TOPIC 7

Review of Some Fundamental Topics Previously Treated

LEARNING OBJECTIVE

Upon completion of this revision exercise, learners will:

- Learners demonstrate ability to review fundamental topics previously treated, including WASSCE past papers.

We have studied various topic in previous grades. Let's review some of the fundamental topics briefly.

7.1. DEVELOPMENT OF CHEMISTRY

The ancient Greeks first proposed that matter consisted of fundamental particles called *atoms*. Chemistry took its present scientific form in the 18th century, when careful quantitative experiments by Lavoisier, Proust, and Dalton resulted in the law of definite proportions, the law of conservation of mass, and the law of multiple proportions, which laid the groundwork for Dalton's atomic theory of matter. In particular, Avogadro's hypothesis provided the first link between the macroscopic properties of a substance (in this case, the volume of a gas) and the number of atoms or molecules present.

There were many persons who contributed to the development of chemistry. The following are some major contributors of chemistry:

1. **Democritus & Leucippus**: They suggested that matter was not infinitely divisible into smaller particles but instead consisted of fundamental, indivisible particles called **atoms**.

2. **Robert Boyle:** He formulated the Boyles's law which describes the relationship between pressure and volume of air.

3. **Antoine Lavoisier**: His most important contribution was the 'law of conservation of mass.'

4. **J. L. Proust**: He stated the 'law of definite proportions.'
5. **John Dalton:** Wrote atomic theory of matter.
6. **Joseph Gay Lussac**: He formulated the Gay-Lussac's Law
7. **Amedeo Avogadro**: He formulated Avogadro's hypothesis.

7.2. UNIT OF MEASUREMENT

Measurement of any physical quantity involves comparison with a certain basic, arbitrarily chosen, internationally accepted reference standard called **unit**. The result of a measurement of a physical quantity is expressed by a number (or numerical measure) accompanied by a unit. The units for the fundamental or base quantities are called **fundamental** or **base units**.

There are seven base quantities. Below are some base quantities along with their units:

- Length – metre
- Time – second
- Thermodynamic Temperature – kelvin
- Amount of substance – mole
- Mass kologram
- Electric current – ampere
- Luminous intensity – candela

The units of all other physical quantities can be expressed as combinations of the base units. Such units obtained for the derived quantities are called **derived units**. Below are some derived quantities along with their units:

- Area – square metre (m^2)
- Density – kg/m^3
- Volume – cubic metre (m^3)
- Electric charge – coulomb (C)

A complete set of units, both fundamental and derived, is called a **system of units**.

The **International System of Units (SI)** based on seven base units is at present internationally accepted unit system and is widely used throughout the world.

The SI units are used in all physical measurements, for both the base quantities and the derived quantities obtained from them. Certain derived units are expressed by means of SI units with special names (such as joule, Newton, watt, etc.).

The SI units have well defined and internationally accepted unit symbols (such as m for metre, kg for kilogram, s for second, A for ampere, N for newton, etc.).

Physical measurements are usually expressed for small and large quantities in scientific notation, with powers of 10. Scientific notation and the prefixes are used to simplify measurement notation and numerical computation, giving indication to the precision of the numbers.

Certain general rules and guidelines must be followed for using notations for physical quantities and standard symbols for SI units, some other units and SI prefixes for expressing properly the physical quantities and measurements.

In computing any physical quantity, the units for derived quantities involved in the relationship(s) are treated as though they were algebraic quantities till the desired units are obtained.

Direct and indirect methods can be used for the measurement of physical quantities. In measured quantities, while expressing the result, the accuracy and precision of measuring instruments along with errors in measurements should be taken into account.

In measured and computed quantities proper significant figures only should be retained. Rules for determining the number of significant figures, carrying out arithmetic operations with them, and rounding off the uncertain digits must be followed.

7.3. MATTER AND ITS PROPERTIES

Matter is something that occupies space and has mass. All the materials present in this world such as air and water, hydrogen and oxygen, sand, coal, iron, milk and oil, rocks and minerals are different kinds of matter.

Properties of Matter

Properties of a matter can be classified in two groups – (i) Physical properties and (ii) Chemical properties.

Physical properties of matter

(i) All matter is made up of tiny particles called atoms, and each atom have mass and have some volumes.

(ii) A physical property is an attribute of matter that is independent of its chemical composition.

(iii) Mass, volume, length, shape, density, colour, hardness, melting, boiling points, temperature, pressure, electrical conductivity, and thermal conductivity are all examples of physical properties.

Chemical properties of matter

(i) *Reactivity*: The tendency of matter to combine chemically with other substances is known as reactivity. Certain materials like potassium are highly reactive, whereas others are extremely inactive.

(ii) *Acidity*: A substance's ability to react with an acid is a chemical property. Some metals form compounds when they react with different acids. Acids react with bases to create water, which neutralizes the acid.

(iii) *Flammability*: The tendency of matter to burn is referred to as flammability. As matter burns, it reacts with oxygen and transforms into various substances. Wood is a flammable matter.

7.4. ATOMIC STRUCTURE

The atomic structure of an element refers to the constitution of its nucleus and the arrangement of the electrons around it. Primarily, the atomic structure of matter is made up of protons, electrons and neutrons.

The protons and neutrons make up the nucleus of the atom, which is surrounded by the electrons belonging to the atom. The atomic number of an element describes the total number of protons in its nucleus.
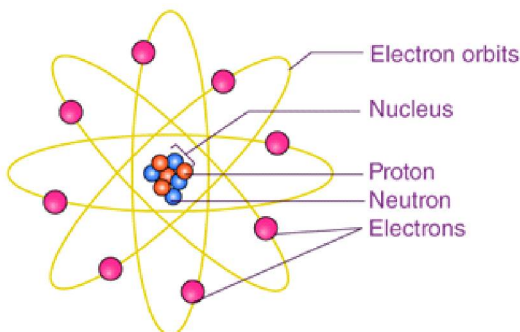

Fig. 7.1 The Structure of an Atom

- J.J. Thomson, a British physicist, discovered electron.
- E. Goldstein, a German scientist, in 1886, discovered proton.
- Ernest Rutherford, a New Zealand physicist, discovered nucleus in 1911.
- Neutrons were discovered experimentally by James Chadwick in 1932.

