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# Chemistry

## Physical Chemistry - 1

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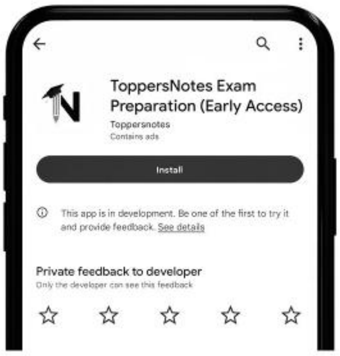


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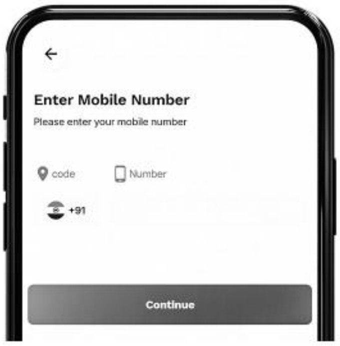


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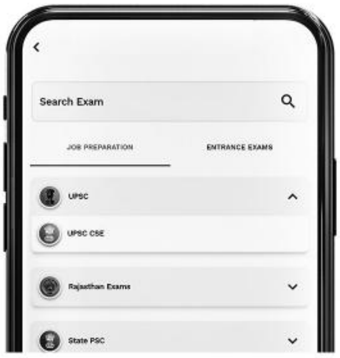


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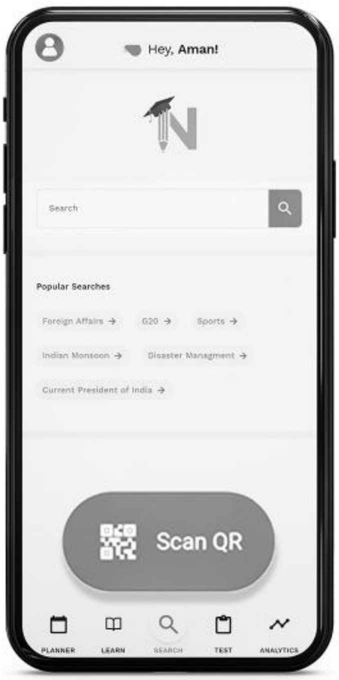


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# CHAPTER 1 Some Basic Concepts of Chemistry

CHAPTER OUTLINE

- Chemistry and its Importance
- Matter
- Avogadro's Law
- Mole Concept
- Mole Fraction
- Physical Quantities and their Measurements
- Laws of Chemical Combinations
- Mass
- Equivalent Weight
- Chemical Equation and Stoichiometry of Chemical Reactions

## CHEMISTRY AND ITS IMPORTANCE

Chemistry is the study of molecules and how they change. It looks at the composition of matter, the changes it goes through, and how changes in composition affect energy. Chemistry is a very important and central part of Science. It is an important part of meeting people's needs for food, health care products, drugs that save lives, etc. Using cis-platin, taxol, and other drugs may make it possible to cure cancer. AZT (Aziodothymidine) is a great thing for people who have AIDS. Even today, we still can't live without antiseptics like Detol, insecticides like DDT and BHC, and painkillers like Paracetamol.

## PHYSICAL QUANTITIES AND THEIR MEASUREMENTS

- Physical properties are mass, length and temperature time electric current.
- To express the measurement of any physical quantity we require its numerical value as well as its unit.
- Hence, the magnitude of a physical quantity can be given as **Magnitude of physical quantity = Its numerical value * Unit.**

| SI Unit | |
|---|---|
| **Measure** | **Unit** |
| Length (l) | Metre (m) |
| Mass (m) | Kilogram (kg) |
| Time (t) | Second (s) |
| Temperature (T) | Kelvin (K) |
| Current (i) | Ampere (A) |
| Intensity (I) | Candela (Cd) |
| Amount of Substance (n) | Mole (mol) |

### Derived Units

| Derived Units | |
|---|---|
| Concentration (C or S) | $= \frac{\text{Mass of solute}}{\text{Mass of solution}}$ $= \frac{\text{mol}}{\text{m}^3} = \text{molm}^{-3}$ |
| Volume (V) | = Length × Height × Breadth $= \text{m} \times \text{m} \times \text{m} = \text{m}^3$ |
| Density (d) | $= \frac{\text{Mass}}{\text{Volume}} = \frac{\text{kg}}{\text{m}^3} = \text{kg m}^{-3}$ |
| Velocity (v) | $= \frac{\text{Distance}}{\text{Time}} = \frac{\text{m}}{\text{sec}}$ $= \text{m sec}^{-1}$ |
| Acceleration (a) | $= \frac{\text{Change in velocity}}{\text{Time}}$ $= \frac{\text{m sec}^{-1}}{\text{sec}} = \text{m sec}^{-2}$ |
| Force (F) | = Mass × Acceleration $= \text{m} \times \text{a}$ $= \text{kg m sec}^{-1} = \text{Newton(N)}$ |

| | |
|---|---|
| Pressure (P) | $= \frac{\text{Force}}{\text{Area}} = \frac{\text{kg m sec}^{-2}}{\text{m}^2}$ $= \text{kg m}^{-1}\text{ sec}^{-2} =$ Pascal (Pa) |
| Work (W) | = Force × Displacement = $\text{F} \times \text{d}$ |
| | $= \text{kg m}^2\text{sec}^{-2} =$ Joule |

**Facts to Remember**

- Plane angle (Radian, that is, 'rad')
- Solid angle (Steradian, that is, 'str')

**Few Prefixes Used for Subsidiary Units**

**Sub multiples**

| | |
|---|---|
| 1 micro | $(\mu) = 10^{-6}$ |
| 1 nano | $(\text{n}) = 10^{-9}$ |
| 1 femto | $(\text{f}) = 10^{-15}$ |
| 1 atto | $(\text{a}) = 10^{-18}$ |
| 1 zepto | $(\text{Z}) = 10^{-21}$ |
| I yocta | $(\text{y}) = 10^{-24}$ |
| 1 Giga | $(\text{G}) = 10^{9}$ |
| 1 Tetra | $(\text{T}) = 10^{12}$ |
| 1 Exa | $(\text{E}) = 10^{18}$ |
| 1 Zetta | $(\text{Z}) = 10^{21}$ |
| 1 Yotta | $(\text{Y}) = 10^{24}$ |

1 litre $= 10^{-3}\text{m}^3 = 1\text{dm}^3$

1 atmosphere = 760 mm or torr

$= 101.325$ Pa or $\text{Nm}^{-2}$

1 bar $= 10^5\text{Nm}^{-2} = 10^5$ Pa

1 calorie $= 4.184$ joule

1eV (electron volt $= 1.602 \times 10^{-19}$ joule

1 joule $= 10^{-7}\text{erg}$

So, $1\text{eV} = 1.602 \times 10^{-12}\text{erg}$

$1\text{cal} > 1\text{ J} > 1\text{erg} > 1\text{eV}$

'Barn' is a unit of area to measure the cross section of nucleus.

