UNIT 1

"From invisible atoms to measurable matter, this unit builds the grammar of Chemistry—equip with tools to quantify, classify, and write the formulas of the world around us."

Some Basic Concepts of Chemistry

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba		Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra		Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts	Og

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Periodic table

Importance of Chemistry, Nature of matter, Properties of matter & their measurement. Laws of chemical combinations, Dalton's Atomic Theory. Names of important elements and their symbol, valency, writing the formula of certain compounds, SI units, drawing the relation between SI and non-SI units, Atomic & molecular masses, Percentage Composition. Writing the dimension for physical quantities like volume, pressure, force, area, viscosity and surface tension

1.1 Introduction:

Chemistry is the branch of science that deals with studies of the composition, properties and interaction of matter and substances or the interaction between individuals and how substances interact with energy. *Antoine Laurent Lavoisier* is called the father of Chemistry.

1.2 Importance of Chemistry:

Chemistry plays a central role in science and is often intertwined with other branches of science like Physics, Biology, Geology etc. Chemistry is part of everything in our lives. Chemistry is essential for meeting our basic needs of food, clothing, shelter, health, energy, and clean air, water, and soil. Chemical principles are important in diverse areas like weather patterns, functioning of brain and operation of a computer. Chemical industries like manufacturing fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metals, alloys and other inorganic and organic chemicals, including new materials, contribute in a big way to the national economy.

The five major branches of Chemistry are organic, inorganic, analytical, physical, and biochemistry. These divide into many sub-branches.

1.3 Importance of Chemistry for Paramedical Courses:

Chemistry is essential in paramedical education, providing the scientific basis for understanding the human body, disease mechanisms, and drug actions. It supports key areas like pharmacology, diagnostics, radiology, and anesthesia. Knowledge of chemical principles ensures accurate testing, safe medication use, and effective sterilization. Overall, chemistry enhances clinical skills, improves patient care, and empowers paramedical professionals to make informed decisions in healthcare settings.

Overall, chemistry equips paramedical students with the scientific foundation needed to make informed decisions, perform accurate diagnostics, and contribute meaningfully to patient treatment and recovery.

- **Core to Understanding Human Biology**
  - Biochemical processes: Chemistry explains how enzymes, hormones, and neurotransmitters function in the body.
  - Cellular reactions: Knowledge of chemical reactions helps paramedics understand metabolism, respiration, and disease mechanisms.
- **Drug and Medication Knowledge**
  - Pharmacology: Chemistry is essential for understanding drug composition, interactions, dosages, and side effects.
  - Sterilization and preservation: Chemical principles guide the use of disinfectants, antiseptics, and preservatives in medical settings.

- **Diagnostic Techniques**
  - Laboratory tests: Blood tests, urine analysis, and other diagnostics rely on chemical reactions and indicators.
  - Imaging and contrast agents: Chemistry helps explain how substances like iodine or barium are used in X-rays and scans.
- **Hygiene and Safety**
  - Infection control: Understanding chemical properties of cleaning agents ensures proper sterilization and hygiene.
  - Handling chemicals safely: Paramedics must know how to store and use chemicals without causing harm.
- **Applied Skills in Paramedical Fields**
  - Medical laboratory technology: Requires strong chemistry knowledge for analyzing samples and interpreting results.
  - Radiology and anesthesia: Involves chemical understanding of gases, contrast media, and drug interactions.

1.4 Nature of matter:

Anything which has mass and occupies space is called matter. It makes up everything around us from the air we breathe to the food we eat and the bodies we live in. It is made up of small particles which have space between them. The matter particles attract each other and are in a state of continuous motion.

- **Basic Characteristics**

Mass: The amount of matter in an object.
Volume: The space that matter occupies.
Inertia: Resistance to changes in motion.

Matter can exist in four physical states namely solid, liquid, gas and plasma
In solids, these particles are held very close to each other in an orderly manner and there is not much freedom of movement. In liquids, the particles are close to each other but they can move around. However, in gases, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.