Neutral atoms have equal numbers of protons and electrons. However, atoms may gain or lose electrons in order to increase their stability and the resulting charged entity is called an ion.

7.4.1. Characteristics of the Sub-Atomic Particles

The characteristic properties of sub-atomic particles are summarized in the following table:

	Electron	Proton	Neutron
Charge	-1.602×10^{-19} C	1.602×10^{-19} C	0
Mass	9.110×10^{-31} kg	1.673×10^{-27} kg	1.675×10^{-27} kg
Relative Mass	1/1840	1	1

7.4.2. Bohr's Model of Hydrogen Atom

Neils Bohr put forth his model of the atom in the year 1915. This is the most widely used atomic model to describe the atomic structure of an element.

Postulates of Bohr's Model:

- The electrons in the hydrogen atom revolve around the nucleus only in certain selected circular orbits. These orbits are associated with definite energies and are called energy shells or energy levels. These are designated as K, L, M, N, ..., etc. shells.
- As long as the electron remains in a particular orbit, it does not lose or gain energy. This means that energy of an electron in a particular orbit remains constant. That is why, these orbits are also called stationary states.
- Electrons can jump to higher levels by absorbing energy and move to lower energy levels by losing or emitting its energy.

7.5. PERIODIC TABLE OR CHEMISTRY

Before the beginning of eighteenth century, only about 30 elements were known and it was quite easy to study and remember their individual properties. However, the situation became difficult with the discovery of large number of elements in the later years. At this stage, the scientists felt the need of some simple method to facilitate the study of the properties of various elements and their compounds. After numerous attempts the scientists were ultimately successful in arranging the elements in such a way so that similar elements were grouped together and different elements were separated. This arrangement of elements is known as classification of elements and it led to the formulation of periodic table. Thus, *periodic table may be defined as the table giving the arrangement of all the known elements according to their properties so that similar elements fall within the same vertical column and dissimilar elements are separated*.

Earlier attempts on classification of elements were based on atomic masses. The formulation of a satisfactory periodic law took place only after 1860.

- **Dobereiner's triad** consider three elements, in which atomic mass of central element is the arithmetic mean to two other elements. For Example, Atomic masses of lithium, sodium and potassium are 7, 23 and 39. The arithmetic mean of 7 and 39 gives 23. But disadvantage is the presence of only few triads.

- **Newland's Law of Octave** say that *elements are arranged in such a way that every eight element has same properties as in the first element*. According to him, only 56 elements are found which is a drawback of Newland Law of Octaves. This was not accepted. It is acceptable only up to calcium. He also placed some unlike elements in the same slot.

- **Lothar Meyer's Curves**: In 1870, the German chemist Julius Lothar Meyer plotted the atomic volumes (i.e., the atomic mass divided by density) of the elements against their atomic masses. From his graph, Lothar Meyer was able to produce a table showing periodic arrangement of elements.

- **Mendeleev** used atomic masses as the basis of arrangement of elements. According to him, elements were arranged in increasing order of their atomic masses. It believes that there was a periodic reappearance in their physical and chemical properties.

Group	I-A	I-B	II-A	II-B	III-A	III-B	IV-A	IV-B	V-A	V-B	VI-A	VI-B	VII-A	VII-B	VIII
Oxide:	R_2O		RO		R_2O_3		RO_2		R_2O_5		RO_3		R_2O_7		RO_4
Hydride:	RH		RH_2		RH_3		RH_4		RH_3		RH_2		RH		
	A	B	A	B	A	B	A	B	A	B	A	B	A	B	
Period 1	H														
Period 2	Li		Be		B		C		N		O		F		
Period 3	Na		Mg		Al		Si		P		S		Cl		
Period 4	K		Ca			*		Ti		V		Cr		Mn	Fe Co Ni
		Cu		Zn	*		*		As		Se		Br		
Period 5	Rb		Sr			Y		Zr		Nb		Mo		Tc	Ru Rh Pd
		Ag		Cd	In		Sn		Sb		Te		I		
Period 6	Cs		Ba			La		Hf		Ta		W			Os Ir Pt
		Au		Hg	Ti		Pb		Bi						

Fig.7.2 Mendeleev's Periodic Table

Limitations of Mendeleev Periodic Table

⇒ Position of isotopes cannot be explained

⇒ Position of hydrogen is not fixed. It is placed in group 1A, though its some properties matches with those of halogens.

- **Modern Periodic Table:** According to modern periodic law, "properties of an element are the periodic function of their increasing atomic number". There are 18 groups and 7 groups.

Fig.7.3 Modern Periodic Table

Trends in Modern Periodic Table:

- *Valency* - On moving left to right in a period, valency increases and then it decreases. But it remains same down in a group.
- *Atomic Size* - It decreases left to right in a period as the nuclear charge increases due to large positive charges on the nucleus. Atomic size increases down in a group due to decrease in nuclear charges and addition of new shell.
- *Metallic Character* - Metallic character decreases from left to right in a period. But non-metallic character increases left to right in a period. And metallic character increases down the group as the size increases it can easily lose electron.
- *Ionization Energy* - It is the energy required to remove an electron from an isolated gaseous atom. Ionization energy increases as we move left to right in a period. This is due to increase in nuclear charge as we move left to right in a period. But down in a group ionization energy decreases due to decrease in nuclear charge but there are some exceptional cases.

7.6. CHEMICAL BONDING

*The attractive force which holds various constituents (atoms, ions, etc.) together in different chemical species is called a **chemical bond**.*

The strength and properties of the chemical bonds formed vary. There are four primary types of chemical bonding. These are:

Ionic Bonding

The forces which hold the ions together in an ionic compound are known as **ionic bonds or electrovalent bonds.** Examples of some ionic compounds incude: Aluminium oxide (Al_2O_3), Sodium hydroxide (NaOH), Common salt (CaCl), etc.

