SAMPLE CONTENT

PERFECT

CHEMISTRY

Vol. I

Weathering of Rocks

Red-orange rock formations owe their colour to high concentration of iron(III) oxide resulted from chemical weathering of the rock.

STD. XI Sci.

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PERFECT CHEMISTRY (Vol. I)

Std. XI Sci

Special Inclusion
- **Insights**
- **Practice Numericals**

Salient Features

- Written as per the latest textbook
- Subtopic-wise segregation for powerful concept building
- Complete coverage of Textual Exercise Questions, Intext Questions, Activities and Textual Examples
- Each chapter contains:
  - **'Insights…'** interesting facts to instill curiosity about the concept
  - **'Numerical Zone'** along with **'Practice Numericals'** and **'Important Formulae'** provided to establish a solid foundation of numerical aspects in the chapter
  - **'Brain Teasers'** section for application of concepts learned in chapter
  - **'Quick Review'** of the chapter for last-minute revision
  - **'Exercise'** to provide more Theory questions, Numericals and MCQs for practice
  - **'Competitive Corner'** to give the glimpse of prominent competitive examinations [MHT-CET, NEET (UG) and JEE (Main)]
  - **'Topic Test'** at the end of each chapter for self-assessment
- Includes important features like
  - **For your knowledge**
  - **Gyan guru**
  - **Connections**
  - **NCERT Corner**
- **Smart Keys:** Multiple study techniques designed to impart holistic learning
  - **Reading between the lines**
  - **Strategy**
  - **Caution**
  - **Smart Check**
- **Q.R. codes** provide:
  - The Video/pdf links boosting conceptual retention
  - Solutions of:
    i. Practice Numericals
    ii. Additional Numericals for Practice
    iii. Competitive corner
    iv. Topic test

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PREFACE

"Everything should be made as simple as possible, but not simpler." - Albert Einstein.

Having this vision in mind, we have created **"Perfect Chemistry Vol. I, Std. XI Sci."** as per the latest textbook of the Maharashtra State Board. It focuses on not just preparing students from an examination point of view but also equipping them to understand and appreciate the beauty of the concepts in chemistry.

Every chapter in this book begins with a brief introduction to the chapter. Following with:

- **Insights…** provided at the start captivate readers with intriguing revelations and thought-provoking observations, setting the stage for an engaging exploration of each new chapter.
- The chapter is **segregated subtopic-wise** and encompasses all textual content in the format of Question and Answers. *Textual Exercise questions*, *Intext questions*, *'Can you tell'*, *'Can you recall'*, *'Try this'*, and '*Activity*' are placed aptly amongst various additional questions in accordance with the flow of the subtopic.
- **Numerical Zone** covers numericals along with their step-wise solutions using log calculation (wherever necessary) at the end of each topic, followed by **Practice Numericals** (solutions to which they are provided through a QR code), which strengthens the numerical aspect of the students.
- **Important Formulae** are placed after covering the last subtopic of the chapter.
- **Exercise** helps the students to gain insight on the various levels of theory and numerical-based questions.
- **Multiple Choice Questions** and **Topic Test** (as per the latest paper pattern) assess the students on their range of preparation and the amount of knowledge of each topic.
- **Quick Review** summarizes the key points in the chapter for last-minute revision.
- The flow chart on the adjacent page will walk you through the **key features** of the book and elucidate how they have been carefully designed to maximize the student learning.

Perfect Chemistry Vol. I, Std. XI Sci. adheres to our vision and achieves several goals: building concepts, developing competence to solve numericals, recapitulation, self-study, self-assessment, and student engagement -– all while encouraging students toward cognitive thinking.

We hope the book benefits the learner as we have envisioned.

Publisher

Edition: Fifth

The journey to create a complete book is strewn with triumphs, failures and near misses. If you think we've nearly missed something or want to applaud us for our triumphs, we'd love to hear from you.
Please write to us on: mail@targetpublications.org

KEY FEATURES

Insights...

'Insights...' *starts each chapter with engaging information and curious facts, sparking your interest right from the beginning.*

For your knowledge

For Your Knowledge presents fascinating information about the concept covered.

Competitive Corner

Competitive Corner includes selective questions from prominent [NEET (UG), JEE (Main) and MHT CET] competitive exams based entirely on the syllabus covered in the chapter.

NCERT Corner

NCERT Corner covers information from NCERT textbook relevant to topic.

Connection

Connections enable students to interlink concepts covered in different chapters.

Brain Teasers

Brain Teasers include challenging questions.

Continued…

KEY FEATURES

Smart Keys

'Smart Keys' comprise a set of remarkable study techniques contrived to benefit students.

Reading between the lines

Reading between the lines provides elaboration or missing fragments of concept which is essential for complete understanding of the concept.

Smart check

Smart Check is a technique to verify the answers. This is our attempt to cross-check the accuracy of the answer.

Caution

Caution helps students to be watchful against commonly made mistakes.

Strategy

Strategy provides a step-by-step process to break a complex numerical problem into simpler parts.

QR Codes

QR codes provide:

i. Access to a video/PDF in order to boost understanding of a concept or activity

ii. Solutions to:
Practice Numericals,
Additional Numericals for Practice,
Competitive corner and Topic test

Important Formulae

Important Formulae includes all of the key formulae in the chapter.

Quick Review

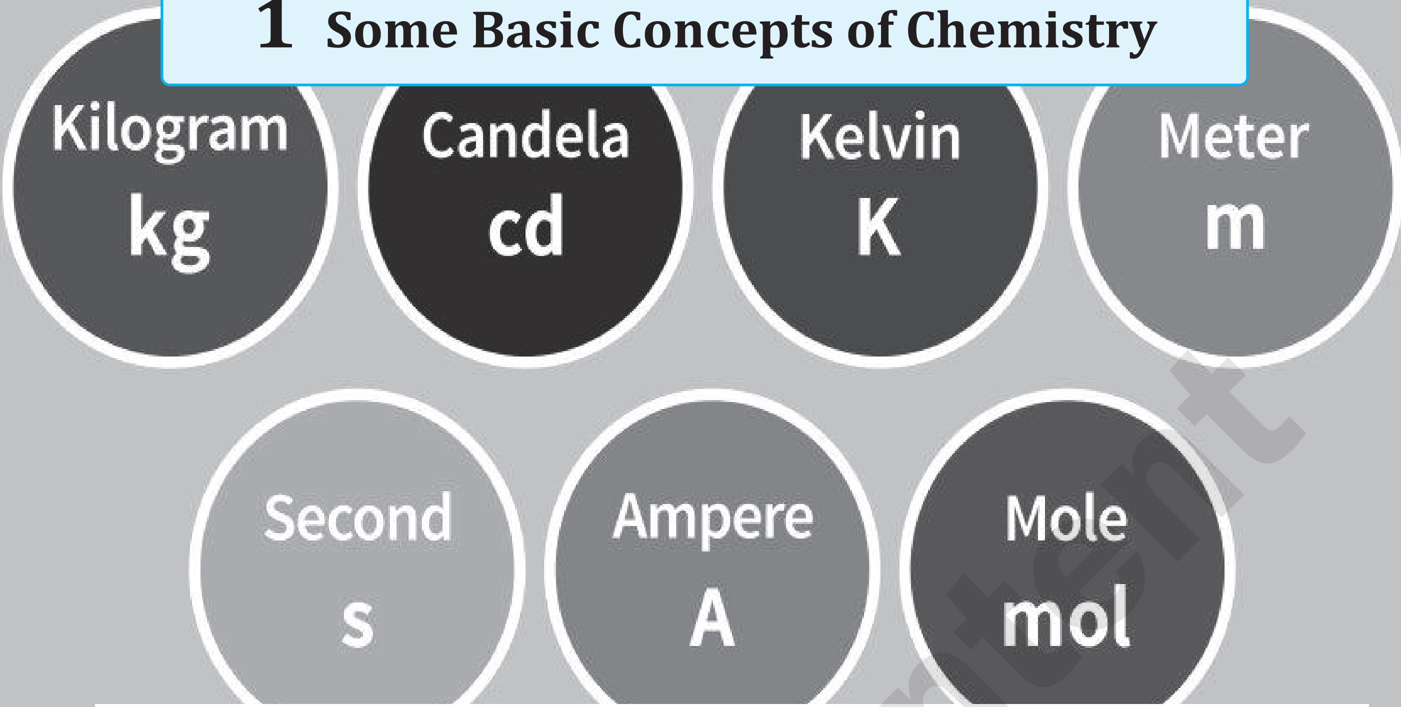
Quick review includes tables / flow chart to summarize the key points in the chapter.

CONTENTS

[Reference: Maharashtra State Board of Secondary and Higher Secondary Education, Pune - 04**]**

Note:
1. * mark represents Textual Exercise question.
2. # mark represents Intext question.
3. + mark represents Textual examples and Numericals.
4. symbol represents textual questions that need external reference for an answer.
5. Chapters 10 to 16 are a part of Perfect Chemistry Vol. II, Std. XI Sci.

1 Some Basic Concepts of Chemistry



This chapter introduces students to the properties and measurement of matter, SI units, various laws of chemical combination, atomic and molecular masses. The chapter also explains an important concept of chemistry – the mole, which is the basic unit used to measure the quantity of a chemical substance. Students should emphasize on solving the numericals in the chapter.
*This chapter is allotted weightage of **5 marks** with option and **3 marks** without option.*

Contents and Concepts

Insights...

1. You would experience weightlessness if you were in the centre of the Earth or in the vacuum because gravitation acceleration (g) is zero.
2. Prefixes are used to modify the size of a unit because the SI units are not always convenient. For example, it is inconvenient to express the mass of a pin in kilograms.
3. Keep in mind that 32 °F is equivalent to 0 °C, and 212 °F is equivalent to 100 °C. Hence, every one degree rise in Celsius scale corresponds 9/5 degree rise in Fahrenheit. The formula that results is °F = 9/5 °C + 32.
4. Atomic masses are usually average values since most elements exist in nature as mixtures of isotopes.
5. For stoichiometric purposes, we can treat carbon as being made of a single type of atom with mass 12.00000, even though natural carbon does not contain a single atom with such mass.
6. Chemists rarely work with single atoms or molecules in our macroscopic world because they are too small to handle with ease. Instead, they usually work with quantities that are large enough to view and handle comfortably. The SI unit 'mole' is used to express quantities of substances.
7. The molar volume depends on the temperature and the pressure, but it is independent of the nature of gas.

Questions and Answers

1.1 Introduction

Q.1. Define chemistry. **[1 Mark]**

Ans: ***Chemistry*** *is the study of matter, its physical and chemical properties and the physical and chemical changes it undergoes under different conditions.*

Q.2. Why is chemistry called a central science? **[2 Marks]**

Ans:

i. Knowledge of chemistry is required in the studies of physics, biological sciences, applied sciences, and earth and space sciences.

ii. Chemistry is involved in every aspect of day-to-day life, i.e. the air we breathe, the food we eat, the fluids we drink, our clothing, transportation and fuel supplies, etc.
Hence, chemistry is called a central science.

Q.3. Give reason: Although chemistry has ancient roots, it has developed as a modern science. **[2 Marks]**

Ans: Technological development in sophisticated instruments have expanded knowledge of chemistry which, now, has been used in applied sciences such as medicine, dentistry, engineering, agriculture and in daily home use products. Hence, due to development and advancement in science and technology, chemistry has developed as modern science.

1.2 Nature of Chemistry

Q.4. How is chemistry traditionally classified? **[2 Marks]**

Ans: Chemistry is traditionally classified into five branches:

i. Organic chemistry
ii. Inorganic chemistry
iii. Physical chemistry
iv. Biochemistry
v. Analytical chemistry

Q.5. Explain the following terms: **[3 Marks]**

i. Organic chemistry
ii. Inorganic chemistry
iii. Physical chemistry

Ans:

i. **Organic chemistry:** It deals with properties and reactions of compounds of carbon.

ii. **Inorganic chemistry:** It deals with the study of all the compounds which are not organic.

iii. **Physical chemistry:** It deals with the study of properties of matter, the energy changes and the theories, laws and principles that explain the transformation of matter from one form to another. It also provides basic framework for all the other branches of chemistry.

***Q.6. Explain: Types of matter (on the basis of chemical composition)** **[4 Marks]**

Ans: Matter on the basis of chemical composition can be classified as follows:

i. **Pure substances:** They always have a definite chemical composition. They always have the same properties regardless of their origin.
e.g. Pure metal, distilled water, etc.
They are of two types:

a. **Elements:** They are pure substances, which cannot be broken down into simpler substances by ordinary chemical changes.
Elements are further classified into three types:

1. Metals:
i. They have a lustre (a shiny appearance).
ii. They conduct heat and electricity.
iii. They can be drawn into wire (ductile).
iv. They can be hammered into thin sheets (malleable).
e.g. Gold, silver, copper, iron. Mercury is a liquid metal at room temperature.

2. Nonmetals:
i. They have no lustre. (except diamond, iodine)
ii. They are poor conductors of heat and electricity. (except graphite)
iii. They cannot be hammered into sheets or drawn into wire, because they are brittle.
e.g. Iodine

3. Metalloids: Some elements have properties that are intermediate between metals and nonmetals and are called metalloids or semimetals. **e.g.** Arsenic, silicon and germanium.

b. **Compounds:** They are the pure substances which are made up of two or more elements in fixed proportion.
e.g. Water, ammonia, methane, etc.

ii. **Mixtures**: They have no definite chemical composition and hence no definite properties. They can be separated by physical methods.
e.g. Paint (mixture of oils, pigment, additive), concrete (a mixture of sand, cement, water), etc.
Mixtures are of two types:

a. **Homogeneous mixture:** In homogeneous mixture, constituents remain uniformly mixed throughout its bulk.
e.g. Solution, in which solute and solvent molecules are uniformly mixed throughout its bulk.

b. **Heterogeneous mixture:** In heterogeneous mixture, constituents are not uniformly mixed throughout its bulk.
e.g. Suspension, which contains insoluble solid in a liquid.

Q.7. Can you tell? *(Textbook page no. 1)*
Which are mixtures and pure substances from the following? **[½ Mark Each]**

i.	**Sea water**	**ii.**	**Gasoline**
iii.	**Skin**	**iv.**	**A rusty nail**
v.	**A page of textbook**	**vi.**	**Diamond**

Ans:

No.	Material	Pure substance or mixture
i.	Sea water	Mixture
ii.	Gasoline	Mixture
iii.	Skin	Mixture
iv.	A rusty nail	Mixture
v.	A page of textbook	Mixture
vi.	Diamond	Pure substance

Q.8. Can you tell? *(Textbook page no. 2)*
Classify the following as element and compound. **[½ Mark Each]**

i.	**Mercuric oxide**	**ii.**	**Helium gas**
iii.	**Water**	**iv.**	**Table salt**
v.	**Iodine**	**vi.**	**Mercury**
vii.	**Oxygen**	**viii.**	**Nitrogen**

Ans:

No.	Material	Element or compound
i.	Mercuric oxide	Compound
ii.	Helium gas	Element
iii.	Water	Compound
iv.	Table salt	Compound
v.	Iodine	Element
vi.	Mercury	Element
vii.	Oxygen	Element
viii.	Nitrogen	Element

***Q.9. Give one example of each:** **[½ Mark Each]**
i. Homogeneous mixture
ii. Heterogeneous mixture
iii. Element
iv. Compound

Ans:
i. Homogeneous mixture: Solution (An aqueous solution of sugar)
ii. Heterogeneous mixture: Suspension (of sand in water)
iii. Element: Gold
iv. Compound: Distilled water

Q.10. Distinguish between: **[1 Mark Each]**
i. Mixtures and pure substances
ii. Mixtures and compounds

Ans:
i.

