Role of chemistry in different spheres of life

What is Chemistry?

- Chemistry is the science of atoms, molecules and their transformations.
- It can also be defined as that branch of science that deals with the study of the composition, properties and interaction of matter.
- Knowledge of chemistry helps us to know:-

1. The properties of different substances like diamond, coal, plastics etc.
2. In preparation of soaps like liquid or powder.
3. The fabrics, plastic toys, paper, pencil, eraser, sugar cubes, petrol or diesel etc.
4. In the manufacture of different types of fertilizers, pesticides and medicines.
5. Answers to questions regarding how many atoms or molecules are there in a given mass of matter.

<u>Chemistry helps us to answer the following questions</u>:-

- To count the number of atoms and molecules in a given mass of matter?
- How to see the atoms and molecules?

<u>Advantages of studying chemistry</u>

1. With a better understanding of chemical principles it has now become possible to design and synthesise new materials having specific magnetic, electric and optical properties.
2. Safer alternatives to environmentally hazardous refrigerants like CFCs (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesised.

<u>Problems which can be tackled with the study of chemistry:-</u>

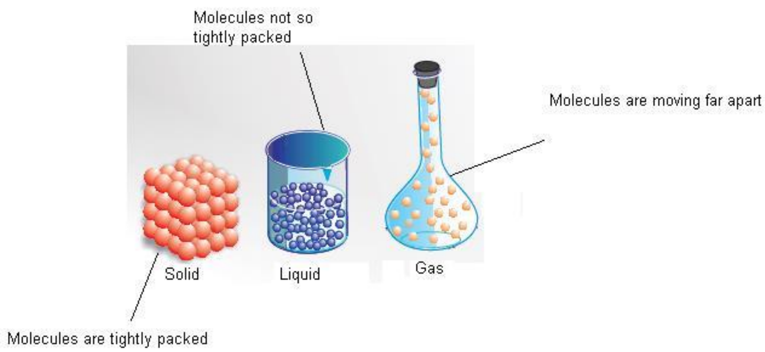
1. Management of Green House gases like methane, carbon dioxide etc.
2. Understanding of bio-chemical processes, use of enzymes for large-scale production of chemicals and synthesis of new exotic materials are some of the intellectual challenges for the future generation of chemists.
3. Understanding of bio-chemical processes, use of enzymes for large-scale production of chemicals.

Matter

- Anything that has mass and occupies space is defined as matter.
    - For example:-soap, notebook, pencil, apple, battery etc.

<u>Physical States of matter</u>

- Matter can exist in 3 physical states namely solid, liquid and gas.



Solids

- In case of solids, the constituent particles are held very close to each other in an orderly fashion and there is not much freedom of movement.
- Solids have definite shape and definite volume.
    - For example: - Apple it has definite shape as well as definite volume.

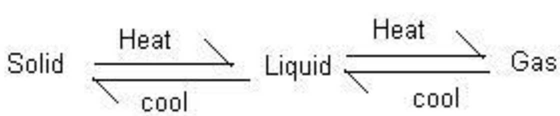
Liquids

- In liquids, the particles are close to each other but they can move around.
- Liquids have definite volume but not definite shape. They take the shape of the container in which they are placed.
    - For example: Water it takes the shape of the tumbler in which it is poured but has volume.

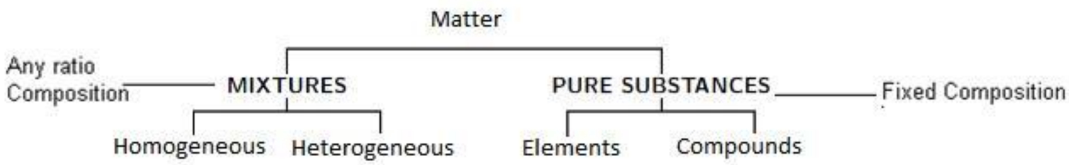
Gases

- In gases, the particles are far apart as compared to those present in solid or liquid states and their movement is easy and fast.
- If container in which they are placed.
    - For example: - Smoke does not have definite shape or volume.

- **Note**:-
- All the 3 states of matter are inconvertible among each other.
- Consider solid if we melt it , it changes into liquid and if it is further heated it becomes
- When the gas is condensed they change to liquid. And when the liquid is freezed it becomes solid.
- By sublimation solid changes into gas and gas by deposition it becomes



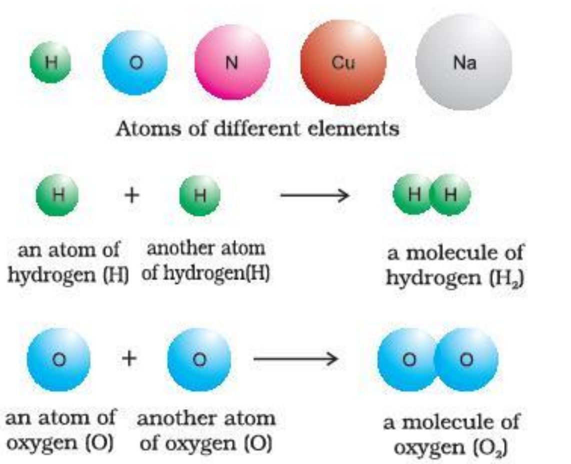
Matter at macroscopic level



Pure Substances

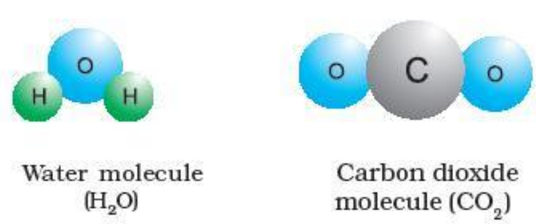
- Pure substances have fixed composition.
  - For example: - Copper, silver, gold, water, glucose are some examples of pure substances.
    - Glucose contains carbon, hydrogen and oxygen in a fixed ratio and thus, like all other pure substances has a fixed composition.
    - Composition of carbon, hydrogen and oxygen in sugar will be always $C_{12}H_{22}O_{11}$.They will be always in fixed ratio.
  - Also, the constituents of pure substances cannot be separated by simple physical methods.
  - Pure substances are further classified into <u>elements</u> and <u>compounds</u>.
  - An <u>element</u> consists of only one type of particles. These particles may be atoms or molecules.
  - For example: - Sodium (Na), copper (Cu), silver (Ag), hydrogen (H), oxygen (O) atoms etc. They contain only one type of atoms.
  - When two or more atoms of different elements combine, the molecule of a <u>compound</u> is obtained.
  - For example: - Water (H_20), Ammonia (NH_3), Sugar, carbon dioxide (CO_2).

