Matter and Laws of Chemical Combinations

General Introduction: Importance and Scope of Chemistry

Chemistry is the branch of science which deals with the study of composition, properties and interaction of matter. In other way, chemistry is also considered as science of atoms and molecules. Chemistry has helped us to meet with our requirements that are necessary for the amelioration of life such as food additives, health care products, plastic, dyes, polymers etc., absence of which would make our life very dull and dreary. Thus, chemistry is arguably the most important science.

Matter

Matter is anything which possesses mass and occupies space. Matter can generally exist in three physical states viz., solid, liquid and gas.

Classification of Matter

Matter can be classified by two ways on the basis of :

(i) Physical state On the basis of physical state, matter may be classified into solid, liquid and gas under the conditions of ordinary temperatures and pressures. Solids have a definite volume and shape, liquids have definite volume but do not have definite shape, gases have neither a definite volume nor a definite shape.

Note

There are two more states—Plasma and Bose-Einstein condensate.

Plasma is a state containing low density ionised gases at very high temperatures. Plasma can be seen in (Aurora borealis) flames, neon lights, clouds of gases and dust around stars.

Bose-Einstein condensate is totally opposite to plasma. The state is conceptualised at super cooled conditions. The super cooled means only a few billionth of a degree above absolute zero. Cornell and Weimann developed BEC at such temperature with rubidium.
(ii) Chemical classification Matter can also be classified on the basis of chemical composition and properties. On this basis, matter can be classified as

Matter

Pure substances
Elements
Compounds
Mixtures
Homogeneous
Heterogeneous

Pure Substances
These have fixed composition and non-variable properties. These cannot be separated into simpler substances by physical methods.

These are further classified into two following classes:

(a) Elements Pure substances which consist of only one type of atoms. e.g. Na, Mg, Al etc.
(b) Compounds Pure substances that are composed of two or more types of atoms, i.e. here different elements present in a fixed proportion by mass. Properties of a compound are different from the properties of the constituent elements. e.g. H₂O, NH₃, NaCl, CaCO₃ etc.

Mixtures
These are the combinations of two or more substances (elements or compounds) which are not chemically combined together and retain their characteristic properties. These may be separated into pure components by applying physical methods.

Mixtures are of two types:

(a) Homogeneous mixture These mixtures have uniform composition throughout and their components are indistinguishable. e.g. a liquid solution of sugar and water etc.
(b) Heterogeneous mixture These mixtures do not have uniform composition throughout and components are distinguishable. e.g. a mixture of sand and salt etc.

Properties of Matter
Every substance (matter) has its characteristic properties. These properties are of the following two types:

(i) Physical properties These properties can be measured without change in composition of substance. e.g. mass, length, temperature, density etc.
(ii) Chemical properties Measurement of these properties require chemical change in substance. e.g. acidity, basicity, combustibility etc.

Measurement of Physical Quantities
Numerical values of physical quantities are expressed in terms of units, which are of the following two types:

(i) Basic or fundamental units Length, mass, time, electric current, thermodynamic temperature, luminous intensity and amount of substance are seven basic fundamental quantities and their units are metre, kilogram, second, ampere, kelvin, candela and mole respectively (in SI system).

(ii) Derived units These units are basically derived from fundamental units. e.g. unit of density is derived from unit mass and volume.

Systems of Measurement
Various systems used for measurement are as follows:

(a) CGS system In this system, centimetre (cm), gram (g) and second (s) are the units of length, mass and time respectively.
(b) MKS system In this system, units of length, mass, time and electric current are metre (m), kilogram (kg), second (s) and ampere (A) respectively.
(c) SI system This is the international system of units and contains seven basic and two supplementary units.

Basic units Basic units include-metre (m) for length, kilogram (kg) for mass, second (s) for time, ampere (A) for electric current, kelvin (K) for thermodynamic temperature, candela (cd) for luminous intensity and mole (mol) for amount of substance.

Supplementary units These are radian (rad) for plane angle and steradian (sr) for solid angle.