1 Barn $= 10^{-28}\text{ m}^2 \approx 10^{-24}\text{ cm}^2$

## PRECISION AND ACCURACY

- The measurements are considered accurate when the average value of the different measurements is closer to the actual value.
- When a measurement is only slightly off from the real value, it is thought to be more accurate.
- Measurements are said to be precise when the values of different measurements are close to each other and to the average value. In fact, precision is just a way to measure how often an experiment can be repeated.

### Uncertainty in measurement and significant figures

- When measuring matter, there are some things that can go wrong with the numbers. We use significant figures to make sure that our measurements are correct.
- The number of significant figures is the number of digits in a number, including the last digit whose value is not known.
- For example, 14.3256±0.0001 has six significant figures.

### Rules to determine significant numbers

- **All of the non-zero digits and the zeros that come between them are important**. For example, the number 6003 has four important digits.
- **Zeros to the left of the first non-zero** digit in a number **are not significant figures.**
  For example, 0.00336 has only three significant figures.
- **If there are zeros to the right of the decimal point in a number that ends in zeros, these zeros are also important figures.**
  For example, 33.600 has five important figures.
- **Zeros at the end of a number without a decimal don't count as significant figures.**
  For example, 12600 only has three significant figures because there are no zeros at the end.
- **The result of dividing or multiplying** by a smaller number must have the **same number of significant figures as the smaller number**.
  For example, 3.331 × 0.011 = 0.036641 ≈ 0.037
- **The result of a subtraction or addition** must have **the same number of significant figures as the least precise term**.
  For example, 5.1 + 7.21 + 8.008 > 20.318 ≈ 20.32.

### Rounding-off non-significant figures

Rounding-off non-significant figures means dropping of the uncertain or non-significant digits in a number. It is possible as follows:

- If the last digit to be rounded is greater than 5, the number before it is increased by one.
  - For example, 3.17 is rounded to 3.2.
- If the last digit to be rounded off is less than 5, the number before it stays the same.
  - For example, 5.12 is rounded to 5.1.
- If the right-most digit to be rounded off is 5, the number before it stays the same if the value is even. But if the number is odd, it is rounded up by one.
  - For example, 4.45 is rounded to 4.4 and 5.35 is rounded to 5.4.

### Exponential notation or scientific notation

- If a number ends in zeros that are not to the right of the decimal point, the zeros do not have to be important.
  - For example, 290 has 2 or 3 significant figures and 19500 has 3,4 or 5 significant figures.
- This confusion can be removed when the values are expressed in terms of scientific notations-
  - For example, 19500 can be written as $1.95 \times 10^4$ (3 significant figures), $1.950 \times 10^4$ (4 significant figures), $1.9500 \times 10^4$ (5 significant figures).
  - In this kind of notation, every number can be written as $\mathrm{N} \times 10^{\mathrm{n}}$ where, $\mathrm{n} =$ Integer, $\mathrm{N} =$ Number with non-zero digit to the left of the decimal point.

## MATTER

- Matter is any kind of thing that has mass and takes up space.
- It can be in three different forms:-
  - Solid
  - Liquid
  - Gas
- At the bulk level, or macroscopic level, we can further divide matter into **pure substances** and **mixtures.**

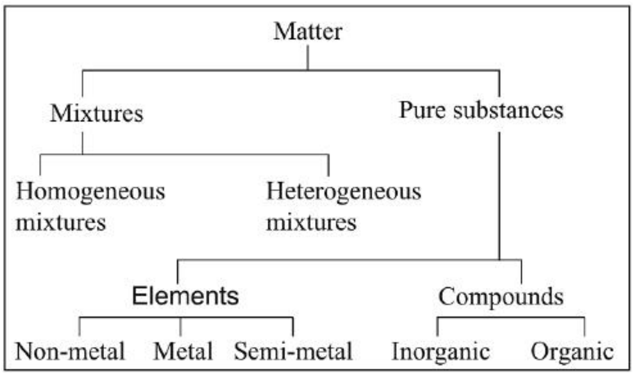



**Figure:** Classification of Matter

| | **Mixture** | **Pure substance** |
|---|---|---|
| **Composition** | Composed of two or more substances which are known as its components or constituents (in any ratio). | Pure substances have fixed compositions. |
| **Separation** | The components can be separated with the help of physical separation methods like filtration, crystallization, distillation. | Their constituents cannot be separated by using simple physical methods of separation. |
| **Types** | Homogeneous and Heterogeneous | Element and Compound |
| **Description** | • **Homogeneous mixture**: all the components undergo complete mixing forming a uniform composition <u>Example</u>- Air or sugar solution. | • **Element**: composed of one type of particle which could either be atoms or molecules. <u>Examples</u>- Na, Cu and Ag have only one type of atoms. |

| | | |
|---|---|---|
| | • **Heterogeneous mixture**: the composition formed due to the mixing of components is not entirely uniform. Example- Grains mixed with dust | • **Compound**: formed by the combination of two or more atoms or different elements. Examples- $H_2O, CO_2$. |

## Dalton's Atomic Theory

- An **atom** is the **smallest particle** of an element that is neutral in nature, keeps all of the element's properties, and participates in a chemical reaction.
- Dalton popularised the term "**atom**" (alamos means undivided).
- Dalton's atomic hypothesis was proposed **on the basis of chemical combination principles.**

**Main assumptions**

- Atoms are the **building blocks** of all matter.
- An atom is the smallest fundamental particle that is not broken up. (A building block for any kind of life).
- An atom **can't be made or broken,** and it can't change into an atom of another element.
- Atoms of the same element are the same size, have the same energy, and have the same properties. Atoms of different elements are different in these ways.
- A molecule is made up of atoms that come together in **whole number ratios**. Because of this, a molecule is the smallest thing that has its own identity.