Representation of atoms and molecules



- The properties of a compound are different from those of its constituent elements.

For example: - Hydrogen (H_2) and Oxygen (O_2) are gases whereas the compound formed by their combination i.e., water (H_20) is a liquid.



- Hydrogen burns with a pop sound and oxygen is a supporter of combustion, but water is used as a fire extinguisher.
- Constituents of a compound cannot be separated into simpler substances by physical methods. They can be separated by chemical methods.
- A <u>mixture</u> contains two or more substances present in it (in any ratio) which are called its components.
- For example:-Air, sugar solution, mixture of pulse and stone.


Water


Mixture Constitutents are mixed in any ratio

- A mixture can be <u>homogeneous</u> or <u>heterogeneous</u>.
- In homogeneous mixture, the components completely mix with each other and its composition is uniform throughout.
- For example: - Air, Sugar solution.
- In heterogeneous mixtures, the composition is not uniform throughout and sometimes the different components can be observed.
- For example: - Mixture of dal mot and bhujia, badam etc.
  - The components of a mixture can be separated by using physical methods such as simple hand picking, filtration, crystallisation, distillation etc.

Properties of Matter
- Every substance has characteristic properties. These properties can be classified into 2 categories: - Physical and Chemical properties.

<u>Physical properties</u>
1. They are those properties which can be measured or observed without changing the identity or the composition of the substance.
2. Some examples of physical properties are colour, odour, melting point, boiling point, density etc.

<u>Chemical properties</u>
1. The chemical properties require a chemical change to occur.
2. Chemical properties are characteristic reactions of different substances; these include acidity or basicity, combustibility etc.

Quantitative property
- Many properties of matter such as length, area, volume, etc., are quantitative in nature.
- Any quantitative observation or measurement is <u>represented by a number</u> followed by units in which it is measured.
  - For example: - length of a room can be represented as 6 m; here 6 is the number and m denotes metre – the unit in which the length is measured.

Two different systems of measurement are as follows:-

1. English System
2. Metric System:

- The metric system which originated in France in late eighteenth century was more convenient as it was based on the decimal system.

The International System of Units (SI)

- The International System of Units (in French Le Systeme International d'Unités – abbreviated as SI) was established by the 11th General Conference on Weights and Measures.
- The SI system has seven base units and they are listed in figure 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.

<u>Base Physical Quantities and their Units</u>

Base Physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

<u>Definitions of SI Base Units</u>

Quantity	Unit	Symbol
Length	meter	m
Mass	Kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	k
Quantity of substance	mole	mol
Luminosity	candle	cd

Prefixes used in SI System

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	µ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

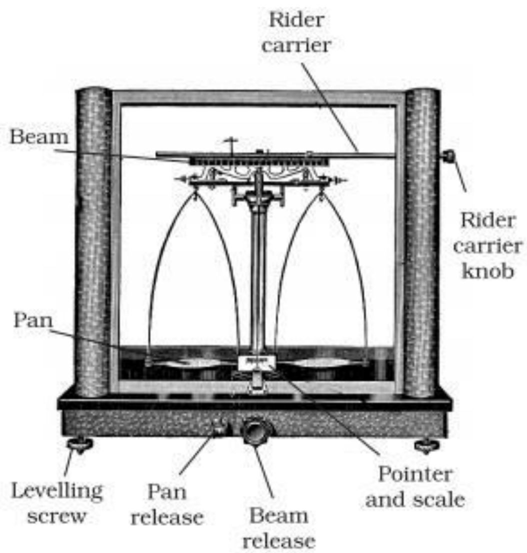
Problem:-

Convert the following into basic units:
(i) 28.7 pm
(ii) 15.15 pm
(iii) 25365 mg
Answer:-
(i) 28.7 pm:
1 pm = 10^{-12} m
28.7 pm = 28.7×10^{-12} m
= 2.87×10^{-11} m

(ii) 15.15 pm:
1 pm = 10^{-12} m
15.15 pm = 15.15×10^{-12} m
= 1.515×10^{-12} m
(iii) 25365 mg:
1 mg = 10^{-3} g
25365 mg = $2.5365 \times 10^{4} \times 10^{-3}$ g
Since,
1 g = 10^{-3} kg
2.5365×101 g = $2.5365 \times 10^{-1} \times 10^{-3}$ kg
25365 mg = 2.5365×10^{-2} kg
Problem:-
If the speed of light is 3.0×10^{8} m s^{-1}, calculate the distance covered by light in 2.00 ns.
Answer:-
According to the question:
Time taken to cover the distance = 2.00 ns
= 2.00×10^{-9} s
Speed of light = 3.0×10^{8} ms^{-1}
Distance travelled by light in 2.00 ns
= Speed of light × Time taken
= $(3.0 \times 10^{8}\ ms^{-1})\ (2.00 \times 10^{-9}\ s)$
= 6.00×10^{-1} m
= 0.600 m

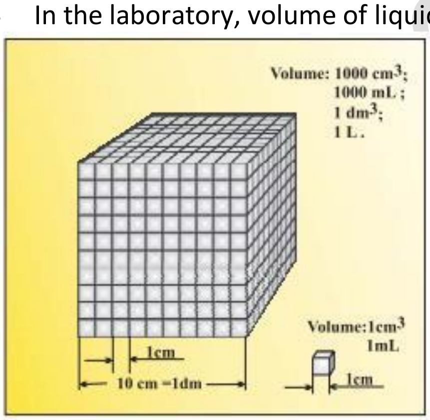
Mass & Weight
- Mass is defined as the amount of matter present in a substance.
- Weight is defined as the force exerted by the gravity on an object.
- The mass of a substance is constant whereas its weight may vary from one place to another due to change in gravity.
- The mass of a substance can be determined very accurately in the laboratory by using an analytical balance.
- The SI unit of mass is kilogram.