Formation of Ionic Bonding

Ionic bonding occurs between a metal and a non-metal. Metal atoms lose their outermost electron(s), forming cations. Non-metal atoms gain electron(s) to fill their outermost shell, forming anions. The electrostatic force of attraction between the oppositely charged ions holds the ions together. For example, **formation of sodium chloride (NaCl)**.

When a hot sodium atom is placed in chlorine gas, a reaction takes place resulting in formation of sodium chloride.

$$Na \longrightarrow Na^+ + e^-$$
$$2, 8, 1 \qquad\qquad 2, 8$$

Sodium ion
(Cation)

$$Cl + e^- \longrightarrow Cl^-$$
$$2, 8, 7 \qquad\qquad 2, 8, 8$$

Chloride ion
(Anion)

$$Na + \cdot \ddot{\underset{\cdot\cdot}{Cl}} : \longrightarrow (Na^+) \left[: \ddot{\underset{\cdot\cdot}{Cl}} : \right]$$
$$2, 8, 1$$
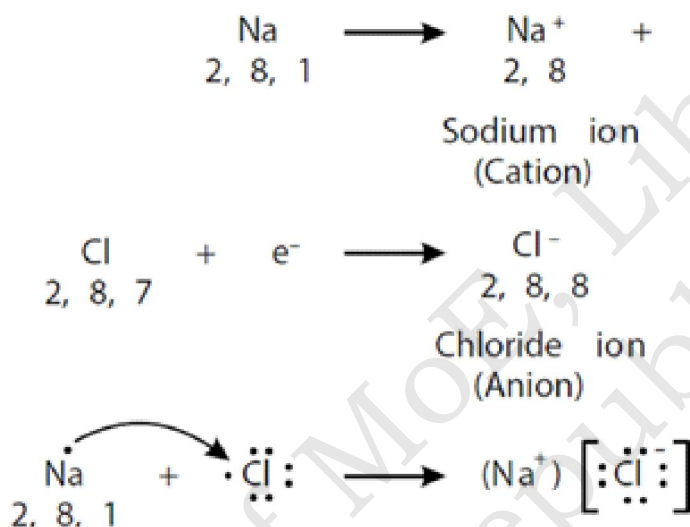
Properties of Ionic Bonding

Some general properties for ionic compounds are:

- Ionic compounds are usually crystalline solids.
- Ionic compounds have high melting and high boiling points.
- Ionic compounds are usually soluble in water but insoluble in organic solvents like petrol and kerosene.
- Ionic compounds conduct electricity when dissolved in water or when melted.

Covalent Bonding

The chemical bond formed by sharing of electrons between two atoms is known as a **covalent bond**. The compounds containing covalent bonds are known as **covalent compounds**. Examples of some covalent compounds include: methane (CH_4), Carbon dioxide (CO_2), Ammonia (NH_3), etc.

Aluminium oxide (Al_2O_3), Sodium hydroxide (NaOH), Common salt (CaCl), etc.

Formation of Covalent Bonding

Covalent bonding between atoms of different elements of non-metals. For example, **formation of methane**.

Carbon atom shares four electrons to form methane.
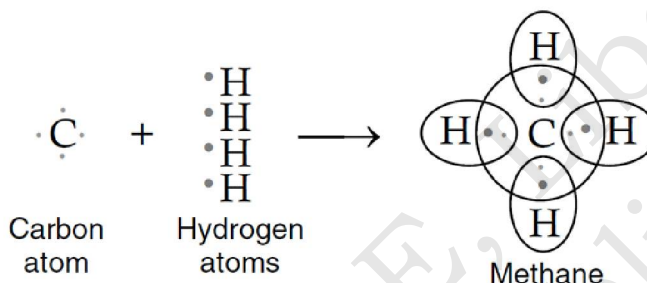
Properties of Covalent Bonding

Some general properties for covalent compounds are:

- Covalent compounds are usually liquids, gases or solids. For example, alcohol, benzene, water and cooking oil are liquids.
- Covalent compounds have usually low melting points and low boiling points.
- Covalent compounds are usually insoluble in water, but they are soluble in organic solvents.

Metallic Bonding

The force which binds various metal atoms together is called **metallic bond**.

Formation of Metallic Bonding

Loreutz proposed the theory of electron gas model or electron sea model for metallic bonding. In this model, the metal is pictured as an array of metal cations in a "sea" of electrons. The atoms in a metallic solid contribute their valence electrons to form a "sea" of electrons that surrounds metallic cations. Delocalised electrons are not held by any specific atom and can move easily throughout the solid. A metallic bond is the attraction between these electrons and the metallic cation.

Properties of Metallic Bonding

Some general properties of metals are:

Metals are good conductors of heat and electricity.

- Metals are lustrous (or shiny).
- Metals are usually strong.
- Metals are ductile.
- Metals are malleable.

## 7.7. CHEMICAL REACTIONS/EQUATIONS AND STOICHIOMETRY

Chemical Reactions

The transformation of chemical substance into another chemical substance is known as **Chemical Reaction**. For example: Rusting of iron, the setting of milk into curd, digestion of food, respiration, etc.

The substances which take part in a chemical reaction are called **reactants**.

The new substances produced as a result of a chemical reaction are called **products**.

Example: The burning of magnesium in the air to form magnesium oxide is an example of a chemical reaction.

$$2Mg(s) + O_2(g) \xrightarrow{\triangle} 2MgO(s)$$
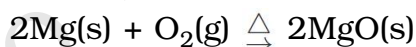
Before burning in air, the magnesium ribbon is cleaned by rubbing with sandpaper. This is done to remove the protective layer of basic magnesium carbonate from the surface of the magnesium ribbon.

