2026
EXAMINATION

CBSE
QUESTION & CONCEPT BANK

Chapter-wise & Topic-wise

CLASS 11

Chapter-wise
CONCEPT MAPS

Important terms, Formulae & Myth Buster
NCERT & SMART SNAPS

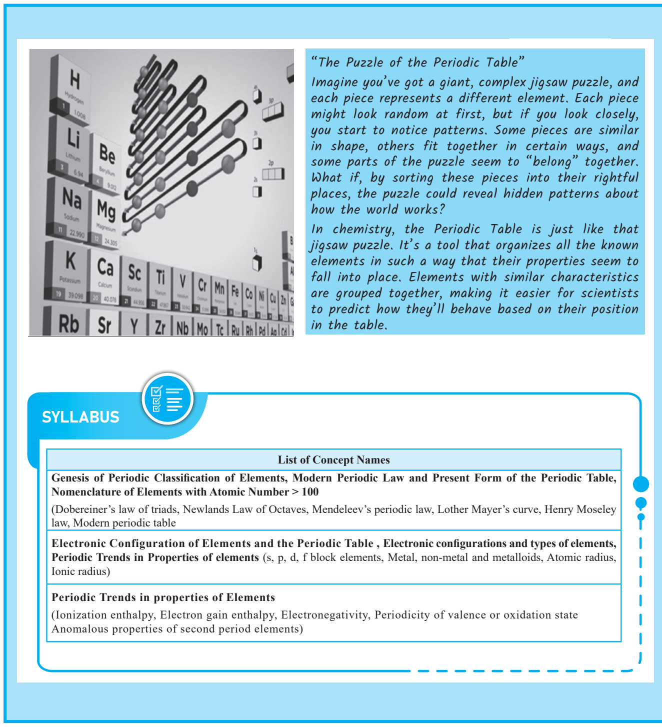
Revision Blue Print & Solved Questions
COMPETENCY FOCUSED

Important Questions with Detailed Explanations
POWER PRACTICE

CHEMISTRY

HOW TO USE THIS BOOK

This book is structured to support your learning journey of preparing for your board exams through a variety of engaging and informative elements. Here's how to make the most of it:



"The Puzzle of the Periodic Table"

Imagine you've got a giant, complex jigsaw puzzle, and each piece represents a different element. Each piece might look random at first, but if you look closely, you start to notice patterns. Some pieces are similar in shape, others fit together in certain ways, and some parts of the puzzle seem to "belong" together. What if, by sorting these pieces into their rightful places, the puzzle could reveal hidden patterns about how the world works?

In chemistry, the Periodic Table is just like that jigsaw puzzle. It's a tool that organizes all the known elements in such a way that their properties seem to fall into place. Elements with similar characteristics are grouped together, making it easier for scientists to predict how they'll behave based on their position in the table.

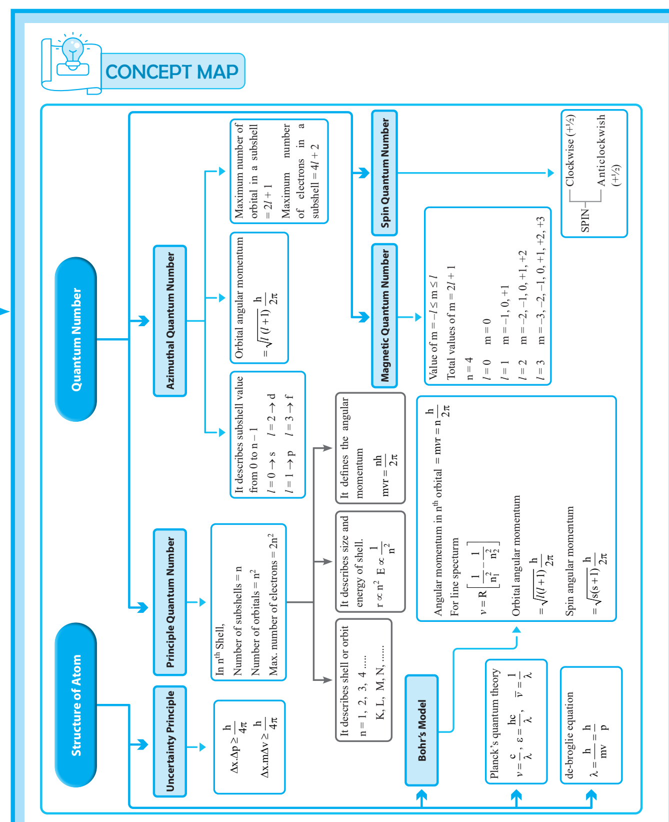
SYLLABUS

List of Concept Names
Genesis of Periodic Classification of Elements, Modern Periodic Law and Present Form of the Periodic Table, Nomenclature of Elements with Atomic Number > 100 (Dobereiner's law of triads, Newlands Law of Octaves, Mendeleev's periodic law, Lother Mayer's curve, Henry Moseley law, Modern periodic table
Electronic Configuration of Elements and the Periodic Table , Electronic configurations and types of elements, Periodic Trends in Properties of elements (s, p, d, f block elements, Metal, non-metal and metalloids, Atomic radius, Ionic radius)
Periodic Trends in properties of Elements (Ionization enthalpy, Electron gain enthalpy, Electronegativity, Periodicity of valence or oxidation state Anomalous properties of second period elements)

Preview

At the start of every chapter, you'll find a thoughtfully chosen image and a quote that captures the main idea and motivation of the topic. This approach aims to get your interest and give you a glimpse of the theme ahead.

Before diving into the details, we outline the syllabus. This helps you prioritize your study focus based on the significance of each section.



Concept Map

The concept map appears to be a comprehensive study aid that outlines key concepts in a structured format, featuring definitions, diagrams, and processes. For a student, it would serve as a visual summary, making complex ideas more accessible and aiding in revision and understanding of concept for their curriculum.

1 INTRODUCTION, GENESIS OF PERIODIC CLASSIFICATION OF ELEMENTS, MODERN PERIODIC LAW AND PRESENT FORM OF THE PERIODIC TABLE, NOMENCLATURE OF ELEMENTS WITH ATOMIC NUMBERS > 100

NCERT Definitions *(Commonly asked in 1 mark)*

- **Periodic Table:** A tabular arrangement of elements in rows (periods) and columns (groups) based on their atomic number and recurring chemical properties.
- **Dobereiner's Triads:** A group of three elements where the atomic mass of the middle element is the average of the atomic masses of the other two elements.
- **Newlands' Law of Octaves:** This law states that when elements are arranged in order of increasing atomic masses, every eighth element had properties similar to the first element akin to the eighth note in a musical scale.
- **Mendeleev's Periodic Law:** Dmitri Mendeleev's law states that the properties of elements are periodic functions of their atomic masses.
- **Modern Periodic Law:** This law states that the properties of elements are periodic functions of their atomic numbers, not atomic masses.
- **Moseley's Law:** Moseley's Law states that the frequency of X-ray spectra emitted by an element is related to its atomic number rather than its atomic mass, which helped rearrange the periodic table in terms of atomic number.
- **Groups:** The vertical columns of the periodic table is known as groups.
- **Periods:** The horizontal rows of the periodic tableis known as periods.

Important Facts

01 The periodic table currently contains 118 identified and named chemical elements.

02 Lothar Meyer had developed a table of the elements that closely resembles the Modern Periodic Table.

03 Dmitri Mendeleev predicted the existence of several elements that were not yet discovered when he published his periodic table.

04 Mendeleev ignored the order of atomic weights, thinking that the atomic measurements might be incorrect, and placed the elements with similar properties together.

05 Lothar Meyer arranged elements in order of increasing atomic weight and atomic volume.

06 Newlands' law of octaves was valid only up to calcium.

NCERT Definitions: *It simplifies complex topics into brief, easy-to-understand explanations.*

Important Facts: *Quick, bullet point facts that are crucial for exams.*

Important Concepts:
Familiarizing with key concepts in advance helps prepare cognitive framework for processing and integrating new information. By highlighting important concepts upfront, students are better equipped to identify connections and relationships between various ideas presented in the chapter.
Important Formulas:
Introducing important formulas upfront brings clarity to the chapter's objectives, guiding students' focus towards essential mathematical principles that will be explored further.

Difference Between: *Side-by-side comparisons to help distinguish similar concepts.*

Important Concepts

- **First Law of Thermodynamics:** It is a fundamental principle of energy conservation in thermodynamics. It states that:"Energy can neither be created nor destroyed; it can only be transferred or transformed from one form to another."
  Mathtematically, the first law is expressed as:
  $\Delta U = q + w$
  where:
  ΔU = Change in internal energy of the system.
  q = Heat added to the system.
  w = Work done on or by the system.

Important Formulae

First law of Thermodynamics	$\Delta U = q + w$
Work done	$w = -\int_{v_i}^{v_f} p_{ex} dV$
Work done in irreversible isothermal expansion	$w_{irr} = -P_{ex}(V_f - V_i)$
Work done in reversible isothermal expansion	$w_{rev} = -2.303nRT \log\left(\frac{V_f}{V_i}\right)$
Internal energy at constant volume	$\Delta U = q_v$
Heat for isothermal irreversible change	$q = -w = p_{ex}(V_f - V_i)$
Heat for isothermal reversible change	$q = 2.303nRT \log\frac{V_f}{V_i}$

Difference Between

State function vs. Path function

State function	Path function
Independent of the path taken.	Dependent on the path taken.
Depends only upon the state of the system.	Depends upon how that state of the system has been achieved.
All paths result in the same value.	Different paths may result in different values.
E.g.: Internal energy (U), enthalpy (H), entropy (S), temperature (T), volume (V), pressure (P), etc.	E.g.: Heat (q), work (w), etc.

Reversible Process vs. Irreversible Process

Feature	Reversible Process	Irreversible Process
Definition	A process that occurs infinitely slowly and can be reversed without any net change in the system or surroundings.	A process that occurs spontaneously and cannot be reversed without leaving changes in the system or surroundings.
Speed	Extremely slow (infinitely small steps).	Fast and spontaneous.
Equilibrium	System remains in equilibrium at every stage.	System is not in equilibrium during the process.
Work Done	Maximum work is obtained from the system.	Less work is obtained due to energy dissipation.
Energy Loss	No energy loss due to friction, turbulence, or heat dissipation.	Energy is lost in the form of heat, sound, or friction.

Real Life Application Based Questions

1. **Alkali metals such as sodium and potassium are essential for biological functions but can be highly reactive when exposed to air or water. How does their position in the periodic table explain both their biological significance and their extreme reactivity?**

Ans. Alkali metals belong to Group 1 and have a single electron in their outermost shell, making them highly reactive as they easily lose this electron. This property makes them crucial in biological systems (e.g., sodium and potassium regulate nerve impulses). However, their high reactivity with water and air requires careful handling to prevent hazardous reactions.

2. **Many modern medical devices, such as MRI machines, use superconducting magnets that contain elements like niobium and titanium. How does the periodic classification of elements help scientists choose the right elements for such applications?**

Ans. The periodic classification helps scientists understand the properties of elements based on their position in the periodic table. Elements in specific groups exhibit predictable conductivity, reactivity, and resistance to temperature variations. Niobium and titanium, being transition metals, possess high conductivity and durability, making them ideal for superconducting magnets used in MRI machines.

Myth Buster

- **Myth: The periodic table has always been based on atomic numbers since its creation.**
  Fact: Initially, elements were arranged by atomic mass (Mendeleev's Periodic Table). It was only after Henry Moseley's discovery in 1913 that the modern periodic law was established, organizing elements by atomic number instead.
- **Myth: The periodic table is fully complete and will not change in the future.**
  Fact: The periodic table continues to evolve. New superheavy elements (Z > 118) are being synthesized, and future discoveries might lead to a new period or even a redesign of element organization based on quantum mechanics.
- **Myth: The periodic table only includes elements found in nature.**
  Fact: Many elements in the periodic table are synthetic and do not exist naturally. For example, Technetium (Tc), Neptunium (Np), and all elements beyond uranium (Z > 92) are artificially created.

