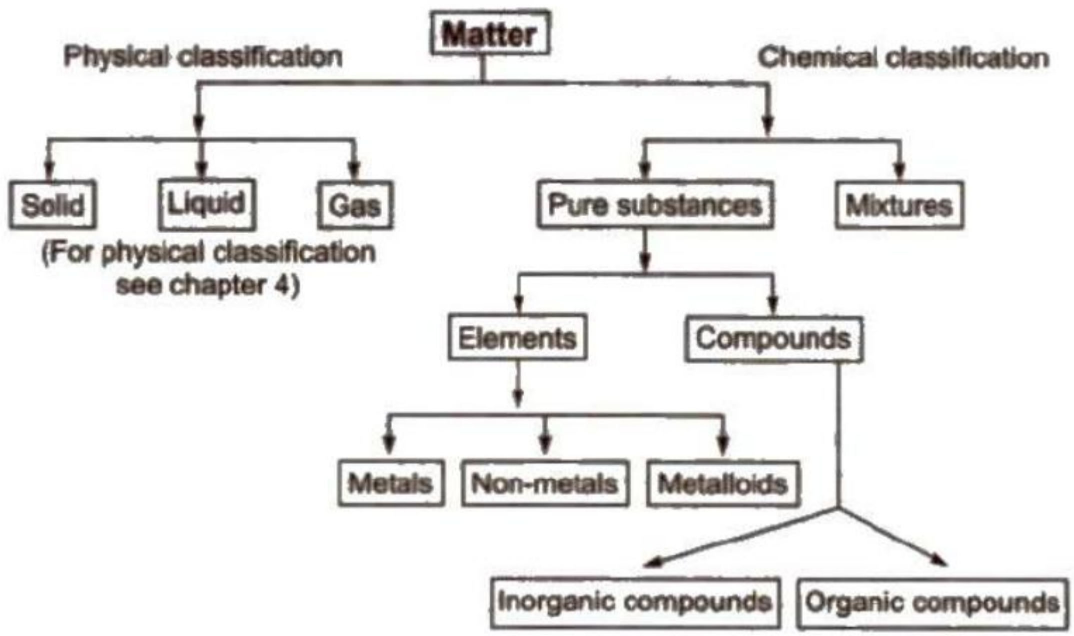
Class 11, Chapter- 1: Some Basic Concept of Chemistry

1. **Matter:** Anything which occupies some space and has some mass is called matter.
2. **Classification of Matter:**



3. Measurement of Matter: The systems used for describing measurements of various physical quantities are

(a) CGS system

(b) FPS system

(c) MKS system

1. International System of Units (SI system- 1960)

Quantity	Unit	Symbol
Length	metre	m
Mass	kilogram	kg
Time	second	S
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	Cd

2. Prefix in SI system

Multiple	Prefix	Symbol
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

4. Volume-- Volume has the units of (length) ^{3}. So volume has units of m3 or cm3 or dm3.

1 L = 1000 mL, 1000 cm3 = 1 dm3

5. Density: mass / volume; SI unit of density = kg/m3 or gm cm–3.
6. Temperature--There are three common scales to measure temperature — °C (degree celsius), °F (degree Fahrenheit) and K (kelvin). Here, K is the SI unit

*F = 9/5(*C) + 32
K = °C + 273.15

Note—Temperature below 0 °C (i.e. negative values) is possible in Celsius scale but in Kelvin scale, negative temperature is not possible.

7. Scientific Notation: In which any number can be represented in the form N × 10n (Where n is an exponent having positive or negative values and N can vary between 1 to 10) like

$$232.508 = 2.32508 \times 10^{2} \quad ; \quad 0.00016 = 1.6 \times 10^{-4}.$$

8. Precision and Accuracy: Precision is simply a measure of reproducibility of an experiment while Accuracy is a measure of the difference between the experimental value or the mean value of a set of measurements and the true value.
In physical measurements, accurate results are generally precise but precise results need not be accurate. In other words good precision does not assure good accuracy. For example

	Measurement 1 (ppm)	**Measurement 2 (ppm)**	**Avg. (ppm)**	**Observation**
Student A	1.95	1.93	1.940	Precise but not accurate
Student B	1.94	2.05	1.995	Neither precise nor accurate
Student C	2.01	1.99	2.000	Precise and accurate measurement

9. Significant Figures and their Rule: Significant figures are meaningful digits which are known with certainty. The uncertainty is indicated by writing the certain digits and the last uncertain digit. Thus, if we write a result as 11.2 ml, we say the 11 is certain and 2 are uncertain. There are certain rules for determining the number of significant-

Sr. No.	**Rule for Significant Figure**	**Example**	**Significant No.**
1	All non zero digit and any zeros b/w non zero digit are significant (count)	285	3
		0.25	2
		300042	6
2	Leading zeros are not significant (don't count).	0.05	1
		0.0062	2
3	Zeros between two non-zero digits are significant	4.006	4

