Class: 9th

Subject: Chemistry

Chapter 4: Stoichiometry

Exercise MCQs:

(i) How many atoms are present in one gram of H_2O?

(a) 1002×10^{23} atoms ✓

(b) 6.022×10^{23} atoms

(c) 0.334×10^{23} atoms

(d) 2.004×10^{23} atoms

(ii) Which is the correct formula of calcium phosphide?

(a) CaP

(b) CaP_2

(c) Ca_2P_3

(d) Ca_3P_2 ✅

(iii) How many atomic mass units (amu) are there in one gram?

(a) 1 amu

(b) 10^{23} amu

(c) 6.022×10^{23} amu ✅

(d) 6.022×10^{22} amu

(iv) Structural formula of 2-hexene is $CH_3–CH=CH–(CH_2)_2–CH_3$. What will be its empirical formula?

(a) C_2H_2

(b) CH

(c) C_6H_{12}

(d) CH_2 ✅

(v) How many moles are there in 25 g of H_2SO_4?

(a) 0.765 moles ✅

(b) 0.51 moles

(c) 0.255 moles (Correct Answer)

(d) 0.4 moles

(vi) A necklace has 6 g of diamonds in it. What are the number of carbon atoms in it?

(a) 6.02×10^{23} ✅

(b) 12.04×10^{2}

(c) 1.003×10^{2}

(d) 3.01×10^{23}

(vii) What is the mass of Al in 204 g of aluminium oxide, Al_2O_3?

(a) 26 g

(b) 27 g

(c) 54 g

(d) 108 g ✅

(viii) Which one of the following compounds will have the highest percentage of the mass of nitrogen?

(a) $CO(NH_2)_2$ ✅

(b) N_2H

(c) NH

(ix) When one mole of each of the following compounds is reacted with oxygen, which will produce the maximum amount of CO_2?

(a) Carbon ✅

(b) Diamond

(c) Ethane (C_2H_2)

(d) Methane (CH_4)

(x) What mass of 95% $CaCO_3$ will be required to neutralize 50 cm^3 of 0.5 M HCl solution?

(a) 9.5 g

(b) 1.25 g

(c) 1.32 g ✅

(d) 1.45 g

Important MCQs:

1. Which of the following elements exists as a molecule in nature?

(a) Na

(b) Fe

(c) O_2 ✅

(d) Ca

2. What is the correct chemical formula of ozone?

(a) O

(b) O_2

(c) O_3 ✅

(d) O_4

3. Which of the following is represented by a formula unit instead of a molecular formula?

(a) H_2O

(b) NH_3

(c) NaCl ✅

(d) CH_4

4. The chemical formula of calcium chloride is:

(a) CaCl

(b) $CaCl_2$ ✅

(c) Ca_2Cl

(d) Ca_2Cl_2

5. Which of the following is a covalent compound?

(a) NaCl

(b) KBr

(c) HCl ✅

(d) $CaCl_2$

6. What does a chemical formula represent in covalent compounds?

(a) Ratio of ions

(b) Atomic number

(c) Molecular mass

(d) Actual number of atoms in a molecule ✅

7. What is the molecular formula of hydrogen peroxide?

(a) H_2O

(b) HO

(c) H_2O_2 ✅

(d) H_2O_4

8. Which of the following compounds has same empirical and molecular formula?

(a) H_2O ✅

(b) H_2O_2

(c) C_6H_6

(d) C_2H_2

9. Which pair of compounds has the same empirical formula?

(a) CH_4 and C_2H_6

(b) H_2O and H_2O_2

(c) C_6H_6 and C_2H_2 ✅

(d) CO_2 and CO

10. Which of the following compounds is composed of ions?

(a) CO_2

(b) CH_4

(c) KBr ✅

(d) NH_3

11. What is the charge on lithium ion?

(a) +2

(b) -1

(c) +1 ✅

(d) -2

12. What is the correct chemical formula of lithium oxide?

(a) LiO

(b) LiO_2

(c) Li_2O ✅

(d) Li_3O_2

13. Which method is commonly used to write the formula of ionic compounds?

(a) Crisscross method ✅

(b) Balancing method

(c) Exchange method

(d) Neutralization method

14. What does the empirical formula show?

(a) Exact number of atoms in molecule

(b) Simplest whole number ratio of atoms ✅

(c) Mass ratio between elements

(d) Total electrons in an element

15.Which of the following expresses the relationship between molecular and empirical formula?

