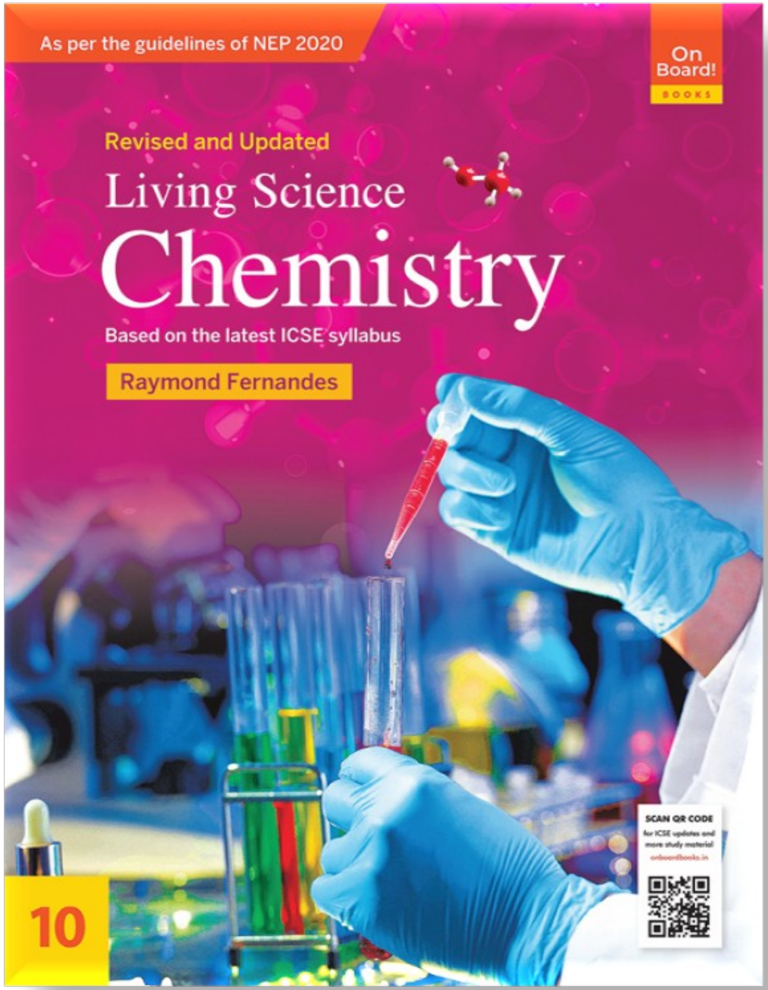
On Board!

ICSE
Living Science
Chemistry

Class 10

Chapter-5 Mole Concept and Stoichiometry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Revised and Updated
Living Science
Chemistry
Based on the latest ICSE syllabus
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LEARNING OBJECTIVES

Gas Laws

- ❖Standard temperature and pressure

Gay-Lussac's Law of Combining Volumes

Avogadro's Law

Relative Atomic Mass and Relative Molecular Mass

- ❖Relative atomic mass and atomic mass
- ❖Relative molecular mass and molecular mass
- ❖Formula unit mass
- ❖Gram atomic mass
- ❖Gram molecular mass

Mole Concept

- ❖Molar volume

Applications of Avogadro's Law

Percentage Composition

Empirical and Molecular Formula

- ❖Calculations based on chemical equations

What is Stoichiometry?

The term is generally used for referring to the calculation of quantities of substances involved in chemical reactions. These calculations are used during chemical analysis and during production of most commonly used chemicals.

Gas Laws

The gaseous state of a substance is defined by three variables, namely, pressure (*P*), volume (*V*), and temperature (*T*). The specific relations between the three variables are called gas laws. The gas laws studied earlier have been summarised in the following table.

