RELATIVE ATOMIC MASS A_r

- Relative atomic mass of an element is the average mass of one atom of **an element** compared with one-twelfth (1/12) of the mass of a carbon-12 atom
  - No unit
  - $Ar = \frac{average\ mass\ of\ one\ atom\ of\ an\ element}{1/12 \times mass\ of\ carbon{-}12\ atom}$
  - Ar = sum of (% abundance × isotopic mass)

Example question: Find the relative atomic mass of Iron ions from 3 isotopes given:

Iron isotopes	Fe-54	Fe-56	Fe-57
% abundance	5.95	91.88	2.17

Answer:

$$Ar = \frac{5.95}{100} \times 54 + \frac{91.88}{100} \times 56 + \frac{2.17}{100} \times 57 \approx 55.9$$

RELATIVE MOLECULAR MASS M_r

- Relative molecular mass of an element is the average mass of **one molecule** compared with one-twelfth (1/12) of the mass of a carbon-12 atom
  - No unit
  - $Mr = \frac{average\ mass\ of\ one\ molecule\ of\ a\ substance}{1/12 \times mass\ of\ carbon{-}12\ atom}$
  - $Mr\ of\ compound = total\ Ar\ of\ all\ atoms$

Example: Calculate the Mr of CaO:

Answer:

$$Mr = 1Ar\ of\ Mg + 1Ar\ of\ O = 1 \times 40 + 1 \times 14 = 54$$

CHEMICAL FORMULA

- Molecular formula is the **actual ratio** of the number of each type of atom in a substance.
  - Eg 2 Iron(III) atoms combine with 2 oxygen atoms to form Fe_2O_3
- Empirical formula is the **simplest ratio** of the number of atoms in a substance
  - Eg molecular formula of glucose ($C_6H_{12}O_6$) has an empirical formula of CH_2O
  - Can be obtained by mass or percentage

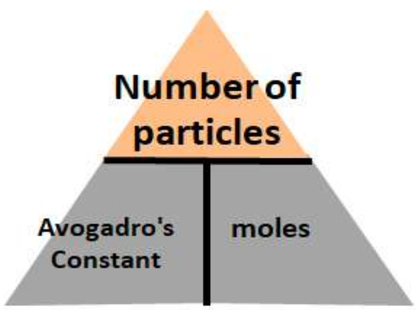
Example: Calculate empirical formula of a compound formed when 16g of oxygen is combined with 103g of lead

Answer:

Step 1	Elements involved	Pb	O
Step 2	Mass/% given	103	16
Step 3	Ar of each element	207	16
Step 4	Step 2/Step 3	0.4976	1
Step 5	Use the least value from step 4 to divide	$\frac{0.4976}{0.4976} = 1$	$\frac{1}{0.4976} = 2$
Step 6	Empirical formula	PbO_2	

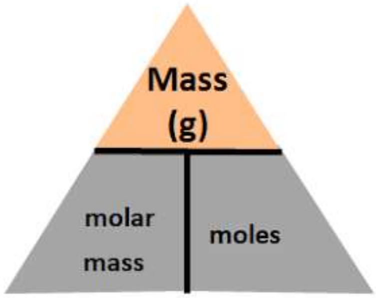
Example:	Empirical formula of the substance is CH_2O. Relative M_r is 180. Determine molecular formula for the substance
Answer:	Step 1: $n = \frac{Mr}{sum\ of\ Ar\ in\ empirical\ formula} = \frac{180}{30} = 6$ Step 2: Molecular formula $(X_aY_b)_n = (CH_2O)_6 = C_6H_{12}O_6$

MOLE CONCEPT



- When the mass of a sample equals to its Ar, quantity of atoms in sample is 6×10^{23}
  - 6×10^{23} is Avogadro's Constant, a fixed quantity
  - Eg Ar of O is 16. A sample of 16g of oxygen contains 6×10^{23} oxygen atoms
  - 1 mole represents a quantity of 6×10^{23}
  - $Mole = \frac{no.of\ particles}{6 \times 10^{23}}$

Example question:	How many moles of water molecules (H2O) are there in 2.0 x 10^24 water molecules?
Answer:	Amount = Number of particles / Avogadro's number = (2.0 x 10^24) / (6.02 x 10^23)≈ 3.32 moles



- Molar mass (M_r) is the mass of 1 mole of substance
- Unit is g/mol or g mol^{-1}
  - $Mole = \frac{mass\ of\ substance}{M_r}$

