SOME BASIC CONCEPTS OF CHEMISTRY

1. CHEMISTRY

Chemistry is defined as the study of the composition, properties and interaction of matter. Chemistry is often called the central science because of its role in connecting the physical sciences, which include chemistry, with the life sciences and applied sciences such as medicine and engineering.

Various branches of chemistry are

1.1 Physical chemistry

The branch of chemistry concerned with the way in which the physical properties of substances depend on and influence their chemical structure, properties, and reactions.

1.2 Inorganic chemistry

The branch of chemistry which deals with the structure, composition and behavior of inorganic compounds. All the substances other than the carbon-hydrogen compounds are classified under the group of inorganic substances.

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.

1.4 Biochemistry

The discipline which deals with the structure and behavior of the components of cells and the chemical processes in living beings is known as biochemistry.

1.5 Analytical chemistry

The branch of chemistry dealing with separation, identification and quantitative determination of the compositions of different substances.

2. MATTER

Matter is defined as any thing that occupies space possesses mass and the presence of which can be felt by any one or more of our five senses.

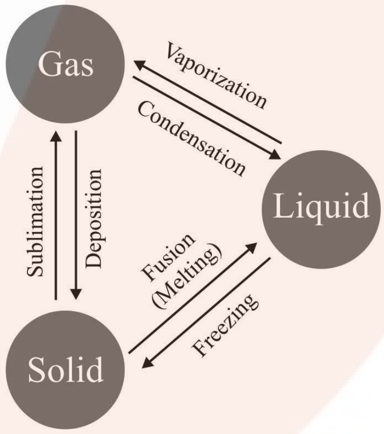
Matter can exist in 3 physical states viz. solid, liquid, gas.

Solid - a substance is said to be solid if it possesses a definite volume and a definite shape, e.g., sugar, iron, gold, wood etc.

Liquid- A substance is said to be liquid, if it possesses a definite volume but no definite shape. They take up the shape of the vessel in which they are put, e.g., water, milk, oil, mercury, alcohol etc.

Gas- a substance is said to be gaseous if it neither possesses definite volume nor a definite shape. This is because they fill up the whole vessel in which they are put, e.g., hydrogen, oxygen etc.

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

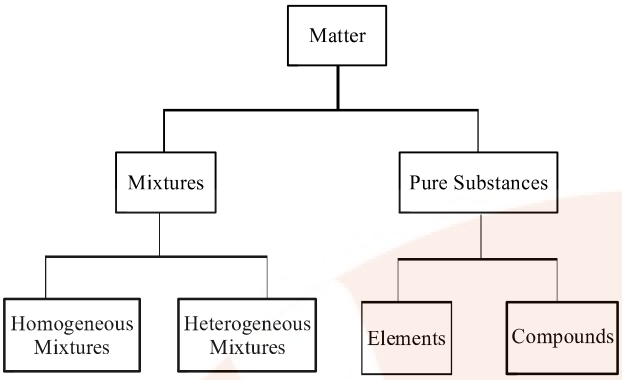


3. CLASSIFICATION OF MATTER AT MACROSCOPIC LEVELL

At the macroscopic or bulk level, matter can be classified as

(a) mixtures (b) pure substances.

These can be further sub-divided as shown below



Classification of matter

(a) Mixtures : A mixture contains two or more substances present in it (in any ratio) which are called its components. A mixture may be homogeneous or heterogeneous.

Homogeneous mixture- in homogeneous mixture the components completely mix with each other and its composition is uniform throughout i.e it consist of only one phase. Sugar solution and air are thus, the examples of homogeneous mixtures.

Heterogeneous mixtures- In heterogeneous mixture the composition is not uniform throughout and sometimes the different phases can be observed. For example, grains and pulses along with some dirt (often stone) pieces, are heterogeneous mixtures.

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 substance is called a pure substance.

Note...

In chemistry, a **substance** is a form of matter that has constant chemical composition and characteristic properties. It cannot be separated into components by physical separation methods, i.e. without breaking chemical bonds. They can be solids, liquids or gases.

Pure substances can be further classified into **elements** and **compounds.**

Element- An element is defined as a pure substance that contains only one kind of particles. Depending upon the physical and chemical properties, the elements are further subdivided into three classes, namely (1) Metals (2) Non-metals and (3) Metalloids.

Compound- A compound is a pure substance containing two or more than two elements combined together in a fixed proportion by mass. Further, the properties of a compound are completely different from those of its constituent elements. Moreover, the constituents of a compound cannot be separated into simpler substances by physical methods. They can be separated by chemical methods.

4. PROPERTIES OF MATTER

Every substance has unique or characteristic properties. These properties can be classified into two categories – physical properties and chemical properties.