	Mixtures	Pure substances
a.	Mixtures have no definite chemical composition.	Pure substances have a definite chemical composition.
b.	Mixtures have no definite properties.	Pure substances always have the same properties regardless of their origin.
e.g.	Paint (mixture of oils, pigment, additive), concrete (a mixture of sand, cement, water), etc.	Pure metal, distilled water, etc.

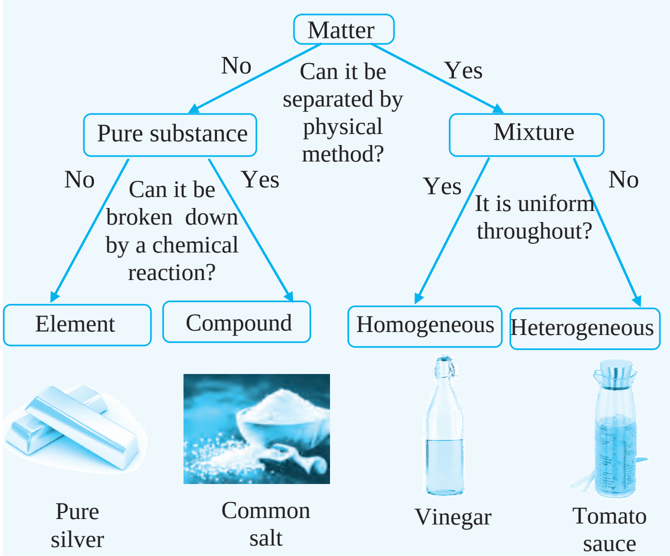
ii.

	Mixtures	Compounds
a.	Mixtures have no definite chemical composition.	Compounds are made up of two or more elements in fixed proportion.
b.	The constituents of a mixture can be easily separated by physical method.	The constituents of a compound cannot be easily separated by physical method.
e.g.	Paint (mixture of oils, pigment, additive), concrete (a mixture of sand, cement, water), etc.	Water, table salt, sugar, etc.

Q.11. What is the difference between element and compound? **[1 Mark]**

Ans: Elements cannot be broken down into simpler substances while compounds can be broken down into simpler substances by chemical changes.

FOR YOUR KNOWLEDGE



Q.12. Explain: States of matter **[2 Marks]**

Ans: There are three different states of matter as follows:

i. **Solid:** Particles are held tightly in perfect order. They have definite shape and volume.

ii. **Liquid:** Particles are close to each other but can move around within the liquid.

iii. **Gas:** Particles are far apart as compared to that of solid and liquid.

These three states of matter can be interconverted by changing the conditions of temperature and pressure.

1.3 Properties of Matter and Their Measurement

Q.13. Explain: Physical and chemical properties **[2 Marks]**

Ans:

i. **Physical properties:** These are properties which can be measured or observed without changing the identity or the composition of the substance.

e.g. Colour, odour, melting point, boiling point, density, etc.

ii. **Chemical properties:** These are properties in which substances undergo change in chemical composition.

e.g. Coal burns in air to produce carbon dioxide, magnesium wire burns in air in the presence of oxygen to form magnesium oxide, etc.

Q.14. How are properties of matter measured? **[2 Marks]**

Ans:

i. Measurement involves comparing a property of matter with some fixed standard which is reproducible and unchanging.

ii. Properties such as mass, length, area, volume, time, etc. are quantitative in nature and can be measured.

iii. A quantitative measurement is represented by a number followed by units in which it is measured.

iv. These units are arbitrarily chosen on the basis of universally accepted standards.

e.g. Length of class room can be expressed as 10 m. Here, 10 is the number and 'm' is the unit 'metre' in which the length is measured.

Q.15. Define: Units **[1 Mark]**

Ans: *The arbitrarily decided and universally accepted standards are called* ***units****.*

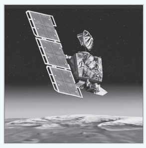
e.g. Metre (m), kilogram (kg).

Q.16. What are the various systems in which units are expressed? **[1 Mark]**

Ans: Units are expressed in various systems like CGS (centimetre for length, gram for mass and second for time), FPS (foot, pound, second) and MKS (metre, kilogram, second) systems, etc.

GG - GYAN GURU

Why are units important?



During calculations, confinement to one single system of unit is advisable. NASA's Mars climate orbiter (first weather satellite for mars) was destroyed due to heat. The mission failed as there was a confusion while estimating the distance between earth and mars in miles and kilometres.

Q.17. What are SI units? Name the fundamental SI units. **[3 Marks]**

Ans: **SI Units:** In 1960, the general conference of weights and measures proposed revised metric system, called International system of Units i.e. SI system (abbreviated from its French name).

The seven fundamental SI units are as given below:

No.	Base physical quantity	SI unit	Symbol
i.	Length	Metre	m
ii.	Mass	Kilogram	kg
iii.	Time	Second	s
iv.	Temperature	Kelvin	K
v.	Amount of substance	Mole	mol
vi.	Electric current	Ampere	A
vii.	Luminous intensity	Candela	cd

Note: Units for other quantities such as speed, volume, density, etc. can be derived from fundamental SI units.

*__Q.18. What is the SI unit of amount of a substance?__ **[1 Mark]**

Ans: The SI unit for the amount of a substance is mole (mol).

Q.19. What is the basic unit of mass in the SI system? **[1 Mark]**

Ans: The basic unit of mass in the SI system is kilogram (kg).

Q.20. Name the following: **[1 Mark Each]**

i. **Full form of CGS unit system**

ii. **Full form of FPS unit system**

iii. **The SI unit of length**

iv. **Symbol used for Candela unit**

v. **SI unit of temperature**

vi. **SI unit of electric current**

Ans:

i. Centimetre Gram Second
ii. Foot Pound Second
iii. Metre (m)
iv. Cd
v. Kelvin (K)
vi. Ampere (A)

NCERT CORNER

Prefixes Used In The SI System:

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

Q.21. Give reason: The mass of a body is more fundamental property than its weight. **[2 Marks]**

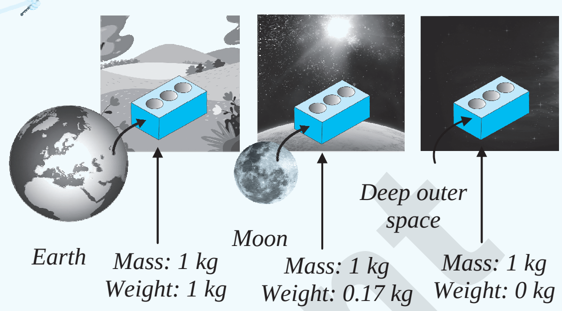
Ans:

i. Mass is an inherent property of matter and is the measure of the quantity of matter of a body.
ii. The mass of a body does not vary with respect to its position.
iii. On the other hand, the weight of a body is a result of the mass and gravitational attraction
iv. Weight varies because the gravitational attraction of the earth for a body varies with the distance from the centre of the earth.

Hence, the mass of a body is more fundamental property than its weight.

GG - GYAN GURU

Weight of an object in deep outer space is ZERO!!



The mass of a body remains constant irrespective its position. However, the weight of a body depends on its position. There is less gravitational pull on moon as compared to earth. Hence, an object will have smaller weight on moon as compared to earth. There is no gravitational force in deep outer space and, so weight is ZERO!!

Q.22. How is gram related to the SI unit kilogram? **[1 Mark]**

Ans: The SI unit kilogram (kg) is related to gram (g) as 1 kg = 1000 g = 10^3 g.

Note: 'Gram' is used for weighing small quantities of chemicals in the laboratories.
Other commonly used quantity is ‘milligram’.
1 kg = 1000 g = 10^6 mg

Q.23. Why are fractional units of the SI units of length often used? Give two examples of the fractional units of length. How are they related to the SI unit of length? **[3 Marks]**

Ans:

i. Some properties such as the atomic radius, bond length, wavelength of electromagnetic radiation, etc. are very small and therefore, fractional units of the SI unit of length are often used to express these properties.
ii. Fractional units of length: Nanometre (nm), picometre (pm), etc.
iii. Nanometre (nm) and picometre (pm) are related to the SI unit of length (m) as follows:
1 nm = 10^{-9} m, 1 pm = 10^{-12} m

Q.24. Define: Volume **[1 Mark]**

Ans: ***Volume*** *is the amount of space occupied by a three-dimensional object.* It does not depend on shape.

Q.25. State the common unit used for the measurement of volume of liquids and gases. **[1 Mark]**

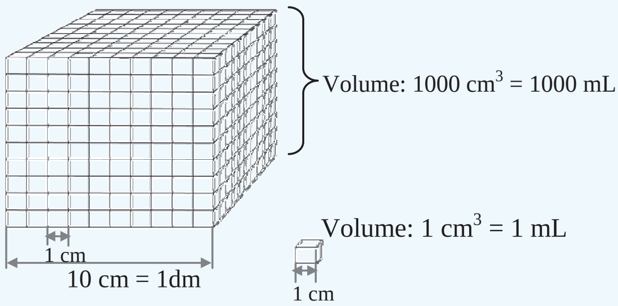
Ans: The common unit used for the measurement of volume of liquids and gases is litre (L).

Q.26. How is the SI unit of volume expressed? **[1 Mark]**

Ans: The SI unit of volume is expressed as (metre)^3 or m^3.

FOR YOUR KNOWLEDGE

The other units used to express volume are dm^3, cm^3, mL, etc. These units are related as follows:
1 L = 1 dm^3 = 1000 mL = 1000 cm^3
1000 cm^3 = 10 cm × 10 cm × 10 cm of volume



Q.27. Name some glassware that are used to measure the volume of liquids and solutions. **[1 Mark]**

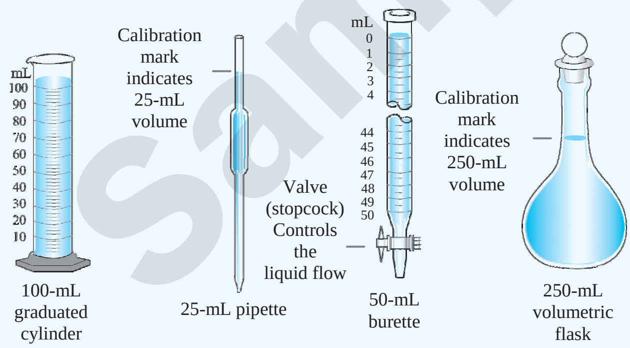
Ans:
i. Graduated cylinder
ii. Burette
iii. Pipette

Q.28. What is a volumetric flask used for in laboratory? **[1 Mark]**

Ans: A volumetric flask is used to prepare a known volume of a solution in laboratory.

FOR YOUR KNOWLEDGE

The calibration of volumetric glass apparatus is shown in the figures given.



Q.29. What is density of a substance? How is it measured? **[2 Marks]**

Ans: Density:

i. ***Density of a substance*** *is its mass per unit volume.* It is the characteristic property of any substance.

ii. It is determined in the laboratory by measuring both the mass and the volume of a sample.

iii. The density is calculated by dividing mass by volume.

Q.30. How is the SI unit of density derived? State CGS unit of density. **[2 Marks]**

Ans:
i. The SI unit of density is derived as follows:
ii. CGS unit of density: g cm^{-3}

$$\text{Density} = \frac{\text{SI unit mass}}{\text{SI unit volume}} = \frac{\text{kg}}{\text{m}^3} = \text{kg m}^{-3}$$

Note: The CGS unit, g cm^{-3} is equivalent to $\frac{\text{g}}{\text{mL}}$ or g mL^{-1}.

Q.31. State three common scales of temperature measurement. **[1 Mark]**

Ans:
i. Degree Celsius (°C)
ii. Degree Fahrenheit (°F)
iii. Kelvin (K)

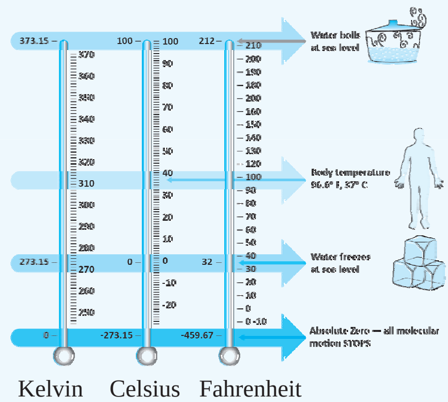
Q.32. State the temperatures in Fahrenheit scale that corresponds to 0 °C and 100 °C. **[1 Mark]**

Ans: The temperature that corresponds to 0 °C is 32 °F and the temperature that corresponds to 100 °C is 212 °F.

FOR YOUR KNOWLEDGE

Thermometers of different temperature scales:

Generally, the thermometer with Celsius scale are calibrated from 0 °C to 100 °C where these two temperatures are respectively the freezing point and the boiling point of water at atmospheric pressure. Human body temperature is 37 °C.



Q.33. Write the expression showing the relationship between: **[2 Marks]**
i. Degree Fahrenheit and Degree Celsius
ii. Kelvin and Degree Celsius

Ans:
i. The relationship between degree Fahrenheit and degree Celsius is expressed as, $°\text{F} = \frac{9}{5}(°\text{C}) + 32$

ii. The relationship between Kelvin and degree Celsius is expressed as, K = °C + 273.15

Numerical Zone

***Q.34. Convert the following degree Celsius temperature to degree Fahrenheit.** **[2 Marks]**

i. **40 °C** **ii.** **30 °C**

Solution:

i.

Given: Temperature in degree Celsius = 40 °C

To find: Temperature in degree Fahrenheit

Formula: $°F = \frac{9}{5}\ (°C) + 32$

Calculation: Substituting 40 °C in the formula,

$$°F = \frac{9}{5}\ (°C) + 32 = \frac{9}{5}(40) + 32$$

$$= 72 + 32 = \mathbf{104\ °F}$$

ii.

Given: Temperature in degree Celsius = 30 °C

To find: Temperature in degree Fahrenheit

Formula: $°F = \frac{9}{5}(°C) + 32$

Calculation: Substituting 30 °C in the formula,

$$F = \frac{9}{5}\ (°C) + 32 = \frac{9}{5}(30) + 32$$

$$= 54 + 32 = \mathbf{86\ °F}$$

Ans: i. The temperature 40 °C corresponds to **104 °F**

ii. The temperature 30 °C corresponds to **86 °F**

Smart Check

Convert your answer back to the original temperature to see whether it matches.

i.

$$°F = \frac{9}{5}\ (°C) + 32$$

$$104 = \frac{9}{5}\ (°C) + 32$$

$$\frac{9}{5}\ (°C) = 104 - 32$$

$$°C = \frac{(104-32)\times 5}{9}$$

$$°C = \frac{72\times 5}{9} = 40\ °C$$

ii.

$$°F = \frac{9}{5}\ (°C) + 32$$

$$104 = \frac{9}{5}\ (°C) + 32$$

$$\frac{9}{5}\ (°C) = 86 - 32$$

$$°C = \frac{(86-32)\times 5}{9}$$

$$°C = \frac{54\times 5}{9} = 30\ °C$$

Q.35. Convert the following degree Fahrenheit temperature to degree Celsius. **[2 Marks]**

i. **50 °F** **ii.** **10 °F**

Solution:

i.