Besides the above units some SI prefixes are very useful in measurements, these prefixes are,

<table>
<thead>
<tr>
<th>Some useful SI prefixes</th>
</tr>
</thead>
<tbody>
<tr>
<td>Multiple</td>
</tr>
<tr>
<td>10⁻²⁴</td>
</tr>
<tr>
<td>10⁻²¹</td>
</tr>
<tr>
<td>10⁻¹⁸</td>
</tr>
<tr>
<td>10⁻¹⁵</td>
</tr>
<tr>
<td>10⁻¹²</td>
</tr>
<tr>
<td>10⁻⁹</td>
</tr>
<tr>
<td>10⁻⁶</td>
</tr>
<tr>
<td>10⁻³</td>
</tr>
<tr>
<td>10¹</td>
</tr>
</tbody>
</table>
Some physical quantities and their derived units

<table>
<thead>
<tr>
<th>Physical quantity</th>
<th>Description</th>
<th>Unit</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Area</td>
<td>Length square</td>
<td>square metre</td>
<td>m²</td>
</tr>
<tr>
<td>Volume</td>
<td>Length cube</td>
<td>cubic metre</td>
<td>m³</td>
</tr>
<tr>
<td>Density</td>
<td>Mass/unit volume</td>
<td>kilogram per cubic metre</td>
<td>kg m⁻³</td>
</tr>
<tr>
<td>Velocity</td>
<td>Distance/unit time</td>
<td>metre per second</td>
<td>ms⁻¹</td>
</tr>
<tr>
<td>Force</td>
<td>Mass x acceleration</td>
<td>Newton</td>
<td>N or kgm²s⁻²</td>
</tr>
<tr>
<td>Work, energy</td>
<td>Force x distance</td>
<td>Joule (Newton metre)</td>
<td>kg m²/s² or kgm²s⁻²</td>
</tr>
<tr>
<td>Acceleration</td>
<td>Speed change/unit time</td>
<td>metre per second square</td>
<td>ms⁻²</td>
</tr>
<tr>
<td>Frequency</td>
<td>Cycles/second</td>
<td>hertz</td>
<td>Hz or s⁻¹</td>
</tr>
<tr>
<td>Electric charge</td>
<td>Current x time</td>
<td>coulomb</td>
<td>C or A·s⁻¹</td>
</tr>
<tr>
<td>Potential difference</td>
<td>—</td>
<td>volt</td>
<td>V or kg m²s⁻³A⁻¹</td>
</tr>
<tr>
<td>Pressure</td>
<td>Force/unit area</td>
<td>pascal</td>
<td>Pa or Nm⁻²</td>
</tr>
<tr>
<td>Electrical resistance</td>
<td>Potential-difference /current</td>
<td>ohm</td>
<td>Ω or VA⁻¹</td>
</tr>
<tr>
<td>Electric conductance</td>
<td>Reciprocal of resistance or mho</td>
<td>Ω⁻¹ or A/V⁻¹</td>
<td></td>
</tr>
</tbody>
</table>

### Some Important Unit Conversions

(a) **Mass**
- 1 Ton = 1000 kg
- 1 Quintal = 100 kg
- 1 kg = 2.205 pound
- 1 kg = 1000 g
- 1 gram = 1000 milligram
- 1 amu = 1.6 × 10⁻²⁵ g

(b) **Length**
- 1 mile = 1760 yard, 1 yard = 3 feet
- 1 foot = 12 inch, 1 inch = 2.54 cm
- 1Å = 10⁻¹⁰ m or 10⁻⁸ cm

(c) **Volume**
- 1 L = 1 dm³ = 10⁻³ m³
- = 10⁶ cm³ = 10³ mL = 10³cc
- 1 mL = 1 cm³ = 10⁻⁶ m³ = 1 cc

(d) **Pressure**
- 1 atmosphere (atm) = 760 torr
- = 760 mm of Hg = 76 cm of Hg
- = 1.103 × 10⁵ pascal (Pa)
- = 1.103 × 10⁵ N/m²

(e) **Energy**
- 1 calorie = 4.184 J = 4.2 J
- 1 J = 10⁷ erg
- 1 L atm (L-atm) = 101.3 J
- 1 electron volt (eV) = 1.602 × 10⁻¹⁹ J

### Precision and Accuracy

Precision refers to the closeness of the set of values obtained from identical measurement of quantity. Accuracy refers to the closeness of a single measurement to its true value. Good accuracy means good precision but reverse is not always true.