4.1 Physical Properties

Physical properties are those properties which can be measured or observed without changing the identity or the composition of the substance. Some examples of physical properties are color, odor, melting point, boiling point, density etc.

4.2 Chemical properties

Chemical properties are those in which a chemical change in the substance occurs. The examples of chemical properties are characteristic reactions of different substances; these include acidity or basicity, combustibility etc.

5. MEASUREMENT

5.1 Physical quantities

All such quantities which we come across during our scientific studies are called Physical quantities. Evidently, the measurement of any physical quantity consists of two parts

(1) The number, and (2) The unit

A **unit** is defined as the standard of reference 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 11^{th} 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.

The SI system has seven base units and they are listed in table given 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 :

Definitions of SI Base Units

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 internationl 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 fraction 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 540×10^{12} hertz and that has a radiant intensity in that direction of 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

Mass of a substance is the amount of matter present in it while **weight** is the force exerted by gravity on an object. The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity. The **SI unit** of mass is the **kilogram** (kg). The **SI derived unit** (unit derived from SI base units) of weight is **newton**.

6.2 Volume

Volume is the quantity of three-dimensional space enclosed by some closed boundary, for example, the space that a substance (solid, liquid, gas, or plasma) or shape occupies or contains. Volume is often quantified numerically using the SI derived unit, the cubic meter.

6.3 Density

The **mass density** or **density** of a material is defined as its mass per unit volume. The symbol most often used for density is 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 that quantitatively expresses the common notions of hot and cold. There are three common scales to measure temperature — °C (degree celsius), °F (degree fahrenheit) and K (kelvin). The temperature on two scales is related to each other by the following relationship:

$$°F = 9/5\,(°C) + 32$$

$$K = °C + 273.15$$

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 put forth by Antoine Lavoisier in 1789.

7.2 Law of Constant / Definite Proportions

The ratio in which two or more elements combine to form a compound remains fixed and is independent of the source of the compound. This law was given by, a French chemist, Joseph Proust.

7.3 Law of Multiple Proportions

When two elements combine to form two or more compounds then the ratio of masses of one element that combines with a fixed mass of the other element in the two compounds is a simple whole number ratio. This law was proposed by Dalton in 1803.

7.4 Law of Reciprocal Proportions

When three elements combine with each other in combination of two and form three compounds then the ratio of masses of two elements combining with fixed mass of the third and the ratio in which they combine with each other bear a simple whole number ratio to each other. This Law was given by **Richter** in 1792.

7.5 Gay Lussac's Law of Gaseous Volumes

This law was given by Gay Lussac in 1808. He observed that when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

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 the smallest part of an element that can 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. = 1/12th 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...

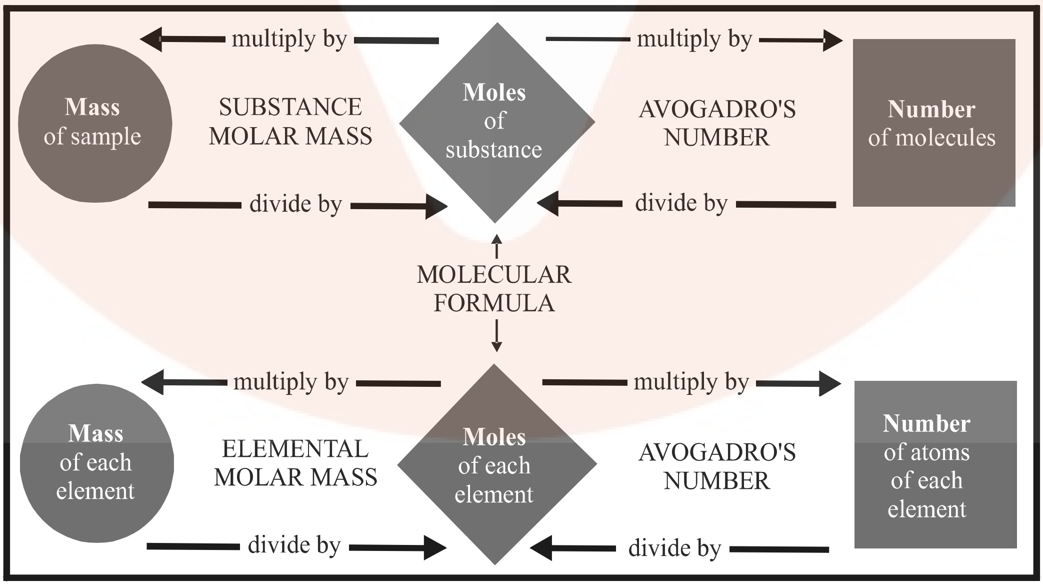
- Mass of 1 atom in amu and mass of 6.022×10^{23} atoms in grams are numerically equal.
- When atomic mass is taken in grams it is also called the molar atomic mass.
- 6.022×10^{23} is also called 1 mole of atoms and this number is also called the **Avogadro's Number.**
- Mole is just a number. As 1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

  1 million = 10^6; 1 mole = 6.022×10^{23}**.**

10. MOLECULES

A group of similar or dissimilar atoms which exist together in nature is known as a molecule. e.g. H_2, NH_3.