Given: Temperature in degree Fahrenheit = 50 °F

To find: Temperature in degree Celsius

Formula: $°F = \frac{9}{5}\ (°C) + 32$

Calculation: Substituting 50 °F in the formula,

$$°F = \frac{9}{5}\ (°C) + 32$$

$$50 = \frac{9}{5}\ (°C) + 32$$

$$°C = \frac{(50\ -\ 32)\ \times\ 5}{9} = \mathbf{10\ °C}$$

ii.

Given: Temperature in degree Fahrenheit = 10 °F

To find: Temperature in degree Celsius

Formula: $°F = \frac{9}{5}\ (°C) + 32$

Calculation: Substituting 10 °F in the formula,

$$°F = \frac{9}{5}\ (°C) + 32$$

$$10 = \frac{9}{5}\ (°C) + 32$$

$$°C = \frac{(10\ -\ 32)\ \times\ 5}{9} = \mathbf{-12.2\ °C}$$

Ans: i. The temperature 50 °F corresponds to **10 °C**.

ii. The temperature 10 °F corresponds to **–12.2 °C**.

Smart Check

Convert your answer back to the original temperature to see whether it matches.

i.

$$°F = \frac{9}{5}\ (°C) + 32$$

$$= \frac{9}{5}\ (10) + 32$$

$$= 18 + 32$$

$$= 50\ °F$$

ii.

$$°F = \frac{9}{5}\ (°C) + 32$$

$$= \frac{9}{5}\ (-12.2) + 32$$

$$= -21.96 + 32$$

$$\approx 10\ °F$$

Practice Numericals

1. A person with a fever has a temperature of 102 °F. What is this temperature in degrees Celsius? **[1 Mark]**

Ans: 38.9 °C

2. A mixture of dry ice and isopropyl alcohol has a temperature of –78 °C. What is this temperature in Fahrenheit? **[1 Mark]**

Ans: 108.4 °F

1.4 Laws of Chemical Combination

Q.36. What is meant by the term 'chemical combination'? **[1 Mark]**

Ans: The process in which the elements combine with each other to form compounds is called **chemical combination**.

Note: The process of chemical combination is governed by five basic laws which were discovered before the knowledge of molecular formulae.

***Q.37. State and explain the law of conservation of mass.** **[3 Marks]**

Ans: Law of conservation of mass:

i. The law of conservation of mass states that, *"Mass can neither be created nor destroyed"* during chemical combination of matter.

ii. Antoine Lavoisier performed many combustion experiments, namely burning of phosphorus and mercury in the presence of air. Both his experiments resulted in increased weight of products.

iii. After several experiments, in burning of phosphorus, he found that the weight gained by the phosphorus was exactly the same as the weight lost by the air. Hence, total mass of reactants = total mass of products.

iv. When hydrogen gas burns and combines with oxygen to form water, the mass of the water formed is equal to the mass of the hydrogen and oxygen consumed. Thus, this is in accordance with the law of conservation of mass.

Q.38. State and explain the law of definite proportions. **[3 Marks]**

Ans: Law of definite proportions:

i. The law states that *"A given compound always contains exactly the same proportion of elements by weight"*.

ii. French chemist, Joseph Proust worked with two samples of cupric carbonate; one of which was naturally occurring cupric carbonate and other was synthetic sample. He found the composition of elements present in both the samples was same as shown below:

Cupric carbonate	% of copper	% of carbon	% of oxygen
Natural sample	51.35	9.74	38.91
Synthetic sample	51.35	9.74	38.91

iii. Thus, irrespective of the source, a given compound always contains same elements in the same proportion.

READING BETWEEN THE LINES

*The validity of this law has been further supported by various experiments. This law is often called as **Law of definite composition**.*

FOR YOUR KNOWLEDGE

The law of definite composition is not true for all types of compounds. It is true for only those compounds which are obtained from one type of isotope.

e.g. Carbon exists in two common isotopes: ^{12}C and ^{14}C. When it forms $^{12}CO_2$, the ratio of masses is 12:32 or 3:8. However, when it is formed from ^{14}C i.e., $^{14}CO_2$, the ratio will be 14:32 i.e., 7:16, which is not same as in the first case.

***Q.39. State the law of multiple proportions.** **[1 Mark]**

Ans: The law states that, *"When two elements A and B form more than one compound, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers"*.

Q.40. State and explain the law of multiple proportions. **[3 Marks]**

Ans: Law of multiple proportions:

i. John Dalton (British scientist) proposed the law of multiple proportions in 1803.

ii. It has been observed that two or more elements may combine to form more than one compound.

iii. The law states that, *"When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers"*.

e.g. Hydrogen and oxygen combine to form two compounds, water and hydrogen peroxide.

$$\underset{2\text{ g}}{\text{Hydrogen}} + \underset{16\text{ g}}{\text{Oxygen}} \longrightarrow \underset{18\text{ g}}{\text{Water}}$$

$$\underset{2\text{ g}}{\text{Hydrogen}} + \underset{32\text{ g}}{\text{Oxygen}} \longrightarrow \underset{34\text{ g}}{\text{Hydrogen peroxide}}$$

Here, the two masses of oxygen (16 g and 32 g) which combine with the fixed mass of hydrogen (2 g) in these two compounds bear a simple ratio of small whole numbers, i.e. 16:32 or 1:2.

Q.41. Show that NO and NO_2 satisfy the law of multiple proportions. **[2 Marks]**

Ans: Nitrogen and oxygen combine to form two compounds, nitric oxide (NO) and nitrogen dioxide (NO_2).

$$\underset{14\text{ g}}{\text{Nitrogen}} + \underset{16\text{ g}}{\text{Oxygen}} \longrightarrow \underset{30\text{ g}}{\text{Nitric oxide}}$$

$$\underset{14\text{ g}}{\text{Nitrogen}} + \underset{32\text{ g}}{\text{Oxygen}} \longrightarrow \underset{46\text{ g}}{\text{Nitrogen dioxide}}$$

Here, the two masses of oxygen (16 g and 32 g) which combine with the fixed mass of nitrogen (14 g) in these two compounds bear a simple ratio of small whole numbers, i.e. 16:32 or 1:2.

This is in accordance with the law of multiple proportions.

Q.42. Show that carbon monoxide and carbon dioxide satisfy the law of multiple proportions. **[2 Marks]**

Ans: Chemical reaction of carbon with oxygen gives two compounds, carbon monoxide (CO) and carbon dioxide (CO_2).

$$\underset{12\text{ g}}{\text{Carbon}} + \underset{16\text{ g}}{\text{Oxygen}} \longrightarrow \underset{28\text{ g}}{\text{Carbon monoxide}}$$

$$\underset{12\text{ g}}{\text{Carbon}} + \underset{32\text{ g}}{\text{Oxygen}} \longrightarrow \underset{44\text{ g}}{\text{Carbon dioxide}}$$

Here, the two masses of oxygen (16 g and 32 g) which combine with the fixed mass of carbon (12 g) in these two compounds bear a simple ratio of small whole numbers, i.e. 16:32 or 1:2.

This is in accordance with the law of multiple proportions.

Q.43. Show that SO_2 and SO_3 satisfy the law of multiple proportions. **[2 Marks]**

Ans: Chemical reaction of sulphur with oxygen gives two compounds, sulphur dioxide (SO_2) and sulphur trioxide (SO_3).

$$\underset{32\text{g}}{\text{Sulphur}} + \underset{32\text{ g}}{\text{Oxygen}} \longrightarrow \underset{64\text{ g}}{\text{Sulphur dioxide}}$$

$$\underset{32\text{g}}{\text{Sulphur}} + \underset{48\text{ g}}{\text{Oxygen}} \longrightarrow \underset{80\text{ g}}{\text{Sulphur trioxide}}$$

Here, the two masses of oxygen (32 g and 48 g) which combine with the fixed mass of sulphur (32 g) in these two compounds bear a simple ratio of small whole numbers, i.e. 32:48 or 2:3.

This is in accordance with the law of multiple proportions.

Q.44. State and explain Gay Lussac's law of gaseous volume with two examples. **[3 Marks]**

Ans: **Gay Lussac's law:**

i. Gay Lussac proposed the law of gaseous volume in 1808.

ii. Gay Lussac's law states that, "*When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all gases are at same temperature and pressure*".

e.g.

a. Under identical conditions of temperature and pressure, 100 mL of hydrogen gas combine with 50 mL of oxygen gas to produce 100 mL of water vapour.

Thus, the simple ratio of volumes is 2:1:2.

$$\underset{\substack{[100\text{ mL}]\\ [2\text{ vol}]}}{\text{Hydrogen}_{(g)}} + \underset{\substack{[50\text{ mL}]\\ [1\text{ vol}]}}{\text{Oxygen}_{(g)}} \longrightarrow \underset{\substack{[100\text{ mL}]\\ [2\text{ vol}]}}{\text{Water}_{(g)}}$$

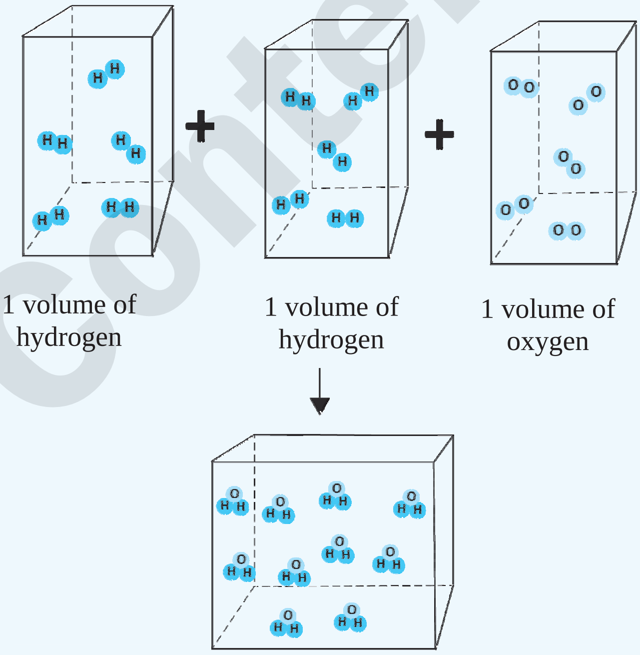
b. Under identical conditions of temperature and pressure, 1 L of nitrogen gas combine with 3 L of hydrogen gas to produce 2 L of ammonia gas. Thus, the simple ratio of volumes is 1:3:2.

$$\underset{\substack{[1\text{ L}]\\ [1\text{ vol}]}}{\text{Nitrogen}_{(g)}} + \underset{\substack{[3\text{ L}]\\ [3\text{ vol}]}}{\text{Hydrogen}_{(g)}} \longrightarrow \underset{\substack{[2\text{ L}]\\ [2\text{ vol}]}}{\text{Ammonia}_{(g)}}$$

FOR YOUR KNOWLEDGE

Gay Lussac's discovery of integer ratio in volume relationship is actually the law of definite proportion by gaseous volumes. Diagrammatic representation of Gay Lussac's law of gaseous volume is as shown below:

(Textbook page no. 6)



2 volumes of water

Q.45. Can you tell? *(Textbook page no. 6)*

If 10 volumes of dihydrogen gas react with 5 volumes of dioxygen gas, how many volumes of water vapour would be produced? **[1 Mark]**

Ans:
$$\underset{[2\text{ vol}]}{2H_{2(g)}} + \underset{[1\text{ vol}]}{O_{2(g)}} \longrightarrow \underset{[2\text{ vol}]}{2H_2O_{(g)}}$$

If 10 volumes of dihydrogen gas react with 5 volumes of dioxygen gas, then 10 volumes of water vapour would be produced.

Q.46. Give two examples which support the Gay Lussac's law of gaseous volume. **[2 Marks]**

Ans:

i. Under identical conditions of temperature and pressure, 1 L of hydrogen gas reacts with 1 L of chlorine gas to produce 2 L of hydrogen chloride gas.

Thus, the ratio of volumes is 1:1:2

This is in accordance with Gay Lussac's law.

Hydrogen +	Chlorine $\longrightarrow$	Hydrogen chloride
[1 L]	[1 L]	[2 L]
[1 vol]	[1 vol]	[2 vol]

ii. Under identical conditions of temperature and pressure, 200 mL sulphur dioxide combine with 100 mL oxygen to form 200 mL sulphur trioxide. Thus, the ratio of volumes is 2:1:2.
This is in accordance with Gay Lussac's law.

Sulphur dioxide +	Oxygen $\longrightarrow$	Sulphur trioxide
[200 mL]	[100 mL]	[200 mL]
[2 vol]	[1 vol]	[2 vol]

FOR YOUR KNOWLEDGE

i. Gay Lussac's law of combining volumes is applicable only to reactions involving gases and not to solids and liquids.

ii. *The volumes of gases in the chemical reaction are not additive.* For example, in case of reaction between hydrogen and chlorine gases it appears to be additive. However, in case of reaction between sulphur dioxide and oxygen, 2 volumes of sulphur dioxide and 1 volume of oxygen, that is, total 3 volumes of reactants get converted into 2 volumes of the product, sulphur trioxide.

iii. Similarly, in case of formation of ammonia, 1 volume of nitrogen and three volumes of hydrogen, that is, total 4 volumes of reactants, react to get converted into 2 volumes of the product, ammonia.

Numerical Zone

***Q.47. 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 verified by these data?** **[3 Marks]**

Solution:

Here, metal oxide is obtained by two different methods; reactions of metal with oxygen and reaction of metal with water vapour (steam).

In first reaction (reaction with oxygen), the mass of oxygen in metal oxide = 3.2 − 2.0 = 1.2 g

$$\% \text{ of oxygen} = \frac{1.2}{3.2} \times 100 = 37.5\% \ ;$$

$$\% \text{ of metal} = \frac{2.0}{3.2} \times 100 = 62.5\%$$

In second reaction (reaction with steam), the mass of oxygen in metal oxide = 2.27 − 1.42
= 0.85 g

$$\% \text{ of oxygen} = \frac{0.85}{2.27} \times 100 = 37.44 \approx 37.5\% \ ;$$

$$\% \text{ of metal} = \frac{1.42}{2.27} \times 100 = 62.56 \approx 62.5\%$$

Therefore, irrespective of the source, the given compound contains same elements in the same proportion. The law of definite proportions states that *"A given compound always contains exactly the same proportion of elements by weight"*. Hence, the **law of definite proportions** is verified by these data.

Ans: The **law of definite proportions** is verified by given data.

***Q.48. 24 g of carbon reacts with some oxygen to make 88 grams of carbon dioxide. Find out how much oxygen must have been used.** **[2 Marks]**

Solution:

Given: Mass of carbon (reactant) = 24 g, mass of carbon dioxide (product) = 88 g

To find: Mass of oxygen (reactant)

Calc ulation: 12 g of carbon combine with 32 g oxygen to form 44 g of carbon dioxide as follows:

Carbon +	Oxygen $\longrightarrow$	Carbon dioxide
12 g	32 g	44 g

Hence, (2 × 12 = 24 g) of carbon will combine with (2 × 32 = **64 g**) of oxygen to give (2 × 44 = 88 g) carbon dioxide.

Ans: Mass of oxygen used = **64 g**

SMART CHECK

You can check the answer by addition; the sum of the masses of carbon and oxygen, 24 + 64 grams must equal the mass of the carbon dioxide, 88 grams.

