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.

Target Publications® Pvt. Ltd.

Written as per the latest textbook prescribed by the Maharashtra State Bureau of Textbook Production and Curriculum Research, Pune.

# PERFECT CHEMISTRY (Vol. I)

## Std. XI Sci.

### Salient Features

- Written as per the new textbook
- Subtopic-wise segregation for powerful concept building
- Complete coverage of Textual Exercise Questions, Intext Questions and Numericals
- Extensive coverage of New Type of Questions
- 'Solved Examples' guide you through every type of problem
- 'Apply Your Knowledge' section for application of concepts
- 'Quick Review' at the end of every chapter facilitates quick revision
- A compilation of all 'Important Formulae'
- 'Competitive Corner' presents questions from prominent competitive examinations
- Reading Between the Lines, Enrich Your Knowledge, Gyan Guru, Connections, NCERT Corner are designed to impart holistic education

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## PREFACE

*"I never teach my pupils; I only attempt to provide the conditions in which they can learn." – Albert Einstein*

**"Chemistry: Std. XI Volume - 1"** forms a part of **'Target Perfect Notes'** prepared as per the new textbook. It focuses on active learning along with making the process of education more interesting and builds up the students' knowledge quotient in the process.

The **subtopic-wise** classified format for each chapter of this book helps the students to comprehend concepts easily. Every chapter begins with the coverage of all textual content in the format of Objectives, Question-Answers, Give Reasons, Numericals, Short Notes, Diagram related questions and a host of other Objective and Subjective type of questions. The questions titled under 'Use your brain power', 'Can you tell', 'Can you recall', 'Problems' and various similar titles pave the way for a robust concept building. For the students to gain a better understanding of the concept lying behind the answer, 'Reading between the lines' (*not a part of the answer*) has been provided as deemed necessary. We have provided QR codes to access a video for a better understanding of the concept.

While ensuring complete coverage of the syllabus in an effortless and easy to grasp format, emphasis is also given on active learning. To achieve this, we have infused several sections such as, ***Gyan Guru, Enrich Your Knowledge, Connections***, ***Reading between the lines*** and ***NCERT Corner***, and additional sections such as**,** ***Apply Your Knowledge, Quick Review, Important Formulae, Exercise*** and ***Competitive Corner.*** The following screenshots will walk you through the core features of this book and elucidate how they have been carefully designed to maximize the student learning.

*GG - Gyan Guru*

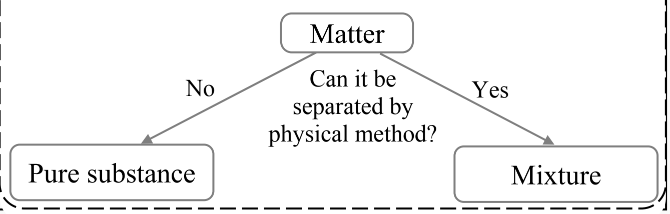
***Your weight in deep outer space is ZERO!!***

*The mass of a body remains constant irrespective its position. However, the weight of a body*

**Gyan Guru** illustrates real life applications or examples related to the concept discussed.
*This is our attempt to link learning to the life.*

**Enrich your knowledge** presents fascinating information about the concept covered.
*This is our attempt to create interest within the students about the concept.*

**Enrich Your Knowledge**




Connections

*You studied in chapter 1 on calculating molar mass of a compound using atomic masses of constituent elements.*

**Connections** enables students to interlink concepts covered in different chapters.
*This is our attempt to enable students to comprehend the subject as a whole.*

**Reading between the lines** provides for concept elaboration
*This is our attempt to help students to understand the underlying concept behind an answer.*

### Reading between the lines

*Many metals show variable oxidation numbers in their compounds. Therefore, in their molecular formulae, the oxidation numbers are often represented by Roman numbers in parentheses after the chemical symbol of the metal.* ***e.g.*** *Au(III)$Cl_3$, Sn(II)$Cl_2$, Hg(II)$Cl_2$, etc.*

### NCERT Corner

**Ostwald's process**

Ostwald's process is used to prepare nitric acid on a large scale.

i. This method is based upon catalytic oxidation of $NH_3$ by atmospheric oxygen.

**NCERT Corner** covers additional information from NCERT textbook relevant to the topic
*This is our attempt to bridge the gap between NCERT and State Board textbook, thereby helping students to prepare for National level competitive examinations.*

**QR code** provides access to videos that boost conceptual understanding.
*This is our attempt to facilitate learning with visual aids.*

***[Note:** Students can scan the adjacent QR code to get conceptual clarity with the aid of a relevant video.**]***

### Apply Your Knowledge

**Q.73. Chalcopyrite ($CuFeS_2$) is a common ore of copper. Since it has low concentration of copper, the ore is first concentrated through froth floatation process. The concentrated ore is then heated strongly with silicon dioxide (silica) and oxygen in a furnace. The product**

**Apply your knowledge** includes challenging questions.
*This is our attempt to take students one step further and challenge their conceptual understanding.*

**Quick review** includes tables/ flow chart to summarize the key points in chapter.
*This is our attempt to help students to reinforce key concepts*

### Quick Review

**Classical theory**
Matter is composed of particles

↓ ***Einstein*** ***Planck***

Energy is Quantized.