The mass of molecules is measured by adding the masses of the atoms which constitute the molecule. Thus, the mass of a molecule can also be represented by the two methods used for measuring the mass of an atom viz. amu and g/mol.



11. CHEMICAL REACTIONS

A chemical reaction is only rearrangement of atoms. Atoms from different molecules (may be even same molecule) rearrange themselves to form new molecules.

Points to remember :

- Always balance chemical equations before doing any calculations
- The number of molecules in a reaction need not to be conserved e.g.

  $N_2 + 3\ H_2 \rightarrow 2\ NH_3$. The number of molecules is not conserved

  If we talk about only rearrangement of atoms in a balanced chemical reaction then it is evident that the mass of the atoms in the reactants side is equal to the sum of the masses of the atoms on the products side. This is the **Law of Conservation of Atoms** and **Law of Conservation of Mass.**

12. STOICHIOMETRY

The study of chemical reactions and calculations related to it is called Stoichiometry. The coefficients used to balance the reaction are called **Stoichiometric Coefficients.**

Points to remember :

- The stoichiometric coefficients give the ratio of molecules or moles that react and **not the ratio of masses.**
- Stoichiometric ratios can be used to predict the moles of product formed only if all the reactants are present in the stoichiometric ratios.

  Practically the amount of products formed is always less than the amount predicted by theoretical calculations

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

If the reactants are not taken in the stoichiometric ratios then the reactant which is less than the required amount determines how much product will be formed and is known as the **Limiting Reagent** and the reactant present in excess is called the **Excess Reagent.** e.g. if we burn carbon in air (which has an infinite supply of oxygen) then the amount of CO_2 being produced will be governed by the amount of carbon taken. In this case, Carbon is the LR and O_2 is the ER.

13. PERCENT YIELD

As discussed earlier, due to practical reasons the amount of product formed by a chemical reaction is less than the amount predicted by theoretical calculations. The ratio of the amount of product formed to the amount predicted when multiplied by 100 gives the percentage yield.

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

14. REACTIONS IN AQUEOUS MEDIA

Two solids cannot react with each other in solid phase and hence need to be dissolved in a liquid. When a solute is dissolved in a solvent, they co-exist in a single phase called the solution. Various parameters are used to measure the strength of a 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:** $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$
- **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

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

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 × 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

When a solution is diluted, more solvent is added, the moles of solute remains unchanged. If the volume of a solution having a Molarity of M_1 is changed from V_1 to V_2 we can write that:

$$\mathbf{M_1V_1 = \text{moles of solute in the solution} = M_2V_2}$$

16. EFFECT OF TEMPERATURE

Volume of the solvent increases on increasing the temperature. But it shows **no effect on the mass of solute** in the solution assuming the system to be closed i.e. no loss of mass.

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

Equivalent concept is a way of understanding reactions and processes in chemistry which are often made simple by the use of Equivalent concept.

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 Valency Factor (Z)

Valency 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.

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

17.3 Equivalents

$$\textbf{No. of equivalents} = \frac{\text{wt. of acid / base taken}}{\text{Eq. wt.}}$$

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

Whenever we have a mixture of multiple acids and bases we can find whether the resultant solution would be acidic or basic by using the equivalent concept. For a mixture of multiple acids and bases find out the equivalents of acids and bases taken and find which one of them is in excess.

19. LAW OF CHEMICAL EQUIVALENCE

According to this law, one equivalent of a reactant combines with one equivalent of the other reactant to give one 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 1eq of D

20. EQUIVALENT WEIGHTS OF SALTS

To calculate the equivalent weights of compounds which are neither acids nor bases, we need to know the charge on the cation or the 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 the equivalent masses of the anion and the cation, it gives us the equivalent mass of the salt. For salts, Z in the total amount of positive or negative charge furnished by 1 mol of the salt.

21. ORIGIN OF EQUIVALENT CONCEPT

Equivalent weight of an element was initially defined as **weight of an element which combines with 1g of hydrogen.** Later the definition wad modified to : **Equivalent weight of an element is that weight of the element which combines with 8g 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 = 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 1L 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 **Normality × 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 gives rise to a lot of confusion and is a common mistake that students make. This is the equation of dilution where the number of equivalents are conserved. Now, since one equivalent of a reactant always reacts with 1 equivalent of another reactant a similar equation is used in problems involving titration of acids and bases. Please do not extend the same logic to molarity.

Relationship between Normality and Molarity

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