	95	96	97	98	99	100
Ans.	2	2	2	3	2	1	3	1	2	1	4	2	4	3	2	2	4	3	1	2
Que.	101	102	103	104	105	106	107	108	109	110										
Ans.	2	4	3	1	4	2	3	2	4	4										

Exercise-II

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	2	4	3	2	1	3	2	1	3	2	3	3	1	3	1	1	4	1	1	1
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	1	4	3	2	1	2	1	4	3	1	2	1	1	4	4	2	1	4	4	3
Que.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	4	3	1	1	4	3	1	4	3	1	4	1	2	2	1	2	1	2	1	1
Que.	61	62	63	64	65	66	67	68	69	70	71	72	73	74	75	76	77	78	79	80
Ans.	4	3	2	2	3	1	1	2	3	3	3	3	1	3	2	2	2	3	4	3

Exercise-III

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	2	3	3	2	4	1	2	1	2	4	1	1	2	2	2	1	1	3	3	2

REDOX REACTION

INTRODUCTION

- Many Chemical reactions involved in transfer of electrons from one chemical substance to another. These electron-transfer reactions are termed as **Oxidation-Reduction** or **Redox reactions.**

Or

- Those reactions which involves oxidation and reduction both simultaneously are known as oxidation and reduction or Redox Reaction.

Or

- Those reaction which involves increase in oxidation number and decrease in oxidation number simultaneously are known as Oxidation & Reduction or Redox Reaction.

OXIDATION AND REDUCTION

- There are four concepts for oxidation and reduction reactions.

(A) Classical concept (B) Modern concept (C) Valency concept (D) Oxidation number concept.

(A) Classical concept:- According to this concept, oxidation and reduction can be explained as:

(a) Oxidation:- Oxidation is a process which involves,

(i) **Addition of oxygen**

Ex. $2Mg + O_2 \longrightarrow 2MgO$ (Oxidation of Magnesium)

(ii) **Removal of hydrogen**

Ex. $H_2S + Cl_2 \longrightarrow 2HCl + S$ (Oxidation of Hydrogen sulphide)

(iii) **Addition of electronegative element**

Ex. $Fe + S \longrightarrow FeS$ (Oxidation of iron)

(iv) **Removal of electropositive element**

Ex. $2NaI + H_2O_2 \longrightarrow 2NaOH + I_2$ (Oxidation of NaI)

(v) **Increment in oxidation state of electropositive element**

Ex. $SnCl_2 + Cl_2 \longrightarrow SnCl_4$ (Oxidation of $SnCl_2$)

Sn^{+2} $\qquad$ Sn^{+4}

(b) Reduction: Reduction is a process which involves,

(i) **Removal of oxygen**

Ex. $CuO + C \longrightarrow Cu + CO$ (Reduction of cupric oxide)

(ii) **Addition of hydrogen**

Ex. $N_2 + 3H_2 \longrightarrow 2NH_3$ (Reduction of Nitrogen)

(iii) **Removal of electronegative element**

Ex. $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$ (Reduction of mercuric chloride)

(iv) **Addition of electropositive element**

Ex. $HgCl_2 + Hg \longrightarrow Hg_2Cl_2$ (Reduction of mercuric chloride)

(v) **Decrement in oxidation state of electropositive element**

Ex. $CuSO_4 + Fe \longrightarrow FeSO_4 + Cu$ (Reduction of $CuSO_4$)

Cu^{+2} $\qquad$ Cu°

(B) Modern concept or Electronic concept

(i) **Oxidation:** According to this concept the process which involves the loss of one or more electrons from an atom or an ion or molecule is called oxidation (de-electronation).

Ex. $M \longrightarrow M^{+n} + ne^-$

$Mg \longrightarrow Mg^{++} + 2e^-$

(ii) **Reduction:-** According to this concept, the process which involves gain of one or more electrons by an atom or ion or molecule is called reduction (electronation).

Ex. $M^{+n_2} + (n_2 - n_1)e^- \longrightarrow M^{+n_1}$

(C) Valency concept of oxidation and reduction

(i) **Oxidation:** According to this concept, it is the process in which increase in (+) ve valency or decrease in (–) ve valency of a substance takes place.

Ex. $2Mg + O_2 \longrightarrow 2MgO$

(0) (2)

(ii) **Reduction:-** According to this concept, the process in which (+) ve valency decreases whereas (–) ve valency increases.

Ex. $2HgCl_2 + SnCl_2 \longrightarrow Hg_2Cl_2 + SnCl_4$

(+2) (+1)

(D) Oxidation number concept

(i) **Oxidation:-** According to this concept, the oxidation no. of an element in a reaction increases.

(ii) **Reduction:-** According to this concept, the oxidation no. of an element in a reaction decreases.

Example

$$6H^+ + 2MnO_4^- + 5H_2S \longrightarrow 2Mn^{+2} + 5S + 8H_2O$$

Mn^{+7} S^{-2} Mn^{+2} S°

Special Note

- Redox reaction involves two half reactions, one involving loss of electron called half oxidation reaction and the other involving gain of electron called half reduction reaction.

$-2e^-$

Oxidation

Ex. $2Hg^{+2} + Sn^{+2} \longrightarrow Hg_2^{+2} + Sn^{+4}$

$-2e^-$

Reduction

OXIDISING AGENTS OR OXIDANTS (O. A.)

It is the substance which accepts electrons in a chemical reaction i.e., electron acceptors are oxidising agent.

(i) Oxidising agents are lewis acids.

(ii) Substances which can oxidises others and reduces themselves.

(iii) Substances which shows the decrement in oxidation number.

Some important oxidising agents or oxidant

- All high electronegative elements like N, O, F, Cl etc.
- All metallic oxides like Li_2O, Na_2O, Na_2O_2, CO_2, CaO, MgO, BaO_2 etc.
- Some nonmetallic oxides like CO_2, SO_2, H_2O_2, O_3.
- All neutral compound or ion in which element shows their highest oxidation number or state, will act as oxidant or oxidising agent like $KMnO_4$, H_2SO_4, $SnCl_4$,H_3PO_4, $K_2Cr_2O_7$, $HClO_4$, $CuCl_2$, HNO_3, H_2SO_5, $FeCl_3$, $HgCl_2$, etc.

REDUCING AGENT OR REDUCTANT (R. A.)

- The substance which donates electrons in a chemical reaction is called reducing agent i.e., electron donors are reducing agents.

(i) Reducing agents are lewis bases.

(ii) Substances which can reduce other and oxidise themselves.

(iii) Substances which show the increment in oxidation number.

Some important reducing agent or reductant

- All metals like, K, Mg, Ca, etc.
- All metallic hydrides like NaH, CaH_2, $LiAlH_4$, $NaBH_4$, AlH_3 etc.
- All hydraacids like HF, HCl, HBr, H_2S, etc.

- Some organic compounds like Aldehyde, formic acid, oxalic acid, tartaric acid.
- All neutral compounds or ions, in which element shows their lowest oxidation no. or state, will act as reductant or reducing agent

  MnO, $HClO$, $HClO_2$, H_3PO_2, HNO_2, H_2SO_3, $FeCl_2$, $SnCl_2$, Hg_2Cl_2, CH_2Cl_2 etc.

Special points

- Some important compound which can act as oxidant and reductant both.

  HNO_2, SO_2, H_2O_2, O_3, Al_2O_3, Cr_2O_3, MnO_2, ZnO, CuO, are called as amphoteric oxide.

OXIDATION NUMBER

- It represents the number of electron gained or lost by atom when it changes into compound from a freestate.
- It is the charge developed on an atom when atom is in combination.
- If electrons are gained by an atom in the formation of compound, oxidation number is given (–) ve sign.
- If electron are lost by an atom in the formation of compound, oxidation number is given (+) ve sign.
- It represents the real charge in case of ionic compounds and represents the imaginary charge in case of covalent compounds.
- Maximum oxidation number of an element is equal to group number which belongs in the periodic table.
- Minimum oxidation number of an element is equal to (group number –8)

I Group elements always shows +1 as oxidation number

II Group elements always shows +2 as oxidation number

III Group elements always shows +3 oxidation number

IV Group shows –4 to +4 oxidation number

V Group shows –3 to +5 oxidation number

VI Group shows –2 to +6 oxidation number

VII Group shows –1 to +7 oxidation number (except F)

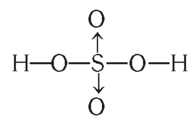
Inert gases always show zero oxidation number

Oxidation number for coordinate bond

(a) When coordinate bond is formed in between low electronegative element and high electronegative element then the e^- donor element show +2 oxidation number whereas e^- acceptor element show –2 oxidation number.

In this type of bond or bonded compounds electronegativity of acceptor > electronegativity of donor.

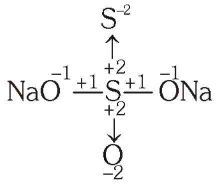
For example in H_2SO_4.



Here 'S' is less electronegative element than O therefore, oxidation number of S = +2 and oxidation number of O = –2 (due to single dative bond).

(b) When coordinate bond is formed between the two same electronegative elements then the e^- donor element show +2 oxidation number where e^- acceptor element show –2 oxidation number. In this type of bond or bonded compound, Electronegativity of acceptor = electronegativity of donor.

For example: In $Na_2S_2O_3$



Here oxidation number of 'S' is +2 because it is e^- donor and the other 'S' is –2, because it is e^- acceptor. (Due to dative bond)

(c) When coordinate bond is formed in between high electronegative element to low electronegative element then no change in oxidation number will be shown by both the elements, which is bonded by coordinate bond e.g. HCN. $\underset{}{H}\overset{+1\ \ -1}{—}\underset{-2}{N^0}\equiv\underset{+2}{C^0}$

OXIDATION STATE

- Oxidation state of an atom is defined as oxidation number per atom for all practical purposes. Oxidation state is often expressed as oxidation number.

The rules to derive oxidation number or oxidation state

(a) The oxidation state of an element in its free state is zero. Example, oxidation state of Na, Cu, I, Cl, O etc. are zero.

(b) Sum of oxidation state of all the atoms in a neutral molecule is zero.

(c) Sum of oxidation state of all the atoms in a complex ion is equal to number of charge present on it.

(d) In complex compounds, oxidation state of some neutral molecules (ligands) is zero. Ex. CO, NO, NH_3,H_2O.

(e) Generally oxidation state of Oxygen is –2 but in H_2O_2 it is –1 & in OF_2 it is +2.

(f) Generally oxidation state of Hydrogen is +1 but in metallic hydrides it is –1.

(g) Generally oxidation state of Halogen atoms is –1 but in interhalogen compounds it changes.

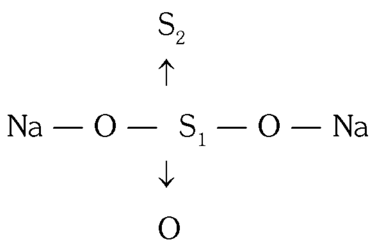
(h) Generally oxidation state of alkali metals is +1 and that of alkaline earth metals is +2.

(i) Oxidation state of transition elements varies from compound to compound. Mn has oxidation state from +1 to +7.

$Mn_2O \rightarrow +1$, $MnO \rightarrow +2$, $Mn_3O_4 \rightarrow +8/3$, $MnO_2 \rightarrow +4$, $Mn_2O_5 \rightarrow +5$, $MnO_4^{-2} \rightarrow +6$, $MnO_4^- \rightarrow +7$.

(j) Oxidation state of an atom may be fractional, negative, zero as well as positive.

Special Note:- Some times same atom in a compound has different oxidation state for example, structures of $Na_2S_2O_3$ is



Here S_1 and S_2 both are sulphur atoms but they have different oxidation state

Oxidation state of S_1 = +6

Oxidation state of S_2 = –2 (It is accepting two electrons from S_1)

Average oxidation state of $S = \frac{+6-2}{2} = +2$

Oxidation state as a periodic property

- Oxidation state of an atom depends upon the electronic configuration of an atom i.e. why it is periodic property.
  (a) I A group of alkali metals show +1 oxidation state.
  (b) II A group or alkaline earth metals show +2 oxidation state
  (c) The maximum normal oxidation state, show by III A group elements is +3. These elements also show +2 and +1 oxidation states.
  (d) Elements of IV A group show their maximum and minimum oxidation states +4 and –4 respectively.
  (e) Non metals show number of oxidation states, the relation between maximum and minimum oxidation states for non metals is equal to (maximum oxidation state – minimum oxidation state = 8).

For example sulphur has maximum oxidation number +6 as being in VI A group element.

Fractional oxidation states

Most of elements show fractional oxidation states. For example oxidation state of oxygen in superoxides of alkali metals (KO_2, SO_2, RbO_2) is –1/2.

e.g. In Fe_3O_4, Fe shows its oxidation state as 8/3 as it is a mixed oxide and can be written as Fe^{II} Fe_2^{III} O_4. Similarly oxidation states of Boron in B_4O_{10} and B_5H_9 is –2.5 and –1.8 respectively.

APPLICATION OF OXIDATION NUMBER

(1) To compare the strength of acid and base:-

Strength of acid $\propto$ Oxidation Number

Strength of base $\propto \dfrac{1}{\text{Oxidation Number}}$

Example: Order of acidic strength in HClO, $HClO_2$, $HClO_3$, $HClO_4$ will be.

Solution: Oxidation Number of chlorine

HClO (Hypo chlorous acid)	+1
$HClO_2$ (Chlorous acid)	+3
$HClO_3$ (Chloric acid)	+5
$HClO_4$ (Perchloric acid)	+7

$\therefore$ Strength of acid α Oxidation Number

So the order will be

$HClO_4 > HClO_3 > HClO_2 > HClO$

(2) To determine the oxidising and reducing nature:-

Group	**Range of Oxidation Number (n – 8) to n**	
I A	+1	
II A	+2	
III A	+1, +3	
IV A	–4 to + 4	
V A	–3 to +5	
VI A	–2 to +6	[Exception → Maximum Oxidation Number of Oxygen is +2]
VII A	–1 to +7	[Exception → Oxidation of F is –1]

- If any compound is in maximum oxidation state (n), then it will act as only oxidant.
- If, compound is in minimum oxidation state (n – 8) then it will act as only reductant.
- If oxidation state is intermediate then compound can act as reductant as well as oxidant both.

Ex. • Oxidation number of S in H_2S = –2 minimum. So act as reducant.

- Oxidation number of N in HNO_2 = +3. So act as reductant as well as oxidant both.
- Oxidising strength $\propto$ Oxidation Number
- Reducing strength $\propto \dfrac{1}{\text{Oxidation Number}}$

(3) To calculate the equivalent weight

Equivalent weights of oxidising and reducing agents.

The equivalent weight of an oxidising agent is that weight which accepts one mole electron in a chemical reaction.

(a) Equivalent weight of an oxidant (get reduced) $= \dfrac{\text{Molecular weight}}{\text{Number of electrons gained by one mole}}$

Ex. In acidic medium

$$6e^- + Cr_2O_7^{2-} + 14H^+ \longrightarrow 2Cr^{3+} + 7H_2O$$

Here atoms which undergoes reduction is Cr its oxidation state is decreasing from +6 to +3

Equivalent weight of $K_2Cr_2O_7 = \dfrac{\text{Molecular weight of } K_2Cr_2O_7}{3\times 2} = \dfrac{\text{Molecular weight}}{6}$

Note: [6 in denominator indicates that 6 electrons were gained by $Cr_2O_7^{2-}$ as it is clear from the given balanced equation]

(b) Similarly equivalent weight of a reducant (gets oxidised) $= \dfrac{\text{Molecular weight}}{\text{Number of electrons lost by one mole}}$

Ex. In acidic medium, $C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^-$

Here, Total electrons lost = 2. So, equivalent weight $= \dfrac{\text{Molecular weight}}{2}$

(c) In different condition a compound may have different equivalent wts. Because, it depends upon the number of electrons gained or lost by that compound in that reaction.

Ex. (i) $MnO_4^- \longrightarrow Mn^{2+}$ (acidic medium)

(+7) (+2)

Here 5 electrons are taken so equivalent weight $= \dfrac{\text{Molecular weight of } KMnO_4}{5} = \dfrac{158}{5} = 31.6$

(ii) $MnO_4^- \longrightarrow Mn^{+2}$ (neutral medium)

(+7) (+2)

Here, only 3 electrons are gained, so equivalent weight $= \dfrac{\text{Molecular weight of } KMnO_4}{3} = \dfrac{158}{3} = 52.7$

(iii) $MnO_4^- \longrightarrow MnO_4^{-2}$ (alkaline medium)

(+7) (+6)

Here, only one electron is gained, so equivalent weight $= \dfrac{\text{Molecular weight of } KMnO_4}{1} = 158$

Note: It is important to note that $KMnO_4$ acts as an oxidant in every medium although with different strength which follows the order –

acidic medium > neutral medium > alkaline medium while, $K_2Cr_2O_7$ acts as an oxidant only in acidic medium as follows

$Cr_2O_7^{2-} \longrightarrow 2Cr^{3+}$

$(2 \times 6) \longrightarrow (2 \times 3)$

Here, 6 electrons are gained so equivalent weight $= \dfrac{\text{Molecular weight of } K_2Cr_2O_7}{6} = \dfrac{294.21}{6} =$ 49.03

(d) It is clear that $KMnO_4$ is better oxidant than $K_2Cr_2O_7$.

The equivalent weight of a reducing agent is the weight which donates one electron in a chemical reaction.

Ex. $2S_2O_3^{--} \longrightarrow S_4O_6^{--} + e^-$

(R. A.)

equivalent weight of $S_2O_3^{--} = \frac{2M}{2} = M$

(4) To determine the molecular formula of compound

Ex. Suppose that there are three atoms A, B, C and their oxidation number are +6, –1, –2, respectively. Then the molecular formula of compound will be.

Sol. Since, the charge on a free compound is zero. So

$+6 = (-1 \times 4) + (-2)$

$+6 = -6$

or $+6 = (-1 \times 2) + (-2 \times 2)$

$= -2 + (-4) = -6$

So molecular formula, AB_4C or AB_2C_2.

Special points:

- When metals participate in chemical reaction, they are always oxidised. Thus metals always behave like reducing agents.
- The acidic nature of non-metal oxide increases with increasing oxidation state of the non-metal.
- The strength of oxy-acid of an element increases with increasing oxidation state of the element.

Strength of oxy-acids of nitrogen increases in the following order.

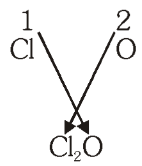
+1	+3	+5
$H_2N_2O_2$	$< HNO_2$	$< HNO_3$
Hyponitrous acid	Nitrous acid	Nitric acid
(weakest acid)		(strongest acid)

- An element in its lowest state in its compound behave like a reducing agent only.
- Metals exist in mixed oxidation states in non-stoichiometric compounds.

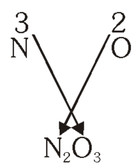
For example: Oxidation state of Fe in $Fe_{0.94}O$ is +2 and +3

- The oxidation state of iron in ferro compounds is +2 while in ferric compounds, it is in +3 oxidation state.
- The formula of the anhydride of an oxy-acid of an element can be calculated as follows :

Oxy-acid $H^{+1}ClO$ Acid-anhydride



Oxy-acid $H^{+1}NO_2$ Acid anhydride



REDOX REACTIONS

- The reactions in which oxidation and reduction both occur simultaneously are called redox reactions.
  - Most of the chemical reactions are redox because if one element is there to lose electrons, other element has to be there to accept them.
  - Any redox reaction may be divided in two parts.
  - Oxidation half reaction.
  - Reduction half reaction.

  Now, we will study some reaction.

S. No.	Reaction	Oxidant (Getting Reduced)	Reductant (Getting Oxidised)
1.	$C + O_2 \longrightarrow CO_2$	O $[0 \longrightarrow -2]$	C $[0 \longrightarrow +4]$
2.	$PbS + 4H_2O_2 \longrightarrow PbSO_4 + 4H_2O$	O $[-1 \longrightarrow -2]$	S $[-2 \longrightarrow +6]$
3.	$Sn + 2F_2 \longrightarrow SnF_4$	F $[0 \longrightarrow -1]$	Sn $[0 \longrightarrow +4]$
4.	$SO_2 + 2H_2O + Cl_2 \longrightarrow 2HCl + H_2SO_4$	Cl $[0 \longrightarrow -1]$	S$[+4 \longrightarrow +6]$
5.	$I_2 + 10HNO_3 \longrightarrow 2HIO_3 + 10NO_2 + 4H_2O$	N $[+5 \longrightarrow +4]$	I $[0 \longrightarrow +5]$
6.	$CuO + H_2 \longrightarrow Cu + H_2O$	Cu $[+2 \longrightarrow 0]$	H $[0 \longrightarrow +1]$
7.	$2KMnO_4 + 3H_2SO_4 + 5H_2S \longrightarrow K_2SO_4 + 2MnSO_4 + 8H_2O + 5S$	Mn $[+7 \longrightarrow +2]$	S $[-2 \longrightarrow 0]$
8.	$H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2$ (Oxygen of H_2O_2)	Ag $[+1 \longrightarrow 0]$	O $[-1 \longrightarrow 0]$
9.	$H_2SO_4 + 2HI \longrightarrow SO_2 + I_2 + 2H_2O$	S $[+6 \longrightarrow +4]$	I $[-1 \longrightarrow 0]$

Types of Redox reaction

(a) Intermolecular redox reaction:- When oxidation and reduction take place separately in the different compounds, called intermolecular redox reaction.

$$SnCl_2 + 2FeCl_3 \longrightarrow SnCl_4 + 2FeCl_2$$

$Sn^{+2} \longrightarrow Sn^{+4}$ (Oxidation)

$Fe^{+3} \longrightarrow Fe^{+2}$ (Reduction)

(b) Intra molecular redox reaction:- During the chemical reaction, if oxidation and reduction take place in single compound then reaction is called intramolecular redox reaction.

$$2KClO_3 \longrightarrow KCl + 3O_2$$

$Cl^{+5}O^{-2}$ $\quad$ Cl^{-1} $\quad$ O°

Reduction ($Cl^{+5} \longrightarrow Cl^{-1}$)

Oxidation ($O^{-2} \longrightarrow O^{\circ}$)

(c) Disproportionation reaction:- When reduction and oxidation takes place on same element of a compound is called disproportionation reaction.

$$H_2O_2 \longrightarrow H_2O + 1/2\ O_2$$

O^{-1} $\quad$ O^{-2} $\quad$ O°

Reduction ($O^{-1} \longrightarrow O^{-2}$)

Oxidation ($O^{-1} \longrightarrow O^{\circ}$)

REDOX REACTION

EXERCISE-I

OXIDATION AND REDUCTION

1. Oxidation involves –
(1) Loss of electron (2) Gain of electron
(3) Increase in valency of negative part (4) All

2. If in a chemical reaction, Fe^{+2} is converted into Fe^{+3}, then Fe^{+2} –
(1) Gains one electron (2) Loses one electron (3) Gains one proton (4) Loses one proton

3. In the equation $PbO_2 + 4HCl \rightarrow PbCl_2 + 2H_2O + Cl_2$, the substance undergoing oxidation is_______.
(1) lead dioxide (2) hydrochloric acid (3) hydrogen (4) lead chloride

4. Which of the following statement is incorrect?
(1) In oxidation, oxygen is added to a substance.
(2) In reduction, hydrogen is added to a substance.
(3) Oxidizing agent is oxidized.
(4) Reducing agent is oxidized.

5. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
Zn undergoes –
(1) Reduction (2) Oxidation
(3) Both oxidation and reduction (4) None of these

OXIDATION NUMBER

6. Oxidation number of Mn in K_2MnO_4 is –
(1) +7 (2) –5 (3) + 6 (4) – 6

7. The oxidation state of I in $H_4IO_6^-$ is :–
(1) +7 (2) –1 (3) +5 (4) +1

8. The oxidation state of osmium (Os) in OsO_4 is :
(1) + 7 (2) +6 (3) + 4 (4) + 8

9. In the change $Ce^{+3} \longrightarrow Ce^{+4}$
(1) Oxidation number increases (2) Oxidation number decreases
(3) Reduction takes place (4) None of the above

10. The oxidation state of C in CO_2 is
(1) –4 (2) +4 (3) +1 (4) –2

11. The oxidation state of Boron in the compound KBF_4 is
(1) +3 (2) +2 (3) + (4) –3

12. Which of the following has minimum oxidation state of chlorine ?
(1) Cl_2 (2) Cl^- (3) ClO_2^- (4) ClO_4^-

13. Oxygen ion structure in its peroxide, superoxide, ozonide form is :
(1) O_2^-, O_2^2, O_3^{-2} (2) O_2^{-2}, O_2^-, O_3^- (3) O_2^{-2}, O^{-2}, O_3^- (4) O_2^-, O_2^{-3}, O_3^{-2}

14. Oxidation number of S in H_2SO_4 is :
(1) 0 (2) 2 (3) 4 (4) 6

15. Oxidation number of Xe in XeF_5^- is :-
(1) +1 (2) +2 (3) +3 (4) +4

16. Which has zero oxidation state?
(1) Cd (2) Cl_2 (3) (1) & (2) (4) None

17. Oxidation state of Cr in $Cr_2O_7^{-2}$ is –

(1) +2 (2) +6 (3) –6 (4) –4

18. What is the oxidation number of the element in free state?

(1) 0 (2) +1 (3) –1 (4) None

19. What is the oxidation number of O in KO_2?

(1) 0 (2) $-\frac{1}{2}$ (3) –1 (4) –2

20. Manganese achieves its highest oxidation state in which of the following compounds–

(1) MnO_3 (2) Mn_3O_4 (3) $KMnO_4$ (4) K_2MnO_4

OXIDISING/REDUCING AGENT AND REDOX REACTION

21. The characteristic property of reductant is –

(1) It accepts electron (2) Gives up electrons (3) Both the above (4) None of the above

22. How many electrons have been transferred in the process : $2Fe^{+3} + 2Hg \longrightarrow Hg_2^{+2} + 2Fe^{+2}$?

(1) 2 electron (2) 3 electron (3) 4 electron (4) 6 electron

23. Which substance act as reducing agent in the following reaction?

$14H^+ + Cr_2O_7^{-2} + 3Ni \longrightarrow 2Cr^{+3} + 7H_2O + 3Ni^{+2}$

(1) H_2O (2) Ni (3) H^+ (4) $Cr_2O_7^{-2}$

24. In the reaction $2Ag + 2H_2SO_4 \rightarrow Ag_2SO_4 + 2H_2O + SO_2$, H_2SO_4 acts as:

(1) Reducing agent (2) Hydrolysing agent (3) Dehydrating agent (4) Oxidizing agent

25. Which quantities are conserved in all oxidation reduction reaction?

(1) Charge only (2) Mass only

(3) Both charge and mass (4) None

26. In acidic medium the oxidation number of Mn in $KMnO_4$ shows the following change –

(1) $+7 \rightarrow +4$ (2) $+7 \rightarrow +2$ (3) $+7 \rightarrow +3$ (4) $+7 \rightarrow +5$

27. In the following reaction :

$3Br_2 + 6CO_3^{-2} + 3H_2O \rightarrow 5Br^- + BrO_3^- + 6HCO_3^-$

(1) Bromine is oxidised and carbonate is reduced

(2) Bromine is reduced and carbonate is oxidised

(3) Bromine is neither reduced nor oxidised

(4) Bromine is reduced and oxidised both

28. MnO_4^- on reduction in acidic medium forms –

(1) MnO_4 (2) Mn^{++} (3) MnO_4^{-2} (4) Mn

29. $KMnO_4$ acts as an oxidant in neutral, basic and acidic medium. Its final product under these conditions would be respectively

(1) MnO_2, MnO_2, Mn^{+2} (2) MnO_4^{-2}, Mn^{+3}, Mn^{+2} (3) MnO_2, MnO_4^{-2}, Mn^{+2} (4) MnO, MnO_2, Mn^{+2}

30. The reaction,

$2H_2O(l) \rightarrow 4H^+ (aq) + O_2(g) + 4e^-$ is :

(1) An oxidation reaction (2) A reduction reaction

(3) A redox reaction (4) A hydrolysis reaction

31. Which of the following equation represents reducing property of SO_2?

(1) $2H_2S + SO_2 \rightarrow 3S + 2H_2O$

(2) $3Fe + SO_2 \rightarrow 2FeO + FeS$

(3) $K_2Cr_2O_7 + SO_2 + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2O$

(4) All of these

32. Select the example of disproportionation reaction

(1) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$ (2) $NH_4NO_3 \rightarrow N_2O + 2H_2O$

(3) $4H_3PO_3 \rightarrow PH_3 + 3H_3PO_4$ (4) $AgCl + 2NH_3 \rightarrow Ag(NH_3)_2Cl$

33. HNO_3 can act as :

(1) Acid (2) Oxidizing agent (2) Reducing agent (4) Both 1 & 2

34. $Ag_2O + H_2O + 2e^- \rightarrow 2Ag + 2OH^-$

(1) Water is oxidised (2) Electrons are reduced

(3) Silver is oxidised (4) Silver is reduced

35. Which of the following is a redox reaction ?

(1) $NaCl + KNO_3 \rightarrow NaNO_3 + KCl$ (2) $CaC_2O_4 + 2\,HCl \rightarrow CaCl_2 + H_2C_2O_4$

(3) $Zn + 2AgCN \rightarrow 2\,Ag + Zn(CN)_2$ (4) $Mg(OH)_2 + 2NH_4Cl \rightarrow MgCl_2 + 2\,NH_4OH$

BALANCING OF REDOX REACTIONS

36. $a\,Zn + b\,NO_3^- + cH^+ \rightarrow dNH_4^+ + e\,H_2O + f\,Zn^{+2}$ a, b, c, d, e and f are –

	a	b	c	d	e	f
(1)	2	4	6	8	4	2
(2)	1	4	10	3	1	4
(3)	4	1	10	1	3	4
(4)	10	4	1	3	4	2

37. What is the value of *n* in the following equation : $Cr(OH)_4^- + OH^- \rightarrow CrO_4^{2-} + H_2O + ne$?

(1) 3 (2) 6 (3) 5 (4) 2

38. $(x)\ MnO^-_4 + (y)\ H_2O_2 + (z)H^+ \rightarrow 2\ Mn^{+2} + 8H_2O + 5\ O_2$

in this reaction, value of (x), (y) and (z) are :

(1) 2, 5, 6 (2) 5, 2, 9 (3) 3, 5, 6 (4) 2, 6, 6

39. Choose the set of coefficients that correctly balances the following equation :

$x\,Cr_2O_7^{2-} + yH^+ + z\,e^- \rightarrow a\,Cr^{+3} + bH_2O$

	x	y	z	a	b
(1)	2	14	6	2	7
(2)	1	14	6	2	7
(3)	2	7	6	2	7
(4)	2	7	6	1	7

40. The number of electrons required to balance the following equation –

$NO_3^- + 4H^+ + e^- \longrightarrow 2H_2O + NO$ is

(1) 5 (2) 4 (3) 3 (4) 2

EQUIVALENT WEIGHT

41. Equivalent mass of oxidising agent in the reaction, $SO_2 + 2H_2S \rightarrow 3S + 2H_2O$ is :

(1) 32 (2) 64 (3) 16 (4) 8

42. As per the equation, $Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$, the equivalent mass of sodium carbonate is