**Modern view about atom**

- An atom can be divided into smaller particles known as subatomic particles.
- It can also combine in non-whole number ratios, as in non-stoichiometric compounds (Berthollide compounds) such as $Fe_{0.93}O$.
- Atoms of the same element differ in mass and mass-related properties, as do isotopes.
- Atoms are rearranged in a chemical reaction.

**Molecule**

- Avogadro came up with the word "molecule."
- It is the smallest unit of matter that can exist on its own and still have all the properties of the substance it is made of.
- Most molecules have a diameter between 4 and 20 Å and a molecular mass between 2 and 1000.

**Berzelius Hypothesis**

- The Berzelius hypothesis says, "*All gases with the same volume have the same number of atoms when the temperature and pressure are the same*."
- This concept, when applied to the law of combining volumes, proves that atoms can be broken up, which goes against Dalton's theory.

# LAWS OF CHEMICAL COMBINATIONS

## Law of Conservation of Mass

- Lavoisier created the concept of the law of conservation of mass in 1774.
- It was verified by Landolt.
- According to this law, "*In a chemical change, the total mass of the products is equal to the total mass of the reactants, that is, mass is neither created nor destroyed*."
  For example: When a solution with calculated weight of $AgNO_3$ and $NaCl$ is mixed, white precipitates of $AgCl$ are formed while $NaNO_3$ remains in solution. The weight of the solution remains the same before and after this experiment.
- It can't be used to explain nuclear reactions.

### Law of Constant Composition or Law of Definite Proportion

- Law of constant composition was created by Proust in 1779.
- It was verified by Star and Richards.
- According to this law, "*A chemical compound always contains same elements combined together in same proportion by mass*."
  For example: $NaCl$ extracted from sea water or achieved from deposits will have $23$ g $Na$ and $35.5$ g of chlorine in its one mole.
- It is doesn't work for non-stoichiometric compounds like $Fe_{0.93}O$.

### Law of Multiple Proportion

- Dalton proposed the Law of multiple proportion in 1804.
- It was verified by Berzilius.
- According to this law, "*Different weights of an element that combine with a fixed weight of another element bear a simple whole number ratio*."
  For example: In case of $CO$, and $CO_2$ weight of oxygen which combines with $12$ g of carbon is in 1:2 ratio.
- It is used when different isotopes of the same element are used to make the same compound. Examples: $H_2O, D_2O$.

### Law of Reciprocal Proportion

- Richter proposed the Law of reciprocal proportion in 1792.
- It was verified by Star.
- According to this law, "*When two different elements undergo combination with same weight of a third element, the ratio in which they combine will either be same or some simple multiple of the ratio in which they combine with each other*."
- Also known as **Law of equivalent proportion** which states "*Elements always combine in terms of their equivalent weight*."

### Law of Combining Volume

- Gay Lussac proposed the Law of combining volume.
- It **applies to gases.**
- According to this law, "*When gases react with each other they bear a simple whole number ratio with one another as well as the product under conditions of same temperature and pressure*."

### Avogadro's law

- Avogadro's law explains **law of combining volumes.**
- According to this law, "*Under similar conditions of temperature and pressure equal volume of gases contain equal number of molecules*."
- Mathematically, **Avogadro's law is $\mathbf{V \propto n}$.**
- It is used in:
  1 Deriving molecular formula of a gas
  2 Determining atomicity of a gas
  3 Deriving a relation Molecular mass $=$ $2 \times$ Vapour Density ($M = 2 \times$ V.D.)
  4 Deriving the gram molecular volume
- **Avogadro number** $(\mathbf{N_0}$ **or** $\mathbf{N_A}) = 6.023 \times 10^{23}$.
- Avogadro number of gas molecules occupies **22.4 litre or $\mathbf{22400\ mL}$ or $\mathbf{cm^3}$ volume at STP.**
- The number of molecules in $1\ cm^3$ of a gas at STP is equal to Loschmidt number, that is, $2.68 \times 10^{19}$.
- **Reciprocal of Avogadro number** is known as **avogram**.

## MASS

Mass can be expressed in terms of atoms or molecules as follows:

1) Atomic mass
2) Molecular mass

## Atomic Mass

- Atomic mass is the **relative mass of an atom** which shows the number of times an atom is **heavier than $\frac{1}{12}$ mass of C-12.**

  **Atomic Mass** $= \frac{\text{Mass of one atom of an element}}{\frac{1}{12}\text{ mass of one C-atom}}$
- The **gram atomic mass (GAM)** or gram atom is the atomic mass of any element measured in grams
- A gram atom has number of atoms of the element.
- **Atomic mass** $= \mathbf{E} \times \mathbf{V}$

  Here, $\mathrm{E}$ = Equivalent weight, $\mathrm{V}$ = Valency
- **Dulong Petit's Law states that:**

  **Atomic mass** $= \frac{6.4}{\text{Specific heat in calories}}$

### Atomic Mass Unit

- It is equal to 1/12 of the mass of one atom of C-12 and is written as a.m.u.

  1 A.m.u $= \frac{\mathbf{1.99 \times 10^{-23}}}{\mathbf{12}} = \mathbf{1.66 \times 10^{-24}}$

  Atomic Mass $= \frac{\text{Mass of one atom of an element}}{\text{1 A.m.u}}$

Here $1.99 \times 10^{-23}$ g is wt. of one C-12-atom.

Average Atomic Mass:

(At. Mass) $Av. = \frac{\mathrm{M_1 \times a + M_2 \times b + M_3 \times C}}{a+b+c}$

Here $M_1, M_2$ and $M_3$ are masses of isotopes and a, b, c are their percentage ratio.

## Molecular Mass

- The total mass of a molecule is its molecular mass, or the number of times it is heavier than 1/12 the weight of a C-12 atom or 1/16 the weight of an O-atom.
- It is non-variable.

### Determination of molecular mass

**A) Vapour density method**

**Mol. mass** $= \mathbf{2} \times$ **V.D.**

**V.D.** $= \frac{\mathbf{W \times 22400}}{\text{Volume at STP (in mL)}}$

Here, $\mathrm{W}$ = Weight of substance in g, V.D. = Vapour density

**B) Diffusion method**

- It is based on **Graham's law of diffusion.**
- Graham's law states that *"The rate of diffusion of different gases, under similar conditions of temperature and pressure are inversely proportional to the square roots of their density (or molecular weights)".*

  Mathematically,

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

Here, $r_1$ and $r_2$ are rates of diffusion/effusion for two species, $M_1$ and $M_2$ are their molecular masses and $d_1$ and $d_2$ are the densities of the 2 species respectively.