(a) Molecular formula = Empirical formula × 2

(b) Molecular formula = Empirical formula ÷ n

(c) Molecular formula = n × Empirical formula ✅

(d) Molecular formula = Empirical formula only

16. What is the value of Avogadro's number?

(a) 3.011×10^{23}

(b) 6.022×10^{22}

(c) 6.022×10^{23} ✅

(d) 1.008×10^{23}

17. Empirical formula represents:

(a) Actual number of atoms

(b) Simplest whole number ratio of atoms ✅

(c) Molar mass

(d) Molecular weight

18. The molecular formula of hydrogen peroxide is:

(a) H_2O

(b) H_2O_2 ✅

(c) HO

(d) H_2O_4

19. The molar mass of O_2 is:

(a) 16 g

(b) 18 g

(c) 32 g ✅

(d) 34 g

20. One mole of NaCl contains:

(a) 6.022×10^{23} atoms

(b) 6.022×10^{23} ions

(c) 6.022×10^{23} formula units ✅

(d) 6.022×10^{23} molecules

21. Molar mass of CO_2 is:

(a) 28 g

(b) 32 g

(c) 44 g ✅

(d) 40 g

22. Which law forms the basis of stoichiometry?

(a) Law of constant composition

(b) Law of multiple proportions

(c) Law of conservation of mass ✅

(d) Law of definite proportions

23. The molar mass of hydrogen atoms is:

(a) 1.008 g ✅

(b) 2.016 g

(c) 16.00 g

(d) 1.000 g

24. The number of particles in one mole of any substance is called:

(a) Molar volume

(b) Atomic mass

(c) Avogadro's number ✅

(d) Empirical ratio

25. Structural formula provides:

(a) Only the number of atoms

(b) Shape of molecule

(c) Arrangement of atoms in a molecule ✅

(d) Empirical mass only

26. Which step is required to deduce a molecular formula from a structural formula?

(a) Count valency

(b) Count number of atoms ✅

(c) Write formula units

(d) Write empirical mass

27. The empirical formula of hydrogen peroxide is:

(a) H_2O

(b) HO ✅

(c) H_2O_2

(d) O_2H

28. Which of the following represents the mass of one mole of a substance?

(a) Atomic mass

(b) Empirical mass

(c) Molecular mass

(d) Molar mass ✅

29. 12 g of carbon contains how many atoms?

(a) 1 mole ✅

(b) 0.5 mole

(c) 2 moles

(d) 6.022×10^{12} atoms

30. If molar mass = 60 and empirical formula mass = 30, the value of n is:

(a) 1

(b) 2 ✅

(c) 30

(d) 60

31. In a chemical equation, the substances which react are called:

(a) Products

(b) Reactants ✅

(c) Molecules

(d) Compounds

32. Which of the following must a chemical equation always follow?

(a) Law of inertia

(b) Law of multiple proportions

(c) Law of conservation of mass ✅

(d) Law of gravity

33. What is used to separate reactants from products in a chemical equation?

(a) Plus sign (+)

(b) Equal sign (=)

(c) Arrow (⇒) ✅

(d) Division sign (÷)

34. The physical state of an aqueous solution is represented as:

(a) (s)

(b) (l)

(c) (g)

(d) (aq) ✅

35. Which of the following is an example of a reversible reaction?

(a) $Zn + H_2SO_4 \Rightarrow ZnSO_4 + H_2$

(b) $N_2 + 3H_2 \rightleftharpoons 2NH_3$ ✅

(c) $H_2 + O_2 \Rightarrow H_2O$

(d) $Na + Cl \Rightarrow NaCl$

36. The mole ratio in the reaction: $Zn + H_2SO_4 \Rightarrow ZnSO_4 + H_2$ is:

(a) 2:1:2:1

(b) 1:2:1:2

(c) 1:1:1:1 ✅

(d) 1:1:2:2

37. Which compound forms a precipitate in the reaction:

$AgNO_3 + NaCl \Rightarrow AgCl + NaNO_3$

(a) $NaNO_3$

(b) NaCl

(c) AgCl ✅

(d) $AgNO_3$

38. A balanced chemical equation ensures:

(a) Equal number of molecules on both sides

(b) Equal mass on both sides ✅

(c) Equal volume on both sides

(d) Equal moles of gases only

39. In the reaction:

$CaCO_3 + 2HCl \Rightarrow CaCl_2 + H_2O + CO_2$,

the total mass of the reactants is:

(a) 111 g

(b) 144 g

(c) 173 g ✅

(d) 100 g

40. In the above reaction, how many grams of $CaCl_2$ are produced?

(a) 44 g

(b) 18 g

(c) 111 g ✅

(d) 73 g

2. Exercise Short Questions:

i. Write down the chemical formula of barium nitride.

Answer:

Barium nitride is formed by combining Ba^{2+} and N^{3-} ions. The correct formula is:

Ba_3N_2

ii. Find out the molecular formula of a compound whose empirical formula is CH_2O and its molar

mass is 180.

