CHAPTER – 1

Some Basic Concepts of Chemistry

2marks:

1. Calculate the molecular mass of the following:

i) H2O2

Answer:

H2O

The molecular mass of water, H2OH2O, can be calculated by the following steps given below:

= (2 × Atomic mass of hydrogen element)

= (2 × 1.0084 u) + (1 × 16.00 u)

= 2.016 u + 16.00 u

= 18.016 u or 18.02 u

ii) CO2

The molecular mass of carbon dioxide, CO2, is calculated below:

= (1 ×× Atomic mass of carbon) + (2 ×× Atomic mass of oxygen)

= (1 × 12.011 u) + (2 × 16.00 u)

= 12.011 u + 32.00 u

= 44.01 u

iii) CH_4

The molecular mass of methane, CH_4, is calculated below in step by step manner:

= (1 ×× Atomic mass of carbon) + (4 ×× Atomic mass of hydrogen)

= (1 × 12.011 u) + (4 × 1.008 u) (1 × 12.011 u) + (4 × 1.008 u)

= 12.011 u +4.032 u

= 16.043 u

2. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4)**.**

Answer:

The given compound in the question is sodium sulphate and its formula is Na2SO4Na2. Its molecular formula is calculated below:

Na_2SO_4=[(2 × 23.0)+(32.066)+4(16.00)]

= 142.066 g

Now, let us find the mass percentage of each element in the given compound using the formula given below:

$$\text{Mass percent of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar compound of the element}} *100$$

∴Mass percent of sodium:

=46.0 g142.066 g × 100

= 32.379

= 32.4%

Now, let us find the mass percentage of sulphur:

=32.066 g142.066 g × 100

= 22.57

= 22.6%

Now, the mass percentage of oxygen:

=64.0 g142.066 g × 100

= 45.049

= 45.05%

3.How much copper can be obtained from 100 g of copper sulphate (CuSO4)**?**

Answer:

In copper sulfate, we can see that there is one atom of copper so, we can say that 1 mole of CuSO4will have 1 mole of copper.

The molar mass of copper sulphate is calculated below:

CuSO4= 63.5 + 32 + (4 × 16)

= 63.5 + 32.0 + 64.0

= 159.5 g

We can say that 159.5 g of CuSO4 will have 63.5 g of copper.

⇒100 g of CuSO4⇒ will contain 63.5 × 100 g1 / 59.563.5

So, the amount of copper that can be obtained from 100 g of CuSO4=63.5 × 100 g1 / 59.563.5

= 39.81 g

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

	% Natural Abundance	Molar Mass
35Cl35Cl	75.77	34.9689
37Cl37Cl	24.23	36.9659

Answer:

The average atomic mass of chlorine is calculated below:

= [(Fractional abundance of 35Cl) (Molar mass of 35Cl) + (Fractional abundance of 37Cl) (Molar mass of 37Cl]

= [{(75.77 / 100) (34.9689u)} + {(24.23 / 100) (36.9659u)}

= 26.4959 + 8.9568

= 35.4527 u

So, the average atomic mass of chlorine is 35.4527 u.

5. In three moles of ethane(C2H6)

calculate the following:

i). Number of moles of carbon atoms.

Answer:

2 Moles of carbon atoms are present in each mole of C2H6

Therefore, number of moles of carbon atoms in 3 moles of C2H6

= 2 × 3 = 6

ii) Number of moles of hydrogen atoms.

Answer:

6 moles of hydrogen atoms are present in each mole of C2H6

Therefore, number of moles of carbon atoms in 3 moles ofC2H6

= 3 × 6 = 18

iii) Number of molecules of ethane.

Answer:

6.023 × 1023

molecules of ethane are present in each mole of C2H6

Therefore, number of molecules in 3 moles of C2H6

=3×6.023×1023=18.069×1023

6. What is the concentration of sugar (C12H22O11)
in mol L-1
if its 20 g are dissolved in enough water to make a final volume up to 2 L?

Answer:

Molarity (M) of the Ans can be calculated by the formula given below:

=Number of moles of solute Volume of solution in Litres

This can be further written as

=Mass of sugar / molar mass of sugar2 L

Putting the values, we get:

=20 g / [(12 × 12) + (1 × 22) + (11 × 16)] g2 L

=20 g/ 342 g2 L

=0.0585mol2L

=0.02925 mol L-1

So, the molar concentration of sugar is =0.02925 mol L-1

7. What is the SI unit of mass? How is it defined?

Answer:

The kilogram is the SI unit of mass in the SI system (kg). An international kilogram prototype is defined as one kilogram.