Mnemonics

- **S-block elements**

H	Li	Na	K	Rb	Cs	Fr
↓	↓	↓	↓	↓	↓	↓
Hey	Li	Na	Kare	Rab	Se	Fariyad

- **P-block elements**

B	Al	Ga	In	Tl
↓	↓	↓	↓	↓
Baingan	Aalu	Gaajar	in	Thela
C	Si	Ge	Sn	Pb
↓	↓	↓	↓	↓
College	Student	Get	Some	Problem

Real-Life Application Based Questions: *Exercises that connect theory with practical scenarios. It will enhance your understanding and relevance of concepts.*

Myth Buster: *Clear up common misconceptions to ensure your understanding is accurate.*

Mnemonics: *Memory aids to help you retain and recall information.*

COMPETENCY BASED SOLVED EXAMPLES

Multiple Choice Questions (1 M)

1. **All the elements in a group in the periodic table have the same;** *(Re)*
   (*a*) atomic number
   (*b*) electronic configuration
   (*c*) atomic weight
   (*d*) number of electrons in the valence shell
2. **IUPAC name of element having atomic number 108 is;** *(Un)*
   (*a*) Unniloctium (*b*) Ununoctium
   (*c*) Nilniloctinium (*d*) Ununoctinium
3. **Which of the following sets of elements follows Newland's octave rule?** *(Re)*
   (*a*) Be, Mg, Ca (*b*) Na, K, Rb
   (*c*) F, Cl, Br (*d*) B, Al, Ga
4. **Atomic weight of Cl is 35.5 and I is 127 . What will be the atomic weight of Br , as per Dobereiners triad rule;** *(Ap)*
   (*a*) 81.25 (*b*) 85
   (*c*) 95 (*d*) 162
5. **The places that were left empty by Mendeleev's were, for;** *(Re)*
   (*a*) Aluminium and silicon
   (*b*) Galium and germinium
   (*c*) Arsenic and antimony
   (*d*) Molybdenum and tungsten

Assertion and Reason (1 M)

Direction: **The following questions consist of two statements – Assertion (A) and Reason (R). Answer these questions by selecting the appropriate option given below:**

(*a*) Both A and R are true, and R is the correct explanation of A.
(*b*) Both A and R are true, 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.

1. **Assertion (A):** Mendeleev's arranged elements in horizontal rows and vertical columns.
   Reason (R): Mendeleev's ignored the order of atomic weight thinking that the atomic measurements might be incorrect. *(Re)*
2. **Assertion (A):** In Modern periodic table, position of isotopes of an element is different.
   Reason (R): Alkali metals as well as hydrogen have one electron in their valence shell. Therefore, hydrogen is placed in the group of alkali metals *(Un)*
3. **Assertion (A):** Döbereiner's Law of Triads could not be applied to all elements known at that time.
   Reason (R): Newlands' Law of Octaves stated that the properties of every eighth element were similar to the first when elements were arranged in increasing order of atomic mass. *(Re)*
4. **Assertion (A):** Newlands' Law of Octaves was rejected because it could not be applied to elements beyond calcium.
   Reason (R): The discovery of noble gases completely fit into Newlands' Law of Octaves without any modifications. *(Un)*
5. **Assertion (A):** Mendeleev's periodic table grouped elements with similar properties together.
   Reason (R): Elements in the same group of Mendeleev's periodic table have the same atomic mass. *(Un)*

Subjective Questions

Very Short Answer Type Questions (2 M)

1. **Elements A, B, C, D and E have the following electronic configurations:**
   A: $1s^2\ 2s^2\ 2p^1$
   B: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$
   C: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
   D: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
   E: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
   Which among these will belong to the same group in the periodic table ? *(Ap)*

Ans. Out of these, elements A and B will belong to the same group of the periodic table because they have same outer electronic configuration, ns^2np^1. *(2 M)*

2. **What would be the IUPAC name and symbol for the element with atomic number 120?** *(Un)*

Ans. The root for 1,2 and 0 are un, bi and nil respectively. The IUPAC name of the element is unbinilium. *(1 M)*
The symbol of the element is ubn. *(1 M)*

3. **What is the basic theme of organisation in the periodic table?** *(Un)*

Ans. The basic theme of organisation of elements in the periodic table is to classify the elements in periods and groups according to their properties. This arrangement makes the study of elements and their compounds simple and systematic. In the periodic table, elements with similar properties are placed in the same group. *(2 M)*

Solved Examples

For each topic, solved examples are provided including tagging of Competencies, PYQs, CBSE SQPs etc that exemplify how to approach and solve questions. This section is designed to reinforce your learning and improve problemsolving skills.

At the end of each chapter, you'll find additional exercises intended to test your grasp of the material. These are great for revision and to prepare for exams.

Answer Key and Explanations including, Nailing the right answer and Key takeaway to know how to write the ideal answer.

Answer Key

MISCELLANEOUS EXERCISE

Multiple Choice Questions (1 M)

1. The valence shell electronic configuration of transition elements is; *(Un)*
   (*a*) ns^1 (*b*) ns^2np^5
   (*c*) $ns^{0-2}\ (n-1)d^{1-10}$ (*d*) $ns^2\ (n-1)d^{10}$
2. Among the elements Ca, Mg, P and Cl the order of increasing atomic radii is; *(Un)*
   (*a*) Mg < Ca < Cl < P (*b*) Cl < P < Mg < Ca
   (*c*) P < Cl < Ca < Mg (*d*) Ca < Mg < P < Cl
3. Amongst the elements with following electronic configuration, which one of them may have the highest ionization energy? *(Un)*
   (*a*) $Ne[3s^23p^2]$ (*b*) $Ar[3d^{10}4s^24p^2]$
   (*c*) $Ne[3s^23p^1]$ (*d*) $Ne[3s^23p^3]$
4. Which period in the periodic table contains the maximum number of elements? *(Un)*
   (*a*) 1st period (*b*) 3rd period
   (*c*) 2nd period (*d*) 6th period
5. Atomic weight of P is 31 and Sb is 120. What will be the atomic weight of As, as per Dobereiners triad rule? *(Ap)*
   (*a*) 151 (*b*) 75.5
   (*c*) 89.5 (*d*) 60
6. Which of the following statement is wrong? *(Re)*
   (*a*) 2nd period contain 8 elements.
   (*b*) 3rd period contains 18 elements.
   (*c*) 1st period contains two non-metals.
   (*d*) In p–block, metal, nonmetal and metalloids are present.

ANSWER KEYS

Multiple Choice Questions

1. (*b*) 2. (*a*) 3. (*a*) 4. (*b*) 5. (*c*) 6. (*d*) 7. (*d*) 8. (*a*) 9. (*c*) 10. (*c*)
11. (*a*) 12. (*a*) 13. (*a*) 14. (*a*) 15. (*a*) 16. (*c*) 17. (*b*) 18. (*b*) 19. (*d*) 20. (*b*)

Assertion and Reason

1. (*b*) 2. (*c*) 3. (*b*) 4. (*b*) 5. (*c*) 6. (*a*) 7. (*a*) 8. (*a*) 9. (*a*) 10. (*d*)

HINTS & EXPLANATIONS

Multiple Choice Questions

1. (*b*) ΔH (change in enthalpy) represents the heat exchanged at constant pressure.
2. (*a*) In an adiabatic process, there is no heat exchange between the system and surroundings: $q = 0, \Delta U = -W$.
3. (*a*) $W = -P\Delta V$
   where: $\Delta V = V_2 - V_1$
   $\Delta V = (1 \times 10^{-2}) - (1 \times 10^{-3})$
   $\Delta V = 9 \times 10^{-3}\ m^3$
   $W = -(1 \times 10^4) \times (9 \times 10^{-3})$
   $W = -90$ J
4. (*b*) $\Delta G = \Delta H - T\Delta S$
5. (*c*) $\Delta G < 0$ (negative) → The reaction is spontaneous.

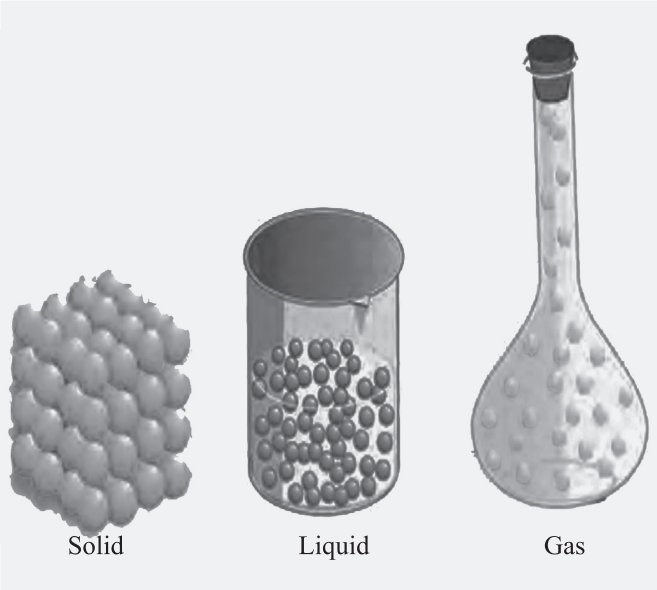
When we reverse the reaction, the sign of the enthalpy change also reverses: $\Delta H = +1304.4$kJ
Since there are 3 moles of H_2, we have:
Number of H-H bonds = 3
Now, we can calculate the bond dissociation energy:
Bond dissociation energy
$= \frac{1304.4kJ}{3} = 434.8kJ$

8. (*a*) For $\Delta H = \Delta U$, Δn_g must be zero, meaning the number of gaseous moles does not change.
   $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$
   $\Delta ng = (2) - (1 + 1) = 0$
9. (*c*) $\Delta U = q + w$
   Work done on the system: +4 kJ
   Heat given out by the system: -1 kJ

CONTENTS

SOME BASIC CONCEPTS OF CHEMISTRY

1

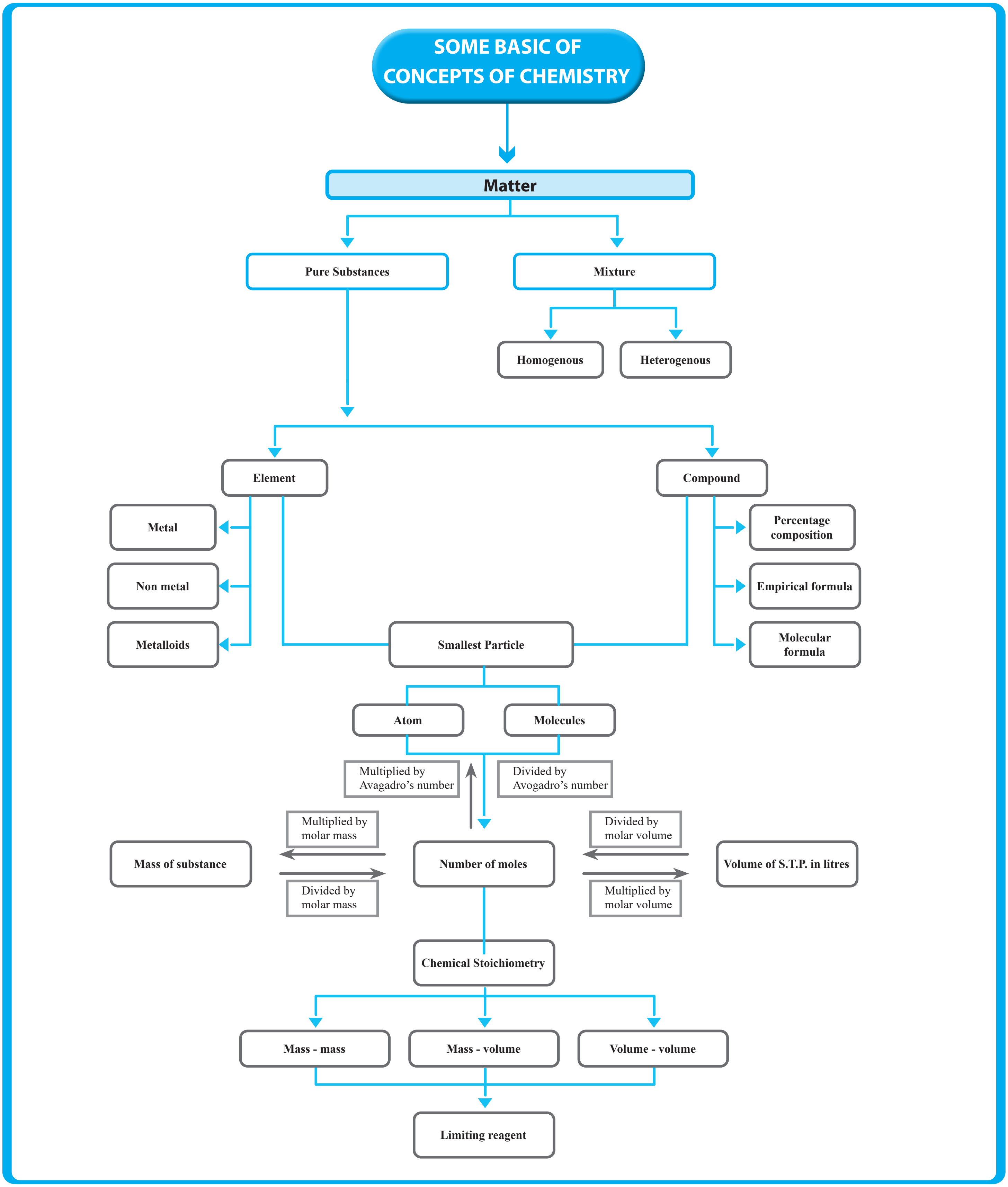


"Some Basic Concepts of Chemistry" serves as the first step toward understanding the vast and intricate field of chemistry. This chapter provides a thorough introduction to the essential concepts and principles that form the foundation of the subject. This chapter lays the foundation for understanding the language of chemistry, focusing on key concepts such as the nature of matter, the laws of chemical combination, the mole concept and the importance of atomic and molecular masses. It emphasizes the significance of measurement and units in chemistry, highlighting the need for precision and accuracy in scientific calculations. The chapter also explores how chemical reactions are quantified, providing a basis for further study in the subject"

SYLLABUS

List of Concept Names
Matter (Nature and Measurement), Uncertainity in Measurement, Laws of Chemical Combination, Atomic and Molecular Masses.
Mole Concept, Stoichiometry Calculations, Concentration Terms.