Step 1: Find molar mass of empirical formula:

C = 12, H_2 = 2, O = 16 ⇒ CH_2O = 30 g/mol

Step 2:

Molar mass of compound ÷ empirical formula mass = 180 ÷ 30 = 6

Answer:

Molecular formula = $(CH_2O)_6$ = $C_6H_{12}O_6$

iii. How many molecules are present in 1.5 g of H_2O?

Step 1: Molar mass of H_2O = 18 g/mol

Step 2: Moles = 1.5 ÷ 18 = 0.0833 mol

Step 3: Molecules = 0.0833 × 6.022 × 10^{23} = 5.02 × 10^{22} molecules

Answer:

Approximately 5.02 × 10^{22} molecules of H_2O

iv. What is the difference between a mole and Avogadro's number?

Answer:

Mole: A unit that represents 6.022×10^{23} particles (atoms, ions, or molecules).

Avogadro's number: The numerical value of particles in one mole, i.e., 6.022×10^{23}.

In short: Mole is a quantity, Avogadro's number is the count of particles in that quantity.

v. Write down the chemical equation of the following reaction:

Reaction: Copper + Sulphuric acid ⇒ Copper sulphate + Sulphur dioxide + Water

Balanced Chemical Equation:

$Cu + 2H_2SO_4 \Rightarrow CuSO_4 + SO_2 + 2H_2O$

important Short Questions:

1. What is meant by the term "formula unit"? Give one example.

Answer:

Formula unit is the simplest ratio of ions present in an ionic compound.

Example: NaCl shows a 1:1 ratio of Na^+ and Cl^- ions.

2. Differentiate between the chemical formula of an element and a compound with examples.

Answer:

- The chemical formula of an element shows one or more atoms of the same element, e.g., O_2, Fe.
- The chemical formula of a compound shows atoms of different elements chemically combined, e.g., H_2O, NaCl.

3. Write the molecular formula and empirical formula of benzene.

Answer:

- Molecular formula of benzene: C_6H_6
- Empirical formula of benzene: CH

4. What does the formula H_2O represent? Explain briefly.

Answer:

The formula H_2O represents a water molecule consisting of two hydrogen atoms bonded with one oxygen atom.

5. Name two covalent compounds and write their molecular formulas.

Answer:

1. Ammonia – NH_3

2. Methane – CH_4

6. What information does an empirical formula provide about a compound?

Answer:

An empirical formula gives the simplest whole-number ratio of atoms present in a compound.

7. Write down the chemical formula of magnesium nitride using the crisscross method.

Answer:

Magnesium ion: Mg^{2+}, Nitride ion: N^{3-}

By crisscross method: Mg_3N_2

8. How is the chemical formula of an ionic compound written? Explain with an example.

Answer:

By writing symbols of cation and anion with their charges and using the crisscross method.

Example: For aluminium oxide – Al^{3+} and O^{2-} 'n Al_2O_3

9. Give an example of a compound that has the same empirical and molecular formula.

Answer:

Water (H_2O) has the same empirical and molecular formula.

10. State the mathematical formula used to find the molecular formula from empirical formula and molar mass.

Answer:

Molecular formula = n × Empirical formula

Where: n = Molar Mass / Empirical Formula Mass

11. How can you deduce the molecular formula from a structural formula?

Answer:

To deduce the molecular formula:

1. Count the number of each type of atom in the structural formula.

2. Write their chemical symbols.

3. Write the number of each atom as subscript.

4. Do not write subscript "1".

Example: Structural formula shows 2 carbon and 6 hydrogen atoms $\Rightarrow$ Molecular formula = C_2H_6

12. What is Avogadro's number?

Answer:

Avogadro's number is 6.022×10^{23}.

It represents the number of atoms, molecules, or ions in one mole of a substance.

13. Why is Avogadro's number important in chemistry?

Answer:

Avogadro's number helps convert atomic masses into gram masses, making chemical calculations easier in laboratory conditions.

14. Define a mole.

Answer:

A mole is a quantity that contains 6.022×10^{23} particles (atoms, molecules, or ions) of a substance.

Example: 1 mole of oxygen (O_2) = 6.022×10^{23} molecules = 32 g

15. What is molar mass?

Answer:

- Molar mass is the mass of one mole of a substance.
- It is expressed in grams per mole (g/mol).
- **Example**: Molar mass of H_2O = 18.016 g/mol

16. How many molecules are present in one mole of CO_2?

Answer:

One mole of CO_2 contains exactly 6.022×10^{23}

molecules.

17. What is the molar mass of water (H_2O)?

Answer:

H = 1.008, O = 16

Molar mass = 2 × 1.008 + 16 = 18.016 g/mol

18. What is the relationship between atomic mass and molar mass?

Answer:

Atomic mass (in amu) of an element is numerically equal to its molar mass (in grams per mole).