Gas law	Variables	Relationship	Equation
Boyle's law	Pressure, Volume	Volume $\propto \frac{1}{\text{Pressure}}$ (at constant temperature)	$P_1V_1 = P_2V_2$
Charle's law	Volume, Temperature	Volume $\propto$ Temperature (at constant pressure)	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gas equation (obtained on applying Boyle's and Charle's law)	Pressure, Volume, Temperature	When above two laws are combined together Pressure $\propto \frac{1}{\text{Volume}}$ $\propto$ Temperature	$\frac{PV}{T} = \text{K (constant)}$ $\Rightarrow \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Standard Temperature and Pressure (S.T.P.)

The volume of a gas changes with change in temperature and pressure. In order to compare the volume of gases, standard conditions of temperature and pressure are employed.

Standard temperature = 0 °C or 273 K

Standard pressure = 10^5 Pa

Gay-Lussac's Law of Combining Volumes

When gases react, they do so in volumes which bear a simple whole number ratio to one another and to the volume of the product if gaseous, provided all the volumes are measured at the same temperature and pressure.

This law is applicable only to gases, and therefore, solid and liquid reactants and products are not considered for chemical calculations even if they are a part of the reaction. Water, however, poses a slightly tricky proposition as it exists as vapour at 100 °C and as a liquid below 100 °C. It is, thus, imperative that the temperature of the reaction is mentioned, especially if water is a part of the reaction.

For example, for the reaction,

$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

measurements are done at 100 °C.

Avogadro's Law

Equal volumes of all gases at the same temperature and pressure contain the same number of molecules. Thus, if the temperature, pressure and volumes of two gases are the same, then, the two gases contain the same number of molecules regardless of their identity. Avogadro's law is valid for all gases, no matter what they are.

This suggests that one litre of oxygen contains the same number of molecules as are present in one litre of nitrogen, provided that they have the same temperature and pressure.

The usefulness of Avogadro's law lies in the fact that it differentiates between an atom and a molecule.

An atom is the smallest particle of an element that participates in a chemical reaction and cannot have an independent existence.

A molecule is the smallest particle of an element or a compound that can have an independent existence.

Relative Atomic Mass and Relative Molecular Mass

Relative atomic mass and atomic mass

The mass of an atom is actually very small because atoms are extremely small. For example, the mass of one atom of hydrogen is 1.66×10^{-24} g, while that of one atom of carbon is 1.99×10^{-23} g. Numbers are so small that it is difficult to use. Therefore, instead of using actual masses, the relative atomic masses are used. It could be done by choosing the mass of the lightest atom as a standard mass and then comparing the mass of atoms of any element relative to it. It is this relative mass which is called the atomic mass or relative atomic mass.

As hydrogen atoms are the lightest of all the atoms of elements, one atom of hydrogen was originally taken as the mass with which all other atoms would be compared and hence, the mass of an atom of hydrogen could be taken exactly as one unit.

According to hydrogen standard, original relative

$$\text{atomic mass} = \frac{\text{Mass of one atom of the element}}{\text{Mass of one atom of hydrogen}}$$

But since many elements do not react with hydrogen, the hydrogen standard was replaced by oxygen standard, as oxygen is more reactive than hydrogen. As an oxygen atom is 16 times heavier than hydrogen atom, the relative atomic mass of an element was expressed as:

$$\text{Relative atomic mass of an element} = \frac{\text{Mass of 1 atom of an element}}{1/16 \times \text{mass of one atom of oxygen}}$$

However, due to the discovery of isotopes of oxygen like O-17 and O-18, the relative atomic mass of an element in terms of oxygen standard cannot be used.

In 1961, C-12 was selected as the most appropriate and stable atom for standard reference. Although C-13 and C-14 were also found as the isotopes of carbon, but their relative abundance is very small so that it can be taken as negligible. Thus, the relative atomic mass of an element can be defined as ‘the number of times an atom is heavier than one-twelfth the mass of a carbon-12 atom’

$$\text{Relative atomic mass of an element} = \frac{\text{Mass of one atom of the element}}{\text{Mass of } \frac{1}{12}\text{th of an atom of carbon-12}}$$

The relative atomic mass is a ratio, and therefore, it has no unit.