### Significant Figures

These are the meaningful digits in a measured or calculated quantity. Total number of digits in a number including the last digit whose value is uncertain is called the number of significant figures.

#### Rules for Determining the Number of Significant Figures

(i) All non-zero digits are significant, e.g. 265 has three significant figures and 0.265 has also three significant figures.

(ii) Zeroes to the left of the first non-zero digit in the number are not significant. e.g. 0.006 has only one significant figure and 0.036 has only two significant figures.

(iii) Zeroes between non-zero digits are significant. e.g. 0.265 has also three significant figures.

(iv) Zeroes to the right of the decimal point are significant. e.g. 6.00, 0.060 and 0.6000 have three, two and four significant figures.

(v) If a number ends in zeroes that are not to the right of a decimal, the zeroes may or may not be significant. e.g. 1200 g may have two, three or four significant figures. Mass of 1200 g may be expressed in scientific notation in the following form

\[ 1.200 \times 10^3 \text{ g (four significant figures)} \]
\[ 1.20 \times 10^3 \text{ g (three significant figures)} \]
\[ 1.2 \times 10^3 \text{ g (two significant figures)} \]

All the zeroes to the right of the decimal point are significant.

(vi) The result of an addition or subtraction should be started to the same number of decimal places as that of the term with least number of decimal places.

\[ 3.523 + 2.2 + 7.34 = 13.063 \]
\[ \text{Reported sum = 13.1} \]

(vii) Result of multiplication and division should be reported to the same number of significant figure as is possessed by least precise term used in the calculations.

\[ 5.452 \times 3.6 = 19.6272 \]
\[ \text{Least number of significant figures = 2} \]
\[ \therefore \text{Reported product = 19} \]
Chemistry

Historical Approach to Matter

Matter is composed of small indivisible particles called ‘a-tomio’ (meaning indivisible) given by a Greek philosopher, Democritus in 460-370 BC.

The first law about matter is given by Lavoisier (law of conservation of mass) in 1785. Afterward, there was further development in 1799 (Law of constant composition-Joseph Proust); in 1803 (law of multiple proportion-John Dalton).

However, to describe the structure of matter which could explain the experimental facts about elements, compounds and mixtures and also the law of chemical combinations, John Dalton, in 1808 published ‘A New system of chemical Philosophy’ in which he proposed the concepts of matter.

Laws of Chemical Combinations

1. Law of conservation of mass (Lavoisier) Matter can neither be created nor destroyed in a chemical reaction. Thus, total mass of reactants = total mass of products.

2. Law of constant composition or definite proportions (J. Proust) A sample of pure compound always consists of the same elements combined in same proportions by mass whatever be it’s source.

3. Law of multiple proportions (John Dalton) If an element forms more than one compound with another element for a given mass of an element, masses of other elements are in the ratio of small integers. e.g. in NH₃ and N₂H₄, the ratio of masses of hydrogen for the same mass of nitrogen (i.e. 14g) is 3 : 2.

4. Law of reciprocal proportions (Ritcher) When two elements combine separately with a fixed mass of a third element, then the ratio of their masses in which they do so is either same or some whole number multiple of the ratio in which they combine with each other.

5. Law of combining volumes or Law of definite proportions by volume (Gay-Lussac) The volume of gaseous reactants and products in a large number of chemical reactions bear a simple ratio provided volumes are measured at same temperature and pressure.

6. Avogadro’s law At constant temperature and pressure, for a given mass of an ideal gas, the volume of the gas is proportional to the number of moles of the gas.

According to this law, \( V \propto n \)

where, \( n \) = number of moles of a gas.