Here the reactants are Mg and O_2, while MgO is the product.

Kinds of Chemical Reactions

Generally, chemical reactions are of two types:

 (i) **Reversible reaction**: The reactions in which products under suitable conditions react to give back reactants are called *reversible reactions*.

(ii) **Irreversible reaction**: The reactions in which the products do not react under any condition to give back reactants are called *irreversible reactions*.

Types of Chemical Reactions

Chemical reactions can be classified in following types:

(i) **Combination Reaction**: Reactions in which two or more reactants combine to form one product are called Combination Reactions. A general combination reaction can be represented by the chemical equation given here:

Examples:

When magnesium is burnt in the air (oxygen), magnesium oxide is formed. In this reaction, magnesium is combined with oxygen.

$$Mg(s) \quad + \quad O_2(g) \rightarrow 2\ MgO(s)$$

Magnesium   Oxygen   Magnesium Oxide
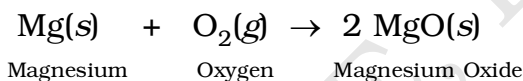
(ii) **Single Replacement Reactions**: Single replacement reactions are those reactions in which one element displaces another element from a compound.

Example:

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

Copper Sulphate   Iron   Iron sulphate   Copper
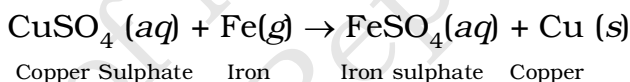
(iii) **Double Replacement Reactions**: Double displacement reactions are those reactions in which two compounds react by exchange of ions to form two new compounds. These reactions are also known as metathesis reactions or exchange reactions.

Example:

$$BaCl_2(aq) \quad + \quad CuSO_4(aq) \quad \rightarrow \quad BaSO_4(s) \quad + \quad CuCl_2(aq)$$

Barium chloride   Copper sulphate   Barium sulphate   Copper chloride
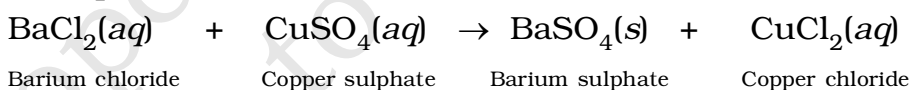
(iv) **Decomposition reactions**: Decomposition reactions are those reactions in which a single reactant breaks down into two (or more) simpler substances (elements or compounds).

Example:

$$2AgCl(s) \xrightarrow{\text{sunlight}} 2Ag(s) + Cl_2(g)$$
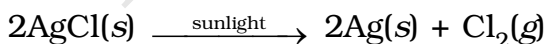
(v) **Combustion Reactions**: This type of reaction refers to the reaction of an element or compound with oxygen. Combustion usually releases a lot of heat energy. It is also referred to as burning.

Example: Burning of coal

$$C(s) + O_2(g) \rightarrow CO_2(g)$$
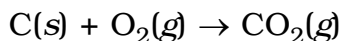
Chemical Equations

Representation of chemical reaction using symbols and formulae of the substance is called *Chemical Equation*.

Example:

$$H_2 \quad + \quad O_2 \quad \rightarrow \quad H_2O$$

Hydrogen   Oxygen   Water

In this equation, H_2 and O_2 are called reactants and H_2O is called the product. The arrow shows the direction of the chemical reaction. Condition, if any, is written generally above the arrow.

Types of Chemical Equation

A chemical equation can be divided into two types:

(i) **Balanced Chemical Equation**: A balanced chemical equation has the number of atoms of each element equal on both sides.

Example:

$$2H_2 \quad + \quad O_2 \quad \rightarrow \quad 2H_2O$$

Hydrogen   Oxygen   Water

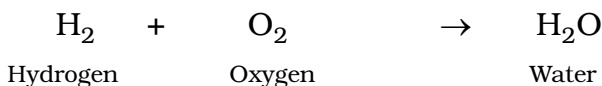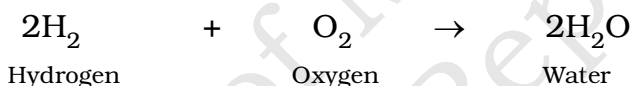In this equation, numbers of atoms of each element in the reactants is equal to the number of atoms of each element present in the product on both sides, so it is a Balanced Chemical Equation.

(ii) **Unbalanced Chemical Equation**: If the number of atoms of each element in reactants is not equal to the number of atoms of each element present in the product, then the chemical equation is called Unbalanced Chemical Equation.

Example:

$$H_2 \quad + \quad O_2 \quad \rightarrow \quad H_2O$$

Hydrogen   Oxygen   Water
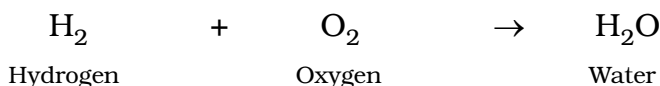
7.8. OXIDATION-REDUCTION REACTION

According to **classical concept** following definitions were proposed to explain the process of oxidation and reduction.

Oxidation is a process of chemical addition of oxygen or any electronegative radical or removal of hydrogen or any electropositive radical.

   Reduction is a process of chemical addition of hydrogen or any electropositive radical or removal of oxygen or any electronegative radical.

   Example: Reaction of H_2S and Cl_2

$$\begin{array}{c}
\overset{\displaystyle \overset{\text{Oxidation}}{\longrightarrow}}{Cl_2 + H_2O \longrightarrow 2HCl + S} \\
\underset{\text{Reduction}}{\longrightarrow}
\end{array}$$

   Here, hydrogen is being removed hydrogen sulphide (H_2S) and is being added to chlorine (Cl_2). Thus, H_2S is oxidised and Cl_2 is reduced.
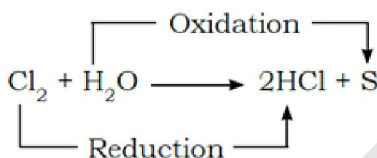
   According to **electronic concept**, the oxidation and reduction processes can be defined as follows oxidation.