**C) Colligative properties method**

**Mathematically,** $\boldsymbol{\pi}\mathbf{V} = \frac{\mathbf{W}}{\mathbf{m}}\mathbf{RT}$

Where, $\pi$ = Osmotic pressure in atm

$\mathrm{V}$ = Volume in litre

$\mathrm{W}$ = Weight in gram

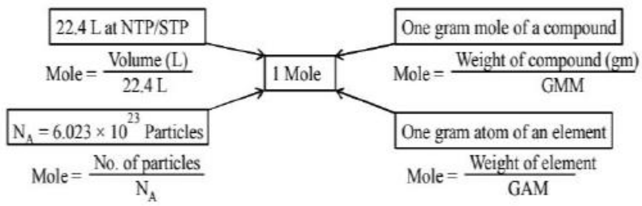
$\mathrm{R}$ = Universal gas constant

$\mathrm{T}$ = Given temperature

$\mathrm{m}$ = Molar mass

## MOLE CONCEPT

- **Mole is related to mass of the substance, volume of the gaseous substance**




Here, GMM= Gram molecular mass

GAM= Gram Atomic mass

**Facts to Remember**

- 1 Mole $= 6.023 \times 10^{23}$ particles.
- 1 Mole of atoms $= 6.023 \times 10^{23}$ Atoms.
- 1 Mole of molecules $= 6.023 \times 10^{23}$ molecules
- 1 Mole of electrons $= 6.023 \times 10^{23}$ electrons.
- The number $6.023 \times 10^{23}$ is called Avogadro number ($N_A$)

A mole of any substance (like $N_2$ ) stands for:

- $6.023 \times 10^{23}$ molecules of $N_2$
- $2 \times 6.023 \times 10^{23}$ atoms of nitrogen
- 28 g of nitrogen
- 22.4 litre of $N_2$ at STP

## To Find Total Number of Identities

- Number of moles " $n$ ",
$$= \frac{\text{Weight in grams}}{\text{Gram molecular weight}} = \frac{\mathbf{W}}{\mathbf{M}}$$
- Weight of a substance containing a definite number of moles = Number of moles × Gram molecular weight ($\mathbf{W} = \mathbf{n} \times \mathbf{MW}$)
- No. of molecules present in a given substance.
$$= \frac{\text{Weight}}{\text{Gram molecular weight}} \times \text{Avogadro number}$$
$$= \frac{\mathbf{W}}{\mathbf{MW}} \times \mathbf{N}$$
- No. of atoms present in a given element
$$= \frac{\text{Weight}}{\text{Gram molecular weight}} \times \text{Avogadro number}$$
- Total Number of Molecules = **mole (n) × $N_A$**
- Total Number of Atoms in one molecule = mole (n) × $N_A$ × Number of atoms present in one molecule
- **Total Number of Electrons in one molecule = mole $(\mathbf{n}) \times \mathbf{N_A} \times$ Number of electron present in one molecule**
- **Total charge on any ion = mole $(\mathbf{n}) \times \mathbf{N_A} \times$ charge on one ion $\times \mathbf{1.6 \times 10^{-19} C}$**

## Equivalent Weight

- The weight of an element or compound is called its "**equivalent weight**" if it combines with or takes the place of 1.008 parts by weight of $H_2$ or 8 parts by weight of $O_2$ or 35.5 parts by weight of $Cl_2$.
- When equivalent weight is written in gram, the number is called **gram equivalent**.
- It depends on what kind of chemical reaction is going on between the two substances.

## Methods to Find Equivalent Weight

**(a) For acids**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Basicity of acid}}$$

**Examples:** $H_3PO_4, E = \frac{M}{3}$

For $H_3PO_3, E = \frac{M}{2}$,

For $H_2SO_4, E = \frac{M}{2}$, for $\begin{array}{c}COOH\\|\\COOH\end{array} \times 2H_2O, E = \frac{M}{2}$

For $H_3PO_2, E = \frac{M}{1}$,

**(b) For bases**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Acidity or number of } \mathbf{OH^-} \textbf{ ions}}$$

**Examples**: $Ca(OH)_2, E = \frac{M}{2}$

For $Al(OH)_3, E = \frac{M}{3}$

**(b) For ions**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Charge on ion}}$$

**Examples**: $SO_4^{2-}, E = \frac{M}{2}$

For $PO_4^{3-}, E = \frac{M}{3}$

**(d) For compounds**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Valency of cation or anion}}$$

**Examples**: $CaCO_3, E = \frac{M}{2}$; $Na_2CO_3, E = \frac{M}{2}$

For $AlCl_3, E = \frac{M}{3}$

**(e) For redox reactions**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Total change in oxidation number}}$$

**Example:** $KMnO_4$

(1) In acidic medium, $E = \frac{M}{5}$

$$2\overset{+7}{KMnO_4} + 3H_2SO_4 \longrightarrow K_2SO_4 + 2\overset{+2}{MnSO_4} + 3H_2O + 5[O]$$

5 units change in oxidation number

(2) In basic medium, $E = \frac{M}{1}$

$$2\overset{+7}{KMnO_4} + 2KOH \longrightarrow 2\overset{+6}{K_2MnO_4} + H_2O + [O]$$

One unit change in oxidation number

(3) In neutral medium, $E = \frac{M}{3}$

$$2\overset{+7}{KMnO_4} + H_2O \longrightarrow 2KOH + 2\overset{+4}{MnO_2} + 3[O]$$

3 units change in oxidation number

**(f) For acidic salts**

$$\mathbf{E} = \frac{\textbf{Molecular weight}}{\textbf{Number of replaceable H-atoms}}$$