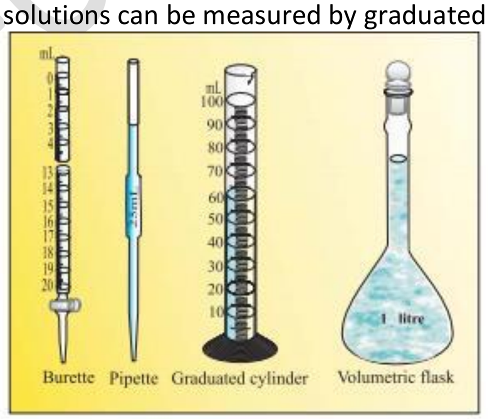
Analytical Balance



Volume

- Volume has the units of (length) ^{3}. So in SI system, volume has units of m^3.
- A common unit, litre (L) which is not an SI unit, is used for measurement of volume of liquids.
- 1 L = 1000 mL, 1000 cm^3 = 1 dm^3.
- In the laboratory, volume of liquids or solutions can be measured by graduated cylinder, burette, pipette etc.





<u>Maintaining the National Standards of Measurement</u>

- Each country has a National Metrology Institute (NMI) which maintains standards of measurements.
- For India it is National Physical Laboratory (NPL), New Delhi.
- This laboratory establishes experiments to realize the base units and derived units of measurement and maintains National Standards of Measurement.
- These standards are periodically inter-compared with standards maintained at other National Metrology Institutes in the world as well as those established at the International Bureau of Standards in Paris.

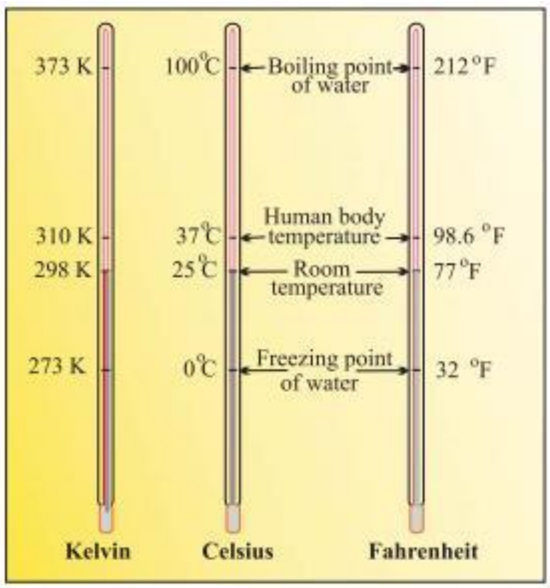
Density & Temperature

- Density of a substance is its amount of mass per unit volume.
- SI unit of density = kg/m^3.

Temperature

- There are three common scales to measure temperature — °C (degree Celsius), °F (degree Fahrenheit) and K (kelvin). Here K is the unit SI unit.
- The temperatures on two scales are related to each other by the following relationship:-
  - $^{O}F = (9/5)(^{O}C) + 32$
- The Kelvin scale is related to Celsius scale as follows:-
  - $K = {}^{O}C + 273.15$
- **Note**: - The temperatures below 0 °C (i.e. negative values) are possible in Celsius scale but in Kelvin scale, negative temperature is not possible.

Thermometers using different temperature scales



Reference Standard

- Measuring devices we use is standardised or calibrated against some reference.
- 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.
- The metre was originally defined as the length between two marks on a Pt-Ir bar kept at a temperature of 0°C (273.15 K)
- In 1960 the length of the metre was defined as 1.65076373×10^6 times the wavelength of light emitted by a krypton laser.
- Although this was a cumbersome number, it preserved the length of the metre at its agreed value.
- The metre was redefined in 1983 by CGPM as the length of path travelled by light in vacuum during a time interval of (1/299 792 458) of a second.

Uncertainty in measurement

- Since the chemistry is the study of atoms and molecules which have extremely low masses and are present in extremely large numbers, a chemist has to deal with numbers as large as 602,200,000,000,000,000,000,000 for the molecules of 2 g of hydrogen gas or as small as 0.00000000000000000000000166 g mass of a H atom.
- This problem is solved by using scientific notation for such numbers, i.e., exponential notation in which any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values and N can vary between 1 to 10.

Precision & Accuracy

- Every experimental measurement has some amount of uncertainty associated with it.
- Everyone wants the results to be precise and accurate.
- Precision and Accuracy are often referred to whenever we talk about measurement.
- Precision refers to the closeness of various measurements for the same quantity. It does not depend on true value.
- Accuracy is the agreement of a particular value to the true value of the result. It depends on the true value.

