Revision Notes

Class 11 Chemistry

Chapter 1 - Some Basic Concepts of Chemistry.

1. CHEMISTRY

Chemistry (derived from the Egyptian word kēme (chem), which means "earth") is a science that studies the composition, structure and properties of matter and the changes it undergoes during chemical reactions. Chemistry is often referred to as core science because it plays a role in linking physical sciences (including chemistry) with life sciences and applied sciences (such as medicine and engineering).

Chemistry is divided into following branches:

1.1 Physical chemistry

The branch of chemistry which deals with macroscopic as well as physical phenomena in a universe. It is generally the impact of physical property on the chemical property as well as structure of a substance.

1.2 Inorganic chemistry

The branch of chemistry that studies compounds that do not contain carbon and hydrogen atoms is called "inorganic chemistry." Simply put, it is the opposite of organic chemistry. Substances that do not have carbon-hydrogen bonds include metals, salts, and chemicals.

1.3 Organic chemistry

The discipline which deals with the study of the structure, composition and the chemical properties of organic compounds is known as organic chemistry. It involves the study of Carbon and its compounds.

1.4 Biochemistry

Biochemistry is that branch of chemistry that explores the chemical processes in organisms and associated with them. It's a laboratory-based science that connects biology and chemistry. By using chemical knowledge and technology, biochemists can understand and solve biological problems

1.5 Analytical chemistry

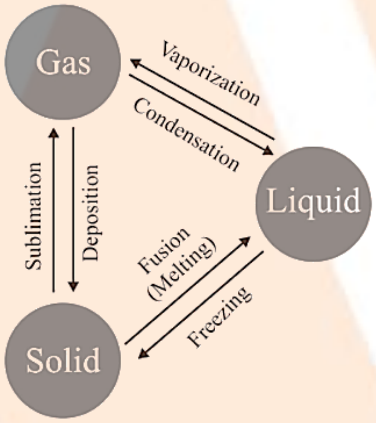
It is the branch of chemistry which uses instruments and analytical techniques to determine structure, functionality and properties of a substance.

2. MATTER

Matter is defined as any thing that have some mass and also occupies a certain volume in a space.
Generally matter is classified into three phases:

- **Solid**- The substance which have a definite shape as well as maintain its volume as per it's shape, also they have least freedom of movement. e.g., sugar, iron, gold, wood etc.
- **Liquid**- A substance is a substance which generally possess the shape of a container but have a fixed volume. Also liquids have the property to flow or to be poured. E.g., water, milk, oil, mercury, alcohol etc.
- **Gas**- Substances which do not have a definite volume as well as definite shape. Gases generally completely fill the container they are kept in. E.g., hydrogen, oxygen etc.

The three states are interconvertible by changing the conditions of temperature and pressure as follows:



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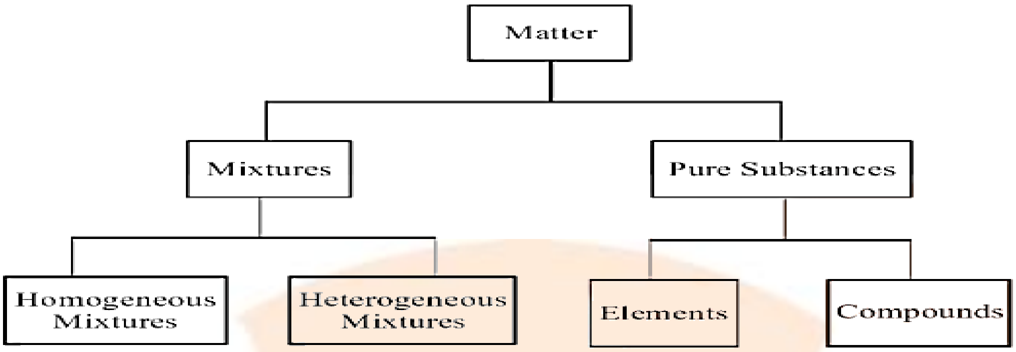
3. CLASSIFICATION OF MATTER AT MACROSCOPIC LEVEL

Matter can further be classified into following at bulk or macroscopic level:
(a) Mixtures (b) Pure Substances.
These can be further classified as shown below:



Classification of matter (taken from original pdf)

(a) Mixtures : A mixture is a substance in which two or more substance are present in any ratio. Primarily, It is of two types: Heterogeneous and Homogeneous mixtures.