Q.49. 32 g of oxygen reacts with some carbon to make 56 grams of carbon monoxide. Find out how much mass must have been used. **[2 Marks]**

Solution:

Given: Mass of oxygen (reactant) = 32 g, mass of carbon monoxide (product) = 56 g

To find: Mass of oxygen (reactant)

Calculation: 12 g of carbon combine with 16 g oxygen to form 28 g of carbon monoxide as follows:

Carbon +	Oxygen $\longrightarrow$	Carbon monoxide
12 g	16 g	28 g

Hence, (2 × 12 = **24 g)** of carbon will combine with (2 × 16 = 32 g) of oxygen to give (2 × 28 = 56 g) carbon monoxide.

Ans: Mass of carbon used = **24 g**

SMART CHECK

You can check the answer by addition; the sum of the masses of carbon and oxygen, 24 + 32 grams must equal the mass of the carbon monoxide, 56 grams.

***Q.50. Calculate the mass of sulphur dioxide produced by burning 16 g of sulphur in excess of oxygen in contact process. (Average atomic mass: S = 32 u, O = 16 u).** [2 Marks]

Solution:

Given: Mass of sulphur (reactant) = 16 g

To find: Mass of sulphur dioxide (product)

Calculation: 32 g of sulphur combine with 32 g oxygen to form 64 g of sulphur dioxide as follows:

Sulphur	+	Oxygen	$\longrightarrow$	Sulphur dioxide
32 g		32 g		64 g

Hence, (0.5 × 32 = 16 g) of sulphur will combine with (0.5 × 32 = 16 g) of oxygen to give (0.5 × 64 = **32 g) sulphur dioxide**.

Ans: Mass of sulphur dioxide produced = **32 g**

SMART CHECK

The ratio of masses of sulphur and oxygen is 1 : 1 in sulphur dioxide. Hence, 16 g sulphur combines with 16 g oxygen to give 32 g sulphur dioxide.

Q.51. Calculate the mass of sulphur trioxide produced by burning 64 g of sulphur in excess of oxygen. (Average atomic mass: S = 32 u, O = 16 u). [2 Marks]

Solution:

Given: Mass of sulphur (reactant) = 64 g

To find: Mass of sulphur trioxide (product)

Calculation: 32 g of sulphur combine with 48 g oxygen to form 80 g of sulphur trioxide as follows:

Sulphur	+	Oxygen	$\longrightarrow$	Sulphur trioxide
32 g		48 g		80 g

Hence, (2 × 32 = 64 g) of sulphur will combine with (2 × 48 = 96 g) of oxygen to give (2 × 80 = **160 g) sulphur trioxide**.

Ans: Mass of sulphur trioxide produced = **160 g**

SMART CHECK

The ratio of masses of sulphur and oxygen is 2 : 3 in sulphur trioxide. Hence, 32 × 2 = 64 g sulphur combines with 32 × 3 = 96 g oxygen to give 160 g sulphur dioxide.

Practice Numericals

1. 12 g of carbon reacts with some oxygen to make 44 grams of carbon dioxide. Find out how much oxygen must have been used. [2 Marks]

Ans: 32 g

2. Calculate the mass of sulphur trioxide produced by burning 16 g of sulphur in excess of oxygen. (Average atomic mass: S = 32 u, O = 16 u). [2 Marks]

Ans: 40 g

1.5 Avogadro's law

***Q.52. State and explain Avogadro's law.** [3 Marks]

Ans:

i. In the year 1811, Avogadro made a distinction between atoms and molecules and thereby proposed Avogadro's law.

ii. Avogadro proposed that, "*Equal volumes of all gases at the same temperature and pressure contain equal number of molecules*".

e.g. Hydrogen gas combines with oxygen gas to produce water vapour as follows:

$Hydrogen_{(g)}$	+	$Oxygen_{(g)}$	$\longrightarrow$	$Water_{(g)}$
[100 mL]		[50 mL]		[100 mL]
[2 vol]		[1 vol]		[2 vol]

According to Avogadro's law, if 1 volume contains n molecules, then 2n molecules of hydrogen combine with n molecules of oxygen to give 2n molecules of water, i.e., 2 molecules of hydrogen gas combine with 1 molecule of oxygen to give 2 molecules of water vapour as represented below:

$Hydrogen_{(g)}$	+	$Oxygen_{(g)}$	$\longrightarrow$	$Water_{(g)}$
[2n molecules]		[n molecules]		[2n molecules]
[2 molecules]		[1 molecule]		[2 molecules]

READING BETWEEN THE LINES

Avogadro could explain the above result by assuming the molecules to be polyatomic that is quite understandable today as hydrogen and oxygen are diatomic molecules.

Q.53. Complete the following table: [½ Mark Each]

Statement	Law
When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers.	----------
Equal volumes of all gases at the same temperature and pressure contain equal number of molecules.	----------

When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.	----------
A given compound always contains exactly the same proportion of elements by weight.	----------

Ans:

Statement	Law
When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers.	**Law of multiple proportions**
Equal volumes of all gases at the same temperature and pressure contain equal number of molecules.	**Avogadro's law**
When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.	**Gay Lussac's law**
A given compound always contains exactly the same proportion of elements by weight.	**Law of definite proportions**

1.6 Dalton's Atomic Theory

Q.54. Explain Dalton's atomic theory. **[2 Marks]**

Ans: John Dalton published "A New System of chemical philosophy" in the year of 1808. He proposed the following features, which later became famous as Dalton's atomic theory.

i. Matter consists of tiny, indivisible particles called atoms.

ii. All the atoms of a given elements have identical properties including mass. Atoms of different elements differ in mass.

iii. Compounds are formed when atoms of different elements combine in a fixed ratio.

iv. Chemical reactions involve only the reorganization of atoms. Atoms are neither created nor destroyed in a chemical reaction.

Dalton's atomic theory could explain all the laws of chemical combination.

Q.55. Give reason: Dalton's atomic theory explains the law of conservation of mass. **[2 Marks]**

Ans:

i. The law of conservation of mass states that, *"Mass can neither be created nor destroyed"* during chemical combination of matter.

ii. According to Dalton's atomic theory, chemical reactions involve only the reorganization of atoms. Therefore, the total number of atoms in the reactants and products should be same and mass is conserved during a reaction.

Hence, Dalton's atomic theory explains the law of conservation of mass.

Q.56. Give reason: Dalton's atomic theory explains the law of multiple proportion. **[2 Marks]**

Ans:

i. The law of multiple proportion states that, *"When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers"*.

ii. According to Dalton's atomic theory, compounds are formed when atoms of different elements combine in fixed ratio.

Hence, Dalton's atomic theory explains the law of multiple proportion.

1.7 Atomic And Molecular Masses

Q.57. Can you recall? *(Textbook page no. 6)*

What is an atom and a molecule? What is the order of magnitude of mass of one atom? What are isotopes? **[3 Marks]**

Ans:

i. *The smallest indivisible particle of an element is called an **atom**.*

ii. *A **molecule** is an aggregate of two or more atoms of definite composition which are held together by chemical bonds.*

iii. Every atom of an element has definite mass. The order of magnitude of mass of one atom is 10^{-27} kg.

iv. ***Isotopes** are the atoms of the same element having same atomic number but different mass number.*

Q.58. Define: Atomic mass unit (amu) **[1 Mark]**

Ans: Atomic mass unit or *amu* *is defined as a mass exactly equal to one twelfth of the mass of one carbon-12 atom.*

***Q.59. How many grams does an atom of hydrogen weigh?** **[1 Mark]**

Ans: The mass of a hydrogen atom is 1.6736×10^{-24} g.

Q.60. How is relative atomic mass of an atom measured? **[3 Marks]**

Ans:

i. The mass of a single atom is extremely small, i.e. the mass of a hydrogen atom is 1.6736×10^{-24} g. Hence, it is not possible to weigh a single atom.

ii. In the present system, mass of an atom is determined relative to the mass of an atom of carbon-12 as the standard. This was decided in 1961 by IUPAC.

iii. The atomic mass of carbon-12 is assigned as 12.00000 atomic mass unit (amu).

iv. The masses of all other elements are determined relative to the mass of an atom of carbon-12 (C-12).

v. The atomic masses are expressed in amu which is exactly equal to one twelfth of the mass of one carbon-12 atom.

vi. The value of 1 amu is equal to 1.6605×10^{-24} g.

READING BETWEEN THE LINES

The exact value of amu was experimentally determined as shown below:

$$1\ amu = \frac{1}{12} \times \text{mass of one C-12}$$

$$= \frac{1}{12} \times 1.992648 \times 10^{-23}\ g$$

$$= 1.66056 \times 10^{-24}\ g$$

Q.61. What is meant by Unified Mass unit? **[1 Mark]**

Ans:

i. Presently, instead of amu, Unified Mass has now been accepted as the unit of atomic mass.

ii. It is called Dalton and its symbol is 'u' or 'Da'.

Q.62. What is average atomic mass? **[1 Mark]**

Ans: The atomic mass of an element which exists as mixture of two or more isotopes is the average of atomic masses of its isotopes. This is called average atomic mass.

***Q.63. Explain: The need of the term average atomic mass.** **[2 Marks]**

Ans:

i. Several naturally occurring elements exist as a mixture of two or more isotopes.

ii. Isotopes have different atomic masses.

iii. The atomic mass of such an element is the average of atomic masses of its isotopes.

iv. For this purpose, the atomic masses of isotopes and their relative percentage abundances are considered.

Hence, the term average atomic mass is needed to express atomic mass of elements containing mixture of two or more isotopes.

READING BETWEEN THE LINES

Carbon has three isotopes. The relative abundance and atomic masses of the isotopes of carbon are as shown in the table below:

Isotopes	*Atomic mass (u)*	*Relative abundance (%)*
^{12}C	12.00000	98.892
^{13}C	13.00335	1.108
^{14}C	14.00317	2×10^{-10}

Average atomic mass of carbon

$= (12.00000 \times 98.892/100) + (13.00335 \times 1.108/100) + (14.00317 \times 2 \times 10^{-10}/100)$

$= (11.86704) + (0.144077) + (0.00000)$

$= 12.01112 = 12.011\ u$

Note: The relative abundance of ^{14}C is very small and hence, its contribution to average atomic mass of carbon is negligible.

FOR YOUR KNOWLEDGE

In the periodic table of elements, the atomic masses mentioned for different elements are actually their average atomic masses. For practical purpose, the average atomic mass is rounded off to the nearest whole number when it differs from it by a very small fraction. *(Textbook page no. 7)*

Element	Isotopes	Average atomic mass	Rounded off atomic mass
Carbon	^{12}C, ^{13}C, ^{14}C	12.011 u	12.0 u
Nitrogen	^{14}N, ^{15}N	14.007 u	14.0 u
Oxygen	^{16}O, ^{17}O, ^{18}O	15.999 u	16.0 u
Chlorine	^{35}Cl, ^{37}Cl	35.453 u	35.5 u
Bromine	^{79}Br, ^{81}Br	79.904 u	79.9 u

GG - GYAN GURU

Isotopes as Detective!!

If an athlete takes a synthetic steroid to enhance performance, how would scientist find out whether the steroid (testosterone) is normally occurring in body or that it has synthetic origin? The naturally occurring steroid in athletes in most countries will have a different $^{13}C/^{12}C$ ratio than synthetic steroid. A scientist with a mass spectrometer can easily detect the difference and thus catch up the illegal drug abuse among athletes!!!

Q.64. Define: Molecular mass **[1 Mark]**

Ans: ***Molecular mass*** *of a substance is the sum of average atomic masses of the atoms of the elements which constitute the molecule.*

OR

Molecular mass *of a substance is the mass of one molecule of that substance relative to the mass of one carbon-12 atom.*

Q.65. How is molecular mass of a substance calculated? Give an example. **[2 Marks]**

Ans: Molecular mass is calculated by multiplying average atomic mass of each element by the number of its atoms and adding them together.

e.g. Molecular mass of carbon dioxide (CO_2) is calculated as follows:

Molecular mass of CO_2 = (1 × average atomic mass of C) + (2 × average atomic mass of O)

= (1 × 12.0 u) + (2 × 16.0 u) = 44.0 u

Q.66. Define: Formula mass **[1 Mark]**

Ans: *The* ***formula mass*** *of a substance is the sum of atomic masses of the atoms present in the formula.*

***Q.67. Explain: Formula mass with an example** **[3 Marks]**

Ans:

i. Definition: *Refer Q. 66.*

ii. In substances such as sodium chloride, positive (sodium) and negative (chloride) entities are arranged in a three-dimensional structure in a way that one sodium (Na^+) ion is surrounded by six chloride (Cl^-) ions, all at the same distance from it and vice versa. Thus, sodium chloride do not contain discrete molecules as the constituent units.

iii. Therefore, NaCl is the formula which is used to represent sodium chloride though it is not a molecule.

iv. In such compounds, the formula (i.e., NaCl) is used to calculate the formula mass instead of molecular mass.

e.g. Formula mass of sodium chloride

= atomic mass of sodium + atomic mass of chlorine

= 23.0 u + 35.5 u = 58.5 u

Q.68. Complete the following table: **[½ Mark Each]**

Column A	Column B
The mass of one hydrogen atom in gram	----------
The exact value of 1 atomic mass unit (amu) in gram	----------
Isotopes of carbon	----------
Formula mass of NaCl	----------

Ans:

Column A	Column B
The mass of one hydrogen atom in gram	$\mathbf{1.6736 \times 10^{-24}}$ **g**
The exact value of atomic mass unit (amu) in gram	$\mathbf{1.66056 \times 10^{-24}}$ **g**
Isotopes of carbon	$\mathbf{^{12}C, ^{13}C, ^{14}C}$
Formula mass of NaCl	**58.5 u**

Q.69. Name the following. **[1 Mark Each]**

i. **An atom which is assigned a mass of exactly 12.00000 u.**

ii. **A unified mass unit which is recently replaced by amu.**

iii. **Elements having different atomic masses but same atomic number.**

iv. **The term used for the mass of one molecule of a substance relative to the mass of one carbon-12 atom.**

Ans:

i. Carbon-12 ii. Dalton

iii. Isotopes iv. Molecular mass

Numerical Zone

+Q.70. Mass of an atom of oxygen in gram is 26.56896 × $\mathbf{10^{-24}}$ g. What is the atomic mass of oxygen in u?

(Problem 1.1 of Textbook page no. 7) **[2 Marks]**

Solution:

Given: Mass of an atom of oxygen in gram is 26.56896×10^{-24} g.

To find: Atomic mass of oxygen in u

Calculation: 1.66056×10^{-24} g = 1 u

∴ 26.56896×10^{-24} g = x

$$\therefore \quad x = \frac{26.56896 \times 10^{-24}\text{ g}}{1.66056 \times 10^{-24}\text{ g/u}} = \mathbf{16.0\ u}$$

Ans: The atomic mass of oxygen in u = **16.0 u**

Caution

Always use the proper units with your numerical answers.

Q.71. Mass of an atom of hydrogen in gram is 1.6736 × $\mathbf{10^{-24}}$ g. What is the atomic mass of hydrogen in u? **[2 Marks]**

Solution:

Given: Mass of an atom of hydrogen in gram is 1.6736×10^{-24} g.