	(count).		
		1.01	3
4	Zeros on the right side of the decimal point are significant.	0.200	3
5	Terminal zeros are not significant if there is no decimal point.	100	1
		1000	1
		290000	2
		100.	3
		100.0	4
6	Counting / Exact numbers of objects, for example, 2 balls or 20 eggs, have infinite significant figures because these can be represented by writing infinite number of zeros after placing a decimal i.e., 2 = 2.000000 or 20 = 20.000000.	2	∞
		20	∞
7	In numbers written in scientific notation, all digits are significant	4.01×10^{2}	3
		6.254×10^{-3}	4

Addition, subtraction, multiplication & division of S. Figures: For that the final result should be reported to the same number of decimal places as that of the term with the least number of decimal places, e.g.,

$$\begin{array}{l} 2.512 \\ 2.2 \\ \underline{5.23} \\ 9.942 = 9.9 \end{array} \quad ; \quad 1.6 \times 1.32 = 2.112 = 2.1$$

Rounding Off the Numerical Results: When a number is rounded off, the number of significant figures is reduced, in this case the last digit is increased by 1 only if the following digit is ≥ 5 and is left as such if the following digit is ≤ 4 like

12.696 = 12.7; 18.35 = 18.4; 13.93 = 13.9

10. Law of Chemical Combinations: The combination of elements to form compounds is governed by the following five basic laws-

1) Law of Conservation of Mass: Its states that matter can neither be created nor destroyed.

2) Law of Definite Proportions (Proust, 1799): According to this law, a chemical compound obtained by different sources always contains same percentage of each constituent element.

According to Proust composition of elements was found same in both sample of cupric carbonate, taken from natural origin and other than synthetic one like

Sample	% of Copper	% of Oxygen	% of Carbon
Natural Sample	51.35	9.74	38.91
Synthetic Sample	51.35	9.74	38.91

3) Law of Multiple Proportions (Dalton, 1803): 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 H_2O and H_2O_2, fixed mass of hydrogen requires oxygen in the ratio of 1 : 2.

$$H_2 + \frac{1}{2} O_2 \rightarrow H_2O$$

$$H_2 + O_2 \rightarrow H_2O_2$$

4) Gay Lussac Law of Gaseous Volume (1808): 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.

$$2H_2 + O_2 \rightarrow 2H_2O$$

2 Volumes   1 Volume   2 Volumes

5) Avogadro Law (1811): It states that equal volumes of all gases under the same conditions of temperature and pressure contain the same number of molecules.

11. Dalton Atomic Theory: This theory was based on laws of chemical combinations. 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 or molecules result from combination of atoms in some simple numerical ratio.

12. Atomic Mass Unit (amu): One atomic mass unit is defined as a mass exactly equal to one twelfth the mass of one C^{12} atom.

$$\textbf{One amu} = \mathbf{1.66056 \times 10^{-24}} \textbf{ gm}$$

Today, 'amu' has been replaced by 'u' which is known as unified mass.

Atomic Mass: Atomic mass of an element is defined as the **average relative atomic mass** of an atom of an element as compared to the mass of an atom of C^{12} taken as 12.

$$\text{Atomic mass} = \frac{\text{mass of an atom}}{1/12 \text{ mass of a carbon atom} (^{12}C)}$$

Mass of one atom of Hydrogen (gm) = 1.6736×10^{-24} gm

Mass of one atom of Hydrogen (amu) = $1.6736 \times 10^{-24} / 1.66056 \times 10^{-24}$ = 1.0078 amu

Gram Atomic Mass: Gram atomic mass is the mass, in gram, of one mole of atoms in a monatomic element. It is numerically equal to the relative atomic mass (atomic weight) in grams or in other words atomic mass of an element expressed in grams is the gram atomic mass.

For example, the atomic mass of oxygen = 16 amu

Therefore gram atomic mass of oxygen = 16 g

Average Atomic Mass: Actually many elements, exist as more than one isotope. So average atomic mass for different isotope of an element like carbon can be calculated by following formula-

$$= (\%RA \text{ of } C^{12} / 100) \times \text{A. mass of } C^{12} \text{ Isotope} + (\%RA \text{ of } C^{13} / 100) \times \text{A. mass of } C^{13}$$

Where RA = Relative Abundance (% occurrence)

Refer related numerical from NCERT P-22, Exercise- 1.9

13. Molecular mass: It is the sum of the masses of all atoms in a molecule for example- Molecular mass of NH_3 = (14 u) + 3 (1.0 u) = 17 u

14. Molar Mass: The mass of one mole of a substance in grams is called its molar mass. For example- Molar mass of NH_3 = (14) + 3 (1.0) = 17 gm / mol

15. Formula mass: Formula mass is the sum of the masses of all atoms in a formula unit of an ionic compound for example- Formula mass of NaCl = 23.0 u + 35.5 u = 58.5 u

16. Mole Concept. Mole is defined as the amount of a substance, which contains the same number of chemical units (atoms, molecules, ions or electrons) as there are atoms in exactly 12 grams of pure carbon-12.