CONCEPT MAP


SOME BASIC OF CONCEPTS OF CHEMISTRY
Matter
Pure Substances
Mixture
Homogenous
Heterogenous
Element
Compound
Metal
Non metal
Metalloids
Percentage composition
Empirical formula
Molecular formula
Smallest Particle
Atom
Molecules
Multiplied by Avagadro's number
Divided by Avogadro's number
Multiplied by molar mass
Divided by molar mass
Mass of substance
Number of moles
Divided by molar volume
Multiplied by molar volume
Volume of S.T.P. in litres
Chemical Stoichiometry
Mass - mass
Mass - volume
Volume - volume
Limiting reagent

1 MATTER, UNCERTAINTY IN MEASUREMENT, LAWS OF CHEMICAL COMBINATION, ATOMIC AND MOLECULAR MASSES

NCERT Definitions *(Commonly asked in 1 mark)*

- **Matter:** Matter is anything that has mass and occupies space.
- **Significant Figures:** The digits in a properly recorded measurement or the total number of figures in a number including the last digit whose value is uncertain are called significant figures, e.g., 180.00 has five significant figures.
- **Law of Conservation of Mass:** Mass can neither be created nor destroyed.
- **Law of Definite Proportions:** A given compound always contains exactly the same proportion of elements by mass.
- **Law of Multiple Proportions:** When two elements combine with each other 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 the gases are at same temperature and pressure.
- **Avogadro's Law:** At the same temperature and pressure, equal volume of all gases should contain equal number of molecules.
- **Weight:** Weight is the force exerted by gravity on an object. The weight of a substance may vary from one place to another due to change in gravity.
- **Precision:** Precision refers to the closeness of various measurements of same quantity.
- **Accuracy:** Accuracy is the agreement of a particular value to the true value of the results.
- **Atomic Mass:** One atomic mass unit is defined as a mass exactly equal to one-twelfth of the mass of one carbon-12 atom.
- **Average Atomic Mass:** The average atomic mass of an element refers to the atomic masses of the isotopes of the element, taking into account the abundances of the element's isotopes.
- **Gram Atomic Mass:** Gram atomic mass is the mass, in grams, of one mole of atoms in a monoatomic chemical element.
- **Molecular Mass:** Molecular mass refers to the mass of a molecule. It is calculated as the sum of atomic masses of all the atoms in a molecule of the substance.

Important Facts

01 The mass of a substance is constant, whereas, it's weight may vary from one place to another due to change in gravity.

02 Gay lussac's law is not obeyed if any reactant and product is a solid or liquid because the volume occupied by them is extremely small as compared to gas.

03 Law of conservation of mass is not applicable to nuclear reactions.

04 SI unit of volume = m^3. ($1\ L = 1 dm^3 = 10^3\ cm^3 = 10^{-3} m^3$)

05 Zeroes at the end or right of a number are significant, provided they are on the right side of the decimal point.

06 Standard temperature pressure (STP): 273.15 K temperature and 1 amospheric pressure.

Important Concepts

Exponential Number or Scientific Notation.

In Chemistry, we come across very large and very small numbers. It is very tedious to write down such numbers in the ordinary way. For example, it is not convenient to write Avogadro constant as 602, 213, 700,000,000,000,000,000. These numbers are usually expressed in a simple way known as exponential form or scientific notation. For example, the number 246.38 may be expressed as:

$$246.38 = 2.4638 \times 10 \times 10, \text{ or } = 2.4638 \times 10^2$$

In general, in scientific notation, a number may be expressed as: $N \times 10^n$, where n is an exponent having positive or negative values and N is a single non-zero digit and lies between 1 to 9.999.... N is called **digit term** and n is called **exponent**.

number with a single non-zero digit → N

number of places decimal point was moved → n

$N \times 10^n$

+ moved left
– moved right

Law of Multiple Proportions

This law was proposed by Dalton in 1803. Law of multiple proportions states that "When two elements combine to form more than one compound, then the masses of one of the elements which combine with a fixed mass of the other element are in a simple whole number ratio."

Oxide	Number of parts by weight of nitrogen	Number of parts by weight of oxygen	Fixed weight of nitrogen (28 parts)
N_2O_3	28	48	28
N_2O_4	28	64	28
N_2O_5	28	80	28

Here, the masses of oxygen (i.e., 48 g, 64 g and 80 g) which combine with a fixed mass of nitrogen (28g) bear a simple ratio, i.e., 3 : 4 : 5.

Important Formulas

- **Density** $= \dfrac{\text{Mass}}{\text{Volume}}$
- $°F = \dfrac{9}{5}(°C) + 32$
- **Kelvin (K)** $K = °C + 273.15$
- **Average Atomic Mass** $= \dfrac{\sum(\text{Mass of Isotopes} \times \%\text{ abundance})}{100}$

Difference Between

Compounds vs. Mixtures

Property	Compounds	Mixtures
Definition	In a compound, two are more elements are combined chemically.	In a mixture, two or more elements or compounds just mix together.
Composition	The compounds contains two or more elements in a fixed ratio by mass. Its composition is always fixed.	The components of a mixture may be present in any ratio. Its composition is variable.
Formula	A compound has a definite formula.	A mixture does not have a definite formula.
Nature	A compound is always homogeneous i.e., has the same composition throughout.	A mixture may be homogeneous or heterogeneous.
Formation	A chemical reaction takes place and therefore, the formation of a compound takes place with absorption or evolution of energy.	No chemical reaction takes place and therefore, the formation of mixture is not accompained by any energy change.
Properties	The properties of a compound are entirely different from those of its constituents.	A mixture shows the properties of its constituents.
Separation	A compound cannot be separated into its constituents by ordinary physical methods. These can be separated by chemical or electrochemical reactions.	A mixture can be separated into its constituents by physical methods (like filtration, evaporation, distillation, sublimation, mechanical separation, etc.)
Physical Properties	A compound has a fixed melting point, boilling point, etc.	A mixture does not have fixed melting point, boiling point, etc.

Homogeneous vs. Heterogeneous mixture

Features	Homogeneous Mixture	Heterogeneous mixture
Composition	Uniform throughout the solution.	Not uniform the solution.
Phases	Single phase.	Two or more phases.
Separation	Difficult	Easier (by filtration, decantation, etc.)
Examples	Sugar solution, air, alloys	Sand-water, oil-water, salad

Atomic Mass vs. Molecular Mass

Property	Atomic Mass	Molecular Mass
Definition	The mass of a single atom of an element.	The sum of the atomic masses of all atoms in a molecule.
Unit	Atomic Mass Unit (amu or u)	Atomic Mass Unit (amu or u)
Formula	Approximate atomic mass = Number of protons + Number of neutrons	Molecular mass = Sum of the atomic masses of all atoms in a molecule
Example	Atomic mass of Oxygen (O) = 16 u	Molecular mass of Oxygen gas (O_2) = 16 × 2 = 32 u
Applied to	Individual atoms of elements.	Covalent molecules, compounds and diatomic elements.

Real Life Application Based Questions

1. **Why is it important to measure the mass of ingredients while cooking?**

Ans. In cooking, precise measurement of ingredients ensures the right taste, texture, and consistency of the food. The concept of mass and mole concept helps in maintaining proper proportions, similar to how chemical reactions require exact reactant amounts for the desired product.

2. **Why do LPG cylinders mention "Net Weight" instead of "Volume"?**

Ans. Liquefied Petroleum Gas (LPG) is stored in a compressed liquid state, and its amount is measured in mass (kg) instead of volume, as volume varies with temperature while mass remains constant.

COMPETENCY BASED SOLVED EXAMPLES

Multiple Choice Questions

(1 M)

1. **Which of the following is an example of a heterogeneous substance?** *(Ap)*
   (*a*) Bottled water
   (*b*) Table salt
   (*c*) Pieces of copper
   (*d*) Candle

2. **A measured temperature on Fahrenheit scale is 200 °F. What will this reading be on Celsius scale?** *(Ap)*
   (*a*) 40° C (*b*) 94° C
   (*c*) 93.3° C (*d*) 30° C

3. **Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data, mark the correct option out of the following statements.**

Students	Readings (i)	Readings (ii)
A	**3.01**	**2.99**
B	**3.05**	**2.95**

(An) **(CBSE CFPQ,2024)**

   (*a*) Results of both the students are neither accurate nor precise.
   (*b*) Results of student A are both precise and accurate.
   (*c*) Results of student B are neither precise nor accurate.
   (*d*) Results of student B are both precise and accurate.

4. **Which of the following is the best example of law of conservation of mass?** *(Un)* **(NCERT Exemplar)**
   (*a*) 12 g of carbon combines with 32 g of oxygen to form 44 g of CO_2.
   (*b*) When 12 g of carbon is heated in a vacuum there is no change in mass.
   (*c*) A sample of air increases in volume when heated at constant pressure but its mass remains unaltered.
   (*d*) The weight of a piece of platinum is the same before and after heating in air.

5. **Irrespective of the source, pure sample, of water always yields 88.89% mass of oxygen and 11.11% mass of hydrogen. This is explained by the law of:** *(Un)*
   (*a*) conservation of mass (*b*) multiple proportions
   (*c*) constant proportions (*d*) constant volume

6. **1 amu is equal to:** *(Re)*
   (*a*) $\frac{1}{14}$ of O-16 (*b*) $\frac{1}{12}$ of C-12
   (*c*) 1 g of H_2 (*d*) 1.66×10^{-23} kg

Assertion and Reason

(1 M)

Direction: The following questions consist of two statements – Assertion (A) and Reason (R). Answer these questions by selecting the appropriate option given below:

(*a*) Both A and R are true and R is the correct explanation of A.
(*b*) Both A and R are true 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.

1. **Assertion (A):** Significant figures for 0.200 is 3 whereas for 200 it is 1.
   Reason (R): Zero at the end or right of a number are significant provided they are not on the right side of the decimal point. *(Re)*

2. **Assertion (A):** One atomic mass unit is defined as one twelfth of the mass of one carbon - 12 atom.
   Reason (R): Carbon-12 isotope is the most abundant isotope of carbon and has been chosen as standard. *(Re)*

3. **Assertion (A):** Volume of a gas is inversely proportional to the number of moles of gas.
   Reason (R): The ratio by volume of gaseous reactants and products is in agreement with their mole ratio. *(Un)*

Subjective Questions

Very Short Answer Type Questions

(2 M)

1. **Express the following upto three significant places:**
   (*a*) the height of a man, 5 feet 9 inches in centimetres (1 inch = 2.54 cm)
   (*b*) one millionth of one.
   (*c*) four thousand
   (*d*) decimal equivalent of 2/3 *(Ap)*

Ans. (*a*) 5 feet 9 inches = 69 inch = 69 × 2.54 = 175.26 cm.
With three significant figures answer is 175 cm. *(½ M)*

(*b*) One millionth of one is $\frac{1}{10^6}$. It can be written as 10^{-6} and upto 3 significant figures as 1.00×10^{-6}. *(½ M)*

(*c*) Four thousand = 4.00×10^3 *(½ M)*

(*d*) $\frac{2}{3} = 0.667$. *(½ M)*

2. **If 6.3 g of $NaHCO_3$ are added to 15.0 g of CH_3COOH solution, the residue is found to weigh 18.0 g. What is the mass of CO_2 released in the reaction ?** *(Un)*

Ans. $$\underset{6.3g}{NaHCO_3}+\underset{15.0g}{CH_3COOH}\longrightarrow \underbrace{CH_3COONa+H_2O}_{\substack{\text{Residue}\\18.0g}}+\underset{xg}{CO_2}$$

Sum of the mass of reactants = Mass of $NaHCO_3$ + Mass of CH_3COOH

= 6.3 + 15.0 = 21.3 g *(1 M)*

Sum of the mass of products = $\begin{matrix}\text{Mass of}\\\text{residue}\end{matrix} + \begin{matrix}\text{Mass of}\\CO_2\end{matrix}$