Example: Carbon: 12 amu ⇒ Molar mass = 12 g/mol

19. If a compound has a molar mass of 58.5 g/mol, how many formula units are there in 58.5 g?

Answer:

There are exactly 6.022×10^{23} formula units in 58.5 g of the compound (i.e., one mole).

20. Write the molecular formula of a compound that contains 2 carbon atoms and 6 hydrogen

atoms.

Answer:

Molecular formula = C_2H_6 (This is the formula of ethane).

21. What is a chemical equation?

A chemical equation is a symbolic representation of a chemical reaction using symbols of elements and formulas of compounds.

22. What are reactants and products?

Reactants are substances that take part in a chemical reaction. Products are substances formed as a result of the reaction.

23. How is a chemical equation written correctly?

Reactants are written on the left-hand side, products on the right-hand side, and an arrow separates them.

24. Why should a chemical equation be balanced?

To obey the law of conservation of mass, meaning no atoms are created or destroyed in the reaction.

25. What are the physical state symbols used in chemical equations?

- (s) for solid
- (l) for liquid
- (g) for gas
- (aq) for aqueous (dissolved in water)

26. What is meant by mole ratio in a chemical equation?

It is the ratio of moles of reactants to moles of products as shown in a balanced chemical equation.

27. What is a reversible reaction? Give an example.

A reversible reaction is one in which products can change back into reactants.

Example:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

28. Give an example of a reaction involving ions.

$AgNO_3(aq) + NaCl(aq) \Rightarrow AgCl(s) + NaNO_3(aq)$

(AgCl is a white precipitate.)

29. What does the following chemical equation tell us?

$Zn(s) + H_2SO_4(aq) \Rightarrow ZnSO_4(aq) + H_2(g)$

It tells that 1 mole of zinc reacts with 1 mole of sulfuric acid to form 1 mole of zinc sulfate and 1 mole of hydrogen gas.

30. Why is experimental verification important before writing a chemical equation?

Because the correct formulas, physical states, and products must be confirmed to ensure an accurate chemical equation.

Important Long Questions:

Q1: Differentiate between the chemical formula of an element and that of a compound. Explain with examples.

Answer:

A chemical formula is a symbolic representation that shows the types and numbers of atoms in an

element or compound.

- The **chemical formula** of an element represents either a single atom or a molecule made up of atoms of the same kind.
- Some elements exist as single **atoms**, represented by their symbols alone, e.g., Na (**Sodium**), Ca (**Calcium**), Fe (Iron).
- Other elements exist as molecules composed of two or more atoms **chemically bonded** together, e.g., O_2 (**Oxygen**), N_2 (**Nitrogen**), H_2 (**Hydrogen**), O_3 (**Ozone**).
- The **chemical formula** of a compound represents two or more different elements chemically combined in a fixed ratio.
- **Ionic compounds** are represented by their formula units, showing the simplest whole-number ratio of ions, e.g., NaCl (Sodium chloride), $CaCl_2$ (Calcium chloride), KBr (Potassium bromide).
- **Covalent compounds** are represented by

molecular formulas, showing the exact number of atoms of each element in a molecule, e.g., H_2O (Water), NH_3 (Ammonia), CH_4 (Methane).

Q2: What is a chemical formula? Describe the types of chemical formulas with suitable examples.

Answer:

A chemical formula is a way of representing a substance using symbols of elements and numbers to show the proportions of atoms in the substance.

There are two main types of chemical formulas:

1. Molecular Formula:

- It shows the actual number of atoms of each element in one molecule of a compound.
- Used mainly for covalent compounds.

Examples:

Water: H_2O (2 hydrogen atoms and 1 oxygen atom)

Methane: CH_4 (1 carbon atom and 4 hydrogen

atoms)

Ammonia: NH_3

2. Formula Unit:

- It shows the simplest whole number ratio of ions in an ionic compound.
- Used mainly for ionic compounds.

Examples:

- Sodium chloride: NaCl (1 sodium ion and 1 chloride ion)
- Calcium chloride: $CaCl_2$ (1 calcium ion and 2 chloride ions)

Difference between Molecular and Empirical Formula:

- The **molecular formula** shows the actual number of atoms in a molecule.
- The **empirical formula** shows the simplest whole number ratio of atoms.
- **For example**, hydrogen peroxide has a molecular formula H_2O_2 but its empirical

formula is HO, which shows the simplest ratio (1:1) of hydrogen to oxygen.

Q3: Define empirical formula and explain how it is different from molecular formula. Give suitable examples.

Answer:

- The empirical formula of a compound is the simplest whole number ratio of atoms of each element present in the compound.
- It does not show the actual number of atoms but rather the smallest ratio.
- Empirical formulas are commonly used to represent ionic compounds because they exist as a lattice of ions.
- The molecular formula shows the exact number of atoms of each element in one molecule of the compound.