For $H_3PO_4$, for example,

$$Ca(OH)_2 + H_3PO_4 \longrightarrow CaHPO_4 + 2H_2O$$

$$E = \frac{M}{2} (M = \text{ Molar Mass of Acid })$$

**Some Other methods**

**(a) Hydrogen displacement method**

$$E = \frac{W \times 11200}{\text{Volume of } H_2 \text{ at NTP}}$$

**(b) Oxide formation method**

$$E = \frac{\text{Wt. of metal}}{\text{Wt. of oxygen}} \times 8$$

Weight of oxygen = Weight of metal oxide - Weight of metal

**(c) Chloride formation method**

$$E = \frac{\text{Wt. of metal}}{\text{Wt. of chloride}} \times 35.5$$

Weight of chloride = Weight of metal chloride – Weight of metal

**(d) Double decomposition method**

$$\frac{\text{Eq. wt. of salt taken}}{\text{Eq. wt. of salt ppt.}} = \frac{\text{Wt. of salt taken}}{\text{Wt. of salt ppt.}}$$

**(e) Metal displacement method**

$$\frac{E_1}{E_2} = \frac{W_1}{W_2}$$

## CHEMICAL EQUATION AND STOICHIOMETRY OF CHEMICAL REACTIONS

- A stoichiometric equation is a description of a chemical reaction that is balanced.
- In a stoichiometric equation, the coefficients of the reactants and products show how much of each one is needed for the reaction to happen.
- In an irreversible reaction, the reactant that is completely used up is called the **limiting reagent**. The remaining reactant is called the **excess reagent**.
  For example, if 20g of calcium is burned in 32g of $O_2$, then Ca is the limiting reagent and $O_2$ is the excess reagent.
- Using stoichiometric calculations, you can figure out if it makes economic sense to make a certain substance or not.
- These stoichiometric calculations are of following four types:
  1. Calculations based on weight-weight relationship
  2. Calculations based on weight-volume relationships
  3. Calculations based on volume-volume relationships
  4. Calculations based on weight-volume-energy relationships
- If you know how much of the reactant is used in a certain reaction, you can figure out how much of the other substance is needed or how much of the product is made.
- For stoichiometric calculations, you must take into account the following steps:
  1. To write a balanced chemical equation, you must use the chemical formulas of the reactants and the products.
  2. Here, the coefficients of a balanced chemical equation show the mole ratio of the reactants and products.
  3. This mole ratio can be changed into a weight-weight ratio (w/w), a weight-volume ratio (w/v), or a volume-volume ratio (v/v). These are called percentage by weight, percentage by volume, and percentage by strength.

### Empirical and Molecular Formulas

**Empirical formula**:

- It is the easiest way to figure out how many atoms of each element are in one molecule of a compound.

- It doesn't show how many atoms of each element are actually in one molecule of the compound.
- Example: Empirical formula of benzene is "CH".

**Calculation of the empirical formula**

Following steps are involved in the calculation of the empirical formula:

(i) First, figure out the proportion of each element there is in the compound in terms of weight.

$$C\% = \frac{12}{44} \times \frac{\text{Wt. of } CO_2}{\text{Wt. of organic comp.}} \times 100$$

$$H\% = \frac{2}{18} \times \frac{\text{Wt. of } H_2O}{\text{Wt. of organic comp.}} \times 100$$

**(a) Duma's method:**

$$N\% = \frac{28}{22400} \times \frac{\text{Volume of } N_2 \text{ at S.T.P.}}{\text{Wt. of organic comp.}} \times 100$$

OR

$$N\% = \frac{\text{Volume of } N_2 \text{ at S.T.P.}}{8 \times \text{Wt. of organic comp.}}$$

**(b) Kjeldahl's Method**

$$N\% = \frac{1.4 \times N \times V}{\text{Wt. of org. compound}}$$

Here $N$ = normality of acid the used to neutralize ammonia

$V$ = volume of acid the used to neutralize ammonia

(ii) Now, the percent of each element is divided by its atomic weight to find out how many atoms of each element there are in total.

(iii) These relative numbers obtained are divided by smallest number to get the simplest ratio numbers.

(iv) If the simplest ratio number is not a whole number, it should be multiplied by a suitable integer to make it a whole number.

(v) The empirical formula of the compound can be found by taking the ratio of these simple whole numbers.

**Molecular formula**:

- It represents the **actual number of atoms of various elements present in** a single molecule of the chemical.
- Example: Molecular formula of benzene is $C_6H_6$.

  Mathematically,

$$\mathbf{n} = \frac{\textbf{Molecular weight}}{\textbf{Empirial formula weight}}$$

- **Molecular formula = Empirical formula $\times$ n.**
- Molecular weight of a substance can be determined by various methods like
  1) Vapour density method
  2) Elevation in boiling point method
  3) Depression in freezing point methods
- **Victor Mayer's method** is used to determine the molecular weight of volatile compounds.
- **Molecular Weight $= 2 \times$ Vapour density**
- **Molecular weight = empirical formula wt. $\times$ '$n$'**.

  **Note:** In some cases, the molecular formula and empirical formula are the same.

**Important Points**

1. Giorgi introduced MKS system.
2. There are infinite number of significant figures in $\pi$.
3. 1 mole of $H_2O \neq 22400$ mL or cc of $H_2O$ (since it is liquid)
   1 mole of $H_2O = 18$cc of $H_2O$ (as density of $H_2O = 1$ g/cc )
4. Mass of one mole of $e^-$ = Mass of one $e^- \times N_A$
   $= 9.1 \times 10^{-31} \times 6.02 \times 10^{23} = 0.55$ mg
5. 20 carat gold is a mixture of 20 parts gold and 4 parts copper by weight.
6. **Efflorescent substances**: Some substances like $CuSO_4 \cdot 5H_2O, Na_2CO_3 \cdot 10H_2O$ have a tendency to lose water in air. These are called **efflorescent substances** and this tendency is called **efflorescence**.

7. **Deliquescent**: Certain solid chemicals, such as NaOH and KOH, have a strong tendency to absorb moisture from the air and become wet. This is referred to as deliquescence.
8. Quicklime (CaO), anhydrous $P_2O_5$, and other hygroscopic compounds absorb moisture from the air.
9. **Isomorphs** are compounds with comparable chemical compositions in the same crystalline form.
   For instance, all alums
   $[M_2SO_4.M_2(SO_4)_3.24H_2O]$
   Here, $M =$ Monovalent $(K)$
   $M =$ Trivalent (Al)
   $FeSO_4.7H_2O$ (Green vitriol) and $ZnSO_4.7H_2O$
10. **Polymorphs** are different crystalline forms of a substance, and this phenomenon is known as polymorphism.
    For example, $ZnS \rightarrow$ Zinc blende
    $\downarrow$
    Wurtzite
11. **Silver Salt Formation Method** is used to find equivalent weight of an acid
    $$\frac{\text{Eq. wt. of RCOOAg}}{108} = \frac{\text{Wt. of RCOOAg}}{\text{Wt. of Ag}}$$
    Eq. wt. of $R-COOH$ = Eq. wt. of RCOOAg - 107
12. **Equivalent weight** $(\mathbf{E}) =$ Weight deposited by 96500 coulombs or 1 Faraday.