To find: Atomic mass of hydrogen in u

Calculation: 1.66056×10^{-24} g = 1 u

∴ 1.6736×10^{-24} g = x

$$\therefore \quad x = \frac{1.6736 \times 10^{-24}\text{ g}}{1.66056 \times 10^{-24}\text{ g/u}} = \mathbf{1.008\ u}$$

Ans: The atomic mass of hydrogen in u = **1.008 u**

***Q.72. The mass of an atom of hydrogen is 1.008 u. What is the mass of 18 atoms of hydrogen?** **[1 Mark]**

Solution:

Mass of 1 atom of hydrogen = 1.008 u

∴ Mass of 18 atoms of hydrogen = 18 × 1.008 u = **18.144 u**

Ans: The mass of 18 atoms of hydrogen = **18.144 u**

Q.73. The mass of an atom of one carbon atom is 12.011 u. What is the mass of 20 atoms of the same isotope? **[1 Mark]**

Solution:

Mass of 1 atom of carbon = 12.011 u

∴ Mass of 20 atoms of same carbon isotope = 20 × 12.011 u = **240.220 u**

Ans: The mass of 20 atoms of same carbon isotope = **240.220 u**

+Q.74. Calculate the average atomic mass of neon using the following data:
(Problem 1.2 of Textbook page no. 8) **[2 Marks]**

Isotope	Atomic mass	Natural Abundance
^{20}Ne	19.9924 u	90.92%
^{21}Ne	20.9940 u	0.26 %
^{22}Ne	21.9914 u	8.82 %

Solution:

Average atomic mass of Neon (Ne)

$$= \frac{(\text{At. mass of } ^{20}\text{Ne} \times \%\text{Abundance}) + (\text{At. mass of } ^{21}\text{Ne} \times \%\text{Abundance}) + (\text{At. mass of } ^{22}\text{Ne} \times \%\text{Abundance})}{100}$$

$$= \frac{(19.9924\ \text{u} \times 90.92) + (20.9940\ \text{u} \times 0.26) + (21.9914\ \text{u} \times 8.82)}{100}$$

= **20.1707 u**

Ans: Average atomic mass of neon = **20.1707 u**

***Q.75. The natural isotopic abundance of ^{10}B is 19.60% and ^{11}B is 80.40%. The exact isotopic masses are 10.13 and 11.009 respectively. Calculate the average atomic mass of boron.** **[2 Marks]**

Solution:

Average atomic mass of Boron (B)

$$= \frac{(\text{At. mass of } ^{10}\text{B} \times \%\text{Abundance}) + (\text{At. mass of } ^{11}\text{B} \times \%\text{Abundance})}{100}$$

$$= \frac{(10.13\ \text{u} \times 19.60) + (11.009\ \text{u} \times 80.40)}{100}$$

= **10.84 u**

Ans: Average atomic mass of boron = **10.84 u**

SMART CHECK

The average mass atomic mass should be near the mass of the isotope with the largest abundance. The abundance of ^{11}B is 80.40%. Hence, the average atomic mass should be close to 11 u. The calculated average atomic mass 10.84 u is close to 11 u. Hence, the answer is correct.

Q.76. Calculate the average atomic mass of argon from the following data: **[2 Marks]**

Isotope	Isotopic mass (u)	Abundance
^{36}Ar	35.96755	0.337%
^{38}Ar	37.96272	0.063%
^{40}Ar	39.9624	99.600%

Solution:

Average atomic mass of argon (Ar)

$$= \frac{(\text{At. mass of } ^{36}\text{Ar} \times \%\text{Abundance}) + (\text{At. mass of } ^{38}\text{Ar} \times \%\text{Abundance}) + (\text{At. mass of } ^{40}\text{Ar} \times \%\text{Abundance})}{100}$$

$$= \frac{(35.96755\ \text{u} \times 0.337) + (37.96272\ \text{u} \times 0.063) + (39.9624\ \text{u} \times 99.60)}{100}$$

= **39.947 u**

Ans: Average atomic mass of argon = **39.947 u**

SMART CHECK

The average mass atomic mass should be near the mass of the isotope with the largest abundance. The abundance of ^{40}Ar is 99.600%. Hence, the average atomic mass should be very close to 39.9624 u. The calculated average atomic mass 39.947 u is close to 39.96 u. Hence, the answer is correct.

Q.77. Calculate the molecular mass of the following in u: **[1 Mark Each]**

i. H_2O **ii.** C_6H_5Cl **iii.** H_2SO_4

Solution:

i. Molecular mass of H_2O = (2 × Average atomic mass of H) + (1 × Average atomic mass of O)
= (2 × 1.0 u) + (1 × 16.0 u) = **18 u**

ii. Molecular mass of C_6H_5Cl = (6 × Average atomic mass of C) + (5 × Average atomic mass of H) + (1 × Average atomic mass of Cl)
= (6 × 12.0 u) + (5 × 1.0 u) + (1 × 35.5 u) = **112.5 u**

iii. Molecular mass of H_2SO_4 = (2 × Average atomic mass of H) + (1 × Average atomic mass of S) + (4 × Average atomic mass of O)
= (2 × 1.0 u) + (1 × 32.0 u) + (1 × 16.0 u) = **98 u**

Ans:
i. The molecular mass of H_2O = **18 u**
ii. The molecular mass of C_6H_5Cl = **112.5 u**
iii. The molecular mass of H_2SO_4 = **98 u**

*Q.78. **Calculate the molecular mass of the following in u:** [1 Mark Each]
i. NH_3 **ii.** CH_3COOH
iii. C_2H_5OH

Solution:
i. Molecular mass of NH_3 = (1 × Average atomic mass of N) + (3 × Average atomic mass of H)
= (1 × 14.0 u) + (3 × 1.0 u) = **17 u**

ii. Molecular mass of CH_3COOH = (2 × Average atomic mass of C) + (4 × Average atomic mass of H) + (2 × Average atomic mass of O)
= (2 × 12.0 u) + (4 × 1.0 u) + (2 × 16.0 u) = **60 u**

iii. Molecular mass of C_2H_5OH = (2 × Average atomic mass of C) + (6 × Average atomic mass of H) + (1 × Average atomic mass of O)
= (2 × 12.0 u) + (6 × 1.0 u) + (1 × 16.0 u) = **46 u**

Ans:
i. The molecular mass of NH_3 = **17 u**
ii. The molecular mass of CH_3COOH = **60 u**
iii. The molecular mass of C_2H_5OH = **46 u**

+Q.79. **Find the mass of 1 molecule of oxygen (O_2) in amu (u) and in grams.**
(Problem 1.3 of Textbook page no. 8) [2 Marks]

Solution:
Molecular mass of O_2 = 2 × 16 u
∴ Mass of 1 molecule = **32 u**
∴ Mass of 1 molecule of O_2 = 32 × 1.66056 × 10^{-24} g
= **53.1379 × 10^{-24} g**

Ans: Mass of 1 molecule in amu = **32 u**
Mass of 1 molecule in grams = **53.1379 × 10^{-24} g**

+Q.80. **Find the formula mass of**
i. **NaCl** **ii.** $Cu(NO_3)_2$
(Problem 1.4 of Textbook page no. 8)
[1 Mark Each]

Solution:
i. Formula mass of NaCl
= Average atomic mass of Na
+ Average atomic mass of Cl
= 23.0 u + 35.5 u = **58.5 u**

ii. Formula mass of $Cu(NO_3)_2$
= Average atomic mass of Cu + 2 × (Average atomic mass of N + Average atomic mass of three O)
= 63.5 + 2 × [14 + (3 × 16)] = **187.5 u**

Ans:
i. Formula mass of NaCl = **58.5 u**
ii. Formula mass of $Cu(NO_3)_2$ = **187.5 u**

CAUTION

You must be very careful when you are counting the number of atoms present in compound $Cu(NO_3)_2$. The subscript 2 after the brackets indicates that there are 2 nitrate ions.

Q.81. **Find the formula mass of:**
i. **KCl** **ii.** **AgCl**
Atomic mass of K = 39 u, Ag =108 u and Cl = 35.5 u. [1 Mark Each]

Solution:
i. Formula mass of KCl
= Average atomic mass of K
+ Average atomic mass of Cl
= 39 u + 35.5 u = **74.5 u**

ii. Formula mass of AgCl
= Average atomic mass of Ag
+ Average atomic mass of Cl
= 108 + 35.5 = **143.5 u**

Ans: i. Formula mass of KCl = **74.5 u**
ii. Formula mass of AgCl = **143.5 u**

Q.82. **Try this** *(Textbook page no. 8)*
Find the formula mass of $CaSO_4$, if atomic mass of Ca = 40.1 u, S = 32.1 u and O = 16.0 u.
[1 Mark]

Solution:
Formula mass of $CaSO_4$
= Average atomic mass of Ca + Average atomic mass of S + Average atomic mass of four O
= 40.1 + 32.1 + (4 × 16.0) = **136.2 u**

Ans: Formula mass of $CaSO_4$ = **136.2 u**

Practice Numericals

1. Calculate the average atomic mass of chromium from the following data: [3 Marks]

Isotope	Isotopic mass (u)	Abundance
^{50}Cr	49.9461	0.0435
^{52}Cr	51.9405	0.8379
^{53}Cr	52.9407	0.0950
^{54}Cr	53.9389	0.0236

Ans: 51.99 u

2. Calculate the molecular mass of the following in u: [3 Marks]
i. CH_3OH ii. NO_2 iii. HNO_3
Ans: i. 32 u ii. 46 u
iii. 63 u

3. Find the formula mass of Na_2SO_4.
(Atomic mass of Na = 23 u, S = 32 u, O = 16 u)
[1 Mark]
Ans: 142 u.

1.8 Mole Concept and Molar Mass

Q.83. **Can you recall?** *(Textbook page no. 8)*
i. **One dozen means how many items?** [1 Mark]
ii. **One gross means how many items?** [1 Mark]

Ans:
i. One dozen means 12 items.
ii. One gross means 144 items.

***Q.84. Explain: Mole concept** **[2 Marks]**

Ans:

i. Even a small amount of any substance contains very large number of atoms or molecules. Therefore, a quantitative adjective *'mole'* is used to express the large number of sub-microscopic entities like atoms, molecules, ions, electrons, etc. present in a substance.

ii. *Thus, **one mole** is the amount of a substance that contains as many entities or particles as there are atoms in exactly 12 g (or 0.012 kg) of the carbon -12 isotope.*

iii. One mole is the amount of substance which contains 6.0221367×10^{23} particles/entities.

Reading Between the Lines

Mass of one carbon-12 atom as determined by mass spectrometer is 1.992648×10^{-23} g.

Mass of one mole of carbon atoms is 12 g.

Hence, Number of atoms in 12 g of carbon-12 =

$$\frac{12 \text{ g / mol}}{1.992648 \times 10^{-23} \text{ g/atom}} = 6.02213 \times 10^{23} \text{ atom/mol}$$

***Q.85. How many particles are present in 1 mole of a substance?** **[1 Mark]**

Ans: The number of particles in one mole is 6.0221367×10^{23}.

For Your Knowledge

i. The name of the unit is **mole** and the symbol for the unit is **mol**.

ii. The number 6.0221367×10^{23} is known as **'Avogadro's Constant (N_A)'** in the honour of Amedo Avogadro.

iii. The number of atoms, molecules, ions or electrons, etc. present in 1 mole of a substance is found to be equal to 6.0221367×10^{23}, which is called **Avogadro Number**.

iv. The number 6.0221367×10^{23} is often rounded to three decimal point as 6.022×10^{23} in calculations.

v. In SI system, mole (Symbol mol) was introduced as seventh base quantity for the amount of a substance.

***Q.86. Explain: Molar mass** **[3 Marks]**

Ans:

i. *The mass of one mole of a substance (element/compound) in grams is called its **molar mass**.*

ii. The molar mass of any element in grams is numerically equal to atomic mass of that element in u.

e.g.

Element	Atomic mass (u)	Molar mass (g mol^{-1})
H	1.0	1.0
C	12.0	12.0
O	16.0	16.0

iii. Similarly, molar mass of polyatomic molecule, in grams is numerically equal to its molecular mass or formula mass in u.

e.g.

Polyatomic substance	Molecular/formula mass (u)	Molar mass (g mol^{-1})
O_2	32.0	32.0
H_2O	18.0	18.0
NaCl	58.5	58.5

***Q.87. Point out the difference between 12 g of carbon and 12 u of carbon.** **[1 Mark]**

Ans: 12 g of carbon is the molar mass of carbon while 12 u of carbon is the mass of one carbon atom.

Numerical Zone

***Q.88. What is the ratio of molecules in 1 mole of NH_3 and 1 mole of HNO_3?** **[2 Marks]**

Solution:

One mole of any substance contains particles equal to 6.022×10^{23}.

1 mole of $NH_3 = 6.022 \times 10^{23}$ molecules of NH_3

1 mole of $HNO_3 = 6.022 \times 10^{23}$ molecules of HNO_3

$$\therefore \quad \text{Ratio} = \frac{6.022 \times 10^{23}}{6.022 \times 10^{23}} = \mathbf{1:1}$$

Ans: The ratio of molecules is = **1:1**

Caution

One mole of any substance contains Avogadro's number of molecules/particles/atoms.

***Q.89. In two moles of acetaldehyde (CH_3CHO) calculate the following:**

i. Number of moles of carbon

ii. Number of moles of hydrogen

iii. Number of moles of oxygen

iv. Number of molecules of acetaldehyde

[2 Marks]

Solution:

Molecular formula of acetaldehyde: C_2H_4O
Moles of acetaldehyde = 2 mol

i. Number of moles of carbon atoms
= Moles of acetaldehyde × Number of carbon atoms
= 2 × 2 = **4 moles of carbon atoms**

ii. Number of moles of hydrogen atoms
= Moles of acetaldehyde × Number of hydrogen atoms
= 2 × 4 = **8 moles of hydrogen atoms**

iii. Number of moles of oxygen atoms
= Moles of acetaldehyde × Number of oxygen atoms
= 2 × 1 = **2 moles of oxygen atoms**

iv. Number of molecules of acetaldehyde
= Moles of acetaldehyde × Avogadro number (N_A)
= 2 mol × 6.022 × 10^{23} molecules/mol
= 12.044 × 10^{23} molecules of acetaldehyde

Ans:

i. Number of moles of carbon, hydrogen and oxygen are **4, 8, 2** respectively.

ii. Number of molecules of acetaldehyde = **12.044 × 10^{23}**

STRATEGY

Given: Moles of CH_3CHO

1. No. of moles of atom 'X' in a given substance is equal to the No. of moles of that substance multiplied by no. of atom 'X' in a molecule of that substance
2. Using the above relation,
Number of moles of C-atoms
= Moles of CH_3CHO × No. of C-atoms
3. Similarly, calculate no. of moles of H-atoms and O-atoms
4. No. of molecules of a given substance
= Moles of that substance × N_A
5. Using this relation, calculate no. of molecules of CH_3CHO

***Q.90. Calculate the number of moles of magnesium oxide, MgO in**
i. 80 g and
ii. 10 g of the compound. **[3 Marks]**
(Average atomic masses of Mg = 24 and O = 16)

Solution:

Given: i. Mass of MgO = 80 g
ii. Mass of MgO = 10 g

To find: Number of moles of MgO

Formula: Number of moles (n)

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$$

Calculation: i. Molecular mass of MgO
= (1 × Average atomic mass of Mg) + (1 × Average atomic mass of O)
= (1 × 24 u) + (1 × 16 u) = 40 u

∴ Molar mass of MgO = 40 g mol^{-1}

Mass of MgO = 80 g
Number of moles (n)

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$$

$$= \frac{80\text{ g}}{40\text{ g mol}^{-1}} = \mathbf{2\ mol}$$

ii. Mass of MgO = 10 g,
Molar mass of MgO = 40 g mol^{-1}
Number of moles (n)

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$$

$$= \frac{10\text{ g}}{40\text{ g mol}^{-1}} = \mathbf{0.25\ mol}$$

Ans:

i. The number of moles in 80 g of magnesium oxide, MgO = **2 mol**

ii. The number of moles in 10 g of magnesium oxide, MgO = **0.25 mol**

SMART CHECK

If the given mass is less than its molar mass, then the number of moles should be less than one. If the given mass is more than its molar mass, then the number of moles should be more than one.