According to law of conservation of mass

Mass of reactants = Mass of products

$21.3 = 18.0 + x$

or $x = 21.3 - 18.0 = 3.3$ g

∴ Mass of CO_2 released = 3.3 g *(1 M)*

3. **Boron occurs in nature in the form of two isotopes having atomic mass 10 u and 11 u. What are the percentage abundances of two isotopes in a sample of boron having average atomic mass 10.8 u?** *(Ap)*

Ans. Let the % abundance of ^{10}B isotope = x %

Then, % abundance of ^{11}B isotope = (100 – x)%

The average atomic mass = $\dfrac{(x\times10)+[(100-x)\times11]}{100}$

$$\therefore \quad \frac{(x\times10)+[(100-x)\times11]}{100}=10.8 \quad (1\text{ M})$$

or $10x + 1100 - 11x = 10.8 \times 100$

$-x = -1100 + 1080$

or $x = 20$

Thus, percentage abundance; ^{10}B = 20 %, ^{11}B = 80% *(1 M)*

4. **2.0 g of a metal burnt in oxygen gave 3.2 g of its oxide. 1.42 g of the same metal heated in steam gave 2.27 g of its oxide. Which law is shown by this data?** *(An)*

Ans. In the first compound,

3.2 g of metal oxide contained 2.0 g of metal

100 g of metal oxide contained metal

$$=\frac{2.0}{3.2}\times100=62.5\text{ g}$$

∴ % Metal in first compound = 62.5% *(1 M)*

In the second compound,

2.27 g of metal oxide contained metal = 1.42 g

100 g of metal oxide contained metal

$$=\frac{1.42}{2.27}\times100=62.55\text{ g}$$

∴ % Metal in second compound = 62.55%

Thus, the percentage of metal in metal oxide obtained from two experiments is nearly same. Hence, the above data illustrate the law of constant composition. *(1 M)*

5. **Suppose a length had been reported to be 31.24 cm. What is the minimum uncertainty in this measurement? What is the difference between 0.006 g and 6.00×10^{-3} g?** *(Re)*

Ans. The minimum uncertainty in this measurement is ± 0.01 cm. *(1 M)*

0.006 g contains one significant digit while 6.00×10^{-3} g contains 3 significant digits. *(1 M)*

Short Answer Type Questions (3 M)

1. **Express the results of the following calculations to the appropriate number of significant figures:**

(i) $\dfrac{3.24\times0.08666}{5.006}$ (ii) $\dfrac{(1.36\times10^{-4})(0.5)}{2.6}$

(iii) $2.64 \times 10^3 + 3.27 \times 10^2$ *(Ap)* (NCERT)

Ans. (i) $\dfrac{3.24\times0.08666}{5.006}=0.05608=0.0561$

In the calculation, the digit with least number of significant figures is 3.24 i.e., 3, therefore, the result should have 3 significant figures. Therefore, the correct answer is 0.0561. The number after 0 is 8 and therefore it is rounded off to 1. *(1 M)*

(ii) $\dfrac{(1.36\times10^{-4})(0.5)}{2.6}=0.2615\times10^{-4}$

$= 0.3 \times 10^{-4}$.

The answer should have one significant figure because the digit with least number of significant figure is 0.5 which has one significant figure. *(1 M)*

(iii) $2.64 \times 10^3 + 3.27 \times 10^2$

or $2.64 \times 10^3 + 0.327 \times 10^3 = 2.967 \times 10^3 = 2.97 \times 10^3$

since 2.64 has two digits after decimal place, the answer should be rounded off to two decimal places. *(1 M)*

2. **Two oxides of a metal contain 27.6% and 30% of oxygen respectively. If the formula of the first compound is M_3O_4 find the formula of the second compound.**

(Un) **(NCERT Intext)**

Ans.

First oxide	Second oxide
Oxygen = 27.6%	Oxygen = 30%
Metal = 72.4%	Metal = 70%

Formula of first oxide = M_3O_4

Suppose the atomic weight of metal = x

Percentage of metal in the compound M_3O_4

$$\frac{3x}{3x+64}\times100$$

$$\therefore \quad \frac{3x}{3x+64}\times100=72.4$$

or $300x = 217.2x + 4633.6$

or $82.8x = 4633.6$ or $x = 56$ *(1½ M)*

Now in the second oxide, metal and oxygen are 70% and 30%. Therefore, their atomic ratio will be

$$M : O$$

$$\frac{70}{56} : \frac{30}{16}$$

$$1.25 : 1.875$$

or $1 : 1.5$

or $2 : 3$

Therefore, formula of the compound = M_2O_3. *(1½ M)*

3. Perform the following calculations and express the result to proper number of significant figures: *(Ap)*

(i) $144.3\ m^2 + (2.54\ m \times 8.4\ m)$

(ii) $(4.05 \times 10^2\ mL) – (0.0225 \times 10^2\ mL)$

(iii) $(3.50 \times 10^2\ cm)\ (4.00 \times 10^6\ cm)$

Ans. (i) $(144.3\ m^2) + (2.54\ m \times 8.4\ m)$

$2.54\ m \times 8.4\ m = 21.336\ m^2$ or $21\ m^2$ (upto 2 significant figures)

$$\begin{array}{r} 144.3\ m^2 \\ \underline{21.0\ m^2} \\ 165.3\ m^2 \text{ or } 165\ m^2 \end{array}$$ *(1 M)*

(ii) $(4.05 \times 10^2\ mL) – (0.0225 \times 10^2\ mL)$

$$\begin{array}{r} 4.05 \times 10^2\ mL \\ \underline{- 0.0225 \times 10^2\ mL} \\ 4.0275 \times 10^2\ mL \text{ or } 4.03 \times 10^2\ mL \end{array}$$

(upto second decimal place as in 4.05) *(1 M)*

(iii) $(3.50 \times 10^2\ cm) \times (4.00 \times 10^6\ cm)$

$= 14.0 \times 10^8\ cm^2$ (upto 3 significant figures) *(1 M)*

Long Answer Type Questions (5 M)

1. Express each of the following in S I units :

(i) 93 million miles (this is the distance between the earth and the sun). (1 mile = 1.60 km)

(ii) 5 feet 2 inches (this is the average height of an Indian female).

(iii) 100 miles per hour (this is the typical speed of Rajdhani Express).

(iv) 0.74 Å (this is the bond length of hydrogen molecule). ($1\ Å = 10^{–10}\ m$)

(v) 46°C (this is the peak summer temperature in Delhi). *(Un)*

Ans. (i) The S I unit of distance is metre (m)

1 mile = 1.60 kilometre = 1.60 × 1000 m

$$\text{Unit factor} = \frac{1.60 \times 1000\text{m}}{1\text{Mile}} = \frac{1.6 \times 10^3\text{ m}}{1\text{Mile}}$$

$$93\text{ million miles} = \frac{93 \times 10^6\ \cancel{\text{miles}} \times 1.6 \times 10^3\text{ m}}{1\ \cancel{\text{mile}}}$$

$= 93 \times 1.6 \times 10^9$ m

$= 148.8 \times 10^9\ m = 1.49 \times 10^{11}$ m. *(1 M)*

(ii) 5 feet 2 inches = 62 inches

1 inch = $2.54 \times 10^{–2}$ m

$$\text{Unit factor} = \frac{2.54 \times 10^{-2}\text{ m}}{1\text{inch}}$$

$$62\text{ inches} = \frac{62\ \cancel{\text{inches}} \times 2.54 \times 10^{-2}\text{ m}}{1\ \cancel{\text{inch}}}$$

$= 62 \times 2.54 \times 10^{–2}$ m

$= 157.48 \times 10^{–2}$ m = 1.57 m. *(1 M)*

(iii) 1 mile = 1.60 km = 1.60×10^3 m

$$\text{Unit factor} = \frac{1.60 \times 10^3\text{ m}}{1\text{mile}}$$

1 hr = 60 × 60 s = 3.6×10^3 s

$$\text{Unit factor} = \frac{3.6 \times 10^3\text{ s}}{1\text{hr}}$$

$$\text{Speed} = \frac{100\text{ miles}}{\text{hr}}$$

$$= \frac{100\ \cancel{\text{miles}}}{\cancel{\text{hr}}} \times \frac{1.60 \times 10^3\text{ m}}{1\ \cancel{\text{mile}}} \times \frac{1\ \cancel{\text{hr}}}{3.6 \times 10^3\text{ s}}$$

$= 44\ ms^{–1}$. *(1 M)*

(iv) $1\ Å = 10^{–10}$ m.

$$\text{Unit factor} = \frac{10^{-10}\text{ m}}{1Å}$$

$$\therefore\ 0.74\ Å = \frac{0.74\ \cancel{Å} \times 10^{-10}\text{ m}}{1\ \cancel{Å}}$$

$= 0.74 \times 10^{–10}$ m or $= 7.4 \times 10^{–11}$ m *(1 M)*

(v) 0°C = 273.15 K

46°C = 273.15 + 46 = 319.15 K *(1 M)*

2. (i) 10 mL of H_2 combines with 5 mL of O_2 to form water. When 200 mL of H_2 at STP is passed over heated CuO, the CuO loses 0.144 g of its weight. Does the above data correspond to the law of constant composition?

(ii) Which one of the following will have the largest number of atoms?

(*a*) 1 g Au (s) (Atomic mass = 197 u)

(*b*) 1 g Na (s) (Atomic mass = 23 u)

(*c*) 1 g Li (s) (Atomic mass = 7 u) *(Un)*

Ans. (i) In the second experiment, 0.144 g weight is lost from CuO. This is due to the reduction of CuO into Cu. In other words, 0.144 g oxygen combined with 200 mL of H_2.

32 g oxygen occupies 22400 mL volume at STP.

$$\therefore\ 0.144\text{ g oxygen will occupy} = 22400 \times \frac{0.144}{32}$$

= 100.8 ml

It means the ratio of H_2 and O_2 in water is 200 : 100.8 = 2 : 1. The same ratio is in first case (10: 5 or 2: 1).

Thus, the data corresponds to the law of constant composition. *(2 M)*

2 MOLE CONCEPT, STOICHIOMETRY CALCULATIONS, CONCENTRATION TERMS

NCERT Definitions

- **Gram Molecular Mass or Gram Molecule:** Gram molecular mass is the mass in grams of one mole of a molecular substance.
- **Formula Mass:** The formula mass of a molecule is the sum of the atomic weights of the atoms in the empirical formula of the compound.
- **Molar Mass:** Molar mass is the mass of one mole of any substance i.e., element or compound.
- **Limiting Reactant:** In a chemical reaction, the reactant which gets consumed first, limits the amount of product formed and is called the limiting reagent. It is present in smaller amount.
- **Percentage composition:** Percentage composition of a compound is the ratio of the amount of each element to the total amount of individual elements present in a compound multiplied by 100.

Important Facts

01 1 amu = 1.66056×10^{-24} g

02 1 Mole of substance = Molar mass of substance = Avogadro's number = 22.4 L volume at STP

03 Molar Stoichiometry calculations are done only for balanced equations.

04 Molarity of a solution decreases on increasing temperature.

05 Molarity of pure water is 55.56 mol L^{-1}.

06 Molality is independent of temperature.

Important Concepts

Mole Concept

Quite commonly, we use different units for counting such as dozen for 12 articles and score for 20 articles irrespective of their nature. In a similar way, chemists use the unit mole for counting atoms, molecules, ions, etc. A mole is a collection of 6.022×10^{23} particles. Thus, a mole is defined as "the amount of substance that contains as many particles or entities (atoms, molecules or ions), as there are atoms in exactly 12 g (or 0.012 kg) of carbon-12 isotope."