## EXPRESSION OF STRENGTH / CONCENTRATION OF SOLUTION

The concentration of the solution, or the amount of substance present in a particular volume, can be represented in several ways:

**1. Mass Percent or Weight Percent (w/W %)**

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

**(i) Weight-weight percent (w/W):**

Weight percent

$$= \frac{\text{Weight of solute (gm)}}{\text{Weight of solution (gm)}} \times 100$$

**(ii) Volume-volume percent $(\mathbf{v/V})$ :**

Volume-volume percentage

$$= \frac{\text{Volume of solute (ml.)}}{\text{Volume of solution (ml.)}} \times 100$$

**(iii) Weight-volume percentage $(\mathbf{w/V})$ :**

Weight - volume percentage

$$= \frac{\text{Weight of solute (gm)}}{\text{Volume of solution (ml)}} \times 100$$

**2. Normality**

- The number of solute gram equivalents dissolved per litre of solution.
- It is denoted by ' $N$ ' :
- $$\text{Normality} = \frac{\text{Number of gram equivalents of solute}}{\text{Volume of solution (lit.)}}$$

  $\because$ Gram equivalents of solute

  $$= \frac{\text{Weight of solute (gm)}}{\text{Equivalent weight of solute}}$$

**3. Mole Fraction**

- Mole fraction is the number of moles of one part of a solution divided by the total number of moles in the solution. In the case of a binary solution with two parts A and B, it is written as X.

  $$X_A = \frac{n_A}{n_A + n_B}$$

  $$X_B = \frac{n_B}{n_A + n_B}$$

  $$X_B + X_B = 1$$

- Mole fraction of solute $(X_2) = \frac{n_2}{n_1 + n_2}$

  Here, $n_1$ and $n_2$ represent number of moles of solvent and solute respectively.
- Mole fraction doesn't change with temperature because both the solute and the solvent are given in terms of their weights.

**4. Molarity**

- Number of moles of solute in 1 litre of solution.

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

**5. Molality**

- Number of moles of solute present in 1 kg of solvent.

$$\text{Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$

**6. ppm (Parts per million)**

- The parts of the component per million parts $(10^6)$ of the solution.

$$\text{ppm} = \frac{w}{w+W} \times 10^6$$

where, $w =$ weight of solute, $W =$ weight of solvent

**Important Formula**

1. Mole $(n) = \frac{W}{M} = \frac{V}{22.4\ L} = \frac{N}{N_A}$

   Here, $W =$ Weight
   $M =$ Molecular weight
   $N =$ Number of atoms / molecules
   $N_A$ = Avagadro no.
   $V =$ Volume in litre

2. Molar mass $(M) = 2 \times$ V.D. (Vapour density)

   $$\text{V.D.} = \frac{W \times 22400}{\text{Volume at STP (in mL)}}$$

   Here, $W =$ Weight of substance in g

3. $\frac{r_1}{r_2} = \sqrt{\left(\frac{m_2}{M_1}\right)}$

   Here $r_1, r_2$ are rates of diffusion for two species while $M_1, M_2$ are their molecular masses respectively.

4. $\pi V = \frac{W}{M} RT$

   Here, $\pi =$ Osmotic pressure in atm
   $V =$ Volume in litre
   $W =$ Weight in gram
   $R =$ Universal gas constant
   $T =$ Given temperature
   M = Molecular Weight

5. Equivalent weight (E)

   $$E = \frac{M}{H^+ \text{ or } OH^- \text{ or Charge}}$$

   $$E = \frac{M}{\text{Change in oxidation number}}$$

   $$E = \frac{M}{\text{Number of replaceable hydrogen atoms}}$$

   $$\frac{E_1}{E_2} = \frac{W_1}{W_2}$$

6. Specific gravity $= \frac{\text{Mass of Liquid}}{\text{Volume of Liquid}}$

7. $M = \frac{\text{\% by mass} \times d \times 10}{\text{M. wt. of solute}}$ ;

   $$N = \frac{\text{\% by mass} \times d \times 10}{\text{Eq. wt. of solute}}$$

   $$m = \frac{M \times 100}{(1000 \times d - M \times \text{M.Wt})}$$

   (Here ' $d$ ' is density of solution in $gcm^{-3}$, $M$ is molarity, $N$ is normality and ' $m$ ' is molality)

8. Vapour Density $= \frac{\text{Mass of } V_{(L)} \text{ of Gas}}{\text{Mass of } V_{(L)} \text{ of } H_2}$

## SOLVED NUMERICALS

### Mole Concept

**1. If 1 Faraday was to be 60230 coulombs instead of 96500 coulombs, what will be the charge on an electron?**

**Solution:**

As One mole electron carries 1 Faraday charge.

$6.023 \times 10^{23}$ electrons carry $= 60230C$

So, 1 electron carries $= \frac{60230C}{6.023 \times 10^{23}} = 1 \times 10^{-19}C$.

**2. Calculate the number of atoms of oxygen present in $88$ g of $CO_2$. What would be the mass of $CO$ having the same number of oxygen atoms?**

**Solution**:

Number of moles of $CO_2 = \frac{88\,g}{44\,g\ mol^{-1}}$

$= 2$ moles

1 mole of $CO_2$ contains 2 moles of oxygen atoms, 2 moles of $CO_2$ will contain 4 moles of oxygen atoms.

Number of oxygen atoms $= 4 \times 6.023 \times 10^{23} = 2.5092 \times 10^{24}$

1 mole oxygen atom is present in 1 mole of CO, 4 moles oxygen atoms are present in 4 moles of CO

Its mass is $4(12 + 16) = 112$ g.

3. **Calculate the total number of electrons present in $1.6$ g of methane.**

**Solution:**

Molecular mass of methane $= 16\ \text{g mol}^{-1}$

$16\ \text{g}\ CH_4$ contains $6.02 \times 10^{23}$ molecules of $CH_4$

$1.6\ \text{g}\ CH_4$ contains $6.02 \times 10^{22}$ molecules of $CH_4$

As one molecule of $CH_4$ contains $(6 + 4) = 10$ electrons, $6.02 \times 10^{22}$ molecules of $CH_4$ will have $10 \times 6.02 \times 10^{22} = 6.02 \times 10^{23}$ electrons.