Data to illustrate Precision and Accuracy

Student	Exp1 output	Exp2 output	Mean	Is Precise	Is Accurate
Ramesh	2.80	3	2.9	No	No
Suresh	2.91	2.90	2.905	Yes	No
Rohit	3.01	2.99	3	Yes	Yes
Rakesh	2.90	3.10	3	No	Yes

Significant figures

- The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures.
- Significant figures are meaningful digits which are known with certainty.
- There are certain rules for determining the number of significant figures. They are as follows:

1. All non-zero digits are significant.
   - For example: - There are 4 significant figures in 328.2 cm.
2. Zeros preceding to first non-zero digit are not significant.
   - For example: - 0.03 has one significant figure and 0.0052 has two significant figures.
3. Zeros between two non-zero digits are significant.
   - For example: - 2.005 have four significant figures.
4. Zeros at the end or right of a number are significant provided they are on the right side of the decimal point.
   - For example: - 200 g has three significant figures. But, if otherwise, the zeros are not significant. For example, 100 have only one significant figure.
5. Exact numbers have an infinite number of significant figures.
   - For example:- In 2 balls or 20 eggs, there are infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal i.e.,2 = 2.000000 or 20 = 20.000000.

Addition & Subtraction

- The result cannot have more digits to the right of the decimal point than either of the original numbers.
  - For example: - (12.11 + 18.0+ 1.012) = 31.122. As 18.0 have only one digit after the decimal point therefore the result will be 31.1, one digit after the decimal point.

Multiplication & Division

- In these operations, the result must be reported with no more significant figures as are there in the measurement with the few significant figures.

- For example: - $2.5 \times 1.25 = 3.125$. Answer will be 3.1 because 2.5 have two significant figures and the result should not have more than two significant figures.

Problem:-
How many significant figures are present in the following?
(i) 0.0025
(ii) 208
(iii) 5005
(iv) 126,000
(v) 500.0
(vi) 2.0034
Answer:-
(i) 0.0025
There are 2 significant figures.
(ii) 208
There are 3 significant figures.
(iii) 5005
There are 4 significant figures.
(iv) 126,000
There are 3 significant figures.
(v) 500.0
There are 4 significant figures.
(vi) 2.0034
There are 5 significant figures.
Problem:-
Round up the following upto three significant figures:
(i) 34.216
(ii) 10.4107
(iii) 0.04597
(iv) 2808
Answer:-
(i) 34.2
(ii) 10.4
(iii) 0.0460
(iv) 2810

Dimensional Analysis
- When calculating, there is a need to convert units from one system to other.
- The method used to accomplish this is called factor label method or unit factor method or dimensional analysis.
    - For example: - A jug contains 2L of milk. Calculate the volume of the milk in m^3.

Problem:-
A jug contains 2L of milk. Calculate the volume of the milk in m^3.
Answer:-
Since 1 L = 1000 cm^3 and 1m = 100 cm which gives
$(1\ m)/(100\ cm) = 1 = (100\ cm/1m)$
To get m^3 from the above unit factors, the first unit factor is taken and it is cubed.
Therefore, $(1m)^3 / (100cm)^3 = 1 = (1m^3)/(10^6\ cm^3)$ (unit factor)

Now 2 L of milk = $2 \times 1000\ cm^3$
The above is multiplied by the unit factor
= $2 \times 1000\ cm^3 \times (1\ m^3/10^6\ cm^3)$
= $2 \times 10^{-3}\ m^3$

Law of chemical combination

- The combination of elements to form compounds is governed by the following five basic laws:-

1. Law of Conservation of Mass.
2. Law of Definite Proportions.
3. Law of Multiple Proportions.
4. Gay Lussac's Law of Gaseous Volumes.
5. Avogadro Law.

Law of Conservation of Mass

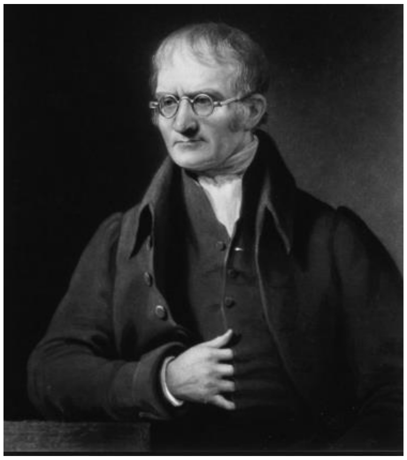
- Law of conservation of mass states that the matter cannot be created nor be destroyed.
- This law was put forth by Antoine Lavoisier in 1789.
- He performed careful experimental studies for combustion reactions for reaching to the above conclusion.

Antoine Lavoisier

Law of Multiple Proportions

- 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.
    - For example: - Hydrogen combines with oxygen to form 2 compounds, water and hydrogen peroxide.
    - H_2(2g) + (1/2)O_2 (16g) à H_2O (18g)
    - H_2(2g) + O_2 (32g) àH_2O_2 (34g)
    - The masses of oxygen O (16g and 32g) combine with the fixed mass of (2g) hydrogen H. Therefore the simple ratio is 16:32 or 1:2.
- This law was given by Dalton in 1803.

John Dalton



Law of Definite Proportions

- According to this law, a given compound always contains exactly the same proportion of elements by weight.
- This law was given by French chemist, Joseph Proust in 1806.
- He observed 2 samples of cupric carbonate.
    - One was of natural origin and another was of synthetic origin.
- He found that the composition of elements present in it was same for both the samples as shown below:-

	% of copper	% of oxygen	% of carbon
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

- Note: - It is sometimes also referred to as law of definite composition.

Joseph Proust

Gay Lussac's Law of Gaseous Volumes

- Gay Lussac's 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.
    - For example: - H (Hydrogen) (100mL) + O (Oxygen) (50mL)à Water (100mL).
    - The volumes of hydrogen (H) and oxygen (O) which combine together (i.e. 100mL and 50mL) bear a simple ratio of 2:1.

Gay Lussac

Avogadro Law

- In 1811, Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.
- He made distinction between atoms and molecules.

Problem:-

The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

Mass of dinitrogen	Mass of dioxygen
(i) 14 g	16 g
(ii) 14 g	32 g
(iii) 28 g	32 g
(iv) 28 g	80 g

(a)Which law of chemical combination is obeyed by the above experimental data? Give its statement.

