CONCEPT MAP

SOME BASIC CONCEPTS OF CHEMISTRY

Class XI

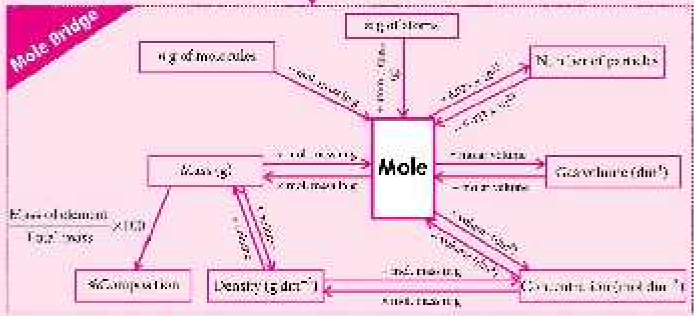
Mole concept is the centre of quantitative calculations in chemistry and the multiple interpretations of this concept allow us to bridge the gap between the submicroscopic world of atoms and molecules and the macroscopic world that we can observe.

Mole Concept

- 1 mole = N_A particles = 6.023×10^{23} particles.
- A mole is defined as the amount of substance that contains the same number of entities (atoms, molecules, ions or other particles), as the number of atoms present in 12 g of the C-12 isotope.
- The number of atoms present in 12 g of C-12 is equal to 6.023×10^{23}.

Have a Look!

- Contrary to belief of chemistry students, 'Avagadros' number was not discovered by Amadeo Avagadro. It is just an honorary name.
- First time number of molecules in any substance was calculated in 1865, by Josef Loschmidt. (2.6×10^{19} molecules in one cm^3 of a gaseous substance).
- Term Avagadros' number was used by Jean Baptiste Perrin in 1909.
- The unit 'mole' was introduced in 1900 by Ostwald and defined this unit in terms of gram.



Empirical and Molecular Formula

Atomic ratio = $\frac{\%age}{\text{At. mass}}$ → Change to → Simplest ratio = $\frac{\text{Atomic ratio}}{\text{minimum atomic ratio}}$ → Change to → Simple whole number ratio = Simplest ratio × integer

Empirical formula $\xrightarrow{\text{Multiply with } n}$ Molecular formula

Stoichiometric Calculations

Limiting Reagent

- The limiting reagent or reactant is the reactant that limits the amount of the other reactant that can combine and the amount of product that can form in a chemical reaction.
- The excess reagent is the substance that is not used up completely in a reaction.
- For example, in combustion of 12 g of carbon in excess of oxygen (i.e., more than 32 g of oxygen), carbon acts as the limiting reagent.

Reactions in Solutions

- Mass % $=\frac{W_B\times100}{W_A+W_B}$
- Molarity (M) $=\frac{W_B\times1000}{M_B\times V\text{ in mL}}$
- Normality (N) $=\frac{W_B\times1000}{GEM_B\times V\text{ in mL}}$
- Molality (m) $=\frac{W_B\times1000}{GMM_B\times W_A\text{ in g}}$
- Mole fraction, $x_2=\frac{n_2}{n_1+n_2}$ and $x_1=\frac{n_1}{n_1+n_2}$

CONCEPT MAP

SOME BASIC CONCEPTS OF CHEMISTRY

Some basic concepts of chemistry provide the base to elucidate the various observations on the physical and chemical properties of matter and the various laws of chemical combinations help in understanding the formation of compounds.

MATTER

Classification

- Matter: Anything that has mass and occupies space.
- Matter exists in three physical states: solid, liquid and gas.
- Based on chemical composition, matter is classified into mixtures and pure substances.
- Mixtures are of two types: homogeneous (uniform composition) and heterogeneous (non-uniform composition).
- Pure substances are elements (one type of particles) and compounds (more than one type of particles).

Measurement

- SI system has seven base units for seven fundamental physical quantities; other units are defined from these units.
- To change the order of magnitude, prefixes are added before the base units.

deci	10^{-1}	deca	10
centi	10^{-2}	hecto	10^{2}
milli	10^{-3}	kilo	10^{3}
micro	10^{-6}	mega	10^{6}
nano	10^{-9}	giga	10^{9}
pico	10^{-12}	tera	10^{12}

- Significant figures:
  - These are all certain digits with last digit uncertain.
  - All non-zero digits are significant.
  - Zeros preceding to first non-zero digit are not significant.
  - Zeros between two non-zero digits are significant.
  - Zeros on the right side of the decimal are significant.
- Some useful conversion factors:
  - Length = 1 Å = 10^{-8} cm = 10^{-10} m, 1 nm = 10^{-9} m, 1 pm = 10^{-12} m
  - Volume = 1 L = 1000 mL = 1000 cm^3 = 1 dm^3 = 10^{-3} m^3
  - Pressure = 1 atm = 760 mm Hg = 101325 Pa, 1 bar = 10^5 N m^{-2} = 10^5 Pa
  - Energy: 1 calorie = 4.184 J, 1 eV = 1.6022×10^{-19} J, 1 J = 10^7 ergs

Laws of chemical combinations

- Law of conservation of mass: Matter can neither be created nor destroyed.
- Law of definite proportions: A given compound always contains exactly the same proportion of elements by weight.
- Law of multiple proportions: 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.
- Gay Lussac's law of gaseous volumes: 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.

Particulate nature of matter

- 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.
- Avogadro's law: Equal volumes of gases at the same temperature and pressure contain equal number of molecules.

IMPORTANT TERMS AND FORMULAE

Mole concept

- Mole: Collection of 6.022×10^{23} (Avogadro number) particles.
- For atomic substances: 1 Mole = Gram atomic mass = 1 Gram atom = 6.022×10^{23} atoms
- For molecular substances: 1 Mole = Gram molecular mass = 1 Gram molecule = 6.022×10^{23} molecules
- For gaseous substances: 1 mole = 22.4 L at STP

Masses

- Atomic mass unit (amu or u): Mass exactly equal to 1/12th of the mass of an atom of C-12 isotope.
- Atomic mass of an element: Average relative mass of its atoms as compared with an atom of C-12 taken as 12.
- Average atomic mass (for isotopes): $\frac{\sum x_iA_i}{\sum x_i}$ where x_i = % abundance, A_i = atomic mass
- Gram atomic mass: Atomic mass of an element expressed in grams.
- Molecular mass: Sum of atomic masses of all the elements present in a molecule.

Percentage composition, empirical and molecular formulae

- Percentage composition: Mass % of an element $=\frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}}\times100$
- Empirical formula: It is the formula showing the simplest whole number ratio of the atoms of various elements present in one molecule of the compound.
- Molecular formula: It is the formula showing exact number of atoms of various elements present in one molecule of the compound.

Molecular formula = n × Empirical formula where, n is a simple whole number.

Limiting reagent

The reactant which gets consumed completely in a reaction and limits the amount of product formed is called limiting reagent.

Methods of expressing concentration

- Mass percent (%) $=\frac{W_{solute}}{W_{solution}}\times100$
- Mole fraction: $x_A=\frac{n_A}{n_A+n_B}$, $x_B=\frac{n_B}{n_A+n_B}$
- Molarity (M) $=\frac{n_B\times1000}{V_{solution}(\text{in mL})}$
- Molality (m) $=\frac{n_B\times1000}{W_{solvent}(\text{in g})}$
- Normality (N) $=\frac{W_B\times1000}{\text{Eq. mass of solute}\times V(\text{in mL})}$