## Important Formulae

1. Celsius to Fahrenheit: $°F = \frac{9}{5}(°C) + 32$
2. Celsius to Kelvin: $K = °C + 273.15$
3. Average atomic mass = $\frac{\text{Sum of (Isotopic mass} \times \text{\% Abundance)}}{100}$

**Important Formulae** includes all of the key formulae in the chapter.
*This is our attempt to enable students to see all the important formulae in the chapter at a glance.*

## Exercise

**Exercise** includes subtopic-wise additional questions and problems
*This is our attempt to provide additional practice questions that involve conceptual application from the topics across the entire chapter.*

**8.1 Hydrogen**

1. Name any two isotopes of hydrogen.

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

2. Write the chemical equation for the preparation of

**Ans:** *Refer Q. 15.*

## Competitive Corner

1. Which among the following is correct for electrolysis of brine solution?
**[MHT CET 2019]**
(A) $Cl_2$ gas is liberated at cathode
(B) Sodium metal is collected at anode
**(C) $H_2$ gas is liberated at cathode**

**Competitive Corner** presents questions from prominent competitive exams based entirely on the syllabus covered in the chapter.
*This is our way of providing students a competitive edge*

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

*A book affects eternity; one can never tell where its influence stops.*

*Best of luck to all the aspirants!*

From,
Publisher

**Edition:** First

# CONTENTS



**Note:**
1. * mark represents Textual question.
2. # mark represents Intext question.
3. +mark intext problems
4. symbol represents textual questions that need external reference for an answer

# 1 Some Basic Concepts of Chemistry

## Contents and Concepts



## 1.1 Introduction

**Q.1. Define chemistry.**
**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?**
**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.**
**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?**
**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:**
**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)**
**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?**
**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.**
**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**

**i. Homogeneous mixture** **ii. Heterogeneous mixture**
**iii. Element** **iv. Compound**

**Ans:**

i. Homogeneous mixture: Solution
ii. Heterogeneous mixture: Suspension
iii. Element: Gold
iv. Compound: Distilled water

**Q.10. Distinguish between**

**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. |

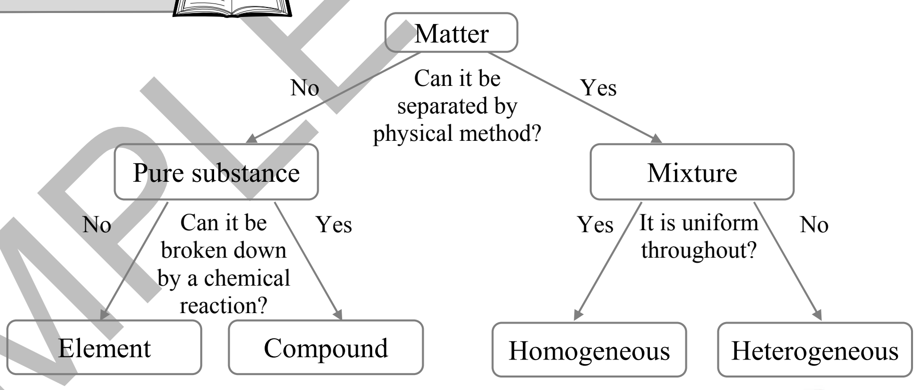
**ii.**

| | **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?**

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

**Enrich Your Knowledge**





Pure silver


Common salt


Vinegar


Tomato sauce

**Q.12. Explain: States of matter**

**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**

**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?**

**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**

**Ans:** *The arbitrarily decided and universally accepted standards are called **units**.*
**e.g.** Metre (m), kilogram (kg).

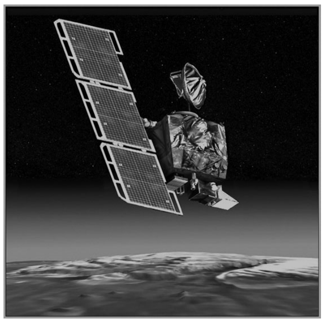
**Q.16. What are the various systems in which units are expressed?**

**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.**

**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?**

**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?**

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

**Q.20. Name the following:**

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

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

**Ans:**

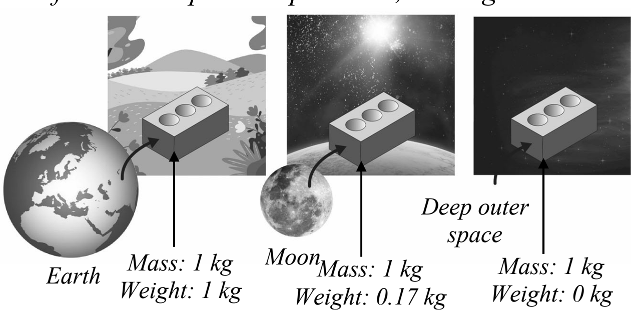
i. Mass is an inherent property of matter and is the measure of the quantity of matter of a body.
ii. The mass of a body does not vary with respect to its position.
iii. On the other hand, the weight of a body is a result of the mass and gravitational attraction
iv. Weight varies because the gravitational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

