CHAPTER ONE

SOME BASIC CONCEPTS OF CHEMISTRY

1. Anything which has mass and occupies space is called matter.

2. Matters exist in three physical states → solid, liquid and gas.

3. In solids, these particles are held very close to each other in an orderly fashion 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.

4. Solids have definite volume and definite shape.

5. Liquids have definite volume but not the definite shape. They take the shape of the container in which they are placed.

6. Gases have neither definite volume nor definite shape. They completely occupy the container in which they are placed.

7. A mixture contains two or more substances present in it which are called its components.

Table 1.1 Base Physical Quantities & their units

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

8. A mixture may be homogeneous or heterogeneous.

9. In a homogeneous mixture the components completely mix with each other and its composition is uniform throughout. Sugar solution and air are thus, the examples of homogeneous mixtures.

10. In heterogeneous mixture the composition is not uniform throughout and sometimes the different components can be observed.
For examples - the mixture of salt and sugar grains and pulses along with some dirt pieces, are heterogeneous mixture.

11. The components of a mixture can be separated by using physical methods such as simple hand picking, filtration, crystallization distillation etc.

12. Pure Substances - have characteristics different from the mixtures. They have fixed composition. Also, the constituents of pure substances cannot be separated by simple physical methods.

13. Element - An element consists of only one type of particles. These particles may be atoms or molecules

14. When two or more atoms of different elements combine, the molecule of a compound is obtained.

15. SI system - The SI system (Systeme International of units - abbreviated as SI) has seven units and

and they are listed in.

16. Mass of a substance - 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.

17. Volume - Volume has the units of $(\text{length})^3$. 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 liquids.

$1 L = 1000 mL$, $1000 m^3 = 1 dm^3$

18. Density - Density of a substance is its amount of mass per unit volume SI units of density kg $kg\,m^{-3}$
This unit is quite large and a chemist often expresses density in $g\,cm^{-3}$

19. There are three common scales to measure temperature. Here, K is the SI unit.

20. The Kelvin Scale is related to celsius scale as follows: $K = °C + 273.15$

21. The °F scale is related to celsius scale as follows
$°F = \frac{9}{5}(°C) + 32$.

22. Scientific notation - In scientific notation any number can be represented in the form $N \times 10^n$ where n is an exponent having positive or negative values & N can vary between 1 to 10.

23. Significant figures - In significant figures are meaningful digits which are known with certainty. The uncertainty is indicated by writing the certain digit and the last uncertain digit.

24. There are certain rules for determining the number of significant figures. These are stated below.

a) All non-zero digits are significant.
b) Zeros preceding to first non-digit are not significant. Such zero indicates the position of decimal point.
c) Zeros between two non-zero digits are significant.
d) Zeros at the end or right of a number are significant provided they are on the right side of the decimal point.
e) In numbers written in scientific notation, all digits are significant e.g. 4.01×10^2 has three significant figures, and 8.256×10^{-3} has four significant figures.

25. Law of Conservation of Mass - Antoine Lavoisier established the law of conservation of mass. It states that matter can neither be created nor destroyed.

26. Laws of Definite proportions - States that a given compound always contains exactly the same proportion of element by weight.

27. Law of Multiple proportions - States that 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.

28. 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.

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

30. Dalton's Atomic theory - In 1808, Dalton published 'A new System of Chemical Philosophy' in which he proposed the following.

a- Matter consists of indivisible atoms.

b- All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.

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

d. chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

Dalton's theory could explain the law of chemical combination.

31. One atomic mass unit - is defined as a mass exactly equal to one twelfth the mass of one carbon - 12 atom.

32. Molecular mass - is the sum of atomic masses of the elements present in a molecule. It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together. Molecular mass - expressed in grams is called gram molecular mass.
Formula mass - Sum of atomic masses of all atoms in a formula unit of the compound.

33. Mole - 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.

34. Empirical formula - An empirical formula represents the simplest whole number ratio of various atoms present in a compound whereas the molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

If the mass per cent of various elements present in a compound is known, its emperical formula can be determined.

Following steps should be followed to determine empirical formula of the compound:

Step-1 Conversion of mass percent of various elements into grams.

Step-2 Convert mass obtained into number of moles.

Step-3 Divide the mole value obtained by the smallest mole value.

Step-4 ~~If~~ If the ratios are not whole numbers, then they may be converted into whole number by multiplying with the suitable coefficient.

Step-5 Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

35. Limiting Reagent - The reactant which gets consumed, limits the amounts of product formed and is, therefore, called the limiting reagent.

36. Mass per cent -

Mass of solute per 100 g of solution = $\frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

It is the amount of solute in grams dissolved per 100g. of solution. e.g. 10% solution of sodium chloride present in 100g of solution.

37. Mole Fraction: It is ratio of number of moles of a particular component to the total number of moles of all the Components.

$$\text{Mole-fraction of solute} = \frac{\text{No. of moles of solute}}{\text{No. of mole of solute} + \text{No. of mole of solvent}}$$

$$x_b = \frac{n_B}{n_A + n_B} = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{W_B}}$$

$$x_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}}$$ in case of dilute solution.

$$\therefore \frac{W_B}{M_B} << \frac{W_A}{M_A}$$

38. Molality (m) – It is defined as number of moles of solute (B) per 1000g or 1 Kg of solve.

$$\text{Molarity}(M) = \frac{\text{No. of moles of Solute}}{\text{Kg. of solvent}} =$$ where W_A is mass of solvent.

39. Molarity (M) – It is expressed as the number of moles of solute per litre of solution.

$$\text{Molarity}(M) = \frac{\text{No. of moles of solute}}{\text{Litres of solution}} = \frac{W_a}{M_a} \times \frac{1000}{\text{Volume of solution in ml}}$$

Where W_B is mass of solvent.