(1) 106 (2) 10.6 (3) 53 (4) 5.3

43. Which of the following have n-factor equal to one

(1) H_3PO_2 (2) KOH (3) HC1 (4) All

44. n-factor of $NaHCO_3$ as a base is

(1) 0 (2) 1 (3) 2 (4) None

45. The Mw of a oxide of an element is 44. The Ew of the oxide of element is 14. The atomic weight of the element is

(1) 14 (2) 28 (3) 42 (4) 56

EXERCISE-II

OXIDATION AND REDUCTION

1. Which is the best description of the behaviour of bromine in the reaction given below?

$H_2O + Br_2 \rightarrow HOBr + HBr$

(1) Both oxidized and reduced (2) Oxidized only

(3) Reduced only (4) Proton acceptor only

2. Which of the following equations is not correctly balanced ?

(1) $Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$ (2) $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$

(3) $A1 + 2H^+ \rightarrow A1^{3+} + H_2$ (4) $Br_2 + 2I^- \rightarrow Br^- + I_2$

3. In the reaction. $Cl_2 + OH^- \rightarrow Cl^- + ClO_3^- + H_2O$ chlorine is ________.

(1) Oxidised (2) Reduced

(3) Oxidised as well as reduced (4) Neither oxidised nor reduced

4. Of the following reactions, only one is a redox reaction. Identify it :-

(1) $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$ (2) $BaCl_2 + MgSO_4 \rightarrow BaSO_4 + MgCl_2$

(3) $2S_2O_7^{2-} + 2H_2O \rightarrow 4SO_4^{2-} + 4H^+$ (4) $Cu_2S + 2FeO \rightarrow 2Cu + 2Fe + SO_2$

5. In the reaction $2FeCl_3 + SnCl_2 \longrightarrow 2FeCl_2 + SnCl_4$,

(1) Fe undergoes oxidation (2) Sn undergoes oxidation

(3) Sn undergoes reduction (4) None of the above

OXIDATION NUMBER

6. The oxidation number of sulphur in $Na_2S_4O_6$ is –

(1) 1½ (2) 2½ (3) 3 (4) 2

7. Which of the following has zero oxidation state of Fe :-

(1) $[Fe(CN)_6]^{-4}$ (2) $[Fe(CN)_6]^{-3}$ (3) $Fe(CO)_5$ (4) All of the above

8. Oxidation number of P in pyrophosphoric acid ($H_4P_2O_7$) is :-

(1) +5 (2) +2 (3) +3 (4) +4

9. Oxidation numbers of A, B and C are +2, +5 and –2 respectively. Possible formula of the compound is :-

(1) $A_2(BC_2)_2$ (2) $A_3(BC_4)_2$ (3) $A_2(BC_3)_2$ (4) $A_3(B_2C)_2$

10. The oxidation states of sulphur in the anions SO_3^{2-}, $S_2O_4^{2-}$ and $S_2O_6^{2-}$ follow the order:

(1) $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$ (2) $SO_3^{2-} < S_2O_4^{2-} < S_2O_6^{2-}$

(3) $S_2O_4^{2-} < S_2O_6^{2-} < SO_3^{2-}$ (4) $S_2O_6^{2-} < S_2O_4^{2-} < SO_3^{2-}$

11. The correct order of acidic strength is –

(1) $HClO_4 < HClO_3 < HClO_2 < HClO$ (2) $HClO_2 < HClO_3 < HClO_4 < HClO$

(3) $HClO_4 < HClO < HClO_2 < HClO_3$ (4) $HClO < HClO_2 < HClO_3 < HClO_4$

12. Oxidation number of N in NH_4NO_3 is

(1) –3 (2) +5 (3) –3 & +5 (4) +3 & –5

13. The zero oxidation state of carbon is observed in the compound :

(1) CH_2O (2) CO_2 (3) CH_4 (4) $CHBr_3$

14. The oxidation state of Ni and NH_3 in $[Ni(NH_3)_6]^{+2}$:

(1) Ni = +2, NH_3 = 0 (2) Ni = +1, NH_3 = – 1/6

(3) Ni = +1, NH_3 = + 1/6 (4) Ni = 0, NH_3 = +2

15. Oxidation state of Fe in K_3 $[Fe(CN)_6]$

(1) +2 (2) +3 (3) 0 (4) None

16. In the conversion $Br_2 \rightarrow BrO_3^-$. Oxidation state of bromine changes from :-

(1) 0 to 5 (2) 1 to 3 (3) 2 to 4 (4) 5 to 1

17. $MnO_2 + 4H^+ + e^- \rightarrow Mn^{+3} + 2H_2O$

In the above reaction oxidation number of Mn changes from :-

(1) + 2 to – 3 (2) + 4 to – 3 (3) + 2 to + 3 (4) + 4 to + 3

18. Oxidation number of nitrogen is –1 in which of the following compounds

(1) NH_3 (2) NH_4OH (3) NH_2OH (4) NH_4^+

19. Oxidation number of P in $Mg_2P_2O_7$ is

(1) +3 (2) +2 (3) +5 (4) –3

20. The oxidation state of nickel in $Ni(CO)_4$ is –

(1) 1 (2) 0 (3) 2 (4) 4

21. The oxidation number of iron in Fe_3O_4 is :

(1) +2 (2) +3 (3) +8/3 (4) +2/3

22. Among the following, identify the species with an atom in +6 oxidation state :

(1) MnO_4^- (2) $Cr(CN)_6^{3-}$ (3) NiF_6^{2-} (4) CrO_2Cl_2

23. Identify the pair of binary compounds in which nitrogen exhibits the lowest and the highest oxidation state:

(1) NH_3 NO_2 (2) NH_3, N_2O_5 (3) N_2, HNO_3 (4) N_2O, N_2O_5

24. In the reaction of H_2SO_4 with H_2S, the change in oxidaition states of sulphur are

(1) +6 to +4 and –2 to 0 (2) +4 to +3 and +2 to 0

(3) +6 to –2 and +6 to 4 (4) +2 to +6 and 0 to –2

25. What is the oxidation number of O in Na_2O_2?

(1) 0 (2) $-\frac{1}{2}$ (3) –1 (4) –2

OXIDISING/REDUCING AGENT AND REDOX REACTION

26. Which of the following oxide cannot work as reducing agent?

(1) ClO_2 (2) SO_2 (3) NO_2 (4) CO_2

27. Which one of the following is a redox reaction?

(1) Reaction of H_2SO_4 and NaOH

(2) Evaporation of water

(3) The formation of ozone by lightning of oxygen of atmosphere

(4) The preparation of nitrogen oxide from atmospheric nitrogen & oxygen

28. In which of the following case, metal is reduced [These are not balance equations]

(1) $[Fe(CN)_6]^{-4} \rightarrow [Fe(CN)_6]^{-3}$ (2) $MnO_4^- \rightarrow MnO_2$

(3) $MnO_4^{-2} \rightarrow MnO_4^{-1}$ (4) $[Cr_2O_7]^{-2} \rightarrow [CrO_4]^{-2}$

29. Which of the following act both as oxidant & reductant?

(1) H_2S (2) SO_3 (3) H_2O_2 (4) F_2

30. Which of the following reaction is spontaneous oxidation–reduction reaction?

(1) $Mn^{+2} + 5Fe^{+3} + 4H_2O \rightarrow MnO_4^- + 5Fe^{+2} + 8H^+$

(2) $MnO_4^- + 5Fe^{+3} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{+2} + 4H_2O$

(3) $MnO_4^- + 5Fe^{+2} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{+3} + 4H_2O$

(4) $Mn^{+2} + 5Fe^{+2} + 4H_2O \rightarrow MnO_4^- + 5Fe^{+3} + 8H^+$

31. In the following reaction :

$P_4 + 3NaOH + 3H_2O \longrightarrow PH_3 + 3NaH_2PO_2$

(1) Phosphorus is oxidised
(2) Phosphorus is oxidised as well as reduced
(3) Phosphorus is reduced
(4) Sodium is oxidised

32. Which of the following substances can act as only reducing agent?

(1) Sulphur (2) Hydrogen sulphide (3) SO_2 (4) SO_3

33. Bleaching powder contains which of the following anions ?

(1) Cl^-, OCl^- (2) ClO_3^-, Cl^- (3) ClO_3^-, OCl^- (4) OCl_2^{-2}

34. In the preparation of Cl_2 from HCl and MnO_2, MnO_2 acts as :-

(1) Catalyst (2) Reducing agent (3) Oxidising agent (4) None

35. In the redox reaction : $2FeCl_3 + SnCl_2 \longrightarrow 2FeCl_2 + SnCl_4$.The oxidant and reductant are respectively :-

(1) $FeCl_3$, $SnCl_4$ (2) $FeCl_3$, $SnCl_2$ (3) $SnCl_2$, $FeCl_3$ (4) $SnCl_2$, $FeCl_2$

36. Check whether the reaction $H_2S+SO_2 \rightarrow H_2O + S$ is a redox reaction or not ? If yes, what will be the oxidant

(1) Yes, SO_2 (2) Yes, S (3) Yes, H_2S (4) No, this is not redox

37. The reaction in which hydrogen peroxide acts as a reducing agent is :

(1) $PbS + 4H_2O_2 \rightarrow PbSO_4 + 4H_2O$
(2) $2KI + H_2O_2 \rightarrow 2KOH + I_2$
(3) $2FeSO_4 + H_2SO_4 + H_2O_2 \rightarrow Fe_2(SO_4)_3 + 2H_2O$
(4) $Ag_2O + H_2O_2 \rightarrow 2Ag + H_2O + O_2$

38. Which of the following is a powerful reducing agent?

(1) HNO_3 (2) Na (3) Cl_2 (4) Cr

39. H_2S is acidic in nature while H_2O is neutral in nature. Give reason.

(1) Oxygen is more electronegative than sulphur
(2) H – S bond is weaker than O – H bond.
(3) Bond angle in H_2S is less than that in H_2O
(4) H – S bond is stronger than O – H bond

40. In which of the following reaction H_2O_2 acts as reducing agent :-

(1) $2FeCl_2 + 2HCl + H_2O_2 \rightarrow 2FeCl_3 + 2H_2O$
(2) $Cl_2 + H_2O_2 \rightarrow 2HCl + O_2$
(3) $2HI + H_2O_2 \rightarrow 2H_2O + I_2$
(4) $H_2SO_3 + H_2O_2 \rightarrow H_2SO_4 + H_2O$

BALANCING OF REDOX REACTIONS

41. Consider the following reaction,

$5H_2O_2 + xClO_2 + 2OH^- \rightarrow x\,Cl^- + yO_2 + 6H_2O$

The reaction is balanced if :

(1) x = 5, y = 2 (2) x = 2, y = 5 (3) x = 4, y = 10 (4) x = 5, y = 5

42. In oxidation reduction reaction

$MnO_4^- + C_2O_4^{-2} + H^+ \rightarrow Mn^{+2} + CO_2 + H_2O$ the balance equation coefficient for MnO_4^-, $C_2O_4^{-2}$, H^+ are :–

(1) 2, 5, 16 (2) 16, 5, 2 (3) 2, 16, 5 (4) 5, 2, 16

43. $2KMnO_4+5H_2S+6H^+ \rightarrow 2Mn^{2+} + 2K^+ + 5S + 8H_2O$. In the above reaction, how many electrons would be involved in the oxidation of 1 mole of reductant?

(1) Two (2) Five (3) Ten (4) One

44. For the redox reaction

$Zn + NO_3^- \rightarrow Zn^{2+} + NH_4^+$ in basic medium, coefficients of Zn, NO_3^- and OH^- in the balanced equation respectively are :

(1) 4, 1, 7 (2) 7, 4, 1 (3) 4, 1, 10 (4) 1, 4, 10

45. Balance gven following half reaction for the unbalanced whole reaction :

$CrO_4^{2-} \rightarrow CrO_2^- + OH^-$ is :

(1) $CrO_4^{-2} + 2H_2O + 3e^- \rightarrow CrO_2^- + 4OH^-$ (2) $2CrO_4^{-2} + 8H_2O \rightarrow CrO_2^- + 4H_2O + 8OH^-$

(3) $CrO_4^{-2} + H_2O \rightarrow CrO_2^- + H_2O + OH^-$ (4) $3CrO_4^{-2} + 4H_2O + 6e^- \rightarrow 2CrO_2^{-1} + 8OH^-$

EQUIVALENT WEIGHT

46. In acidic medium potassium dichromate acts as an oxidant according to the equation $Cr_2O_7^{-2} + 14H^+ + 6e^- \longrightarrow 2Cr^{+3} + 7H_2O$. What is the equivalent weight of $K_2Cr_2O_7$? (molecular weight=M)

(1) M (2)M/2 (3) M/3 (4) M/6

47. 1.0 g of a monobasic acid when completely acted upon Mg gave 1.301 g of anhydrous Mg salt. Equivalent weight of acid is

(1) 35.54 (2) 36.54 (3) 17.77 (4) 18.27

48. The Ew of H_3PO_4, in the reaction is : $Ca(OH)_2 + H_3PO_4 \rightarrow CaHPO_4 + 2H_2O$

(Ca = 40, P = 31, O = 16)

(1) 49 (2) 98 (3) 32.66 (4) 147

49. What weight of a metal of equivalent weight 12 will give 0.475 g of its chloride?

(1) 0.12 g (2) 0.24 g (3) 0.36 g (4) 0.48 g

50. $N_2 + 3H_2 \rightarrow 2NH_3$.

Molecular weight of NH_3 and N_2 are x_1 and x_2, respectively. Their equivalent weights are y_1 and y_2 respectively. Then $(y_1 - y_2)$ is

(1) $\left(\frac{2x_1 - x_2}{6}\right)$ (2) $(x_1 - x_2)$ (3) $(3x_1 - x_2)$ (4) $(x_1 - 3x_2)$

EXERCISE-III

QUESTIONS ASKED IN PREVIOUS EXAMS

1. The order of corrosion of metals, namely aluminium, iron, tin and zinc is

(1) Fe > Sn > Al > Zn (2) Sn > Fe > Al > Zn

(3) Al > Zn > Fe > Sn (4) Fe > Zn > Sn > Al

2. An electrochemical cell constructed for the reaction: $Cu^{2+}_{(aq)} + M_{(s)} \longrightarrow Cu_{(s)} + M^{2+}_{(aq)}$ has an E° = 0.75 V. The standard reduction potential for $Cu^{2+}_{(aq)}$ is 0.34 V. What is the standard reduction potential for $M^{2+}_{(aq)}$?

(1) 1.09 V (2) 0.410 V (3) –0.410 V (4) –1.09 V

3. $Ag(s) + NO_3^-(aq) + H^+(aq) \longrightarrow Ag^+(aq) + NO(g) + H_2O(\ell)$

When the oxidation reduction equation above is balanced, what is the coefficient for H^+ (aq) ?

(1) 1 (2) 2 (3) 3 (4) 4

4. Which equation represents an oxidation reduction reaction ?

(1) $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ (2) $H_2SO_4 + Ca(OH)_2 \longrightarrow CaSO_4 + 2H_2O$

(3) $MgCrO_4 + BaCl_2 \longrightarrow MgCl_2 + BaCrO_4$ (4) $Zn(NO_3)_2 + Na_2CO_3 \longrightarrow 2NaNO_3 + ZnCO_3$

5. In the chemical reaction

$3Br_2 + 6CO_3^{2-} + 3H_2O \rightarrow BrO_3^- + 5Br^- + 6HCO_3^-$

(1) bromine is reduced and water is oxidized (2) bromine is oxidized and carbonate is reduced

(3) bromine is both oxidized and reduced (4) bromine is neither oxidized nor reduced

6. What will be the volume of Cl_2 at STP produced during electrolysis of $MgCl_2$ which produces 6.5 g Mg (At. wt. of Mg = 24.3g, Cl = 35.5g)

(1) 5.099 litre (2) 5.99 litre (3) 12.02 litre (4) 3.099 litre

7. In the reaction

(i) $BaO_2(s) + O_3(g) \rightarrow BaO(s) + 2O_2(g)$ (ii) $ZnS(s) + 4O_3(g) \rightarrow ZnSO_4(s) + 4O_2(g)$

(1) O_3 is reduced both in (i) & (ii) (2) O_3 is oxidized both (i) & (ii)

(3) O_3 is oxidized in (i) & reduced in (ii) (4) O_3 is reduced in (i) & oxidized in (ii)

8. Which of the following is an oxidation - reduction reaction ?

(1) $H_3O^+(aq) + CO_3^{2-}(aq) \rightarrow HCO_3^-(aq) + H_2O$

(2) $HNO_3(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + NO_3^-(aq)$

(3) $Mg(s) + F_2(g) \rightarrow MgF_2(s)$

(4) $Pb(NO_3)_2(aq) + 2NaCl(aq) \longrightarrow PbCl_2(s) + 2NaNO_3(aq)$

8. What is the voltage for this cell when $[Cu^{2+}]$ = 1.0 M and $[Cr^{3+}]$ = 0.010 M ?

$2Cr(s) + 3Cu^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 3Cu(s)$ E° = 0.43 V

(1) 1.2 (2) 0.87 (3) 0.47 (4) 0.39

10. What is the ratio of number of electrons gained by acidified $KMnO_4$ and acidified $K_2Cr_2O_7$ in the reaction given below

$KMnO_4 \xrightarrow{H^+} Mn^{2+}$

$K_2Cr_2O_7 \xrightarrow{H^+} Cr^{3+}$

(1) 5 : 6 (2) 6 : 5 (3) 3 : 5 (4) 5 : 3

ANSWER KEY

Exercise-I

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	1	2	2	3	2	3	1	4	1	2	1	2	2	4	4	3	2	1	2	3
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	2	1	2	4	3	2	4	2	3	1	3	3	4	4	3	3	1	1	2	3
Que.	41	42	43	44	45															
Ans.	3	1	4	2	1															

Exercise-II

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	1	3	3	4	2	2	3	1	2	1	4	3	1	1	2	1	4	3	3	2
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	3	4	2	1	3	4	4	2	3	2	2	2	1	3	2	1	4	2	2	2
Que.	41	42	43	44	45	46	47	48	49	50										
Ans.	2	1	1	3	1	4	2	1	1	1										

Exercise-III

Que.	1	2	3	4	5	6	7	8	9	10
Ans.	3	3	4	1	3	2	3	4	3	1

CHEMICAL EQUILIBRIUM

- The most important characteristic property of a reversible reaction is that, it always attains a state of chemical equilibrium.
- Consider a general reversible reaction in a closed vessel

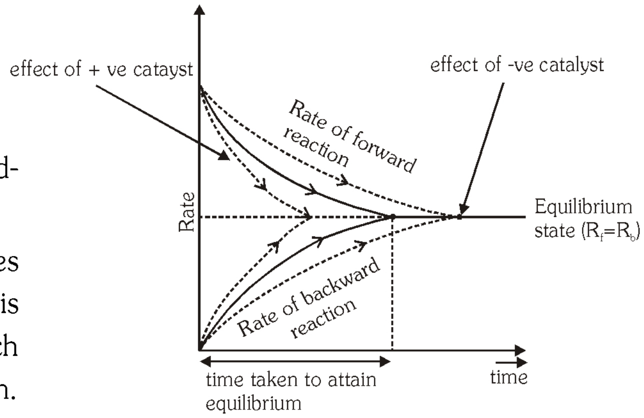
$A + B \underset{R_b}{\overset{R_f}{\rightleftharpoons}} C + D$ where R_f = forward rate of reaction

R_b = backward rate of reaction

- Initially rate of forward reaction is maximum and rate of backward reaction is zero, but as the concentration of products increases the rate of backward reaction also starts increasing and rate of forward reaction decreases due to decrease in concentration of reactants.
- At a certain stage, rate of forward reaction becomes equal to the rate of backward reaction called equilibrium state.

At equilibrium state

- Rate of forward reaction (R_f) = Rate of backward reaction (R_b)
- The concentrations of the reactants and products do not change with time.
- At this stage, number of moles of substances produced per second in the forward reaction is equal to the number of moles of substances which disappear per second in the backward reaction.



Characteristics of equilibrium

- Chemical equilibrium is dynamic in nature i.e. the reaction although appears to be stopped but actually takes place in both the directions with the same speed.
- Chemical equilibrium can be approached from both sides

$2HI \rightleftharpoons H_2 + I_2$ or $H_2 + I_2 \rightleftharpoons 2HI$

- At equilibrium, each reactant and product has a fixed concentration and this is independent of the fact whether the reaction starts from forward direction or backward direction with the reactant or with the product.
- Equilibrium is not affected by the presence of catalyst. The catalyst only helps in attaining equilibrium rapidly.
- The measurable properties of the system like temperature, concentration, colour, density etc. don't undergo any change with time at the chemical equilibrium conditions.
- Homogeneous equilibrium is the equilibrium in which the reactants and products are in the same phase.

$$CH_3COOC_2H_5(\ell) + H_2O(\ell) \overset{H^+}{\rightleftharpoons} CH_3COOH(\ell) + C_2H_5OH(\ell)$$

- Heterogeneous equilibrium is the equilibrium in which the reactants and products are in two or more phases.

$Zn(s) + CO_2(g) \rightleftharpoons ZnO(s) + CO(g)$

Active mass

- The term active mass means the concentration of the reactants expressed in moles per litre (molar concentration) or the pressure of the reacting gas in atmosphere. In case of gases and solutions, the molar concentration means the number of gram molecules present per litre.
- Active mass is usually expressed by enclosing the symbol of the reactant in square bracket [ ].

$$\text{Active mass} = \frac{\text{Number of gram moles of the substance}}{\text{Volume in litres}}$$

$$= \frac{\text{Weight of substance (in grams)}}{\text{Molecular weight }(M_w) \times \text{Volume (in litres)}}$$

$$= \frac{w}{M_w \times V(\text{litre})} = \frac{w \times 1000}{M_w \times V(\text{mL})}$$

- The active mass of solids and pure liquids is a constant quantity (unity) because it is an intensive property i.e. number of molecules present per unit volume do not change because density and molecular weight of solids and pure liquids are constant. But it does not apply for gaseous substances because for them number of molecules present per unit volume change with change in volume of vessel.

LAW OF MASS ACTION

- The law of mass action was given by **Guldberg and Waage (1864)**. It states that the rate of a chemical reaction is directly proportional to the product of active masses of the reacting substances.
- Consider the reaction

$$A + B \longrightarrow \text{Products}$$

According to Law of Mass Action,

Rate at which A reacts $\propto$ [A]

Rate at which B reacts $\propto$ [B]

Rate at which A and B react together $\propto$ [A] [B] = k[A] [B]

Where k is a constant of proportionality and is called **velocity constant**.

- Again, consider the reaction

$$2A + 3B \longrightarrow \text{Products}$$

It can be written as

$$A + A + B + B + B \longrightarrow \text{Products}$$

$\therefore$ Rate at which A reacts $\propto [A][A] = [A]^2$

Similarly, rate at which B reacts $\propto [B][B][B] = [B]^3$

$\therefore$ Rate of reaction between A and B $\propto [A]^2 [B]^3$

Hence for the most general reaction

$$aA + bB + cC + \ldots\ldots \longrightarrow \text{Products}$$

Rate of reaction $\propto [A]^a [B]^b [C]^c \ldots\ldots$

LAW OF CHEMICAL EQUILIBRIUM

- Consider a reversible homogeneous chemical reaction which has attained equilibrium state at a particular temperature : $A + B \rightleftharpoons C + D$
- Let the active masses of A, B, C and D be [A] [B] [C] and [D] respectively at equilibrium.

According to law of mass action

Rate of forward reaction $(R_f) \propto [A]\,[B]$

Rate of backward reaction $(R_b) \propto [C]\,[D]$

$$R_f = k_f [A]\,[B] \quad \text{and} \quad R_b = k_b [C]\,[D]$$

Where k_f and k_b are forward and backward rate or velocity constants respectively.

At equilibrium state :

$R_f = R_b$

$k_f [A] [B] = k_b [C] [D]$

$$\frac{k_f}{k_b} = \frac{[C][D]}{[A][B]}$$

$$K_C = \frac{[C][D]}{[A][B]}; \qquad \because \; K_C = \frac{k_f}{k_b}$$

- K_c is known as equilibrium constant. K_c has a definite value for every chemical reaction at particular temperature.

- For a general reaction $m_1A + m_2B + m_3C \rightleftharpoons n_1M + n_2N + n_3O$

$$K_C = \frac{[M]^{n_1}[N]^{n_2}[O]^{n_3}}{[A]^{m_1}[B]^{m_2}[C]^{m_3}}$$

- The equilibrium constant at a given temperature is the ratio of the rate constants of forward and backward reactions. $K = \frac{k_f}{k_b}$

Forms of K

Volume (K_C) | Partial pressure (K_P)

For reaction $A + B \rightleftharpoons C + D$

$$K_C = \frac{[C][D]}{[A][B]} \qquad K_p = \frac{P_C \times P_D}{P_A \times P_B}$$

Units **[ ]** = Moles/litre P = atmosphere

RELATION BETWEEN K_P AND K_C

- Consider a reversible homogeneous chemical equilibrium reaction

$m_1A + m_2B \rightleftharpoons n_1C + n_2D$

- According to law of mass action (L.M.A.)

$$K_C = \frac{[C]^{n_1}[D]^{n_2}}{[A]^{m_1}[B]^{m_2}} \quad \text{and} \quad K_P = \frac{(P_C)^{n_1}(P_D)^{n_2}}{(P_A)^{m_1}(P_B)^{m_2}}$$

$$\boxed{K_P = K_C (RT)^{\Delta n}}$$

$\Delta n = (n_1 + n_2) - (m_1 + m_2)$

= total number of moles of gaseous products – total number of moles of gaseous reactants

- The K_C is expressed by the units $(\text{mole litre}^{-1})^{\Delta n}$ and K_p by $(\text{atm})^{\Delta n}$.
- Three cases may arise:

(a) When $\Delta n = 0$ $K_p = K_C (RT)^0 = K_C$

For example: (i) $N_2 + O_2 \rightleftharpoons 2NO$ (ii) $H_2 + I_2 \rightleftharpoons 2HI$

- K_C and K_P are unitless in this case.

(b) **When Δn = +ve,** $K_P > K_C$

For example: (i) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ $\begin{pmatrix} K_C \to \text{mole litre}^{-1} \\ K_P \to \text{atm} \end{pmatrix}$

(ii) $2NH_3 \rightleftharpoons N_2 + 3H_2$ $\begin{pmatrix} K_C \to \text{mole}^2\,\text{litre}^{-2} \\ K_P \to \text{atm}^2 \end{pmatrix}$

(c) **When Δn = –ve,** $K_P < K_C$

For example: (i) $N_2 + 3H_2 \rightleftharpoons 2NH_3$ $\begin{pmatrix} K_C \to \text{mole}^{-2}\,\text{litre}^{2} \\ K_P \to \text{atm}^{-2} \end{pmatrix}$

(ii) $PCl_3 + Cl_2 \rightleftharpoons PCl_5$ $\begin{pmatrix} K_C \to \text{mole}^{-1}\text{litre}^{1} \\ K_P \to \text{atm}^{-1} \end{pmatrix}$

Special points

(i) If $T = \frac{1}{R}$ then $K_P = K_C\left(R \times \frac{1}{R}\right)^{\Delta n}$

or $K_P = K_C\,(1)^{\Delta n}$ or $K_P = K_C$ For any value of Δn

(ii) Possible values of K: $K = \frac{[\text{Product}]}{[\text{Reactant}]}$

K = 1 when [Product] = [Reactant]

K > 1 when [Product] > [Reactant]

K < 1 when [Product] < [Reactant]

Application of K

(a) Stability of reactants and products:

Stability of reactants increases when K decreases

Stability of products increases when K increases

Ex : In the following reaction which one oxide is more stable.

$2XO\,(g) \rightleftharpoons X_2(g) + O_2(g)$; $K_1 = 1 \times 10^{24}$

$2XO_2\,(g) \rightleftharpoons X_2(g) + 2O_2(g)$; $K_2 = 2.5 \times 10^{10}$

$\because$ $K_1 > K_2$ So the stability of $XO_2 > XO$

(b) Time taken to attain equilibrium increases when K decreases

FACTORS AFFECTING THE EQUILIBRIUM CONSTANT

(a) **The mode of representation of the reaction:** Consider the reversible chemical equilibrium reaction

$A + B \rightleftharpoons C + D$

The equilibrium constant for the reaction : $K_C = \frac{[C][D]}{[A][B]}$

If the reaction is reversed : $C + D \rightleftharpoons A + B$

The equilibrium constant for the reaction is : $K'_C = \frac{[A][B]}{[C][D]}$

The equilibrium constant K'_c is actually the reciprocal of K_C

Thus, the two equilibrium constants are related as $\to$ $\boxed{K'_C = \frac{1}{K_C}}$

(b) Stoichiometry of the reaction: When a reversible reaction can be written with the help of two or more stoichiometric equation then the value of equilibrium constant will be numerically different in these cases.

For example the dissociation of NO_2 can be represented as :

$2NO_2 \rightleftharpoons N_2 + 2O_2$ ....(1) $\qquad K_C = \dfrac{[N_2][O_2]^2}{[NO_2]^2}$

(i) If reaction (1) is divided by 2

Then the reaction becomes : $NO_2 \rightleftharpoons \frac{1}{2}N_2 + O_2 \qquad K_C' = \dfrac{[N_2]^{\frac{1}{2}}[O_2]}{[NO_2]}$

Thus, the two equilibrium constants are related as $K_C' = \sqrt{K_C}$

So if reaction is divided by n then $\boxed{K_C' = (K_C)^{\frac{1}{n}}}$

(ii) If reaction (1) is multiplied by 2

Then the reaction becomes: $4NO_2 \rightleftharpoons 2N_2 + 4O_2$

$K_C' = \dfrac{[N_2]^2[O_2]^4}{[NO_2]^4}$

Thus, the two equilibrium constants are related as $K_C' = (K_C)^2$

So if reaction is multiplied by n then $\boxed{K_C' = (K_C)^n}$

(c) If a reaction can be expressed as the sum of two or more reactions then overall K_C will be equal to the product of the individual equilibrium constant of the reactions.

Ex. $SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g) \rightarrow K_1$

$NO_2(g) \rightleftharpoons NO(g) + \frac{1}{2}O_2(g) \rightarrow K_2$

then $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g) \rightarrow K$

So $\boxed{K = K_1 \times K_2}$

(d) Temperature: The value of equilibrium constant changes with the change of temperature. If K_1 and K_2 be the equilibrium constants of a reaction at absolute temperatures T_1 and T_2 and ΔH is the change in enthalpy then:

$\log K_2 - \log K_1 = \dfrac{\Delta H}{2.303R}\left[\dfrac{1}{T_1} - \dfrac{1}{T_2}\right]$ (According to vant hoff equation)

If the temperature T_2 is higher than T_1 then, $\dfrac{1}{T_1} - \dfrac{1}{T_2} > 0$ So $\log K_2 - \log K_1 = \dfrac{\Delta H}{2.303R}$

(i) When ΔH = +ve (endothermic reaction)

$\log K_2 - \log K_1 > 0$ or $\log K_2 > \log K_1$ $\qquad K_2 > K_1$

The value of equilibrium constant increases when temperature increases in case of endothermic reactions.

(ii) When ΔH = –ve (exothermic reaction)

$\log K_2 - \log K_1 < 0$

$\log K_2 < \log K_1$

$K_2 < K_1$

The value of equilibrium constant decreases when temperature increases in the case of exothermic reactions.

Special points

The value of equilibrium constant is independent of the following factors :

(a) Concentration of reactants and products.

(b) Pressure

(c) Volume

(d) The presence of a catalyst.

(e) Presence of inert materials.

Degree of dissociation

$$\alpha = \frac{x}{a} \times 100$$

Where α= Degree of dissociation in percentage

x= Number of dissociated moles

a= Given initial number of moles

At equilibrium, the degree of dissociation is directly proportional to the pressure.

LE-CHATELIER'S PRINCIPLE

- According to this principle, if a system at equilibrium is subjected to a change of concentration, pressure or temperature then the equilibrium is shifted in such a way as to nullify the effect of change.