The atomic mass is expressed in atomic mass unit. One atomic mass unit (u) is equal to 1/12 th of the mass of one atom of carbon-12.

$$1\ \text{u} = \frac{\text{Mass of one atom of carbon-12}}{12}$$

$$= \frac{1.9924\times10^{-23}}{12}\ \text{g}$$

$$1\ \text{u} = 1.66\times10^{-24}\ \text{g or } 1.66\times10^{-27}\ \text{kg}$$

Note: Atomic mass of an element is not always a whole number. It may be fractional. This is due to the presence of isotopes of an element. The isotopes of a particular element differ in their atomic masses as well as the relative proportions. Therefore, for the elements existing as isotopes, the atomic mass is considered as average atomic mass, which is equal to the average of atomic mass of its natural isotopes. For example, chlorine exists as two isotopes, $_{17}^{35}Cl$ and $_{17}^{37}Cl$ in the ratio of 3 : 1 and its average atomic mass is 35.43 u.

Relative molecular mass and molecular mass

The relative molecular mass of an element or a compound is defined as the number of times one molecule of an element or a compound is heavier than 1/12 th of the mass of an atom of carbon-12. For example, a molecule of oxygen (O2) is 32 times heavier than 1/12th of the mass of an atom of carbon. Therefore, the relative molecular mass of oxygen is 32 u.

The molecular mass is the average mass of one molecule of a compound. The molecular mass of a substance is obtained as the sum of the atomic masses of all the atoms in a molecule of the compound. It is, therefore, the molecular mass, i.e., mass of a molecule is expressed in atomic mass unit.

For example, the molecular mass of glucose, i.e., $C_6H_{12}O_6$ can be calculated as follows:

$$\left[\begin{pmatrix}\text{Atomic mass}\\ \text{of carbon}\end{pmatrix}\times 6\right] + \left[\begin{pmatrix}\text{Atomic mass}\\ \text{of hydrogen}\end{pmatrix}\times 12\right] + \left[\begin{pmatrix}\text{Atomic mass}\\ \text{of oxygen}\end{pmatrix}\times 6\right]$$

$$= (12 \times 6) + (1 \times 12) + (16 \times 6)$$

$$= 72 + 12 + 96 = 180 \text{ u}$$

Formula unit mass

For the substances containing ions, i.e. for ionic compounds, the term formula unit mass is used instead of molecular mass. This is because ionic compounds exist as a cluster of ions and not as molecules. The formula unit mass is the sum of the atomic masses of all the atoms present in a formula unit of compound.

Gram atomic mass

The amount of an element whose mass in grams is numerically equal to its atomic mass is called gram atomic mass of that element. In other words, the atomic mass of an element expressed in grams is called its gram atomic mass. For example,

Atomic mass of oxygen = 16 u

So, Gram atomic mass of oxygen = 16 g

Gram molecular mass

The molecular mass of a substance expressed in grams is called its gram molecular mass. Molecular mass of oxygen, O_2 = 32 u

So, gram molecular mass of oxygen, O_2 = 32 g

Mole Concept

Atoms and molecules are expressed in terms of mole. The mole is the standard chemical unit used to measure the quantity of a substance in terms of number and mass.

One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g of the C-12 isotope. The mole is equal to 6.022×10^{23} particles (atoms, molecules or ions) of a substance.

1 mole of atoms = 6.022×10^{23} atoms

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

1 mole of oxygen atoms (O)

= 6.022×10^{23} oxygen atoms

1 mole of oxygen molecules (O_2)

= 6.022×10^{23} oxygen molecules

This number of 6.022×10^{23}, which represents a mole, is known as Avogadro number (NA or L).

Mole of atoms: One mole of atoms of an element has a mass equal to the gram atomic mass of the element.