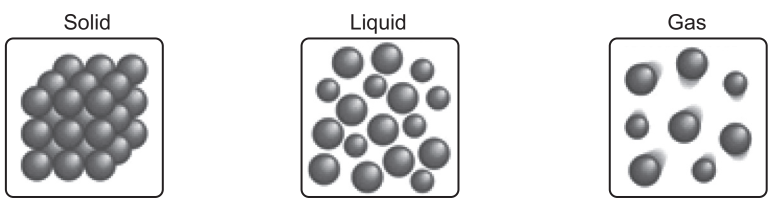


Fig . 1.1 : States of Matter

Because of such arrangement of particles, different states of matter exhibit the following characteristics:

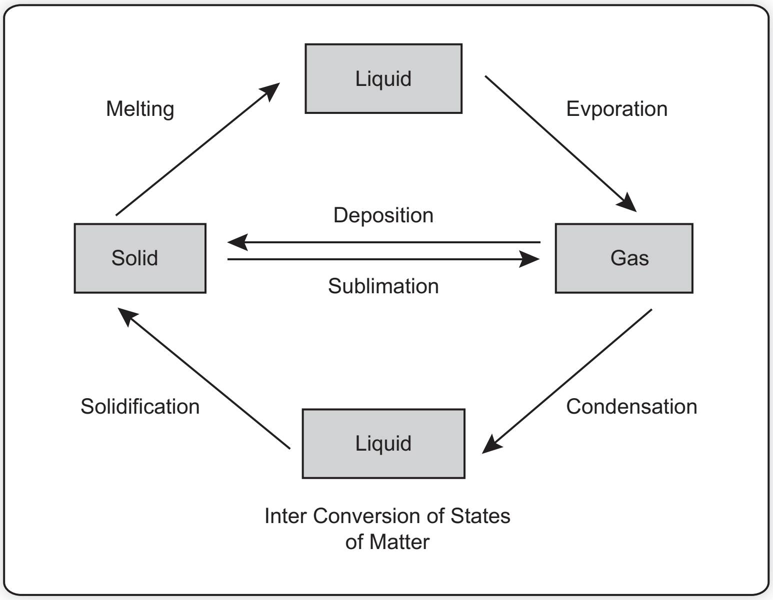
i. **Solid**: Fixed shape and volume; particles are tightly packed.

ii. **Liquid**: Fixed volume but no fixed shape; particles move more freely. They take the shape of the container in which they are placed.

iii. **Gas**: No fixed shape or volume; particles move rapidly and are far apart. They completely occupy the container in which they are placed.

iv. **Plasma**: Ionized gas found in stars and lightning.

The three states of matter are interconvertible by changing the conditions of temperature and pressure.



Flow Chart - 1.1 : Interconversion of States of Matter

1.5 Elements:

Elements are pure substances made of only one type of atom which can neither be decomposed into nor built from simpler substances by ordinary physical and chemical methods. They are the basic building blocks of all matter.

Each element has a **name**, a **symbol**, and an **atomic number**. They are organized in the **Periodic Table** based on their properties. The number of elements known till date is 118. An element can be a metal, a non-metal or a metalloid.

Hydrogen is the most abundant element in the universe. Oxygen (46.6%), a non-metal, is the most abundant element in the earth crust. Aluminum is the most abundant metal in the earth crust.

1.6 Symbols:

A symbol is an abbreviation or shortened form for the full name of an element and compound They make communication in science precise and efficient.

Every element has its own symbol. The present system of symbols was introduced by Berzelius. Each element is represented by one or two letters. The first letter is always capitalized; the second (if present) is lowercase.

Examples:

- **H** – Hydrogen
- **O** – Oxygen
- **Na** – Sodium
- **Cl** – Chlorine
- **Fe** – Iron

Compound Symbols

Compounds are written using the symbols of their constituent elements.

Examples:

H_2O – Water (2 hydrogen atoms + 1 oxygen atom)

NaCl – Sodium chloride (table salt)

CO_2 – Carbon dioxide

Other Common Symbols

(s) – Solid

(l) – Liquid

(g) – Gas

(aq) – Aqueous (dissolved in water)

$\rightarrow$ – Indicates a chemical reaction

Δ – Heat is applied **(Table - 1.1)**

1.7 Compounds:

It is also the form of matter which can be formed by combining two or more elements in fixed proportions. Unlike mixtures, compounds have unique properties and cannot be separated by physical means. It can be decomposed into its constituent elements **only** by suitable chemical methods. They have fixed ratios.

e.g., Water (H_2O) is made of hydrogen and oxygen in the ratio 1: 8 by mass.