+Q.91. Calculate the number of moles and molecules of urea present in 5.6 g of urea.
(Problem 1.5 of Textbook page no. 9) **[3 Marks]**

Solution:

Given: Mass of urea = 5.6 g

To find: The number of moles and molecules of urea

Formulae: i. Number of moles

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$$

ii. Number of molecules
= Number of moles × Avogadro's constant

Calculation: Mass of urea = 5.6 g
Molecular mass of urea, NH_2CONH_2
= (2 × Average atomic mass of N)
+ (4 × Average atomic mass of H)
+ (1 × Average atomic mass of C)
+ (1 × average atomic mass of O)
= (2 × 14 u) + (4 × 1 u) + (1 × 12 u) + (1 × 16 u)
= 60 u

$\therefore$ Molar mass of urea = 60 g mol^{-1}

Number of moles

$= \dfrac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \dfrac{5.6 \text{ g}}{60 \text{ g mol}^{-1}}$ = **0.09333 mol**

CALCULATION USING LOG TABLE

$\dfrac{5.6}{60}$

$= \text{Antilog}_{10} [\log_{10} (5.6) - \log_{10} (60)]$

$= \text{Antilog}_{10} [0.7482 - 1.7782]$

$= \text{Antilog}_{10} [\bar{2}.9700]$

$= 0.09333$

Now,

Number of molecules of urea

= Number of moles × Avogadro's constant

= 0.09333 mol × 6.022 × 10^{23} molecules/mol

= 0.5616 × 10^{23} molecules *(Using log table)*

= **5.616 × 10^{22} molecules**

CALCULATION USING LOG TABLE

0.09333 × 6.022

$= \text{Antilog}_{10} [\log_{10} (0.09333) + \log_{10} (6.022)]$

$= \text{Antilog}_{10} [\bar{2}.9698 + 0.7797]$

$= \text{Antilog}_{10} [\bar{1}.7495] = 0.5616$

Ans: Number of moles of urea = **0.0933 mol**

Number of molecules of urea = **5.616 × 10^{22} molecules**

SMART CHECK

If the given mass is less than its molar mass, then the number of molecules should be less than Avogadro's constant. If the given mass is more than its molar mass, then the number of molecules should be more than Avogadro's constant.

***Q.92. Calculate the number of moles and molecules of acetic acid present in 22 g of it.** **[3 Marks]**

Solution:

Given: Mass of acetic acid = 22 g

To find: The number of moles and molecules of acetic acid

Formulae: i. Number of moles $= \dfrac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

ii. Number of molecules = Number of moles × Avogadro's constant

Calculation: Mass of acetic acid = 22 g

Molecular mass of acetic acid, CH_3COOH

= (2 × Average atomic mass of C)

+ (4 × Average atomic mass of H)

+ (2 × Average atomic mass of O)

= (2 × 12 u) + (4 × 1 u) + (2 × 16 u)

= 60 u

$\therefore$ Molar mass of acetic acid = 60 g mol^{-1}

Number of moles

$= \dfrac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \dfrac{22 \text{ g}}{60 \text{ g mol}^{-1}}$ = **0.367 mol**

Now,

Number of molecules of acetic acid

= Number of moles × Avogadro's constant

= 0.367 mol × 6.022 × 10^{23} molecules/mol

= **2.210 × 10^{23} molecules**

Ans: Number of moles = **0.367 mol**

Number of molecules of acetic acid

= **2.210 × 10^{23} molecules**

***Q.93. Calculate the number of atoms in each of the following (Given: Atomic mass of I = 127 u).**

i. 254 u of iodine (I) ii. 254 g of iodine (I) **[3 Marks]**

Solution:

i. 254 u of iodine (I) = x atoms,

Atomic mass of iodine (I) = 127 u

$\therefore$ Mass of one iodine atom = 127 u

$\therefore$ $x = \dfrac{254 \text{ u}}{127 \text{ u}}$ = **2 atoms**

ii. 254 g of iodine (I),

Atomic mass of iodine = 127 u

$\therefore$ Molar mass of iodine = 127 g mol^{-1}

Now, Number of moles $= \dfrac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \dfrac{254 \text{ g}}{127 \text{ g mol}^{-1}} = 2 \text{ mol}$

Now, Number of atoms

= Number of moles × Avogadro's constant

= 2 mol × 6.022 × 10^{23} atoms/mol

= 12.044 × 10^{23} atoms

= **1.2044 × 10^{24} atoms**

Ans:

i. Number of iodine atoms in 254 u = **2 atoms**

ii. Number of iodine atoms in 254 g

= **1.2044 × 10^{24} atoms**

Q.94. Calculate the number of atoms in each of the following:

i. 64 u of oxygen (O)

ii. 42 g of nitrogen (N) **[3 Marks]**

Solution:

i. 64 u of oxygen (O) = x atoms,
Atomic mass of oxygen (O) = 16 u

∴ Mass of one oxygen atom = 16 u

∴ $x = \frac{64 \text{ u}}{16 \text{ u}}$ = **4 atoms**

ii. 42 g of nitrogen (N),
Atomic mass of nitrogen = 14 u

∴ Molar mass of nitrogen = 14 g mol^{-1}

Now, $\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{42 \text{ g}}{14 \text{ g mol}^{-1}} = 3 \text{ mol}$

Now, Number of atoms
= Number of moles × Avogadro's constant
= 3 mol × 6.022 × 10^{23} atoms/mol
= 18.07 × 10^{23} atoms
= **1.807 × 10^{24} atoms**

Ans:

i. Number of oxygen atoms in 64 u = **4 atoms**

ii. Number of nitrogen atoms in 42 g
= **1.807 × 10^{24} atoms**

+Q.95. Calculate the number of atoms in each of the following:

i. 52 moles of Argon (Ar)

ii. 52 u of Helium (He)

iii. 52 g of Helium (He)

(Problem 1.6 of Textbook page no. 10) **[3 Marks]**

Solution:

i. 52 moles of Argon
1 mole Argon atoms = 6.022 × 10^{23} atoms of Ar

∴ Number of atoms
= 52 mol × 6.022 × 10^{23} atoms/mol
= **313.144 × 10^{23} atoms of Argon**

ii. 52 u of Helium
Atomic mass of He = mass of 1 atom of He = 4.0 u
4.0 u = 1 He

∴ 52 u = x

∴ $x = 52 \text{ u} \times \frac{1 \text{ atom of He}}{4.0 \text{ u}}$ = **13 atoms of He**

iii. 52 g of He
Molar mass of He = 4.0 g mol^{-1}

$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{52 \text{ g}}{4.0 \text{ g mol}^{-1}} = 13 \text{ mol}$

Number of atoms of He = Number of moles
× Avogadro's constant
= 13 mol × 6.022 × 10^{23} atoms/mol
= **78.286 × 10^{23} atoms of He**

Ans:

i. Number of argon atoms in 52 moles
= **313.144 × 10^{23} atoms of Argon**

ii. Number of helium atoms in 52 u
= **13 atoms of He**

iii. Number of helium atoms in 52 g
= **78.286 × 10^{23} atoms of He**

***Q.96. Calculate number of atoms is each of the following.**
(Average atomic mass: N = 14 u, S = 32 u)

i. 0.4 mole of nitrogen

ii. 1.6 g of sulphur **[3 Marks]**

Solution:

i. 0.4 mole of nitrogen (N)
Number of atoms of N = Number of moles
× Avogadro's constant
= 0.4 mol × 6.022 × 10^{23} atoms/mol
= **2.4088 × 10^{23} atoms of N**

ii. 1.6 g of Sulphur (S), Molar mass of sulphur
= 32 g mol^{-1}

$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{1.6 \text{ g}}{32 \text{ g mol}^{-1}} = 0.05 \text{ mol}$

Number of atoms of S = Number of moles
× Avogadro's constant
= 0.05 mol × 6.022 × 10^{23} atoms/mol
= 0.3011 × 10^{23} atoms
= **3.011 × 10^{22} atoms of S**

Ans:

i. Number of nitrogen atoms in 0.4 mole
= **2.4088 × 10^{23} atoms of N**

ii. Number of sulphur atoms in 1.6 g
= **3.011 × 10^{22} atoms of S**

***Q.97. A student used a carbon pencil to write his homework. The mass of this was found to be 5 mg. With the help of this calculate**

i. The number of moles of carbon in his homework writing.

ii. The number of carbon atoms in 12 mg of his homework writing. **[2 Marks]**

Solution:

i. 5 mg carbon = 5 × 10^{-3} g carbon, Atomic mass of carbon = 12 u

∴ Molar mass of carbon = 12 g mol^{-1}

$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{5 \times 10^{-3} \text{ g}}{12 \text{ g mol}^{-1}}$

= **4.167 × 10^{-4} mol**

CALCULATION USING LOG TABLE

$\left[\frac{5}{12}\right]\times 10^{-3}$

$= \text{Antilog}_{10}\,[\log_{10}(5) - \log_{10}(12)] \times 10^{-3}$

$= \text{Antilog}_{10}\,[0.6990 - 1.0792] \times 10^{-3}$

$= \text{Antilog}_{10}\,[\bar{1}.6198] \times 10^{-3}$

$= 4.167 \times 10^{-4}$

ii. 12 mg carbon = 12×10^{-3} g carbon

Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{12 \times 10^{-3}\ \text{g}}{12\ \text{g mol}^{-1}}$

$= 1 \times 10^{-3}$ mol

Number of atoms = Number of moles × Avogadro's constant

Number of atoms of carbon

$= 1 \times 10^{-3}$ mol × 6.022×10^{23} atoms/mol

= 6.022×10^{20} atoms

Ans: Number of moles of carbon in his homework writing = **4.167×10^{-4} mol**

Number of atoms of carbon in 12 mg homework writing = **6.022×10^{20} atoms**

***Q.98. Calculate the number of atoms of hydrogen present in 5.6 g of urea, $(NH_2)_2CO$. Also calculate the number of atoms of N, C and O.** **[4 Marks]**

Solution:

Given: Mass of urea = 5.6 g

To find: The number of atoms of hydrogen, nitrogen, carbon and oxygen

Calculation: Molecular formula of urea: $(NH_2)_2CO$

Molar mass of urea = 60 g mol^{-1}

Number of moles

$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{5.6\ \text{g}}{60\ \text{g mol}^{-1}} = 0.0933$ mol

∴ Moles of urea = 0.0933 mol

Number of atoms = Number of moles × Avogadro's constant

Now, 1 molecule of urea has total 8 atoms, out of which 4 atoms are of H, 2 atoms are of N, 1 of C and 1 of O.

∴ Number of H atoms in 5.6 g of urea

$= (4 \times 0.0933)$ mol × 6.022×10^{23} atoms/mol

= 2.247×10^{23} atoms of hydrogen

∴ Number of N atoms in 5.6 g of urea

$= (2 \times 0.0933)$ mol × 6.022×10^{23} atoms/mol

= 1.124×10^{23} atoms of nitrogen

∴ Number of C atoms in 5.6 g of urea

$= (1 \times 0.0933)$ mol × 6.022×10^{23} atoms/mol

= 0.562×10^{23} atoms of carbon

∴ Number of O atoms in 5.6 g of urea

$= (1 \times 0.0933)$ mol × 6.022×10^{23} atoms/mol

= 0.562×10^{23} atoms of oxygen

Ans: 5.6 g of urea contain **2.247×10^{23} atoms of H, 1.124×10^{23} atoms of N, 0.562×10^{23} atoms of C and 0.562×10^{23} atoms of O.**

STRATEGY

Given: Mass of substance

1. Use formula:

   No. of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

2. Find molar mass of the substance
3. Substitute the values and calculate number of moles (n)
4. Use formula:

   Number of 'X' atoms = No. of 'X' atoms in a molecule of the substance × No. of moles of the substance × N_A

5. Substitute the values and find the number of atoms.

Q.99. Calculate the number of atoms of 'C', 'H' and 'O' in 72.5 g of isopropanol, C_3H_7OH (molar mass = 60 g mol^{-1}). **[3 Marks]**

Solution:

Given: Mass of isopropanol(C_3H_7OH) = 72.5 g

To find: The number of atoms of C, H, O

Calculation: Molecular formula of isopropanol, is C_3H_7OH.

Molar mass of C_3H_7OH = 60 g mol^{-1}

Number of moles

$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

$= \frac{72.5\ \text{g}}{60\ \text{g mol}^{-1}} = 1.208$ mol

∴ Moles of isopropanol = 1.21 mol

Number of atoms = Number of moles × Avogadro's constant

Now, 1 molecule of isopropanol has total 12 atoms, out of which 8 atoms are of H, 3 of C and 1 of O.

∴ Number of C atoms in 72.5 g isopropanol

$= (3 \times 1.208)$ mol × 6.022×10^{23} atoms/mol

= 2.182×10^{24} atoms of carbon

$\therefore$ Number of 'H' atoms in 72.5 g isopropanol
= (8 × 1.208) mol × 6.022 × 10^{23} atoms/mol
= **5.819 × 10^{24} atoms of hydrogen**

$\therefore$ Number of 'O' atoms in 72.5 g isopropanol = (1 × 1.208) mol × 6.022 × 10^{23} atoms/mol
= **7.274 × 10^{23} atoms of oxygen**

Ans: 72.5 g of isopropanol contain **2.182 × 10^{24} atoms of C, 5.819 × 10^{24} atoms of H and 7.274 × 10^{23} atoms of O.**

*__Q.100. Arjun purchased 250 g of glucose ($C_6H_{12}O_6$) for Rs 40. Find the cost of glucose per mole.__ **[3 Marks]**

Solution:

Given: Mass of urea = 250 g, cost for 250 g glucose = Rs 40, molecular formula of glucose = $C_6H_{12}O_6$

To find: Cost per mole of glucose

Calculation: Molecular formula of glucose is ($C_6H_{12}O_6$).

Molecular mass of glucose
= (6 × Average atomic mass of C)
+ (12 × Average atomic mass of H)
+ (6 × Average atomic mass of O)
= (6 × 12 u) + (12 × 1 u) + (6 × 16 u)
= 180 u

$\therefore$ Molar mass of glucose = 180 g mol^{-1}

Number of moles

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$$

$$= \frac{250\text{ g}}{180\text{ g mol}^{-1}} = \frac{250}{180}\text{ mol}$$

Now, $\frac{250}{180}$ mol of glucose cost = Rs 40

1 mol glucose cost = x

$\therefore$ $x = \frac{40 \times 180}{250}$ = **Rs 28.8/mol of glucose**

CALCULATION USING LOG TABLE

$$\frac{40 \times 180}{250}$$

= Antilog_{10} [$\log_{10}$ (40) + $\log_{10}$ (180) – $\log_{10}$ (250)]
= Antilog_{10} [1.6021 + 2.2553 – 2.3979]
= Antilog_{10} [1.4595] = 28.80

Ans: The cost of glucose per mole is **Rs 28.8**.