7. Match the following prefixes with their multiples:

	Prefixes	***Multiples***
(a)	***femto***	***10***
(b)	***giga***	***10−15***
(c)	***mega***	***10−6***
(d)	***deca***	***109***
(e)	***micro***	***106***

Answer:

	Prefixes	Multiples

(a)	femto	10−15
(b)	giga	109
(c)	mega	106
(d)	deca	10
(e)	micro	10^{-6}

8. What do you mean by significant figures?

Answer.

Significant figures are the meaningful digits which are known with certainty. Significant figures indicate uncertainty in experimented value.

e.g.: The result of the experiment is 15.6 mL in that case 15 is certain and 6 is uncertain. The total significant figures are 3.

9. Express the following in the scientific notation:

Answer:

If we want to express the numbers in scientific notation, we must first put the number in decimal form and multiply it by 10 with some power. These are given below:

0.0048

Answer:

0.0048=4.8×10−3

234,000

Ans: 234,000=2.34×105

8008

Ans: 8008=8.008×103

500.0

Ans: 500.0=5.000×102

6.0012

Ans: 6.0012=6.0012×100

10. How many significant figures are present in the following?

Answer:

There are some rules to find the number of significant figures and by following the rules the significant figures are given below:

0.0025

Ans:

There are 2 significant figures.

208

Ans: There are 3 significant figures.

5005

Ans: There are 4 significant figures.

126,000

Ans: There are 3 significant figures.

500.0

Ans: There are 4 significant figures.

2.0034

Ans: There are 5 significant figures.

11. Round up the following up to three significant figures.

Answer:

These can be written as:

34.216

Ans: 34.216 = 34.2

10.4107

Ans: 10.4107 = 10.4

0.04597

Ans: 0.04597 = 0.0460

2808

Ans: 2808 = 2810

12. The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

	Mass of dioxygen	***Mass of dinitrogen***
(i)	***16 g***	***14 g***

(ii)	*32 g*	*14 g*
(iii)	*32 g*	*28 g*
(iv)	*80 g*	*28 g*

Which law of chemical combination is obeyed by the above experimental data?

Give its statement.

(b) Fill in the blanks in the following conversions:

(i) 1 km = …………………. mm = …………………. pm

(ii) 1 mg = …………………. kg = …………………. ng

(iii) 1 mL = …………………. L = …………………. dm ^{3}

Answer:

Here if we fix the mass of dinitrogen at 14g, then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 16g, 32g, 32g, and 80g.

The masses of dioxygen bear a whole number ratio of 1:2:2:5.

Hence, the given experimental data obeys the Law of Multiple Proportions.

13. If the speed of light is 3.0 × 108 ms-1

calculate the distance covered by light in 2.00 ns.

Answer:

From the question we can see that the time taken to cover the distance is 2.00ns. This can be written as:

$=2.00 \times 10\text{-}9s$

We know the speed of light= 3.0×108 ms-1

So, the distance travelled by light in 2.00ns will be:

=Speed of light × Time taken

$= (3.0 \times 108ms\text{-}1)\ (2.00 \times 10\text{-}9s)$

$=6.00 \times 10\text{-}1m$

= 0.600 m

14. How are 0.50 mol Na2CO3 and 0.50 M Na2CO3 different?

Answer:

The molar mass of Na2CO3 is given below:

Na2CO3 = (2 × 23) + 12.00 + (3 × 16)

= 106 g/mol

So, 1 mole of Na2CO3 means 106 g of Na2CO3

Therefore, for 0.5 mol of Na_2CO_3 can be calculated as:

0.5 mol of Na_2CO_3 = 106 g1 mole × 0.5 mol Na_2CO_3

=53g Na_2CO_3

= 0.50 M Na_2CO_3=0.50mol/L Na_2CO_3

15. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapor would be produced?

Answer:

Let us first write the reaction between dihydrogen and dioxygen. The reaction will be:

$2H_2(g)+O_2(g)\rightarrow 2H_2O(g)$

Dioxygen reacts with two volumes of dihydrogen to generate two volumes of water vapour.

As a result, ten volumes of dihydrogen will react with five volumes of dioxygen to generate ten volumes of water Vapours.

16. Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	***Molar mass***	***Abundance***

^{36}AR	***35.96755*** g mol^{-1}	***0.337 %***
^{38}AR	***37.96272*** g mol^{-1}	***0.063 %***
^{40}AR	***39.9624 g mol^{-1}***	***99.600 %***

Answer:

Molar mass of Argon:

= [(35.96755×0.337100) + (37.96272×0.063100) + (39.9624×99.600100)]= [0.121 + 0.024 + 39.802]

$(G^{mol)-1}$= 39.947

4marks:

1.Define mole and Avogadro's number. How are they related?

Answer:

The mole is a unit of measurement used in chemistry to express amounts of a chemical substance. Avogadro's number (6.022 x 10^23) is the number of entities (atoms, molecules, ions, etc.) in one mole of a substance. The relationship is that one mole of any substance contains Avogadro's number of entities.

2.Explain the concept of molar mass. Calculate the molar mass of water (H_2O).

Answer:

Molar mass is the mass of one mole of a substance, expressed in grams per mole. For water (H_2O), the molar mass is calculated as follows:

Molar Mass of H_2O

= (Mass of 2H atoms) + (Mass of 1O atom) Molar Mass of H_2O=

(Mass of 2H atoms) + (Mass of 1O atom)

= (2×Atomic Mass of H) + (Atomic Mass of O)

= (2×Atomic Mass of H) + (Atomic Mass of O)

=(2×1.01g/mol) + (16.00g/mol)

=(2×1.01g/mol) + (16.00g/mol)

=18.02g/mol

=18.02g/mol

3.Define empirical and molecular formulas. Provide an example

to illustrate the difference between the two.

Answer:

Empirical formula represents the simplest whole-number ratio of atoms in a compound, while the molecular formula shows the actual number of atoms of each element in a molecule. For example, glucose has an empirical formula CH_2O, but its molecular formula is $C_6H_{12}O_6$. The empirical formula shows the simplest ratio (1:2:1), while the molecular formula provides the exact number of atoms.

4.Describe the law of definite proportions with an example.

Answer:

The law of definite proportions states that a compound always contains the same elements in the same proportion by mass. For instance, in water (H_2O), the mass ratio of hydrogen to oxygen is always 1:8, regardless of the source or preparation method.

5.Discuss the postulates of Dalton's atomic theory. How have these postulates been modified with advancements in atomic theory?

Answer:

Dalton's atomic theory postulated that elements are composed of indivisible atoms, atoms of different elements combine in simple whole-number ratios to form compounds, and chemical reactions involve the rearrangement of atoms. With advancements, it was found that atoms are divisible, isotopes exist, and subatomic particles (protons, neutrons, electrons) play crucial roles, modifying some of Dalton's postulates.

6.Explain the concept of limiting reactants in a chemical reaction. Provide an example and calculate the limiting reactant.

Answer:

Limiting reactant is the substance that is completely consumed in a chemical reaction, limiting the amount of product formed. For example, consider the reaction

$2A+3B \rightarrow C$. If 4 moles of A and 5 moles of B are present, B is the limiting reactant. The calculation involves determining the moles of each reactant and comparing them to the stoichiometric coefficients.

7.Define the terms molarity and molality. Differentiate between them with examples.

Answer:

Molarity (M) is the concentration of a solution expressed as moles of solute per liter of solution, while molality (m) is the concentration expressed as moles of solute per kilogram of solvent. For instance, a 2 M solution of NaCl means 2 moles of NaCl in 1 liter of solution, while a 2 m solution means 2 moles of NaCl in 1 kg of solvent. Molarity depends on volume, while molality depends on mass.

7marks:

1. Define the term 'mole' and explain how it is used in stoichiometry. Provide an example.

Answer:

The mole is a fundamental unit in chemistry used to quantify entities like atoms, molecules, or ions. In stoichiometry, the mole is crucial for understanding the quantitative relationships between reactants and products in a chemical reaction. Avogadro's number (6.022 x 10^23) is the number of entities in one mole of a substance.

For example, consider the reaction

2H2+O2→2H2O.

Here, two moles of hydrogen molecules (2H2) react with one mole of oxygen molecules (2O2) to produce two moles of water molecules (2H2O). This ratio is determined by the coefficients in the balanced chemical equation.

2.Discuss Dalton's atomic theory and its limitations. How did subsequent discoveries modify the original theory?

Answer:

Dalton's atomic theory, proposed in the early 19th century, suggested that atoms are indivisible, identical, and participate in chemical reactions through rearrangement. However, later advancements in the 19th and 20th centuries led to modifications of Dalton's postulates.