A mole represents a collection of 6.022×10^{23}(Avogadro's number) chemical units.

1 mole of atoms = 6.022×10^{23} atoms (Avogadro No.)

1 mole of molecule = 6.022×10^{23} molecule

$$\text{Mole} = \text{mass (gm) of substance} / \text{Molar mass of substance}$$

$$\textbf{10. No. of Atoms in Element} = \frac{\text{mass of element (gm)} \times N_A}{\text{Molar mass (gm/mole)}}$$

$$\textbf{No. of Atoms in Element} = \text{mass of element (u)} \times N_A$$

$$\textbf{11. \% Mass of element in the compound} = \frac{\textbf{Mass of that element in the compound x 100}}{\textbf{Molar mass of the compound}}$$

20. Empirical and Molecular Formula: Empirical formula is the simplest formula of a compound giving simplest whole number ratio of atoms present in one molecule, e.g., CH is empirical formula of benzene (C6H6).

Molecular formula is the actual formula of a compound showing the total number of atoms of constituent elements, e.g., C6H6 is molecular formula of benzene.

Empirical Formula x n = Molecular Formula;

n = molecular formula mass / empirical formula mass

Where, n is simple whole number having values 1, 2, 3,… etc., and can be calculated as

21. Stoichiometry and their Calculations: The relative proportions in which the reactants react and the products are formed, is called stoichiometry (from the Greek word meaning 'to measure an element'.)

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2 + 2H_2O\ (g)$$

Thus, according to the above chemical reaction,

- 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 $H_2O(g)$
- 22.7 L of $CH_4(g)$ reacts with 45.4 L of O_2 (g) to give 22.7 L of CO_2 (g) and 45.4 L 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).

From these relationships, the given data can be interconverted as follows :

mass $\rightleftharpoons$ moles $\rightleftharpoons$ no. of molecules

Limiting reagent: It is the reactant which is completely consumed during the reaction.

Excess reagent: It is the reactant which is not completely consumed and remains unreacted during the reaction.

Calculate the mass of reagent or product =

$$\frac{\text{Mass of Reagent- A} \quad \times \quad \text{Molar Mass of Reagent-B or Product}}{\text{Molar mass of Reagent –A}}$$

How many grams of HCl react with 5.0 g MnO_2 according to the equation :

$$4HCl\ (aq) + MnO_2\ (s) \rightarrow 2H_2O\ (l) + MnCl_2\ (aq) + Cl_2\ (g)$$

Soln: 146 gm/mole 86.94 gm/mole

= 5/86.94 x 146 = 8.40 gms

Refer related numerical from NCERT P-24, Exercise- 1.4, 1.23, 1.24, 1.26

22. Concentration of solution:

$M_1V_1 = M_2V_2$

3M NaCl soln means Molarity (Conc^n) of NaCl = 3 mole/ litre = 3 x 58.5 gm (mass of NaCl) in 1 liter solution

1 mole HCl = 1 x 36.5 = 36.5 gm of HCl

5 mole NH_3 = 5 x 17 = 85 gm of NH_3 NH_3

$$\text{\% Mass of Solute} = \frac{\text{mass of solute} * 100}{\text{Mass of solute (gm) + Mass of Solvent (gm)}}$$

$$\text{Mole fraction } (x_B) = \frac{\text{No. of moles of solute } (n_B)}{\text{No. of mole of solute } (n_B) + \text{No. of moles of Solvent } (n_A)}$$

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

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

Where: W_B = Mass of solute; M_B = Molar mass of solute
W_A = Mass of solvent; M_A = Molar mass of solvent

$$\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (liter)}}$$

$$\text{No. of moles of solute} = \frac{\text{mass of solute (gm)}}{\text{Molar mass (m.wt.) of solute (gm/mole)}}$$

$$\text{Molarity} = \frac{\text{mass of solute (gm) x 1000}}{\text{Molar mass (m.wt.) of solute (gm /mole) x Vol. of soln (ml)}}$$

Some important formulas: Chapter- 1: Some Basic Concept of Chemistry

1. International System of Units (SI system- 1960)

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Temperature	kelvin	K
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Electric current	ampere	A
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Multiple	Prefix	Symbol
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10^{-6}	micro	µ
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10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
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14. Molarity

15. Molality:

16. Normality:

17. Mole Fraction:

18. %Mass of solute:

19. Dilution law: $M_1V_1 = M_2V_2$

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