***Your weight 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?**

**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 mg = 1000 g = $10^6$ kg**]***

**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?**

**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**

**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.**

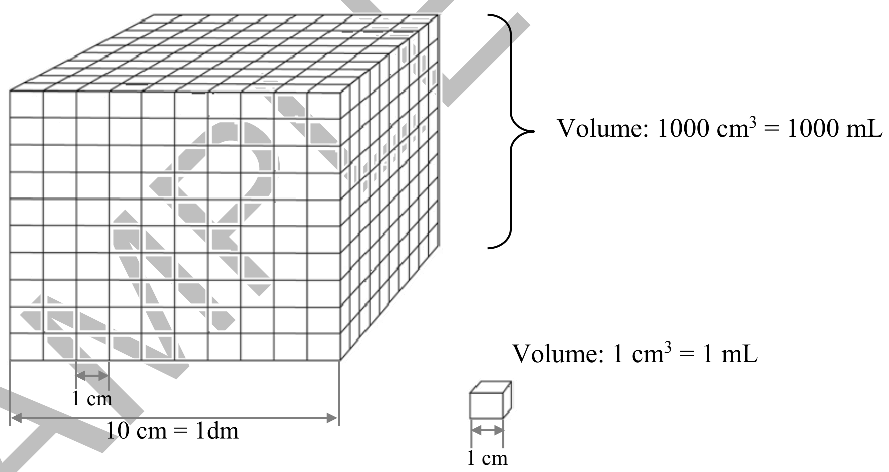
**Ans:** The common unit used for the measurement of volume of liquids and gases is litre (L).

**Q.26. How is the SI unit of volume expressed?**

**Ans:** The SI unit of volume is expressed as $(\text{metre})^3$ or $m^3$.

**Enrich 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.**

**Ans:**

i. Graduated cylinder

ii. Burette

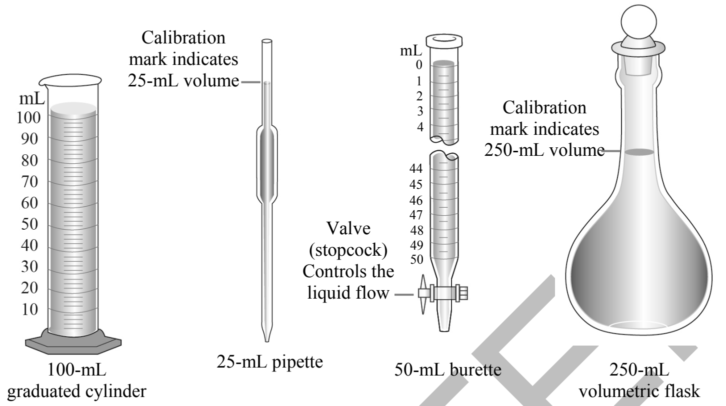
iii. Pipette

**Q.28. What is a volumetric flask used for in laboratory?**

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

**Enrich 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?**

**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.**

**Ans:**

i. The SI unit of density is derived as follows:

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

ii. CGS unit of density: g $cm^{-3}$

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

**Q.31. State three common scales of temperature measurement.**

**Ans:**

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

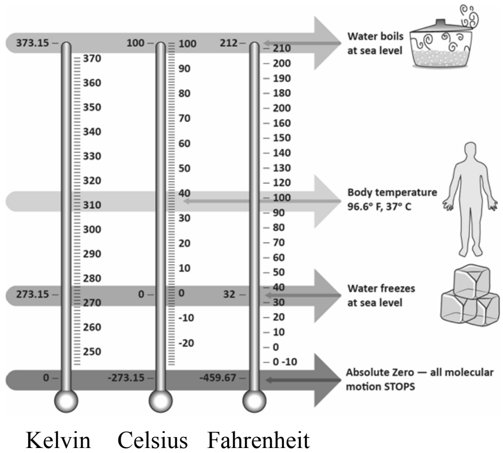
**Q.32. State the temperatures in Fahrenheit scale that corresponds to 0 °C and 100 °C.**

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

**Enrich 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:**
**i. Degree Fahrenheit and Degree Celsius** **ii. Kelvin and Degree Celsius**

**Ans:**

i. The relationship between degree Fahrenheit and degree Celsius is expressed as,

$°F = \frac{9}{5}(°C) + 32$

ii. The relationship between Kelvin and degree Celsius is expressed as,

K = °C + 273.15

## Solved Examples

***Q.34. Convert the following degree Celsius temperature to degree Fahrenheit.**
**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

**= 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

**= 86 °F**

**Ans:** i. The temperature 40 °C corresponds to **104 °F**.
ii. The temperature 30 °C corresponds to **86 °F**.

**Q.35. Convert the following degree Fahrenheit temperature to degree Celsius.**
**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,

$°F = \frac{9}{5}(°C) + 32$

$50 = \frac{9}{5}(°C) + 32$

$°C = \frac{(50 - 32) \times 5}{9}$

= **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**.

## 1.4 Laws of chemical combination

**Q.36. What is a chemical combination?**
**Ans:**
i. The process in which the elements combine with each other to form compounds is called **chemical combination**.
ii. 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.**
**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 who is often referred to as the father of modern chemistry performed careful experimental studies for various combustion reactions, namely burning of phosphorus and mercury in the presence of air.
iii. Both his experiments resulted in increased weight of products.
iv. 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.
v. 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.**
**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****.*

**Enrich 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.**

**Ans:** 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"*.

**Q.40. State and explain the law of multiple proportions.**

**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. Verify the law of multiple proportions for the chemical reaction between nitrogen and oxygen.**

**Ans:** Nitrogen and oxygen combine to form two compounds, nitric oxide and nitrogen dioxide.

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

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

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. Verify the law of multiple proportions for the chemical reaction between carbon and oxygen.**

**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. Verify the law of multiple proportions for the chemical reaction between sulphur and oxygen.**