Dalton’s Atomic Theory

On the basis of laws of chemical combinations, Dalton (1803) proposed atomic theory. The main postulates of this theory are

1. Matter is made up of indivisible and indestructible particles called atoms.

2. Atoms are neither be created nor destroyed in the course of an ordinary chemical reaction.

3. Atoms combine with each other to form compounds in simple whole number ratio.

4. Atom is the smallest portion of matter which takes part in chemical reaction.

5. All atoms of an element have identical mass and similar chemical properties.

Atom

An atom may be defined as “the smallest particle of the matter which cannot exist independently but participates in chemical reactions”. Atom is further made up of subatomic particles electrons, protons and neutrons.
Molecule
A molecule is the smallest particle of matter which can exist independently but can take part in a chemical reaction.
Atoms combine to form molecules
(i) Molecules may be monotomic (contain only one atom e.g. Na, Fe), diatomic (contain 2 atoms, e.g. H₂, O₂), triatomic (contain 3 atoms, e.g. O₃), and polyatomic (contain more than three atoms, e.g. S₈, P₄ etc.).
(ii) Molecules may be of the following two types:
   (a) Homatomic molecules These molecules contain atoms of one element. e.g. H₂, O₂, N₂, Cl₂, S₈ etc.
   (b) Heteroatomic molecules These molecules contain atoms of different elements. These are the molecules of the compounds. e.g. CO, HCl, H₂SO₄, HNO₃, H₂O₂ etc.

Different Types of Mass
Atomic Mass
On the present atomic mass scale, the mass of 12 atomic mass unit (amu) is chosen as standard and is arbitrarily assigned the mass of 12 atomic mass unit (amu). It is defined as the number which indicates how many times the mass of one atom of the element is heavier as compared to 1/12th part of the mass of ¹²C.
Thus, one amu or u (unified mass) is equal to exactly ¹/₁₂ th of the mass of ¹²C atom.
\[ 1u = \frac{1}{12} \times \frac{12}{6.022 \times 10^{23}} = 1.66 \times 10^{-24} g \]
If an element exists in the isotopic forms having atomic masses \( m_1, m_2 \) and \( m_3 \) in ratio \( x, y \) and \( z \) respectively, the average atomic mass
\[ M_{av} = \frac{m_1 \times x + m_2 \times y + m_3 \times z}{x + y + z} \]

Molecular Mass
It is the sum of atomic masses of the elements contained in one molecule. It is obtained by multiplying the atomic mass of each element by the number of their atoms and adding them together.
e.g. molecular mass of C₂H₆ is 2 \times 12.011 + 6 \times 1.008 = 30.07 u

Formula Mass
Formula mass of a substance is the sum of the atomic masses of all atoms in the formula unit of the compound. e.g. formula mass of C₂H₆ is 2 \times 12 + 6\times1 = 30 u.

Equivalent Mass
(or Equivalent Weight)
The number of parts of substance that combines directly or indirectly 1.008 parts by mass of hydrogen or 35.5 parts by mass of chlorine or 8 parts by mass of oxygen is called the equivalent mass of the substance.

Gram Equivalent Weight
An amount of a substance equal to its equivalent weight expressed in grams.

Tips for the Calculation of Equivalent Weight
Equivalent weight of metal
\[ \frac{\text{mass of metal}}{\text{mass of H₂ gas (displaced)}} \times 1.008 \]
Equivalent weight of metal
\[ \frac{\text{mass of metal}}{\text{mass of O₂ (combined)}} \times 8.0 \]
Equivalent weight of metal
\[ \frac{\text{mass of metal}}{\text{mass of chlorine (combined)}} \times 35.5 \]
Equivalent weight of acid
\[ \frac{\text{molecular mass of acid}}{\text{basicity}} \]
Equivalent weight of base
\[ \frac{\text{molecular mass of base}}{\text{acidity}} \]
Equivalent weight of salt
\[ \frac{\text{molecular mass of salt}}{\text{total positive or negative charge of metal atoms}} \]
Equivalent weight of a substance that undergoes oxidation/reduction
\[ \frac{\text{molecular mass}}{\text{change in oxidation number}} \]

Gram Atomic Mass (GAM) and Gram Molecular Mass (GMM)
Gram atomic mass is the quantity of the element whose mass in gram is numerically equal to its atomic mass. Gram atomic mass is also called one gram-atom of the element.
e.g. 1 gram-atom of Na = gram atomic mass of Na = 23 g
Similarly, quantity of a substance whose mass in gram is numerically equal to its molecular mass is called Gram Molecular Mass (GMM).
1 gram molecule of H₂O = gram molecular mass of H₂O = 18 g
Molecular formula