Differences:

- Aspect Empirical Formula Molecular Formula

- Represents Simplest ratio of atoms Actual number of atoms in molecule
- Used for Ionic compounds, some covalent Covalent compounds
- Example $H_2O_2 \Rightarrow HO$, C_6H_6 'n CHH_2O_2, C_6H_6

Examples:

- Hydrogen Peroxide:
- Molecular formula: H_2O_2
- Empirical formula: HO

Benzene:

- Molecular formula: C_6H_6
- Empirical formula: CH

Water:

- Molecular formula and empirical formula are the same: H_2O
- Some different compounds may share the same empirical formula:
- Benzene (C_6H_6) and Acetylene (C_2H_2) both have empirical formula CH.

Q4: What is Avogadro's Number? Explain its significance in chemistry with suitable examples.

Avogadro's Number is defined as the number of particles (atoms, molecules, or ions) present in one mole of a substance. Its value is approximately 6.022×10^{23}. This means that one mole of any substance contains exactly 6.022×10^{23} fundamental units of that substance.

Significance in Chemistry:

- Avogadro's Number provides a bridge between the microscopic world of atoms and molecules and the macroscopic world we can measure in the laboratory.
- It allows chemists to count particles by weighing them. Instead of counting individual atoms (which is impossible because they are extremely small), chemists measure the amount of substance in moles.
- It helps in converting atomic mass units (amu) to grams, making it easier to calculate masses

involved in chemical reactions.

Example:

- Consider carbon. One atom of carbon has an atomic mass of 12 amu. But 12 amu is an extremely small mass. When we take 6.022×10^{23} atoms of carbon (1 mole), the total mass becomes 12 grams. So, 1 mole of carbon atoms weighs 12 grams.

In a chemical reaction like:

- 2 moles of carbon atoms ($2 \times 6.022 \times 10^{23}$ atoms) react with 1 mole of oxygen molecules (6.022×10^{23} molecules) to form 2 moles of carbon monoxide molecules.
- This helps to relate the number of particles with the measurable mass in grams.

Q5: Explain the relationship between the number of atoms/molecules and their masses using Avogadro's Number, including how 1 gram is related to 6.022×10^{23} amu.

Atoms and molecules have extremely small masses

measured in atomic mass units (amu). For example, 1 carbon atom weighs about 12 amu, which is too small to measure directly. Avogadro's Number allows us to scale these tiny masses up to amounts that can be measured in grams.

Relationship:

- 1 atomic mass unit (amu) is defined as 1/12th the mass of one carbon-12 atom.
- The mass of 1 mole of atoms (6.022×10^{23} atoms) of any element in grams is numerically equal to its atomic mass in amu.
- This means, 1 gram of a substance contains approximately (6.022×10^{23}) ÷ (atomic mass) atoms or molecules.

Using Avogadro's Number:

- The total mass of 6.022×10^{23} atoms of an element = atomic mass in grams.
- For example, oxygen molecule (O_2) has a molecular mass of 32 amu. So, 1 mole (6.022×10^{23} molecules) weighs 32 grams.

How 1 gram is related to 6.022×10^{23} amu:

- 1 gram of a substance contains Avogadro's Number of atomic mass units.
- In other words, 1 gram = 6.022×10^{23} amu.
- This allows us to convert the microscopic scale (amu) to the laboratory scale (grams) easily.

Q6: Define the term "mole". How is Avogadro's number related to the concept of a mole? Explain with examples.

Answer:

A mole is a unit in chemistry that represents a specific quantity of particles, whether they are atoms, molecules, or ions. One mole contains Avogadro's number of particles, which is 6.022×10^{23}. This is similar to how a "dozen" means 12 items, or a "gross" means 144 items. The mole allows chemists to count huge numbers of tiny particles in a practical way by measuring their mass.

Examples:

- **One mole** of **carbon atoms** contains 6.022×10^{23} atoms and weighs exactly 12 grams.
- **One mole** of **oxygen** molecules (O_2) contains 6.022×10^{23} **molecules** and weighs 32 grams.
- **One mole** of **sodium chloride** (NaCl) contains 6.022×10^{23} formula units and weighs 58.5 grams.
- This **relationship** connects the **microscopic** world of **atoms** and **molecules** with the macroscopic world of grams and laboratory measurements.

Q7: What is molar mass? How does molar mass relate to atomic mass and molecular mass? Give examples to support your answer.

Answer:

The molar mass of a substance is the mass of one mole of that substance, expressed in grams per mole (g/mol). It numerically equals the atomic or molecular mass but is measured in grams instead of atomic mass units (amu).

- The **atomic mass** of an element is the mass of one atom, measured in amu.
- The **molecular mass** is the sum of atomic masses of all atoms in a molecule.

Relation:

- For an **element**, the **molar mass** (in grams) is equal to its atomic mass (in amu).
- For a **molecule**, the **molar mass** (in grams) is equal to its molecular mass (in amu).