(b)Fill in the blanks in the following conversions:

(i) 1 km = ... mm = ... pm

(ii) 1 mg = ... kg = ... ng

(iii) 1 mL =..................... L = ... dm^3

Answer:-

(a)

If we fix the mass of dinitrogen at 28 g, then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 32 g, 64 g, 32 g, and 80 g.

The masses of dioxygen bear a whole number ratio of 1:2:2:5. Hence, the given experimental data obeys the law of multiple proportions. The law states that if two elements combine to form more than one compound, then the masses of one element that combines with the fixed mass of another element are in the ratio of small whole numbers.

(b)

(i) 1 km = 1 km × (1000m/1km) x (100cm/1m) x (10mm/1cm)

Therefore 1 km = 10^6 mm

1 km = 1 km × (1000m/1km) x (1pm/10^{-12}m)

Therefore 1 km =10^{15} pm

Hence, 1 km = 10^6 mm = 10^{15} pm

(ii) 1 mg = 1 mg × (1 g/1000mg) x (1kg/1000g)

⇒ 1 mg = 10^{-6} kg

1 mg = 1 mg × (1 g /1000mg) x (1 ng /10^{-9}g)

⇒ 1 mg = 10^6 ng

1 mg = 10^{-6} kg = 10^6 ng

(iii) 1 mL = 1 mL × (1L/1000mL)

⇒ 1 mL = 10^{-3} L

1 mL = 1 cm^3 = 1 cm^3 = 1 x ((1 dm x 1dm x 1dm) / (10cm x 10cm x10 cm)) cm^3

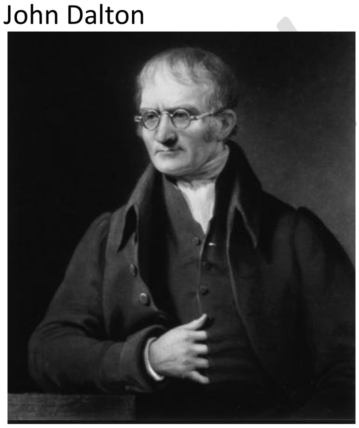
⇒ 1 mL = 10^{-3} dm^3

1 mL = 10^{-3} L = 10^{-3} dm^3

Dalton's Atomic Theory

- Matter consists of indivisible atoms.
- All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
- Compounds are formed when atoms of different elements combine in a fixed ratio.
- Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

John Dalton



Atomic Mass

- Atomic mass is the mass of the atom.
- Nowadays sophisticated techniques e.g. mass spectrometry is used to determine the atomic mass of the atom.

- But in 19th century atomic mass was determined by calculating the mass of the atom relative to hydrogen by using stoichiometry.

History of Atomic Mass

- In 1803 John Dalton determined the atomic mass or atomic weight.
- Atomic weight was originally defined as relative to that of lightest element hydrogen taken as 1.
- Oxygen was taken as base as its atomic mass was 16(whole number).All the observations were based on stoichiometry.
- Chemists picked naturally occurring Oxygen, which is a mixture of isotopes Oxygen-16, Oxygen-17 and Oxygen-18 (Atomic mass 16.008).
- Physicists picked up carbon 12 as base based on mass spectrometry & mass of one carbon -12 atom is a whole number.
- Carbon-12 was agreed to be taken as standard in 1961 because the mass of one carbon -12 atom is a whole number.
- The masses of all other atoms are taken relative to mass of 12C.

Average Atomic Mass

- Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance (per cent occurrence).

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2 10^{-10}	14.00317

- From the above data, the average atomic mass of carbon will come out to be :
    - (0.98892) (12 u) + (0.01108) (13.00335 u) + (2 × 10^{-12}) (14.00317 u)
    - = 12.011 u

Problem:-

Calculate the atomic mass (average) of chlorine using the following data:-

Isotopes of Chlorine	% Natural Abundance	Molar Mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Answer:-

Atomic mass of first isotope = 34.9689
Natural abundance of first isotope = 75.77% or 0.757
Atomic mass of second isotope= 36.9659
Natural abundance of second isotope= 24.23% or 0.242
Now average atomic mass of chlorine
= [34.9689x 0.757 + 36.9659x 0.242]/ (0.757+0.242) =35.4521
So, the average atomic mass of chlorine = 35.4527 u

Problem:-

Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	Isotopic molar mass	Abundance
^{36}Ar	35.96755 $gmol^{-1}$	0.337%
^{38}Ar	37.96272 $gmol^{-1}$	0.063%

^{40}Ar 39.9624 $gmol^{-1}$ 99.600%

Answer:-

Molar mass of argon =

Atomic mass of ^{36}Ar = 35.96755 & abundance = 0.337

Or

Total atomic mass of ^{36}Ar = 35.96755 * 0.337= 0.121g/mol

Atomic mass of ^{38}Ar = 37.96272 & abundance = 0.063

Or

Total atomic mass of ^{38}Ar = 37.96272 * 0.063= 0.024g/mol

Atomic mass of ^{40}Ar =39.9624 & abundance = 99.600

Or

Total atomic mass of ^{40}Ar = 39.9624 * 99.600= 39.802g/mol

Therefore molar mass of argon = total mass of ^{36}Ar+ total mass of ^{38}Ar + atomic mass of ^{40}Ar

= 0.121 + 0.024 + 39.802= 39.947 g/mol

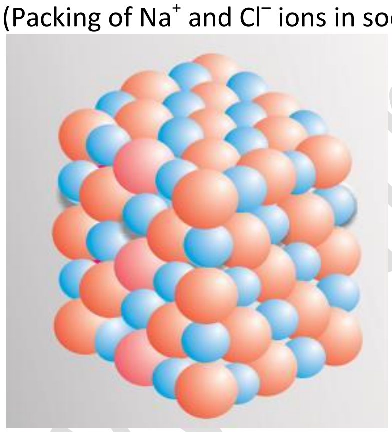
Molecular Mass

- Molecular mass is the sum of atomic masses of the elements present in a molecule.
- For example:-Molecular mass of methane CH_4 = (12.011 u) + 4 (1.008 u) = 16.043 u.