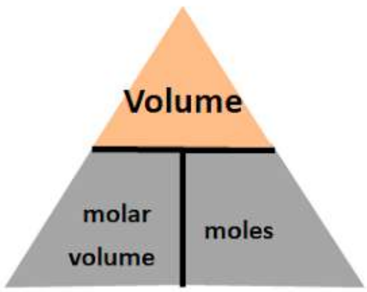
Example question:	A sample of a compound has a molar mass of 18.01528 g/mol. How many moles of this compound are present in a 90.0764 g sample?
Answer:	Number of moles = Mass (g) / Molar mass (g/mol) = 90.0764 g / 18.01528 g/mol ≈ 5 moles

Example question:	A sample of a compound contains 0.25 moles. If the mass of the sample is found to be 15.5 grams, what is the molar mass of the compound?
Answer:	Molar mass (g/mol) = Mass (g) / Number of moles = 15.5 g / 0.25 moles = 62 g/mol

Example question: Calculate the mass of oxygen gas (O_2) needed to completely react with 8.0 grams of hydrogen gas (H_2) to form water (H_2O).

Answer:

Step	Description	Calculation	Result
1	Write the balanced equation	$2H_2 + O_2 \rightarrow 2H_2O$	
2	Calculate the molar mass	Molar mass of H_2 = 2.02 g/mol Molar mass of O_2 = 32.00 g/mol Molar mass of H_2O = 18.02 g/mol	
3	Calculate moles of H_2	Moles of H_2 = Mass (g) / Molar mass (g/mol) = 8.0 g / 2.02 g/mol	3.96 moles
4	Use stoichiometry to find moles of O_2	Moles of O_2 = Moles of H_2 = 3.96 moles	
5	Calculate the mass of O_2	Mass of O_2 = Moles of O_2 × Molar mass of O_2= 3.96 moles × 32.00 g/mol	126.72g



- Molar volume is 1 mole of **any gas** occupying a volume of 24dm^3 at standard room temperature and pressure (r.t.p)
- Unit is dm^3/mol or dm^3 mol^{-1}
    - $Mole = \frac{volume\ of\ gas}{24cm^3/mol}$
- Remember to convert cm^3 to dm^3 by ÷1000

Example: Calculate the volume of 2.5 moles of N_2

Answer:

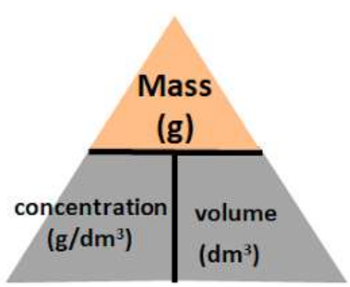
Volume = moles × Molar volume (dm^3/mol) = 2.5 × 24 = 60 dm^3

Example: Calculate the number of moles of

(a) 0.12dm^3 of CO_2 and (b) 200cm^3 of NH_3

Answer:

(a) Number of moles = Volume / molar volume = 0.12/24 = 0.0050mol
(b) Number of moles = Volume / molar volume = (200/1000)/24 = 0.00833mol



- Concentration is the amount of a solute in a given volume of solvent
- Depends on amount of solute (in mols or grams) or volume of solvent (in dm^3)
    - $conc.(mol/dm^3) = \frac{no.moles\ of\ solute\ (mol)}{volume\ of\ solvent\ (dm^3)}$
    - $conc.(g/dm^3) = \frac{mass\ of\ solute\ (g)}{volume\ of\ solvent\ (dm^3)}$

Example: A solution contains 0.5 moles of sodium chloride (NaCl) dissolved in 250 cm^3 of water. Calculate the concentration of NaCl in the solution in mol/dm^3

Answer:

Volume of solvent = 250 cm^3 = 250/1000 dm^3 = 0.25 dm^3

Concentration (mol/dm^3) = Number of moles of solute / Volume of solvent = 0.5 moles / 0.25 dm^3 = 2 mol/dm^3

Example: You have 20 grams of glucose (C6H12O6) dissolved in 500 cm^3 of water. Calculate the concentration of glucose in the solution in g/dm^3

Answer:

Volume of solvent = 500 cm^3 = 500/1000 dm^3 = 0.5 dm^3

Concentration (g/dm^3) = Mass of solute (g) / Volume of solvent (dm^3)

Concentration = 20 g / 0.5 dm^3 = 40 g/dm^3

STOICHIOMETRY

- Stoichiometry demonstrates the relationship between quantities of substances in a chemical reaction

Example:

The decomposition of sodium hydrogen carbonate by heating is:

$2\ NaHCO_3\ (s) \rightarrow Na_2CO_3\ (s) + H_2O\ (g)$

If 42g of sodium hydrogen carbonate were decomposed, determine the mass of residue left after the reaction.