Table - 1.1 : Examples of Common Elements with symbols

Element Name	Symbol	Use in Paramedical Field
Oxygen	O	Essential for respiration
Hydrogen	H	Found in water and body fluids
Carbon	C	Backbone of organic molecules
Calcium	Ca	Important for bones and teeth
Iron	Fe	Vital for blood (hemoglobin)
Sodium	Na	Regulates fluid balance
Potassium	K	Helps nerve and muscle function
Chlorine	Cl	Used in disinfectants
Iodine	I	Needed for thyroid function
Nitrogen	N	Found in proteins and DNA

Examples of Common Compounds

- **Water (H_2O)** – Vital for life and body functions.
- **Carbon dioxide (CO_2)** – Produced during respiration.
- **Glucose ($C_6H_{12}O_6$)** – A sugar used for energy.
- **Sodium chloride (NaCl)** – Common table salt.
- **Calcium carbonate ($CaCO_3$)** – Found in bones and teeth.

Compounds can be of two types based on their origin:

(i) **Inorganic compounds-** Previously, it was believed that these compounds are derived from non-living sources, like rocks and minerals. An inorganic compound is typically a chemical compound that lacks carbon–hydrogen bonds.

(ii) **Organic compounds-** According to earlier scientists, these compounds are derived from living sources like plants and animals, or these remain buried under the earth (e.g., petroleum). According to modern concept, organic compounds are any compounds that contain carbon and hydrogen which are called as hydrocarbons and their derivatives.

Table - 1.2 : Classification Based on Origin

Type	Description	Example
Inorganic Compounds	Derived from non-living sources	NaCl (salt)
Organic Compounds	Derived from living sources	CH_4(methane)

Types of Compounds based on type of bond involved:

Compounds can be classified based on the type of chemical bond that holds their atoms together.

Ionic compounds are formed through the transfer of electrons from one atom to another, resulting in oppositely charged ions that attract each other; a common example is sodium chloride (NaCl), commonly known as table salt.

Covalent compounds, on the other hand, arise from the sharing of electrons between atoms, as seen in water (H_2O), where hydrogen and oxygen atoms share electrons to form stable molecules.

Metallic compounds involve metal atoms sharing a "sea" of delocalized electrons, which allows for properties like conductivity and malleability; brass, an alloy of copper (Cu) and zinc (Zn), is a typical example of this type

Table - 1.3 : Classification Based on Bond

Type	Description	Example
Ionic Compounds	Formed by transfer of electrons	NaCl (salt)
Covalent Compounds	Formed by sharing of electrons	H_2O (water)
Metallic Compounds	Involve metal atoms sharing electrons	Brass (Cu+Zn)

1.8 Valence:

It is defined as the combining capacity of the element. It helps determine how atoms join together to form molecules and compounds by gain, lose, or share of electrons. The valency of an element is related to the electronic configuration of its atom and usually determined by the electrons present in **the outermost shell** (valence electrons). Atoms try to achieve a **stable configuration** (usually 8 electrons in the outer shell—called the octet rule). Elements with 1, 2, or 3 valence electrons tend to **lose** them. Elements with 5, 6, or 7 valence electrons tend to **gain** or **share** electrons **(Table - 1.4)**

Valency helps predict:

- How elements bond (ionic or covalent)
- The formulas of compounds (e.g., H_2O, CO_2)
- Chemical reactions and equations

1.9 Formula:

A formula is an abbreviation or shortened form for the full name of a compound. The position of the numbers in chemical formulas provides specific information about a compound. The number preceding symbol of an element indicates no of atoms or molecules. Subscript numbers in chemical formula represent the number of atoms or molecules immediately preceding the subscript.

Table - 1.4 : Examples of Valency

Element	Symbol	No. of valence electrons	Valency	Reason
Hydrogen	H	1	1	Needs 1 electron to complete shell
Calcium	Ca	2	2	Loses 2 electrons
Aluminium	Al	3	3	Loses 3 electrons
Carbon	C	4	4	Shares 4 electrons
Nitrogen	N	5	3	Needs 3 electrons to complete shell
Oxygen	O	6	2	Needs 2 electrons to complete shell
Chlorine	Cl	7	1	Needs 1 electrons to complete shell
Argon	Ar	8	0	Needs no electron to complete shell