Dalton's theory failed to account for the existence of subatomic particles. J.J. Thomson's discovery of the electron and Rutherford's gold foil experiment, which showed that atoms have a small, dense nucleus, challenged Dalton's idea of indivisibility. Moreover, the existence of isotopes (atoms of the same element with different masses) contradicted Dalton's assumption of identical atoms.

These subsequent discoveries reshaped our understanding of atomic structure, leading to the modern atomic theory that considers atoms as divisible entities with subatomic particles.

3.Explain the concept of empirical and molecular formulas. Provide a step-by-step example to determine the empirical formula from given data.

Answer:

Empirical formulas represent the simplest whole-number ratio of atoms in a compound, while molecular formulas show the actual number of atoms. To determine the empirical formula:

Find the moles of each element: Convert the mass of each element to moles using its molar mass.

Divide by the smallest mole value: Determine the smallest mole value and divide each mole value by it.

Convert to the simplest ratio: Ensure the resulting ratio is in the simplest whole-number ratio.

For example, consider a compound composed of 4.8 g of carbon and 1.6 g of hydrogen. The molar masses of carbon and hydrogen are approximately 12 g/mol and 1 g/mol, respectively.

Moles of carbon (nC) = 4.8g/12g/mol=0.4mol4.8g/12g/mol=0.4mol

Moles of hydrogen (nH) = 1.6g/1g/mol 1.6mol 1.6g/1g/mol=1.6mol

Divide by the smallest mole value (nC): 0.4mol/0.4mol=10.4mol/0.4mol=1and 1.6mol/0.4mol=41.6mol/0.4mol=4The empirical formula is 4CH4

This demonstrates the process of determining the simplest ratio of atoms in a compound based on given data.

4.Explain the concept of limiting reactants in a chemical reaction. Provide a step-by-step example to determine the limiting reactant.

Answer:

The limiting reactant in a chemical reaction is the substance that is entirely consumed, limiting the amount of product formed. To identify the limiting reactant:

Determine moles of each reactant: Convert the given masses of reactants to moles using their respective molar masses.

Use stoichiometry: Write the balanced chemical equation and determine the moles of the product that each reactant can produce.

Identify the limiting reactant: The reactant that produces the least amount of product is the limiting reactant.

For example, in the reaction

$2A+3B \rightarrow C$, if you have 4 moles of A and 5 moles of B, calculate the moles of product for each:

Moles of C from A: $4/2 \times 1=2$ moles

Moles of C from B: $5/3 \times 1=5/3$moles

Since B produces fewer moles of C, B is the limiting reactant.

5.Define the terms molarity and molality. Differentiate between them with examples.

Answer:

Molarity (M):

Molarity is the concentration of a solution expressed as moles of solute per liter of solution.

Formula: =moles of solute/liters of solution

Example: A 1 M solution of NaCl means 1 mole of NaCl in 1 liter of solution.

Molality (m):

Molality is the concentration expressed as moles of solute per kilogram of solvent.

Formula:

m=moles of solute/kilograms of solvent

Example: A 2 m solution of NaCl means 2 moles of NaCl in 1 kg of solvent.

Difference:

Molarity depends on volume and changes with temperature, while molality depends on mass and remains constant with temperature changes.

6.Discuss the concept of empirical and molecular formulas using a specific example.

Answer:

Empirical Formula:

Empirical formulas represent the simplest whole-number ratio of atoms in a compound.

Example: In glucose ($C_6H_{12}O_6$), the empirical formula is CH_2O as it represents the simplest ratio.

Molecular Formula:

Molecular formulas represent the actual number of atoms of each element in a molecule.

Example: In glucose, the molecular formula is $C_6H_{12}O_6$, indicating the actual number of carbon, hydrogen, and oxygen atoms.

Difference:

The empirical formula shows the simplest ratio, while the molecular formula provides the exact number of atoms.

7.Explain the laws of multiple proportions with an example. How do these laws contribute to our understanding of chemical combinations?

Answer: Laws of Multiple Proportions:

These laws state that when two elements combine to form different compounds, the masses of one element that combine with a fixed mass of the other are in simple whole-number ratios.

Example: Consider carbon monoxide (CO) and carbon dioxide (CO_2). The ratio of the mass of oxygen combining with a fixed mass of carbon is 1:2 in CO and 1:1 in CO_2.

Contribution to Understanding:

These laws contribute by revealing the consistent patterns in the combination of elements, emphasizing the role of whole-number ratios in compound formation.