(a) Change in concentration

- In an equilibrium increasing the concentrations of reactants results in shifting the equilibrium in favour of products while increasing concentrations of the products results in shifting the equilibrium in favour of the reactants.

(b) Change in pressure

- When the pressure on the system increases, the volume decreases proportionately i.e. the total number of moles present per unit volume increases. According to Le-Chatelier's principle, the equilibrium shifts in that direction in which there is decrease in number of moles.
- If there is no change in number of moles of gases in a reaction then a pressure change does not affect the equilibrium.

(c) Change in temperature

If the temperature of the system at equilibrium is increased then reaction will proceed in that direction in which heat can be used. Thus increase in temperature will favour the forward reaction for endothermic reaction.

- Similarly, increase in temperature will favour the backward reaction for exothermic reactions.

Special point

- Le-Chatelier's principle is applicable for both chemical and physical equilibrium.

PHYSICAL EQUILIBRIUM

Physical reaction

- Those reactions in which change in only physical properties of substance takes place without any chemical change is called physical reaction.

Examples

(a) Ice-water system (melting of ice)

Melting of ice is accompained by absorption of heat (endothermic) and decrease in volume

$$\underset{\substack{(H_2O)\\ \text{1gram (1.09 mL)}}}{\text{Ice (s)}} \rightleftharpoons \underset{\substack{(H_2O)\\ \text{1gram (1.01 mL)}}}{\text{water } (\ell)}$$

Hence both increase of temperature and pressure will favour the melting of ice into water.

(b) Water-water vapour system (Vaporisation of water)

Vaporisation of water is an endothermic reaction and condensation of vapour into water is an exothermic reaction

$$\underset{(H_2O)}{\text{Water } (\ell)} \rightleftharpoons \underset{(H_2O)}{\text{Vapour (g)}}$$

- The equilibrium shifts towards right side when the temperature is increased so rise in temperature will increase the vapour.
- The equilibrium shifts towards left side when the pressure is increased (i.e. volume is decreased) so increase in pressure will favour the rate of condensation of vapour into water.
- Thus favourable conditions for conversion of water into vapour are high tempepratute and low pressure.

(c) Solubility of gases

Gas (g) + Water (ℓ) $\rightleftharpoons$ Aqueous solution (ℓ)
(Solute) (Solvent) (Solution)

• Effect of pressure $\rightarrow$ Solubility of gases in a solvent increases with a decrease in volume, and with increasing pressure.

REACTION QUOTIENT (Q)

- Consider a general homogeneous reversible reaction :

$$A + B \rightleftharpoons C + D$$

- Reaction Quotient (Q) $= \frac{[C][D]}{[A][B]}$, (Applied in any condition)
- Equilibrium constant K = $\frac{[C][D]}{[A][B]}$, (Applied only in equilibrium state)

(i) When Q = K then reaction is in equilibrium state.
(ii) When Q < K then rate of forward reaction increases.
(iii) When Q > K then rate of backward reaction increases.

CHEMICAL EQUILIBRIUM

EXERCISE-I

EQUILIBRIUM

1. True for following general reversible equilibrium : $A + B \rightleftharpoons AB$

(1) Both reaction is endothermic

(2) Both reaction is exothermic

(3) Association reaction is endothermic and dissociation is exothermic

(4) Association reaction is exothermic and dissociation is endothermic.

2. A chemical reaction Catalysed by catalyst X. In this reaction X.

(1) does not effected equilibrium constant. (2) decrease rate constant of reaction

(3) decrease enthalpy of reaction (4) increase Activation energy

3. $x \rightleftharpoons y$ reaction is said to be in equilibrium, when :-

(1) Only 10% conversion x to y is just

(2) Complete conversion of x to y has taken place

(3) Conversion of x to y is only 50% complete

(4) The rate of change of x to y is just equal to the rate of change of y to x in the system

4. A chemical reaction is in equilibrium when

(1) Formation of product is minimum

(2) Reactions are completely transformed into products

(3) Rates of forward and backward reactions are equal

(4) Equal amounts of reactants and products are present

5. Which of the following is a characteristic of a reversible reaction?

(1) It never proceeds to completion

(2) It can be completed in finite time

(3) It proceeds only in the forward direction

(4) Number of moles of reactants and products are equal

LAWS OF MASS ACTION AND EQUILIBRIUM CONSTANTS

6. Equilibrium constant for the reaction : $2NO(g) + Cl_2(g) \rightleftharpoons 2NOCl(g)$,is correctly given by expression

(1) $K = \dfrac{[2NOCl]}{[2NO][Cl_2]}$ (2) $K = \dfrac{[NOCl]^2}{[NO]^2[Cl_2]}$ (3) $K = \dfrac{[NO]^2[Cl_2]^2}{[NO]^2 + [Cl_2]}$ (4) $K = \dfrac{[NO]^2 + [Cl]^2}{[NOCl]}$

7. Given - $SO_2 + 1/2\ O_2 \rightleftharpoons SO_3$; K_1, $2SO_2 + O_2 \rightleftharpoons 2SO_3$; K_2. Where K_1 and K_2 are the equilibrium constant. The expression showing correct relationship between K_1 and K_2 is:

(1) $K_2 = K_1$ (2) $K_2 = K_1^2$ (3) $K_2 = \dfrac{1}{k_1}$ (4) $K_2 = \dfrac{1}{k_1^2}$

8. For following reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$ value of K_C depends on

(1) Initial concentration of reactant (2) Pressure

(3) Temperature (4) All of these

9. Which of the following statement true for endothermic reaction $N_2 + O_2 \rightleftharpoons 2NO$

(1) K does not depend on temperature (2) K increases with temperature

(3) K changes with concentration (4) K decreases with pressure decreases

10. The following equilibrium are given

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ -------------- K_1

$N_2 + O_2 \rightleftharpoons 2NO$ -------------- K_2

$H_2 + \frac{1}{2}O_2 \rightleftharpoons H_2O$ -------------- K_3

The equilibrium constant of the reaction $2NH_3 + \frac{5}{2}O_2 \rightleftharpoons 2NO + 3H_2O$, in terms of K_1, K_2 and K_3 is :

(1) $\frac{K_1K_2}{K_3}$ (2) $\frac{K_1K_3^2}{K_2}$ (3) $\frac{K_2K_3^3}{K_1}$ (4) $K_1K_2K_3$

11. $N_2 + O_2 \rightleftharpoons 2NO - 33.2$ kcal equilibrium of reaction does not effected by :-

(1) Temperature (2) Concentration (3) Pressure (4) None of these

12. For reaction $HI \rightleftharpoons \frac{1}{2} H_2 + \frac{1}{2} I_2$ value of K_c is 1/8 then value of K_c for $H_2 + I_2 \rightleftharpoons 2HI$.

(1) $\frac{1}{64}$ (2) 64 (3) $\frac{1}{8}$ (4) 8

13. For the reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3$, equilibrium constant is K, then for $2N_2 + 6H_2 \rightleftharpoons 4NH_3$ will be :-

(1) $\sqrt{K}$ (2) K^2 (3) K/2 (4) $\sqrt{K} + 1$

14. For the reaction $N_2 + O_2 \rightleftharpoons 2NO$, K_C is 100 then K_C for reaction. $2NO \rightleftharpoons N_2 + O_2$ will be

(1) 0.01 (2) 0.1 (3) 10 (4) 100

15. K for the synthesis of HI is 50. What is K for its dissociation

(1) 50 (2) 5 (3) 0.2 (4) 0.02

16. If K_C is 41 for, $N_2 + 3H_2 \rightleftharpoons 2NH_3$ then for $NH_3 \rightleftharpoons \frac{1}{2} N_2 + \frac{3}{2} H_2$ K_C will be

(1) 41 (2) $\sqrt{41}$ (3) 20.5 (4) $\frac{1}{\sqrt{41}}$

17. The equilibrium constant for the reaction : $N_2 + 3H_2 \rightleftharpoons 2NH_3$ is K. The equilibrium constant for $\frac{1}{2}N_2 + \frac{3}{2}H_2 \rightleftharpoons NH_3$ will be

(1) $\frac{K}{2}$ (2) 2K (3) $\sqrt{K}$ (4) K^2

18. The equilibrium constant for the gaseous reaction $N_2 + O_2 \rightarrow 2NO$ is K. The equilibrium constant for the formation of one mole of NO will be

(1) K/2 (2) K (3) $\sqrt{K}$ (4) 2K

19. For the reaction

$$PCl_3 (g) + Cl_2 (g) \rightleftharpoons PCl_5 (g)$$

The value of K_C at 250°C is 26. The value of K_p at this temperature will be :

(1) 0.61 (2) 0.57 (3) 0.83 (4) 0.46

20. The ratio in between the molar concentration of products to that of the reactants to the power their respective stoichiometries is called

(1) equilibrium coefficient (2) equilibrium constant

(3) dissociation constant (4) none of the above

21. For the equation : $pP + qQ \rightarrow rR + sS$

The value of the equilibrium constant, K is Written as

(1) $K = [P]^p[Q]^q / [R]^r + [S]^s$ (2) $K = [R]^r \times [S]^s / [P]^P$

(3) $K = [R]^r [S]^s / [P]^P [Q]^q$ (4) None of the above

22. In a reversible reaction : $A \underset{K_2}{\overset{K_1}{\rightleftharpoons}} B$, if the initial concentrations of A and B are a and b (in mole per litre) and the equilibrium concentrations are (a–x) and (b+x) respectively; express 'x' in terms of K_1, K_2, a and b.

(1) $\frac{K_1a - K_2b}{K_1 + K_2}$ (2) $\frac{K_1a - K_2b}{K_1 - K_2}$ (3) $\frac{K_1a - K_2b}{K_1K_2}$ (4) $\frac{K_1a + K_2b}{K_1K_2}$

23. In reaction A + B $\rightleftharpoons$ AB,concentration of A is doubled then rate of reaction is-

(1) Doubled (2) Halved (3) Unchanged (4) Four time

24. Rate of reaction for

$CCl_3CHO + NO \rightleftharpoons CHCl_3 + NO + CO$ reaction is $K[CCl_3CHO][NO]$ then the unit of K is-

(1) litre^2 mole^{-2} second^{-1} (2) mole litre^{-1} second^{-1}

(3) litre mole^{-1} second^{-1} (4) second^{-1}

25. 3.1 moles of $FeCl_3$ and 3.2 moles of NH_4SCN are added to one litre of water. At equilibrium 3.0 moles of $FeSCN^{2+}$ are formed. The equilibrium constant K_C of the reacion $Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$ will be :

(1) 6.66×10^{-3} (2) 0.30 (3) 3.30 (4) 150

26. Law of mass action is proposed by-

(1) Guldberg and waage (2) Bondstein (3) Berthlott (4) Grahm

27. In a reversible equilibrium reaction rate of both side reaction is –

(1) Equal (2) Different (3) One side more (4) Not definite

28. In reversible reactions, at equilibrium

(1) the rate of forward reaction is faster than rate of backward reaction.

(2) the rate of forward reaction is slower than rate of backward reaction.

(3) the rate of forward and the backward reaction is equal.

(4) none of the above.

29. Which statements are correct for a reaction at equilibrium ?

I. The forward and reverse reactions both continue.

II. The rates of the forward and reverse reactions are equal.

III.The concentration of reactants and products are equal

(1) I and II only (2) I and III only (3) II and III only (4) I, II and III

30. 64 g of HI are present in a 2 litre vessel. The active mass of HI is

(1) 0.5 (2) 0.25 (3) 1.0 (4) None of these

APPLICATIONS OF EQUILIBRIUM CONSTANTS

31. For a reaction : $2A(g) + B(g) \rightleftharpoons C(g) + D(g)$, the value of K_p will be :

(1) $K_p = \frac{n_Cn_D}{n_A^2n_B}\frac{V}{RT^2}$ (2) $K_p = \frac{n_Cn_D}{n_A^2n_B}\frac{P}{RT}$ (3) $K_p = \frac{n_Cn_D}{n_A^2n_B}\frac{RT^2}{V}$ (4) $K_p = \frac{n_Cn_D}{n_A^2n_B}\frac{V}{RT}$

32. For reaction $2X(g) + Y(g) \rightleftharpoons 2Z(g)$ value of equilibrium constant is 2.25, if value of X and Z concentration are 2 moles/litre and 3 moles/litre respectively then concentration of Y at equilibrium is

(1) 1.0 mole/litre (2) 2.25 moles/litre (3) 2.0 moles/litre (4) 5.0 moles/litre

33. If 4 moles of A and 4 moles of B are mixed at equilibrium and produced 2 moles of C, according following reaction A + B $\rightleftharpoons$ C + D. Then value of equilibrium constant is-

(1) 4 (2) 1 (3) $\frac{1}{2}$ (4) $\frac{1}{4}$

34. In which of the following process reaction is fastest complete.

(1) $K = 10$ (2) $K = 1$ (3) $K = 10^3$ (4) $K = 10^{-2}$

35. The value of equilibrium constant for reaction $H_2 + I_2 \rightleftharpoons 2HI$ in which equilibrium concentration of HI, H_2 and I_2 is 0.7, 0.1, and 0.1 moles/litre is :

(1) 0.36 (2) 36 (3) 49 (4) 0.4

36. Dissociation of 2gm moles of PCl_5 produces 0.4 moles of Cl_2, if volume of container is 1 litre then equilibrium constant is :-

(1) 1×10^{-1} (2) 1×10^{-2} (3) 2×10^{-1} (4) 1×10^{-6}

37. At constant temperature one mole of SO_3 is taken in one litre flask as per following equilibrium $2SO_3 \rightleftharpoons 2SO_2 + O_2$; If 0.6 moles of SO_2 produce at equilibrium then value of equilibrium constant is :-

(1) 0.36 (2) 0.45 (3) 0.54 (4) 0.675

38. One mole of PCl_5 are present in a one litre Container and at equilibrium 0.4 moles of Cl_2 is produce. Then the value of K_C is-

(1) 2.6×10^{-1} (2) 2.6×10^{-2} (3) 2.6 (4) None of these

39. In a one litre Container equilibrium mixture of following reaction $2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$ is moles of H_2S, H_2 and S_2 is 0.5, 0.1 and 0.4 respectively. Then equilibrium constant of above reaction is :-

(1) 0.004 moles/litre (2) 0.080 moles/litre

(3) 0.016 moles/litre (4) 0.08 moles/litre

40. For the equilibrium process $x + y \rightleftharpoons xy$. If the concentration of x and y is doubled, then equilibrium constant.

(1) Become twice (2) Become half (3) Unchanged (4) Become thrice

41. Under equilibrium condition if the concentration of reactant and product is double, K_C would become remain :–

(1) Half (2) Twice (3) One fourth (4) Unchanged

42. In a reaction, equilibrium proceeds towards reactants then K will be :-

(1) $K >> 1$ (2) $K << 1$ (3) $K = 0$ (4) $K = 1$

43. For a reaction at equilibrium rate constant of forward and backward reaction are 2.5×10^{-4} and 7.5×10^{-4} respectively. Then equilibrium constant is :-

(1) 3 (2) 1/3 (3) 1 (4) None

44. In the reaction $A + 2B \rightleftharpoons C + 3D$ scientist found that 5 g C formed by complete reaction of 10 g A and 20 g of B then quantity of D is :-

(1) 2.5 g (2) 25 g

(3) 30 g (4) Knowledge incomplete

45. Which of the following is favours of forward reactions :–

(1) $Q = K_c$ (2) $Q > K_c$

(3) $Q < K_c$ (4) None

FACTORS AFFECTING EQUILIBRIA

46. A reversible reaction in equilibrium, if changed one factor effect on reaction is -
(1) Rate of forward and backward reaction increase.
(2) Rate of forward and backward reaction decrease.
(3) Rate of that reaction increases or decreases which partially offsets the change while reaching a new state of equilibrium.
(4) No difference occur.

47. By Le chatelier's principle we get following information
(1) Equilibrium may be shifted by change in concentration of the reactants.
(2) Dissociation constant of weak acids.
(3) Energy change in reaction.
(4) Equilibrium constant of reaction.

48. Reaction PCl_5 (g) $\rightleftharpoons$ PCl_3(g) + Cl_2(g) is equilibrium in closed container. If PCl_5 is added at fixed temperature then effect on equilibrium concentration of Cl_2(g) is :-
(1) Decrease (2) Increase
(3) Unchanged (4) Nothing say without value of K_p

49. $N_2 + 3H_2 \rightleftharpoons 2NH_3$ $\Delta H = -ve$
If temperature of above equilibrium reaction increases then equilibrium will be-
(1) Shift Right side (2) Shift left side (3) Unchanged (4) Nothing say.

50. According to Le-chatelier's principle which of following factor affect chemical system in equilibrium
(1) Only concentration (2) Only pressure
(3) Only temperature (4) Concentration, pressure or temperature

51. Which of the following equilibrium will shift towards right side due to increase in temperature
(1) CO(g) + H_2O(g) $\rightleftharpoons$ CO_2(g) + H_2(g) (2) $2SO_2$(g) + O_2(g) $\rightleftharpoons$ $2SO_3$ (g)
(3) H_2O (g) $\rightleftharpoons$ H_2(g) + $\frac{1}{2}$ O_2(g) (4) 4HCl (g) + O_2(g) $\rightleftharpoons$ $2H_2O$(g) + $2Cl_2$(g)

52. In which reaction equilibrium moves in left hand side when pressure is increased :-
(1) $H_2 + Cl_2 \rightleftharpoons 2HCl$ (2) 2Mg (s) + O_2(g) $\rightleftharpoons$ 2MgO(s)
(3) $2H_2O \rightleftharpoons 2H_2 + O_2$ (4) $N_2 + 3H_2 \rightleftharpoons 2NH_3$

53. The Beans are cooked earlier in pressure cooker, because :
(1) Boiling point increase with increasing pressure.
(2) Boiling point decrease with increasing pressure.
(3) Extra pressure of pressure cooker, softens the beans.
(4) Internal energy is not lost while cooking in pressure cooker.

54. In any system, after adding catalyst :-
(1) Equilibrium concentration is increase.
(2) Does not affect equilibrium concentration.
(3) Decrease equilibrium concentration.
(4) Rate of forward reaction increase and backward reaction is decrease.

55. In a reversible reaction, a catalyst will affect :-
(1) The rate of reverse reaction
(2) The rate of forward reaction
(3) Neither the rate of forward reaction nor the rate of reverse reaction
(4) The rate of forward reaction and reverse reaction equally

56. For which equilibrium reaction low pressure is favourable condition-

(1) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$
(2) $N_2 + 3H_2 \rightleftharpoons 2NH_3$
(3) $2SO_2 + O_2 \rightleftharpoons 2SO_3$
(4) $N_2 + O_2 \rightleftharpoons 2NO$

57. If pressure is applied on the given chemical equilibrium, then equilibrium will shift towards $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$:-

(1) Forward (2) Backward (3) Slow down (4) Will not shift

58. For the equilibrium

$N_2+3H_2 \rightleftharpoons 2NH_3 + 22.4$ kcal, favourable condition for proceeding in forward direction

(1) Increasing pressure
(2) Decreasing pressure
(3) Increasing temperature
(4) Removal of H_2

59. In the manufacture of NH_3 by Haber's process which condition give maximum yield–

(1) High temperature, High pressure and high concentration of reactants
(2) High temperature, low pressure and low concentration of reactants
(3) Low temperature, high pressure and high concentration of reactants
(4) Low temperature, low pressure and low concentration of reactants

60. In the reaction : $CH_4 + H_2O \rightleftharpoons CO + 3H_2$ The increase of pressure will

(1) favour the forward reaction
(2) favour the reverse reaction
(3) stop the reaction
(4) will have no effect.

61. Which of the following reaction is not affected by pressure.

(1) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$
(2) $H_2 + I_2 \rightleftharpoons 2HI$
(3) $N_2O_4 \rightleftharpoons 2NO_2$
(4) $N_2 + 3H_2 \rightleftharpoons 2NH_3$

62. If an inert gas is added to the equilibrium mixture of the dissociation of PCl_5 in a closed vessel,

(1) the concentration of Cl_2 will increase
(2) the concentration of PCl_3 will increase
(3) the concentration of PCl_5 will increase
(4) the equilibrium concentrations will remain unaffected.

63. If pressure is increased on the equilibrium $N_2 + O_2 \rightleftharpoons 2NO$, the equilibrium will :

(1) shift in the forward direction
(2) shift in the backward direction
(3) remain undisturbed
(4) may shift in the forward or backward direction

64. In $2HI \rightleftharpoons H_2 + I_2$ the forward reaction is affected by change in :

(1) catalyst (2) pressure (3) volume (4) temperature

65. For the reaction,

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g),$$

$$\Delta H = -57.2\text{kJ mol}^{-1} \text{ and } Kc = 1.7 \times 10^{16}$$

Which of the following statement is **INCORRECT**?

(1) The equilibrium constant is large suggestive of reaction going to completion and so no catalyst is required.
(2) The equilibrium will shift in forward direction as the pressure increase.
(3) The equilibrium constant decreases as the temperature increases.
(4) The addition of inert gas at constant volume will not affect the equilibrium constant.

EXERCISE-II

EQUILIBRIUM

1. The reaction : $3Fe(s)+4H_2O(g) \rightleftharpoons Fe_3O_4(s)+4H_2(g)$, is reversible if it is carried out

(1) at constant pressure (2) at constant temperature

(3) in an open vessel (4) in a closed vessel.

2. Which of the following statement is false in case of equilibrium state -

(1) There is no apparent change in measurable properties with time.

(2) It is dynamic in nature.

(3) It can be attained from either side of the reaction.

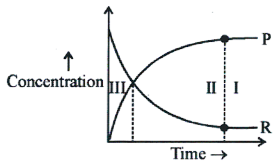
(4) It can be attained from the side of reactants only.

3. The state of chemical equilibrium of a reversible reaction is not influenced by

(1) pressure (2) catalyst

(3) concentration of the reactants (4) temperature.

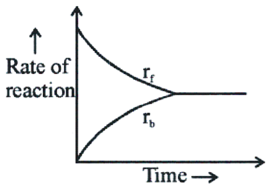
4. For the reaction: $R \rightleftharpoons P$, Variation of concentration is plotted against time:

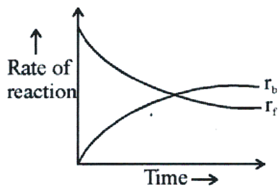


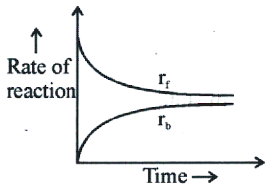
Which of the following regions show(s) state of equilibrium?

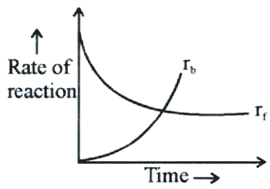
(1) III (2) II (3) I (4) both II and III

5. Which of the following represents the rate of forward reaction (r_f) and rate of backward reaction (r_b) at equilibrium?

(1) 

(2) 

(3) 

(4) 

LAWS OF MASS ACTION AND EQUILIBRIUM CONSTANTS

6. In gaseous reaction

$2NO \rightleftharpoons N_2 + O_2 \ \Delta H = -43.5$ kcal.

Which of the following statement true for reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

(1) Value of K does not depend on T (2) Value of T increases with decreases K

(3) Value of K decrease with decreases T (4) K change with number

7. The value of K_c at 721 K temperature for following reaction is 50. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ if equilibrium concentration of H_2 and I_2 is 0.5M then value of K_p is.

(1) 0.02 (2) 0.2 (3) 50.0 (4) 50/RT

8. K_P for the following reaction at 700 K is 1.3×10^{-3} atm^{-1}. The K_C at same temperature of the reaction $2SO_2 + O_2 \rightleftharpoons 2SO_3$ will be :-

(1) 1.1×10^{-2} (2) 3.1×10^{-2} (3) 5.2×10^{-2} (4) 7.4×10^{-2}

9. For reaction $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$, K_C at 427°C is 3×10^{-6} L^{-1} mol. The value of K_P is nearly :-

(1) 7.50×10^{-5} (2) 2.50×10^{-5} (3) 2.50×10^{-4} (4) 1.75×10^{-4}

10. For the reaction :

$CH_4(g) + 2O_2(g) \rightleftharpoons CO_2(g) + 2H_2O(l)$ $\Delta_r H = -170.8 \text{ kJ mol}^{-1}$

Which of the following statements is not true:-

(1) At equilibrium, the concentrations of $CO_2(g)$ and $H_2O(l)$ are not equal

(2) The equilibrium constant for the reaction is given by $K_p = \frac{[CO_2]}{[CH_4][O_2]}$

(3) Addition of $CH_4(g)$ or $O_2(g)$ at equilibrium will cause a shift to the right

(4) The reaction is exothermic

11. Equilibrium constant for following gaseous reaction is :– $H_2 + I_2 \rightleftharpoons 2HI$

(1) $K_p > K_c$ (2) $K_p < K_c$ (3) $K_p = K_c$ (4) $K_p = (1)/K_p$

12. For reaction $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ equilibrium constant (K_p) is changed

(1) By pressure (2) By catalyst (3) By temperature (4) Quantity of H_2 and I_2

13. Which of the following expression is correct for the process : $N_2 + 3H_2 \rightleftharpoons 2NH_3 + Q$ kcal

(1) $K_p = K_c$ (2) $K_p = K_c(RT)^{-1}$ (3) $K_p = K_c(RT)^{-2}$ (4) $K_p = K_C(RT)^2$

14. For the reversible reaction

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ at 500^0C, the value of K_P is 1.44×10^{-5} when partial pressure is measured in atmospheres. The corresponding value of K_C, with concentration in mole litre^{-1} is

(1) $1.44 \times 10^{-5} / (0.082 \times 500)^{-2}$ (2) $1.44 \times 10^{-5} / (8.314 \times 773)^{-2}$

(3) $1.44 \times 10^{-5} / (0.082 \times 773)^{2}$ (4) $1.44 \times 10^{-5} / (0.082 \times 773)^{-2}$

15. What is the relation between K_p and K_c

(1) $K_c = K_p(RT)^{\Delta n}$ (2) $K_p = K_c(RT)^{\Delta n}$ (3) $K_c = K_p(RT)^{-\Delta n}$ (4) 2 & 3 both

16. For reaction, $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$ which of the following is correct:-

(1) $K_p = K_c^2$ (2) $K_p = K_c$ (3) $K_p = \sqrt{K_c}$ (4) $K_p = \frac{1}{K_c}$

17. $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$

According to above reaction, partial pressure of CO_2 and CO are 4 and 8 respectively then find out K_p of the above reaction :-

(1) 6 (2) 2 (3) 16 (4) 32

18. For which of the following reaction $K_P = K_C$:–

(1) $H_2 + I_2 \rightleftharpoons 2HI$ (2) $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ (3) $2NH_3 \rightleftharpoons 3H_2 + N_2$ (4) $2SO_2 + O_2 \rightleftharpoons 2SO_3$

19. In the reaction : $N_2 + 3H_2 \rightarrow 2NH_3$ if the concentration of nitrogen and hydrogen is doubled, the equilibrium constant will

(1) also be doubled (2) be halved (3) remain constant (4) become four times

20. Which of following reaction is irreversible if container is closed :-

(1) $H_2 + I_2 \rightleftharpoons 2HI$ (2) $AgNO_3 + NaCl \rightleftharpoons AgCl(\downarrow) + NaNO_3$

(3) $CaCO_3 \rightleftharpoons CaO + CO_2$ (4) $O_2 + 2SO_2 \rightleftharpoons 2SO_3$

21 The yield of product can be increased by :-

(1) Decreasing the concentration of reactant (2) Increasing the concentration of product

(3) By escaping out the gaseous product. (4) Carying the reaction in a closed container/vessel

22. Increase in temperature in a reversible equilibrium reaction favours –

(1) Forward reaction only (2) Backward reaction only

(3) Either forward or backward reaction (4) Neither forward nor backward reaction

23. For the reversible reaction $A + 2B \rightleftharpoons C + 3D$ at equilibrium. The partial pressure of A, B, C and D are 0.20, 0.10, 0.30 and 0.70 atmosphere respectively. Then evaluate K_P

(1) 11.25 (2) 51.45 (3) 0.05 (4) 3.75

24. Ratio of active masses of 22 g CO_2,3 gH_2 and 7 gN_2 in a gaseous mixture :-

(1) 22 : 3 : 7 (2) 0.5 : 3 : 7 (3) 1 : 3 :1 (4) l : 3 : 0.5

25. For which reaction at 298 K, the value of $\frac{K_p}{K_c}$ is maximum and minimum respecitvely:–

(a) $N_2O_4 \rightleftharpoons 2NO_2$ (b) $2SO_2 + O_2 \rightleftharpoons 2SO_3$ (c) $X + Y \rightleftharpoons 4Z$ (d) $A + 3B \rightleftharpoons 7C$

(1) d, c (2) d, b (3) c, b (4) d, a

APPLICATIONS OF EQUILIBRIUM CONSTANTS

26. The rate of forward reaction is two times that of reverse reaction at a given temperature and identical concentration, K_{eq} is :

(1) 1.5 (2) 2.0 (3) 0.5 (4) 2.5

27. If in an evacuated vessel of volume 1 dm^3 initially 8 moles of $AB_3(g)$ are taken and at equilibrium 2 moles of $A_2(g)$ are present then find out K_C :

$2AB_3(g) \rightleftharpoons A_2(g) + 3B_2(g)$

(1) 72 mol^2/L^2 (2) 36 mol^2/L^2 (3) 27mol^2/L^2 (4) 3 mol^2/L^2

28. For the reaction, $A+B \rightleftharpoons C + D$, $K_c = 9$. If A and B are taken in equal amounts, then amount of C at equilibrium is

(1) 1 (2) 0.25 (3) 0.75 (4) None of these

29. In a reaction $PCl_5 \rightleftharpoons PCl_3 + Cl_2$ degree of dissociation is 30%. If initial moles of PCl_5 is one then total moles at equilibrium is :

(1) 1.3 (2) 0.7 (3) 1.6 (4) 1.0

30. If x moles of N_2O_4 dissociate then the total number of moles at equilibrium will be :-

$N_2O_4 \rightleftharpoons 2NO_2$

(1) 1 (2) 2 (3) 1 + x (4) $(1- x)^2$

31. 4.5 moles each of hydrogen and iodine heated in a sealed ten litre vessel. At equilibrium, 3 moles of HI were found. The equilibrium constant for $H_2(g)+I_2(g) \rightleftharpoons 2HI(g)$ is :–

(1) 1 (2) 10 (3) 5 (4) 0.33

32. 40% of a mixture of 0.2 mole of $N_2(g)$ and 0.6 mole of $H_2(g)$ react to give $NH_3(g)$ according to the equation : $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$, at constant temperature and pressure. Then the ratio of the final volume to the initial volumes of gases are as :

(1) 4 : 5 (2) 5 : 4 (3) 7 : 10 (4) 8 : 5

33. The dissociation of CO_2 can be expressed as $2CO_2(g) \rightleftharpoons 2CO(g)+O_2(g)$. If 2 moles of CO_2 are taken initially and 40% of the CO_2 is dissociated. What is the total number of moles at equilibrium

(1)2.4 (2) 2.0 (3) 1.2 (4) 5

34. At 250°C for 50% dissociation of PCl_5 required pressure is-

(1) Equal to K_P (2) Three times of K_P (3) Double of K_P (4) Four times of K_P

35. One mole of N_2O_4 in a 1 litre flask decomposes to attain the equilibrium $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ At the equilibrium the mole fraction of NO_2 is 1/2. Hence K_C will be :

(1) 1/3 (2) 1/2 (3) 2/3 (4) 1

FACTORS AFFECTING EQUILIBRIA

36. In a given reaction : $2X(g) + Y(g) \rightleftharpoons 2Z(g) + 80$ kcal
Which combination of pressure and temperature gives the highest yield of Z at equilibrium
(1) 1000 atm and 200°C (2) 500 atm and 500°C
(3) 1000 atm and 100°C (4) 500 atm and 100°C

37. Consider the reaction equilibrium : $2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\,SO_{3(g)}$; $\Delta H^o = -198$ kJ
On the basis of Le-Chatelier's principle, the condition favourable for the forward reaction is-
(1) Lowering the temperature and increasing the pressure
(2) Any value of temperature as well as pressure
(3) Lowering of temperature as well as pressure
(4) Increasing temperature as well as pressure

38. In which of the following equilibrium reactions, the equilibrium would shift to right side, if total pressure is decreased :–
(1) $N_2(g) + 3H_2(g) \rightleftharpoons 2\,NH_3(g)$ (2) $H_2(g) + I_2(g) \rightleftharpoons 2\,HI(g)$
(3) $N_2O_4(g) \rightleftharpoons 2\,NO_2(g)$ (4) $N_2(g) + O_2(g) \rightleftharpoons 2\,NO_2(g)$

39. Reaction $2BaO_2(s) \rightleftharpoons 2BaO(s) + O_2(g)$; $\Delta H = +$ ve. At equilibrium condition, Pressure of O_2 is depends on :–
(1) Increase mass of BaO_2 (2) Increase mass of BaO
(3) Increase temperature at equilibrium (4) Increase mass of BaO_2 and BaO both

40. In the following reaction $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ at constant temperature rate of backward reaction is increase by
(1) Inert gas mixed at constant volume (2) Cl_2 gas mixed at constant volume
(3) Inert gas mixed at constant pressure (4) PCl_5 mixed at constant volume.