For example,

- Carbon-di-oxide, formula is CO_2, one carbon atom and 2 oxygen atoms in a compound Carbon-di-oxide.
- The chemical name for common salt is Sodium chloride. Symbol for Sodium is Na and symbol for Chlorine is Cl. Both sodium and chlorine have one valency. Therefore, chemical formula for sodium chloride is NaCl.
- The chemical name for dry cleaning solvent is Carbon Tetrachloride. Symbol for Carbon is C and for Chlorine is Cl. Valence of carbon is 4, and for chlorine is 1. Therefore, chemical formula for Carbon Tetrachloride is CCl_4.
- Similarly, formula for aluminum sulphate is $Al_2(SO_4)_3$, Symbol for aluminum is Al and for sulphate is SO_4. Valence for Aluminum is 3 and sulphate is 2 **(Table - 1.5).**

1.9 Mixtures:

These are made up of two or more pure substances. In chemistry, a **mixture** refers to a physical combination of two or more pure substances where each substance retains its own chemical identity and properties. Unlike compounds, mixtures do not involve chemical bonding between components. The substances in a mixture can be present in any proportion and can be separated using physical methods such as filtration, evaporation, or decanting.

Mixtures are broadly classified into two types: homogeneous and heterogeneous.

- A homogeneous mixture has a uniform composition throughout and appears as a single phase. The individual components are not visibly distinct.

Table - 1.5 : Examples of Common Compounds

Compound Name	Formula	Composition & Explanation
Carbon Dioxide	CO_2	1 Carbon + 2 Oxygen atoms. Carbon valency = 4, Oxygen = 2
Sodium Chloride	NaCl	1 Sodium + 1 Chlorine. Both have valency 1'
Carbon Tetrachloride	CCl_4	1 Carbon + 4 Chlorine. Carbon valency = 4, Chlorine = 1
Aluminum Sulphate	$Al_2 (SO_4)_3$	2 Aluminum (valency 3) + 3 Sulphate ions (valency 2)
Water	H_2O	2 Hydrogen + 1 Oxygen. Hydrogen valency = 1, Oxygen = 2
Ammonia	NH_3	1 Nitrogen + 3 Hydrogen. Nitrogen valency = 3
Calcium Carbonate	$CaCO_3$	1 Calcium + 1 Carbon + 3 Oxygen. Used in chalk and limestone
Glucose	$C_6 H_{12} O_6$	6 Carbon + 12 Hydrogen + 6 Oxygen. A simple sugar
Sulfuric Acid	$H_2 SO_4$	2 Hydrogen + 1 Sulfur + 4 Oxygen. Strong industrial acid

Examples:

- Salt dissolved in water
- Sugar in tea
- Air (mixture of gases)
- Brass (copper + zinc alloy)

On the other hand, a heterogeneous mixture has a non-uniform composition, and its different parts are visibly distinct. These mixtures are easier to separate.

Examples:

- Oil and water
- Sand and iron filings
- Salad
- Cereal in milk

Mixtures can also be categorized based on the size of their particles.

- A **solution** is a type of mixture where the particles are very small and completely dissolved is a type of mixture where the particles are very small and completely dissolved, such as saltwater or vinegar in water.
- A **colloid** contains intermediate-sized particles that do not settle under gravity; examples include milk, blood, and smoke.
- A **suspension** has large particles that settle over time, like muddy water or sand in water. contains intermediate-sized particles that do not settle under gravity; examples include milk,

blood, and smoke. A **suspension** has large particles that settle over time, like muddy water or sand in water.

The properties of mixtures are quite flexible. The components can be mixed in any ratio, and there is no energy change during their formation. Unlike pure substances, mixtures do not have fixed melting or boiling points; These vary depending on the composition. Mixtures can exist in solid, liquid, or gaseous states, making them a versatile concept in both laboratory and industrial chemistry.

1.10 Properties of matter and their Measurements

Every substance has unique properties.

These properties can be classified into two categories – Physical properties and Chemical properties.

Physical properties are those properties which can be measured or observed without changing the identity or the composition of the substance.

Examples:

- Mass
- Volume
- Density
- Color
- Odor
- Melting and boiling points
- Electrical conductivity
- Hardness

These properties help us identify and classify substances. For instance, water has a density of 1 g/cm^3 and boils at 100°C under standard conditions.

Chemical properties are those properties which chemical properties a substance reacts or changes into a new substance.