1 mole of oxygen atoms

= Gram atomic mass of oxygen

= 16 g

Mole of molecules: One mole of molecules of a substance has a mass equal to the gram molecular mass of the substance.

One mole of oxygen molecules contains 6.022×10^{23} molecules and weighs 32 g.

For ionic compounds: For ionic compounds, the term 'formula unit mass' is used in place of molecular mass and the term 'formula unit' is used in place of 'molecule'. 1 mole of ionic compound contains 6.022×10^{23} formula units of the compound and has a mass equal to their formula unit mass.

Molar Volume

The volume occupied by 1 mole of any gas at S.T.P. (i.e. at 273 K and 10^5 Pa) is called molar volume. This is equal to 22.4 L.

Note: Refer to **Table 5.2 for** Mole relationships for some gases.

Applications of Avogadro's Law

1. The law helped in initiating the crossover of statements involving **volumes of gases** to **molecules of gases**. "1 **volume** of hydrogen combines with 1 **volume** of chlorine to form 2 **volumes** of hydrogen chloride." This could now be referred to as 1 **molecule** of hydrogen combines with 1 **molecule** of chlorine to form 2 **molecules** of hydrogen chloride.

2. Modification of the Atomic Theory: The law modified the Atomic Theory by making a distinction between atoms and molecules.

3. Deduction of Gay-Lussac's law

4. Determining the atomicity of elementary gases:

- The molecules of noble gases contain only one atom. So, noble gases are termed monoatomic.
- The molecules of gases like oxygen and nitrogen contain two atoms. Therefore, oxygen and nitrogen molecules are termed diatomic.
- The molecules of ozone contain three atoms. Therefore, ozone is termed triatomic.

5. Deriving the relationship between atomic weight and vapour density: The vapour density (V.D.) of a gas is defined as the ratio between the mass of a certain volume of the gas to the mass of the same volume of hydrogen gas under similar conditions of temperature and pressure. Thus,

$$\text{V.D.} = \frac{\text{Mass of } V \text{ mL of a substance in the gaseous state}}{\text{Mass of } V \text{ mL of hydrogen under similar conditions}}$$

According to Avogadro's hypothesis, equal volumes of gases under similar condition should contain equal number of molecules. Therefore, if V mL of any gas contains n molecules, then,

$$\text{V.D.} = \frac{\text{Mass of } n \text{ molecules of the gaseous substance}}{\text{Mass of } n \text{ molecules of hydrogen}}$$

$$\text{or} \quad \text{V.D.} = \frac{\text{Mass of 1 molecule of the gaseous substance}}{\text{Mass of 1 molecule of hydrogen}}$$

6. Determining atomic weights

In 1858, Cannizzaro used Avogadro's law to calculate atomic and molecular weights. At S.T.P., 1 litre of H_2 weighs 0.0899 g and 1 litre of O_2 weighs 1.4290 g. By Avogadro's law, since their volumes are equal their number of molecules should also be equal and therefore 1 molecule of hydrogen will be 0.0899/1. 4290 times heavier than a molecule of oxygen.

7. Determining molecular formulae

To determine the molecular formula of ammonia, it was decomposed and the volumes of the gaseous products were determined.

8. Derivation of the mole concept

In general, a mole is that amount of a substance which contains as many entities (atoms, molecules, or ions) as there are atoms in 0.012 kg or 12 g of Carbon-12.

1 mole of a gas occupies 22.4 litres at S.T.P. This is also called molar volume.

1 mole of a gas also contains 6.022×10^{23} molecules (or atoms in case of monoatomic gases). This number is known as Avogadro's number.

Percentage Composition

The mass percentage of each constituent element present in any compound is called its percentage composition.