- **Homogeneous mixture**- Two substances are mixed to form a mixture such that there exist one single uniform phase i.e. composition of substances present is uniform. Sugar solution and air are thus, the examples of homogeneous mixtures.
- **Heterogeneous mixtures-** Two or more substances are mixed which result in non-uniform composition throughout the mixture. Some of the examples are suspensions, mixture of two solids suppose salt and sugar.

Note: Any distinct portion of matter that is uniform throughout in composition and properties is called a Phase.

(b) Pure substances:- A material containing only one type of particle is called a pure substance.

Note: In chemistry, Form of matter having constant chemical composition and chemical properties and they cannot be separated into component by physical methods.

Pure substances are further divided as given below:

- **Element-** An element is defined as a pure substance that contains only one kind of atoms and cannot be further broken down. The elements are further split into three classes based on their physical and chemical properties i.e. (1) Metals (2) Non- metals and (3) Metalloids.

- **Compound-** A compound is a pure material that consists of two or more elements mixed in a defined mass proportion. Furthermore, a compound's qualities are distinct from those of its constituting elements. Moreover, the constituents of a compound cannot be separated into simpler substances by physical methods. They can only be separated by chemical methods.

4. PROPERTIES OF MATTER

Unique or characteristic properties is depicted by every substance.
Physical properties and chemical properties are the two types of properties that are observed.

4.1 Physical Properties

Physical properties are those that may well be measured or observed without affecting the substance's identity or composition. Colour, fragrance, melting point, boiling point, density, and other physical qualities are some of the examples.

4.2 Chemical properties

Chemical properties are the properties of specific substances that can be observed in chemical reactions. Some of the main chemical properties include flammability, toxicity, heat of combustion, pH, radioactive decay rate, and chemical stability.

5. MEASUREMENT

5.1 Physical quantities

Physical quantities are quantities which we encounter during our scientific study. Any physical quantity can be measured in two parts:
(1) The number, and (2) The unit: Unit is defined as the reference standard chosen to measure any physical quantity.

5.2 S.I. UNITS

The International System of Units (in French Le Systeme International d'Unités – abbreviated as SI) was established by the eleventh General Conference on Weights and Measures (CGPM from Conference Generale des Poids at Measures). The CGPM is an inter governmental treaty organization created by a diplomatic treaty known as Meter Convention which was signed in Paris in 1875.

There are seven base units in SI system as listed below
These units pertain to the seven fundamental scientific quantities. The other physical quantities such as speed, volume, density etc. can be derived from these quantities. The definitions of the SI base units are given below:

Unit of length	Metre	The metre is the length of the path travelled by light in vacuum during a time interval of $1/299\ 792\ 458$ of a second.
Unit of mass	Kilogram	The kilogram is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
Unit of time	Second	The second is the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium -133 atom.
Unit of electric current	Ampere	The ampere is that constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross-section, and placed 1 metre apart in vacuum, would produce between these conductors a force

		equal to 2×10^{-7} newton per metre of length.
Unit of thermodynamic Temperature	Kelvin	The kelvin, unit of thermodynamic temperature, is the temperature fraction $\frac{1}{273}$. 16 of the thermodynamic temperature of the triple point of water.
Unit of amount of substance	mole	1. The mole is the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon–12 ; its symbol is "mol.". 2. When the mole is used, the elementary entities must be specified and may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.
Unit of luminous intensity	Candela	The candela is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of

		frequency 540×1012 hertz and that has a radiant intensity in that direction of $\frac{1}{683}$ watt per steradian.

Note: The mass standard is the kilogram since 1889. It has been defined as the mass of platinum-iridium (Pt-Ir) cylinder that is stored in an airtight jar at International Bureau of Weights and Measures in Sevres, France. Pt-Ir was chosen for this standard because it is highly resistant to chemical attack and its mass will not change for an extremely long time.

6. SOME IMPORTANT DEFINITION

6.1 Mass and Weight

The mass of a substance is the amount of substance present in it and the weight is the force exerted by gravity on the object. The mass of matter is constant, and due to changes in gravity, its weight can vary from place to place. The SI unit of mass is kilogram (kg). The SI derived unit of weight (the derived unit of the SI base unit) is Newton.

6.2 Volume

Volume is the amount of three-dimensional space surrounded by certain closed boundaries, for example, the space occupied or contained by a substance (solid, liquid, gas, or plasma) or shape. Volume is usually quantified numerically using SI derived units (cubic meters).

6.3 Density

The mass density or density of a material is defined as its mass per unit volume. The density is represented by the symbol ρ (the lower case Greek letter rho). SI unit of density is $kg\ m^{-3}$.