Examples:

- Reactivity with acids or bases
- Flammability
- Combustibility
- Acidity or basicity
- Oxidation (eg, rusting of iron)

Chemical properties are observed during chemical reactions. For example, iron reacts with oxygen to form rust (Fe_2O_3)

1.11 Units

Units are Measurement of Properties. Accurate measurement is essential in chemistry to understand and compare substances. Many properties of matter such as length, area, volume, etc., are quantitative in nature. Any quantitative measurement is represented by a number followed by units in which it is measured.

Units are of two types:

(i) **Basic units:**

The basic or fundamental units are those of length (m), Mass (kg), time (s), electric current (A), thermodynamic temperature (K), amount of substance (mol) and luminous intensity (cd).

(ii) **Derived units**:

Derived units are basically derived from the fundamental units, e.g., unit of density is derived from units of mass and volume. The systems used for describing measurements of various physical quantities are

(a) **CGS system-** It is based on centimeter, gram and second as the units of length, mass and time respectively.

(b) **FPS system-** A British system which used foot (ft). Pound (lb) and second (s) as the fundamental units of length, mass and time.

(c) **MRS system-** Uses meter (m), kilogram (kg) and second (s) respectively for length, mass and time; ampere (A) was added later on for electric current.

(d) **SI system** - (1960) International system of units and contains following seven basic units:

The SI system has seven base units and they are listed in Table

Table - 1.6 : SI Units

No	Physical Quantity	Symbol for quality	Name of the S.I unit	Symbol of the S.I unit	Description
1.	Length	l	meter	m	Measures distance or size
2.	Mass	m	kilogram	kg	Measures amount of matter
3.	Time	t	second	s	Measures duration
4.	Current	I	ampere	A	Measures flow of electric charge
5.	Temperature	T	kelvin	K	Measures thermal energy
6.	Amount of substance	n	mole	mol	Measures number of particles
7.	Luminous Intensity	I_v	candela	cd	Measures brightness of light

1.12 Atomic & molecular masses:

Atomic mass is the mass of a single atom, usually mass units **Daltons (Da)**. Atomic mass is approximately equal to the sum of **protons and neutrons** in the nucleus. It represents the average mass of all naturally occurring isotopes of an element, weighted by their abundance.

The atomic mass or the mass of an atom is based on C-12 as the standard. ^{12}C is assigned a mass of exactly 12 atomic mass unit (amu) and masses of all other atoms are given relative to this standard. One atomic mass unit is defined as a mass exactly equal to 1/12 the mass of one C-12 atom.

And 1 amu = 1.66056×10^{-24} g

Mass of an atom of hydrogen = 1.6736×10^{-24} g

Thus, in terms of amu, the mass of hydrogen atom = $\frac{1.6736X10^{-24}\ g}{1.66056\ 10^{-24}\ g}$ = 1.008 amu

Similarly, the mass of oxygen - 16 (^{16}O) atom would be 15.995 amu.

- Hydrogen (H): ~1.008 amu
- Carbon (C): ~12.01 amu
- Oxygen (O): ~16.00 amu

Molecular mass is the sum of atomic masses of the elements present in a molecule. It tells us how heavy the molecule is compared to the carbon-12 standard.

Molecular Mass = ∑(Atomic Mass of each atom X Number of atoms)

- Molecular mass of water (H_2O) = 2 × atomic mass of H + 1 × atomic mass of O
  = 2 (1.008) + 16.00 = 18.02 amu
- Molecular mass of (CO_2) = 1 × atomic mass of C +2 × atomic mass of O
  = 1 (12) + 2(16) = 44 amu
- Molecular mass of (NaCl) = 1 × atomic mass of Na+1 × atomic mass of Cl
  = 1 (23) + 1(35.5) = 58.5 amu
- Molecular mass of (CCl_4) = 1 × atomic mass of C+4 × atomic mass of Cl
  = 1 (12) + 4(35.5) = 154amu

1.14 Percentage Composition:

Percentage composition is the term used to describe the percentage by mass of each element in a compound. The percentage composition of elements can be calculated as follows:

$$\text{Mass \% of an element} = \frac{\text{Mass of that element in the compound} \times 100}{\text{Molar mass of the compound}}$$

For example

- Water (H_2O) which contains hydrogen and oxygen, the percentage composition of both these elements can be calculated

Molar mass of water = 18.02 g

Mass % of hydrogen = $\frac{2 \times 1.008 \times 100 = 11.18}{18.02}$

Mass % of oxygen = $\frac{16.00 \times 100}{}$ = 88.79

- Ethanol (C_2H_5OH) which contains carbon, hydrogen and oxygen, the percentage composition of these elements can be calculated