Formula Mass

- Some substances such as sodium chloride do not contain discrete molecules as their constituent units. E.g. NaCl.
- In such compounds, positive (sodium) and negative (chloride) entities are arranged in a three-dimensional structure as shown in the figure.
    - For example: - Formula mass of sodium chloride (NaCl) = atomic mass of sodium (Na) + atomic mass of chlorine (Cl) = 23.0 u + 35.5 u = 58.5 u.

(Packing of Na^+ and Cl^- ions in sodium chloride NaCl)



Problem:-

Calculate the molecular mass of the following:

(i) H_2O (ii) CO_2 (iii) CH_4

Answer:-

(i) H_2O:

The molecular mass of water, H_2O

= (2 × Atomic mass of hydrogen) + (1 × Atomic mass of oxygen)

= [2(1.0084) + 1(16.00 u)]

= 2.016 u + 16.00 u
= 18.016
= 18.02 u
(ii) CO_2:
The molecular mass of carbon dioxide, CO_2
= (1 × Atomic mass of carbon) + (2 × Atomic mass of oxygen)
= [1(12.011 u) + 2 (16.00 u)]
= 12.011 u + 32.00 u
= 44.01 u
(iii) CH_4:
The molecular mass of methane, CH_4
= (1 × Atomic mass of carbon) + (4 × Atomic mass of hydrogen)
= [1(12.011 u) + 4 (1.008 u)]
= 12.011 u + 4.032 u
= 16.043 u

Mole Concept and Molar Masses

- Atoms and molecules are extremely small in size but their numbers are very large even in a small amount of any substance.
- In order to handle such large numbers, a unit of similar magnitude is required.
- We use unit dozen to denote 12 items; score for 20 items and so one. To count the entities at microscopic level mole concept was introduced.
- In SI system, mole was introduced as seventh base quantity for the amount of a substance.
- One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C isotope.
- The mass of a carbon–12 atom was determined by a mass spectrometer and found to be equal to (1.992648×10^{-23}) g.
- We know that one mole of carbon weighs 12g,the number of atoms in it is equal to:-
- (12 g / mol ^{12}C)/(1.992648×10^{-23} g/^{12}C atom)
- = 6.0221367×10^{23} atoms/mol.
- This number of entities in 1 mol is so important that it is given a separate name and symbol.
- It is known as ‘Avogadro constant’, denoted by N_A in honour of Amedeo Avogadro.
- If a number is written without using the powers of ten 602213670000000000000000, so many entities (atoms, molecules or any other particle) constitute one mole of a particular substance.
- 1 mol of hydrogen atoms = 6.022×10^{23}
- 1 mol of water molecule = 6.022×10^{23}
- 1 mol of sodium chloride = 6.022×10^{23} units of NaCl.

1 mole of various substances

Molar mass
- The mass of one mole of a substance in grams is called its molar mass.
- The molar mass in grams is numerically equal to atomic /molecular/formula mass in u.
- Molar mass of oxygen (O) =16.02g.
- Molar mass of water (H_2O) = 18.02g.
- Molar mass of carbon (C) = 12g.
- Molar mass of sodium chloride (NaCl) = 58.5g.

Percentage Composition
- The percentage composition of a given compound is defined as the ratio of the amount of each element to the total amount of individual elements present in the compound multiplied by 100.
  - For example:-Consider H_2O molecule. Molar mass of Hydrogen (H) = 2g, and Molar mass of Oxygen (O) = 16g.
    - Consider 18g of H_2O it contains 2g of H and 16g of O.
    - Mass % of H = (2×1.008×100)/(18.02) = 11.18%
    - Mass % of O = (16.00×100)/(18.02)= 88.79%
  - <u>**Note**</u>: - By using information from percentage composition we can calculate empirical formula.

Empirical Formula for Molecular formula
- An empirical formula represents the simplest whole number ratio of various atoms present in a compound.
- Molecular formula shows the exact number of different types of atoms present in a molecule of a compound.
- If the mass percent of various elements present in a compound is known, then its empirical formula can be determined.
- Molecular formula can further be obtained if the molar mass is known.

<u>**Problem:-**</u>

A compound contains 4.07 % hydrogen, 24.27 % carbon and 71.65 % chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas?

<u>**Answer**</u>:-

Step 1:- Conversion of mass per cent to grams.

Since we are having mass per cent, it is convenient to use 100 g of the compound as the starting material. Thus, in the

100 g sample of the above compound,

4.07g hydrogen is present, 24.27g carbon is present and 71.65 g chlorine is present.

Step 2:- Convert into number moles of each element

Divide the masses obtained above by respective atomic masses of various elements.

Moles of hydrogen = (4.07 g/1.008 g) = 4.04

Moles of carbon = (24.27 g/12.01g) = 2.021
Moles of chlorine = (71.65 g/35.453 g) = 2.021
Step 3:- Divide the mole value obtained above by the smallest number
Since 2.021 is smallest value, division by it gives a ratio of 2:1:1 for H: C: Cl.