Examples:

- **Hydrogen atom atomic mass** = 1.008 amu, so molar mass of hydrogen atoms = 1.008 g/mol.
- **Hydrogen molecule** (H_2) **molecular mass** = 2 × 1.008 = 2.016 amu, so molar mass of hydrogen molecules = 2.016 g/mol.
- **Oxygen molecule** (O_2) molecular mass = 32 amu, so molar mass = 32 g/mol.

Q8: Define a chemical equation and explain its main components with an example. Why is it

important to balance a chemical equation?

A chemical equation is a symbolic representation of a chemical reaction using chemical formulas and symbols. It shows the substances that react (called reactants) on the left side and the substances produced (called products) on the right side, separated by an arrow (‘n) indicating the direction of the reaction.

Main components of a chemical equation:

- **Reactants**: These are the starting substances that undergo change in the reaction. They are written on the left side of the equation.
- **Products**: These are the new substances formed as a result of the reaction, written on the right side.
- **Arrow** (⇒): This separates reactants from products and shows the direction of the reaction.
- **Coefficients**: Numbers placed before formulas to indicate the mole ratio of reactants and

products.

- Physical state symbols: These indicate the physical state of substances; (s) for solid, (l) for liquid, (g) for gas, and (aq) for aqueous (dissolved in water).

Example:

The reaction between zinc and sulfuric acid is represented as:

$Zn\ (s) + H_2SO_4\ (aq) \Rightarrow ZnSO_4\ (aq) + H_2\ (g)$

- Zn and H_2SO_4 are reactants.
- $ZnSO_4$ and H_2 are products.

Importance of balancing a chemical equation:

Balancing a chemical equation ensures that the law of conservation of mass is obeyed — meaning the total number of atoms of each element remains the same on both sides of the equation. This is essential because atoms cannot be created or destroyed in a chemical reaction. Balancing helps us know the exact proportion (mole ratio) of reactants and products involved, which is crucial for

quantitative chemical calculations and for understanding the reaction properly.

Q9: Discuss the law of conservation of mass in relation to chemical equations. How does it affect the way chemical equations are written?

The law of conservation of mass states that mass can neither be created nor destroyed in a chemical reaction. This means the total mass of the reactants must equal the total mass of the products.

In terms of chemical equations, this law requires that:

- The number of atoms of each element must be the same on both sides of the equation.
- No atom is lost or gained during the reaction; atoms are only rearranged.
- Because of this, chemical equations must be balanced by adjusting coefficients in front of formulas so that the number of atoms for each element is equal on both sides. This ensures

the equation accurately represents what happens in reality.

For example, in the reaction of aluminum with oxygen to form aluminum oxide:

Unbalanced:

$Al + O_2 \Rightarrow Al_2O_3$

Balanced:

$4Al + 3O_2 \Rightarrow 2Al_2O_3$

Here, the number of Al and O atoms are equal on both sides, obeying the law of conservation of mass.

Q10: What is a chemical reaction? Describe the different types of chemical reactions with suitable examples.

A chemical reaction is a process in which one or more substances (reactants) change into new substances (products) with different properties. During a chemical reaction, chemical bonds are broken and new bonds are formed.

Types of chemical reactions:

1. **Combination** (Synthesis) **Reaction**: Two or more substances combine to form a single product.

Example:

$2H_2\,(g) + O_2\,(g) \Rightarrow 2H_2O\,(l)$

2. **Decomposition Reaction:** A single compound breaks down into two or more simpler substances.

Example:

$2H_2O_2\,(aq) \Rightarrow 2H_2O\,(l) + O_2\,(g)$

3.**Displacement** (**Single Replacement**) **Reaction**: One element replaces another element in a compound.

Example:

$Zn\,(s) + 2HCl\,(aq) \Rightarrow ZnCl_2\,(aq) + H_2\,(g)$

4. **Double Displacement** (Metathesis) **Reaction**: Exchange of ions between two compounds to form new compounds.

Example:

$AgNO_3\,(aq) + NaCl\,(aq) \Rightarrow AgCl\,(s) + NaNO_3\,(aq)$

5. **Combustion Reaction**: A substance reacts with

oxygen, releasing energy in the form of heat and light.

Example:

$CH_4\ (g) + 2O_2\ (g) \Rightarrow CO_2\ (g) + 2H_2O\ (g)$

6. **Reversible Reaction**: Reactions that can proceed in both forward and backward directions, reaching equilibrium.

Example:

$N_2\ (g) + 3H_2\ (g) \rightleftharpoons 2NH_3\ (g)$

3. Constructed Response Questions:

i. Different compounds will never have the same molecular formula but they can have the same empirical formula. Explain.