41. In equilibrium reaction
$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ $\Delta H = -3000$ calories, which factor favours dissociation of HI :-
(1) Low temperature (2) High Pressure (3) High temperature (4) Low pressure.

42. In a gaseous reaction $A + 2B \rightleftharpoons C +$ heat : forward process is favoured by -
(1) Low temperature and low pressure. (2) low pressure.
(3) High pressure and low temperature (4) High pressure and high temperature

43. At given temperature for reaction
$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$
concentration of $CO_2(g)$ is increased by
(1) catalyst (2) inert gas
(3) decrease volume of container (4) increase quantity of CO(g)

44. In following reaction $N_2 + 3H_2 \rightleftharpoons 2NH_3 + Q$ kcal. Favourable condition for maximum production of NH_3 is:
(1) High temperature & high pressure. (2) High temperature & low pressure
(3) Low temperature & high pressure (4) Low temperature & low pressure.

45. In the following reaction
$SO_2Cl_2(g) \rightleftharpoons SO_2(g) + Cl_2(g)$
if inert gas is insert at equilibrium then the which of the following statement is correct for this reaction at constant volume :–
(1) High quantity of Cl_2 is produce
(2) Decrease concentration of SO_2
(3) High quantity of SO_2Cl_2 is formed
(4) Does not change in concentration of SO_2Cl_2, SO_2 and Cl_2

EXERCISE-III

QUESTIONS ASKED IN PREVIOUS EXAMS

1. A sealed container at a certain temperature is half full of water. The temperature of the container is increased and maintained till equilibrium is re-established. Which statement is correct when the equilibrium is re-established at the higher temperature ?
(1) The rate of vaporization is greater than the rate of condensation
(2) The amount of water vapour is greater than the amount of liquid water
(3) The amount of water vapour is greater than what is was at the lower temperature.
(4) The rate of condensation is greater than the rate of vaporization.

2. Which statements are correct for a reaction at equilibrium ?
I. The forward and reverse reactions both continue.
II. The rates of the forward and reverse reactions are equal.
III. The concentration of reactants and products are equal
(1) I and II only (2) I and III only (3) II and III only (4) I, II and III

3. Increase in pressure shifts the equilibrium of the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
(1) to produce more N_2 (2) to produce more NH_3
(3) to reduce the temperature of the reaction (4) to produce more H_2

4. When pressure is applied to equlibrium system, Ice (s) $\rightleftharpoons$ H_2O (ℓ) which of the following will happen ?
(1) More ice will be formed (2) Water will evaporate
(3) More water will be formed (4) There will be no change

5. The equilibrium constant for the reaction $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$ is 4.0×10^{-4} at 2000 K. In presence of a catalyst equilibrium is attained 10 times faster. Therefore equilibrium constant in presence of catalyst at 2000 K is
(1) 40×10^{-4} (2) 4×10^{-4} (3) 4×10^{-2} (4) none of these

6. A mixture of 0.3 moles of H_2 and 0.3 moles of I_2 are allowed to react in 10 litre evacuated flask at 500°C. The reaction is $H_2 + I_2 \rightleftharpoons 2HI$ Value of k is found to be 64. The amount of I_2 present at equilibrium is
(1) 0.15 mole (2) 0.06 mole (3) 0.03 mole (4) 0.2 mole

7. Which of the following will change the value of the equilibrium constant for the reaction ?
$N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$
(1) Add more N_2 (2) Increase of pressure
(3) Use a smaller reaction vessel (4) Increase the temperature

8. Assertion (A): If the volume of the vessel is doubled then for the following reaction
$A(g) \rightleftharpoons U(g) + C(g)$
Equilibrium constant is decreased
Reason (R): Equlibrilum constant $Kc = X^2/(x-X)V$.
(1) Both (A) and (R) are true and (R) is the correct explanation of (A).
(2) Both (A) and (R) are true and (R) is not the correct explanation of (A).
(3) (A) is true but (R) is false
(4) (A) is false but (R) is true

9. For the reaction : $aE + bF \rightleftharpoons cG + dH$
The expression of equilibrium constant can be written as $Kc = [G]^C [H]^d/[E]^a [F]^b$
Where all the concentrations are equilibrium concentrations. Before approaching equilibrium, the same concentration ratio is called reaction quotient Qc.
For the reaction system to reach equilibrium
(1) Q_c must increase in the reaction (2) Q_c must decrease in the reaction
(3) $Q_c = K_c$ (4) Q_c = zero

10. In Haber's process 0.240 mole of Nitrogen, 3.9 moles of hydrogen are taken which lead to the formation of 7.8 moles product in a 3.00 litres of reaction vessel at 375°C. Considering that equilibrium constant at this temperature is 41.2 Calculate the value of reaction quotient (Q) and predict whether the reaction is in equilibrium or it will proceed in either direction

(1) Q = 38.62 and reaction will be in equilibrium.
(2) Q = 19.31 and reaction will proceed in forward direction.
(3) Q = 38.62 and reaction will proceed in forward direction.
(4) Q = 19.31 and reaction will proced in backward direction.

11. At 5 atm pressure PCl_5 gas dissociates by 10%. What will be the value of Kp at same tamperature?

(1) 0.0 –5 atm (2) 0.050 atm (3) 0.9 atm (4) 0. atm

12. Consider the following reaction : 4 $PCl_3(g) \rightarrow P_4(g) + 6Cl_2(g)$. If the initial concentration of $PCl_3(g)$ is 1.0 M, and "x" is the equilibrium concentration of $P_4(g)$, what is the correct equilibrium relation?

(1) $Kc = 6x^7$ (2) $Kc = 6x^7/(1.0 - x)^4$ (3) $Kc = (x)(6x)^6/(1.0 - 4x)^4$ (4) $Kc = x^7/(1.0 - x)$

13. The equilibrium constant for the gaseous reaction $N_2 + O_2 \rightarrow 2NO$ is K. The equilibrium constant for the formation of one mole of NO will be

(1) K/2 (2) K (3) $\sqrt{K}$ (4) 2K

14. In the Haber process for the synthesis of ammonia, amount of ammonia, amount of ammonia formed will be more if

(1) pressure is decreased and temperature is increased
(2) pressure is increased and temperature is decreased
(3) both pressure and temperature are increased
(4) both pressure and temperature are decreased

15. The following reaction is at equlibrium in a given cylinder of constant volume.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Addition of argon gas to the cylinder will.....

(1) reduce the formation of NH_3 with no change in temperature of the equilibrium system.
(2) Increase the formation of NH_3 with increase in temperature.
(3) reduce the formation of NH_3 with increase in temperature.
(4) made up of specialized muscle tissues and located near the heart.

ANSWER KEY

Exercise-I

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	4	1	4	3	1	2	2	3	2	3	3	2	2	1	4	4	3	3	1	2
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	3	1	1	3	4	1	1	3	1	2	4	1	2	3	3	1	4	1	3	3
Que.	41	42	43	44	45	46	47	48	49	50	51	52	53	54	55	56	57	58	59	60
Ans.	4	2	2	2	3	3	1	2	2	4	3	3	1	2	4	1	4	1	3	2
Que.	61	62	63	64	65															
Ans.	2	4	3	4	1															

Exercise-II

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20
Ans.	4	4	2	3	1	3	3	4	4	2	3	3	3	4	2	2	3	1	3	2
Que.	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	37	38	39	40
Ans.	3	3	2	4	2	2	3	3	1	3	1	1	1	2	3	3	1	3	3	2
Que.	41	42	43	44	45															
Ans.	3	3	4	3	4															

Exercise-III

Que.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	3	1	2	3	2	2	4	4	1	3	2	3	3	2	3

SPACE FOR ROUGH WORK

ACIDS, BASES & SALTS

INTRODUCTION

About three million (30 lakh) chemical compounds are known till now. These compounds may be classified on the basis of their properties such as colour, taste, solubility and chemical behaviour. The earliest classification of compounds into acids, bases and salts was primarily based on their taste and action on litmus.

CLASSIFICATION OF ACIDS

On the basis of their Sources

On the basis of their **sources**, acids can be classified in two categories:

1. Organic acids   2. Inorganic acids

1. Organic acids

The acids which are usually obtained from plants are known as **organic acids.** Oxalic acid [$(COOH)_2$], acetic acid (CH_3COOH) etc. are very common examples of organic acids. Some other organic acids with their natural sources are given in Table.

Some Organic Acids with Their Natural Sources

S.No.	*Organic acid*	*Natural sources*	*S.No.*	*Organic acid*	*Natural sources*
1	Acetic acid	Vinegar	7	Oleic acid	Olive oil
2	Citric acid	Citrus fruits (like orange and lemon)	8	Stearic acid	Fats
3	Butyric acid	Rancid butter	9	Amino acid	Proteins
4	Formic acid	Sting of bees and ants	10	Uric acid	Urine
5	Lactic acid	Sour milk	11	Tartaric acid	Tamarind
6	Malic acid	Apples	12	Oxalic acid	Tomatoes

It may be noted that all organic acids contain carbon as one of their constituting elements. These are weak acids and therefore, do not ionise completely in their aqueous solutions. Since these acids do not ionise completely in their aqueous solutions, therefore, their solutions contains both ions as well as undissociated molecules. For example, formic acid aqueous solution contains H_3O^+, $HCOO^-$ as well as undissociated HCOOH molecules.

$$\underset{\text{Formic acid}}{HCOOH} + H_2O \rightleftharpoons \underset{\text{Hydronium ion}}{H_3O^+} + \underset{\text{Formate ion}}{HCOO^-}$$

List of Common Organic Acids

Name	*Type*	*Chemical Formula*	*Where found or used*
Formic acid	Organic acid	$HCOOH$ (CH_2O_2)	Found in the stings of ants and bees, used in tanning leather and in medicines for treating gout.
Acetic acid	Organic acid	CH_3COOH ($C_2H_4O_2$)	found in vinegar, used as solvent in the manufacture of dyes and
Lactic acid	Organic acid	$CH_3CH(OH)COOH$ ($C_3H_6O_3$)	Responsible for souring of milk and present in curd.
Benzoic acid	Organic acid	C_6H_5COOH	Used as food preservative.
Citric acid	Organic acid	$C_6H_8O_7$	Present in lemons, oranges and other citric fruits.
Tartaric acid	Organic acid	$C_4H_6O_6$	Present in tamarind.

2. Inorganic Acids

The acids which are usually obtained from minerals are known as **inorganic acids.** Since the acids are obtained from minerals, therefore, also called **mineral acids.** Some common examples of inorganic acids are : Hydrochloric acid (HCl), Sulphuric acid (H_2SO_4), Nitric acid (HNO_3) etc.

List of Common Inorganic Acids

Name	*Type*	*Chemical Formula*	*Where found or used*
Phosphoric acid	Mineral acid	H_3PO_4	Used in anti-rust paints and in fertilizers.
Carbonic acid	Mineral acid	H_2CO_3	In soft drinks and lends fizz.
Hydrochloric acid	Mineral acid	HCl	In stomach as gastric juice, used in cleaning metal surfaces and in tanning
Nitric acid	Mineral acid	HNO_3	Used in the manufacture of explosives and fertilizers.
Sulphuric acid	Mineral acid	H_2SO_4	Commonly used in car batteries, in the manufacture of fertilizers, detergents etc.

On the basis of their strength

1. Strong Acids

The acids which undergo almost complete ionisation in a dilute aqueous solution, thereby producing a high concentration of hydronium ions (H_3O^+) are known as **strong acids.**

Example of strong acids : Some examples of strong acids are :

(i) Hydrochloric acid (HCl) (ii) Sulphuric acid (H_2SO_4) (iii) Nitric acid (HNO_3)

All these three mineral acids are considered to be strong acids because they ionise almost completely in their dilute aqueous solutions.

2.Weak Acids

The acids which undergo partial or incomplete ionisation in a dilute aqueous solution, thereby producing a low concentration of hydronium ions (H_3O^+) are known as **weak acids.**

Examples of weak acids

Some examples of weak acids are :

(i) Acetic acid (CH_3COOH)
(ii) Formic acid (HCOOH)
(iii) Oxalic acid [$(COOH)_2$]
(iv) Carbonic acid (H_2CO_3)
(v) Sulphurous acid (H_2SO_3)
(vi) Hydrogen sulphide (H_2S)
(vii) Hydrocyanic acid (HCN)

The aqueous solution of weak acids contain both ions as well as undissociated molecules.

On the basis of Concentration of the Acid

By the term **concentration,** we mean the amount of water present in the given sample of acid solution in water.

1. Concentrated Acid

The sample of an acid which contains very small or no amount of water is called a **concentrated acid.**

2. Dilute Acid

The sample of an acid which contains far more amount of water than its own amount is known as a **dilute acid.**

It must be mentioned here that concentration of an acid simply tells the amount of water in the acid. It may not be confused with strength of an acid, which is a measure of concentration of hydronium ion it produces in aqueous solution.

PHYSICAL PROPERTIES OF ACIDS

1. Taste

Acids have a **sour** taste. Since acids like sulphuric acid and nitric acid are highly corrosive in nature, therefore these are generally not tasted.

2. Physical state

Some acids are solids while others are liquids at room temperature. Physical states of some common acids are given in Table

Physical State of Some Common Acids

S.No.	Acid	Physical state at room temperature	S.No.	Acid	Physical state at room temperature
1	Oxalic acid [$(COOH)_2$]	Solid	6	Hydrochloric acid (HCl)	Volatile liquid
2	Boric acid (H_3BO_3)	Solid	7	Nitric acid (HNO_3)	Volatile liquid
3	Acetic acid (CH_3COOH)	Liquid	8	Sulphurous acid (H_2SO_3)	Volatile liquid
4	Formic acid ($HCOOH$)	Liquid	9	Sulphuric acid (H_2SO_4)	Non-volatile liquid
5	Carbonic acid (H_2CO_3)	Volatile liquid			

3. Effect on indicators

Acids change the colour of **blue litmus to red,** carbonic acid (H_2CO_3), which is a weak mineral acid, turns **blue litmus pink.**

4. Effect on skin

All strong mineral acids have a corrosive action on skin and cause painful burns. For example, concentrated sulphuric acid stains the skin black while concentrated nitric acid and concentrated hydrochloric acid stains the skin to yellow and amber colours respectively. However, carbonic acid (H_2CO_3) is an exception and is not corrosive in nature.

5. Electrical conductivity

All mineral acids are good conductors of electricity and conduct electricity in their aqueous solutions. On electrolysis, they decompose liberating **hydrogen** gas at **cathode.**

CLASSIFICATION OF ALKALIS AND BASES

On the basis of strength

1. Strong alkalis or bases.

The alkalis or bases which undergo almost complete ionisation in aqueous solution to produce high concentration of hydroxyl (OH^-) ions are known as **strong alkalis or strong bases.**

Example of strong alkalis or bases- Some example of strong alkalis or bases are: Sodium hydroxide ($NaOH$), Potassium hydroxide (KOH) and Barium hydroxide [$Ba(OH)_2$] etc.

$$\left.\begin{array}{l} NaOH(aq.) \longrightarrow Na^+(aq.) + OH^-(aq.) \\ KOH(aq.) \longrightarrow K^+(aq.) + OH^-(aq.) \\ Ba(OH)_2(aq.) \longrightarrow Ba^{2+}(aq.) + 2OH^-(aq.) \end{array}\right\} \text{(Almost completely ionised)}$$

2. Weak alkalis or bases

The alkalis or bases which undergo only partial ionisation in aqueous solution to produce a relatively low concentration of hydroxyl (OH^-) ions are known as **weak alkalis or weak bases.**

Example of weak alkalis or bases.

Some examples of weak alkalis or bases are: Ammonium hydroxide (NH_4OH), Calcium hydroxide [$Ca(OH)_2$], Magnesium hydroxide [$Mg(OH)_2$] etc.

$$NH_4OH\ (aq.) \rightleftharpoons NH_4^+\ (aq) + OH^-\ (aq)$$

$$Ca(OH)_2\ (aq.) \rightleftharpoons Ca^{2+}\ (aq) + 2OH^-\ (aq)$$

$$Mg(OH)_2\ (aq.) \rightleftharpoons Mg^{2+}\ (aq) + 2OH^-\ (aq)$$

Since these alkalis are not ionising completely, therefore, there is a dynamic equilibrium between the undissolved alkali and the ions produce by it.

On the basis of their concentration

By the term **concentration,** we mean the amount of water present in the given sample of alkali solution in water. On the basis of concentration, the alkalis can be classified as under :

1. Concentrated alkali-A solution of alkali having a relatively high percentage of alkali in its aqueous solution is known as **concentrated alkali.**

2. Dilute alkali- A solution of alkali having a relatively low percentage of alkali in its aqueous solution is known as a **dilute alkali.**

Some Common Bases

Name	*Commercial name*	*Chemical formula*	*Uses*
Sodium hydroxide	Caustic soda	$NaOH$	In manufacture of soap, refining of petroleum, paper, pulp etc.
Potassium hydroxide	Caustic potash	KOH	In alkaline storage battery, manufacture of soap, absorbing CO_2 gas etc.
Calcium hydroxide	Slaked lime	$Ca(OH)_2$	In manufacture of bleaching powder, softening of hard water etc.
Magnesium hydroxide	–	$Mg(OH)_2$	As an antacid to remove acidity from stomach.
Aluminium hydroxide	–	$Al(OH)_3$	As foaming agent in fire extinguishers.
Ammonium hydroxide	–	NH_4OH	In removing grease stains from clothes, and in cleaning window panes.

Distinction Between an Alkali and a Base- It may be kept in mind that a base which is soluble in water is called an alkali. This means that all alkalis are bases but all bases are not alkalis. For example, ferric hydroxide [$Fe(OH)_3$] and cupric hydroxide [$Cu(OH)_2$] are bases but not termed as alkalis because they are insoluble in water.

PHYSICAL PROPERTIES OF BASES

Important physical properties of bases or alkalis are :

1. Taste

They are sharp **bitter** in taste.

2. Feeling of touch

They give a feeling of soapy touch i.e. when their solutions are felt between the fingers, one feels a slippery sensation.

Have you ever thought why caustic potash or caustic soda give a feeling of soapy touch?

One of the important properties of alkalis is that they react with oils to form soap. Since our skin contains oil in the form of fat, therefore, when we touch caustic soda or caustic potash, it reacts with the oil present in our skin to form soap which gives a feeling of soapy touch.

3. Effect on indicators

Bases change the colour of **red litmus to blue**

4. Effect on skin

All alkalis have a mild corrosive action on skin.

INDICATORS

An indicator is a chemical compound which has one colour in an acidic solution and a different colour in a basic solution or a neutral solution. For example :

(i) Litmus is an indicator which is red in an acidic solution but it has blue colour in a basic solution.

(ii) Methyl orange, methyl red, and phenolphthalein are some other examples of indicators.

These indicators change colour when the nature of the solution is changed (from acidic to basic or vice versa).Therefore they are also known as **visual indicators**.

Types of Indicators

1. Olfactory Indicators - It may be noted that visual indicators are not only means to check that a given substance is acidic or basic simply noticing by change in colour. There are some substances whose odour changes in acidic or basic solutions. These are called **olfactory indicators**. The commonly used olfactory indicators are raw onion, vanilla extract and clove oil. Just as visual indicators change colour in response to acidic or basic solution, an olfactory indicator will change either its odour or odour intensity with change in acidic or basic nature of solution.

2. Acid - base indicators- These are the indicators, which indicate the presence of acid or base in an aqueous solution. Litmus, methyl orange and phenolphthalein are commonly used indicators in the chemistry laboratory.

Action of Acids with Indicators

1. **Action of acids on litmus :** An aqueous solution of an acid (like hydrochloric acid) changes blue litmus to red. However acids do not affect red litmus.
2. **Action of acids on phenolphthalein :** An acidic solution remains colourless when phenolphthalein indicator is added to it.
3. **Action of acids on methyl orange :** An acidic solution changes yellow colour of methyl orange to red or pink.

Change of Colour of Indicators in Different Acids

Name of Indicator	***Original Colour of Indicator***	***Colour of Indicator in the Acid***			
		HCl	H_2SO_4	HNO_3	CH_3COOH
1. Litmus	Blue	Red	Red	Red	Red
2. Methyl orange	Yellow	Red	Red	Red	Pink
3. Phenolphthalein	Colourless	Colourless	Colourless	Colourless	Colourless

Action of bases with indicators

Water soluble bases affect indicators in the following ways :

1. **Action of bases on litmus-** An aqueous solution of a base (like sodium hydroxide) changes red litmus to blue. However, bases do not affect blue litmus.
2. **Action of bases on phenolphthalein-** A basic solution turns pink when phenolphthalein indicator is added to it.
3. **Action of bases on turmeric solution-** It is a yellow solution. When a base is added to it, a brown colour is produced. Turmeric (haldi) spots on kitchen cloth turn brown-red when washed with soap. It is because soap contains alkali which interacts with turmeric.
4. **Action of bases on methyl orange-** Bases do not change the colour of methyl orange indicator.

Note. Water insoluble bases do not affect indicators.

Change of Colour of Indicators in NaOH Solution

Name of Indicator	***Original Colour of Indicator***	***Colour of Indicator in NaOH Solution***
1. Litmus	Red	Blue
2. Methyl orange	Yellow	No change
3. Phenolphthalein	Colourless	Pink
4. Turmeric solution	Yellow	Reddish brown

CHEMICAL PROPERTIES OF ACIDS AND BASES

Chemical Properties of acids

1. Reaction of acids with metals

$$Zn + 2HCl \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2} + H_2$$

$$Fe + H_2SO_4 \longrightarrow \underset{\text{Ferrous sulphate}}{FeSO_4} + H_2$$

$$Mn + \underset{\text{(Dilute, 1\%)}}{2HNO_3} \longrightarrow \underset{\text{Manganese nitrate}}{Mn(NO_3)_2} + H_2$$

Do you know why we use a very dilute solution of nitric acid in the last reaction?

This is because nitric acid is an oxidising agent which oxidises hydrogen to form water.

From above reactions, we conclude that metals combine with acids to form salts and hydrogen. Now, if you compare the rate of evolution of hydrogen gas in the above reactions, you will observe that some metals react vigorously at room temperature, while other metals react slowly.

3. Reaction of acids with metallic oxides

All metal oxides react with dilute mineral acids to form their respective metallic salts and water. This is supported by the following reactions :

(i) $$\underset{\text{Sodium oxide}}{Na_2O} + 2HNO_3\,(dil.) \longrightarrow \underset{\text{Sodium nitrate}}{2NaNO_3} + H_2O$$

(ii) $$\underset{\text{Copper (II) oxide}}{CuO} + 2HCl\,(dil.) \longrightarrow \underset{\text{Copper (II) chloride}}{CuCl_2} + H_2O$$

4. Reaction of acids and bases

All metal hydroxides (also known as bases) react with dilute mineral acids to form their respective salts and water only. This reaction is also known as **acid-base neutralization reaction.** For example:

$$\underset{\text{Potassium hydroxide}}{2KOH\,(aq.)} + H_2SO_4(aq.) \rightarrow \underset{\text{Potasium sulphate}}{K_2SO_4\,(aq.)} + H_2O(l)$$

Chemical Properties of bases

Important properties of alkalis or bases are discussed below :

1. Action on metals.

$$\underset{\text{(Zinc)}}{Zn} + 2NaOH \longrightarrow \underset{\text{Sodium zincate}}{Na_2ZnO_2} + H_2\,(g)\uparrow$$

2. Action of air

Some of the alkalis like sodium hydroxide (NaOH) are highly **hygroscopic** in nature, i.e., they absorb moisture from air and gets dissolved in it. They also absorb carbon dioxide from the air to form their respective carbonates. For example,

(i) $$\underset{\text{Sodium hydroxide}}{2NaOH} + CO_2 \longrightarrow \underset{\text{Sodium carbonate}}{Na_2CO_3} + H_2O$$

(ii) $$\underset{\text{Potassium hydroxide}}{2KOH} + CO_2 \longrightarrow \underset{\text{Potassium carbonate}}{K_2CO_3} + H_2O$$

3. Action of acids

They neutralise the acids to form salt and water. For example,

(i) $$\underset{\text{Sodium hydroxide}}{NaOH} + \underset{\text{Hydrochloric acid}}{HCl} \longrightarrow \underset{\text{Sodium chloride}}{NaCl} + H_2O$$

(ii) $$\underset{\text{Ferrous hydroxide}}{Fe\,(OH)_2} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Ferrous chloride}}{FeCl_2} + 2H_2O$$

HOW STRONG ARE ACID OR BASE SOLUTIONS? (pH SCALE)

Acids and bases may be either strong or weak. We can compare their relative strengths on the basis of the Arrhenius theory. According to the theory, more the number of H^+ ions released by acid in water, stronger is the acid. Similarly, more the number of OH^- ions released by base in water, stronger is the base.

Dilution of acids and bases with water decreases the concentration of $H^+(aq)$ or $OH^-(aq)$ ions in the acidic and basic solution respectively.

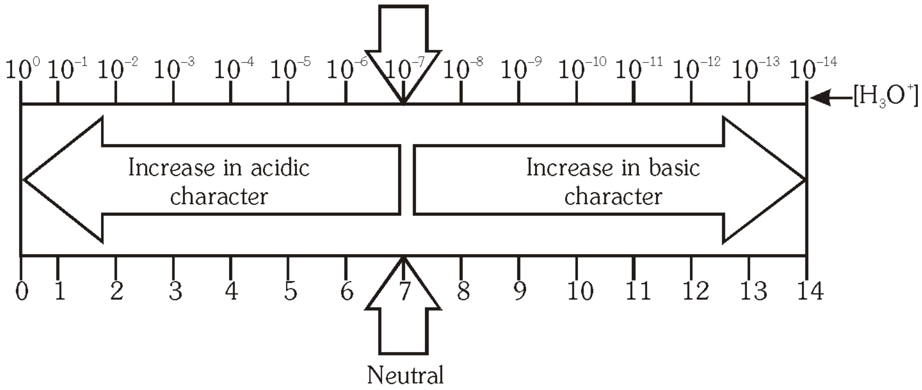
If we find quantitatively, the amount of $H^+(aq)/OH^-(aq)$ ions present in a solution, we can judge how strong an acid or a base is?

We can do this by the help of a **universal indicator**, which is a mixture of several indicators. The universal indicator shows different colours at different concentration of hydrogen ions or pH values in solution.

pH Scale

S.P.L. Sorenson, a Danish chemist in 1909 introduced the concept of measuring the concentration of hydrogen ions [$H^+(aq)$] in a particular solution. The 'p' in pH stands for 'potenz' in German, meaning power.

On the pH scale we can measure pH from "0" (very acidic) to 14 (very alkaline).



pH scale

(i) Solutions with pH = 0-3 are strongly acidic, with pH = 3–5 are moderately acidic while with pH = 5–7 are weakly acidic.

(ii) Solution with pH = 7–9 are weakly basic, with pH = 9–12 are moderately basic while with pH = 12–14 are strongly basic.

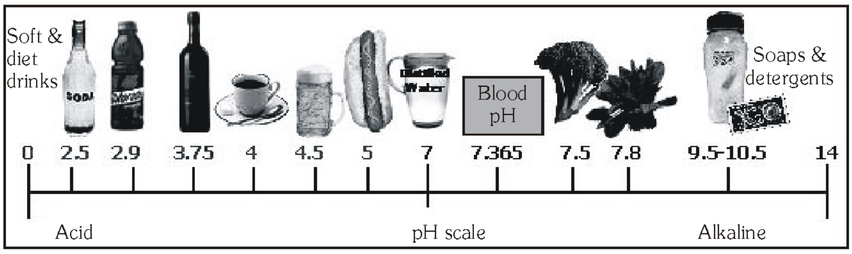
(iii) If pH = 7, then the solution is neutral.

pH value of solutions

S.No.	Solution	Colour of pH paper	Approximate pH value	Nature of substance
1	Saliva (before meal)	Blue	7.4	Basic
2	Saliva (after meal)	Yellow	5.8	Acidic
3	Lemon juice	Red	2.3	Acidic
4	Colourless aerated drink	Red-orange	3	Acidic
5	Carrot juice	Yellowish green	6	Acidic
6	Coffee	Yellow	5	Acidic
7	Tomato juice	Orange	4.1	Acidic
8	Tap water	Greenish blue	7.4	Basic
9	1M NaOH	Dark violet	14	Basic
10	1M HCl	Dark red	0	Acidic

Conclusion

Solutions with pH < 7 but close to 7 are weakly acidic, solutions with pH << 7 are strongly acidic, solution with pH > 7 but close to 7 are weakly basic, solutions with p>>7 are strongly basic.



pH of some common substances.

Importance of pH in every day life

1. Plants and animals are pH sensitive

The pH plays an important role in the survival of animals, including human being. Our body works well with in a narrow pH range of 7.0 to 7.8. The aquatic animals like fish can survive in river water within a narrow range of pH change.

Example : When the pH of rain water is less than 5.6, it is called acid rain. Acid rain can lower the pH of river water to such an extent and make it so acidic that the survival of aquatic animals becomes difficult or may even kill the aquatic animals.

2. Soil pH and plants

The growth of plants in a particular soil is related to its pH. Actually, different plants prefer different pH range for their growth. It is therefore, quite important to provide the soil with proper pH for their healthy growth. Plants require a specific pH for their growth. If the soil is **too acidic** or **too basic** the plants grow badly or do not grow at all.

3. Importance of pH in our digestive system

As we know our stomach produces gastric juice which contains large amount of hydrochloric acid (pH about 1.4). The acid so produced does not harm the stomach walls, but kills germs and bacteria which enter in our digestive system along with food, thus in a way it protects us from diseases and helps in digestion. Sometimes excess of acid is produced in the stomach due to overeating or eating spicy foods. This stage is called **acidity**. To get relief from this pain, we take tablets known as **antacids**. These contain bases to neutralise the excess acids. Example : Magnesium hydroxide (milk of magnesia), $Mg(OH)_2$

4. pH change as the cause of tooth decay

Generally, the pH in the mouth is more than 7, as the saliva produced in the mouth is basic in nature. However, when we take food, some food particles remain in the mouth after eating and bacteria present in the mouth produce acids by degradation of food particles.