- **Oxidation** is a process in which an atom or a group of atoms taking part in chemical reaction loses one or more electrons. The loss of electrons results in the increase of positive charge or decrease of negative charge of the species.

$$2Cl^- \quad \rightarrow \quad Cl_2 \quad + \quad e^-$$

   Chloride ion   Chlorine molecule

- The species which undergo the loss of electrons during the reactions are called **reducing agents** or reductants. Cl^-, Fe^{2+} and Cu are reducing agents in the above examples.

- **Reduction** is a process in which an atom or a group of atoms taking part in chemical reaction gains one or more electrons. The gain of electrons results in the decrease of positive charge or increase of negative charge of the species.

$$Ag^+ \quad + \quad e^- \quad \rightarrow \quad Ag$$
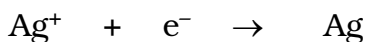
   (*Decrease of positive charge*)

- The chemical reactions involving reduction-oxidation are called *redox reactions*.

- The species which undergo gain of electrons during the reactions are called **oxidising agents or oxidants**. In the above examples, Ag^+, Fe^{3+} ions, Br_2 molecule are oxidising agents.

7.9. STATES OF MATTER

Matter can be classified into three categories depending upon its physical state, namely: solids, liquids and gases.

1. **Solids**: The matter in solid state possesses a definite volume, a definite shape and a definite mass. Solids are rigid and almost incompressible. Some common examples are: table, chair, common salt, silver, ice, diamond, etc.

2. **Liquids**: The matter in liquid state possesses a definite volume, a definite mass, but no definite shape. Liquids are also almost incompressible but are not rigid. In fact, they acquire the shape of the container. Some examples are: milk, water, alcohol, etc.

3. **Gases**: The matter in gaseous state neither has definite volume nor definite shape but it has definite mass. It acquires the shape and volume of the container. Some common examples are: air, oxygen, hydrogen, sulphur dioxide, etc. Gases are highly compressible. For example, natural gas in compressed form is used as fuel (Compressed Natural Gas—CNG) in internal combustion engines.

Change of States of Matter

Under specific conditions, matter changes its state. For example, when ice is heated, it converts into water (solid to liquid change). When water is heated, it converts into steam (liquid to gas change). The steam when cooled, converts into liquid (gas to liquid change). Thus, depending on the temperature, a substance can exist in all the three physical states—solid, liquid, and gas.
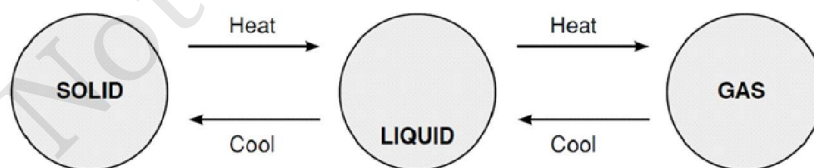

Fig.7.4

7.10. SOLUTION

A solution is a homogeneous mixture of two or where pure substances. Solution retains the properties of with solvent and solute even if the solution is homogeneous.

A solution is made up of two components: i.e. a *solute* and a *solvent*. **Solvent** is the component of solution present in large amount and its physical state is same as that of the resulting solution.

Solute is the component of solution present in lesser amount than that of solvent and its physical state may or may not same as that of the resulting solution.

Concertration of Solutions

Concentration of solution refers to *the amount of solute present in the given quantity of solution or solvent.* The concentration of the solution may be expressed in any of the following ways:

1. Mass Percentage (% mass)

Mass percentage may be defined as the *number of parts by mass of solute per hundred parts by mass of solution.* For example, a 5% (aqueous) solution of sugar by mass means that per 100 g of solution contain 5 g of sugar.

Let WB be the mass of solute (B) and WA be the mass of solvent (A), then

$$\text{Mass percentage of B} = \frac{W_B}{W_A + W_B} \times 100$$

2. Volume Percentage (% volume)

This mode of concentration is used in case of solutions when *solutes* and *solvents* are both *liquids*. Volume percentage may be defined as *the number of parts by volume of solute per hundred parts by volume of solution.* For example, a 25% solution of ethyl alcohol (by volume) means that 100 cm^3 of the solution contain 25 cm^3 of ethyl alcohol and 75 cm^3 of water.

Let VA and VB be the volumes of component A and B, then

Volume percentage of B

$$\frac{V_B}{V_A + V_B} \times 100$$

3. Normality (N)

Normality of a solution is defined as *the number of gram-equivalents of the solute present per litre or per cubic decimeter of the solution. It may also be expressed in terms of milliequivalents of solute present in one mililitre of solution.* It is represented by N.

Mathematically,

$$\text{Normality (N)} = \frac{\text{Gram-equivalents of solute (B)}}{\text{Volume of solution in litre}}$$

Gram-equivalents of solute (B) represents the ratio of its mass in gram (W_B) to its equivalent mass (E_B). Also, the ratio of the mass of solute in gram (W_B) to the volume of the solution in litre (V_L) represents the **strength** of the solution. Thus,

$$\textbf{Normality (N)} = \frac{W_B(g)}{E_B \times V_{(L)}}$$

4. Molarity (M)

Molarity of a solution is defined as *the number of gram mole of the solute present per litre or per cubic decimeter of the solution.* It may also be expressed in terms of millimol of solute present per mililitre of solution. It is represented by M.

Mathematically,

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Volume of solution in litres}}$$

$$= \frac{\text{Mass of solute in gram}}{\text{Molar mass of solute} \times \text{Volume of solution (L)}}$$

Now, moles of the solute (n_B) represents the ratio of mass of solute in gram (W_B) to its molar mass (M_B). Thus, molarity can be represented by the expressions.

$$M = \frac{n_B}{V_{(L)}} \text{ or } \frac{W_B(g)}{M_B \times V_{(L)}}$$

$$\text{or } \frac{\text{Strength (gL}^{-1})}{M_B} \text{ or } \frac{W_B(g) \times 1000}{M_B \times V_{(mL)}}$$

5. Molality (m)

Molality of a solution may be defined as *the number of gram mole of the solute present per kilogram (1000 g) of the solvent.* It is represented by *m*.