- This statement is actually incorrect as stated — it is the opposite:
- Different compounds can have the same molecular formula (called isomers), but they always have the same empirical formula if they

have the same molecular formula.

- Molecular formula shows the exact number of atoms of each element in a molecule (e.g., C_2H_6O).
- Empirical formula shows the simplest whole number ratio of atoms in a compound (e.g., CH_3O for C_2H_6O).

Explanation:

- **Different compounds** may have the same molecular formula but different structures; these are called isomers. For **example**, ethanol (C_2H_6O) and dimethyl ether (C_2H_6O) have the same molecular formula but different properties because their atoms are arranged differently.
- However, **empirical formula** only shows the simplest ratio of elements. Different compounds can have the same empirical formula if their elements are in the same ratio but their molecular formulas differ by a multiple.

For example:

Glucose: $C_6H_{12}O_6$ (molecular formula), empirical formula: CH_2O

Formaldehyde: CH_2O (both molecular and empirical formula)

So compounds with molecular formulas like $C_6H_{12}O_6$ and CH_2O have the same empirical formula but different molecular formulas.

ii. Write down the chemical formulas of the following compounds:

- Calcium phosphate: $Ca_3(PO_4)_2$
- Aluminium nitride: AlN
- Sodium acetate: CH_3COONa or $C_2H_3O_2Na$
- Ammonium carbonate: $(NH_4)_2CO_3$
- Bismuth sulphate: $Bi_2(SO_4)_3$

iii. Why does Avogadro's number have an immense importance in chemistry?

Avogadro's number (6.022×10^{23}) represents the number of particles (atoms, molecules, or ions) in

one mole of a substance. Its importance lies in:

Relating microscopic and macroscopic worlds: It connects the number of tiny particles to measurable amounts we can handle in the lab.

Mole concept: It allows chemists to count particles by weighing substances, making it easier to understand and calculate reactions.

Stoichiometry: It helps in calculating how much of each reactant is needed or product formed in chemical reactions.

Universal constant: It applies to all substances, enabling a standard way to express quantities in chemistry.

iv. When 8.657g of a compound were converted into elements, it gave 5.217g of carbon, 0.962g of hydrogen, and 2.478g of oxygen. Calculate the percentage of each element present in this compound.

To calculate the percentage of each element in the compound, we use the formula:

$$\text{Percentage of element} = \left(\frac{\text{Mass of element}}{\text{Total mass of compound}} \right) \times 100$$

Given:

Total mass of compound = 8.657 g

Mass of Carbon (C) = 5.217 g

Mass of Hydrogen (H) = 0.962 g

Mass of Oxygen (O) = 2.478 g

Step-by-step Calculation:

1. **Carbon:**

$$\frac{5.217}{8.657} \times 100 = 60.26\%$$

1. **Hydrogen:**

$$\frac{0.962}{8.657} \times 100 = 11.11\%$$

1. **Oxygen:**

$$\frac{2.478}{8.657} \times 100 = 28.63\%$$

Final Answer:

- Carbon = **60.26%**
- Hydrogen = **11.11%**
- Oxygen = **28.63%**

V. How can you calculate the masses of the products formed in a reversible reaction?

To calculate the masses of products formed in a reversible reaction, follow these steps:

1. Write the balanced chemical equation for the reaction, including the reversible arrow ($\rightleftharpoons$), to know the mole ratio of reactants and products.

2. Determine the amount (mass or moles) of reactants given initially.

3. Convert the masses of reactants into moles using their molar masses.

4. Use the mole ratio from the balanced equation to calculate the number of moles of products formed.

5. Consider the equilibrium position if the reaction is not complete; at equilibrium, the reaction does

not proceed to full completion. You may need additional data such as equilibrium concentrations or extent of reaction.

6. Convert the moles of products back to masses by multiplying by their respective molar masses.

Example outline:

If the reaction is:

$A + B \rightleftharpoons C + D$

- If you start with known masses of A and B, calculate their moles.
- Use the balanced equation mole ratio to find how many moles of C and D are formed.
- If given equilibrium data, adjust the moles accordingly.
- Convert moles of C and D to grams to get product masses.

4. Descriptive Questions:

i. Which conditions must be fulfilled before writing a chemical equation for a reaction?

Before writing a chemical equation, the following conditions must be fulfilled:

- The reactants and products must be clearly identified through experiments.
- The correct chemical formulas of all reactants and products must be known.
- The physical states (solid, liquid, gas, aqueous) of all substances should be determined.
- The reaction must obey the law of conservation of mass, so the equation should be balanced with equal atoms on both sides.
- The mole ratio between reactants and products must be established.
- The direction of the reaction should be indicated (irreversible $\Rightarrow$ or reversible $\rightleftharpoons$).

ii. Explain the concepts of Avogadro's number and mole.