Practice Numericals

1. Antacid tablets contain 0.5 g of calcium. Calculate the moles of calcium in each tablet **[1 Mark]**

Ans: 0.0125 mol

2. A particular helium balloon contains 0.40 moles of He. How many grams and molecules of helium are in the balloon? **[2 Marks]**

Ans: i. 1.6 g
ii. 2.4 × 10^{23} atoms

1.9 MOLES AND GASES

*__Q.101. Explain: Molar volume of gas__ **[2 Marks]**

Ans:

i. It is more convenient to measure the volume rather than mass of the gas.

ii. It is found from Avogadro law that one mole of any gas occupies a volume of 22.4 dm^3 at **standard temperature (0 °C) and pressure (1 atm) (STP)**.

iii. The volume of 22.4 dm^3 at STP is known as **molar volume** of a gas.

iv. The relationship between number of moles and molar volume can be expressed as follows:

Number of moles of a gas (n)

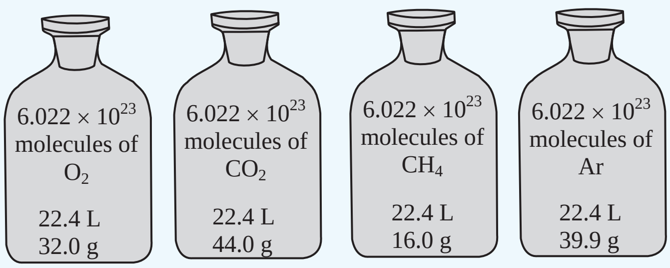
$$= \frac{\text{Volume of the gas at STP}}{\text{Molar volume of the gas}}$$

$$= \frac{\text{Volume of the gas at STP}}{22.4\text{ dm}^3\text{ mol}^{-1}}$$

Note: IUPAC has recently changed the standard pressure to 1 bar. Under these new STP conditions the molar volume of a gas is 22.71 L mol^{-1}.

FOR YOUR KNOWLEDGE

One mole of various gas (of different molar masses) occupy 22.4 L at STP and contain Avogadro's number of molecules.

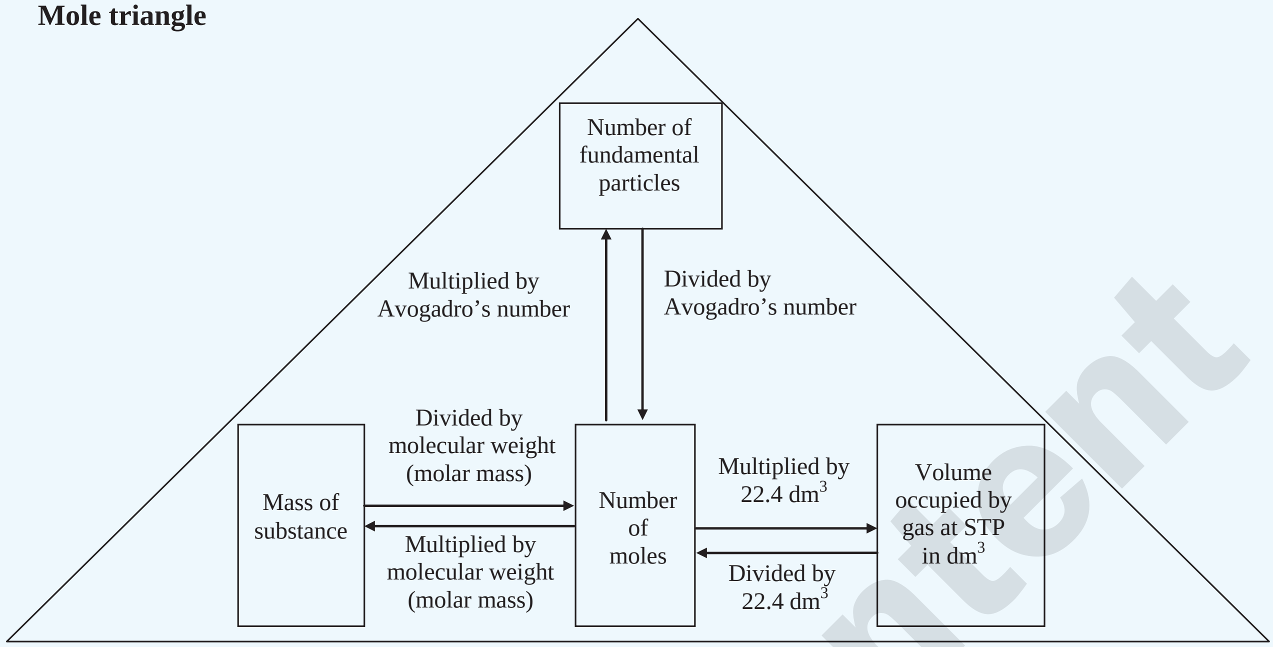


*__Q.102. What is meant by molar volume of a gas?__ **[1 Mark]**

Ans: The volume occupied by one mole of a gas at standard temperature (0 °C) and pressure (1 atm) (STP) is called as molar volume of a gas. The molar volume of a gas at STP is 22.4 dm^3.

FOR YOUR KNOWLEDGE

Mole triangle



*Q.103. **Activity:**

Collect information of various scientists and prepare charts of their contribution in chemistry.

Ans:

Scientists		Contributions
Joseph Louis Gay-Lussac (1778 – 1850) (French chemist and physicist)	i. ii. iii.	Formulated the gas law. Collected samples of air at different heights and recorded temperatures and moisture contents. Discovered that the composition of atmosphere does not change with increasing altitude.
Amedeo Avogadro (1776 – 1856) (Italian scholar)	i.	Published article in French journal on determining the relative masses of elementary particles of bodies and proportions by which they enter combinations.
	ii.	Published a research paper titled "New considerations on the theory of proportions and on determination of the masses of atoms."

*[**Note:** Students are expected to find out contributions of other scientists on their own.]*

Numerical Zone

*Q.104. **Calculate number of moles of hydrogen in 0.448 litre of hydrogen gas at STP.** **[2 Marks]**

Solution:

Given: Volume of hydrogen at STP = 0.448 L

To find: Number of moles of hydrogen

Formula: Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

Calculation: Molar volume of a gas = 22.4 dm^3 mol^{-1}

= 22.4 L at STP

Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

$$= \frac{0.448 \text{ L}}{22.4 \text{ L mol}^{-1}} = \mathbf{0.02\ mol}$$

Ans: Number of moles of hydrogen = **0.02 mol**

SMART CHECK

Convert your answer back to the original volume to see whether it matches.

Volume of the gas at STP

= Number of moles of a gas (n)

× Molar volume of a gas

= 0.02 mol × 22.4 dm^3 mol^{-1} = **0.448 dm^3**.

+Q.105. **Calculate the number of moles and molecules of ammonia (NH_3) gas in a volume 67.2 dm^3 of it measured at STP.**

(Problem 1.7 of Textbook page no.10) **[2 Marks]**

Solution:

Given: Volume of ammonia at STP = 67.2 dm^3

To find: Number of moles and molecules of ammonia

Formulae: i. Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

ii. Number of molecules = Number of moles $\times 6.022 \times 10^{23}$ molecules mol^{-1}

Calculation: Molar volume of a gas = 22.4 dm^3 mol^{-1} at STP.

Number of moles (n)

$$= \frac{\text{Volume of the gas at STP}}{\text{Molar volume of gas}}$$

Number of moles of NH_3

$$= \frac{67.2\ dm^3}{22.4\ dm^3\ mol^{-1}} = \mathbf{3.0\ mol}$$

Number of molecules = Number of moles $\times 6.022 \times 10^{23}$ molecules mol^{-1}

= 3.0 mol $\times 6.022 \times 10^{23}$ molecules mol^{-1}

= **18.066×10^{23} molecules**

Ans: Number of moles of ammonia = **3.0 mol**

Number of molecules of ammonia = **18.066×10^{23} molecules**

***Q.106. What is volume of carbon dioxide, CO_2 occupying by**

i. 5 moles and

ii. 0.5 mole of CO_2 gas measured at STP. **[2 Marks]**

Solution:

Given: i. Number of moles of CO_2 = 5 mol

ii. Number of moles of CO_2 = 0.5 mol

To find: Volume at STP

Formula: Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

Calculation: Molar volume of a gas = 22.4 dm^3 mol^{-1} at STP.

Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

∴ i. Volume of the gas at STP

= Number of moles of a gas (n) × Molar volume of a gas

= 5 mol × 22.4 dm^3 mol^{-1}

= **112 dm^3**

ii. Volume of the gas at STP

= Number of moles of a gas (n) × Molar volume of a gas

= 0.5 mol × 22.4 dm^3 mol^{-1}

= **11.2 dm^3**

Ans:

i. Volume of 5 mol of CO_2 = **112 dm^3**

ii. Volume of 0.5 mol of CO_2 = **11.2 dm^3**

SMART CHECK

Convert your answer back to the original moles to see whether it matches.

i. Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}} = \frac{112}{22.4} = 5 \text{ moles}$$

ii. Number of moles of a gas (n)

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}} = \frac{11.2}{22.4} = 0.5 \text{ mole}$$

Q.107. Try this *(Textbook page no. 10)*

Calculate the volume in dm^3 occupied by 60.0 g of ethane at STP. **[2 Marks]**

Solution:

Given: Mass of ethane at STP = 60.0 g

To find: Volume of ethane

Formulae: i. Number of moles

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of the substance}}$$

ii. Number of moles

$$= \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

Calculation: Molar volume of a gas = 22.4 dm^3 mol^{-1} at STP

Molecular mass of ethane = 30 g mol^{-1}

Number of moles

$$= \frac{\text{Mass of a substance}}{\text{Molar mass of the substance}}$$

$$= \frac{60.0\ g}{30\ g\ mol^{-1}} = 2 \text{ mol}$$

Number of moles of a gas (n)

$$= \frac{\text{Volume of the gas at STP}}{\text{Molar volume of a gas}}$$

∴ Volume of the gas at STP = Number of moles of a gas (n) × Molar volume of a gas

= 2 mol × 22.4 dm^3 mol^{-1} = **44.8 dm^3**

Ans: Volume of ethane = **44.8 dm^3**

Q.108. 3.40 g of ammonia at STP occupies volume of 4.48 dm^3. Calculate molar mass of ammonia. **[2 Marks]**

Solution:

Given: Mass of ammonia = 3.40 g

Volume at STP = 4.48 dm^3

To Find: Molar mass of ammonia

Calculation: Let '*x*' grams be the molar mass of NH_3.

Molar volume of a gas = 22.4 dm^3 mol^{-1} at STP.

Volume occupied by 3.40 g of NH_3 at S.T.P = 4.48 dm^3

Volume occupied by '*x*' g of NH_3 at S.T.P = 22.4 dm^3

$\therefore \quad x = \dfrac{22.4 \times 3.40}{4.48} = \mathbf{17.0\ g\ mol^{-1}}$.

Ans: Molar mass of ammonia is **17.0 g mol^{-1}.**

*__Q.109. Calculate the mass of potassium chlorate required to liberate 6.72 dm^3 of oxygen at STP. Molar mass of $KClO_3$ is 122.5 g mol^{-1}.__ **[3 Marks]**

Solution:

The molecular formula of potassium chlorate is $KClO_3$.

Required chemical equation:

$$\underset{[2\ \text{moles}]}{2KClO_3} \longrightarrow 2KCl + \underset{[3\ \text{moles}]}{3O_2 \uparrow}$$

2 moles of $KClO_3$ = 2 × 122.5 = 245 g

3 moles of O_2 at STP occupy = (3 × 22.4 dm^3) = 67.2 dm^3

Thus, 245 g of potassium chlorate will liberate 67.2 dm^3 of oxygen gas.

Let '*x*' gram of $KClO_3$ liberate 6.72 dm^3 of oxygen gas at S.T.P.

$\therefore \quad x = \dfrac{245 \times 6.72}{67.2} = \mathbf{24.5\ g}$

Ans: Mass of potassium chlorate required = **24.5 g**

CONNECTIONS

You will study in chapter 2 about chemical reactions and stoichiometric calculations

Practice Numericals

1. Calculate number of moles of hydrogen in 0.896 litre of helium gas at STP. **[2 Marks]**

Ans: 0.04 mol

2. Calculate the volume in dm^3 occupied by 34.0 g of methane at STP. **[2 Marks]**

Ans: 44.8 dm^3

Brain Teasers

Q.110. Veg puffs from a particular bakery have an average mass of 27.0 g, whereas egg puffs from the same bakery have an average mass of 40 g.

i. Suppose a person buys 1 kg of veg puff from the bakery. Calculate the number of veg puffs he receives.

ii. Determine the mass of egg puffs (in kg) that will contain the same number of eggs puffs as in one kilogram of veg puffs.

Solution:

i. Mass of a veg puff = 27.0 g = 0.027 kg

$\therefore$ Number of veg puffs in 1 kg = 1 / 0.027 = **37**

ii. One kilogram of veg puffs contains 37 veg puffs.

Mass of 37 egg puffs = 37 × 0.040 = **1.48 kg**

Ans:

i. **37** veg puffs in 1 kg of puff.

ii. Mass of 37 egg puffs is **1.48 kg**

Quick Review

➢ **Classification of matter (On basis of chemical composition):**

Matter
Occupies space and has mass

- **Pure substances**
  Have a definite chemical composition
  - **Elements**
    Cannot be broken down into simpler substances by ordinary chemical changes **e.g.** Gold, silver etc.
  - **Compounds**
    Two or more elements combine in fixed proportion
    e.g. Water, ammonia, etc.
- **Mixtures**
  Have no definite chemical composition and hence no definite properties
  - **Homogeneous mixture**
    Constituents are uniformly mixed throughout its bulk **e.g.** Solution
  - **Heterogeneous mixture**
    Constituents are not uniformly mixed throughout the bulk. **e.g.** Suspension

➢ **Laws of chemical combination:**

Laws of chemical combination

- **Law of conservation of mass**
  Mass can neither be created nor destroyed.
- **Law of definite proportion**
  A given compound always contains exactly the same proportion of elements by weight.
- **Law of multiple proportion**
  When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers.
- **Gay Lussac's law of gaseous volume**
  When gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.
- **Avogadro's law**
  Equal volumes of all gases at the same temperature and pressure contain equal number of molecules.