**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.**

**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.

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

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

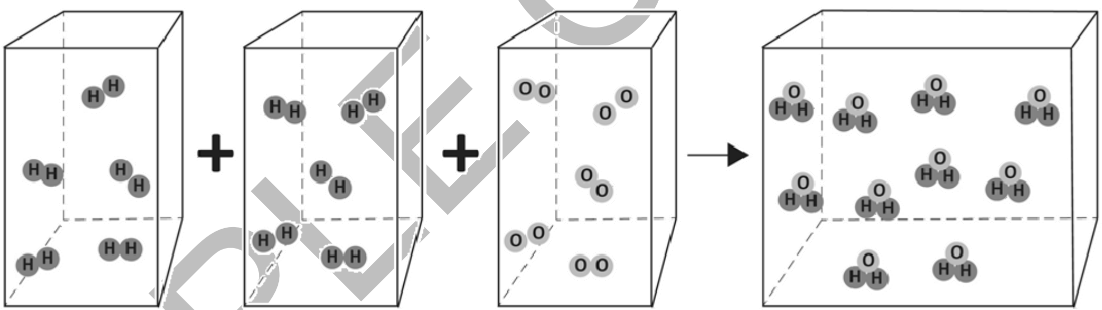
b. Under identical conditions of temperature and pressure, 1 L of nitrogen gas combine with 3 L of hydrogen gas to produce 2 L of ammonia gas.

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

Thus, the simple ratio of volumes is 1:3:2.

**Enrich 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:




**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?**

**Ans:** 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.**

**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.

$$\underset{\substack{[1\text{ L}]\\[1\text{ vol}]}}{\text{Hydrogen}} + \underset{\substack{[1\text{ L}]\\[1\text{ vol}]}}{\text{Chlorine}} \longrightarrow \underset{\substack{[2\text{ L}]\\[2\text{ vol}]}}{\text{Hydrogen chloride}}$$

Thus, the ratio of volumes is 1:1:2
This is in accordance with Gay Lussac's law.

ii. Under identical conditions of temperature and pressure, 200 mL sulphur dioxide combine with 100 mL oxygen to form 200 mL sulphur trioxide.

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

Thus, the ratio of volumes is 2:1:2.
This is in accordance with Gay Lussac's law.

**Enrich 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.

***Q.47. State and explain Avogardro's law.**

**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.48. Match the following:**

| | Law | | Statement |
|---|---|---|---|
| i. | Law of definite proportions | a. | 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. | Gay Lussac's law | b. | Equal volumes of all gases at the same temperature and pressure contain equal number of molecules |
| iii. | Law of multiple proportions | c. | 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 |
| iv. | Avogardro's law | d. | A given compound always contains exactly the same proportion of elements by weight |

**Ans:** i – d, ii – c, iii – a, iv – b

## Solved Examples

***Q.49. 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?**

***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.50. 24 g of carbon reacts with some oxygen to make 88 grams of carbon dioxide. Find out how much oxygen must have been used.**

***Solution:***

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

*To find:* Mass of oxygen (reactant)

*Calculation:* 12 g of carbon combine with 32 g oxygen to form 44 g of carbon dioxide as follows:

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

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**

**Q.51. 32 g of oxygen reacts with some carbon to make 56 grams of carbon monoxide. Find out how much mass must have been used.**

***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:

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

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**

***Q.52. 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).**

***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**

**Q.53. 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).**

***Solution:***

*Given:* Mass of sulphur (reactant) = 64 g

*To find:* Mass of sulphur dioxide (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**

## 1.5 Dalton's atomic theory

**Q.54. State and explain Dalton's atomic theory.**

**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.**

**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.**

**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.

## 1.6 Atomic and molecular masses

**Q.57. Can you recall?** *(Textbook page no. 6)*
**What is an atom and molecule? What is the order of magnitude of mass of one atom? What are isotopes?**

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

***Q.59. How many grams does an atom of hydrogen weigh?**
**Ans:** The mass of a hydrogen atom is $1.6736 \times 10^{-24}$ g.

**Q.60. How is relative atomic mass of an atom measured?**
**Ans:**

i. The mass of a single atom is extremely small, i.e. the mass of a hydrogen atom is $1.6736 \times 10^{-24}$ 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 international agreement.

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 twelth of the mass of one carbon-12 atom.

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

**Reading between the lines**

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

$$1\ amu = \frac{1}{12} \times mass\ of\ one\ C\text{-}12$$
$$= \frac{1}{12} \times 1.992648 \times 10^{-23}\ g$$
$$= 1.66056 \times 10^{-24}\ g$$

**Q.61. What is meant by Unified Mass unit?**
**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?**
**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.**
**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 | $2 \times 10^{-10}$ |

*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.]*

**Enrich 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.

| 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**

**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 example.**

**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**

**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**

**Ans:**

i. 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.

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

iii. 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:**

| 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 | $^{12}C$, $^{13}C$, $^{14}C$ |
| Formula mass of NaCl | **58.5 u** |

**Q.69. State TRUE or FALSE. If false, correct the statement.**

**i. An atom of carbon-12 is assigned a mass of exactly 1.00 u.**

**ii. Recently, amu has been replaced by unified mass unit called Dalton.**

**iii. Isotopes have same atomic mass.**

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

**Ans:**

i. False
An atom of carbon-12 is assigned a mass of exactly 12.00000 u.

ii. True

iii. False
Isotopes have different atomic masses.

iv. True

## Solved Examples

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

***Solution:***

Mass of an atom of oxygen in gram is $26.56896 \times 10^{-24}$ g.