In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.
Step 4:- Write empirical formula by mentioning the numbers after writing the symbols of respective elements.
CH_2Cl is, thus, the empirical formula of the above compound.
Step 5:- Writing molecular formula
(a) Determine empirical formula mass
Add the atomic masses of various atoms present in the empirical formula.
For CH_2Cl, empirical formula mass is
(12.01) + (2 × 1.008) + (35.453)
= 49.48 g
(b) Divide Molar mass by empirical formula mass
n = (Molar mass/Empirical formula)
n=2
(c) Multiply empirical formula by n obtained above to get the molecular formula
Empirical formula = CH_2Cl, n = 2. Hence molecular formula is $C_2H_4Cl_2$.
Problem:-
Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).
Answer:-
The molecular formula of sodium sulphate is (Na_2SO_4)
Molar mass of (Na_2SO_4) = [(2 × 23.0) + (32.066) + 4 (16.00)]
= 142.066 g
Mass percent of an element:-
(Mass of that element in the compound/ Molar mass of the compound) x 100
Therefore, Mass percent of sodium:
= (46.0g/142.066g)/100
=32.379
=32.4%
Mass percent of sulphur:
= (32.066g)/ (142.066g) x 100
=22.57
=22.6%
Mass percent of oxygen:
= (64.0g)/ (142.066g) x 100
=45.049
=45.05%
Problem:-
Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.
Answer:-
% of iron by mass = 69.9 % [Given]
% of oxygen by mass = 30.1 % [Given] Relative moles of iron in iron oxide:
Relative moles of oxygen in iron oxide:

(% of iron by mass)/ (Atomic mass of iron)
= (69.9)/ (55.85)
Simplest molar ratio of iron to oxygen:
= 1.25: 1.88
= 1: 1.5
2: 3
Therefore, the empirical formula of the iron oxide is Fe_2O_3.

Stoichiometry

- The word 'stoichiometry' is derived from two Greek words - stoicheion (meaning element) and metron (meaning measure).
- Stoichiometry deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.
- This is done using balance chemical equation.

Chemical Reaction

- A chemical reaction takes place whenever a chemical change occurs.
    - For example: - Magnesium (Mg) + Oxygen (O) à Magnesium oxide (MgO).
- Reactants are the substances which undergo chemical change in the reaction.
- Products are the new substances, formed during the reaction.
    - For example: - Magnesium (Mg) + Oxygen (O) à Magnesium oxide (MgO). In this equation Magnesium and Oxygen are reactants and Magnesium oxide is the product formed.

Balance Chemical Reaction

- According to law of conservation of mass; mass can neither be created nor be destroyed in a chemical reaction. That is, the total mass of the elements present in the products of a chemical reaction has to be equal to the total mass of the elements present in the reactants.
- Number of atoms on each element remains the same, before and after a chemical reaction.
    - For example:- (a) $2Mg + O_2$ à $2MgO$
    - (b) $Zn + H_2SO_4$ à $ZnSO_4 + H_2$

Balance Chemical equation in Stoichiometry

- $CH_4(g) + 2O_2(g)$ à $CO_2(g) + 2H_2O(g)$.
- The above reaction gives the information as follows:-
    - One mole of CH_4 (g) reacts with two moles of O_2 (g) to give one mole of CO_2 (g) and two moles of H_2O (g).
    - One molecule of CH_4 (g) reacts with 2 molecules of O_2 (g) to give one molecule of CO_2 (g) and 2 molecules of H_2O (g).
    - 16 g of CH_4 (g) reacts with 2×32 g of O_2 (g) to give 44 g of CO_2 (g) and 2×18 g of H_2O (g).

Problem:-
Calculate the amount of water (g) produced by the combustion of 16 g of methane.
Answer:-
The balanced equation for combustion of methane is:
$CH_4(g) + 2O_2(g)$ à $CO_2(g) + 2H_2O(g)$
(i) 16 g of CH_4 corresponds to one mole.
(ii) From the above equation, 1 mol of CH_4 (g) gives 2 mol of H_2O (g).
2 mol of water (H_2O) = 2 × (2+16) = 2 × 18 = 36 g
1 mol H_2O = 18 g H_2O ⇒ (18g H_2O)/ (1 mol H_2O) = 1
Hence 2 mol H_2O × (18g H_2O)/ (1 mol H_2O)

= 2 × 18 g H_2O = 36 g H_2O

Limiting Reagent

- In a chemical reaction, reactant which is present in the lesser amount gets consumed after sometime and after that no further reaction takes place whatever be the amount of the other reactant present. Hence, the reactant which gets consumed, limits the amount of product formed and is, therefore, called the limiting reagent.

Problem:-

50.0 kg of N_2 (g) and 10.0 kg of H_2 (g) are mixed to produce NH_3 (g).
Calculate the NH_3 (g) formed. Identify the limiting reagent in the production of NH3 in this situation.

Answer:-

A balanced equation for the above reaction is written as follows:
Calculation of moles:
N_2 (g) + $3H_2$ (g) $\rightleftharpoons$ $2NH_3$ g
moles of N_2
= 50.0 kg N_2 × (1000 g N)/ (1 kgN_2) x (1molN_2)/ (28.0g N_2)
= 17.86×10^2 mol
moles of H_2
= 10.00 kg H_2 × (1000g H_2)/ (1kg H_2)/ (1molH_2)/ (2.016g H_2)
= 4.96×10^3 mol
According to the above equation, 1 mol
N_2 (g) requires 3 mol H_2 (g), for the reaction.
Hence, for 17.86×10^2 mol of N_2, the moles of H_2 (g) required would be
17.86×10^2 mol N_2 × (3 molH_2 (g))/ (1mol N_2 (g)
= 5.36×10^3 mol H_2
But we have only 4.96×10^3 mol H_2. Hence, dihydrogen is the limiting reagent in this case. So NH_3 (g) would be formed only from that amount of available dihydrogen i.e., 4.96×10^3 mol
Since 3 mol H_2 (g) gives 2 mol NH_3 (g)
(4.96×10^3 mol H_2 (g)) × (2molNH_3 (g))/ (3molH_2 (g))
= 3.30×10^3 mol NH_3 (g)
=3.30×10^3 mol NH_3 (g) is obtained.
If they are to be converted to grams, it is done as follows:
1 mol NH_3 (g) = 17.0 g NH_3 (g)
3.30×10^3 mol NH_3 (g) × (17.0 g NH_3 (g))/ (1mol NH_3 (g))
= $3.30\times10^3\times17$g NH_3 (g)
= 56.1×10^3g NH_3
= 56.1 kg NH_3

Problem:-

Calculate the amount of carbon dioxide that could be produced when
(i) 1 mole of carbon is burnt in air.
(ii) 1 mole of carbon is burnt in 16 g of dioxygen.
(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Answer:

The balanced reaction of combustion of carbon can be written as:
(i) As per the balanced equation, 1 mole of carbon burns in1 mole of dioxygen (air) to produce 1 mole of carbon dioxide.