6.4 Temperature

Temperature is a physical property of matter, which quantitatively expresses the common concepts of heat and cold. There are three common scales for measuring temperature: **(Celsius)**, **(Fahrenheit)** and **(Kelvin)**.The temperature on two scales is related to each other by the following relationship:

$$°F = 9/5\left(°C\right) + 32$$

$K = {}^\circ C + 273.15$

7. LAW OF CHEMICAL COMBINATION

7.1 Law of conservation of mass

"In a chemical reaction the mass of reactants consumed and mass of the products formed is same, that is mass is conserved." This is a direct consequence of law of conservation of atoms. This law was given by Antoine Lavoisier in 1789.

7.2 Law of Constant / Definite Proportions

The law of definite proportions states that the mass proportions of the elements in a composite sample are always the same. It was given by, a French chemist, Joseph Proust.

7.3 Law of Multiple Proportions

The law of multiple proportions states that when two elements are combined to form more than one compound, the weight of one element is proportional to the fixed weight of the other element as a whole number. This law was proposed by Dalton in 1803.

7.4 Law of Reciprocal Proportions

The law states that if two different elements are combined with the fixed mass of the third element, their combined mass ratio is the same, or a simple multiple of their combined mass. This Law was proposed by Richter in 1792 .

7.5 Gay Lussac's Law of Gaseous Volumes

The law developed by Gay Lussac in 1808 establishes that "the relationship between the volume of a gaseous reactant and a product can be represented by a simple whole number."

7.6 Avogadro Law

In 1811 , Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

8. DALTON'S ATOMIC THEORY

In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following:

1. Matter consists of indivisible atoms.
2. All the atoms of a given element have identical properties including

identical mass. Atoms of different elements differ in mass.

3. Compounds are formed when atoms of different elements combine in a fixed ratio.

4. Chemical reactions involve reorganization of atoms. These are neither created nor destroyed in a chemical reaction.

9. ATOM

Atom is defined as the **smallest unit** that retains the properties of an element as well as participate in a chemical reaction {Note: This definition holds true only for non-radioactive reactions}

9.1 Mass of an Atom

There are two ways to denote the mass of atoms.

9.2 Method 1

Atomic mass can be defined as a mass of a single atom which is measured in atomic mass unit (amu) or unified mass (u) where

1 a.m.u. $= \frac{1}{12}$ th of the mass of one C^{12} atom

9.3 Method 2

Mass of 6.022×10^{23} atoms of that element taken in grams. This is also known as molar atomic mass.

Note:

- Mass of 1 atom in amu and mass of 6.022×10^{23} atoms in grams are numerically equal.
- When atomic mass is taken in grams it is also called the molar atomic mass.
- 6.022×10^{23} is also called 1 mole of atoms and this number is also called the Avogadro's Number.
- Mole is just a number. As 1 dozen $= 12$;

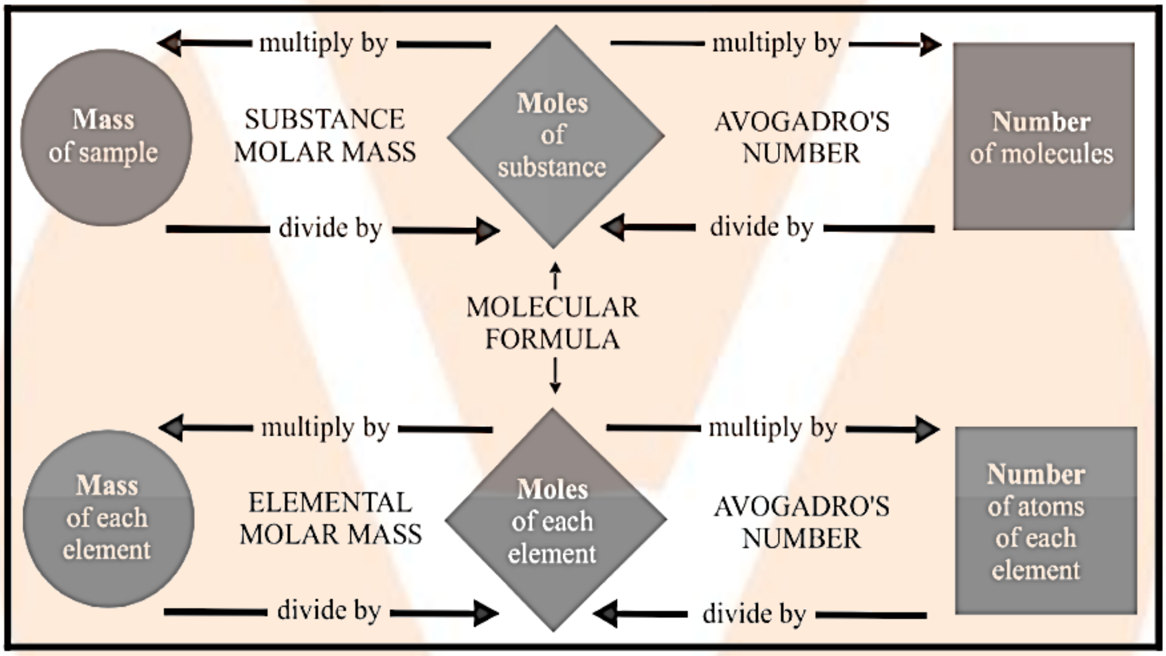
  1 million $= 10^{6}$;