$1.66056 \times 10^{-24}$ g = 1 u

$\therefore$ $26.56896 \times 10^{-24}$ g = $x$

$\therefore$ $x = \dfrac{26.56896 \times 10^{-24}\ \text{g}}{1.66056 \times 10^{-24}\ \text{g/u}}$ = **16.0 u**

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

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

***Solution:***

Mass of an atom of hydrogen in gram is $1.6736 \times 10^{-24}$ g.

$1.66056 \times 10^{-24}$ g = 1 u

∴ $1.6736 \times 10^{-24}$ g = $x$

∴ $x = \dfrac{1.6736 \times 10^{-24}\ \text{g}}{1.66056 \times 10^{-24}\ \text{g/u}}$ = **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?**

***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?**

***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:**

| 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)

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

$$= \frac{(19.9924\ \text{u} \times 90.92) + (20.9940\ \text{u} \times 0.26) + (21.9914\ \text{u} \times 8.82)}{100} = \mathbf{20.1707\ u}$$

**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.**

***Solution:***

Average atomic mass of Boron (B)

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

$$= \frac{(10.13\ \text{u} \times 19.60) + (11.009\ \text{u} \times 80.40)}{100} = \mathbf{10.84\ u}$$

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

**Q.76. Calculate the average atomic mass of argon from the following data:**

| Isotope | Isotopic mass (g mol$^{-1}$) | 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)

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

$$= \frac{(35.96755\ \text{u} \times 0.337) + (37.96272\ \text{u} \times 0.063) + (39.9624\ \text{u} \times 99.60)}{100} = \mathbf{39.947\ g\ mol^{-1}}$$

**Ans:** Average atomic mass of argon **= 39.947 g mol$^{-1}$**

**Q.77. Calculate the molecular mass of the following in u:**

**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:**

**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.**

***Solution:***

Molecular mass of $O_2$ = 2 × 16 u

∴ Mass of 1 molecule = **32 u**

∴ Mass of 1 molecule of $O_2$ = $32 \times 1.66056 \times 10^{-24}$ g = $\mathbf{53.1379 \times 10^{-24}}$ **g**

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

**+Q.80. Find the formula mass of**

**i.** **NaCl** **ii.** **$Cu(NO_3)_2$**

***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**

**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.**

***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.**

***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**

## 1.7 Mole concept and molar mass

**Q.83. Can you recall?** *(Textbook page no. 8)*

**i. One dozen means how many items?** **ii. One gross means how many items?**

**Ans:**

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

***Q.84. Explain: Mole concept**

**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 $6.0221367 \times 10^{23}$ particles/entities.

**Reading between the lines**

*Mass of one carbon-12 atom as determined by mass spectrometer is $1.992648 \times 10^{-23}$ g.*
*Mass of one mole of carbon atoms is 12 g.*
*Hence, Number of atoms in 12 g of carbon-12*

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

$= 6.02213 \times 10^{23}$ *atom/mol*

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

**Ans:** The number of particles in one mole is $6.0221367 \times 10^{23}$.

**Enrich Your Knowledge**

i. The name of the unit is **mole** and the symbol for the unit is **mol**.
ii. The number $6.0221367 \times 10^{23}$ is known as **'Avogadro's Constant'** 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 $6.0221367 \times 10^{23}$, which is called **Avogadro Number**.
iv. The number $6.0221367 \times 10^{23}$ is often rounded to three decimal point as $6.022 \times 10^{23}$ 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**

**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.**

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

## Solved Examples

***Q.88. What is the ratio of molecules in 1 mole of $NH_3$ and 1 mole of $HNO_3$?**

***Solution:***

One mole of any substance contains particles equal to $6.022 \times 10^{23}$.
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 $HNO_3$

$\therefore$ Ratio = $\frac{6.022 \times 10^{23}}{6.022 \times 10^{23}}$ = **1:1**

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

***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**

***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
$= 2 \times 1$
**= 2 moles of oxygen atoms**

iv. Number of molecules of acetaldehyde = Moles of acetaldehyde × Avogadro number
$= 2 \text{ mol} \times 6.022 \times 10^{23}$ molecules/mol
**= $12.044 \times 10^{23}$ molecules of acetaldehyde**

***Q.90. Calculate the number of moles of magnesium oxide, MgO in**
**i. 80 g and** **ii. 10 g of the compound.**
**(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

*Formulae:* 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)
$= (1 \times 24 \text{ u}) + (1 \times 16 \text{ u})$
$= 40 \text{ u}$

∴ Molar mass of MgO = $40 \text{ g mol}^{-1}$

Mass of MgO = 80 g

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

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

= **2 mol**

ii. Mass of MgO = 10 g, Molar mass of MgO = $40 \text{ g mol}^{-1}$

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

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

= **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**

**+Q.91. Calculate the number of moles and molecules of urea present in 5.6 g of urea.**

***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)
$= (2 \times 14 \text{ u}) + (4 \times 1 \text{ u}) + (1 \times 12 \text{ u}) + (1 \times 16 \text{ u}) = 60 \text{ u}$

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

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**

Now,

Number of molecules of urea = Number of moles × Avogadro's constant

$= 0.0933 \text{ mol} \times 6.022 \times 10^{23}$ molecules/mol

$= 0.5619 \times 10^{23}$ molecules

$= \mathbf{5.619 \times 10^{22}}$ **molecules**

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

Number of molecules of urea = $\mathbf{5.619 \times 10^{22}}$ **molecules**

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

***Solution:***

*Given:* Mass of acetic acid = 22 g

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

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

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

*Calculation:* Mass of 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

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

$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{22 \text{ g}}{60 \text{ g mol}^{-1}} = \mathbf{0.367 \text{ mol}}$