Molecular formula of ethanol is: C_2H_5OH

Molar mass of ethanol is: (2×12.01 + 6×1.008+1x 16.00) g= 46.068 g

Mass per cent of Carbon = $\frac{24.02g \times 100 = 52.14\%}{46.068g}$

Mass per cent of Hydrogen = $\frac{6 \times 1.008 \times 100 = 13.13\%}{46.068g}$

Mass per cent of Oxygen = $\frac{16 \times 100}{46.068g}$ = 34.73%%

1.15 Dimension for physical quantities:

Dimensions represent nature of a physical quantity in terms of the fundamental units. They show how a quantity is derived from base units like mass (M), length (L), time (T), temperature (Θ), electric current (I), luminous intensity (J), and amount of substance (N). The dimension of a physical quantity is defined as the powers to which the fundamental quantities are raised in order to represent that quantity.

Dimensions are expressed in square brackets: $[M^aL^bT^c]$.They help verify equations, convert units, and understand relationships between quantities. **(Table - 1.7)**

Dimensions of common derived quantities

1. Area- It is Products of length and breadth

   Area= Length x Breadth

   Dimension for Area [A] is $[L^2]$

2. Volume- It is the amount of space occupied by an object.

   Volume = Length x Area

   Dimension for Volume [v] is $[L^3]$

Table - 1.7 : Fundamental quantities and their dimensions

Quantity	Symbol	Dimension
Mass	M	[M]
Length	L	[L]
Time	T	[T]
Temperature	Θ	[Θ]
Electric Current	I	[I]
Luminous Intensity	J	[J]
Amount of Substance	N	[N]

3. Velocity –It is change in distance per unit time
   Velocity=distance/ time
   Dimension for velocity [v] is $[LT^{-1}]$
4. Acceleration- It is change in velocity per unit time
   Acceleration= Velocity /time
   Dimension for Acceleration[a] = $[LT^{-2}]$
5. Force – It is product of mass and acceleration
   Force = Mass x Acceleration
   Dimension for Force [F] is $[MLT^{-2}]$
6. Pressure-It is force per unit area
   Pressure =Force/Area
   Dimension for Pressure [P] is $[ML^{-1}T^{-2}]$
7. Coefficient of viscosity (η)= $\dfrac{\text{Force x distance between the layers}}{\text{Area x velocity}}$

   Dimension for Viscosity= $\dfrac{[MLT^{-2}]\,[L]}{[L^{2}]\,[LT^{-1}]} = [ML^{-1}T^{-1}]$
8. Surface Tension (T) = Force / Length.
   Dimension for Surface tension = $[MLT^{-2}]/[L]=[MT^{-2}]$ **(Table - 1.8)**

1.16 Laws of Chemical Combinations:

When two or more atoms of different elements combine, a compound is formed. The combination of different elements to form compounds is governed by certain basic rules. These rules are called

Table - 1.8 : Dimensions

Physical Quantity	Formula / Relation	Dimensional Formula
Area	Length × Breadth	$[L^2]$
Volume	Length^3	$[L^3]$
Speed	Displacement / Time	$[L T^{-1}]$
Acceleration	Velocity / Time	$[L T^{-2}]$
Force	Mass × Acceleration	$[M L T^{-2}]$
Pressure	Force / Area	$[M L^{-1} T^{-2}]$
Work / Energy	Force× Distance	$[M L^2 T^{-2}]$
Power	Work / Time	$[M L^2 T^{-3}]$
Density	Mass / Volume	$[M L^{-3}]$
Momentum	Mass × Velocity	$[M L T^{-1}]$
Frequency	1 / Time	$[T^{-1}]$

as laws of chemical combinations. The combination of elements to form compounds is governed by the following five basic laws:

1. **Law of conservation of mass** (Lavoisier, 1774)

   This law was put forth by *Antoine Lavoisier*. This law states that during any physical or chemical change, mass is neither created nor destroyed during a chemical reaction, meaning the total mass of reactants equals the total mass of products. It does not hold good for nuclear reactions.

2. **Law of definite proportions** (Proust, 1799)

   This law was given by, a French chemist, *Joseph Proust*. According to this law, a chemical compound obtained by different sources always contains same percentage of each constituent element.