In order to determine the number of atoms in 12 g of C-12, the mass of a carbon-12 was determined by a mass spectrometer. It was found to be equal to 1.992648×10^{-23} g. Knowing that 1 mol of carbon weighs 12 g, the number of atoms in 1 mol of C-12 is:

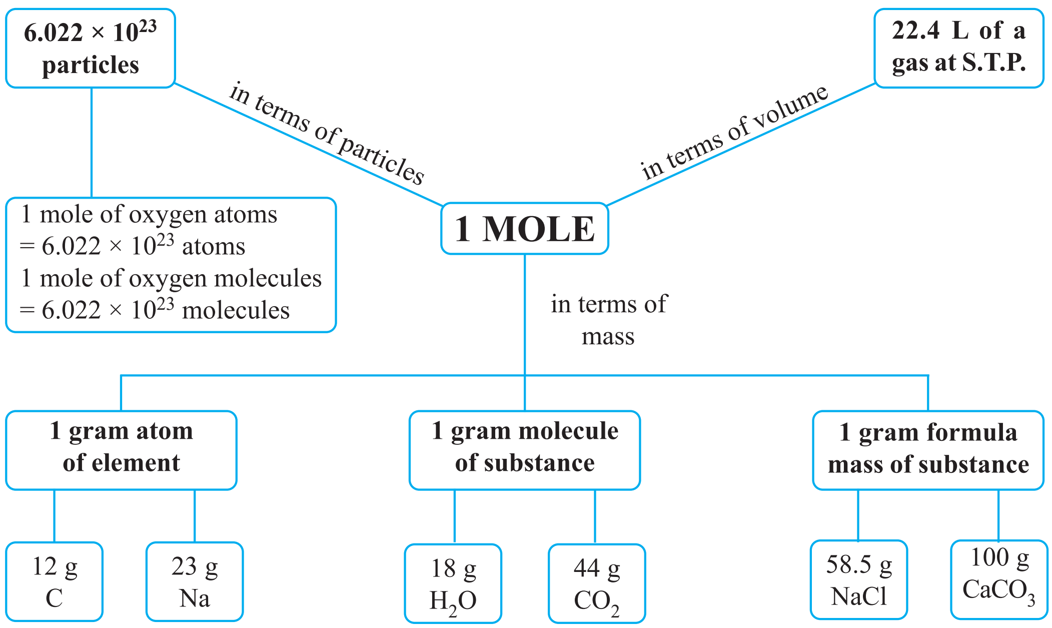
$$= \frac{12 \text{ g/mol of C-12}}{1.992648 \times 10^{-23} \text{ g/atom of C}}$$

$$= 6.022137 \times 10^{23} \text{ atoms/mol}$$

The number 6.022×10^{23} is called Avogadro number or Avogadro constant. It is denoted by N_A in honour of nineteenth century Italian scientist, Amedeo Avogadro. In other words, a mole is an Avogadro number of particles. For example,

1 mole of hydrogen atoms = 6.022×10^{23} hydrogen atoms.

1 mole of hydrogen molecules = 6.022×10^{23} hydrogen molecules.



Limiting Reactant

In many cases, the substances in a mixture are not present in exactly the same amount as required by the balanced chemical equation. In such situations, one reactant is in excess over the other. The reactant which is present in lesser amount gets consumed after sometime and after that no further reaction takes place even though the other reactant is present. Thus, the reactant which gets completely consumed in a reaction is called limiting reactant or reagent.

The concentration of the limiting reactant limits the amount of products formed. The other reactants present in quantity greater than those needed to react with the quantity of the limiting reagent present be left unreacted. It is also called excess reagent. For example, in the reaction

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

if the reaction mixture contains 2 mol of H_2 and 2 mol of O_2, then only 1 mol of O_2 will be used up and 1 mol of O_2 will be left over. Therefore, in this case, hydrogen is limiting reactant. Oxygen is excess reactant.

Molarity of a solution

It is the number of moles of the solute dissolved per litre of the solution. It is represented as 'M'. Thus, a solution which contains one gram mole of the solute dissolved per litre of the solution, is regarded as one molar solution.

For example, 1M Na_2CO_3 (molar mass = 106 g/mol) solution has 106 g of the solute present per litre of the solution.

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litres}}$$

It is convenient to express volume in cm^3 or mL so that

$$\text{Molarity} = \frac{\text{Moles of solute} \times 1000}{\text{Volume of solution (in mL or cm}^3\text{)}} \qquad (\because 1 \text{ litre} = 1000 \text{ mL})$$

Thus, the units of molarity are moles per litre (mol L^{-1}) or moles per cubic decimetre (mol dm^{-3}). The symbol M is used for mol L^{-1} or mol dm^{-3} and it represents molarity.

Important Formulas

- **Number of moles** $= \dfrac{\text{Given mass}}{\text{Molar mass}}$
- **Number of moles of a gas** $= \dfrac{\text{Volume of the gas (STP)}}{22.4 \text{ L}}$
- **Number of molecules** = Number of moles × N_A
- **Percentage Composition:**

  $$\text{Mass \% of the element} = \frac{\text{Mass of element in a molecule of the compound}}{\text{Molecular mass of the compound}} \times 100$$
- **Relationship between Empirical and Molecular Formulae:**

  $$\text{Molecular formula} = n \times \text{Empirical formula, where } n = \frac{\text{Molecular formula mass}}{\text{Empirical formula mass}}$$
- **Mass percentage** $= \dfrac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$
- **Mole fraction:**

  $$\mathbf{x_A} = \frac{\text{Number of moles of A}}{\text{Number of moles of solution}} = \frac{n_A}{n_A + n_B}$$

  $$\mathbf{x_B} = \frac{\text{Number of moles of B}}{\text{Number of moles of solution}} = \frac{n_B}{n_A + n_B}$$
- **Molarity (M)** $= \dfrac{\text{Number of moles of solute}}{\text{Volume of solution in litres}}$
- **Molality (m)** $= \dfrac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$

Difference Between

Molality vs. Molarity

Property	Molality (m)	Molarity (M)
Definition	Moles of solute per kilogram of solvent	Moles of solute per liter of solution
Formula	m = moles of solute/kg of solvent	M = moles of solute/litre of solution
Unit	mol/kg	mol/L
Depends on Temperature?	No, since mass does not change with temperature	Yes, because volume can expand or contract with temperature

Real Life Application Based Questions

1. **How does the concept of the mole help in medicine?**

Ans. Medicines are prescribed in specific dosages based on the molar mass of compounds. For example, a doctor prescribes medicine based on the required number of molecules (moles) to produce the desired effect without causing toxicity.

2. **How does understanding molar mass helps in agriculture (fertilizers)?**

Ans. Farmers use fertilizers based on the amount of nitrogen, phosphorus or potassium present in a compound. For example, urea ($CO(NH_2)_2$) has a molar mass of 60 g/mol and knowing this helps in calculating the correct dosage for crops.

3. **How does percentage composition helps in food labelling?**

Ans. The percentage composition of nutrients like proteins, carbohydrates, fats and minerals in packaged food helps consumers understand the nutritional value and choose products accordingly.

4. **How does stoichiometry help in making fireworks?**

Ans. Fireworks contain compounds like potassium nitrate (KNO_3) and aluminum powder in precise amounts to ensure controlled explosions and vibrant colours. If the proportions are incorrect, the fireworks may not burn properly or could become dangerous.

COMPETENCY BASED SOLVED EXAMPLES

Multiple Choice Questions (1 M)

1. **What is the average atomic mass of bromine from the following data : (abundance is in %)** *(Ap)*

Isotope	**Mass**	**Abundance**
$_{79}Br$	**78.9183361**	**50.69**
$_{81}Br$	**80.916289**	**49.31**

(*a*) 79.9 u (*b*) 76.6 u
(*c*) 75.9 u (*d*) 69.9 u

2. **Number of atoms in 558.5 g Fe (at. mass of Fe = 55.85 u) is:** *(An)* **(CBSE, 2023)**

(*a*) twice that in 60 g carbon
(*b*) 6.023×10^{22}
(*c*) half that in 8 g He
(*d*) $558.5 \times 6.023 \times 10^{23}$

3. **Empirical formula of hydrocarbon containing 80% carbon and 20% hydrogen is :** *(Un)* **(CBSE APQ, 2023)**

(*a*) CH_3 (*b*) CH_4
(*c*) CH (*d*) CH_2

4. **In the reaction**

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$$

When 1 mole of ammonia and 1 mole of O_2 are made to react to completion, *(Un)* **(NCERT Exemplar)**

(*a*) 1.0 mole of H_2O is produced.
(*b*) 1.0 mole of NO will be produced.
(*c*) all the oxygen will be consumed.
(*d*) all the ammonia will be consumed.

Assertion and Reason (1 M)

Direction: The following questions consist of two statements – Assertion (A) and Reason (R). Answer these questions by selecting the appropriate option given below:

(*a*) Both A and R are true and R is the correct explanation of A.
(*b*) Both A and R are true 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.

1. **Assertion (A):** Equal moles of different substances contain same number of constituent particles.

Reason (R): Equal weights of different substances contain the same number of constituent particles. *(Re)*

2. **Assertion (A):** The empirical mass of ethene is half of its molecular mass.

Reason (R): The empirical formula represents the simplest whole number ratio of various atoms present in a compound. *(Re)*

3. **Assertion (A):** The sum of mole fractions of all the components of a solution is unity.

Reason (R): Mole fraction is temperature dependent mode of concentration. *(Re)*

4. **Assertion (A):** One mole of SO_2 contains double the number of molecules present in one mole of O_2.

Reason (R): Molecular weight of SO_2 is two times to that of O_2. *(Un)*

Subjective Questions

Very Short Answer Type Questions (2 M)

1. **(i) Calculate the mass of an atom of silver (atomic mass = 108 u), and** *(Un)*

(ii) 1 molecule of naphthalene ($C_{10}H_8$)

Ans. (i) Mass of 6.022×10^{23} atoms of silver = 108 g

$$\text{Mass of 1 atom of silver} = \frac{108}{6.022\times10^{23}} = 1.793 \times 10^{-22} g$$ *(1 M)*

(ii) Molecular mass of naphthalene ($C_{10}H_8$)

$= 10 \times 12 + 8 \times 1 = 128$ u

Mass of 6.022×10^{23} molecules of naphthalene = 128 g

$$\text{Mass of 1 molecule of naphthalene} = \frac{128}{6.022\times10^{23}}$$

$= 2.12 \times 10^{-22}$ g *(1 M)*

2. **Calculate the mass percentage composition of copper pyrite ($CuFeS_2$). (Atomic mass of Cu = 63.5 u, Fe = 55.8 u and S = 32 u)** *(Ap)* **(CBSE, 2023)**

Ans. Molar mass of $CuFeS_2 = 63.5 + 55.8 + 2 \times 32 = 183.3$ g/mol

Mass of Cu = 63.5 g

$$\text{Mass percentage of Cu} = \frac{63.5}{183.3}\times100 = 34.64\%$$

Mass of iron = 55.8 g

$$\text{Mass percentage of iron} = \frac{55.8}{183.3}\times100 = 30.44\%$$

Mass of sulphur = 2 × 32 = 64 g

$$\text{Mass percentage of sulphur} = \frac{64}{183.3}\times100 = 34.91\%$$ *(2 M)*

3. **2.46 g of sodium hydroxide (molar mass = 40 g/mol) are dissolved in water and the solution is made to 100 cm³ in a volumetric flask. Calculate the molarity of the solution.** ***(Un)***

Ans. Amount of NaOH = 2.46 g

Volume of solution = 100 cm³

$$\text{Moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{Molar mass}}$$

$$= \frac{2.46}{40} = 0.0615 \qquad \textit{(1 M)}$$

$$\text{Molarity} = \frac{\text{Mass of NaOH}}{\text{Volume of solution}} \times 1000$$

$$= \frac{0.0615}{100} \times 1000$$

$$= 0.615 \text{ M.} \qquad \textit{(1 M)}$$

4. **Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.** ***(Un)***

Ans. $$x_{C_2H_5OH} = \frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + n_{H_2O}} = 0.040 \qquad ...(i)$$

As the solution is dilute,

Number of moles of water in 1 L of water

$$= \frac{1000\text{g}}{18\text{g mol}^{-1}} = 55.55 \text{ moles}$$

Substituting $n_{H_2O} = 55.55$ in eq. (i), we get

$$\frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + 55.55} = 0.040$$

or, $0.96\ n_{C_2H_5OH} = 55.55 \times 0.040$

or, $n_{C_2H_5OH} = 2.31$ mol

Hence, molarity of the solution = 2.31 M *(2 M)*

5. **In the reaction 2A + 4B → 3C + 4D, when 5 moles of A react with 6 moles of B, then**
   (i) which is the limiting reagent?
   (ii) calculate the amount of C formed. ***(Un)***

Ans. In the reaction, 2A + 4B → 3C + 4D

(i) Limiting reagent

2 moles of A react with 4 moles of B

$$5 \text{ moles of A will react with} = \frac{4}{2} \times 5$$

$$= 10 \text{ moles of B}$$

Since in the reaction only 6 moles of B are there, hence B is the limiting reagent. *(1 M)*

(ii) Amount of C formed

Since B is the limiting reagent, reaction will proceed according to B.