4. **How many atoms of carbon has a young man given to his bride-to-be if the engagement ring contains 0.5 carat diamond? ($1$ carat $= 200$mg)**

**Solution:**

Mass of diamond (C) = 0.5 × 200mg

= 100 mg = 100 × $10^{-3}$ g

= 0.1 g

Number of mole of C $= \frac{0.1\,g}{12\,g\ mol^{-1}}$

$= \frac{1}{120}$ mole

Number of C atoms $= \frac{1}{120} \times 6.023 \times 10^{23}$

$= 5.02 \times 10^{21}$

5. **A mixture of aluminium and zinc weighing $1.67$ g was completely dissolved in acid and the evolved 1.69 litres of hydrogen gas was measured at $273$ K and one atmosphere pressure. What was the mass of aluminium in the original mixture?**

**Solution**:

Let the mass of aluminium in the sample be 'A' g.

The mass of Zn = (1.67 – A)g

The volume of $H_2$ at NTP given by Al

$$= \frac{3 \times 22.4 \times A}{2 \times 27}\ L \qquad (1)$$

The volume of $H_2$ at NTP given by Zn

$$= \frac{(1.67 - A)22.4}{65.4}\ L \qquad (2)$$

From equation (1) and (2),

$$\frac{3 \times 22.4 \times A}{54} + \frac{(1.67 - A)22.4}{65.4} = 1.69$$

142.2 × A = 176.26

A = 1.248g

6. **Find the equivalent mass of $H_3PO_4$ in the reaction:**

$\mathbf{Ca(OH)_2 + H_3PO_4 \longrightarrow CaHPO_4 + 2H_2O}$

**Solution:**

As in this reaction only two hydrogen atoms are replaced so its equivalent mass will be given by the following expression: Equivalent mass of $H_3PO_4$

$$= \frac{\text{Molecular mass of } H_3PO_4}{2}$$

$$= \frac{98}{2} = 49$$

7. **How many years would it take to spend Avogadro number of rupees at the rate of 10 lac rupees per second?**

**Solution:**

Avogadro number $= 6.023 \times 10^{23}$

Total rupees $= 6.023 \times 10^{23}$

Rate of spending = 10 lac rupees/s = $10^6$ Rs/s

Number of years to spend all the rupees

$$= \frac{6.023 \times 10^{23}\ \text{Rs}}{10^6 \times 60 \times 60 \times 24 \times 365\ \text{Rs/year}}$$

= $1.90988 \times 10^{10}$ years

**8. $2.68 \times 10^{-3}$ moles of a solution containing an ion $A^{n+}$ required $1.61 \times 10^{-3}$ moles of $MnO_4^-$ for the oxidation of $A^{n+}$ to $AO_3^-$ in an acidic medium. What is the value of ?**

**Solution:**

$1.61 \times 10^{-3}$ M$KMnO_4 \equiv 2.68 \times 10^{-3}$ M solution of $A^{n+}$

$$M/5KMnO_4 = \frac{2.68 \times 10^{-3} M}{1.61 \times 10^{-3}} \times \frac{M}{5}$$

$$0.33M = \frac{M}{5-n}$$

$$5-n = \frac{1}{0.33} = 3$$

n = 2

## Concentration terms

**9. $50$ mL of $10$ N$H_2SO_4$, $25$ mL of $12$ NHCl and $40$ mL of $5$ N $HNO_3$ are mixed and the volume of the mixture is made $1000$ mL by adding water. Find the normality of the resulting solution.**

**Solution:**

$N_1 V_1 + N_2 V_2 + N_3 V_3 = N_R V_R$

$10 \times 50 + 12 \times 25 + 5 \times 40 = N_R \times 1000$

$500 + 300 + 200 = N_R \times 1000$

$N_R = 1$

Hence, the normality of resulting solution is 1.

**10. The formula weight of an acid is 82. In a titration, $100\ cm^3$ of a solution of this acid containing $39.0$ g of the acid per litre were completely neutralized by $95.0\ cm^3$ of aqueous NaOH containing $40.0$ g of NaOH per litre. What is the basicity of the acid?**

**Solution:**

Normality of NaOH = 1

Normality of acid $= \frac{1 \times 95}{100} = 0.95$

Suppose the equivalent mass of the acid is E.

$$\frac{39}{E} = 0.95$$

E = 41

Therefore, basicity $= \frac{82}{41} = 2$.

**11. One g of impure $Na_2CO_3$ is dissolved in water and the solution is made upto $50$ mL. To $50$ mL of this made up solution, $50$ mL of $0.1$ N HCl is added and the mixture after shaking well required $10$ mL of $0.16$ N sodium hydroxide solution for complete neutralization. Calculate the per cent purity of the sample of $Na_2CO_3$.**

**Solution:**

Strength of the $Na_2CO_3$ solution $= 4\ g\ L^{-1}$

Suppose the normality of $Na_2CO_3$ solution $= N_x$

As after mixing $Na_2CO_3$ and HCl solution, NaOH solution is added, so, according to the normality equation

$$50 \times N_x + 0.16 \times 10 = 50 \times 0.1$$
$$N_x = 0.068\ N$$

**Strength ($gL^{-1}$) = Normality × Equivalent mass**

$= 0.068 \times 53 = 3.6\ gL^{-1}$

So, purity of $Na_2CO_3 = \frac{3.6 \times 100}{4}$

= 90%

**12. A small amount of $CaCO_3$ completely neutralizes $525$ mL of $0.1$ N HCl and no acid is left in the end. After converting all calcium chloride to $CaSO_4$, how much plaster of paris can be obtained?**

**Solution:**

525 mL of 0.1 N HCl = 525 mL of 0.1 N $CaCl_2$

= 525 mL of 0.1 N plaster of paris

Molecular mass of plaster of paris = 145

Equivalent mass of plaster of paris $= \frac{145}{2} = 72.5$

Mass of plaster of paris in 525 mL of 0.1 N solution

$$= \frac{N \times E \times V}{1000}$$

$$= \frac{0.1 \times 72.5 \times 525}{1000}$$

= 3.806 g

## Calculations Based on Reactions

**13. Metallic tin in the presence of $HCl$ is oxidized by $K_2Cr_2O_7$ solution to stannic chloride. What volume of deci-normal dichromate solution would be reduced by $1$ g of n ?**