- Avogadro's Number (NA) is the number of particles (atoms, molecules, ions) present in one mole of a substance. Its value is 6.022×10^{23} particles/mol.
- A mole is a unit representing a specific number of particles, just like a dozen means 12 items. One mole contains exactly Avogadro's number of particles.
- The mole helps chemists count particles by weighing them because counting individual atoms or molecules is impossible. For example, one mole of carbon atoms contains 6.022×10^{23} carbon atoms and weighs 12 grams.

iii. How many grams of CO_2 will be produced when we react 10 g of CH_4 with excess of O_2?

Given reaction:

$CH_4 + 2O_2 \Rightarrow CO_2 + 2H_2O$

Step 1: Calculate moles of CH_4:

- Molar mass of CH_4 = 12 + (4 × 1) = 16 g/mol
- Moles of CH_4 = 10 g ÷ 16 g/mol = 0.625 mol

Step 2: Use mole ratio to find moles of CO_2:

- From the equation, 1 mole CH_4 produces 1 mole CO_2
- So, 0.625 moles CH_4 produces 0.625 moles CO_2

Step 3: Calculate mass of CO_2:

- Molar mass of CO_2 = 12 + (2 × 16) = 44 g/mol
- Mass of CO_2 = 0.625 × 44 = 27.5 g

Answer: 27.5 grams of CO_2 will be produced.

iv. How many moles of coal (carbon) are needed to produce 10 moles of CO?

Given reaction:

$3C + O_2 + H_2O \Rightarrow H_2 + 3CO$

- From the equation, 3 moles of C produce 3 moles of CO. So, mole ratio of C to CO is 1:1.
- To produce 10 moles of CO, moles of C needed

= 10 moles.

Answer: 10 moles of coal (carbon) are needed.

v. How much SO_2 (in grams) is needed to produce 10 moles of sulphur?

Given reaction:

$2H_2S + SO_2 \Rightarrow 2H_2O + 3S$

- **From the equation:**
- 1 mole of SO_2 produces 3 moles of S
- Mole ratio SO_2 : S = 1 : 3

To produce 10 moles of sulphur:

- Moles of SO_2 needed = (10 × 1) ÷ 3 = 3.33 moles
- Molar mass of SO_2 = 32 + (2 × 16) = 64 g/mol
- Mass of SO_2 = 3.33 × 64 = 213.12 g

Answer: 213.12 grams of SO_2 are needed.

vi. How much ammonia is needed in grams to produce 1 kg of urea fertilizer?

Given reaction:

$2NH_3 + CO_2 \Rightarrow CO(NH_2)_2 + H_2O$

Molar mass of urea ($CO(NH_2)_2$) = 12 + 16 + (2 × (14 + 2)) = 60 g/mol

1 kg = 1000 g

- **Moles** of urea produced = 1000 ÷ 60 = 16.67 moles
- From **equation**, 2 moles NH_3 produce 1 mole urea
- Moles of NH_3 needed = 16.67 × 2 = 33.34 moles
- **Molar mass** of NH_3 = 14 + (3 × 1) = 17 g/mol
- **Mass** of NH_3 = 33.34 × 17 = 566.78 g

Answer: 566.78 grams of ammonia are needed.

vii. Calculate the number of atoms in the following:

(a) 3 g of H_2

- Molar mass of H_2 = 2 g/mol
- Moles of H_2 = 3 ÷ 2 = 1.5 moles

- Number of molecules = $1.5 \times 6.022 \times 10^{23} = 9.033 \times 10^{23}$ molecules
- Number of atoms = molecules × 2 = $9.033 \times 10^{23} \times 2 = 1.8066 \times 10^{24}$ atoms

(b) 3.4 moles of N_2

- Number of molecules = $3.4 \times 6.022 \times 10^{23} = 2.047 \times 10^{24}$ molecules
- Number of atoms = molecules × 2 = $2.047 \times 10^{24} \times 2 = 4.094 \times 10^{24}$ atoms

(c) 10 g of $C_6H_{12}O_6$ (glucose)

- Molar mass of $C_6H_{12}O_6$ = (6 × 12) + (12 × 1) + (6 × 16) = 180 g/mol
- Moles = 10 ÷ 180 = 0.0556 moles
- Number of molecules = $0.0556 \times 6.022 \times 10^{23} = 3.35 \times 10^{22}$ molecules
- Number of atoms in one molecule = 6 + 12 + 6 = 24 atoms
- Total atoms = $3.35 \times 10^{22} \times 24 = 8.04 \times 10^{23}$ atoms

Note:

This chapter is designed to provide a solid foundation of knowledge, with the goal of deepening understanding and encouraging further exploration of the subject. The content has been carefully selected to support effective learning and inspire students to engage with the topic more deeply.

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Purpose: To contribute to education by offering insightful, valuable content that enhances learning and understanding.

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