This acid lowers the pH in the mouth, tooth decay starts when the pH of acid formed in the mouth falls below 5.5 and the enamel get corroded.

The bacterias present in the mouth break down the sugar, that we eat, into acids. Lactic acid is one of these. The formation of these acids causes decrease in pH. The acids react with calcium phosphate and the enamel coating slowly breaks.

Therefore, to prevent tooth decay, it is advised to clean the mouth and use toothpastes which are generally basic, for cleaning the teeth. It neutralise the excess acid and prevent tooth decay.

5. Self defence by animals and plants through chemical warfare

The sting of the honey bee contains formic acid, this acid causes a lot of irritation and pain. The pain can be reduced by applying baking soda paste on the affected region as the acid gets neutralised.

In plant kingdom **nettle** (Bichu Booti) is a herbaceous plant which grows in wild. The nettle leaves have stinging hair. When a person happens to touch the leaves of a nettle plant accidently, the stinging hair of nettle leaves inject methanoic acid (HCOOH) into the skin of the person causing burning pain. The nettle sting, being acidic can be neutralised by rubbing baking soda on the skin. Nature provides remedy for the nettle sting in the form of a 'dock' plant, which often grows besides the nettle plants. The leaves of dock plant contain some basic chemicals which neutralises methanoic acid.

Some naturally occurring acids

Natural source	Acid	Natural source	Acid
Vinegar	Acetic acid	Sour milk (Curd)	Lactic acid
Orange	Citric acid	Lemon	Citric acid
Tamarind	Tartaric acid	Ant sting	Methanoic acid
Tomato	Oxalic acid	Nettle sting	Methanoic acid

CONCEPTS OF ACIDS AND BASES

1. ARRHENIUS CONCEPT

Acid: According to this concept, those substances which produce free H^+ ions in aqueous solution are called acids.

Ex. HCl, HNO_3, H_2SO_4, H_3PO_4, H_2CO_3, H_2S, CH_3COOH etc.

$$HCl + H_2O \longrightarrow H^+(aq.) + Cl^-(aq.)$$

$$H_2SO_4 + H_2O \longrightarrow 2H^+(aq.) + SO_4^{-2}(aq.)$$

$$CH_3COOH + H_2O \longrightarrow CH_3COO^-(aq.) + H^+(aq.)$$

Base: Those substances which produce free OH^- ions in aqueous solution are called bases.

Ex. $NaOH$, KOH, $Cs(OH)$, $Rb(OH)$, NH_4OH, $Ba(OH)_2$, $Ca(OH)_2$ etc.

$$NaOH + H_2O \longrightarrow Na^+ (aq.) + OH^- (aq.)$$

$$Ba(OH)_2 + H_2O \longrightarrow Ba^{+2} (aq.) + 2OH^- (aq.)$$

Nature of water: According to this concept nature of water is neutral and act as a solvent.

Neutralisation Reaction : Those reactions in which acid and base react together to form water molecule are called neutralisation reactions.

i.e. $Na^+ + OH^- + H^+ + Cl^- \longrightarrow NaCl + H_2O$

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

Degree of ionisation

The fraction of the total number of molecules which dissociate into ions is called the degree of dissociation or degree of ionization and is usually represented by α.

i.e., $\boxed{\alpha = \dfrac{\text{No. of moles dissociated}}{\text{Total no. of moles taken}}}$

Advantage

This concept explains the acids and bases practically. i.e. to find out the pH, ionisation constant, hydrolysis constants, heat of neutralisation etc.

Disadvantage

(i) It explains the behaviour of acids and bases only in aqueous medium.

(ii) It does not explain the stability of proton (H^+).

(iii) It does not explain the conjugate acid-base theory.

(iv) It does not explain the acidic and basic behaviour of aprotic acid and base for example - SO_2, CO_2, $SiCl_4$, $AlCl_3$ etc.

2. BRONSTED-LOWRY CONCEPT

It is based upon the exchange of proton.

Acid : According to this concept those substances which have tendency to donate the proton (H^+) by any method in any solvent are called acids.

Ex. (i) HCl, HNO_3, H_2SO_4, H_2CO_3, H_2S, CH_3COOH, H_3PO_3 etc.

(ii) HS^-, HCO_3^-, HSO_4^-, $H_2PO_4^-$, HPO_4^{-2}, H_2O etc.

(iii) NH_4^+, H_3O^+, PH_4^+, $CH_3COOH_2^+$ etc.

(vi) $[Al(H_2O)_6]^{+3}$, $[Ag(H_2O)_2]^{+1}$, $[Fe(H_2O)_6]^{+3}$ etc.

(a) HCl (Acid) + H_2O (Solvent) $\longrightarrow$ $H_3O^+ + Cl^-$

(b) HS^- (Acid) + H_2O (Solvent) $\longrightarrow$ $H_3O^+ + S^{-2}$

(c) NH_4^+ (Acid) + H_2O (Solvent) $\longrightarrow$ $NH_3 + H_3O^+$

(d) $[Al(H_2O)_6]^{+3}$ (Acid) + H_2O (Solvent) $\longrightarrow$ $[Al(H_2O)_5OH]^{+2} + H_3O^+$

Base : Those substances which have tendency to accept the proton by any method in any solvent are called bases.

Ex. (i) $NaOH$, KOH, $Rb(OH)$, $Cs(OH)$, $Ba(OH)_2$, $Ca(OH)_2$, NH_4OH, $Al(OH)_3$ etc.

(ii) HS^-, HCO_3^-, HSO_4^-, $H_2PO_4^-$, HPO_4^{-2}, H_2O etc.

(iii) NH_3, RNH_2, R_2NH, R_3N, $C_6H_5NH_2$, C_5H_5N, $H_2N\text{-}NH_2$ etc.

(iv) O^{-2}, SO_4^{-2}, CO_3^{-2}, Cl^-, Br^-, I^-, CN^- etc.

(a) $NaOH$ (Base) + H_2O (Solvent) $\longrightarrow$ $OH^- + NaOH_2^+ \xrightarrow[\text{Unstable}]{} Na^+ + H_2O$

(b) HS^- (Base) + H_2O (Solvent) $\longrightarrow$ $H_2S + OH^-$

(c) NH_3 (Base) + H_2O (Solvent) $\longrightarrow$ $NH_4^+ + OH^-$

(d) CO_3^{-2}(Base) + H_2O (Solvent) $\longrightarrow$ $HCO_3^- + OH^-$

Nature of water : According to this concept nature of water is amphoteric or amphiprotic i.e. water can act as both acid and base.

(i) HCl (Acid)+H_2O (Base) $\longrightarrow$ $Cl^-+H_3O^+$ (ii) NH_3 (Base)+H_2O (Acid) $\longrightarrow$ $NH_4^++OH^-$

Neutralisation Reaction: According to this concept those reaction in which acid and base react together and convert into their respective conjugate acid and base are called neutralisation reactions.

Ex.

$$\underset{\text{Acid}}{HCl} + \underset{\text{Base}}{NH_3} \rightleftharpoons \underset{\text{Conjugate Base}}{Cl^-} + \underset{\text{Conjugate Acid}}{NH_4^+}$$

Strength of acids and bases: This concept explain the strength of acid and base depending upon the basis of relative tendency to accept or donate the proton.

Ex.

(i) HCl	+	H_2O	$\rightleftharpoons$	Cl^-	+	H_3O^+
Strong acid		Strong base		Weak base		Weak acid
(ii) CH_3COOH	+	H_2O	$\rightleftharpoons$	CH_3COO^-	+	H_3O^+
Weak acid		Weak base		Strong base		Strong acid
(iii) NH_3	+	H_2O	$\rightleftharpoons$	NH_4^+	+	OH^-
Weak base		Weak acid		Strong acid		Strong base
(iv) HCl	+	CH_3COOH	$\rightleftharpoons$	Cl^-	+	$CH_3COOH_2^+$
Weak acid		Weak base		Strong base		Strong acid
(v) HCl	+	NH_3	$\rightleftharpoons$	Cl^-	+	NH_4^+
Strong acid		Strong base		Weak base		Weak acid
(vi) CH_3COOH	+	NH_3	$\rightleftharpoons$	CH_3COO^-	+	NH_4^+
Strong acid		Strong base		Weak base		Weak acid

Special points

Types of solvent

(a) Protogenic or acidic solvent: They have a tendency to generate or donate the protons.

Ex. CH_3COOH, HCl, HNO_3, H_2SO_4 etc.

(b) Protophilic or basic solvent: They have a tendency to accept the protons.

Ex. NH_3, CH_3OH, RNH_2 etc.

(c) Amphiprotic or amphoteric solvents: They have the tendency to accept or donate the protons or they are either protophilic or protogenic.

Ex. H_2O, HS^-, C_2H_5OH, HSO_4^-, HCO_3^- etc.

(d) Aprotic solvents: They neither donate nor accept the protons.

Ex. C_6H_6, CCl_4, $CHCl_3$, BrF_3, NO_2, $COCl_2$ etc.

Advantage

(i) It explain the behaviour of acids and bases in any type of solvent.

(ii) It explain the stability of proton (H^+).

Disadvantage

(i) It does not explain the acid and base practically.

(ii) According to this concept, every acid and base reaction is a neutralisation reaction which is not true.

(iii) It also fails to explain the acidic and basic behaviour of aprotic acids and bases.

For ex. SO_2, SO_3, CO_2, $AlCl_3$, $SiCl_4$, etc.

3. LEWIS CONCEPT

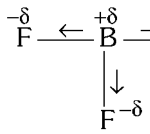
Lewis acid: According to this concept those species which have self tendency to accept the lone pair of electrons are called acids. i.e. Lewis acid is an electron pair acceptor (electrophilic).

$$\underset{\text{Acid}}{H^+} + \underset{\text{Base}}{:\ddot{\underset{..}{O}}{}_{\times}H^-} \longrightarrow H:\ddot{\underset{..}{O}}{}_{\times}H$$

Classification of Lewis acids

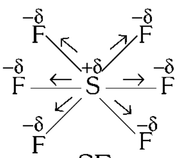
(i) Compounds whose central atom have an incomplete octet (electron deficient)

Ex. BF_3, BBr_3, BCl_3, BI_3, $B(CH_3)_3$, $B(OH)_3$, $AlCl_3$, $GeCl_3$ etc.

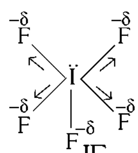


(ii) Compounds whose central atom have vacant d-orbitals and can accept one or more pair of lone electrons.

Ex. SF_4, SF_6, $SnCl_2$, $SnCl_4$, PX_3, PX_5, GeX_4, TeX_4, IF_5, IF_7, etc.



SF_6

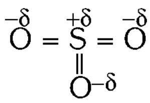


IF_5

(iii) Molecules with a multiple bond between atoms of different electronegativities.

Ex. CO_2, SO_2, SO_3 etc.

$\overset{-\delta}{O} = \overset{+\delta}{C} = \overset{-\delta}{O}$ $\qquad$ $\overset{-\delta}{O} = \overset{+\delta}{S} = \overset{-\delta}{O}$



$$\underset{\text{Lewis acid}}{O = C = O} \quad + \quad \underset{\text{Lewis base}}{OH^-} \longrightarrow {}^-O - \underset{\substack{|\\ OH}}{C} = O \longrightarrow HCO_3^-$$

$$\underset{\text{Lewis acid}}{SO_3} \quad + \quad \underset{\text{Lewis base}}{H_2O} \longrightarrow H_2SO_4$$

(iv) Cations

Ex. Ag^+, Na^+, Li^+, Al^{+3}, Be^{+2}, Mg^{+2}, I^+, Cl^+, H^+ etc.

False cations

Ex. 1. NH_4^+, H_3O^+, $CH_3COOH_2^+$, PH_4^+ etc.

2. $Ag^+ + 2 :NH_3 \rightleftharpoons [Ag(:NH_3)_2]^+$

3. $Cu^{+2} + 2 (:H_2O)_2 \longrightarrow [Cu(:H_2O)_4]^{+2}$

(v) Elements which have six electrons in their outermost shell or valence shell.

O–Family (O, S, Se, Te) Po – Radioactive element

$SO_3^{-2} + \ddot{O}: \longrightarrow [SO_3 \longrightarrow O]^{-2}$

Base Acid

Lewis base: Those species which have self tendency to donate the lone pair of electrons are called bases. i.e. a base is an electron pair (lone pair) donor (nucleophile).

Classification of Lewis bases

(i) Those species whose central atom have lone pair of electrons and have self tendency to donate them –

Ex. $\ddot{N}H_3$, $R-\ddot{N}H_2$, $R_2-\ddot{N}H$, $C_6H_5-\ddot{N}H_2$, $C_5H_5-\ddot{N}$, $H_2\ddot{N}-\ddot{N}H_2$,

$H-\ddot{\underset{..}{O}}-H$, $R-\ddot{\underset{..}{O}}-H$, $R-\ddot{\underset{..}{O}}-R$, $R-\ddot{\underset{..}{S}}-R$, $R-\ddot{\underset{..}{S}}-H$, etc.

(ii) **Anions**

Ex. O^{-2}, SO_4^{-2}, CO_3^{-2}, Cl^-, Br^-, F^-, I^-, N^{-3}, P^{-3}, NO_3^-, S^{-2}, CN^-, etc.

Nature of water: According to this concept nature of water is basic i.e. water is a base.

Neutralisation Reaction: According to this concept those reactions in which acid and base react together to form a co-ordinate bond between them are called neutralisation reactions.

Ex.

$F_3B + \ddot{N}H_3 \rightarrow [F_3B \leftarrow NH_3]$

Lewis acid Lewis base

$H^+ + \ddot{N}H_3 \rightarrow [H \leftarrow NH_3]^+ \rightarrow NH_4^+$

$H^+ + H_2\ddot{O} \rightarrow \left[H \leftarrow O\begin{matrix} H \\ H \end{matrix}\right]^+ \rightarrow H_3O^+$

$Ag^+ + 2\ddot{N}H_3 \rightarrow [H_3\ddot{N} \rightarrow Ag \leftarrow \ddot{N}H_3]^+ \rightarrow [Ag(NH_3)_2]^+$

Strength of acids and bases: This concept explain the strength of acids and bases depending upon the basis of self tendency, to accept or donate the lone pair of electrons (e^-).

Ex. SO_2 < SO_3 (Strong acid)

Advantage

(a) This concept explains the acid and base without any solvent.

(b) This concept explains those acids and bases which are not explained by other concepts.

Disadvantage

(a) This concept fails to explain the acid and base practically.

(b) According to this concept, formation of co-ordinate bond in neutralisation reaction of acid and base is a slow process, whereas neutralisation reaction of acids and bases is actually fast process.

(c) This concept fails to explain the nature of real acids and bases

Ex. HCl, HNO_3, H_2SO_4, NaOH, KOH, etc.

Special point

All the Lewis bases are Bronsted bases but all the Lewis acids are not Bronsted acids.

CLASSIFICATION OF ACIDS AND BASES

A. Classification of all the acids

(a) Oxy acids

The acids which contain oxygen is called oxy acids.

In oxy acids generally $\rightarrow$ Nonmetal + O – H

$$\text{Nonmetal} — \text{O} \not\!— \text{H}$$

(i) When nonmetals are different then –

Acidic character $\propto$ Δ Electronegativity (E.N.)

H_3BO_3	H_2CO_3	HNO_3	—	—
	H_2SiO_3	HPO_3	H_2SO_4	$HClO_4$
		$HAsO_3$	H_2SeO_4	$HBrO_4$
			H_2TeO_4	HIO_4

Electro negativity $\downarrow$
Acidic character $\downarrow$

Electro negativity $\uparrow$
Acidic character $\uparrow$

- $H_3BO_3 < H_2CO_3 < HNO_3 < H_2SO_4 < HClO_4$
- $HIO_4 < HBrO_4 < HClO_4$
- $HAsO_3 < HPO_3 < HNO_3$

(ii) When non metals are same then : Acidic character $\propto$ oxidation number (O.N.)

- $H\overset{+3}{N}O_2 < H\overset{+5}{N}O_3$
- $H_2\overset{+4}{S}O_3 < H_2\overset{+6}{S}O_4$
- $HO\overset{+1}{Cl} < H\overset{+3}{Cl}O_2 < H\overset{+5}{Cl}O_3 < H\overset{+7}{Cl}O_4$
    - $H_3\overset{+5}{P}O_4 < H_3\overset{+3}{P}O_3 < H_3\overset{+1}{P}O_2$ (exception)

(b) Hydrides or Hydra acids

The acids containing hydrogen is called hydra acids.

		Base $\downarrow$	Neutral $\downarrow$	Acid $\downarrow$
B_2H_6	CH_4	NH_3	H_2O	HF
	SiH_4	PH_3	H_2S	HCl
		AsH_3	H_2Se	HBr
			H_2Te	HI

Size increases
Bond length increases
Bond strength decreases
Acidic character increases

ΔEN $\uparrow$ Polar properties $\uparrow$ Acidic character $\uparrow$

Ex. (i) HF < HCl < HBr < HI

(ii) $NH_3 < PH_3 < AsH_3$

(iii) $CH_4 < NH_3 < H_2O < HF$

(c) Acidic anhydrides

The acids which are formed by loss of one or more water molecules is called Acidic anhydrides.

Ex. $H_2SO_3 \xrightarrow{-H_2O} SO_2$

$H_2SO_4 \xrightarrow{-H_2O} SO_3$

$2HNO_2 \xrightarrow{-H_2O} N_2O_3$

$2HNO_3 \xrightarrow{-H_2O} N_2O_5$

$2HClO_4 \xrightarrow{-H_2O} Cl_2O_7$

$2H_3PO_4 \xrightarrow{-3H_2O} P_2O_5$

(HPO_3)

$2 \times CH_3COOH \xrightarrow{-H_2O} (CH_3CO)_2O$

$$CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-O-H + H-O-\underset{\underset{\displaystyle O}{\|}}{C}-CH_3 \longrightarrow CH_3-\overset{\overset{\displaystyle O}{\|}}{C}-O-\overset{\overset{\displaystyle O}{\|}}{C}-CH_3 \longrightarrow (CH_3CO)_2O$$

Special point

Acidic order of acidic anhydrides is same as the acidic order of their oxy acids.

Ex. $N_2O_3 < N_2O_5$ (Acidic order)

$Cl_2O < Cl_2O_3 < Cl_2O_5 < Cl_2O_7$ (Acidic order)

(d) For Cations :- Acidic character $\propto$ Charge density (+)

$$\propto \frac{1}{\text{Size of cation}}$$

$\propto$ Electronegativity (E.N.)

$\propto$ Oxidation number (O.N.)

Ex. (i) $Li^+ < Ca^{+2} < Al^{+3}$

(ii) $Li^+ > Na^+ > K^+ > Rb^+ > Cs^+$

(iii) $Be^{+2} > Mg^{+2} > Ca^{+2} > Sr^{+2} > Ba^{+2}$

B. Classification of all the bases

(a) Hydroxides

The bases containing hydroxyl ion (OH^-) is called hydroxides.

$M \not{-} O - H$

(Metal)

$$\text{Basic character} \propto \frac{1}{\text{Electronegativity of metal}}$$

LiOH	$Be(OH)_2$		$B(OH)_3$
NaOH	$Mg(OH)_2$		$Al(OH)_3$
KOH	$Ca(OH)_2$	d-block	$Ga(OH)_3$
RbOH	$Sr(OH)_2$		$In(OH)_3$
CsOH	$Ba(OH)_2$		$Tl(OH)_3$

Electronegativity $\downarrow$
Basic strength $\uparrow$

Electronegativity $\uparrow$ Basic strength $\downarrow$

CsOH is the strongest base.

Ex. (i) $LiOH < NaOH < KOH < RbOH < CsOH$

(ii) $Be(OH)_2 < Mg(OH)_2 < Ca(OH)_2 < Sr(OH)_2 < Ba(OH)_2$

(b) Basic anhydrides

The bases which are formed by loss of one or more water molecules is called basic anhydrides.

Ex. $2NaOH \xrightarrow{-H_2O} Na_2O$

$2LiOH \xrightarrow{-H_2O} Li_2O$

$Ba(OH)_2 \xrightarrow{-H_2O} BaO$

$2Al(OH)_3 \xrightarrow{-3H_2O} Al_2O_3$ (Amphoteric oxide)

Special point :- Basic order of basic anhydrides is same as the basic order of their metal hydroxides.

(c) For anions

Basic character $\propto$ Charge density (–)

$$\propto \frac{1}{\text{Size of anion}}$$

$$\propto \frac{1}{\text{Electronegativity}}$$

$$\propto \frac{1}{\text{Oxidation number}}$$

Ex. (i) $F^- < O^{2-} < N^{-3}$

(ii) $F^- > Cl^- > Br^- > I^-$

STRENGTH OF ACIDS AND BASES

- According to Arrhenius concept, an acid is defined as a substance which when dissolved in water gives H^+ ions and a base is defined as a substance which when dissolved in water gives OH^- ions. Thus, greater the number of H^+ ions produced in the aqueous solution, the stronger is the acid. Similarly, greater the number of OH^- ions produced in the aqueous solution, stronger is the base.
- Now, as greater is the dissociation constant of the weak acid (K_a), greater is the amount of H^+ (aq) produced, therefore stronger is the acid. Thus, K_a values give a measure of the relative strengths of the weak acids. Similarly, K_b values give a measure of the relative strengths of the weak bases.
- The relative strengths of two weak acids can, however, be compared in a quantitative manner as follows:

Suppose the weak acid is represented as HA. Suppose the initial concentration of HA is C moles/litre and α is its degree of dissociation. Then we can write

	HA + Water	$\rightleftharpoons$ H^+ (aq)	+ A^- (aq)
Initial conc.	C	0	0
Conc. at eqm.	$C - C\alpha$ or $C(1-\alpha)$	$C\alpha$	$C\alpha$

Applying the law of chemical equilibrium, we get

$$K_a = \frac{[H^+(aq)]\,[A^-(aq)]}{[HA]}$$

where K_a is called the dissociation constant (or ionization constant) of the acid.

$$\therefore \quad K_a = \frac{C\alpha . C\alpha}{C(1-\alpha)} = \frac{\alpha^2 C}{1-\alpha}$$

As the degree of dissociation (α) of a weak acid is very small, we can take $1 - \alpha \simeq 1$.

$$K_a = \alpha^2 C \quad \text{or} \quad \alpha = \sqrt{\frac{K_\alpha}{C}}$$

Thus, if two acids of equimolar concentration are taken (so that C is constant),

$$\frac{\alpha_1}{\alpha_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}}$$

But degree of dissociation is a measure of the strength of an acid. Hence, we can say

$$\boxed{\frac{\text{Strength of the acid } HA_1}{\text{Strength of the acid } HA_2} = \sqrt{\frac{K_{a_1}}{K_{a_2}}}}$$

Thus, the relative strengths of two acids of equimolar concentration can be compared by taking square root of the ratio of their dissociation constants.

Similarly, the relative strengths of two weak bases of equimolar concentrations can be compared by taking the square root of the ratio of their dissociation constants, i.e.,

$$\boxed{\frac{\text{Strength of base } (BOH)_1}{\text{Strength of base } (BOH)_2} = \sqrt{\frac{K_{b_1}}{K_{b_2}}}}$$

where K_{b_1} and K_{b_2} are the dissociation constants of the two bases.

- It must be remembered that since the ionization of an acid or a base increases with dilution, the strength of the acid or the base increases with dilution.
- Note that K_a and K_b are taken as dimensionless quantities because the standard state concentration of all the species involved is taken as 1 mol L^{-1}.

For Polyprotic acids and Polyacidic bases

Some acids like sulphuric acid (H_2SO_4), carbonic acid (H_2CO_3), oxalic acid ($(COOH)_2$), phosphoric acid (H_3PO_4) etc. contain more than one ionizable proton. Such acids are called polybasic or polyprotic acids. They ionize in steps. For example, the dibasic acid (H_2X) may ionize in two steps as follows:

$H_2X\ (aq) \rightleftharpoons H^+\ (aq) + HX^-\ (aq)$

$HX^-\ (aq) \rightleftharpoons H^+\ (aq) + X^{2-}\ (aq)$

Their ionization constants, K_{a_1} and K_{a_2} called the first and the second ionization constant respectively, will be given by the equations

$$K_{a_1} = \frac{[H^+][HX^-]}{[H_2X]} \qquad K_{a_2} = \frac{[H^+][X^{2-}]}{[HX^-]}$$

- The overall equilibrium may be obtained by adding the stepwise dissociation equilibria. Thus, for the above process, the overall equilibrium will be

$H_2X\ (aq) \rightleftharpoons 2H^+\ (aq) + X^{2-}\ (aq) \qquad K_a = \dfrac{[H^+]^2[X^{2-}]}{[H_2X]}$ Evidently, $K_a = K_{a_1} \times K_{a_2}$

- Similarly, a tribasic acid like H_3PO_4 has three ionization constants. The overall ionization constant (K_a) will be the product of the ionization constants of the three steps, i.e.,

$$K_a = K_{a_1} \times K_{a_2} \times K_{a_3}$$

- It may be noted that $K_{a_1} > K_{a_2} > K_{a_3}$, the reason for this lies in the fact that it is more difficult to remove a positively charged proton from a negative ion due to electrostatic forces. Greater the charge on the negative ion, more difficult it becomes to remove a proton.
- Similarly, polyacidic bases also ionize in steps with ionization constants K_{b_1}, K_{b_2}, K_{b_3} etc.
- A strong acid has a weak conjugate base and vice versa.

Thus, the relative strengths of some acids and their conjugate bases may be represented as follows

Strongest $\longrightarrow$ weakest

Acid : $HIO_3 > HNO_2 > HF > HCOOH > HCN > H_2O$

Conjugate base: $IO_3^- < NO_2^- < F^- < HCOO^- < CN^- < OH^-$

Weakest $\longleftarrow$ Strongest

- Conversely, if we know with respect to water the relative strengths of two acids or bases involved in the conjugate pair in the acid-base reaction, we can find out whether the forward reaction will be favoured or the backward reaction will be favoured.

Factors affecting strength of hydra acids

- Though the strength of acids depends upon a number of factors but the main factor on which the dissociation of an acid depends is the strength and polarity of the H - A bond. Weaker the H - A bond, more easily it dissociates to give H^+ ion and hence stronger is the acid.
- Similarly, greater the polarity of the H–A bond, i.e, larger the electronegativity difference between the atoms H and A, more easily the bond breaks and hence greater is the acidity.
- For elements in the same group of periodic table, the bond strength dominates over the polar nature. Thus, down the group, as size of the atoms increases, bond strength decreases and hence acidic strength increases.
- For example, for element of Group 17.

$$HF << HCl << HBr << HI$$
$$\longrightarrow$$

Size of atom increases

Bond strength decreases

Acidic strength increases

- Similarly, for elements of Group 16, acidic strength increases as

$$H_2O < H_2S < H_2Se < H_2Te$$

- For the elements along the same period, however, the polarity of the H- A bond decides the acidic strength. With increase in the electronegativity of A the polarity of the bond increases and hence the acidic strength also increases. For example,

$$CH_4 < NH_3 < H_2O < HF$$
$$\longrightarrow$$

Electronegativity increases polarity increases

Acidic strength increases

DISSOCIATION CONSTANT AND IONIC PRODUCT OF WATER

- Pure water is poor conductor of electricity.

This shows that water is a weak electrolyte, i.e., it is ionized to a very small extent as

$$H_2O \rightleftharpoons H^+ + OH^- \quad \text{or} \quad H_2O + H_2O \longrightarrow H_3O^+ + OH^-$$

- This ionization is called self-ionization of water.

Applying the law of chemical equilibrium to the above equilibrium, we get

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

where **K is the dissociation constant of water.**

- Since water is ionized to a very small extent, this means that out of millions of water molecules, only a few are dissociated into H^+ and OH^- ions. Thus, the concentration of unionized water molecules, i.e., $[H_2O]$ remains almost constant

$$[H_2O] = \text{constant}$$

Equation (i) may be rewritten as

$$[H^+]\ [OH^-] = K[H_2O] \qquad = \text{constant} \times \text{constant}$$

= K_w another constant, called **ionic product of water**

- Ionic product of water may be defined as the product of the molar concentration of H^+ ions and OH^- ions. As H^+ ions in water (or in an aqueous solution) exist as H_3O^+ ions, therefore, ionic product may also be defined as the product of molar concentration of H_3O^+ and OH^- ions,

  i.e., $K_w = [H^+]\,[OH^-]$ or $[H_3O^+]\,[OH^-]$

  It may be noted that ionic product of water is constant only at constant temperature. Its value at 25°C (298 K) is found to be 1.008×10^{-14}. However, for simplicity, the value of K. at 298 K is usually taken as

$$\boxed{K_w = 1.00 \times 10^{-14}}$$

- **Effect of temperature on K_W**

  The ionic product of water (K_W) increases with increase of temperature. This is obviously because of the fact that with increase of temperature, the degree of ionization of water increases. In other words, more of H_2O molecules dissociate into H^+ (or H_3O^+) ions and OH^- ions. Thus the concentration of H (or H_3O^+) ions and OH^- ions increases and hence the ionic product also increases.

$$\boxed{K_W \propto \text{Temperature}}$$

H_3O^+ Ion and OH^- Ion Concentrations in Aqueous Solutions of Acids and Bases

- As already explained above, for pure water at 298 K

$$[H_3O^+] = [OH^-] = 10^{-7}\text{ M}$$

- Now, if some acid is added to the pure water, then

$$[H_3O^+] > 10^{-7}\text{ M}$$

- However, experiments show that the equation $K_W = [H_3O^+]\,[OH^-]$ is still valid. Thus, the $[OH^-]$ decreases and may be calculated as

$$[OH^-] = \frac{K_W}{[H_3O^+]}$$

Again, if some base is added to pure water, then $[OH^-] > 10^{-7}$ M

But the equation $K_W = [H_3O^+]\,[OH^-]$ is still found to hold good. Hence, $[H_3O^+]$ decreases and may be calculated as

$$[H_3O^+] = \frac{K_W}{[OH^-]}$$

Special point

The increase or decrease of the H_3O^+ ion concentration in an aqueous solution of an acid or a base may be explained qualitatively on the basis of **Le Chatelier's principle** (or common ion effect) as follows:

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

If some acid is added to pure water, H_3O^+ ion concentration increases, therefore the equilibrium shifts in the backward direction (or ionization of H_2O is suppressed). Thus, the OH^- ion concentration decreases. Similarly, if a base added, OH^- ion concentration increases. Again, the equilibrium shifts in the backward direction (or the ionization H_2O, is suppressed) and hence the H_3O^+ ion concentration decreases.

pH — A Convenient Representation of H_3O^+ Ion Concentration

Sorensen, in 1909 suggested a convenient method of expressing the H_3O^+ ion concentration in terms of pH. The symbol has been taken from the Danish word **potenz de hydrogen ion** which means power of Hydrogen ions.

pH may be **defined** as negative logarithm of hydronium ion concentration.