Mathematically, Molality (m)

$$= \frac{\text{Mole of solute}}{\text{Mass of solvent (kg)}}$$

$$= \frac{\text{Mass of solute (g)}}{(\text{Molar mass of solute}) \times (\text{Mass of solvent (kg)})}$$

or $\dfrac{n_B}{W_A(\text{kg})}$ or $\dfrac{W_B(\text{g})}{M_B \times W_A\,(\text{kg})}$

or $\dfrac{W_B(\text{g}) \times 1000}{M_B \times W_A\,(\text{g})}$

Some Important Formulae

Some very significant formulae about different modes of concentrations of solution are being given below for making the calculations easy.

1. Relationship between Molarity and Normality. For a solution of molarity M, the mass of solute per litre of the solution is M × (molar mass) g. The number of gram equivalents of solute per litre

of the solution, *i.e.*, normality (N) is given as mass $\dfrac{M \times (\text{molar mass})}{\text{Equivalent mass}}$.

Thus, N = $\dfrac{M \times (\text{molar mass})}{\text{Equivalent mass}}$

4" **N × Equivalent mass = M × Molar mass**

For **Acids;** *Normality = Molarity × Basicity*

For **Bases;** *Normality = Molarity × Acidity*

Basicity is the number of H^+ ions furnished by each molecule of acid in solution.

Acidity is the number of OH^- ions furnished by each molecule of base or number base or number of H^+ ion which react with one molecule of base in solution.

2. Relationship between Molarity (M) and Mass Percentage (%)

For a solution of $p\%$ by mass, p g of solute is present in 100 g of solution. If d is the density of solution, then volume of the solution is $100/d\ cm^3$ Thus, molarity is given as

Molarity (M) = $\dfrac{p(g) \times 1000}{M_B \times 100/d}$

or $\mathbf{M} = \dfrac{\%\,(\mathbf{mass}) \times \mathbf{10} \times \mathbf{d}}{\mathbf{M_B}}$

3. Relationship between Molarity and Molality

For a solution of molarity M, the mass of solute per litre of the solution is $(M \times M_B)$ g

Mass of 1000 cm3 of solution is $(1000 \times d)$ g

Mass of solvent is $(1000\ d - M \times MB)$ g

Molality (m) of solution = $\dfrac{\text{Moles of solute}}{\text{Mass of solvent (kg)}}$

or $\mathbf{m} = \dfrac{\mathbf{M} \times \mathbf{1000}}{(\mathbf{1000}\,d - \mathbf{M} \times \mathbf{M_B})}$

4. Mole Fraction

Mole fraction may be defined as the *ratio of number of moles of one component to the total number of moles of all the components (solvent and solute) present in the solution*. It is denoted by the letter x followed by the subscript representing the component. Let us suppose that a solution contains the components A and B and suppose that WA g of A and WB g of B are present in it.

Moles of A (n_A) = $\dfrac{W_A}{M_A}$;

Moles of B (n_B) = $\dfrac{W_B}{M_B}$

M_A and M_B are molar masses of A and B respectively .

Total number of moles of A and B = $n_A + n_B$

Mole fraction of A, $(\boldsymbol{x_A})$ = $\dfrac{\mathbf{n_A}}{\mathbf{n_A} + \mathbf{n_B}}$

Mole fraction of B, $(x\text{B}) = \dfrac{n_B}{n_A + n_B}$

The sum of the mole fractions of solute and solvent in binary solution is **unity** *as shown below.*

$$x_A + x_B = \dfrac{n_A}{n_A + n_B} + \dfrac{n_B}{n_A + n_B} = 1.$$

• **Relationship between Molality (m) and Mole fraction of Solute (x_B)**

$$x_B = \dfrac{m \times M_A}{1000 + m \times M_A}$$

Also $m = \dfrac{1000 x_B}{x_A \times M_A}$

• **Relationship between Molarity (M) and Mole fraction of Solute (x_B)**

$$x_B = \dfrac{M \times M_A}{M(M_A - M_B) + 1000 d}$$

Also $M = \dfrac{1000\, d\, x_B}{x_A \times M_A + x_B \times M_B}$

In the above expressions, the subscript A refers to solvent and B refers to solute.

7.11. ACIDS, BASES AND SALTS

Electrolytes may be acids, bases or salts.

1. **Acids**: Acids are sour in taste. Vinegar contains acetic acid whereas citrus fruits such as orange, lemon, etc., contain citric acid. Tamarind contains tartaric acid. Acids turn blue litmus paper red.

2. **Bases**: Bases are bitter in taste. They turn red litmus paper blue. Caustic soda (NaOH), washing soda ($Na_2CO_3.10H_2O$) and aqueous ammonia are commonly used bases.

3. **Salt**: Salts are made up of oppositely charged ions. They form ions. Sodium chloride (NaCl), Calcium chloride ($CaCl_2$) and Potassium dichromate ($K_2Cr_2O_7$) are commonly used salts.

Acid-Base Titrations

It is a method of quantitative analysis for determining the concentration of an acid or base by exactly neutralizing it with a standard solution of acid or base having known concentration. The progress of titration (process of neutralization) is monitored by using pH indicator.

The acid-base titration is useful in case of strong or weak acid or base.

In titration acid is treated with base and a base is treated with acid. **End point of titration** is decided by an indicator.

End point of titration is a stage that shows the completion of reaction. The end point is detected by a physical change that is either produced by solution itself or by addition of a reagent known as indicator.

The end point or the equivalence point reaches when the stoichiometric amount of acid has been added to the alkali solution. At this point all the alkali has been neutralised and the solution contains salt and water only.