Now,

Number of molecules of acetic acid = Number of moles × Avogadro's constant

$= 0.367 \text{ mol} \times 6.022 \times 10^{23}$ molecules/mol

$= \mathbf{2.210 \times 10^{23}}$ **molecules**

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

Number of molecules of acetic acid = $\mathbf{2.210 \times 10^{23}}$ **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)**

***Solution:***

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

Atomic mass of iodine (I) = 127 u

∴ Mass of one iodine atom = 127 u

∴ $x = \frac{254 \text{ u}}{127 \text{ u}} = \mathbf{2}$ **atoms**

ii. 254 g of iodine (I)

Atomic mass of iodine = 127 u

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

Now,

$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{254 \text{ g}}{127 \text{ g mol}^{-1}} = \mathbf{2 \text{ mol}}$

Now,

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

$= 2 \text{ mol} \times 6.022 \times 10^{23}$ atoms/mol

$= 12.044 \times 10^{23}$ atoms

$= \mathbf{1.2044 \times 10^{24}}$ **atoms**

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

ii. Number of iodine atoms in 254 g= $\mathbf{1.2044 \times 10^{24}}$ **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)**

***Solution:***

i. 64 u of oxygen (O) = $x$ atoms

Atomic mass of oxygen (O) = 16 u

∴ Mass of one oxygen atom = 16 u

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

ii. 42 g of nitrogen (N)

Atomic mass of nitrogen = 14 u

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

Now,

$$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{42 \text{ g}}{14 \text{ g mol}^{-1}} = \mathbf{3\ mol}$$

Now,

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

$= 3 \text{ mol} \times 6.022 \times 10^{23}$ atoms/mol

$= 18.07 \times 10^{23}$ atoms

$= \mathbf{1.807 \times 10^{24}}$ **atoms**

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

ii. Number of nitrogen atoms in 42 g = $\mathbf{1.807 \times 10^{24}}$ **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)**

***Solution:***

i. 52 moles of Argon

1 mole Argon atoms = $6.022 \times 10^{23}$ atoms of Ar

∴ Number of atoms = 52 mol × $6.022 \times 10^{23}$ atoms/mol

$= \mathbf{313.144 \times 10^{23}}$ **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 = $x$

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

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

$= 13 \text{ mol} \times 6.022 \times 10^{23}$ atoms/mol

$= \mathbf{78.286 \times 10^{23}}$ **atoms of He**

**Ans:** i. Number of argon atoms in 52 moles = $\mathbf{313.144 \times 10^{23}}$ **atoms of Argon**

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

iii. Number of helium atoms in 52 g = $\mathbf{78.286 \times 10^{23}}$ **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**

***Solution:***

i. 0.4 mole of nitrogen (N)

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

$= 0.4 \text{ mol} \times 6.022 \times 10^{23}$ atoms/mol

$= \mathbf{2.4088 \times 10^{23}}$ **atoms of N**

ii. 1.6 g of Sulphur (S)

Molar mass of sulphur = 32 g $mol^{-1}$

Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{1.6\ g}{32\ g\ mol^{-1}}$ = 0.05 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.**

***Solution:***

i. 5 mg carbon = 5 × $10^{-3}$ g carbon

Atomic mass of carbon = 12 u

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

Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{5 \times 10^{-3}\ g}{12\ g\ mol^{-1}}$ = **4.167 × $10^{-4}$ mol**

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

Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{12 \times 10^{-3}\ g}{12\ g\ mol^{-1}}$ = 1 × $10^{-3}$ mol

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

Number of atoms of carbon = 1 × $10^{-3}$ 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.**

***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: $CO(NH_2)_2$

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

Number of moles = $\frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{5.6\ g}{60\ g\ mol^{-1}}$ = 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 × 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 × 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 × 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 × 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.**

**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 $mol^{-1}$).**

***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 $mol^{-1}$

$$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{72.5\text{ g}}{60\text{ g mol}^{-1}} = 1.208\text{ mol}$$

∴ 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 × 1.208) mol × 6.022 × $10^{23}$ atoms/mol
= **2.182 × $10^{24}$ atoms of carbon**

∴ 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**

∴ 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.**

***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

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

$$\text{Number of moles} = \frac{\text{Mass of a substance}}{\text{Molar mass of a substance}} = \frac{250\text{ g}}{180\text{ g mol}^{-1}} = 1.389\text{ mol}$$

Now,

1.389 mol of glucose cost = Rs 40

1 mol glucose cost = $x$

∴ $x = \frac{\text{Rs } 40}{1.389\text{ mol}}$ = **Rs 28.8/mol of glucose**

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

## 1.8 Moles and gases

***Q.101. Explain: Molar volume of gas**

**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.

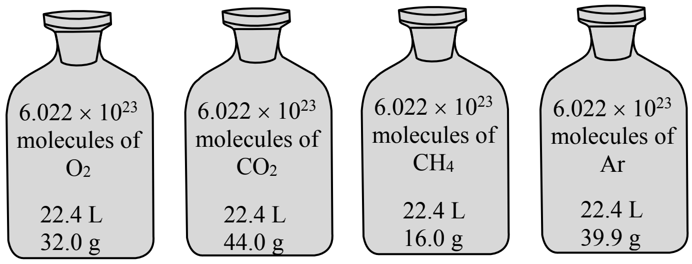
iv. The relationship between number of moles and molar volume can be expressed as follows:

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

***[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 $mol^{-1}$**]***

**Enrich Your Knowledge**

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

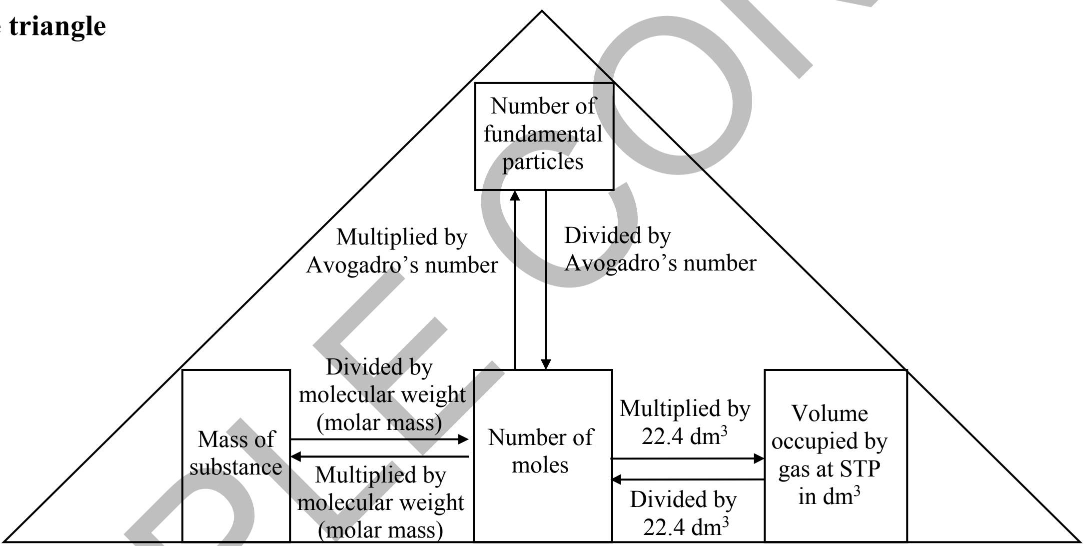



***Q.102. What is meant by molar volume of a gas?**

**Ans:** The volume occupied by one mole of a gas at **s**tandard 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 $dm^3$.

**Enrich 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)<br>(French chemist and physicist) | i.<br>ii.<br><br>iii. | Formulated the gas law.<br>Collected samples of air at different heights and recorded temperatures and moisture contents.<br>Discovered that the composition of atmosphere does not change with increasing altitude. |
| **Amedeo Avogadro** (1776 – 1856)<br>(Italian scholar) | i.<br><br><br>ii. | Published article in French journal on determining the relative masses of elementary particles of bodies and proportions by which they enter combinations.<br>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.**]***

## Solved Examples

***Q.104. Calculate number of moles of hydrogen in 0.448 litre of hydrogen gas at STP.**

***Solution:***

*Given:* Volume of hydrogen at STP = 0.448 L

*To find:* Number of moles of hydrogen

*Formula:* $\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$

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

$$\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

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

$$= \mathbf{0.02\ mol}$$

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

**+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.**

***Solution:***

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

*To find:* Number of moles and molecules of ammonia

*Formulae:* i. $\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$

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

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

$$\text{Number of moles (n)} = \frac{\text{Volume of the gas at STP}}{\text{Molar volume of gas}}$$

$$\text{Number of moles of } NH_3 = \frac{67.2\ dm^3}{22.4\ dm^3\ mol^{-1}} = \mathbf{3.0\ mol}$$

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

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

= $\mathbf{18.066 \times 10^{23}}$ **molecules**

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

Number of molecules of ammonia = $\mathbf{18.066 \times 10^{23}}$ **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.**

***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:* $\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$

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

$$\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

∴ 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 $dm^3$**

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

**Calculate the volume in $dm^3$ occupied by 60.0 g of ethane at STP.**

***Solution:***

*Given:* Mass of ethane at STP = 60.0 g

*To find:* Volume of ethane

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

ii. $\text{Number of moles} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$

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

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

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

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

$$\text{Number of moles of a gas (n)} = \frac{\text{Volume of the gas at STP}}{\text{Molar volume of a gas}}$$

$\therefore$ 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.**

***Solution:***

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 $dm^3$

$\therefore$ $x = \frac{22.4 \times 3.40}{4.48}$ **= 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 $mol^{-1}$.**

***Solution:***

The molecular formula of potassium chlorate is $KClO_3$.

Required chemical equation:

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

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.

$\therefore$ $x = \frac{245 \times 6.72}{67.2}$ = **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.*

## Apply Your Knowledge

**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

∴ 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
  - **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 × $6.022 \times 10^{23}$
6. Number of moles (n) = $\frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$
   $= \frac{\text{Volume of a gas at STP}}{22.4 \text{ dm}^3 \text{ mol}^{-1}}$

## Exercise

### 1.2 Nature of chemistry

1. What are pure substances? Give two examples.
**Ans:** *Refer Q.6. (i)*

2. What are metalloids?
**Ans:** *Refer Q.6. (i-a-3)*

3. What is a homogeneous mixture?
**Ans:** *Refer Q.6. (ii-a)*

4. Give one example of each:
i. Heterogeneous mixture
ii. Compound
iii. Element
iv. Homogeneous mixture
**Ans:** *Refer Q.6.*

5. Explain classification of matter.
**Ans:** *Refer Q.6.*

### 1.3 Properties of matter and their measurement

6. Give SI unit of:
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.
**Ans:** *Refer Q.17. (ii)*

*8. Give two examples to explain Gay-Lussac's law of gaseous volume.
**Ans:** *Refer Q.44. (ii)*