4 moles of B give 3 moles of C

$$6 \text{ moles of B will give} = \frac{3}{4} \times 6 = 4.5 \text{ moles of C} \qquad \textit{(1 M)}$$

Short Answer Type Questions **(3 M)**

1. **Calculate the number of atoms of each type in 5.3 g of Na_2CO_3. (At. mass of Na = 23 u, C = 12 u and O = 16 u)** ***(Un)***

Ans. Gram molecular mass of Na_2CO_3

= 2 × 23 + 12 + 3 × 16 = 106.0 g

106.0 g of Na_2CO_3 = 1 mol

$$5.3 \text{ g of } Na_2CO_3 = \frac{1 \times 5.3}{106.0} = 0.05 \text{ mol} \qquad \textit{(1 M)}$$

Now, 1 mol of Na_2CO_3 contains = 2 × 6.022 × 10^{23} sodium atoms

∴ 0.05 mol of Na_2CO_3 contains = 0.05 × 2 × 6.022 × 10^{23} sodium atoms = 6.022 × 10^{22} Na atoms *(½ M)*

1 mol of Na_2CO_3 contains = 6.022 × 10^{23} carbon atoms

0.05 mol of Na_2CO_3 contains = 0.05 × 6.022 × 10^{23} carbon atom = 3.01 × 10^{22} carbon atoms *(½ M)*

1 mol of Na_2CO_3 contains = 3 × 6.022 × 10^{23} oxygen atoms

∴ 0.05 mol of Na_2CO_3 contains = 0.05 × 3 × 6.022 × 10^{23} oxygen atoms = 9.03 × 10^{22} oxygen atoms. *(1 M)*

2. **The cost of table salt (NaCl) and sugar ($C_{12}H_{22}O_{11}$) are ₹12 per kg and ₹36 per kg respectively. Calculate their cost per mole.** ***(Un)***

Ans. (*a*) Cost of table salt per mole :

Molar mass of table salt (NaCl) = 23 + 35.5 = 58.5 g/mol

Now, 1000 g of NaCl cost = ₹12

$$\therefore 58.5 \text{ g of NaCl will cost} = \frac{12}{1000} \times 58.5 = ₹0.702$$

= 70 paise (approx.) *(1½ M)*

(*b*) Cost of sugar per mole:

Molar mass of sugar ($C_{12}H_{22}O_{11}$)

= 12 × 12 + 22 × 1 + 11 × 16 = 342 g/mol

Now, 1000 g of sugar cost = ₹36

$$342 \text{ g of sugar will cost} = \frac{36}{1000} \times 342 = ₹12.312$$

= ₹12 (approx) *(1½ M)*

3. **The molecular mass of an organic compound is 78 u and its percentage composition is 92.4% C and 7.6% H. Determine the molecular formula of the compound.** ***(Un)***

Ans. Calculation of empirical formula

Element	C	H
Percentage	92.4	7.6
Atomic mass	12.0	1
Moles of atoms	$\frac{92.4}{12.0} = 7.7$	$\frac{7.6}{1} = 7.6$
Mole ratio or atomic ratio	$\frac{7.7}{7.6} = 1.01$	$\frac{7.6}{7.6} = 1$
Simplest whole no. ratio	1	1

The simplest whole number ratio of C : H is 1 : 1

∴ The empirical formula of the compound is CH.

Calculation of molecular formula of the compound.

Empirical formula mass = 1 × 12 + 1 × 1 = 13 u

Molecular mass = 78 u

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{78}{13} = 6$$

Thus, molecular formula of the compound

= 6 × (CH) = C_6H_6. *(3 M)*

4. If 20.0 g of $CaCO_3$ is treated with 20.0 g of HCl, how many grams of CO_2 can be produced according to the reaction : ***(Un)***

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

Ans. First of all we have to calculate the limiting reactant.

$CaCO_3(s)$	+ $2HCl(aq)$	$\rightarrow CaCl_2(aq) + H_2O(l)$
1 mol	2 mol	$+ CO_2(g)$
40+12+3×16	2×(1+35.5)	
= 100 g	= 73 g	

According to above equation,

100 g of $CaCO_3$ require 73 g of HCl

20 g of $CaCO_3$ require $= \frac{73}{100} \times 20 = 14.6$ g

But amount of HCl actually present = 20.0 g

Therefore, $CaCO_3$ is limiting reactant and HCl is excess reactant. *(1½ M)*

Now let us calculate the amount of CO_2 produced when entire quantity of limiting reactant reacts.

$$CaCO_3(s) + 2HCl(g) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

1 mol	1 mol
100 g	44 g
20 g	?

100 g of $CaCO_3$ produces CO_2 = 44 g

20 g of $CaCO_3$ will produce $CO_2 = \frac{44}{100} \times 20 = 8.80$ g

(1½ M)

Long Answer Type Questions (5 M)

1. (*a*) A solution of oxalic acid, $(COOH)_2.2H_2O$ is prepared by dissolving 0.63 g of the acid in 250 mL of the solution. Calculate (i) molarity ***(Un)***

(*b*) A solution of glucose (MM = 180 g/mol) in water is labelled as 10% (w/w). The density of the solution is 1.20 g mL^{-1}. Calculate

(i) molality

(ii) molarity, and

(iii) mole fraction of each component in solution. ***(Un)***

Ans. (*a*) (i) Calculation of molarity

Molar mass of oxalic acid, $(COOH)_2.2H_2O$

= 2 (12 + 32 + 1) + 2 × 18

= 126 g mol^{-1} *(½ M)*

Moles of oxalic acid $= \frac{0.63}{126} = 0.005$ mol

Volume of solution = 250 mL

Molarity $= \frac{0.05 \times 1000}{250} = 0.02$ M *(1 M)*

(*b*) 10% (w/w) solution of glucose means that 10 g of glucose is present in 100 g of solution or in 90 g of water.

(i) Calculation of molality

Mass of glucose = 10 g

Moles of glucose $= \frac{10}{180} = 0.0556$

(Molar mass of glucose = 180 g/mol)

Mass of water = 90 g

$\therefore$ Molality $= \frac{\text{Moles of glucose}}{\text{Mass of water}} \times 1000$

$= \frac{0.0556}{90} \times 1000 = 0.618$ m *(1 M)*

(ii) Calculation of molarity

Moles of glucose = 0.0556

Volume of solution $= \frac{\text{Mass}}{\text{Density}}$

$= \frac{100}{1.20} = 83.3$ mL *(½ M)*

Molarity $= \frac{\text{Moles of glucose}}{\text{Vol. of solution}} \times 1000$

$= \frac{0.0556}{83.3} \times 1000 = 0.667$ M. *(1 M)*

(iii) Calculation of mole fraction of components

Moles of glucose = 0.0556

Moles of water $= \frac{90}{18} = 5.0$

Total moles = 5.0 + 0.0556 = 5.0556

Mole fraction of glucose $= \frac{0.0556}{5.0556} = 0.011$.

Mole fraction of water $= \frac{5.0}{5.0556} = 0.989$. *(1 M)*

2. (i) Butyric acid contains only C, H and O. A 4.24 mg sample of butyric acid is completely burnt. It gives 8.45 mg of carbon dioxide and 3.46 mg of water. What is the mass percentage of each element in butyric acid? ***(Ap)***

(ii) The molecular mass of butyric acid was determined by experiment to be 88 u. What is the molecular formula? ***(Un)***

MISCELLANEOUS EXERCISE

Multiple Choice Questions (1 M)

1. Which of the following is dependent of temperature ? ***(Un)***

(*a*) Molarity
(*b*) Molality
(*c*) Mole fration
(*d*) Mass percentage

2. 4 g of NaOH (MM = 40 g/mol) dissolved in 100 mL solution. Molarity of the solution is: ***(Un)***

(*a*) 1 M (*b*) 10 M
(*c*) 0.1 M (*d*) 4 M

3. Which has the maximum number of molecules among the following? (At. mass C = 12 u, O = 16 u and S = 32 u) ***(Ap)***

(*a*) 44 g of CO_2
(*b*) 44 g of O_2
(*c*) 8 g of H_2
(*d*) 64 g of SO_2

4. 10 mol of Zn react with 10 mol of HCl. Calculate the number of moles of H_2 produced. ***(Un)***

$Zn + 2HCl \rightarrow ZnCl_2 + H_2$

(*a*) 5 mol (*b*) 10 mol
(*c*) 20 mol (*d*) 2.5 mol

5. The number of oxygen atoms in 4.4 g of CO_2 is approximately: ***(Ap)***

(*a*) 1.2×10^{23} (*b*) 6×10^{22}
(*c*) 6×10^{23} (*d*) 12×10^{23}

6. The molarity of a solution obtained by mixing 750 mL of 0.5 M HCl with 250 ml of 2 M HCl will be: ***(Un)***

(*a*) 0.975 M (*b*) 0.875 M
(*c*) 1.00 M (*d*) 1.175 M

7. Number of atoms of He in 100 u of He is: (Atomic mass of He is 4 u) ***(Ap)***

(*a*) 25 (*b*) 50
(*c*) 100 (*d*) 400

8. 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of the solution is: ***(Un)***

(*a*) 0.02 M (*b*) 0.01 M
(*c*) 0.001 M (*d*) 0.1 M

9. A gaseous hydrocarbon gives upon combustion, 0.72 g of H_2O and 3.08 g of CO_2. The empirical formula of the hydrocarbon is: (MM of H_2O = 18 g/mol and CO_2 = 44 g/mol): ***(Un)***

(*a*) C_6H_5 (*b*) C_7H_8
(*c*) C_2H_4 (*d*) C_3H_4

10. If Avogadro number, is changed from 6.022×10^{23} mol^{-1} to 6.022×10^{20} mol^{-1}, this would change: ***(Un)***

(*a*) the mass of one mole of carbon.
(*b*) the ratio of chemical species to each other in a balanced equation.
(*c*) the ratio of elements to each other in a compound.
(*d*) the definition of mass in units of grams.

11. When 22.4 litres of $H_2(g)$ is mixed with 11.2 litres of $Cl_2(g)$, each at S.T.P, the moles of HCl(g) formed is equal to: ***(Un)***

(*a*) 1 mol of HCl(g)
(*b*) 2 mol of HCl(g)
(*c*) 0.5 mol of HCl(g)
(*d*) 1.5 mol of HCl(g)

12. 25.3 g of sodium carbonate, Na_2CO_3 is dissolved in enough water to make 250 mL of solution. If sodium carbonate dissociates completely, molar concentration of sodium ion, Na^+ and carbonate ions, CO_3^{2-} are respectively (Molar mass of Na_2CO_3 = 106 g mol^{-1}) ***(Un)***

(*a*) 0.955 M and 1.910 M
(*b*) 1.910 M and 0.955 M
(*c*) 1.90 M and 1.910 M
(*d*) 0.477 M and 0.477 M

13. An element, X has the following isotopic composition:

^{200}X : 90% ^{199}X : 8.0% ^{202}X : 2.0%

The weighted average atomic mass of the naturally occurring element X is closest to: ***(Ap)***

(*a*) 201 amu (*b*) 202 amu
(*c*) 199 amu (*d*) 200 amu

14. Percentage of Se in peroxidase anhydrous enzyme is 0.5% by weight (At. mass = 78.4) then minimum molecular weight of peroxidase anhydrous enzyme is: ***(Ap)***

(*a*) 1.568×10^4
(*b*) 1.568×10^3
(*c*) 15.68
(*d*) 2.136×10^4

15. Given the numbers: 161 cm, 0.161 cm, 0.0161 cm. The number of significant figures for the three numbers is: ***(Ap)***

(*a*) 3, 3 and 4 respectively.
(*b*) 3, 4 and 4 respectively.
(*c*) 3, 4 and 5 respectively.
(*d*) 3, 3 and 3 respectively.

16. The total number of valence electrons in 4.2 g of N^{3-} ion is: (N_A is the Avogadro's number) ***(Ap)***

(*a*) 2.1 N_A (*b*) 4.2 N_A
(*c*) 1.6 N_A (*d*) 3.2 N_A

Assertion and Reason (1 M)

Direction: The following questions consist of two statements – Assertion (A) and Reason (R). Answer these questions by selecting the appropriate option given below:

(*a*) Both A and R are true and R is the correct explanation of A.

(*b*) Both A and R are true, 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.