From the end point we can find the volume of the base (V_b) having molarity M_b required to neutralize a certain volume of the acid (V_a) of molarity M_a. Now, number of moles in V mL of a solution of molarity M is given by M × V/1000. For the titrations involving monobasic acid and monoacid base, one mole of base is required to completely neutralize one mole of the acid.

Therefore,

$$\frac{M_a V_a}{1000} = \frac{M_b V_b}{1000} \text{ or } M_a V_a = M_b V_b$$

7.12. SOLUBILITY OF SUBSTANCES

Solubility is defined as the amount of solute that can be dissolved in 100 gram of the solvent at the given conditions. Another common way of expressing solubility of a substance at a given temperature is maximum quantity of solute in moles (or gram) that can dissolve in solvent to form 1 dm^3 of solution. The units of solubility are mol dm^{-3} or g dm^{-3} or g per 100 g of solvent. The solubility of gases in water is also expressed in terms of its mole fraction. Solubility depends on nature of

solute, and solvent as well as the conditions of temperature and pressure.

Solubility of Solid in liquid

When a solid solute is added to solvent, dissolution of solute occurs. The solute particles go into solution and acquire random movement. The concentration of the solution gradually increases as more and more solute dissolves.

A solution in which no more solute can be dissolved in the solvent is called **saturated solution.** The amount of solutes dissolved in a solvent is known as **solubility.** *The* **solubility** *of solids is generally is expressed as the amount of solid that can be dissolved in 100 grams of the solvent at a given temperature.*

7.13. CHEMICAL ENERGETICS

Chemical energy is the energy stored in bonds of a chemical compound. This chemical energy may be released during a chemical reaction as in exothermic reactions. Sometimes energy input is required in chemical energy as in endothermic reactions.

Examples of stored chemical energies are, batteries, biomass, natural gas, petroleum and coal.

Systems & Surroundings

One of the most important aspects of chemical energetics is to define the context for studying the changes in energy exchange. The context may be defined in terms of –

- **System**: The object under observation.
- **Surroundings**: Everything other than the system.

Kinetics & Thermodynamics

A reaction can be fundamentally studied according to two viewpoints.

Kinetics

Kinetics deal with the questions regarding the reaction rate and the intermediate steps involved in going from initial reactant(s) to final product(s).

The Laws of Thermodynamics

A. The 1st law of Thermodynamics

The first law of thermodynamics states that the change in internal energy of a system equals the net heat transfer into the system minus the net work done by the system.

$$\Delta U = Q - W$$

ΔU is the change in internal energy U of the system. Q is the net heat transferred into the system—that is, Q is the sum of all heat transferred into and out of the system. W is the net work done by the system.

B. The 2nd Law of Thermodynamics

Heat transfer occurs spontaneously from higher-to lower-temperature bodies but never spontaneously in the reverse direction.

Alternatively, it is impossible for heat transfer from a reservoir to completely convert to work in a cyclical process in which the system returns to its initial state.

In simpler terms, the second law can also be stated as "*all spontaneous changes cause an increase in the universe's entropy,*" wherein the term entropy means the randomness. Essentially, the transfer of energy from the system to the surroundings makes the components in the latter more energetic and hence more random.

C. The 3rd law of thermodynamics

The third law states the condition wherein there would be no randomness of motion of the components of a system. Hence, the entropy of a pure, perfect crystalline substance at 0 K is zero.

This is an ideal condition as the temperature of 0 Kelvin signifies no temperature and hence, no heat in the system to randomize the movement of the components.

Enthalpy – the total heat content of a system

The sum of a system's internal energy and the mathematical product of its pressure and volume is known as enthalpy (H). It represents the total heat content of a system in its normal state.

Hess's Law of Constant Heat Summation

The exchange of energy during chemical reactions can be used to find out the enthalpy of a particular step in a multiple-step reaction. Since

enthalpy is a state function, this is an obvious outcome. However, the method has become popular as Hess's law of constant heat summation. It can get stated formally as follows – The standard enthalpy of a reaction is equal to the standard enthalpies of all the steps if the reaction were to occur in multiple stages.

Fig. 7.5
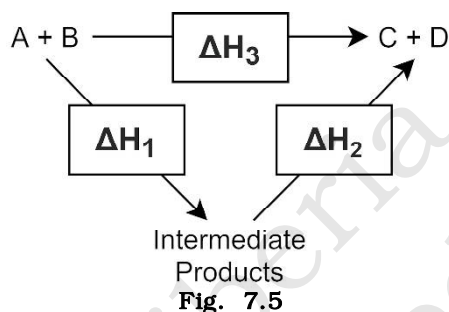
The important thing to note in this case is that all the steps must occur under identical temperature and pressure. The requirement for every step to be conducted in a laboratory, however, does not apply and the steps may be simply theoretical. The benefit of this process is in the theoretical estimation of the standard enthalpy of a reaction if the data for the standard enthalpy of reaction for the individual steps are known. Usually, these steps are very simple reactions involving simple molecules.

7.14. NUCLEAR CHEMISTRY

The spontaneous emission of active radiations by certain elements like uranium is called radioactivity and the elements are called **radioactive elements**.

The radioactive rays emitted when passed through strong electric on magnetic fields, are resolved into:

- **α rays**: The rays which deflected slightly towards negative plate were named as a rays.
- **β rays**: The rays which deflected towards a positive plate were named as â rays.
- **ψ rays**: The rays which remained undeflected were named as ã rays.

Half-life period is defined as the time required for the disappearance of half of the amount of the radioactive substance originally present.

7.15. CHEMICAL EQUILIBRIUM

Equilibrium is the state of a process in which the properties like temperature, pressure, and concentration etc of the system do not show any change with passage of time.

In all processes which attain equilibrium, two opposing processes are involved. Equilibrium is attained when the rates of the two opposing processes become equal.

There are two types of equilibrium. These are: physical equilibrium and chemical equilibrium.

