CONCEPT MAP

SOME BASIC CONCEPTS OF CHEMISTRY

Introduction of most fundamental and important tools of chemistry which help in various calculations.

Measurement and uncertainty in measurement

SI Units

- SI system has seven base units pertaining to seven fundamental scientific quantities:

Physical quantity	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)

- SI system allows the use of prefixes to indicate the multiples or submultiples of a unit:

deci - 10^{-1}	deka - 10
cent - 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}

Scientific Notation

- A number is represented as $x \times 10^n$
- n is −ve if decimal is moved towards right and n is +ve if it is moved towards left.

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.

Dimensional Analysis

- Required unit = Given value × conversion factor
- Some useful conversion factors:

Length – $1\,\text{Å} = 10^{-8}\,\text{cm} = 10^{-10}\,\text{m}$
$1\,\text{nm} = 10^{-9}\,\text{m}$, $1\,\mu\text{m} = 10^{-17}\,\text{m}$

Volume – $1\,\text{L} = 1000\,\text{mL}$
$= 1000\,\text{cm}^3 = 1\,\text{dm}^3 = 10^{-3}\,\text{m}^3$

Pressure – $1\,\text{atm} = 760\,\text{mm of Hg} = 760\,\text{torr}$
$= 101325\,\text{Pa}$
$1\,\text{bar} = 10^5\,\text{N m}^{-2} = 10^5\,\text{Pa}$

Energy – $1\,\text{calorie} = 4.184\,\text{J}$
$1\,\text{eV} = 1.6022 \times 10^{-19}\,\text{J}$
$1\,\text{J} = 10^7\,\text{ergs}$

Laws of chemical combinations and Dalton's atomic theory

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.

Avogadro's Law

Equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

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.

Stoichiometry

- It deals with calculations based upon chemical equations.
- Various steps involved in calculations are:
  - Write balanced chemical equation.
  - Write the relative number of moles or relative masses of reactants and products involved below their formulae.
  - In case of gases write 22.4 L at STP in place of 1 mole.
  - Apply unitary method to make required calculations.

Limiting Reagent

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

Mole concept, Masses and Stoichiometry

Masses

- **Atomic mass unit (amu or u) :** Mass exactly equal to $1/12^{th}$ of the mass of an atom of C-12 isotope.
- **Atomic mass of an element :** Average relative mass of its atoms as compared to an atom of C-12.
- **Average atomic mass :** Given for isotopes. $= \frac{\sum X_i A_i}{\sum X_i}$ 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.

Mole Concept

Mole : Collection of 6.022×10^{23} particles
In case of
atomic substances:
1 mole = Gram atomic mass = 1 gram atom
$= 6.022 \times 10^{23}$ atoms
molecular substances:
1 mole = Gram molecular mass
= 1 gram molecule
$= 6.022 \times 10^{23}$ molecules
gaseous substances:
1 mole = 22.4 L at STP

Percentage Composition

It shows mass of a constituent in 100 parts of a compound.

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

Empirical Formula

- It is the simplest whole number ratio of different atoms present in a compound.
- Steps to obtain empirical formula:

Molar ratio $= \frac{\text{\% age}}{\text{At. mass}}$ → Change to

Simplest molar ratio $= \frac{\text{Molar ratio}}{\text{Minimum molar ratio}}$ → Change to

Simplest whole number ratio = Simplest ratio × Integer → Write the number value before the symbols of elements → Empirical formula

Molecular Formula

- It is the formula showing exact number of atoms present in a molecule.
- Molecular formula = n × empirical formula

Reactions in Solutions

- Mass percent (%) $= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$
- Mole fraction (x_A) $= \frac{n_A}{n_A + n_B}$, $x_B = \frac{n_B}{n_A + n_B}$
- Molarity (M) $= \frac{w_2 \times 1000}{M_2 \times V\,(\text{in mL})}$
- Molality (m) $= \frac{w_2 \times 1000}{M_2 \times w_1\,(\text{in g})}$

HAVE A LOOK!

- Mass is the quantity of matter contained in the substance and is constant whereas weight varies from place to place.
- Exact numbers have an infinite number of significant figures.
- Molar volume of a gas is 22.7 L at 1 bar and 0°C.
- The number of molecules in 1 mL of a gas at STP is known as **Loschmidt number.**