It may be written as

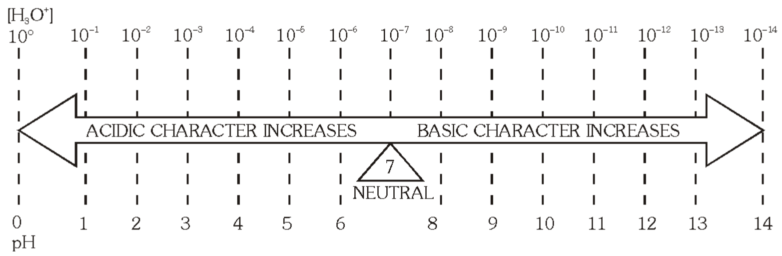
$$pH = -\log\,[H_3O^+] \text{ or } \quad pH = \log\frac{1}{[H_3O^+]} \text{ Hence,}$$

pH may also be defined as logarithm of the reciprocal of H_3O^+ ion concentration.

pH scale

Theoretically, a solution may have any H_3O^+ ion concentration and hence may have any pH value. However, practically, the solutions having H_3O^+ ion or OH^- ion concentration more than 10^0 (= 1 M) and hence OH^- ion or H_3O^+ ion concentration, less than 10^{-14} M are not common. This is because such solutions are not easy to obtain and moreover as the concentration increases, the degree of dissociation of the acid or the base decreases enormously. Hence, the pH range is taken as 0 to 14 for most of the practical purposes.

The acidity or the alkalinity of a solution may be represented on the pH scale as follows :



- **Relationship between pH and pOH**

Just as pH = – log [H^+], similarly, pOH = –log [OH^-]

Since [H_3O^+] [OH^-] = K_W = 10^{-14} at 25° C

$\therefore$ log [H_3O^+] + log [OH^-] = log K_W

= log 10^{-14} = –14 or –log [H_3O^+] –log [OH^-] = –log K_W = 14

or **pH + pOH = pK = 14**

- **Relationship between K_a and K_b, or pK_a and pK_b values** (i.e., between ionization constant of a weak acid and that of its conjugate base).

Consider the weak acid HF. Its conjugate base is F. They dissociate as follows :

$$HF \rightleftharpoons H^+ + F^- \qquad K_a = \frac{[H^+][F^-]}{[HF]} \quad ....(i)$$

$$F^- + H_2O \rightleftharpoons OH^- + HF \qquad K_b = \frac{[OH^-][HF]}{[F^-]} \quad ....(ii)$$

Multiplying eqns, (i) and (ii), we get

$K_a \times K_b = [H^+]\,[OH^-] = K_w$ i.e. $K_a \times K_b = K_w$

$\therefore$ $\log K_a + \log K_b = \log K_w$ or $-\log K_a - \log K_b = -\log K_w$

i.e. $\mathbf{pK_a + pK_b = pK_w}$

Limitations of pH scale

(i) The pH values of the solutions do not give the exact idea of their relative strength. For example,

(a) A solution of pH = 1 has hydrogen ion concentration 100 times than that of a solution of pH=3, and not 3 times.

(b) A 4×10^{-5} N HCl solution is twice concentrated as compared to 2×10^{-5} N HCl solution but the pH values of these solutions are 4.4 and 4.7 respectively, and not double.

(ii) pH = 0 is obtained in 1 N solution of a strong acid and for concentration 2 N, 3 N, 10 N etc. the value can be negative too. Similarly, for a concentrated solution of a strong base (> 1 M), pH can be greater than 14.

(iii) A 10^{-8} M solution of acid cannot have pH = 8. The value will be close to but < 7.

- Greater the value of ionization constant of the acid or the base [K_a or K_b), smaller is the value of pK_a and pK_b and stronger is the acid or base, and vice versa. Thus, larger the value of pK_a or pK_b, weaker is the acid or the base.

Measurement of pH

- Though accurate measurement of pH of a solution is done with the help of an instrument, called pH-meter, approximate pH (within the range of 0.5) can be determined with the help of pH papers which show different colours when dipped in solutions of different pH.

- **pH values of a few common substances are given below:**

Saturated NaOH sol.	14	Black coffee	5.0
0.1 M NaOH sol.	13	Tomato juice	4.2
Lime water, $Ca(OH)_2$	10.5	Vinegar and Soft drinks	3
Milk of magnesia, $Mg(OH)_2$	10	Lemon juice	2.2
Sea water and white of egg	7.8	Gastric juice	1.2
Human blood and Tears	7.4	1 M HCl sol.	0
Milk	6.8		
Human saliva	6.4		

SALTS

Take small amount of Sodium hydroxide solution in a beaker and add an equivalent amount of Hydrochloric acid solution. Do you know what will happen? Sodium hydroxide and hydrochloric acid react with each other to produce Sodium chloride and water as given below.

$$\underset{\text{Sodium hydroxide}}{NaOH} + \underset{\text{Hydrochloric acid}}{HCl} \longrightarrow \underset{\text{Sodium chloride}}{NaCl} + \underset{\text{Water}}{H_2O}$$

This process in which an acid reacts with a base to give salt and water is known as neutralisation reaction. In simple words, it is a process in which H^+ (aq) ion of an acid reacts with OH^- (aq) ion of the base to produce water.

$$\underset{\text{(From acid)}}{H^+(aq)} + \underset{\text{(From base)}}{OH^-(aq)} \longrightarrow \underset{\text{Water}}{H_2O\,(l)}$$

Similarly, if equivalent amounts of Copper hydroxide and Sulphuric acid are mixed, we get Copper sulphate and water as:

$$\underset{\text{Copper hydroxide}}{Cu(OH)_2} + \underset{\text{Sulphuric acid}}{H_2SO_4} \longrightarrow \underset{\text{Copper sulphate}}{CuSO_4} + \underset{\text{Water}}{2H_2O}$$

Sodium chloride formed in the first example and Copper sulphate obtained in the second example dissociates in water giving their constituent ions

$$NaCl + \text{Water} \longrightarrow Na^+(aq) + Cl^-(aq)$$

$$CuSO_4 + \text{Water} \longrightarrow Cu^{2+}(aq) + SO_4^{2-}(aq)$$

Such ionic compounds which when dissolved in water dissociate to yield positive ions other than hydrogen ions (H^+) and negative ions other than hydroxyl ions (OH^-) are called salts.

KCl, NH_4Cl, $ZnCl_2$, $NiCl_2$, K_2SO_4, Na_2SO_4, $NiSO_4$ are some other common examples of such salts. A salt can also be defined in another way.

A salt is a compound formed by partial or complete replacement of ionisable H atoms of an acid by a metallic ion or by an electropositive ion. For example,

$$H_2SO_4 + NaOH \longrightarrow \underset{\text{Sodium hydrogensulphate}}{NaHSO_4} + H_2O$$

(Partial replacement : Only one hydrogen atom is replaced)

$$H_2SO_4 + 2NaOH \longrightarrow \underset{\text{Sodium sulphate}}{Na_2SO_4} + 2H_2O$$

(Complete replacement : Both the hydrogen atoms are replaced)

In this case, both Sodium hydrogensulphate and Sodium sulphate are known as salts.

Let us now take up HCl and add NH_4OH to it.

$$HCl + NH_4OH \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl} + H_2O$$

In this case, hydrogen atom of HCl has been replaced by electropositive NH_4^+ ion and not the metallic ion, therefore, NH_4Cl is also known as a salt.

Classification of Salt (based on their mode of formation)

Salts can be classified into different categories. Let us discuss them one by one.

1. Normal Salts : The salts which are obtained by complete replacement of the ionisable hydrogen atoms of an acid by a metallic or an ammonium ion are known as normal salts. For example, normal salts NaCl and Na_2SO_4 are formed by the complete replacement of ionisable hydrogen atoms of HCl and H_2SO_4 respectively

$$\underset{}{HCl} + NaOH \longrightarrow \underset{\substack{\text{Sodium chloride} \\ \text{(Normal salt)}}}{NaCl} + H_2O$$

$$H_2SO_4 + 2NaOH \longrightarrow \underset{\substack{\text{Sodium sulphate} \\ \text{(Normal salt)}}}{Na_2SO_4} + 2H_2O$$

There is a long list of normal salts some of which are given in Table

Table: Some Examples of Normal Salts with their Parent Acids

S.No.	Parent Acid	Normal Salts
1.	Hydrochloric acid (HCl)	NaCl, KCl, $MgCl_2$, $AlCl_3$, $ZnCl_2$, $CaCl_2$ and NH_4Cl.
2.	Nitric acid (HNO_3)	$NaNO_3$, KNO_3, $Mg(NO_3)_2$,$Al(NO_3)_3$, $Zn(NO_3)_2$, $Ca(NO_3)_2$.
3.	Sulphuric acid (H_2SO_4)	Na_2SO_4, K_2SO_4, $MgSO_4$, $Al_2(SO_4)_3$, $ZnSO_4$, $CaSO_4$.
4.	Acetic acid (CH_3COOH)	CH_3COONa, CH_3COOK, $(CH_3COO)_2$ Ca, $(CH_3COO)_2$ Pb, CH_3COONH_4.
5.	Carbonic acid (H_2CO_3)	Na_2CO_3, K_2CO_3, $MgCO_3$, $ZnCO_3$, $CaCO_3$, $(NH_4)_2CO_3$.
6.	Sulphurous acid (H_2SO_3)	Na_2SO_3, K_2SO_3, $MgSO_3$, $ZnSO_3$, $CaSO_3$, $(NH_4)_2$ SO_3.
7.	Phosphoric acid (H_3PO_4)	Na_3PO_4, K_3PO_4, $Mg_3(PO_4)_2$, $Zn_3(PO_4)_2$, $Ca_3(PO_4)_2$, $(NH_4)_3$ PO_4.

2. Acid salts: The salts which are obtained by the partial replacement of ionisable hydrogen atoms of a polybasic acid by a metal or an ammonium ion are known as acid salts.

These are usually formed when insufficient amount of the base is taken for the neutralisation of the acid. For example, when insufficient amount of NaOH is taken to neutralise H_2SO_4, we get an acid salt $NaHSO_4$.

$$H_2SO_4 + \underset{\text{(Insufficent amount)}}{NaOH} \longrightarrow \underset{\substack{\text{Sodium hydrogensulphate} \\ \text{(Acid salt)}}}{NaHSO_4} + H_2O$$

In this case, only one hydrogen atom out of two has been replaced by sodium atom. Since there is one more hydrogen atom in $NaHSO_4$ which can be replaced, therefore, it further reacts with another molecule of NaOH to produce Na_2SO_4 which is a normal salt.

$$\underset{\substack{\text{Sodium hydrogensulphate} \\ \text{(Acid salt)}}}{NaHSO_4} + NaOH \longrightarrow \underset{\substack{\text{Sodium sulphate} \\ \text{(Normal salt)}}}{Na_2SO_4} + H_2O$$

Acid salts ionise in aqueous solution to produce hydronium ions (H_3O^+), therefore, they exhibit all the properties of acids.

Some other examples of acid salts are given in Table.

Some Acid Salts with Their Parent Acids

S.No.	Parent Acid	Acid Salts
1	Sulphuric acid (H_2SO_4)	$NaHSO_4$, $KHSO_4$, $Ca(HSO_4)_2$
2	Carbonic acid (H_2CO_3)	$NaHCO_3$, $KHCO_3$, $Ca(HCO_3)_2$, $Mg(HCO_3)_2$
3	Sulphurous acid (H_2SO_3)	$NaHSO_3$, $KHSO_3$, $Ca(HSO_3)_2$, $Mg(HSO_3)_2$
4	Phosphoric acid(H_3PO_4)	NaH_2PO_4, Na_2HPO_4, KH_2PO_4, K_2HPO_4, $Ca(H_2PO_4)_2$, $CaHPO_4$

3. Basic Salts : The salts which are formed by partial replacement of hydroxyl (–OH) groups of a di-or a triacidic base by an acid radical are known as basic salts.

These are usually formed when an insufficient amount of acid is taken for the neutralisation of the base. For example, when insufficient amount of HCl is added to Lead hydroxide, Basic lead chloride [Pb(OH)Cl] is formed

$$\underset{\substack{\text{Lead hydroxide}\\ \text{(Diacidic base)}}}{Pb(OH)_2} + HCl \longrightarrow \underset{\substack{\text{Basic Lead chloride}\\ \text{(Basic salt)}}}{Pb(OH)Cl} + H_2O$$

Basic salts, for example, Pb(OH)Cl further reacts with HCl to form normal salts

$$\underset{\substack{\text{Basic Lead Chloride}\\ \text{(Basic salts)}}}{Pb(OH)Cl} + HCl \longrightarrow \underset{\substack{\text{Lead chloride}\\ \text{(Normal salt)}}}{PbCl_2} + H_2O$$

Some other important examples of basic salts are :

(i) Basic copper chloride, Cu(OH)Cl.

(ii) Basic copper nitrate, $Cu(OH)NO_3$

(iii) Basic lead nitrate, Pb $(OH)NO_3$.

4. Double Salts : The salts which are obtained by the crystallisation of two simple salts from a mixture of their saturated solutions are known as double salts.

For example, a double salt **potash alum** [$K_2SO_4 . Al_2(SO_4)_3 . 24H_2O$] is prepared by mixing saturated solutions of two simple salts, K_2SO_4 and $Al_2(SO_4)_3$ and crystallizatiion of the mixture

$$\underset{\substack{\text{Potassium sulphate}\\ \text{(Simple salt)}}}{K_2SO_4} + \underset{\substack{\text{Aluminium sulphate}\\ \text{(Simple salt)}}}{Al_2(SO_4)_3} + 24H_2O \xrightarrow{\text{Crystallisation}} \underset{\substack{\text{Potash alum}\\ \text{(Double salt)}}}{K_2SO_4 . Al_2(SO_4)_3 . 24H_2O}$$

Some other examples of double salts are :

(i) Mohr's Salt, $FeSO_4 .(NH_4)_2SO_4 . 6H_2O$,

(ii) Dolomite, $CaCO_3 . MgCO_3$,

(iii) Carnallite, $KCl. MgCl_2.6H_2O$

5. Mixed Salts: The salts which contain two or more acidic or basic radicals are called mixed salts. For example, Sodium potassium carbonate ($NaKCO_3$) is a mixed salt containing two basic radicals sodium and potassium. Similarly, Bleaching powder, Ca (OCl)Cl is also a mixed salt containing two acid radicals OCl^- and Cl^-.

Some other important examples of mixed salts are :

Sodium potassium sulphate ($NaKSO_4$) (containing two basic radicals), Disodium potassium phosphate (Na_2KPO_4) (containing two basic radicals).

CHEMICALS FROM COMMON SALT

Sodium chloride (Common salt/table salt)

We know that hydrochloric acid and sodium hydroxide combine with each other to form sodium chloride (NaCl) which in common language is also known as common salt. This is the salt which you sprinkle on your salads and use in your kitchens. Common salt is an ionic compound of sodium and chlorine $\mathbf{(Na^+Cl^-)_n}$.

The main source of common salt (sodium chloride) is the sea water. Sea water contains about 3.5% of soluble salts, the most common of which is sodium chloride (2.7 to 2.9%). Saline water of inland lakes, such as Sambhar lake in Rajasthan is also a good source of common salt. Sodium chloride is also found as rock salt. Beds of rock salt were formed when lakes/seas dried up in past.

Common salt act as raw material for making various materials of daily use. Let us discuss some of them.

Sodium hydroxide

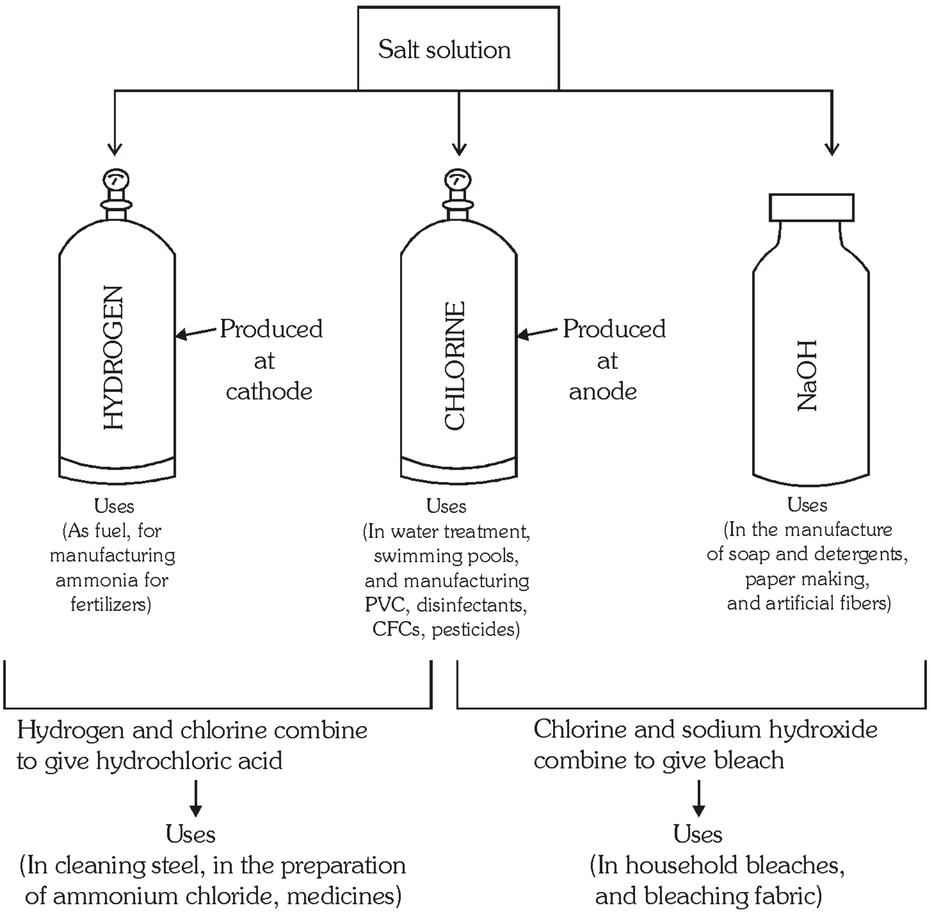
Commercially, sodium hydroxide is also called **caustic soda** because of its corrosive action on animal and vegetable tissues.

Chlor-alkali process for obtaining sodium hydroxide

When we pass electricity through a solution of sodium chloride, commonly called **brine**. It decomposes to form sodium hydroxide according to the following equation:

$$2NaCl(aq) + 2H_2O(l) \xrightarrow[\text{Current}]{\text{Electric}} 2NaOH(aq) + Cl_2(g) + H_2(g)$$

On electrolysis, chlorine gas is formed at anode and hydrogen at cathode, sodium hydroxide solution is formed near the cathode. All these products are commercially important.



Products of chlor-alkali process.

Bleaching powder

Chemically, bleaching powder is generally represented by the formula, $CaOCl_2$ (called, calcium oxychloride).

We know that chlorine is produced during the electrolysis of aqueous sodium chloride **(brine)**. This chlorine gas is used for the manufacture of bleaching powder. Bleaching powder is produced by the action of chlorine on dry slaked lime [$Ca(OH)_2$]. Bleaching powder is represented as $CaOCl_2$, though the actual composition is quite complex.

$$\underset{\text{Slaked lime}}{Ca(OH)_2} + Cl_2 \longrightarrow \underset{\text{bleaching powder}}{CaOCl_2} + H_2O$$

The plant generally used for the manufacture of bleaching powder is known as **'Hasenclever's plant'**.

Uses of bleaching powder

(i) For bleaching cotton and linen in the textile industry, for bleaching wood pulp in paper factories and for bleaching washed clothes in laundry.

(ii) As an oxidising agent in many chemical industries.

(iii) For disinfecting drinking water to make it free of germs.

(iv) In rendering wool unshrinkable.

(v) In the manufacture of chloroform.

(vi) In laundry for bleaching washed clothes.

Baking soda

The chemical name of baking soda is sodium hydrogencarbonate or sodium bicarbonate. Baking soda (or sodium bicarbonate) is represented by the formula $NaHCO_3$. The soda commonly used in the kitchen for making tasty crispy pakoras is baking soda. Sometime it is added for faster cooking. It is produced using sodium chloride as one of the raw materials.

In laboratory, baking soda can be prepared by passing excess of CO_2 gas through the saturated solution of sodium carbonate. As a result, sodium hydrogen carbonate is formed. Being sparingly soluble in water, it gets precipitated. The precipitate is separated, washed & dried without heating.

$$Na_2CO_3 + CO_2 + H_2O \longrightarrow 2NaHCO_3 \text{ (ppt.)}$$

On commercial scale, when washing soda is prepared by **Solvay process** or **ammonia-soda process** $NaHCO_3$ is formed as an intermediate product. In this method a solution of sodium chloride (called "brine") saturated with ammonia is allowed to react with CO_2, sodium hydrogen carbonate & ammonium chloride are formed.

$$\underset{\text{Common salt}}{NaCl(aq)} + H_2O(l) + CO_2(g) + NH_3(g) \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl(aq)} + \underset{\text{Sodium hydrogen carbonate}}{NaHCO_3(s)}$$

It can be used to neutralise an acid because it is mild non-corrosive base due to the presence of HCO_3^- ion.

The following reaction takes place when it is heated during cooking.

$$2NaHCO_3(s) \xrightarrow{\text{Heat}} Na_2CO_3(s) + H_2O(g) + CO_2(g)$$

Uses of sodium hydrogen carbonate ($NaHCO_3$)

(i) For making **baking powder** which is a mixture of baking soda (sodium hydrogen carbonate) and a mild edible acid like tartaric acid. When baking powder is mixed with water, the following reaction takes place.

$$NaHCO_3 + \underset{\text{From acid}}{H^+} \longrightarrow CO_2 + H_2O + \text{Sodium salt of the acid}$$

Carbon dioxide so produced during the reaction is responsible for making the bread and cake to rise making them soft and spongy.

(ii) As an ingredient in antacids. Being alkaline, it neutralises excess acid in the stomach and provides relief.

(iii) It is used in soda-acid fire extinguisher.

Washing soda (sodium carbonate)

The chemical formula of washing soda is $Na_2CO_3 \cdot 10H_2O$, (sodium carbonate decahydrate). Washing soda is manufactured by **Solvay process**. This process is also known as **ammonia soda process**. The raw materials needed for the process are sodium chloride, lime stone ($CaCO_3$) and ammonia (NH_3). The reactions involved are

Step-I

$$\underset{\text{Common salt}}{NaCl(aq)} + H_2O(l) + CO_2(g) + NH_3(g) \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl(aq)} + \underset{\text{Sodium hydrogen carbonate}}{NaHCO_3(s)}$$

The CO_2 required in this reaction is obtained by heating limestone.

$$\underset{\text{Limestone}}{CaCO_3} \xrightarrow{\text{Heat}} \underset{\text{Quicklime}}{CaO} + CO_2(g)$$

Step-II

Dry sodium hydrogen carbonate is heated strongly to produce sodium carbonate.

$$2NaHCO_3(s) \xrightarrow{\text{Heat}} \underset{\text{Anhydrous sodium carbonate (soda ash)}}{Na_2CO_3\ (s)} + H_2O(l) + CO_2(g)$$

Step-III

Washing soda ($Na_2CO_3 \cdot 10H_2O$) is obtained by recrystallisation from a saturated solution of soda ash (Na_2CO_3).

$$Na_2CO_3(s) + H_2O(l) \longrightarrow Na_2CO_3\ (aq) \xrightarrow{\text{crystallisation}} \underset{\text{Washing soda}}{Na_2CO_3 \cdot 10H_2O}$$

Uses of washing soda

(i) Washing soda (or sodium carbonate) is used for washing clothes (laundry purposes).

(ii) Washing soda is used for softening hard water.

(iii) Sodium carbonate (soda ash) is used for the manufacture of detergents.

(iv) Sodium carbonate is used for the manufacture of many important compounds, such as borax ($Na_2B_4O_7$), hypo ($Na_2S_2O_3 \cdot 5H_2O$), etc.

(v) Sodium carbonate is also used in paper, glass, soap and paint industries.

Plaster of Paris ($CaSO_4 . \frac{1}{2}H_2O$)

Plaster of Paris is hemihydrate (hemi means half and hydrate means water) of calcium sulphate. Its molecular formula is $CaSO_4 . \frac{1}{2}H_2O$ or $(CaSO_4)_2.H_2O$.

In Plaster of Paris one molecule of water is shared by two formula units of $CaSO_4$.

1. Preparation of Plaster of Paris

Plaster of Paris is obtained by heating gypsum ($CaSO_4.2H_2O$) in a kiln at 373K (or 100°C).

$$\underset{\text{Gypsum}}{2[CaSO_4.2H_2O](s)} \xrightarrow[\text{heat}]{373K,(100°C)} \underset{\text{Plaster of Paris}}{(CaSO_4)_2 \cdot H_2O(s)} + 3H_2O(g)$$

OR

$$\underset{\text{Gypsum}}{CaSO_4.2H_2O(s)} \xrightarrow[\text{heat}]{373K,(100°C)} \underset{\text{Plaster of Paris}}{CaSO_4 \cdot \frac{1}{2}H_2O(s)} + \frac{3}{2}H_2O(g)$$

During the preparation of Plaster of Paris, temperature should be controlled carefully. Otherwise, anhydrous calcium sulphate ($CaSO_4$) will be formed. Anhydrous calcium sulphate does not set into hard mass when mixed with water. So, if temperature is not controlled carefully, the Plaster of Paris obtained will have poor setting property.

$$CaSO_4 \cdot 2H_2O(s) \xrightarrow[\text{373 K}]{\text{More than}} \underset{\text{Dead burnt plaster}}{CaSO_4(s)} + 2H_2O(g)$$

2. Properties of **Plaster of Paris**

(i) Plaster of Paris is a white, odourless powder.

(ii) At ordinary room temperature, Plaster of Paris absorbs water and a large amount of heat is liberated.

(iii) When mixed with a limited amount of water (50% by mass), it forms a plastic mass, evolves heat and quickly sets to a hard porous mass within minutes. This is called the **setting process**.

During setting, a slight expansion in volume occurs. It is due to this that it fills the mould completely and gives sharp impression. The reaction during process is

$$\underset{\text{Plaster of Paris}}{CaSO_4 \cdot \frac{1}{2}H_2O(s)} + \underset{\text{Water}}{\frac{3}{2}H_2O(\ell)} \rightarrow \underset{\text{Gypsum (Hard mass)}}{CaSO_4 \cdot 2H_2O(s)}$$

Uses of Plaster of Paris

(i) Plaster of Paris is used in making casts and patterns for moulds and statues.

(ii) Plaster of Paris is used as cement in ornamental casting and for making decorative materials.

(iii) Plaster of Paris is used as a fire proofing material and for making chalks.

(iv) Plaster of Paris is used in hospitals for immobilising the affected part in case of bone fracture or strain.

(v) Plaster of Paris (POP) is used to fill small gaps on walls & roofs.

ARE THE CRYSTALS OF SALTS REALLY DRY?

Crystals of some salts contain certain amount of associated water.

The water associated with the crystal (or molecule) of any salt is called water of crystallisation.

The salt containing water of crystallisation are called **hydrated salts**.

Water of crystallisation

It is the fixed number of water molecules present in one formula unit of a crystalline salt, e.g..,

Blue vitriol	$CuSO_4 \cdot 5H_2O$	Green vitriol	$FeSO_4 \cdot 7H_2O$
Glauber's salt	$Na_2SO_4 \cdot 10H_2O$	White vitriol	$ZnSO_4 \cdot 7H_2O$
Gypsum	$CaSO_4 \cdot 2H_2O$	Epsom salt	$MgSO_4 \cdot 7H_2O$

Note : 1. Water of crystallisation in a hydrated salt is a part of crystal structure of salt and no free water molecules exist in salt.

2. Hydrated salt on dissolution in water shows exothermic dissolution.

During the crystallisation of the salt from its aqueous solution, **water of crystallisation** gets associated with the molecules of a salt. The water of crystallisation associated with the molecules of the salt can easily be removed by heating the crystals of the salt to a definite temperature.

$$\underset{\text{washing soda}}{Na_2CO_3.10H_2O} \xrightarrow{\text{below 373k}} \underset{\text{sodium carbonate monohydrate}}{Na_2CO_3.H_2O} + 9H_2O$$

$$Na_2CO_3.H_2O \xrightarrow{\text{above 373k}} \underset{\text{soda ash}}{Na_2CO_3} + H_2O$$

The presence of water of crystallisation in a hydrated salt like hydrated copper (II) sulphate ($CuSO_4.5H_2O$) can be demonstrated by the following experiment.

ACTIVE CHEMISTRY

Copper sulphate crystals (blue in colour) on strongly heating in a dry boiling tube over the flame of a burner for some time, loses water of crystallization and turns to a white powdery substance anhydrous copper sulphate. The tiny water droplets appear on the inner walls of the boiling tube in figure due to liberation of water molecules on heating which were associated as water of crystallisation in hydrated copper sulphate.

$$\underset{\text{Hydrated copper sulphate (Blue)}}{CuSO_4.5H_2O(s)} \xrightarrow{\text{Heat}} \underset{\text{Anhydrous copper sulphate (white)}}{CuSO_4(s)} + \underset{\text{(water)}}{5H_2O(\ell)}$$

On cooling the boiling tube the white powder is taken out on a glass. Now on adding few drops of water on this white powder, the original blue colour of hydrated copper sulphate is restored due to the conversion of anhydrous copper sulphate to hydrated copper sulphate.

$$\underset{\text{Anhydrous copper sulphate (white)}}{CuSO_4(s)} + \underset{\text{Water}}{5H_2O(\ell)} \longrightarrow \underset{\text{Hydrated copper sulphate (Blue)}}{CuSO_4.5H_2O(s)}$$

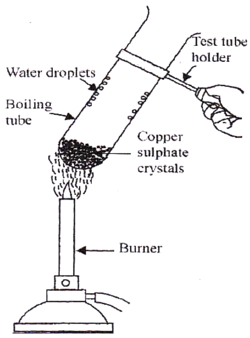


Figure : Action of heat on hydrated copper sulphate crystals

Note: 1. Anhydrous copper sulphate is white. Hydrated copper sulphate, $CuSO_4.5H_2O$ is blue. The property can be used to identify whether pure liquid is water or not.

2. Hydrated copper sulphate is $[Cu(H_2O)_4]SO_4.H_2O$ and contains only one water molecules as hydrated water.

The phenomenon in which a crystalline hydrated salt say ($Na_2CO_3 \cdot 10H_2O$)., on exposure to the atmosphere, loses partly or completely ($Na_2CO_3.H_2O$ or Na_2CO_3), its water of crystallisation, is called **efflorescence** and such a salt is referred to as an **efflorescent salt**.

Efflorescence involves the loss of crystalline nature of the salt, and therefore a crystalline salt crumbles into an amorphous powder.

For example, sodium sulphate decahydrate is an efflorescent salt. When exposed to atmosphere, it slowly loses its water of crystallisation. As a result, it loses the crystallline form and changes to a white powder.

$$Na_2SO_4.7H_2O \xrightarrow[\Delta]{\text{Dry air } 7H_2O} \underset{\text{(white powder)}}{Na_2SO_4} + 7H_2O$$

1. $\underset{\text{Sodium carbonate decahydrate}}{Na_2CO_3.10H_2O}$

$$\underset{\text{(decahydrate)}}{Na_2CO_3.10H_2O} \xrightarrow{\text{Dry air}} \underset{\text{(monohydrate)}}{Na_2CO_3.H_2O} + 9H_2O$$

2. $\underset{\text{Sodium sulphate decahydrate}}{Na_2SO_4.10H_2O}$

$$\underset{\text{(decahydrate) (crystalline)}}{Na_2SO_4.10H_2O} \longrightarrow \underset{\text{Anhydrous sod. sulphate (powder)}}{Na_2SO_4} + 10H_2O$$

3. $\underset{\text{Magnesium sulphate heptahdrate}}{MgSO_4.7H_2O}$

$$\underset{\text{(heptahydrate)}}{MgSO_4.7H_2O} \xrightarrow{\text{Dry air}} \underset{\text{(Monohydrate)}}{MgSO_4.H_2O} + 6H_2O$$

Note: Efflorescent nature of salts noticed when vapour pressure of H_2O is higher than vapour pressure H_2O in air.