Percentage of an element in a compound

$$= \frac{\text{Total mass of the element in one molecule of a compound}}{\text{Gram molecular mass of the compound}} \times 100$$

Empirical and Molecular Formulae

Empirical Formula

The simplest formula of a substance which gives the relative number of atoms of each element present in a molecule of the substance is called **empirical formula**. The empirical formula of a compound gives the simplest whole number ratio between the number of atoms of all the elements present in the compound. For example, a compound contains carbon, hydrogen and oxygen in the ratio 1 : 2 : 1. Its empirical formula will be CH_2O. It indicates the simplest ratio between the carbon, hydrogen and oxygen atoms in its molecule. The actual formula of this compound is $C_6H_{12}O_6$.

Empirical formula mass: The sum of the atomic masses of all the atoms present in the empirical formula is called the **empirical formula mass**.

Molecular formula

A symbolic representation of a molecule of a compound in terms of symbol and actual number of atoms of each element present in one molecule of the compound is called its **molecular formula**.

Molecular formula = n × Empirical formula

where, n = Molecular mass / Empirical formula mass

Determination of molecular formula

To calculate the molecular formula, the relative molecular mass of the compound is required. This will be provided in the sum directly or in the form of the vapour density of the compound or in any other form that will lead to the calculation of the molecular mass.

Step 1: Calculate the empirical formula mass

Step 2: Calculate the molecular mass Molecular mass = 2 × Vapour density

Step 3: Multiplication factor n = Molecular mass / Empirical formula mass

Step 4: Molecular formula = (Empirical formula)n

Note: The value of n must be rounded off to nearest whole number.

Calculations Based on Chemical Equations

A chemical equation should represent a true chemical reaction and it should be a balanced equation, i.e. the number of atoms of each element on both the sides should be equal.

Information conveyed by a chemical equation

A chemical equation gives the following types of information:

1. **Qualitative information:** This gives the information regarding the

a. names of the reactants that take part in the reaction.

b. names of the products formed in the reaction.

2. **Quantitative information:** This gives the information regarding the

a. number of molecules or atoms of reactants and products.

b. number of moles of reactants and products.

c. volumes of gaseous reactants and products measured at S.T.P.

Steps for solving the problems based on chemical equations

To solve the problems based on chemical equations, follow the given steps.

Step 1: Write down the balanced chemical equation.

Step 2: Calculate the gram atomic mass or gram molecular mass or formula unit mass of each reactant and product below their formulae.

Step 3: Identify the substances whose quantities are given or are to be found out. Write down the actual quantities of the substances given.

SUMMARY

1. **Gay-Lussac's law of combining volumes:** When gases react, they do so in volumes which bear a simple whole number ratio to one another and to the volume of the product if gaseous, provided all the volumes are measured at the same temperature and pressure.

2. **Avogadro's law:** Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

3. **Atomicity:** Atomicity is defined as the number of atoms in one molecule of an element.

4. **Vapour density:** It is the ratio of the mass of a certain volume of a gas or vapour to the mass of an equal volume of hydrogen; all measurements made under the same conditions of temperature and pressure.

5. **Relative atomic mass:** It is the ratio of the mass of an atom of an element to 1/12 th the mass of an atom of carbon-12.

6. Relative molecular mass: It is the ratio of the mass of one molecule of a substance to 1/12 th the mass of an atom of carbon-12.

7. Gram atomic mass: It is the relative atomic mass expressed in grams.

8. Gram molecular mass: It is the relative molecular mass expressed in grams.

9. Molar volume: It is the volume occupied by one mole of a gas at S.T.P.

10. Mole: One mole of an atom is the relative atomic mass of that atom expressed in grams. One mole of a molecule is the relative molecular weight of that element expressed in grams. In general, one mole is amount of a substance which contains as many entities (atoms, ions or molecules) as there atoms in 0.012 kg or 12 g of carbon-12.

11. **Molecular formula** = (Empirical formula)$_n$

12. $\text{V.D.} = \dfrac{\text{Mass of a certain volume of gas}}{\text{Mass of the same volume of hydrogen}}$

THANK YOU