Important Formulae

1. Celsius to Fahrenheit: $°F = \frac{9}{5}(°C) + 32$

2. Celsius to Kelvin: $K = °C + 273.15$

3. Average atomic mass = $\frac{\text{Sum of (Isotopic mass} \times \text{\% Abundance)}}{100}$

4. Number of moles (n) = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}}$

5. Number of molecules
   = Number of moles × Avogadro number
   = Number of moles × 6.022×10^{23}

6. Number of moles (n) = $\frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$
   $= \frac{\text{Volume of a gas at STP}}{22.4\ dm^3\ mol^{-1}}$

Exercise

Theory Questions for Practice

1.2 Nature of chemistry

1. What are pure substances? Give two examples. **[2 Marks]**

Ans: *Refer Q.6. (i)*

2. What are metalloids? **[1 Mark]**

Ans: *Refer Q.6. (i-a-3)*

3. What is a homogeneous mixture? **[1 Mark]**

Ans: *Refer Q.6. (ii-a)*

4. Explain classification of matter. **[4 Marks]**

Ans: *Refer Q.6.*

5. Give one example of each: **[½ Mark Each]**
   i. Heterogeneous mixture
   ii. Compound
   iii. Element
   iv. Homogeneous mixture

Ans: *Refer Q.9.*

1.3 Properties of matter and their measurement

6. Give SI unit of: **[½ Mark Each]**
   i. Temperature ii. Mass
   iii. Length

Ans: *Refer Q.17.*

1.4 Laws of chemical combination

7. State and explain the law of definite proportion. **[3 Marks]**

Ans: *Refer Q.38.*

8. i. State the law of conservation of mass.
   ii. Explain the law of multiple proportions with reference to carbon monoxide and carbon dioxide.
   iii. State Gay-Lussac's law of gaseous volume. **[4 Marks]**

Ans:
i. *Refer Q.37.(i)* ii. *Refer Q.42.*
iii. *Refer Q.44. (ii) (Statement)*

9. Give two examples to explain Gay-Lussac's law of gaseous volume. **[2 Marks]**

Ans: *Refer Q.46.*

1.5 Avogadro law

10. State Avogadro's law. **[1 Mark]**

Ans: *Refer Q.52. (ii)*

1.6 Dalton's atomic theory

11. What were the basic assumptions of Dalton's theory? **[2 Marks]**

Ans: *Refer Q.54.*

12. What happens during a chemical reaction according to Dalton's atomic theory? **[1 Mark]**

Ans: *Refer Q.54. (iv)*

1.7 Atomic and molecular masses

13. Why is it impossible to measure the mass of a single atom? **[1 Mark]**

Ans: *Refer Q.60. (i)*

14. Explain the term molecular mass with an example. **[2 Marks]**

Ans: *Refer Q.64 and Q.65.*

1.8 Mole concept and molar mass

15. Define one mole. **[1 Mark]**

Ans: *Refer Q.84. (ii)*

16. Define molar mass. **[1 Mark]**

Ans: *Refer Q.86. (i)*

Additional Numericals for Practice

1.3 Properties of matter and their measurement

1. Convert the following degree Celsius temperature to degree Fahrenheit. **[2 Marks]**

i. 50 °C ii. 85 °C

Ans: i. 104 °F ii. 185 °F

1.7 Atomic and molecular masses

2. Calculate the atomic mass (average) of chlorine using the following data: **[2 Marks]**

	% Natural abundance	Atomic mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Ans: 35.4528 u

3. Find the formula mass of $Pb(Cr_2O_7)_2$ and $AgNO_3$. (Atomic mass of Pb = 207 u, Cr = 52 u, Ag = 108 u, O = 16 u, N = 14 u) **[2 Marks]**

Ans. 639 u, 170 u

1.8 Mole concept and molar mass

4. In five moles of acetic acid (CH_3COOH), calculate the following:

i. Number of moles of carbon
ii. Number of moles of hydrogen
iii. Number of moles of oxygen
iv. Number of molecules of acetic acid **[4 Marks]**

Ans:

i. 10 mol ii. 20 mol
iii. 10 mol
iv. 3.011×10^{24} molecules

5. Calculate the number of moles of NaOH in

i. 60 g and
ii. 20 g of the compound.

(Average atomic masses of Na = 23, O = 16, H = 1) **[2 Marks]**

Ans:

i. 1.5 mol ii. 0.5 mol

6. Calculate the number of moles and molecules of urea present in 30 g of urea. **[3 Marks]**

Ans:

i. 0.5 mol
ii. 3.011×10^{23} molecules

7. Calculate the number of molecules in 28 g of nitrogen gas, 64 g of oxygen gas and 72 g of water. **[3 Marks]**

Ans: Nitrogen gas - 6.022×10^{23} molecules
Oxygen gas - 1.2044×10^{24} molecules
Water - 2.4088×10^{24} molecules

8. How many atoms of sulphur are present in 0.1 mole of S_8 molecules? **[2 Marks]**

Ans: 4.82×10^{23} atoms

9. Calculate the mass (in grams) of : **[2 Marks]**

i. 3 moles of H_2O ii. 5 moles of He

Ans:

i. 54 g ii. 6 g

1.9 Moles and gases

10. Calculate number of moles of ethane in 5.6 L of ethane gas at STP. **[2 Marks]**

Ans: 0.25 mol

11. How many moles of nitrogen gas are there in a 16,500 mL sample of nitrogen gas at STP? **[2 Marks]**

Ans: 0.25 mol

12. Calculate the volume in dm^3 of the following at STP: **[2 Marks]**

i. 6 moles of oxygen gas
ii. 1.6 g of oxygen gas

Ans:

i. 134.4 dm^3 ii. 1.12 dm^3

13. Calculate the number of moles and molecules of ammonia (NH_3) gas in a volume 89.6 dm^3 of it measured at STP. **[2 Marks]**

Ans:

i. 4.0 mol
ii. 2.4088×10^{24} molecules

Multiple Choice Questions

[1 Mark Each]

1. The branch of chemistry which deals with carbon compounds is called _______ chemistry.
(A) organic (B) inorganic
(C) carbon (D) bio

2. A/an _______ is a simple combination of two or more substances in which the constituent substances retain their separate identities.
(A) compound (B) mixture
(C) element (D) All of these

3. Which one of the following is NOT a mixture?
(A) Paint
(B) Gasoline
(C) Liquefied Petroleum Gas (LPG)
(D) Distilled water

*4. SI unit of the quantity electric current is _______.
(A) Volt (B) Ampere
(C) Candela (D) Newton

*5. Which of the following temperature will read the same value on celsius and Fahrenheit scales?
(A) – 40° (B) + 40°
(C) – 80° (D) – 20°

6. The sum of the masses of reactants and products is equal in any physical or chemical reaction. This is in accordance with _______.
(A) law of multiple proportion
(B) law of definite composition
(C) law of conservation of mass
(D) law of reciprocal proportion

*7. A sample of pure water, whatever the source always contains _______ by mass of oxygen and 11.1 % by mass of hydrogen.
(A) 88.9 (B) 18 (C) 80 (D) 16

8. A sample of calcium carbonate ($CaCO_3$) has the following percentage composition:
Ca = 40 %; C = 12 %; O = 48 %
If the law of definite proportions is true, then the weight of calcium in 4 g of a sample of calcium carbonate from another source will be _______.
(A) 0.016 g (B) 0.16 g
(C) 1.6 g (D) 16 g

*9. Which of the following compounds CANNOT demonstrate the law of multiple proportions?
(A) NO, NO_2 (B) CO, CO_2
(C) H_2O, H_2O_2 (D) Na_2S, NaF

10. Two elements, A and B, combine to form two compounds in which 'a' g of A combines with 'b_1' and 'b_2'g of B respectively. According to law of multiple proportion _______.
(A) $b_1 = b_2$
(B) b_1 and b_2 bear a simple whole number ratio
(C) a and b_1 bear a whole number ratio
(D) no relation exists between b_1 and b_2

11. At constant temperature and pressure, two litres of hydrogen gas react with one litre of oxygen gas to produce two litres of water vapour. This is in accordance with _______.
(A) law of multiple proportion
(B) law of definite composition
(C) law of conservation of mass
(D) law of gaseous volumes

*12. In the reaction $N_2 + 3H_2 \longrightarrow 2NH_3$, the ratio by volume of N_2, H_2 and NH_3 is 1 : 3 : 2. This illustrates the law of _______.
(A) definite proportion
(B) reciprocal proportion
(C) multiple proportion
(D) gaseous volumes

13. One mole of oxygen molecule weighs _______.
(A) 8 g (B) 32 g
(C) 16 g (D) 6.022×10^{23} g

*14. How many g of H_2O are present in 0.25 mol of it?
(A) 4.5 (B) 18 (C) 0.25 (D) 5.4

15. The mass of 0.002 mol of glucose ($C_6H_{12}O_6$) is _______.
(A) 0.20 g (B) 0.36 g
(C) 0.50 g (D) 1.80 g

*16. Which of the following has the largest number of atoms?
(A) 1 g $Au_{(s)}$ (B) 1 g $Na_{(s)}$
(C) 1 g $Li_{(s)}$ (D) 1 g $Cl_{2(g)}$

17. Which of the following is CORRECT?
(A) 1 mole of oxygen atoms contains 6.0221367×10^{23} atoms of oxygen.
(B) 1 mole of water molecules contains 6.0221367×10^{23} molecules of water.
(C) 1 mole of sodium chloride contains 6.0221367×10^{23} formula units of NaCl.
(D) All of these

18. 180 g of glucose ($C_6H_{12}O_6$) contains _______ carbon atoms.
(A) 1.8×10^{23} (B) 1.8×10^{24}
(C) 3.6×10^{23} (D) 3.6×10^{24}

19. The number of molecules present in 8 g of oxygen gas is _______.
(A) 6.022×10^{23} (B) 3.011×10^{23}
(C) 12.044×10^{23} (D) 1.505×10^{23}

*20. Which of the following has maximum number of molecules?
(A) 7 g N_2 (B) 2 g H_2
(C) 8 g O_2 (D) 20 g NO_2

21. The number of molecules in 22.4 cm^3 of ozone gas at STP is _______.
(A) 6.022×10^{20} (B) 6.022×10^{23}
(C) 22.4×10^{20} (D) 22.4×10^{23}

*22. The number of molecules in 22.4 cm^3 of nitrogen gas at STP is _______.
(A) 6.022×10^{20} (B) 6.022×10^{23}
(C) 22.4×10^{20} (D) 22.4×10^{23}

23. 11.2 cm^3 of hydrogen gas at STP, contains _______ moles.
(A) 0.0005 (B) 0.01
(C) 0.029 (D) 0.5

24. The mass of 224 mL of hydrogen gas at STP is _______.
(A) 0.02 g (B) 0.224 g
(C) 2.24 g (D) 20.0 g

25. 4.4 g of an unknown gas occupies 2.24 L of volume under STP conditions. The gas may be_______.
(A) CO_2 (B) CO (C) O_2 (D) SO_2

Answers to Multiple Choice Questions:

1.	(A)	2.	(B)	3.	(D)	4.	(B)
5.	(A)	6.	(C)	7.	(A)	8.	(C)
9.	(D)	10.	(B)	11.	(D)	12.	(D)
13.	(B)	14.	(A)	15.	(B)	16.	(C)
17.	(D)	18.	(C)	19.	(D)	20.	(B)
21.	(A)	22.	(A)	23.	(A)	24.	(A)
25.	(A)						

Competitive Corner

1. Amongst the following statements, that which was not proposed by Dalton was _______. **[JEE (Main) 2020]**
(A) all the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
(B) matter consists of indivisible atoms.
(C) when gas combine or reproduced in a chemical reaction they do so in a simple ratio by volume provided all gases are at the same T and P.
(D) chemical reaction involve reorganization of atoms. These are neither created nor destroyed in a chemical reaction.

2. "A given compound always contains exactly the same proportion of elements by weight" is a statement of _______. **[MHT CET 2021]**
(A) Law of combining volumes of gases
(B) Law of conservation of mass
(C) Law of multiple proportion
(D) Law of definite proportion

3. What is the density of water in kg dm^{-3} if its density in g cm^{-3} is 0.863 **[MHT CET 2022]**
(A) 7.86 (B) 0.863 (C) 8.63 (D) 4.60

4. What volume of $CO_{2(g)}$ at STP is obtained by complete combustion of 6 g carbon? **[MHT CET 2023]**
(A) 22.4 dm^3 (B) 11.2 dm^3
(C) 5.6 dm^3 (D) 2.24 dm^3

5. Which of the following pair of compounds demonstrates the law of multiple proportions? **[MHT CET 2023]**
(A) CH_4, CCl_4 (B) BF_3, NH_3
(C) CO, CO_2 (D) NO_2, CO_2

6. The **right** option for the mass of CO_2 produced by heating 20 g of 20% pure limestone is (Atomic mass of Ca = 40)
[$CaCO_3 \xrightarrow{1200\,K} CaO + CO_2$] **[NEET (UG) 2023]**
(A) 1.76 g (B) 2.64 g
(C) 1.32 g (D) 1.12 g

7. 1 gram of sodium hydroxide was treated with 25 mL of 0.75 M HCl solution, the mass of sodium hydroxide left unreacted is equal to **[NEET (UG) 2024]**
(A) Zero mg (B) 200 mg
(C) 750 mg (D) 250 mg

8. The highest number of helium atoms is in **[NEET (UG) 2024]**
(A) 4 g of helium
(B) 2.271098 L of helium at STP
(C) 4 mol of helium
(D) 4 u of helium

Answers to Competitive Corner:

1.	(C)	2.	(D)	3.	(B)	4.	(B)
5.	(C)	6.	(A)	7.	(D)	8.	(C)

Topic Test

Time: 1 Hour 30 Min **Total Marks: 25**

SECTION A

Q.1. Select and write the correct answer: **[04]**

i. In sodium chloride crystal, one Na^+ ion is surrounded by _______ Cl^- ions.
(A) three (B) four (C) six (D) eight

ii. The number of molecules in 11.2 cm^3 of nitrogen gas at STP is _______.
(A) 3.011×10^{20} (B) 3.011×10^{23} (C) 22.4×10^{20} (D) 22.4×10^{23}

iii. The SI unit of luminous intensity is _______.
(A) Volt (B) Ampere (C) Candela (D) Newton

iv. How many g of H_2O are present in 0.3 mol of it?
(A) 3.6 (B) 18 (C) 0.25 (D) 5.4

Q.2. Answer the following: **[03]**

i. Find the molecular mass of H_2SO_4. (given: atomic mass of H = 1 u, O = 16 u, S = 32 u)

ii. State the law of conservation of mass.

iii. Give the SI unit of density.

SECTION B

Attempt any Four: **[08]**

Q.3. How are mixtures classified?

Q.4. In two moles of acetaldehyde (CH_3CHO), calculate the following:
i. Number of moles of carbon ii. Number of moles of hydrogen

Q.5. Give reason: The mass of a body is more fundamental property than its weight.

Q.6. Write a short note on mole concept.

Q.7. 45.4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed. The reaction is given below:
$2N_{2(g)} + O_{2(g)} \longrightarrow 2N_2O_{(g)}$
Which law is being obeyed in this experiment? Write the statement of the law?

Q.8. Complete the following table.

No.	Substance	Element or compound
i.	Helium gas	------
ii.	Nitrogen gas	------
iii.	Water	------
iv.	Table salt	------

SECTION C

Attempt any Two: **[06]**

Q.9. Explain the term formula mass with an example.

Q.10. Calculate the number of moles and molecules of acetic acid present in 22 g of it.

Q.11. State and explain Avogadro's law.

SECTION D

Attempt any One: **[04]**

Q.12. i. State the law of multiple proportions.

ii. Calculate number of atoms in 1.6 g of sulphur. (Average atomic mass: S = 32 u)

iii. What is the ratio of molecules in 1 mole of NH_3 and 1 mole of HNO_3?

Q.13. i. Explain: The need of the term average atomic mass.

ii. Calculate number of moles of hydrogen in 4.48 litre of hydrogen gas at STP.

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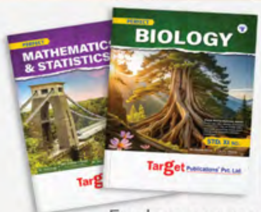

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