1. Assertion (A): A solution of table salt in a glass of water is homogeneous

Reason (R): A solution having same composition throughout is heterogeneous. (*Un*)

2. Assertion (A): The molecular weight of oxygen is 32 amu.

Reason (R): The atomic weight of oxygen is 16 amu. (*Un*)

3. Assertion (A): No of moles of H_2 in 0.224 L of hydrogen is 0.01 moles.

Reason (R): 22.4 L of H_2 at STP contain 6.023×10^{23} moles. (*Un*)

4. Assertion (A): Atomic mass of Na is 23 u.

Reason (R): An atom of sodium is 23 times heavier than 1/12th mass of C-12 isotope. (*Re*)

5. Assertion (A): Number of atoms of He in 60 u is 15.

Reason (R): Atomic weight of He is 4 u. (*Re*) (NCERT Exemplar)

Subjective Questions

Very Short Answer Type Questions (2 M)

1. What is the difference between molecules and compounds? Give examples of each. Classify following as pure substances and mixtures : air, glucose, gold, sodium and milk. (*Re*)
2. State Avogadro's law and law of conservation of mass. (*Re*)
3. Define significant figures. How many significant figures are there in:
   (i) 3.070 and (ii) 0.0025 (*Ap*)
4. The body temperature of a normal healthy person is 37°C. Calculate its value in °F. (*Ap*)
5. Which has more number of atoms? 1.0 g Na or 1.0 g Mg (*An*)
6. Calculate the number of moles in the following masses. (Atomic mass of Fe = 56 u and Ca = 40 u) (*An*)
   (i) 7.85 g of Fe (ii) 7.9 mg of Ca
7. Calculate the percent of carbon, hydrogen and oxygen in ethanol (C_2H_5OH). (Atomic mass of C = 12 u and O = 16 u) (*Un*)
8. 56 kg of $N_2(g)$ and 10 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. Calculate the number of moles of ammonia gas formed. (Atomic mass of N = 14 u, H = 1 u) (*Un*)
   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
9. Calculate the molarity of NaOH in the solution prepared by dissolving its 4 g in enough water to form 250 mL of the solution. (Atomic mass of Na = 23 u and O = 16 u) (*Un*)
10. Calculate the amount of water (in grams) produced by the combustion of 16 g of methane. (*Un*)

Short Answer Type Questions (3 M)

1. State the law of Constant Composition. Illustrate with an example. (*Re*)
2. Define homogeneous and heterogeneous mixture with examples. (*Re*)
3. Calculate: (*Ap*)
   (i) Mass in gram of 5.8 mol N_2O
   (ii) Number of moles in 8.0 g of O_2
   (iii) Molar mass of 11.2 L at STP weigh 8.0 g
4. 16 g of an ideal gas SO_x occupies 5.6 L at STP. What is its molecular mass? What is the value of X? (*Un*)
5. The reaction $2C + O_2 \rightarrow 2CO$ is carried out by taking 24.0 g of carbon and 96.0 g of O_2. Find out: (*Un*)
   (i) Which reactant is left in excess?
   (ii) How much of it is left?
   (iii) How many grams of the other reactant should be taken so that nothing is left at the end of the reaction?
6. Potassium superoxide, KO_2 is used in rebreathing gas masks to generate oxygen.
   $4KO_2(s) + 2H_2O(l) \rightarrow 4KOH(s) + 3O_2(g)$
   If a reaction vessel contains 0.15 mol KO_2 and 0.10 mol H_2O, what is the limiting reagent?
   How many moles of oxygen can be produced? (*Un*)
7. Convert the following into SI units: (*Ap*)
   (i) 28.7 pm
   (ii) 15.15 pm
   (iii) 25365 mg
8. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass): (*Un*)
   (i) Express this in percent by mass
   (ii) Determine the molality of chloroform in the water sample.

Long Answer Type Questions **(5 M)**

1. (*a*) Define the terms: ***(Re)***

(i) Empirical formula

(ii) Molecular formula

(*b*) An organic compound on analysis gave the following data: C = 57.82%, H = 3.6% and the rest oxygen. If the molecular mass was found to be 166 u then calculate its empirical formula and molecular formula. ***(Un)***

2. (*a*) What is the difference between molarity and molality?

(*b*) The molarity of a solution of sulphuric acid is 1.35 M. Calculate its molality. (The density of acid solution is 1.02 g cm^{-3}). ***(Un)***

3. (*a*) Define the terms: ***(Re)***

(i) Limiting reagent

(ii) Mole

(*b*) 20 g of $CaCO_3$ and 20 g of H_2SO_4 react to give $CaSO_4$ along with water and CO_2.

(i) Determine the limiting reagent for the above reaction.

(ii) How much $CaSO_4$ will be formed?

(iii) If 1 mole of gas occupies 22.4 L at STP then calculate the volume of CO_2 evolved in the above reaction.

[Ca = 40 u, C = 12 u, O = 16 u, H = 1 u, S = 32 u]

4. (*a*) Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction given below:

$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

What mass of $CaCl_2$ will be formed when 250 ml of 0·76 M HCl reacts with 1000 g of $CaCO_3$? Name the limiting reagent. Calculate the number of moles of $CaCl_2$ formed in the reaction. ***(Un)***

Case Based Questions

Case Based-I

Recent studies have revealed that the simplest form of matter is atoms and elements may also be defined as the pure substance which is made of one kind of atoms. Examples are carbon, sulphur, hydrogen, oxygen, etc. A compound is also a pure substance like element but it is made up of two or more elements. For example, in sodium chloride the two elements sodium and chlorine are present in the ratio of 23:35.5 by mass. Both elements and compounds are pure substances. But on mixing two or more substances in any ratio, mixture results. For example, air is a mixture of different gasses like nitrogen, oxygen, carbon dioxide, water vapour, etc. Further, mixtures are divided into two categories: homogeneous and heterogeneous.

(i) Classify the following as pure substances or mixtures: Graphite and iodized table salt. ***(Re)***

(ii) Why is tap water considered a mixture while distilled water as a compound? ***(Re)***

(iii) Why is the gaseous state of ammonia regarded as gas while that of water as vapours? ***(Re)***

Case Based-II

Atoms and molecules are so small in size that it is neither possible to count them individually nor possible to determine their mass. These are counted collectively in terms of Avogadro's number. The mass of Avogadro's number of atoms and molecules is known as gram atomic mass and gram molecular mass respectively. The volume occupied by Avogadro's number of molecules of a gas or vapour is known as molar volume.

(i) Mass of CO_2 (MM = 44 g/mol) is 88 g. The number of atoms of oxygen present in it, is: ***(Ap)***

(*a*) 2.41×10^{24} (*b*) 1.2×10^{23}
(*c*) 1.4×10^{23} (*d*) 2.41×10^{23}

(ii) Calculate the molecular mass of cane sugar ($C_{12}H_{22}O_{11}$). (Atomic mass of C = 12 u, O = 16 u, H = 1 u) ***(Ap)***

(*a*) 350 g (*b*) 361 g
(*c*) 342 g (*d*) 345 g

(iii) If N_A is Avogadro's number, then the number of valence electrons in 70 g of nitride ions (N^{3-}) is: ***(Ap)***

(*a*) 42 N_A (*b*) 40 N_A
(*c*) 16 N_A (*d*) 45 N_A

(iv) Choose the correct mass (in grams) of 11.2 L of N_2 at STP. ***(Ap)***

(*a*) 13 g (*b*) 14.5 g
(*c*) 14 g (*d*) 15 g

Case Based-III

A binary solution is made up of two liquids that are entirely miscible with each other. In a binary solution, the component with the lowest concentration is known as the solute, while the component with the highest concentration is known as the solvent. One mole of the solute present in 1L of solution is called 1 molar solution. A 1 molal solution is one in which one mole of solute is dissolved in one kilogram of solvent. The number of moles of a given component to the total number of moles in the solution is referred to as the mole fraction.

(i) 6.02×10^{20} molecules of urea are present in 250 mL of its solution. The concentration of the solution is: ***(Un)***

(*a*) 0.02 M (*b*) 0.004 M
(*c*) 0.001 M (*d*) 0.1 M

(ii) What will be the mole fraction of glycol $C_2H_4(OH)_2$ in a solution containing 45 g of water and 56 g of glycol? (Atomic mass of C = 12 u, O = 16 u and H = 1u) ***(Un)***

(*a*) 0.31 (*b*) 0.50
(*c*) 0.26 (*d*) 0.10

(iii) The value of molality for pure water is: ***(Re)***

(*a*) 55.55 (*b*) 52.6
(*c*) 52 (*d*) 25

(iv) What is the correct advantage for using molality over molarity? ***(Re)***

(*a*) Molarity does not depend upon temperature.
(*b*) Molality does not depend upon temperature.
(*c*) Molality depend on temperature.
(*d*) None of the above

Case Based-IV

The reactants react according to the balanced chemical equation. Quite often, these are not present in the same proportions as is required by the equation; some may be present in a lesser amount while the others may be present in excess than the stoichiometric amounts. The reactant which is present in a lesser quantity is known as a limiting reagent or limiting reactant since it limits the participation of the other reactants in the reaction and also the product of the reaction. For example, in the combustion of methane with oxygen, methane is the limiting reactant because oxygen (present in air) is always available more than the amount of methane. Amount of carbon dioxide and water formed also depends upon the amount of methane and not oxygen.

(i) Find the number of moles of lead (II) chloride formed as a result of the reaction between 6.5 g of PbO and 3.2 g of HCl. (Atomic mass of Pb = 207.2 u) **(*Un*)**

$PbO + 2HCl \rightarrow PbCl_2 + 2H_2O$

(ii) 14g hydrogen and 80 g oxygen were filled in a steel vessel and exploded. The amount of water produced in the reaction will be? **(*Un*)**

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

(iii) Why is it necessary to balance a chemical equation? **(*Un*)**

ANSWER KEYS

Multiple Choice Questions

1. (*a*) **2.** (*a*) **3.** (*c*) **4.** (*a*) **5.** (*a*) **6.** (*b*) **7.** (*a*) **8.** (*b*) **9.** (*b*) **10.** (*a*)
11. (*a*) **12.** (*b*) **13.** (*d*) **14.** (*a*) **15.** (*d*) **16.** (*c*)

Assertion and Reason

1. (*c*) **2.** (*b*) **3.** (*c*) **4.** (*a*) **5.** (*a*)

HINTS & EXPLANATIONS

Multiple Choice Questions

1. (*a*) Molarity is temperature dependent as it includes volume.

2. (*a*) $M = \frac{m_{solute} \times 1000}{MM_{solute} \times V_{(solution)(mL)}} = \frac{4 \times 1000}{40 \times 100} = 1\ M$

3. (*c*) Number of molecules $= \frac{m \times N_A}{MM}$

4. (*a*) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

1 mol Zn reacts with 2 mol HCl

$\therefore$ 10 mol Zn reacts with 20 mol HCl

Hence, HCl is limiting reagent.

Now, 2 mol HCl produces 1 mol H_2

So, 10 mol HCl produces 5 mol H_2.

5. (*a*) Number of O-atoms $= \frac{m}{MM} \times N_A \times$ atomicity

$= \frac{4.4}{44} \times 6 \times 10^{23} \times 2 = 1.2 \times 10^{23}$

6. (*b*) $M_1V_1 = M_2V_2 + M_3V_3$

$M_1 \times 1000 = (0.5 \times 750) + (2 \times 250)$

$M_1 = 0.875\ M$

7. (*a*) number of He atom $= \frac{100}{4} = 25$.

8. (*b*) $M = \frac{6.02 \times 10^{20} \times 1000}{6.02 \times 10^{23} \times 100} = 0.01\ M$.

9. (*b*) Moles of $CO_2 = \frac{3.08}{44} = 0.07$ mol.

Moles of $H_2O = \frac{0.72}{18} = 0.04$ mol.

Ratio of $CO_2 = \frac{C}{H} = \frac{0.07}{0.08} = \frac{7}{8}$

Empirical formula = C_7H_8

10. (*a*) Mass of 1 mol (6.022 × 1023 atoms) of carbon = 12 g

If Avogadro number is changed to 6.022×10^{20} atoms then mass of 1 mol of carbon

$\frac{12 \times 6.022 \times 10^{20}}{6.022 \times 10^{23}} = 12 \times 10^{-3}$ g

11. (*a*) 1 mole = 22.4 litres at S.T.P.

$$n_{H_2} = \frac{22.4}{22.4} = 1\text{mol};\ n_{Cl_2} = \frac{11.2}{22.4} = 0.5 \text{ mol}$$

Reaction is as,

	$H_{2(g)}$	$Cl_{2(g)}$	$\rightarrow 2HCl_{(g)}$
Initial	1 mol	0.5 mol	0
Final	(1 – 0.5)	(0.5 – 0.5)	(2 × 0.5)
	= 0.5 mol	= 0 mol	1 mol

Here, Cl_2 is limiting reagent. So, 1 mole of $HCl_{(g)}$ is formed.

12. (*b*) Given that molar mass of Na_2CO_3 = 106 g

Molarity of solution = $\frac{25.3 \times 1000}{106 \times 250}$

= 0.9547 M = 0.955 M

$Na_2CO_3 \rightarrow 2Na^+ + CO_3^{-2}$

$[Na^+] = 2[Na_2CO_3] = 2 \times 0.955 = 1.910$ M

$[CO_3^{-2}] = [Na_2CO_3] = 0.955$ M.

13. (*d*) Average isotopic mass of X

$$= \frac{200 \times 90 + 199 \times 8 + 202 \times 2}{90 + 8 + 2}$$

$$= \frac{18000 + 1592 + 404}{100} = 199.96 \text{ amu} \simeq 200 \text{ amu}.$$

14. (*a*) In peroxidase anhydrous enzyme 0.5% Se is present means, 0.5 g Se is present in 100 g of enzyme. In a molecule of enzyme one Se atom must be present. Hence 78.4 g Se will be present in

$$= \frac{100}{0.5} \times 78.4 = 1.568 \times 10^4.$$

15. (*d*) Zeros placed left to the number are never significant, therefore the no. of significant figures for the numbers. 161 cm, 0.161 cm and 0.0161 cm are same, i.e. 3.