  1 mole $= 6.022 \times 10^{23}$

10. MOLECULES

The smallest particle of a substance that has all the physical and chemical properties of the substance. A molecule consists of one or more atoms. e.g. H_2, NH_3.

The mass of a molecule is measured by adding the masses of the atoms that compose it. Therefore, the mass of a molecule can also be expressed by the two methods used to measure the mass of an atom i.e. amu and g / mol.



11. CHEMICAL REACTIONS

A process that involves rearranging the molecular or ionic structure of a substance, which is different from a change in physical form or a nuclear reaction.

Points to remember:

- Equations must be balanced before proceeding for calculations.
- We do not need to conserve the number of molecules in a reaction. e.g $N_2 + 3\,H_2 \rightarrow 2NH_3$. If we only discuss the rearrangement of atoms in the equilibrium chemical reaction, and the number of molecules is not conserved, then it is clear that the mass of the atoms on the reaction side is equal to the sum of the masses of the atoms on the reaction side. This is the law of conservation of atoms and the law of conservation of

mass.

12. STOICHIOMETRY

The study of chemical reactions and related calculations is called stoichiometry. The coefficient used to balance the reaction is called the stoichiometric coefficient.

Points to remember :

- The stoichiometric coefficients is the ratio of moles of molecules of atoms that reacts not the mass.
- Only when all reactants are present in a stoichiometric ratio can the stoichiometric ratio be used to predict the number of moles of product formed. The actual quantity of product formed is always less than the quantity predicted by the theoretical calculation

12.1 Limiting Reagent (LR) and Excess Reagent (ER)

If the reagents are not used in a stoichiometric ratio, then the less than the required amount of reagent determines how much product will be formed, which is called the limiting reagent, and the reagent in excess is called the excess reagent. For example, if we burn carbon in the air (it has an unlimited supply of oxygen), the amount of CO_2 produced will depend on the amount of carbon inhaled. In this case, Carbon is the **limiting reactant** and O_2 is the **excess reactant**.

13. PERCENT YIELD

As mentioned above, due to practical reasons, the amount of product formed by the chemical reaction is less than the amount predicted by the theoretical calculation. When multiplied by 100, the relationship between the amount of product formed and the predicted amount gives the percentage yield

$$\text{Percentage Yield}=\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\times 100$$

14. REACTIONS IN AQUEOUS MEDIA

The two solids cannot react with each other in the solid phase, so they must be dissolved in the liquid. When solutes dissolve in a solvent, they coexist in a single phase called a solution. Several parameters are used to measure the strength of the solution. The strength of a solution denotes the amount of solute which is contained in the solution.

The parameters used to denote the strength of a solution are:

- Mole fraction X: moles of a component / Total moles of solution.
- Mass% : Mass of solute (in g) present in 100g of solution.
- Mass/Vol : Mass of solute (in g) present in 100mL of solution
- v/v : Volume of solute/volume of solution {only for liq-liq solutions}
- g/L : Wt. of solute (g) in 1L of solution
- ppm : $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10$
- Molarity (M) : $\frac{\text{moles of solute}}{\text{volume of solution (L)}}$
- Molality (m) : $\frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$

IMPORTANT RELATIONS

1. Relation between molality (m) Molarity (M), density (d) of solution and molar mass of solute (M_O)

d : density in g/mL

M_O : molar mass in g mol^{-1}

$$\text{Molality, } m = \frac{M \times 1000}{1000d - MM_O}$$

2. Relationship between molality (m) and mole fraction (X_B) of the solute

$$m = \frac{X_B}{1 - X_B} \times \frac{1000}{M_X}$$

$$m = \frac{1 - X_A}{X_A} \times \frac{1000}{M_A}$$

Points to remember:

- Molarity is the most common unit of measuring strength of solution.
- The product of Molarity and Volume of the solution gives the number of moles of the solute, $n = M \times V$

- All the formulae of strength have amount of solute. (weight or moles) in the numerator.
- All the formulae have amount of solution in the denominator except for molality (m).