**Solution**:

$$\overset{+6}{K_2Cr_2O_7} + \overset{0}{Sn} \longrightarrow \overset{+3}{Cr^{+3}} + \overset{+4}{SnCl_4}$$

$n_{K_2Cr_2O_7} = +3-(+6)\times 2 = 6$

$n_{Sn} = (+4-0)\times 1 = 4$

From law of equivalence

Equivalents of $K_2Cr_2O_7$ = Equivalents of Sn

$$\frac{1}{10}\times V_{K_2Cr_2O_7} = \frac{1}{119}\times 4$$

$V_{K_2Cr_2O_7} = 337L$

**14. Find the weight of iron which will be converted into its oxide by the action of $18$ g of steam.**

**Solution:**

The reaction is

$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$

As 4 moles steam reacts with 3 moles Fe

So 1 mole (18g) steam reacts with 3/4 moles Fe

$= \frac{3}{4}\text{mole}\times 56\text{g mol}^{-1}$

= 42g Fe.

**15. The mineral haematite is $Fe_2O_3$. Haematite ore contains unwanted material called gangue in addition of $Fe_2O_3$. If $5.0$ kg of ore contains $2.78$ kg of Fe, what per cent of ore is $Fe_2O_3$ ?**

**Solution:**

$2Fe \equiv Fe_2O_3$

2 × 55.85g

2 × 55.85g Fe is present in 159.7g $Fe_2O_3$

2.78 kg Fe is present in

$$= \frac{159.7g\times 2.78kg}{2\times 55.85g}$$

= 3.97 kg $Fe_2O_3$

As 5 kg ore contains = 3.97 kg $Fe_2O_3$

100 kg ore contains $= \frac{3.97\times 100}{5}$

So, = 79.4 kg $Fe_2O_3$

Thus, percentage of $Fe_2O_3$ in ore = 79.4%.

**16. What should be the weight of $NaNO_3$ to make $50$ mL of an aqueous solution so that it contains $70$mg Na mL$^{-1}$ ?**

**Solution:**

Molecular mass of $NaNO_3 = 23 + 14 + 3\times 16$

$= 85 \text{ g mol}^{-1}$

23mg Na is present in 85mg of $NaNO_3$

70mg Na is present in $= \frac{85\times 70}{23}$

= 258.7 mg $NaNO_3$

1 mL solution contains 258.7 mg $NaNO_3$

50 mL solution contains

$$= \frac{258.7mg\times 50mL}{1mL}$$

= 13935 mg

= 13.935 g.

**17. Suppose the two carbonates are $MCO_3$ and $M_2CO_3$ As M = 13.6% by wt. so $W_{Sample}$ = 2.5g. What will be the % of M?**

**Solution:**

$W_m = \frac{13.6}{100}\times 2.5 = 0.34g$

$n_{CO_2} = \frac{W_{CO_2}}{44} = \frac{1.32}{44} = 0.03$

$W_{CO_3^{2-}} = 0.03\times 60 = 1.8g$

Hence % of $M = \frac{2.5-0.34-1.8}{2.5}\times 100 = 14.4\%$

## IMPORTANT QUESTIONS REGARDING NEET AND AIIMS

In the following questions, two statements Assertion (A) and Reason (R) are given. Mark

*(a) If A and R both are correct and R is the correct explanation of A;*

*(b) If A and R both are correct but R is not the correct explanation of A;*

*(c) A is true but R is false;*

*(d) A is false but R is true;*

*(e) A and R both are false.*

1. (A) : One mole of $NaCl$ contains $6.023 \times 10^{23}$ molecules of sodium chloride.
   (R) : $58.5$ g of $NaCl$ also contains $6.023 \times 10^{23}$ molecules of $NaCl$.

2. (A) : $22.4$ L of $N_2$ at NTP and $5.6$ L $O_2$ at NTP contain equal number of molecules.
   (R) : Under similar conditions of temperature and pressure all gases contain equal number of molecules.

3. (A) : Number of g-molecules of $SO_2Cl_2$ in $13.5$ g of sulphuryl chloride is 0.2.
   (R) : Gram molecules is equal to those molecules which are expressed in gram.

4. (A) : In $CO$ molecule 12 parts by mass of carbon combine with 16 parts by mass of oxygen and in $CO_2$, 12 parts by mass of carbon combine with 32 parts by mass of oxygen.
   (R) : When two elements combine separately with a fixed mass or a third element, then the ratio of their masses in which they do so is either the same or whole number multiple of the ratio in which they combine with each other.

5. In the reaction
   (A) : $2KOH + H_3PO_4 \rightarrow K_2HPO_4 + 2H_2O$
   Equivalent weight of the acid $H_3PO_4$ is equal to $\frac{M}{2}$ [M = mol. wt. of $H_3PO_4$]
   (R) : As $E = \frac{M}{n-\text{ factor}}$

6. (A) : Molarity of a solution and molality of a solution both change with density
   (R) : Density of the solution changes when percentage by mass of solution changes.

7. (A) : The percentage of nitrogen in urea is $46.6\%$.
   (R) : Urea is ionic compound.

8. (A) : $0.28$ g of $N_2$ has equal volume as $0.44$ g of another gas at same conditions of temperature and pressure.
   (R) : Molecular mass of another gas is $44 \text{ g mol}^{-1}$.

9. (A) : $\ln MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$, $MnO_4^-$ acts as oxidizing agent and $Fe^{2+}$ acts as reducing agent.
   (R) : The reactions involving simultaneous loss or gain of electron among the reacting species are called oxidation reduction reactions.

10. (A) : Equivalent mass of a base which contains one mole of replaceable $OH^-$ ion in a molecule.
    (R) : It is the mass of a base which completely reacts with one gram equivalent mass of an acid.

11. (A) : Decomposition of $H_2O_2$ is a disproportionation reaction.
    (R) : $H_2O_2$ molecule simultaneously undergoes both oxidation and reduction.

12. (A) : On increasing temperature, normality decreases like molarity.
    (R) : On increasing temperature, volume increases as a result, normality decreases.

13. (A) : The molality of the solution does not change with change in temperature.
    (R) : The molality is expressed in units of moles per $1000$ g of solvent.