They provided critical evidence supporting the atomic theory and helped develop a more comprehensive understanding of chemical combinations.

8.Discuss the role of the mole concept in chemical calculations. Provide an example illustrating its application.

Answer:

Role of Mole Concept:

The mole concept is essential for quantifying chemical reactions, enabling the conversion between the mass of a substance and the number of entities (atoms, molecules) it contains.

Example:

Consider the reaction $2H2+O2 \rightarrow 2H2O$.

If you have 4 moles of hydrogen (H_2), determine the moles of oxygen needed using the stoichiometry of the reaction.

The balanced equation indicates that 1 mole of O_2 reacts with 2 moles of H_2.

Therefore, for 4 moles of H_2, you would need

$4 \times 1/2 = 2$ moles of O_2.

This example demonstrates how the mole concept facilitates precise calculations involving reactants and products in chemical reactions.

Fill in the blanks:

1. Avogadro's number is the number of entities in one ___________ of a substance.

Answer: mole

2. The ___________ theory assumes that matter is composed of small, indivisible particles called atoms.

Answer: atomic

3. The mass of one mole of a substance is known as its __________ mass.

Answer: molar

4. The __________ is the reactant that limits the amount of product formed in a chemical reaction.

Answer: limiting reactant

5. Molarity is expressed as moles of solute per __________ of solution.

Answer: liter

6. The empirical formula of a compound represents the simplest __________ ratio of its elements.

Answer: whole-number

7. Laws of multiple proportions state that when elements combine to form different compounds, the masses of one element are in simple whole-number __________.

Answer: ratios

8. Molality is the concentration expressed as moles of solute per __________ of solvent.

Answer: kilogram

9. In the context of the mole concept, Avogadro's number is approximately __________ x 10^23.

Answer: 6.022

10. The modern atomic theory considers atoms as divisible entities with subatomic __________.

Answer: particles

Multiple choice:

1. What is the basic unit used to quantify the amount of substance in chemistry?

A. Gram

B. Mole

C. Liter

D. Newton

Answer:

B. Mole

2. According to Dalton's atomic theory, atoms are:

A. Indivisible

B. Divisible

C. Identical

D. A and C

Answer:

D. A and C

3. The molar mass of a substance is defined as:

A. The mass of one molecule

B. The mass of one mole of the substance

C. The mass of one atom

D. The mass of one liter of the substance

Answer:

B. The mass of one mole of the substance

4. In a chemical reaction, the reactant that limits the amount of product formed is called the:

A. Excess reactant

B. Catalyst

C. Limiting reactant

D. Intermediate

Answer:

C. Limiting reactant

5. The concentration of a solution expressed as moles of solute per liter of solution is called:

A. Molality

B. Molarity

C. Normality

D. Mass percent

Answer:

B. Molarity

6. The simplest whole-number ratio of atoms in a compound is represented by its:

A. Molecular formula

B. Structural formula

C. Empirical formula

D. Chemical formula

Answer:

C. Empirical formula

7. According to the laws of multiple proportions, when elements combine to form different compounds, the masses of one element are in:

A. Complex ratios

B. Whole-number ratios

C. Fractional ratios

D. Variable ratios

Answer:

B. Whole-number ratios

8. Molality is a measure of concentration expressed in terms of:

A. Moles of solute per liter of solution

B. Moles of solute per kilogram of solvent

C. Mass of solute per volume of solution

D. Moles of solute per mole of solvent

Answer:

B. Moles of solute per kilogram of solvent

9. Avogadro's number represents:

A. The number of molecules in one mole

B. The number of atoms in one mole

C. The number of ions in one mole

D. All of the above

Answer:

D. All of the above

10. The modern atomic theory considers atoms as divisible entities with subatomic particles, including:

A. Electrons

B. Protons

C. Neutrons

D. All of the above

Answer: D. All of the above

Summary:

The "Basic Concepts of Chemistry" chapter in delves into fundamental principles that serve as the cornerstone of the subject. It introduces the concept of matter, classifying it into elements, compounds, and mixtures, emphasizing the atomic and molecular nature of substances. Dalton's Atomic Theory is explored, proposing that atoms are indivisible and participate in reactions through rearrangement. The chapter delves into atomic and molecular masses, introducing Avogadro's number and the mole concept.

Students learn about empirical and molecular formulas, with practical applications in determining formulas through chemical analysis. Stoichiometry is introduced, focusing on limiting reactants and product yields in chemical reactions. The concentration of solutions is discussed, covering molarity and molality calculations