15. DILUTION LAW

The moles of a solution do not change when it is diluted. If the volume of a solution having a Molarity of M_1 is diluted from V_1 to V_2 by adding more solvent we can write that:

$M_1V_1 =$ moles of solute in the solution $= M_2V_2$

16. EFFECT OF TEMPERATURE

The volume of solvent increases with increasing temperature. But assuming that the system is closed, that is, there is no mass loss, and it has no effect on the quality of the solute in the solution. The formulae of strength of solutions which do not involve volume of solution are unaffected by changes in temperature.

e.g. molality remains unchanged with temperature. Formulae involving volume are altered by temperature e.g. Molarity.

17. INTRODUCTION TO EQUIVALENT CONCEPT

The concept of equivalence is a way to understand chemical reactions and processes, and the concept of equivalence is usually used for simplification.

17.1 Equivalent Mass

"The mass of an acid which furnishes 1 mol $H+$ is called its Equivalent mass."

"The mass of the base which furnishes 1 mol $OH-$ is called its Equivalent mass."

17.2 Valence Factor

(n-factor(represented by Z))

Valence factor: is the number of $H+$ ions supplied by 1 molecule or mole of an acid or the number of $OH-$ ions supplied by 1 molecule or 1 mole of the base.

Mass Equivalent, $E = \dfrac{\text{Molecular Mass}}{Z}$

17.3 Equivalents

No. of equivalents $= \dfrac{\dfrac{\text{wt. of acid}}{\text{base taken}}}{\text{Eq. wt}}$.

Note: It should be always remembered that 1 equivalent of an acid reacts with 1 equivalent of a base.

18. MIXTURE OF ACIDS AND BASES

If we have a mixture of multiple acids and bases, we can use the concept of equivalence to determine whether the resulting solution is acidic or alkaline. For a mixture of multiple acids and bases, find out the equivalents of the acids and bases used, and find out which one is in excess.

19. LAW OF CHEMICAL EQUIVALENCE

According to this law, an equivalent of a reactant is combined with an equivalent of another reactant to obtain an equivalent of each product. For, example in a reaction $aA + bB \rightarrow cC + dD$ irrespective of the stoichiometric coefficients, 1 eq. of A reacts with 1 eq. of B to give 1 eq. each of C and 1 eq of D

20. EQUIVALENT WEIGHTS OF SALTS

To calculate the equivalent weight of a compound that is neither an acid nor a base, we need to know the charge of the cation or anion. The mass of the cation divided by the charge on it is called the equivalent mass of the cation, and the mass of the anion divided by the charge on it is called the equivalent mass of the anion. When we add equal masses of anions and cations, it gives us equal amounts of salt. For salts, the valence factor is the total amount of positive or negative charges provided by the 1 mol of salt.

21. ORIGIN OF EQUIVALENT CONCEPT

The equivalent weight of an element was originally defined as the weight of the element combined with 1 grams of hydrogen. Subsequently, the definition was modified as follows: the equivalent weight of an element is the weight of the element combined with 8 grams of oxygen. Note: Same element can have multiple equivalent weights depending upon the charge on it. e.g. Fe^{2+} and Fe^{3+}

22. EQUIVALENT VOLUME OF GASES

Equivalent volume of gas is the volume occupied by 1 equivalent of a gas at STP.

Equivalent mass of gas $= \text{molecular mass}/Z$.

Since 1 mole of gas occupies $22.4L$ at STP therefore 1 equivalent of a gas will occupy $22.4/Z$ L at STP. e.g. Oxygen occupies $5.6L$, Chlorine and Hydrogen occupy $11.2L$.

23. NORMALITY

The normality of a solution is the number of equivalents of solute present in 1 L of the solution.

$$N = \frac{\text{equivalents of solute}}{\text{volume of solution } (L)}$$

The number of equivalents of solute present in a solution is given by $\text{Normality} \times \text{Volume } (L)$.

On dilution of the solution the number of equivalents of the solute is conserved and thus, we can apply the formula : $N_1V_1 = N_2V_2$

Caution:

Please note that the above equation can cause a lot of confusion and is a common mistake students make. This is the dilution equation for the conservation of the equivalent number. Now, since one reactant equivalent always reacts with one equivalent of another reactant, similar equations are used in problems involving acid and base titration. Don't extend the same logic to molarity.

Relationship between Normality and Molarity is given by:

$N = M \times Z$; where 'Z' is the Valency factor