The mass of one mole of a substance in grams is known as its molar mass. The molar mass in grams is numerically equal to atomic or molecular formula mass in u. Units of molar mass are $9 \text{ mol}^{-1}$ or $\text{kg mol}^{-1}$.

e.g. Molar mass of NaCl = $23 + 35.5 = 58.5 \text{ g mol}^{-1}$

Empirical and Molecular Formula

Empirical formula is a simple whole number ratio of the elements constituting a molecule whereas the molecular formula is actual number or ratio of elements in a molecule.

Molecular formula = $(\text{empirical formula}) \times n$

where, $n = 1, 2, 3, \ldots$, etc.

Enhancer 3 Match the following.

<table>
<thead>
<tr>
<th>A</th>
<th>B</th>
<th>C</th>
<th>D</th>
</tr>
</thead>
<tbody>
<tr>
<td>A. Number of atoms in 0.5 mole atom of nitrogen.</td>
<td>1. $6.023 \times 10^{23}$</td>
<td></td>
<td></td>
</tr>
<tr>
<td>B. Avogadro’s Number</td>
<td>2. $2 \times 6.023 \times 10^{23}$ atoms</td>
<td></td>
<td></td>
</tr>
<tr>
<td>C. 46 gm Na contains</td>
<td>3. $3.01 \times 10^{23}$ atoms</td>
<td></td>
<td></td>
</tr>
<tr>
<td>D. Molecular formula</td>
<td>4. $n \times (\text{empirical formula})$</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Codes

<table>
<thead>
<tr>
<th>(a)</th>
<th>(b)</th>
<th>(c)</th>
<th>(d)</th>
</tr>
</thead>
<tbody>
<tr>
<td>3 1 2 4</td>
<td>1 3 2 4</td>
<td>1 3 4 2</td>
<td>4 2 1 3</td>
</tr>
</tbody>
</table>

Solution (a) $A \rightarrow (3) \Rightarrow \text{Number of moles} \times N_A$

(b) $B \rightarrow (1) \Rightarrow N_A = 6.023 \times 10^{23}$ particle

(c) $C \rightarrow (2) \Rightarrow \frac{46}{23} \times 6.023 \times 10^{23} = 2 \times 6.023 \times 10^{23}$ atom

(d) $D \rightarrow (4) \Rightarrow \text{Molecular formula} = n$

Solved Examples

**Example** The equivalent weight of a metal is 9 and vapour density of its chloride is 59.25. The atomic weight of metal is

\[
\begin{align*}
\text{(a)} & \quad 23.9 \\
\text{(b)} & \quad 27.3 \\
\text{(c)} & \quad 36.3 \\
\text{(d)} & \quad 48.3
\end{align*}
\]

**Solution**

\[
\begin{align*}
\text{Vapour density of metal chloride} & = 59.25 \\
\therefore & \quad \text{Molecular weight of metal chloride} \\
& = 2 \times \text{vapour density} \\
& = 2 \times 59.25 = 118.50 \\
\therefore & \quad \text{Valency of metal} = \frac{\text{mol. wt. of metal chloride}}{\text{eq. wt. of metal} + 35.5} \\
& = \frac{118.5}{9 + 35.5} = 2.66
\end{align*}
\]

**Example** Number of atoms of He in 100 u of He (atomic weight of He is 4) are

\[
\begin{align*}
\text{(a)} & \quad 25 \\
\text{(b)} & \quad 100 \\
\text{(c)} & \quad 50 \\
\text{(d)} & \quad 100 \times 6 \times 10^{-23}
\end{align*}
\]

**Solution**

\[
\begin{align*}
\text{Mass of one He atom} & = 4 \text{ u} \\
\therefore & \quad 4 \text{ u} = 1 \text{He atom} \\
\text{or} & \quad 100 \text{ u} = \frac{100}{4} = 25 \text{ He atoms}
\end{align*}
\]