(ii) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mole of carbon to give 22 g of carbon dioxide. Hence, it is a limiting reactant.
(iii) According to the question, only 16 g of dioxygen is available. It is a limiting reactant.
Thus, 16 g of dioxygen can combine with only 0.5 mole of carbon to give 22 g of carbon dioxide.

Problem:-

In a reaction: - $A + B_2 \rightarrow AB_2$
Identify the limiting reagent, if any, in the following reaction mixtures.
(i) 300 atoms of A + 200 molecules of B
(ii) 2 mol A + 3 mol B
(iii) 100 atoms of A + 100 molecules of B
(iv) 5 mol A + 2.5 mol B
(v) 2.5 mol A + 5 mol B

Answer:-

A limiting reagent determines the extent of a reaction. It is the reactant which is the first to get consumed during a reaction, thereby causing the reaction to stop and limiting the amount of products formed.
(i) According to the given reaction, 1 atom of A reacts with 1 molecule of B. Thus, 200 molecules of B will react with 200 atoms of A, thereby leaving 100 atoms of A unused. Hence, B is the limiting reagent.
(ii) According to the reaction, 1 mole of A reacts with 1 mole of B. Thus, 2 mole of A will react with only 2 mole of B. As a result, 1 mole of A will not be consumed. Hence, A is the limiting reagent.
(iii) According to the given reaction, 1 atom of A combines with 1 molecule of B. Thus, all 100 atoms of A will combine with all 100 molecules of B. Hence, the mixture is stoichiometric where no limiting reagent is present.
(iv) 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of B will combine with only 2.5 mole of A. As a result, 2.5 mole of A will be left as such. Hence, B is the limiting reagent.
(v) According to the reaction, 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of A will combine with only 2.5 mole of B and the remaining 2.5 mole of B will be left as such. Hence, A is the limiting reagent.

Reactions in Solution

- A majority of reactions in the laboratories are carried out in solutions.
- The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways:-

1. Mass per cent or weight percent (w/w%)
2. Mole fraction
3. Molarity
4. Molality

Mass per cent or weight percent

- Mass percentage is one way of representing the concentration of an element in a compound or a component in a mixture.
- Mass percentage is calculated as the mass of a component divided by the total mass of the mixture, multiplied by 100%.
    - For example: - Consider a compound of H_2O which contains 2 moles of H_2 and 1 mole of O_2.
        - Or it can be also stated as 2 x 1gram of H and 16grams of O.
        - Therefore Mass % of H_2=(2/18) x 100% and Mass% of O_2=(16/18)x100%

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Problem:-

A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute.

Answer:-

Mass of percent of A = (Mass of A)/ (Mass of solution) x100

= (2g)/ (2g of A + 18g of water) x100

= (2g/20g)/100

= 10 %

Mole Fraction

- It is the ratio of number of moles of a particular component to the total number of moles of the solution.
  - For example: - Consider a substance 'A' dissolves in substance 'B' and their moles are n_A and n_B
  - Mole fraction of A
    - = (No. of moles of A)/(No. of moles of solution)
    - $(n_A)/(n_A + n_B)$
  - Mole fraction of B
    - =(No. of moles of B)/(No. of moles of solution)
    - $(n_B)/(n_A + n_B)$

Molarity

- Molarity is defined as the number of moles of the solute in 1 litre of the solution.
- It is widely used unit and is denoted by 'M'.
- Molarity (M) = (No. of moles in solute)/(Volume of solution in litres)

Problem:-

A 4 g sugar cube (sucrose: $C_{12}H_{22}O_{11}$) is dissolved in a 350 ml teacup filled with hot water. What is the molarity of the sugar solution?

Answer:-

Equation of molarity:-

M = (m/V)

Where M is molarity (mol/L)

m = number of moles of solute

V = volume of solvent (Litres)

For each of the atoms to get the total grams per mole:

$C_{12}H_{22}O_{11}$ = (12) (12) + (1) (22) + (16) (11)

$C_{12}H_{22}O_{11}$ = 144 + 22+ 176

$C_{12}H_{22}O_{11}$ = 342 g/mol

To get the number of moles in a specific mass, divide the number of grams per mole into the size of the sample:

(4 g)/ (342 g/mol) = 0.0117 mol

Molality

- It is defined as the number of moles of solute present in 1 kg of solvent.
- It is denoted by m.
  - Thus Molality(m) =
  - (No. of moles of solute)/(Mass of solvent in kg)

Problem:-

The density of 3 M solution of NaCl is 1.25 g mL^{-1}. Calculate molality of the solution.

Answer:-

M = 3 mol L^{-1}

Mass of NaCl in 1 L solution = 3 × 58.5 = 175.5 g

Mass of 1L solution = 1000 × 1.25 = 1250 g

(Since density = 1.25 g mL^{-1})

Mass of water in solution = 1250 –175.5
= 1074.5g
Molality= (No. of moles of solute)/ (Mass of solvent in kg)
= (3 mol)/ (1.0745kg)
= 2.79 m.

Problem:-

How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

Answer:-

1 mole of $CuSO_4$ contains 1 mole of copper.
Molar mass of $CuSO_4$ = (63.5) + (32.00) + 4(16.00)
= 63.5 + 32.00 + 64.00
= 159.5 g
159.5 g of $CuSO_4$ contains 63.5 g of copper.
⇒ 100 g of $CuSO_4$ will contain (63.5 x 100g)/ (159.5) of copper.
Therefore Amount of copper that can be obtained from 100 g $CuSO_4$
= (63.5 x 100)/ (159.5) =39.81 g