Physical Equilibrium

If the opposing processes involve only physical changes, the equilibrium is called *Physical Equilibrium*. Following are some physical equilibrium.

- **Solid–liquid Equilibrium**. *Example*: $H_2O(s) \rightleftharpoons H_2O(l)$
- **Liquid–Gas Equilibrium**. *Example*: $H_2O(l) \rightleftharpoons H_2O(g)$
- **Solid–Solution Equilibrium**. *Example*: Salt (Solid)   Salt (in solution)
- **Gas–Solution equilibrium**. *Example*: $CO2(g) \rightleftharpoons CO_2$(in solution)

Chemical Equilibrium

If the opposing processes are chemical reactions, the equilibrium is called *Chemical Equilibrium*.

- **Reversible reaction:** A reaction in which not only the reactants react to form the products under certain conditions but also the products react to form reactants under the same conditions

  Examples: $3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g)$
- **Irreversible reaction:** A reaction cannot take place in the reverse direction, i.e. the products formed do not react to give back the reactants under the same condition.

  Example: $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(g)$

  Generally, a chemical equilibrium is represented as aA+bB $\rightleftharpoons$ cC+b D where A, B are reactants and C, D are products.

7.16. REACTION RATES

The *rate of reaction* or *reaction rate* is the speed at which reactants are converted into products. When we talk about chemical reactions, it is a given fact that rate at which they occur varies by a great deal. Some chemical reactions are nearly instantaneous, while others usually take some time to reach the final equilibrium.

This article aims to help students learn about and understand what exactly is the rate of reaction for a given chemical compound.
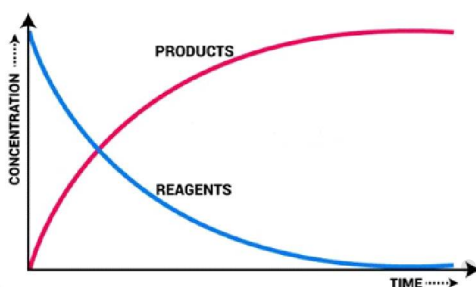
Rate of Reaction Formula

Let's take a traditional chemical reaction.

$$\alpha\,A + b\,B \rightarrow p\,P + q\,Q$$
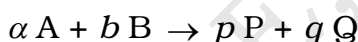
Capital letters (A and B) denote reactants and the (P and Q) denote products, while small letters (a,b,p,q) denote Stoichiometric coefficients.

As per IUPAC's Gold Book, the rate of reaction r occurring in a closed system without the formation of reaction intermediates under isochoric conditions is defined as:

$$r = -\frac{1}{a}\frac{d[A]}{dt} = -\frac{1}{b}\frac{d[B]}{dt} = \frac{1}{p}\frac{d[P]}{dt} = \frac{1}{q}\frac{d[Q]}{dt}$$

Here, the negative sign is used to indicate the decreasing concentration of the reactant.

7.17. ELECTROCHEMISTRY

- Electrochemistry is the branch of chemistry which deals with the relationship between electrical energy and chemical energy and inter-conversion of one form into another.
- An electrochemical cell consists of two metallic electrodes dipped in electrolytic solutions. The cells are of two types:
  (*a*) Electrolytic cells (*b*) Galvanic cells
- A galvanic cell consists of two half cells. Each half cell contains an electrolytic solution and a metallic electrode. The electrode at which- oxidation takes place is called an anode and the electrode at which reduction takes place is called the cathode. The half-cells are separated from each other by means of a porous pot or a salt bridge.

- The passage of current from one electrode to the other indicates the existence of potential difference between them. This difference of potential which causes current to flow from the electrode of higher negative potential is called the electromotive force (emf).

GLOSSARY

- **Accuracy**: The quality or state of being correct.
- **Chemistry:** The branch of science that studies the preparation, properties, structure and reactions of material substances.
- **Combustion:** The process where a substance burns in the presence of Oxygen, giving off heat and light in the process.
- **Isotopes:** Atom of an element having different mass numbers.
- **Group**: A vertical column of elements in the periodic table.
- **Reduction**: A process involving decrease in oxidation number by gain of electrons.
- **Solubility:** The amount of solute that can be dissolved in 100 gram of the solvent at the given conditions.

EXERCISE

A. Multiple Choice Questions

1. How many base units are there?
   - (a) Three
   - (b) Five
   - (c) Seven
   - (d) None of these

2. .... is something that occupies space and has mass.
   - (a) Matter
   - (b) Thought
   - (c) Love
   - (d) None of these

3. How many base units are there?
   - (a) Three
   - (b) Five
   - (c) Seven
   - (d) None of these

4. Who discovered electrons?
   - (a) E. Rutherford
   - (b) J.J. Thomson
   - (c) Neils Bohr
   - (d) None of these

5. Which of these is an ionic compound?
   - (a) NaCl
   - (b) H_2O

(c) NH_3 (d) CO_2

6. The process of reduction involves
   (a) addition of oxygen   (b) addition of hydrogen
   (c) removal of hydrogen   (d) None of these

7. Which of the following a base?
   (a) NaCl   (b) $CaCl_2$
   (c) NaOH   (d) $K_2Cr_2O_7$

8. This is the time required for the disappearance of half of the amount of the radioactive substance.
   (a) Life period   (b) Half-life period
   (c) Both (a) & (b)   (d) None of these.

B. Fill in the Blanks

1. A solution in which no more solute can be dissolved in the solvent is called .... solution.
2. .... is the energy stored in bonds of a chemical compound.
3. A reaction which cannot take place in the reverse direction is called the .... reaction.
4. .... is the speed at which reactants are converted into products.
5. .... is the branch of chemistry which deals with the relationship between electrical energy and chemical energy and inter-conversion of one form into another.

C. Answer the Following Questions

1. Define chemistry.
2. What are the seven base quantities? List them along with their units?
3. Who discovered nucleus of an atom?
4. What property did Mendeleev use to arrange elements in the periodic table?
5. What do you mean by chemical bonding? What are its different types?