(a) Crystalline substances which absorb more amount of moisture from the atmosphere and pass in saturated solution state, (e.g . sodium hydroxide, magnesium sulphate, calcium chloride etc: are called **deliquescent substances** and the phenomenon is known as **deliquesence**.

(b) Substances which readily absorb moisture from the atomosphere. e.g., quick lime (CaO) are called **hygroscopic substances** and the phenonenon is called **hygroscopy**.

(c) Both hygroscopic and deliquescent substances absorbs moisture because vapour pressure of H_2O in atmosphere is more than vapour pressure of H_2O in them.

Deliquescent substances :

(i) Anhydrous magnesium chloride ($MgCl_2$) (ii) Anhydrous zinc chloride ($ZnCl_2$)

(iii) Anhydrous ferric chloride ($FeCl_3$)

Hygroscopic substances :

(i) Conc. Sulphuric Acid (H_2SO_4) (ii) Phosphorus pentoxide (P_2O_5)

(iii) Calcimn oxide or Quick lime (CaO) (iv) Silica gel.

Note: Hygroscopic substances are used as drying agents for drying gases

SALT HYDROLYSIS

- Salt hydrolysis is defined as the process in which a salt reacts with water to give back the acid and the base.

$$\text{Salt + Water} \longrightarrow \text{Acid + Base}$$

or $$\underset{\text{Salt}}{BA} + \underset{\text{Water}}{H_2O} \longrightarrow \underset{\text{Acid}}{HA} + \underset{\text{Base}}{BOH}$$

- Thus, salt hydrolysis may be considered as reverse of neutralization.
- All salts are strong electrolytes and thus ionize completely in the aqueous solution. If the acid (HA) produced is strong and the base (BOH) produced is weak, we can write the above equation as,

$$B^+ + A^- + H_2O \longrightarrow H^+ + A^- + BOH$$

or $$B^+ + H_2O \longrightarrow H^+ + BOH$$

- Thus, in this case the cation reacts with water to give an acidic solution. This is called **cationic hydrolysis**.

Again, if the acid produced is weak and the base produced is strong, we can write

$$B^+ + A^- + H_2O \longrightarrow HA + B^+ + OH^-$$

or $$A^- + H_2O \longrightarrow HA + OH^-$$

- Here, the anion reacts with water to give the basic solution. This is called **anionic hydrolysis**.

Hence,

Salt hydrolysis may be defined as the reaction of the cation or the anion of the salt with water to produce acidic or basic solution.

Thus, depending upon the relative strengths of the acid and the base produced, the resulting solution is acidic, basic or neutral. On this basis, the salts are divided into four categories:

I. Salts of strong acids and strong bases

Examples are:

NaCl, $NaNO_3$, Na_2SO_4, KCl, KNO_3, K_2SO,,

As an illustration, let us discuss the hydrolysis of NaCl. We may write :

$$NaCl + H_2O \longrightarrow NaOH + HCl$$

or $$Na^+ + Cl^- + H_2O \longrightarrow Na^+ + OH^- + H^+ + Cl^-$$

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

Thus, it involves only ionization and no hydrolysis. Further, in the resulting solution, $[H^+] = [OH^-]$, So the solution is neutral. Hence, it can be generalized that the salts of strong acids and strong bases do not undergo hydrolysis and the resulting solution is neutral.

II. Salts of weak acids and strong bases

Examples are:

CH_3COONa, Na_2CO_3, K_2CO_3, Na_3PO_4 etc.

As an illustration, the hydrolysis of sodium acetate (CH_3COONa) may be represented as follows :

$$CH_3COONa + H_2O \rightleftharpoons CH_3COOH + NaOH$$

or $\quad CH_2COO'' + Na^+ + H_2O \rightleftharpoons CH_3COOH + Na^+ + OH^-$

or $\quad CH_3COO^- + H_2O \longrightarrow CH_3COOH + OH^-$

As it produces OH^- ions, the solution of such a salt is alkaline in nature.

III. Salts of strong acids and weak bases

Examples are:

NH_4Cl, $CuSO_4$, NH_4NO_3, $AlCl_3$, CaCl etc.

As an illustration, the hydrolysis of NH_4Cl may be represented as follows:

$$NH_4Cl + H_2O \rightleftharpoons NH_4OH + HCl$$

or $\quad NH_4^+ + Cl^- + H_2O \rightleftharpoons NH_4OH + H^+ + Cl^-$

or $\quad NH_4^+ + H_2O \longrightarrow NH_4OH + H^+$

As it produces H^+ ions, the solution of such a salt is acidic in character.

IV. Salts of weak acids and weak bases

Examples are :

CH_3COONH_4, $(NH_4)_2CO_3$, $AlPO_4$ etc.

As an illustration, the hydrolysis of ammonium acetate maybe represented as follows :

$$CH_3COONH_4 + H_2O \rightleftharpoons CH_3COOH + NH_4OH$$

or $\quad CH_3COO^- + NH_4^+ + H_2O \longrightarrow CH_3COOH + NH_4OH$

Thus, it involves both anionic and cationic hydrolysis.

The resulting solution may be neutral or slightly acidic or basic depending upon the relative degrees of ionization of the weak acid and the weak base produced. In the present example, the acid (CH_3COOH) and the base (NH_4OH) formed are almost equally weak. Hence, the resulting solution is almost neutral.

Hydrolysis constant

The general equation for the hydrolysis of a salt (BA) may be written as

$$BA + H_2O \rightleftharpoons HA + BOH$$

Applying the law of chemical equilibrium, we get

$$\frac{[HA][BOH]}{[BA][H_2O]} = K, \text{ the equilibrium constant.}$$

Since water is present in excess in the aqueous solution, its concentration $[H_2O]$ may be regarded as constant, so that we have

$$K_h = \frac{[HA][BOH]}{[BA]}$$

where K_h is called the hydrolysis constant.

Degree of Hydrolysis

The degree of hydrolysis of a salt is defined as the fraction (or percentage) of the total salt which is hydrolysed.

i.e. $\quad h = \dfrac{\text{No. of moles of the salt hydrolysed}}{\text{Total no. of moles of the salt taken}}$

BUFFER SOLUTIONS AND BUFFER ACTION

- A buffer solution is defined as a solution which resists any change in its pH value (i.e., whose pH remains practically constant) even when small amounts of the acid or the base are added to it.

Types of Buffer solutions

There are two types of buffer solutions. These are :

(i) Solutions of single substances. The solution of the salt of a weak acid and a weak base, e.g. ammonium acetate (CH_3COONH_4) acts as a buffer.

(ii) Solutions of Mixtures. These are further of two types:

(a) Acidic Buffer. It is the solution of a mixture of a weak acid and a salt of this weak acid with a strong base (e.g., $CH_3COOH + CH_3COONa$).

(b) Basic Buffer. It is the solution of a mixture of a weak base and a salt of this weak base with a strong acid. (e.g., $NH_4OH + NH_4Cl$).

Buffer Action

- The property of a buffer solution to resist any change in its pH value even when small amounts of the acid or the base are added to it called 'Buffer action'.

Buffer Action of Acidic Buffer

- For example, let us consider the buffer action of the acidic buffer containing CH_3COOH and CH_3COONa. Acetic acid dissociates to a small extent whereas sodium acetate is almost completely dissociated in the aqueous solution as follows:

$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$ ....(i)

$CH_3COONa \longrightarrow CH_3COO^- + Na^+$ ....(ii)

- By common ion effect', the ionization of CH_3COOH is further suppressed. Thus, in the solution, there are excess of acetate (CH_3COO^-) ions and a small amount of H_3O^+ ions.
- When a few drops of an acid are added to the above mixture solution, the H_3O^+ ions given by the acid combine with the CH_3COO^- ions to form weakly ionized molecules of CH_3COOH.

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

Thus, the H_3O^+ ion concentration and hence the pH of the solution remains almost constant.

- Similarly, when a few drops of a base are added, the OH^- ions given by the base combine with the H_3O^+ ions already present to form weakly ionized molecules of H_2O.

$H_3O^+ \quad + \quad OH^- \quad \longrightarrow \quad 2H_2O$

[From (i)] [Given by the base]

- As the H_3O^+ ions are consumed, the equilibrium (i) shifts towards right (according to Le Chatelier's principle). Thus, more of CH_3COOH dissociates to make up the loss of H_3O^+ ions.
- Hence, the H_3O^+ ion concentration or the pH of the solution does not change.

Buffer Action of Basic Buffer

- The buffer action of a basic buffer, e.g. $NH_4OH + NH_4Cl$ may be explained as follows :

NH_4OH dissociates to a small extent whereas NH_4Cl dissociates completely in the aqueous solution as follows:

$NH_4OH \rightleftharpoons NH_4^+ + OH^-$ ......(i)

$NH_4Cl \longrightarrow NH_4^+ + Cl^-$ ....(ii)

- By common ion effect, the ionization of NH_4OH is further suppressed. Thus, in the solution, there are excess of NH_4^+ ions and a small amount of OH^- ions.
- When a few drops of a base are added, the OH^- ions given by it immediately combine with NH_4^+ ions to form the weakly ionized NH_4OH.

$NH_4^+ + OH^- \longrightarrow NH_4OH$

Thus, the H_3O^+ ion concentration or the pH of the solution remains unaffected.

- Similarly, when a small amount of an acid is added, the H_3O^+ ions given by it combine with the OH^- ions already produced by NH_4OH in equilibrium (i).

$H_3O^+ \quad + \quad OH^- \longrightarrow 2H_2O$

[Given by acids] [Given by (i)]

- As the OH^- ions are consumed, the equilibrium (i) shifts in the forward direction. Thus, more of NH_4OH dissociates to produce more of OH^- ions which makes up the loss of OH^- ions. Hence, the OH^- ion concentration and therefore the H_3O^+ ion concentration or the pH of the solution remains fairly constant.

Importance of Buffer Solutions

(1) **In biological processes**

The pH of our blood is maintained constant (at about 7.4) inspite of various acid and base-producing reactions going on in our body, In the absence of its buffer nature, we could not eat a variety of foods and spices.

The buffer action is due to the presence of carbonic acid (H_2CO_3), bicarbonate ion (HCO_3^-) and carbon dioxide (CO_2) in the blood.

(2) **In Industrial processes**

The use of buffers is an important part of many industrial processes, e.g.,

(i) in electroplating;

(ii) in the manufacture of leather, photographic materials and dyes:

(iii) in the preservation of food articles.

(3) **In analytical chemistry**

(i) in the removal of acid radicals such as phosphate, oxalate and borate which interfere in the precipitation of radicals of Group III (Al^{3+}, Fe^{3+} and Cr^{3+}) of salt analysis;

(ii) in complexometric titrations, e.g., metal ions with EDTA (Ethylene Diamine Tetra Acetate).

(iii) to calibrate the pH meters;

(4) **In bacteriological research**, culture media are generally buffered to maintain the pH required for the growth of the bacteria being studied.

CONDUCTIVITY OF SUBSTANCES

According to conductivity substances are of two types:

(1) Non-Conductor - Those substances which do not show the flow of current or electricity.

Ex. Non - metals, plastic, rubber, wood etc.

Exception – Graphite is a non-metal but show conductivity due to motion of free electrons.

(2) Conductors – Those substances which show conductivity or flow of current are called conductors and these are of two types :

(a) Metallic conductor – Those conductor which show conductivity due to motion of free electrons.

Ex. All metals, Graphite etc.

(b) Ionic conductors / Electrolytes – Those conductor which show conductivity due to movement of free ions.

Ions are in free state in the solutions of ionic compounds. On passing electric current, through the solution, ions move towards oppositely charged electrodes, i.e., the cation moves towards cathode (negative electrode) and the anion moves towards anode (positive electrode). Due to this reason, they are called cations and anions respectively. The current flows through the solution due to the movement of the ion.

Ex. HCl in water, NaCl in water etc.

SOLUBILITY

Definition

(i) At constant temperature, the maximum number of moles of solute which can be dissolved in a solvent to obtain 1 litre of solution (i.e. saturated solution) is called solubility.

$$S = \frac{\text{Number of moles of solute}}{\text{Volume of solution (L)}} \quad \text{or} \quad S = \frac{x}{M_w \times V_L} \text{ mol L}^{-1}$$

(ii) The maximum weight of solute which can be dissolved in a given amount of solvent is known as solubility.

Special points

- Solubility does not depend on amount of substances and volume of solution but it depends on the following–

  (i) Temperature (ii) Presence of common ion (iii) Nature of solvent

- Solubility can be expressed in terms of molarity.

Solubility Product (K_{sp})

• At constant temperature product of concentrations of ions in a saturated solution of substance is called solubility product of that substance. (Saturated solution is that solution in which further dissolution of even a small amount of salt is not possible).

Ex. (i) $AgCl \rightleftharpoons Ag^+ + Cl^-$

According to Law of mass action.

$$K = \frac{[Ag^+][Cl^-]}{[AgCl]}$$

Solubility product at saturation in terms of concentration of ions $K_{sp} = [Ag^+]\,[Cl^-]$

Solubility product in terms of solubility $K_{sp} = (S)\,(S)$

$$\boxed{K_{sp} = S^2}$$

Ex. (ii) Ksp for $CaCl_2$

$$CaCl_2 \rightleftharpoons Ca^{+2} + 2Cl^-$$

According to Law of mass action.

$$K = \frac{[Ca^{2+}][Cl^-]^2}{[CaCl_2]}$$

Solubility product in terms of concentration of ions $Ksp = [Ca^{+2}]\,[Cl^-]^2$

In terms of solubility $K_{sp} = (S)\,(2S)^2$

$$\boxed{K_{sp} = 4S^3}$$

Difference between solubility product and ionic product:

Both ionic product and solubility product represent the product of the concentrations of the ions in the solution, each raised to the power equal to the number of ions as represented by the dissociation of one molecule of the substance. However, they differ in the following two aspects:

(i) The term ionic product has a broad meaning since it is applicable to all types of solutions, may be unsaturated or saturated. On the other hand, the solubility product has restricted meaning since it applies only to a saturated solution in which there exists dynamic equilibrium between in which there exists a dynamic equilibrium between the undissolved salt and the ions present in solution. Thus, the solubility product is, in ionic product for a saturated solution.

(ii) The solubility product of a salt is constant at constant temperature whereas ionic product depends upon the concentration of ions in the solution.

ACIDS, BASES & SALTS

EXERCISE-I

INTRODUCTION

1. An acidic solution could have a pH of

(1) 7 (2) 10 (3) 3 (4) 14

2. Which is the correct set of acid properties :

(1) Sour taste, corrosive, change litmus from red to blue

(2) Sour taste, slippery, change litmus from blue to red

(3) Sweet taste, slippery, change litmus from blue to red

(4) Sour taste, corrosive, change litmus from blue to red

3. Vinegar, fruit juice, and cola are examples of:

(1) Strong acids (2) Weak acids (3) Strong bases (4) Weak bases

4. Which of the following is a base?

(1) Tomatoes juice (2) Salt (3) Caustic soda (4) Vinegar

5. Generally how does a base taste?

(1) Sour (2) Bitter (3) Sweet (4) Salty

6. All the sour tasting substances contain :

(1) Bases (2) Acids (3) Salts (4) Anyone of these

7. Which of the following is a mineral acid?

(1) Hydrochloric acid (2) Nitric acid (3) Sulphuric acid (4) All of these

8. The acid present in vinegar is

(1) Tartaric acid (2) Acetic acid (3) Sulphuric acid (4) Citric acid

9. The sour taste of lemon is due to the presence of

(1) Tartaric acid (2) Acetic acid (3) Sulphuric acid (4) Citric acid

10. The fizzy drinks you drink contain

(1) Hydrochloric Acid (2) Nitric acid (3) Sulphuric Acid (4) Carbonic acid

11. The acid present in the car batteries is

(1) Hydrochloric acid (2) Sulphuric Acid (3) Phosphoric acid (4) Ethanoic Acid

12. Which of the following oxide does not dissolve in water like oxide of silver?

(1) Sodium oxide (2) Potassium oxide (3) Iron oxide (4) Calcium oxide

13. The chemical name for lime water is

(1) Hydrogen peroxide (2) Milk of magnesia (3) Calcium hydroxide (4) Ammonia solution

14. The pH of blood is :–

(1) Less than 6 (2) Greater than 7 and less than 8

(3) Greater than 8 and less than 9 (4) Greater than 10

15. The pH of gastric juice is normally :–

(1) Less than 1.2

(2) Less than 0

(3) Greater than 1 and less than 3

(4) Less than 1 and greater then zero

THEORIES OF ACIDS AND BASES

16. Which substance can be classified as an Arrhenius acid?
(1) HCl (2) NaCl (3) LiOH (4) KOH

17. There are alternate acid base theories that define an acid as any species that can
(1) donate a proton (2) donate an electron (3) accept a proton (4) accept an electron

18. Which of the following is a soluble base?
(1) Aluminium oxide (2) Copper(II) oxide (3) Sodium hydroxide (4) Iron(II) hydroxide

19. Which of the following cannot act both as Bronsted acid and Bronsted base?
(1) Na_2CO_3 (2) OH^- (3) HCO_3^- (4) NH_3

20. The conjugate acid of NH_2^- is :
(1) NH_4^+ (2) NH_4OH (3) NH_3 (4) N_2H_4

21. A substance which can act both an acid and a base is :
(1) Allotropic (2) Amphoteric (3) Isotopic (4) Amorphous

22. Basicity of H_3PO_3 and H_3PO_2 are respectively :-
(1) 1 and 2 (2) 2 and 3 (3) 3 and 2 (4) 2 and 1

23. Which of the following is not a Lewis base?
(1) Ag^+ (2) NH_3 (3) NO_3^- (4) Br^-

24. Which of the following is not a base according to any of the theories ?
(1) $Mg(OH)_2$ (2) NH_3 (3) $H_2PO_4^-$ (4) BF_3

25. Which of the following can act both as a Bronsted acid and a Bronsted base?
(1) O^{2-} (2) HCl (3) HSO_4^- (4) Na_2CO_3

pH SCALE

26. What is the pH of a 0.00001 molar HCl solution?
(1) 1 (2) 9 (3) 5 (4) 4

27. What is the pH of a solution with a hydronium ion concentration of 0.01 mole per litre?
(1) 1 (2) 2 (3) 10 (4) 14

28. When HCl(aq) is exactly neutralized by NaOH(aq), the hydrogen ion concentration in the resulting mixture is
(1) Always less than the concentration of the hydroxide ions
(2) Always greater than the concentration of the hydroxide ions
(3) Always equal than the concentration of the hydroxide ions
(4) Sometimes greater and sometimes less than the concentration of the hydroxide ions

29. Lake water has a pH value of (3) Which substance could be added in to neutralise the pH value of the acidic water present without leaving an alkaline solution.
(1) Aqueous sodium hydroxide (2) Calcium carbonate
(3) Aqueous ammonia (4) Acetic acid

30. Which one of the following burns in air to form an oxide which, when shaken with water, gives a solution with a pH greater than 7?
(1) Carbon (2) Copper (3) Nitrogen (4) Calcium

31. The pH of a solution is defined as the
(1) The positive logarithm (base 10) of hydrogen ions concentration in moles per litres
(2) The negative logarithm (base 10) of hydronium ions concentration in moles per litre
(3) The negative logarithm of (base 5) of hydroxyl ions concentration in moles per litre
(4) None of the above

32. The pH value of an acid is
(1) Less than seven (2) Equal to fifteen (3) Equal to seven (4) In between seven and ten

33. For pure water, the value of $[H_3O^+]$ is
(1) 10^{-9} (2) 10^{-8} (3) 10^{-7} (4) 10^{-6}

34. A 10^{-6} M HCl solution is diluted to 100 times. The pH of the diluted solution would be
(1) Between 6 to 7 (2) Between 7 to 8 (3) Equal to 7 (4) Equal to 10

35. The pH of a 10^{-8} M solution of HCl in water is
(1) 8 (2) –8 (3) between 7 and 8 (4) between 6 and 7

36. What will be the H^+ ion concentration, when 4 g NaOH dissolved in 1000 mL of water?
(1) 10^{-1} (2) 10^{-13} (3) 10^{-4} (4) 10^{-10}

37. Ionic product of water at 25°C is :
(1) 10^{-14} (2) 10^{-7} (3) 10^{-10} (4) 2×10^{-7}

38. At 25^0 C, OH^- ion concentration in 10^{-5} M HCl(aq.) will be :-
(1) Zero (2) 10^{-9} M (3) 10^{-5}M (4) 10^{-2}M

39. HCl............. in water
(1) Ionizes (2) Decompose (3) Dissociates (4) Associates

40. The units of ionic product of water (K_W) are:-
(1) mol^{-1} L^{-1} (2) mol^{2-} L^{-2} (3) mol^{2-} L^{-1} (4) mol^2 L^{-2}

41. Choose the correct relation:-
(1) $\frac{pH + pOH}{14} = 7$ (2) pH + pOH = 14 (3) pOH = 14 + pH (4) pH = 14 + pOH

42. 0.001 N KOH solution has the pH :-
(1) 10^{-1} (2) 3 (3) 11 (4) 2

43. 10^{-6} M HCl is diluted to 100 times. Its pH is :-
(1) 6.0 (2) 8.0 (3) 6.95 (4) 9.5

44. The pH of solution is increased from 3 to 6. Its H^+ ion concentration will be :-
(1) Reduced to half (2) Doubled
(3) Reduced by 1000 times (4) Increased by 1000 times

45. The pH of a solution is defined by the equation:-
(1) $pH = -\log [H_3O^+]$ (2) $pH = \log \frac{1}{[H_3O^+]}$ (3) $[H^+] = 10^{-pH}$ (4) All

46. In a solution of pH = 5, more acid is added in order to reduce the pH = 2. The increase in hydrogen ion concentration is:-
(1) 100 times (2) 1000 times (3) 3 times (4) 5 times

47. An aqueous solution whose pH = 0 is :-
(1) Basic (2) Acidic (3) Neutral (4) Amphoteric

48. Pure water is kept in a vessel and it remains exposed to atmospheric CO_2 which is absorbed. Then the pH will be
(1) Greater than 7 (2) Less than 7
(3) 7 (4) Depends on ionic product of water

49. An aqueous solution of HCl is 10^{-9} M HCl. The pH of the solution should be:-
(1) 9 (2) Between 6 and 7 (3) 7 (4) Unpredictable

50. Calculate pH of a solution whose 100 mL contains 0.2 g NaOH dissolved in it?
(1) 10.699 (2) 11.699 (3) 12.699 (4) 13.699

SALTS AND ITS TYPES

51. Which formula represents a salt?

(1) KOH (2) KCl (3) CH_3OH (4) CH_3COOH

52. Which of the following statements is true concerning acids and bases?

(1) Acids and bases don't react with each other

(2) Acids mixed with bases neutralize each other

(3) Acids mixed with bases make stronger bases

(4) Acids mixed with bases make stronger acids

53. What is the nature of a salt of strong acid and strong base ?

(1) Acidic (2) Basic (3) Neutral (4) None of the above

54. The acid present with sodium bicarbonate in baking powder is

(1) Tartaric Acid (2) Acetic Acid (3) Sulphuric Acid (4) Citric Acid

55. A salt formed when all the replaceable hydrogen atoms are replaced by a metal is called

(1) Normal salt (2) Anhydrous salt (3) Acidic salt (4) Amphoteric salt

56. A salt which has water of crystallization is called

(1) Acidic salt (2) Salt hydrate (3) Anhydrous salt (4) Double salt

57. In neutralisation reaction

(1) An acid reacts with sodium salts (2) An alkali reacts with ammonium salts

(3) An acid reacts with an alkali (4) Alkali reacts with sodium salts

58. The heat of neutralisation is the highest in the following case :-

(1) A strong acid and a weak base (2) A weak acid and a strong base

(3) A strong acid and a strong base (4) A weak acid and a weak base

59. The process of neutralisation invariably results in the production of :-

(1) H^+ ions (2) OH^- ions

(3) Both H^+ and OH^- ions (4) Molecules of water

60. A salt 'X' is dissolved in water of pH = 7. The resulting solution becomes alkaline in nature. The salt is made up of:-

(1) A strong acid and strong base (2) A strong acid and weak base

(3) A weak acid and weak base (4) A weak acid and strong base

IONIZATION OF ACID AND BASES

61. Given the equation: $H^+ + OH^- \rightleftharpoons H_2O$

Which type of reaction does the equation represent?

(1) Esterification (2) Decomposition (3) Hydrolysis (4) Neutralization

62. As the hydrogen ion concentration of an aqueous solution increases, the hydroxide ion concentration of this solution will

(1) Decrease (2) Increase (3) Remain the same (4) None

63. Which of the following statement is not true of dilute hydrochloric acid?

(1) It liberates ammonia gas from ammonium chloride on warming

(2) It liberates carbon dioxide from carbonates

(3) It ionises in water to form hydrogen ions

(4) It reacts with active metals to produce hydrogen

64. Which of the following is true? The strength of an acid depends on:

(1) It's concentration
(2) It's pH
(3) It's solubility in water
(4) It's degree of ionisation in water

65. Which of the following observation can be obtained from this reaction?

$H_2SO_4(aq) + CuCO_3(s) \rightarrow H_2O\ (l) + CO_2(g) + CuSO_4(aq)$

(1) An effervescence of colourless gas with a pungent smell is observed
(2) Copper(II) carbonate reacts with the acid and disappears, forming a blue solution
(3) Copper(II) carbonate reacts with the acid and disappears, forming a colourless solution
(4)Copper(II) carbonate reacts with the acid and disappears, forming a pale green solution

66. The oxides of non–metals like sulphur dissolve in water to form

(1) An acid (2) A base (2) A neutral solution (4) Both (1) and (2)

67. The carbon dioxide is soluble in water. The acid formed when it dissolves in water is called

(1) Carboxylic acid (2) Hydrochloric acid (3) Carbonic acid (4) Formic acid

68. The pH of 0.1M aqueous solution of a weak acid (HA) is 3. What is its degree of dissociation?

(1) 1% (2) 10% (3) 50% (4) 25%

69. Among the following hydroxides, the one which has the lowest value of K_b at ordinary temperature (about 25°C) is :

(1) $Mg\ (OH)_2$ (2) $Ca\ (OH)_2$ (3) $Ba\ (OH)_2$ (4) $Be\ (OH)_2$

70. The hydrogen ion concentration in a solution of weak acid of dissociation constant Ka and concentration C is nearly equal to :-

(1) $\sqrt{K_a.C}$ (2) $\sqrt{\frac{K_a}{C}}$ (3) C/K_a (4) $K_a.C$

71. When 0.01 M HCl solution is added in aqueous solution of acetic acid

(1) CH_3COO^- molar conc. is decreased
(2) CH_3COOH molar conc. is decreased
(3) CH_3COO^- molar conc. is increased
(4) NO change in conc. of any species

72. Degree of dissociation for a weak electrolyte is :

(1) Reduces with dilution
(2) Increases with dilution
(3) Reduces or increases with dilution
(4) Not effected by dilution

73. Which of the following is a true statement :

(1) The ionisation constant and ionic product of water are same.
(2) Water is a strong electrolyte.
(3) The value of ionic product of water is less than that of its ionisation constant.
(4) At 298K, the number of H^+ ions in a litre of water is 6.23×10^{16}.

74. pH of 0.001M acetic acid would be $pK_a = 4.75$:-

(1) 2 (2) > 3 (3) 7 (4) 14

75. The pH of a 0.02 M ammonia solution which is 5% ionised will be :-

(1) 2 (2) 11 (3) 5 (4) 7

HYDROLYSIS OF SALT

76. Which of the following will not undergo hydrolysis in water?
(1) Sodium sulphate (2) Ammonium sulphate
(3) Calcium sulphate (4) All the salts will hydrolyse

77. pH of the solution of $HCOONH_4$ is 6.48 this can be explained by :-
(1) Hydrolysis of both cation and anion (2) Hydrolysis of cation
(3) Hydrolysis of anion (4) Hydrolysis of water

78. Which of the following salt undergo hydrolysis?
(1) CH_3COONa (2) KNO_3 (3) NaCl (4) K_2SO_4

79. Which salt will not undergo hydrolysis?
(1) KCl (2) Na_2SO_4 (3) NaCl (4) All

80. Which of the following salt undergo hydrolysis in water?
(1) Na_3PO_4 (2) CH_3COONa (3) $NaNO_3$ (4) Both (1) and (2)

BUFFER SOLUTION

81. The pH of blood is maintained by CO_2 and H_2CO_3 in the body and chemical constituents of blood. This phenomenon is called :-
(1) Colloidal (2) Buffer action (3) Acidity (4) Salt balance

82. A solution which resists any change in pH value on addition of small amount of an acid or base is called
(1) Amphoteric solution (2) Buffer solution (3) Acidic solution (4) Alkaline solution

83. A buffer solution contains
(1) A weak acid and its salt with a strong base
(2) A weak base and its salt with a strong acid
(3) Both (1) and (2)
(4) Neither (1) nor (2)

84. If a buffer solution of ammonium chloride and ammonium hydroxide is diluted with water, its pH value will
(1) Become seven (2) Increase (3) Decrease (4) Remain the same

85. The buffer solution play an important role in :-
(1) Increasing the pH value (2) Decreasing the pH value
(3) Keeping the pH constant (4) Solution will be neutral

SOLUBILITY PRODUCT AND ITS APPLICATION

86. The correct representation of solubility product of SnS_2 is :-
(1) $[Sn^{4+}][S^{2-}]^2$ (2) $[Sn^{4+}][S^{2-}]$ (3) $[Sn^{4+}][2S^{2-}]$ (4) $[Sn^{4+}][2S^{2-}]^2$

87. The solubility of A_2X_3 is y mol dm^{-3}. Its solubility product is :
(1) $6\ y^4$ (2) $64\ y^4$ (3) $36\ y^5$ (4) $108\ y^5$

88. When HCl gas is passed through saturated solution of common salt, pure NaCl get precipitated because :-
(1) In aqueous solution Cl^- obtained from dissociation of HCl which reduces the solubility product of NaCl.
(2) More water is needed to dissociate HCl gas.
(3) Impurities of NaCl solution dissolves in HCl.
(4) Ionic product of $[Na^+]\ [Cl^-]$ > solubility product of NaCl.