16. (*c*) Each nitrogen atom has 5 valence electrons, therefore total number of electrons in N_3^- ion is 16. Since the molecular mass of N_3 is 42, therefore total number of valence electrons in 4.2 g of N_3^- ion.

$$= \frac{4.2}{42} \times 16 \times N_A = 1.6\, N_A.$$

Assertion and Reason

1. (*c*) A solution having same composition throughtout is homogenous solution.
2. (*b*) Atomic weight of oxygen is 16 amu. Atomicity of molecular oxygen is 2, So its molecular weight is 32 amu.
3. (*c*) At STP, 22.4 L of H_2 contains 6.023×10^{23} molecules.
4. (*a*) Both assertion and reason are true and reason is correct explanation of assertion.
5. (*a*) Number of atom of He = $\frac{60 \text{ u}}{4 \text{ u}} = 15$.

Subjective Questions

Very Short Answer Type Questions

1. **(*a*) Molecule :** Two or more atoms of same elements combine to form molecule. e.g. O_2, N_2, Cl_2, etc. *(½ M)*

Compound : Two or more atoms of different elements combine to form compound. e.g. H_2O, NH_3, CO_2, etc. *(½ M)*

(*b*) Pure substances : Glucose, Gold, Sodium *(½ M)*

Mixture : Air, Milk *(½ M)*

Mistakes 101: What not to do!

Milk is a mixture of proteins and other substances and should not be confused as a pure substance.

2. According to Avogadro's law, "Equal volumes of gases at the same temperature and pressure should contain equal number of molecules". *(1 M)*

According to law of conservation of mass "In a chemical reaction, total mass of reactants is equal to total mass of products." *(1 M)*

Key Takeaways

Always write the law as stated and do not reframe the statement.

3. Significant figures are meaningful digits which are known with certainty plus one which is estimated or uncertain. The uncertainty is indicated by writing the certain digits and the last uncertain digit. *(1 M)*

(i) 4 (ii) 2 *(1 M)*

4. $°F = \frac{9}{5}(°C) + 32$

$= \frac{9}{5} \times 37 + 32 = 66.6 + 32 = 98.6$ °F *(2 M)*

Key Takeaways

Make sure to write the formula.

5. Number of atoms in 1.0 g Na = $\frac{6.022 \times 10^{23} \times 1}{23}$

$= 2.618 \times 10^{22}$ atoms *(1 M)*

Number of atoms in 1.0 g Mg = $\frac{6.022 \times 10^{23} \times 1}{24}$

$= 2.509 \times 10^{22}$ atoms

So, 1.0 g Na has more number of atoms. *(1 M)*

But we have 0.10 moles of H_2O

$\therefore$ KO_2 is the limting reagent. (2 M)

Again, according to reaction,

4 moles of KO_2 produces 3 moles of O_2

0.15 moles of KO_2 will produce $= \frac{3}{4} \times 0.15$ moles of O_2

$= 0.1125$ moles of O_2 (1 M)

7. (i) $\therefore$ 1 pm $= 10^{-12}$ m

$\therefore$ 28.7 pm $= 28.7 \times 10^{-12}$ m or 2.87×10^{-11} m (1 M)

(ii) 1 pm $= 10^{-12}$ m

$\therefore$ 15.15 pm $= 15.15 \times 10^{-12}$ m or 1.515×10^{-11} m (1 M)

(iii) $\therefore$ 1 mg $= 10^{-6}$ kg

$\therefore$ 25365 mg $= 25365 \times 10^{-6}$ kg

$= 2.5365 \times 10^{-2}$ kg (1 M)

8. (i) 15 ppm means 15 parts in 10^6 parts

$\therefore$ % by mass $= \frac{15 \times 100}{10^6} = 1.5 \times 10^{-3}\%$ (1 M)

(ii) Molar mass of

$CHCl_3 = 12 + 1 + 3(35.5) = 119.5$ g/mol

1.5×10^{-3} % means 1.5×10^{-3} g chloroform is found in 100 g sample.

$$\text{Molality (m)} = \frac{w \times 1000}{MM \times W_{(sol)}} = \frac{1.5 \times 10^{-3} \times 10^3}{119.5 \times 100}$$

$= 0.000125$

$= 1.25 \times 10^4$ M (2 M)

Long Answer Type Questions

1. (*a*) (*i*) Empirical formula represents simple whole number ratio of atoms of all elements present in a molecule of the compound. (1 M)

(*ii*) Molecular formula gives the actual number of atoms of various elements present in one molecule of compound. (1 M)

(*b*)

Element	%	At. Mass	Moles	Simplest Atomic Ratio	Simplest whole no. ratio
C	57.82	12	4.82	2	$2 \times 2 = 4$
H	3.6	1	3.6	1.49	$1.49 \times 2 = 3$
O	38.58	16	2.41	1	$1 \times 2 = 2$

Empirical formula $= C_4H_3O_2$

Empirical formula mass $= 12 \times 4 + 1 \times 3 + 16 \times 2 = 83$ u

Molecular mass = 166 u

Molecular formula $= (C_4H_3O_2)_2 = C_8H_6O_4$ (3 M)

Key Takeaways

Write only the definition. Avoid unnecessary explanation. Read the question carefully. Do not get confused between empirical formula and molecular formula.

2. (i)

Molarity	Molality
It is defined as the number of moles of the solute in 1 litre of the solution.	It is defined as the number of moles of solute present in 1 kg of solvent.
It is affected by the change in temperature.	It is not affected by the change in temperature.

(2 M)

(ii) Molarity = 1.35 M

Molar mass of $H_2SO_4 = 2 + 32 + 64 = 98$ g/mol

Mass of H_2SO_4 in 1 L solution $= 1.35 \times 98 = 132.3$ g

Mass of 1 L solution $= 1000 \times 1.02$

$= 1020$ g (Density $= 1.02$ g cm^{-3})

Mass of water in solution

$= 1020 - 132.3 = 887.7$ g $= 0.888$ kg

$$\text{Molality} = \frac{n_{solute}}{m_{(solvent)}\text{kg}} = \frac{1.35}{0.888} = 1.52 \text{ m}$$ (3 M)

3. (*a*) (*i*) Limiting reagent: Limiting reagent is the reactant which is completely used up first when a reaction goes to completion. (1 M)

(*ii*) Mole: It is the amount of substance which contains same number of particles (atoms, molecules or ions) as the number of atoms present in 12 g of Carbon-12. (1 M)

(*b*) $CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

(*i*) 1 mole $CaCO_3$ requires 1 mole H_2SO_4

100 g $CaCO_3$ require 98 g H_2SO_4

20 g $CaCO_3$ will require

$= \frac{98}{100} \times 20 = 19.6$ g H_2SO_4

Hence, $CaCO_3$ is the limiting reagent. (1 M)

(*ii*) 1 mole, $CaCO_3$ produces 1 mole $CaSO_4$

100 g $CaCO_3$ produce 136 g $CaSO_4$

20 g $CaCO_3$ will produce

$= \frac{136}{100} \times 20 = 27.2$ g $CaSO_4$ (1 M)

(*iii*) 1 mole $CaCO_3$ gives 1 mole CO_2

0.2 moles of $CaCO_3$ will give 0.2 moles of CO_2

$= 0.2 \times 22.4$ L

$= 4.48$ L of CO_2 (1 M)

4. Number of moles of HCl = 250 mL

Mass of $CaCO_3$ = 1000 g

Number of moles of $CaCO_3 = \frac{1000}{100} = 10$ mol

According to given equation, 1 mol of $CaCO_3$ (s) requires 2 mol of HCl (aq). Hence, for the reaction of 10 mol of $CaCO_3$ (s), number of moles of HCl required would be:

$$10 \text{ mol } CaCO_3 \times \frac{2 \text{ mol HCl}}{1 \text{ mol } CaCO_3} = 20 \text{ mol HCl}$$ (2 M)

But we have only 0.19 mol HCl (aq), hence, HCl (aq) is a limiting reagent. So amount of $CaCl_2$ formed will depend on the amount of HCl available. Since, 2 mol HCl (aq) forms 1 mol of $CaCl_2$, therefore, 0.19 mol of HCl (aq) would give:

$$0.19 \text{ mol HCl (aq)} \frac{1 \text{ mol } CaCl_2}{2 \text{ mol HCl}} = 0.095 \text{ mol}$$

or 0.095 molar mass of $CaCl_2 = 0.095 \times 111 = 10.54$ g

(3 M)

Case Based Questions

Case Based-I

(i) Graphite-Pure substance (Element). Iodized table salt-Mixture (Heterogeneous) *(1 M)*

(ii) Tap water constitutes some impurities such as dust particles which are normally mixed with it and not combined chemically. In tap water, the constituents are not present in a fixed ratio and hence, it is a mixture. Distilled water contains only water molecules since it is free from impurities, it is therefore, considered as a compound.

(1½ M)

(iii) Only the gaseous states of those substances are regarded as vapours which are liquid at room temperature. Since ammonia exists as a gas at room temperature. Hence, its gaseous state is called gas while water is a liquid at room temperature. Hence, its gaseous state is called vapours. *(1½ M)*

Case Based-II

(i) (*a*) 2.41×10^{24}

44.0 g of CO_2 contains oxygen atoms

$= 2 \times 6.022 \times 10^{23}$

Now, 88.0 g of CO_2 contains oxygen atoms

$= 2 \times 2 \times 6.022 \times 10^{23}$

$= 2.41 \times 10^{24}$ atoms *(1 M)*

(ii) (*c*) 342 g

Molecular mass of cane sugar $C_{12}H_{22}O_{11}$

$= (12 \times 12) + (22 \times 1) + (16 \times 11) = 342$ g *(1 M)*

(iii) (*b*) 40 N_A

Moles of N^{3-} ion $= \frac{70}{14} = 5$ mol

No. of valence electrons in one N^{3-} ion $= 5 + 3 = 8$

Total no. of electrons $= 5 \times 8 \times N_A$ *(1 M)*

(iv) (*c*) 14 g

22.4 L of N_2 at STP weighs = 28.0 g

11.2 L of N_2 at STP weighs $= \frac{28}{22.4} \times 11.2$

$= 14.0$ g *(1 M)*

Case Based-III

(i) (*b*) 0.004 M

$$\text{Number of moles} = \frac{\text{molecules of urea}}{\text{Avogardro's number}}$$

$$= \frac{6.02 \times 10^{20}}{6.02 \times 10^{23}} = 10^{-3}$$

Molarity = Number of moles of solute/ volume of solution

$$= \frac{10^{-3}}{250} \times 1000 = 0 .004 \text{ M}$$ *(1 M)*

(ii) (*c*) 0.26

$$\text{Mole fraction of glycol} = \frac{\text{No. of moles of glycol}}{\text{no. of moles glycol + no. of moles of water}}$$

$$= \frac{\frac{56}{62}}{\frac{56}{62} + \frac{45}{18}} = \frac{0.9}{0.9 + 2.5} = 0.26$$ *(1 M)*

(iii) (*a*) 55.55

Molality of water means number of moles of water in 1000 g of water which is $\frac{1000}{18} = 55.5$ mol *(1 M)*

(iv) (*b*) Molality does not depend upon temperature.

Molality is favored over molarity as the unit of concentration because molality is a function of temperature and changes with temperature but molarity is independent of temperature so it stays the same. The mass of the solvent is also independent of temperature so it remains constant. *(1 M)*

Case Based-IV

(i) $\underset{1 \text{ mole}}{PbO} + \underset{2 \text{ moles}}{2HCl} \rightarrow \underset{1 \text{ mole}}{PbCl_2} + \underset{2 \text{ moles}}{2H_2O}$

$$= \frac{6.5}{223.2} \text{ mol} = \frac{3.2}{36.5} \text{ mol}$$

= 0.029 mol   0.087 mol

So, PbO is the limiting reactant.

= 0.029 mol of $PbCl_2$ is formed. *(1 M)*

(ii) $H_2(g) + \frac{1}{2} O_{2(g)} \rightarrow H_2O_{(l)}$

1 mole   0.5 mole   1 mole

14 g of $H_2 = \frac{14}{2}$ mole = 7 mol

80 g of O_2

$= \frac{80}{32}$ mole = 2.5 mol

So, O_2 is the limiting reagent

Since 0.5 mole of oxygen from water = 1 mol

So, 2.5 mol of oxygen form water $= \frac{1}{0.5} \times 2.5$

= 5 mol *(2 M)*

(iii) A chemical equation has to be balanced in order to satisfy the law of conservation of mass. According to the law, there is no change in mass when the reactants change into the products. Therefore, the chemical equation has to be balanced. *(1 M)*