3. **Law of multiple proportions** (Dalton, 1803)

   This law was proposed by *Dalton*. According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers, e.g., in NH_3 and N_2H_4, fixed mass of nitrogen requires hydrogen in the ratio 3: 2.

4. **Gay Lussac's law of gaseous volumes** (1808)

   This law was given by *Gay Lussac*. It states that under similar conditions of temperature and pressure, whenever gases react together, the volumes of the reacting gases as well as products (if gases) bear a simple whole number ratio.

5. **Avogadro's hypothesis** (1811)

   It states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.

   Together, these laws provide a consistent framework for understanding chemical reactions and the composition of substances.

1.16 Dalton's Atomic Theory (1803)

Dalton's Atomic Theory, proposed in 1808, laid the foundation for modern chemistry by introducing the concept that all matter is composed of indivisible atoms. John Dalton, an English chemist and physicist, formulated his atomic theory to explain the laws of chemical combination.

Its basic postulates are

1. All substances are made up of tiny indivisible particles, called atoms.
2. In each element, the atoms are all alike and have the same mass. The atoms of different elements differ in mass.
3. Atoms can neither be created nor destroyed during any physical or chemical change
4. Compounds result from combination of atoms in simple numerical ratio

Dalton's theory successfully explained the **Law of Conservation of Mass**, **Law of Definite Proportions**, and **Law of Multiple Proportions**, offering a coherent framework for understanding chemical reactions.

However, it had limitations. For instance, Dalton believed atoms were indivisible, but later discoveries of subatomic particles—electrons, protons, and neutrons—proved otherwise. He also assumed that all atoms of an element have the same mass, which was contradicted by the discovery of **isotopes**—atoms of the same element with different masses. Additionally, Dalton's theory couldn't explain the existence of **isobars** (atoms of different elements with the same mass) or **allotropes** (different forms of the same element, like diamond and graphite).

Despite its shortcomings, Dalton's Atomic Theory remains a milestone in the history of science, marking the transition from philosophical speculation to empirical understanding of matter.

QUESTIONS

I. Questions Carrying Two Marks

1. Define matter and classify it based on its physical states.
2. State the Law of Conservation of Mass.
3. Write the chemical formula for calcium carbonate and sodium chloride.
4. Give the SI unit for mass and temperature.
5. Calculate the molecular mass of H_2O

II. Questions Carrying Three Marks

6. Explain the significance of chemistry in daily life with two examples.
7. Differentiate between physical and chemical properties of matter with examples.
8. State and explain the Law of Definite Proportions.
9. Convert 1atmosphere pressure to Pascals, given 1 atm=101325 Pa
10. Determine the percentage composition of carbon in methane (CH_4)

III. Questions Carrying Five Marks

11. Discuss the main postulates of Dalton's Atomic Theory and its limitations.
12. Derive the dimensional formula for pressure and viscosity.
13. A compound contains 4.07% hydrogen, 24.27% carbon, and71.65% chlorine. Determine its empirical formula.
14. Explain the importance of SI units in scientific measurements and provide examples of base SI units.
15. Expain Atomic weight and Molecular weight with suitable example.

IV. Choose the Correct Answer

i) Which of the following is an example of matter?

a) Light
b) Sound
c) Air
d) Heat

ii) Which law states that mass is conserved in a chemical reaction?

a) Law of Definite Proportions
b) Law of Conservation of Mass
c) Law of Multiple Proportions
d) Dalton's Atomic Theory

iii) How is molecular mass calculated?

a) By adding the atomic masses of all atoms in a molecule

b) By dividing the atomic masses of all atoms in a molecule

c) By multiplying the atomic masses of all atoms in a molecule

d) By subtracting the atomic masses of all atoms in a molecule

iv) Who proposed the atomic theory stating that atoms are indivisible and indestructible?

a) John Dalton

b) Dmitri Mendeleev

c) J.J. Thomson

d) Ernest Rutherford

v) How is percentage composition calculated?

a) By dividing the mass of each element by the total mass of the compound and multiplying by 100.

b) By dividing the atomic number by the atomic mass.

c) By adding the masses of all elements in the compound.

d) By subtracting the mass of the compound from the mass of the element.

V. Match the Following

A	B
a) Pressure	i) Kilogram
b) Viscosity	ii) Pascal
c) Mass	iii) Pascal -Second
d) Current	iv) Kelvin
e) Temperature	v) Ampere