9. i. State the law of conservation of mass.
ii. Explain the law of multiple proportions with reference to carbon monoxide and carbon dioxide.
iii. Validate Gay Lussac's law of combining volume of gases using an example.
**Ans:**
i. *Refer Q.37.* ii. *Refer Q.42.*
iii. *Refer Q.46.*

*10. State Avogadro's law.
**Ans:** *Refer Q.47. (ii)*

### 1.5 Dalton's atomic theory

11. What were the basic assumptions of Dalton's theory?
**Ans:** *Refer Q.54.*

12. What happens during a chemical reaction according to Dalton's atomic theory?
**Ans:** *Refer Q.54. (iv)*

### 1.6 Atomic and molecular masses

13. Why is it impossible to measure the mass of a single atom?
**Ans:** *Refer Q.60. (i)*

14. Calculate the atomic mass (average) of chlorine using the following data:

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

**Ans:** 35.4528 g $mol^{-1}$

15. Calculate the molecular mass of the following in u:
i. $CH_3OH$ ii. $NO_2$
iii. $HNO_3$
**Ans:**
i. 32 u ii. 46 u
iii. 63 u

16. Find the formula mass of $Na_2SO_4$.
(Atomic mass of Na = 23 u, S = 32 u, O = 16 u)
**Ans.** 142 u.

### 1.7 Mole concept and molar mass

17. Define one mole.
**Ans:** *Refer Q.84. (ii)*

18. How many atoms of sulphur are present in 0.1 mole of $S_8$ molecules?
**Ans:** $4.82 \times 10^{23}$ atoms

19. Calculate the mass of the following:
i. 0.25 mole of iron
ii. 2.5 moles of ammonia
iii. 250 molecules of sodium chloride
**Ans:** Iron: $1.4 \times 10^{-2}$ kg; ammonia: $4.25 \times 10^{-2}$ kg and sodium chloride: $2.429 \times 10^{-23}$ kg

20. Calculate the number of molecules in 28 g of nitrogen, 64 g of oxygen and 72 g of water.

**Ans:** Nitrogen - $6.022 \times 10^{23}$ molecules
Oxygen - $1.2044 \times 10^{24}$ molecules
Water - $2.4088 \times 10^{24}$ molecules

21. 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)

**Ans:**
i. 1.5 mol  ii. 0.5 mol

22. Calculate the number of moles and molecules of urea present in 30 g of urea.

**Ans:**
i. 0.5 mol
ii. $3.011 \times 10^{23}$ molecules

## 1.8 Moles and gases

23. Calculate the volume in litres of the following gases at STP:
i. 1.6 g of oxygen
ii. $3.5 \times 10^{-3}$ kg of nitrogen
iii. $85 \times 10^{-3}$ kg of hydrogen sulphide

**Ans:**
i. 1.12 L  ii. 2.8 L
iii. 56 L

## Multiple Choice Questions

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) 88.8 (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) $6.022 \times 10^{23}$ 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. Which of the following has the largest number of atoms?
(A) 1g $Au_{(s)}$ (B) 1g $Na_{(s)}$
(C) 1g $Li_{(s)}$ (D) 1g $Cl_{2(g)}$

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

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

*18. 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 $NO_2$

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

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

21. 11.2 $cm^3$ of hydrogen gas at STP, contains_______ moles
(A) 0.0005 (B) 0.01
(C) 0.029 (D) 0.5

22. 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. | (C) | 16. | (D) |
| 17. | (D) | 18. | (B) | 19. | (B) | 20. | (B) |
| 21. | (D) | 22. | (A) | | | | |

## Competitive Corner

1. The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is _______. **[NEET (UG) 2019]**
(A) 20 **(B) 30** (C) 40 (D) 10

**Hint:** $N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$
3 mol $H_2$ = 2 mol $NH_3$
$\therefore$ 30 mol $H_2$ = 20 mol $NH_3$

2. The combining ratios of hydrogen and oxygen in water and hydrogen peroxide are 1:8 and 1:16. Which law is illustrated in this example? **[MHT CET 2019]**
(A) Law of definite proportions (B) Gay Lussac's law of combining volumes of gases
(C) Law of conservation of mass **(D) Law of multiple proportions**

3. What is the SI unit of density? **[MHT CET 2018]**
(A) $g\ cm^{-3}$ (B) $g\ m^{-3}$ **(C) $kg\ m^{-3}$** (D) $kg\ cm^{-3}$

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

4. Which symbol replaces the unit of atomic mass, amu? **[MHT CET 2018]**
**(A) u** (B) A (C) M (D) n

5. The most abundant elements by mass in the body of a healthy human adult are:
Oxygen (61.4 %), Carbon (22.9 %), Hydrogen (10.0 %) and Nitrogen (2.6 %). The weight which a 75 kg person would gain if all $^1H$ atoms are replaced by $^2H$ atoms is _______. **[JEE (Main) 2017]**
**(A) 7.5 kg** (B) 10 kg (C) 15 kg (D) 37.5 kg

**Hint:** Total weight of person = 75 kg
Mass due to $^1H$ atoms = $75 \times \frac{10}{100}$ = 7.5 kg
Mass of $^2H$ atom is twice that of $^1H$ atom.
All $^1H$ atoms are replaced by $^2H$ atoms.
Hence, mass increase is twice i.e., by 7.5 kg.