89. Solubility of a M_2S salt is 3.5×10^{-6} then find out solubility product :-
(1) 1.7×10^{-6} (2) 1.7×10^{-16} (3) 1.7×10^{-18} (4) 1.7×10^{-12}

90. Relation between solubility & solubility product of A_xB_y type salt is :-
(1) $K_{sp} = x^x\ y^y\ s^{x-y}$ (2) $K_{sp} = S^{x+y}$ (3) $K_{sp} = x^y y^x S^{x+y}$ (4) $K_{sp} = x^x\ y^y\ S^{x+y}$

EXERCISE-II

INTRODUCTION

1. Which solution will change red litmus to blue?
(1) HCl(aq) (2) NaCl(aq) (3) CH_3OH(aq) (4) NaOH(aq)

2. The acids turn the blue litmus paper
(1) Red (2) Green (2) White (4) Black

3. The gas evolved when an acid reacts with a metal is
(1) Carbon Dioxide (2) Sulphur Dioxide (3) Hydrogen (4) Nitrogen

4. One can test the hydrogen gas in laboratory as it
(1) Turns blue litmus paper violet (2) Turns red litmus paper blue
(3) Gives cloudy suspension with lime water (4) Burns with a popping sound when ignited

5. Which of the following metals react with dilute acids?
(1) Copper (2) Gold (3) Silver (4) Calcium

6. Which of the following metals will not react with dilute acids?
(1) Magnesium (2) Zinc (3) Lead (4) Sodium

7. It is not advisable to store the substances like lemon achar carrying acidic content in metallic containers like that of iron because
(1) Metals are corrosive in nature
(2) Acids present react with the metal lining
(3) Metals change the colour of the substance stored in them
(4) Metals acts as fungi and decompose them

8. The gas evolved when a mineral acid reacts with a carbonate is
(1) Hydrogen (2) Carbon monoxide (3) Carbon dioxide (4) Sulphur dioxide

9. The chemical which can act as antacid is
(1) Magnesium hydroxide (2) Sodium hydrogen carbonate
(3) Sodium chloride (4) Both (1) and (2)

10. The methyl orange gives '———' colour with alkalis.
(1) Brown (2) Blue (3) Red (4) Yellow

11. Which of the following is an indicator to help distinguish between acids and bases?
(1) Phenolphthalein (2) Litmus paper
(3) Universal indicator solution (4) All of these

12. An agent which readily destroys the skin is called
(1) Caustic (2) Noble (3) Mild (4) Amorphous

13. The oxides of metals are usually
(1) Acidic Oxides (2) Basic Oxides
(3) Amphoteric Oxides (4) Neutral Oxides

14. The number of hydrogen ions obtainable from a molecule of an acid is called
(1) Acidity (2) Basicity (3) Hydrogenicity (4) Anhydricity

15. In a neutral solution
(1) Number of hydrogen ions are more than the number of hydroxyl ions
(2) Number of hydrogen ions are less than the number of hydroxyl ions
(3) Number of hydrogen ions equals the number of hydroxyl ions
(4) There are no hydroxyl ions

16. When an acid reacts with an alkali, the products are
(1) Carbon dioxide and water (2) Salt and water
(3) Salt and carbon dioxide (4) Salt and hydrogen

17. Which indicator works in the pH range 8 – 9.8?
(1) Phenolphthalein (2) Methyl orange (3) Methyl red (4) Litmus

18. The oxides of non–metals like sulphur dissolve in water to form
(1) an acid (2) a base (2) a neutral solution (4) Both (1) and (2)

19. One can test the hydrogen gas in laboratory as it
(1) turns blue litmus paper violet (2) turns red litmus paper blue
(3) gives cloudy suspension with lime water (4) burns with a popping sound when ignited

20. A substance which can act both an acid and a base is :
(1) allotropic (2) amphoteric (3) isotopic (4) amorphous

THEORIES OF ACID AND BASES

21. Which 0.1 M solution will turn phenolphthalein pink?
(1) HBr(aq) (2) CO_2(aq) (3) LiOH(aq) (4) CH_3OH(aq)

22. The process by which the sulphuric acid is commercially produced is called
(1) Contact process (2) Haber's process (3) Solvay process (4) Bosch process

23. A mineral acid does not have the atom of '——' as the central atom.
(1) Nitrogen (2) Sulphur (3) Hydrogen (4) Carbon

24. The nitric acid is manufactured by
(1) Haber's process (2) Solvay process (3) Contact process (4) Ostwald's process

25. The acid which also acts as a dehydrating agent is
(1) Sulphuric acid (2) Hydrochloric acid (3) Nitric acid (4) Carboxylic acid

26. If hot concentrated sulphuric acid is mixed with sugar, the product would be
(1) Hydrogen gas and water (2) Carbon dioxide and water
(3) Sugar charcoal and water (4) None of these

27. In the reaction : $CH_3COOH + HCl \rightarrow Cl^- + CH_3COOH_2^+$
The conjugate acid of acetic acid is
(1) Cl^- (2) HCl (3) H_3O^+ (4) $CHCOOH_2^+$

28. The strongest conjugate base is :
(1) NO_3^- (2) OH^- (3) SO_4^{2-} (4) CH_3COO^-

29. The molecule BF_3 is a :
(1) Bronsted acid (2) Lewis acid (3) Lewis base (4) Salt

30. Which one of the following is a Lewis acid?
(1) HCl (2) H_2O (3) C_2H_4 (4) anhydrous $AlCl_3$

31. Lewis base is
(1) Electron pair donor (2) Electron pair acceptor
(3) Proton donor (4) Electron acceptor

32. Si $(OH)_4$ is :
(1) an acid (2) a base (3) a salt (4) an amphoteric oxide

33. What is the conjugate base of H_2O?
(1) O_2 (2) OH^- (3) O^- (4) O^{2-}

34. In the reaction, $NH_3(aq) + H_2O\ (l) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$ the ammonium ion is the

(1) Lewis Base (2) Conjugate Acid (3) Conjugate Base (4) None of these

35. The conjugate base of $(CH_3)_2\ NH_2^+$ is–

(1) CH_3NH_2 (2) $(CH_3)_2N^+$ (3) $(CH_3)_3N$ (4) $(CH_3)_2NH$

36. In the reaction : $NH_3 + BF_3 \longrightarrow [NH_3 : \rightarrow BF_3]$, BF_3 is : -

(1) a Lewis acid (2) a Lewis base

(3) neither a Lewis acid nor a Lewis base (4) a Lewis acid and also Lewis base

37. Which of the following is not a Lewis base?

(1) $C_2H_5 - NH_2$ (2) $C_2H_5 - OC_2H_5$ (3) CH_3OH (4) BF_3

38. The species among the following, which can act as an acid and a base is

(1) HSO_4^- (2) SO_4^{2-} (3) H_3O^+ (4) Cl^-

39. Which of the following acts as both Bronsted acid and Arrhenius acid ?

(1) Cu^{+2} (2) SO_2 (3) Fe^{+3} (4) HCl

40. In the reaction, $H_2PO_4^- + H_2O(l) \rightarrow H_3O^+\ (l) + HPO_4^{2-}(l)$ the hydrogen phosphate ion is a:

(1) Lewis Base (2) Conjugate Acid (3) Conjugate Base (4) None of these

pH SCALE

41. For $\frac{N}{10}$ H_2SO_4, pH value is :-

(1) 1 (2) 0.586 (3) 0.856 (4) None

42. 8 g NaOH is dissolved in one litre of solution, the molarity of the solution is:-

(1) 0.2 M (2) 0.4 M (3) 0.02 M (4) 0.8 M

43. When equal volumes of two liquids are mixed a solution with a value of pH 7 is formed. Which of the following are the pH values of the two liquids ?

(1) 1st: pH 5, 2nd: pH 2 (2) 1st: pH 5, 2nd: pH 9

(3) 1st: pH 6, 2nd: pH 1 (4) 1st: pH 14, 2nd: pH 7

44. The pH of an aqueous solution of sodium acetate is

(1) More than seven (2) Less than seven (3) Equal to seven (4) More than ten

45. What would be the hydrogen ion concentration of 0.006 M benzoic acid ($K_a = 6 \times 10^{-5}$)

(1) 0.6×10^{-4} (2) 6×10^{-4} (3) 6×10^{-5} (4) 3.6×10^{-4}

46. Ionic product of water is equal to :-

(1) Dissociation constant of water $\times [H_2O]$ (2) Dissociation constant of water $\times [H^+]$

(3) Product of $[H_2O]$ and $[H^+]$ (4) Product of $[OH^-]^2$ and $[H^+]$

47. A solution of $MgCl_2$ in water has pH :

(1) < 7 (2) > 7 (3) 7 (4) 14.2

48. If pH = 3.31, then find out $[H^+]$ (Approximate)

(1) 3.39×10^{-4} (2) 5×10^{-4} (3) 3.0×10^{-3} (4) None

49. Which one of the following has highest pH?

(1) Distilled water (2) 1 M NH_3

(3) 1 M NaOH (4) Water saturated with chlorine

50. The pH of aqueous solution of sodium acetate is

(1) 7 (2) Very low (3) > 7 (4) < 7

51. The highest pH value is of :-

(1) 0.1 M NaCl (2) 0.1 M NH_4Cl

(3) 0.1 M CH_3COONa (4) 0.1 M CH_3COONH_4

52. pH of K_2S solution is:-
(1) 7 (2) Less than 7 (3) More than 7 (4) 0

53. Given :-
(a) 0.005 M H_2SO_4 (b) 0.1 M Na_2SO_4 (c) 10^{-2} M NaOH (d) 0.01 M HCl
Choose the correct code having same pH :-
(1) a, c, d (2) b, d (3) a, d (4) a, c

54. If 100 mL of pH = 3 and 400 mL of pH = 3 is mixed, what will be the pH of the mixture
(1) 3.2 (2) 3.0 (3) 3.5 (4) 2.8

55. A solution has pOH equal to 13 at 298 K. The solution will be:-
(1) Highly acidic (2) Highly basic (3) Moderately basic (4) Unpredictable

56. The pH of the solution containing 10 mL of a 0.1M NaOH and 10 mL of 0.05M H_2SO_4 would be
(1) Zero (2) 1 (3) >7 (4) 7

57. The pH of the solution produced when an aqueous solution of strong acid pH 5 is mixed with equal volume of an aqueous solution of strong acid of pH 3 is :-
(1) 3.3 (2) 3.5 (3) 4.5 (4) 4.0

58. pH of 10^{-1} M formic acid is:-
(1) 1 (2) > 1 (3) < 1 (4) 13

59. H_2X is a dibasic acid which dissociates completely in water. Which one of the following is the molarity of an aqueous solution of this acid which has a pH of 1?
(1) 0.1 (2) 0.05 (3) 0.2 (4) 0.5

60. The formula weight of H_2SO_4 is 98. The weight of the acid in 400mL of 0.1 M solution is:-
(1) 2.45 g (2) 3.92 g (3) 4.90 g (4) 9.8 g

61. If $[OH^-] = 5.0 \times 10^{-5}$ M then pH will be :-
(1) 5 - log 5 (2) 9 + log 5 (3) log 5 - 5 (4) log 5 - 9

62. The pH of a 10^{-10} M NaOH solution is nearest to :-
(1) 10 (2) 7 (3) 4 (4) – 10

63. The hydrogen ion concentration in a given solution is 6×10^{-4} M. Its pH will be :-
(1) 6 (2) 3.22 (3) 4 (4) 2

64. What would be he pH of 0.001M acetic acid would be
(1) 2 (2) > 3 (3) 7 (4) 14

65. The hydrogen ion concentration of a wine is 1.4×10^{-4} M. What is the pH of the wine ?
(Given :- log 1.4 = 0.1461)
(1) 2.15 (2) 2.85 (3) 3.85 (4) 4.14

SALTS AND ITS TYPES

66. Solution of sodium carbonate is :-
(1) Strong acid (2) Weak base (3) Strong base (4) weak acid

67. In an acidic salt
(1) All the replaceable hydrogen atoms in an acid are replaced by metal
(2) A part of replaceable hydrogen atoms in an acid are replaced by metal
(3) No hydrogen is replaced by metal
(4) None of the above

68. At 90°C pure water has $[H_3O^+] = 10^{-6}$ mol L^{-1}. The value of K_W at 90°C is :-
(1) 10^{-6} (2) 10^{-12} (3) 10^{-14} (4) 10^{-8}

69. Ionic product of water increases with :–
(1) Decreasing pressure (2) Increasing H^+
(3) By mixing OH^- (4) Increasing temperature

70. Y is heated with aqueous sodium hydroxide and ammonia gas is formed. Y is ________.
(1) Hydrochloric acid (2) Copper (II) oxide
(3) Ammonium sulphate (4) Sodium nitrate

71. The salts which in weakly alkaline or neutral solutions form insoluble precipitates are mostly
(1) Acidic Salts (2) Basic Salts (3) Amphoteric Salts (4) Double Salts

72. Which aqueous solution having maximum pH value
(1) NaOH (2) $NaHSO_3$ (3) Na_2CO_3 (4) NaCl

73. Which of the following is a strong electrolyte :-
(1) NH_4OH (2) $Ca(NO_3)_2$ (3) HCN (4) H_2SO_3

74. Aqueous solution of ammonium acetate is :
(1) Weakly acidic (2) Weakly basic (3) Strongly acidic (4) Almost neutral

75. An aqueous solution of CH_3COONa will be :-
(1) Acidic (2) Alkaline (3) Neutral (4) None of these

76. Aqueous solution of $Al_2(SO_4)_3$ is :–
(1) Basic & acidic (2) Neutral (3) Basic (4) Acidic

77. An aqueous solution of $FeCl_3$ is :–
(1) basic (2) acidic (3) neutral (4) amphoteric

78. Which of the following is not an acidic salt?
(1) $NaHSO_4$ (2) HCOONa (3) NaH_2PO_3 (4) None of them

79. Which of the following is an acid salt?
(1) Na_2S (2) Na_2SO_3 (3) $NaHSO_3$ (4) Na_2SO_4

80. Consider :-
(a) $FeCl_3$ in water - Basic (b) NH_4Cl in water - Acidic
(c) Ammonium acetate in water - Acidic (d) Na_2CO_3 in water - Basic
Which is/are not correctly matched:-
(1) b and d (2) b only (3) a and c (4) d only

81. The pH of 0.1 M solution of the following salts increases in order :-
(1) $NaCl < NH_4Cl < NaCN < HCl$ (2) $NaCN < NH_4Cl < NaCl < HCl$
(3) $HCl < NaCl < NaCN < NH_4Cl$ (4) $HCl < NH_4Cl < NaCl < NaCN$

82. A solution of H_3BO_3 and borax is called :-
(1) Acidic buffer (2) Basic buffer (3) Basic solution (4) None of these

83. When sodium acetate (CH_3COONa) is added to aqueous solution of acetic acid (CH_3COOH), the-
(1) the pH value becomes zero (2) pH value remains unchanged
(3) pH value decreases (4) pH value increases

84. Which of the following salts does not undergo hydrolysis ?
(1) KCN (2) KCl (3) NH_4NO_3 (4) $FeCl_3.6H_2O$

85. Which of the following salts is neutral in water?
(1) KCl (2) NH_4NO_3 (3) NH_4CN (4) NH_4F

IONIZATION OF ACID AND BASES

86. Ionisation constant of CH_3COOH is 1.7×10^{-5} and concentration of H^+ ions is 3.4×10^{-4}. Then find out initial concentration of CH_3COOH molecules :-

(1) 3.4×10^{-4} (2) 3.4×10^{-3} (3) 6.8×10^{-4} (4) 6.8×10^{-3}

87. The ionization of NH_4OH is suppressed by :

(1) NaOH (2) NH_4Cl (3) Both (1) and (2) (4) NaCl

88. Addition of conc. HCl to saturated $BaCl_2$ solution precipitates $BaCl_2$, because:-

(1) It follows from Le Chatelier's principle

(2) of common-ion-effect

(3) ionic product $[Ba^{++}][Cl^-]$ remains constant in saturated solution

(4) At constant temperature, the product $[Ba^{2+}]$ $[Cl^-]^2$ remains constant in a saturated solution.

89. An acid HA dissociates in the following manner $HA \rightleftharpoons H^+ + A^-$

It has concentration 1 M and pH = 5 then find out dissociation constant :-

(1) 1×10^{-10} (2) 1×10^{-5} (3) 5×10^{-5} (4) 5

90. Correct relation is :-

(1) $pK_a \times pK_b = pK_w$ (2) $pK_a + pK_b = pK_w$ (3) $pK_a/pK_b = pK_w$ (4) $pK_a - pK_b = pK_w$

91. The amount of acetic acid present in 100 mL of 0.1M solution is :-

(1) 0.30 g (2) 3.0 g (3) 0.60 g (4) None

92. The common ion effect is shown by which of the following sets of solutions?

(1) $BaCl_2 + BaNO_3$ (2) NaCl + HCl (3) $NH_4OH + NH_4Cl$ (4) None

93. Basic strength of NH_4OH in presence of NH_4Cl

(1) Increases (2) Remains unchanged

(3) Decreases (4) Some times increases or sometimes decreases

94. The pH of a 0.02 M ammonia solution which is 5% ionised will be

(1) 2 (2) 11 (3) 5 (4) 7

95. The pH of H_2SO_4 = 2. Its molar concentration is

(1) 0.01 (2) 0.005 (3) 0.02 (4) 0.05

POLYBASIC ACIDS AND POLYACIDIC BASES

96. The dissociation constant of H_2S and HS^- are respectively 10^{-7} and 10^{-13}. The pH of 0.1 M aqueous solution of H_2S will be

(1)4 (2)3 (3)5 (4)2.5

97. The pH of 0.1 M $NaHCO_3$ is (Given K_{a1} and K_{a2} for H_2CO_3 are 6.38 and 10.26 respectively)

(1) 1.7 (2) 5.68 (3) 7.00 (4) 8.32

98 pK_{a1} and pK_{a2} of H_2CO_3 are 6.38 and 10.26 respectively. The pH of 1 M and 0.1 M $NaHCO_3$ are

(1) 8.32, 7.32 (2) 7.32, 8.32 (3) 8.32, 8.32 (4) 7.32, 7.32

99. pK_{a1}, pK_{a2} and pK_{a3} of H_3PO_4 are respectively x, y and z. pH of 0.1 M Na_2HPO_4 solution is

(1) 1 (2) $\frac{1}{2}(x+y)$ (3) $\frac{1}{2}(y+z)$ (4) $\frac{1}{2}(x+y+z)$

100. pH of $Ca(OH)_2$ solution is 12, Millimoles of $Ca(OH)_2$ present in 100 mL of solution will be

(1) l (2) 0.5 (3) 0.05 (4)5

HYDROLYSIS OF SALT

101. The compound whose 0.1 M solution is basic is

(1) Ammonium acetate (2) Ammonium chloride

(3) Ammonium sulphate (4) Sodium acetate

102. The pK_a of a weak acid, HA, is 4.80. The pK_b of a weak base, BOH. is 4.78. The pH of an aqueous solution of the corresponding salt, BA, will be

(1) 8.58 (2) 4.79 (3) 7.01 (4) 9.22]

103. If pK_b for CN^- at 25°C is 4.7. The pH of 0.5M aqueous NaCN solution is

(1) 12 (2) 10 (3) 11.5 (4) 11

104. 500 mL of 0.2 M acetic acid are added to 500 mL of 0.30 M sodium acetate solution. If the dissociation constant of acetic acid is 1.5×10^{-5} then pH of the resulting solution is

(1) 5.0 (2) 9.0 (3) 3.0 (4) 4.0

105. A solution contains 0.2M NH_4OH and 0.2M NH_4Cl. If 1.0 mL of 0.001 M HCI is added to it. What will be the $[OH^-]$ of the resulting solution $[K_b = 2 \times 10^{-5}]$

(1) 2×10^{-5} (2) 5×10^{-10} (3) 2×10^{-3} (4) None of these

BUFFER SOLUTION

106. Which of the following is not a buffer solution?

(1) $\frac{H_2CO_3}{HCO_3^-}$ (2) $\frac{CH_3COOH}{CH_3COONa}$ (3) $\frac{NH_4OH}{NH_4Cl}$ (4) $\frac{NH_4OH}{CH_3COOH}$

107. Which of the following would produce a buffer solution when mixed in equal volumes?

(1) 1 M CH_3 COOH and 0.5 M NaOH (2) 1 M CH_3 COOH and 0.5 M HCI

(3) 1 M NH_4 OH and 0.5 M NaOH (4) 1 M NH_4Cl and 0.5 M HCl

108. Which can act as buffer?

(1) NH_4OH + NaOH (2) HCOOH + CH_3COONa

(3) 40 mL 0.1 M NaCN + 20 mL of 0.1 M HCl (4) None of them

109. In a buffer solution the ratio of concentration of NH_4CI and NH_4OH is 1 : 1 when it changes in 2: 1 what will be the value of pH of buffer

(1) Increase (2) Decrease (3) No effect (4) None of these

110. pK_b for NH_4OH at certain temperature is 4.74. The pH of basic buffer containing equimolar concentration of NH_4OH and NH_4Cl will be

(1) 7.74 (2) 4.74 (3) 2.37 (4) 9.26

SOLUBILITY PRODUCT AND ITS APPLICATION

111. The precipitation occur if ionic concentration is :

(1) Less than solubility product (2) More than solubility product

(3) Equal to solubility product (4) Either '1' or '3'

112. The units of solubility product of sliver chromate (Ag_2CrO_4) will be

(1) $mol^2\ L^{-2}$ (2) $mol^3\ L^{-3}$ (3) $mol\ L^{-1}$ (4) $mol^{-1}\ L$

113. The molar solubility of silver sulphate is 1.5×10^{-2} mol L^{-1}. The solubility product of the salt will be

(1) 2.25×10^{-4} (2) 1.35×10^{-5} (3) 1.7×10^{-6} (4) 3.0×10^{-3}

114. If S_0, S_1, S_2 and S_3 are the solubilities of AgCl in water, 0.01 M $CaCL_2$, 0.01 M NaCl and 0.5 M $AgNO_3$ solutions, respectively, then which of the following is true?

(1) $S_0 > S_2 > S_1 > S_3$ (2) $S_0 = S_2 = S_1 > S_3$ (3) $S_3 > S_1 > S_2 > S_0$ (4) $S_0 > S_4 > S_3 > S_1$

115. Given K_{sp} (AgI) = 8.5×10^{-17}. The solubility of AgI in 0.1 M KI solution is

(1) 0.1 M (2) 8.5×10^{-16} M (3) 8.5×10^{-17} M (4) 8.5×10^{-18} M

EXERCISE-III

QUESTIONS ASKED IN PREVIOUS EXAMS

1. In more acidic solution, Fe(II) does not precipitate as it's sulphide. It is due to :-
(1) High concentration of sulphide ion (2) Low concentration of sulphide ion
(3) Acidic medium (4) Fe (II) is not in II group

2. A solution of $FeCl_3$ in water acts as acidic due to:-
(1) Acidic impurities (2) Ionisation (3) Hydrolysis of Fe^{3+} (4) Dissociation

3. What is the suitable indicator for titration of NaOH and oxalic acid?
(1) Methyl orange (2) Methyl red (3) Phenolphthalein (4) Starch solution

4. Phenolphthalein does not act as an indicator for the titration between :-
(1) KOH and H_2SO_4 (2) NaOH and CH_3COOH
(3) Oxalic acid and $KMnO_4$ (4) $Ba(OH)_2$ and HCl

5. Which solution, of concentration 0.10 mol dm^{-3}, has the highest pH value ?
(1) HCl(aq) (2) $MgCl_2$(aq) (3) NaCl(aq) (4) $AlCl_3$(aq)

6. Which compound dissolves in water to form an aqueous solution that can conduct an electric current ?
(1) CCl_4 (2) C_2H_5OH (3) CH_3COOH (4) CH_4

7. The pH of solution X is 1 and that of Y is 2. Which statement is correct about the hydrogen ion concentrations in the two solutions ?
(1) $[H^+]$ in X is half that in Y (2) $[H^+]$ is X is twice that in Y
(3) $[H^+]$ in X is one tenth of that in Y. (4) $[H^+]$ in X is ten times that in Y.

8. 50 cm^3 of 1.5 M NaOH is titrated against 2M HCl. The pH of the reaction system after the addition of 35 ml of HCl will be :
(1) 0.84 (2) 1.23 (3) 12.76 (4) 7.95

9. 0.1 mole of a weak acid HA and 0.1 mole of its sodium salt are dissolved in water and the solution is made upto 100 cm^3. If $K_a = 1.8 \times 10^{-5}$, the pH of the solution is :
(1) 2.4 (2) 3.70 (3) 4.74 (4) 5.00

10. Which of the following solution will have pH close to 1.0 ?
(1) 100 mL 0.1 M HCl + 100 mL 0.1 M NaOH
(2) 55 mL 0.1 M HCl + 45 mL 0.1 M NaOH
(3) 10 mL 0.1 M HCl + 90 mL 0.1 M NaOH
(4) 75 mL 0.2 M HCl + 45 mL 0.2 M NaOH

11. The pKa of aspirin, a weak acid is 3.5. The pH of gastric juice in the human stomach is between 2 and 3, while that in the small intestine is about 8. Aspirin will be .....
(1) Unionized in the small intestine and stomach
(2) Completely ionized in the small intestine and stomach
(3) Ionized in the small intestine and almost unionized in the stomach
(4) Ionized in the stomach and almost unionized in the small intestine in the stomach

12. The compound whose 0.1 M solution is basic is
(1) Ammonium acetate (2) Ammonium chloride
(3) Sodium acetate (4) Sodium sulphate

13. The pH of solution X is 2 and that of Y is 4. Which statement is correct about the hydrogen ion concentrations in the two solution ?
(1) $[H^+]$ in X is half that in Y (2) $[H^+]$ in X is twice that in Y
(3) $[H^+]$ is X is ten times of that in Y (4) $[H^+]$ in X is hundred times that in Y

14. One mole of oxalic acid is equivalent to :
(1) 0.5 mole of NaOH (2) 1 mole of NaOH (3) 1.5 mole of NaOH (4) 2 mole of NaOH

15. Out of the following the salt/s that has pH value higher that 7.5 is/are

(1) Sodium hydrogen carbonate (2) Sodium sulphate

(3) Sodium chloride (4) Sodium carbonate

16. Copper sulphate is the salt of

(1) A weak acid and a strong base (2) A weak acid and a weak base

(3) A strong acid and a weak base (4) A strong acid and a strong base

17. A solution having pH 2 mixed with excess of solution of washing soda, the pH of mixture is

(1) 4 (2) 5 (3) 6 (4) 8

18. 1.000 mL of 0.1000 mol L^{-1} hydrochloric acid was diluted to 100.0 mL with deionised water. 10.00 mL of this solution was diluted to 100.0 mL again using deionised water. What is the pH of the final solution ?

(1) 2 (2) 3 (3) 4 (4) 5

19. The pH of a 0.025 M solution of KOH is

(1) 1.60 (2) 3.69 (3) 10.31 (4) 12.40

20. A white salt is readily soluble in water and gives colorless solution with pH of about 9. The salt would be.....

(1) NH_4NO_3 (2) CH_3COONa (3) CH_3COONH_4 (4) $CaCO_3$

21. Which one is not an acid salt ?

(1) NaH_2PO_4 (2) Na_3PO_4 (3) NaH_2PO_3 (4) None of these

22. Sumit went to a fun with his friends Amit and Rohit. Rohit and Amit were scared to sit on a merry go round and preferred to stroll around. Sumit was very excited when he came down the merry go round. How will this change the pH of his blood ?

(1) increases (2) decrease

(3) no change in pH (4) pH level gets adjusted at 7

23. Oxides are acidic, basic or amphoteric based on their metallic or non-metallic character. Which one of the following oxides reacts with both HCl and and NaOH ?

(1) CaO (2) ZnO (3) SO_2 (4) CO_2

24. Esha performed a simple experiment to distinguish strong acid from weak acid. For this she performed experiments with universal indicator, using tamarind and the acid present in gastric juice and she recorded her observations. Which of the following finding did she observe ?

	Acid present in gastric		**Acid present in**	
	Strength	Colour of universal indicator	Strength	Colour of universal indicator
(1)	weak	Red	Strong	Red
(2)	weak	Yellow	Weak	Green
(3)	strong	Light red	Weak	Yellow
(4)	strong	Green	strong	Blue

25. Amol took 10 mL of 2.2×10^{-5} M hydrochloric acid solution. He then diluted it to 1 litre. He found that the pH of diluted solution is

(1) 4.7 (2) 6.1 (3) 4.5 (4) 6.65

26. Assertion (A) Sodium carbonate can be titrated against sulphuric acid by using either phenolphthalein or methyl orange as indicator.

Reason: The volume of sulphuric acid required to produce colour for two indicators is different.

(1) Both (A) and (R) are true and (R) is the correct explanation of (A).

(2) Both (A) and (R) are true and (R) is not the correct explanation of (A).

(3) (A) is true but (R) is false.

(4) (A) is false but (R) is true.

27. There are 3 containers X, Y and Z. X contains 10 ml of milk. Y contains 5 ml of milk same in container Z, mixed with 5 ml of water. All 3 containers have pH value of 6.5. P amount of Acetic acid is added to container X, Q amount to Y and R amount to Z. Such that the final pH value in the three containers is

(1) $P < Q < R$ (2) $P < R = Q$ (3) $P = Q = R$ (4) $P < R < Q$

28. The pH of blood is maintained within the range 7.36 – 7.242 by

(1) CH_3COONH_4 (2) HCO_3^- / CO_3^{2-}

(3) CH_3COONa/CH_3COOH (4) CH_3COOH

29. Which among the following salts will not change the pH of water on addition.

(1) Sodium cyanide (2) Sodium chloride (3) Sodium bicarbonate (4) Sodium carbonate

30. Sati was studying neutralisation reaction. She accidentally dropped 'x' grams of a sodium carbonate monohydrate into a 100 mL solution of HCl whose concentration was 0.25 M. 10 mL from this final solution was taken and titrated against 0.05 M NaOH solution to yield a titre value of 20 mL. What is the value of 'x'.

(1) 186 mg (2) 93 mg (3) 1860 mg (4) 930 mg

ANSWER KEY

Exercise-I

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	3	4	2	3	2	2	4	2	4	4	2	3	3	2	3	1	1	3	1	3
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	2	4	1	4	3	3	2	3	3	4	2	1	3	1	1	2	1	2	1	4
Que.	**41**	**42**	**43**	**44**	**45**	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	2	3	3	3	3	2	2	2	2	3	2	2	3	1	1	2	3	3	4	4
Que.	**61**	**62**	**63**	**64**	**65**	**66**	**67**	**68**	**69**	**70**	**71**	**72**	**73**	**74**	**75**	**76**	**77**	**78**	**79**	**80**
Ans.	4	1	1	4	2	1	3	1	3	2	1	2	4	2	2	1	2	1	4	4
Que.	**81**	**82**	**83**	**84**	**85**	**86**	**87**	**88**	**89**	**90**										
Ans.	2	2	3	4	3	1	4	4	2	4										

Exercise-II

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	4	1	3	4	4	3	2	3	4	4	4	1	2	2	3	2	1	1	4	2
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**	**31**	**32**	**33**	**34**	**35**	**36**	**37**	**38**	**39**	**40**
Ans.	3	1	3	4	1	3	4	2	2	4	1	1	4	2	4	1	4	1	4	2
Que.	**41**	**42**	**43**	**44**	**45**	**46**	**47**	**48**	**49**	**50**	**51**	**52**	**53**	**54**	**55**	**56**	**57**	**58**	**59**	**60**
Ans.	1	1	2	1	2	1	1	2	3	3	3	3	3	2	1	4	1	2	2	2
Que.	**61**	**62**	**63**	**64**	**65**	**66**	**67**	**68**	**69**	**70**	**71**	**72**	**73**	**74**	**75**	**76**	**77**	**78**	**79**	**80**
Ans.	2	2	2	2	3	3	2	2	4	3	2	1	2	4	2	4	2	2	3	3
Que.	**81**	**82**	**83**	**84**	**85**	**86**	**87**	**88**	**89**	**90**	**91**	**92**	**93**	**94**	**95**	**96**	**97**	**98**	**99**	**100**
Ans.	4	1	4	2	1	4	3	2	1	2	3	3	3	2	2	1	4	3	3	2
Que.	**101**	**102**	**103**	**104**	**105**	**106**	**107**	**108**	**109**	**110**	**111**	**112**	**113**	**114**	**115**					
Ans.	4	3	3	1	1	4	1	3	2	4	2	2	2	1	2					

Exercise-III

Que.	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**	**11**	**12**	**13**	**14**	**15**	**16**	**17**	**18**	**19**	**20**
Ans.	2	3	3	3	3	3	4	3	3	4	3	3	4	4	4	3	4	3	4	2
Que.	**21**	**22**	**23**	**24**	**25**	**26**	**27**	**28**	**29**	**30**										
Ans.	2	3	2	3	4	4	3	2	2	4										

SPACE FOR ROUGH WORK

SPACE FOR ROUGH WORK