Answer:

To find residue: find mass of NaHCO3 then find the difference between the initial mass of NaHCO3 and the produced Na2CO3 mass.

Step	Description	Calculation	Result
1	Calculate the molar mass of $NaHCO_3$	Molar mass of $NaHCO_3$ = 23 + 1 + 12 + (3 × 16)	84
2	Calculate moles of $NaHCO_3$	Moles of $NaHCO_3$ = Mass (g) / Molar mass (g/mol)	0.5 moles
3	Use stoichiometry to find moles of Na_2CO_3	Moles of Na_2CO_3 = (1/2) × Moles of $NaHCO_3$	0.25 moles
4	Calculate mass of residue (Na_2CO_3)	Mass of residue (g) = Moles of Na_2CO_3 × Molar mass (g/mol)	15.5 g

Example: Methane is used to produce hydrogen by reacting with steam:

$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$

Calculate the volume of hydrogen obtained from $100cm^3$ of methane at r.t.p

Answer:

If related substances involving are gases, ratio deduced from balanced chemical equation can be used directly to determine ratio of their volumes

$100cm^3$ of methane given and to determine hydrogen gas

	Methane	Hydrogen
Mole ratio	1	3
Volume ratio	$100cm^3$	$300cm^3$

A volume of $300cm^3$ of hydrogen gas is obtained from $100cm^3$ of methane

- Limiting reagent is the reagent that limits the amount of product formed
    - Will react completely in a chemical reaction
    - Act as reference for mole ratio during chemical calculations
- Excess reagent is the reagent which is left unreacted after the reaction

Example: When 4.0 moles of hydrogen gas (H_2) react with 6.0 moles of oxygen gas (O_2), what is the limiting reagent and the mass of water (H_2O) produced?

Answer:

Step	Description	Calculation	Result
1	Write the balanced chemical equation	$2 H_2 + O_2 \rightarrow 2 H_2O$	
2	Calculate the moles of H_2 and O_2 available	Moles of H_2 = 4.0 moles Moles of O_2 = 6.0 moles	
3	Use stoichiometry to find moles of H_2O from H_2 and O_2	Moles of H_2O from H_2 = 4.0 moles × (2 moles H_2O / 2 moles H_2) = 4.0 moles Moles of H_2O from O_2 = 6.0 moles × (2 moles H_2O / 1 mole O_2) = 12.0 moles	
4	Determine the limiting reagent	The limiting reagent is the one that produces fewer moles of H_2O	H_2 is limiting
5	Calculate the mass of H_2O produced from the limiting reagent	Mass of H_2O from H_2 = 4.0 moles × (18.02 g/mol)	72.08 g (H_2O)

Example: In the reaction: $2H_2 + O_2 \rightarrow 2H_2O$, if you have 10.0 grams of hydrogen gas (H_2) and 16.0 grams of oxygen gas (O_2), determine the limiting reactant and the mass of water (H_2O) formed.

Answer:

Step	Description	Calculation	Result
1	Write the balanced chemical equation	$2H_2 + O_2 \rightarrow 2H_2O$	
2	Calculate the molar mass	Molar mass of H_2 = 2.02 g/mol Molar mass of O_2 = 32.00 g/mol Molar mass of H_2O = 18.02 g/mol	
3	Calculate moles of H_2 and O_2	Moles of H_2 = Mass (g) / Molar mass (g/mol) = 10.0 g / 2.02 g/mol Moles of O_2 = 16.0 g / 32.00 g/mol	4.95 moles 0.5 moles
4	Determine the limiting reagent	Calculate moles of H_2O produced from each reactant. Moles of H_2O from H_2 = 4.95 moles × (2 moles H_2O / 2 moles H_2) Moles of H_2O from O_2 = 0.5 moles × (2 moles H_2O / 1 mole O_2)	4.95 moles 1 mole
5	Identify the limiting reactant	The limiting reactant is the one that produces fewer moles of H_2O.	O_2 is the limiting reagent
6	Calculate the mass of H_2O produced from the limiting reagent	Mass of H_2O from O_2 = Moles × Molar mass of H_2O = 1 mole × 18g/mol	18g

PERCENTAGE YIELD

- Percentage yield is how much product is actually formed out of a theoretical maximum
- Many reactions usually do not reach completion due to factors such as impurities
- $percentage\ yield\ = \frac{actual\ yield}{theoretical\ yield} \times 100\%$
  - Actual yield is the amount of product obtained from experiment
  - Theoretical yield is the amount of product from stoichiometric calculation

Example: In a chemical reaction, 20.0 g of sodium hydroxide (NaOH) is reacted with excess hydrochloric acid (HCl) to produce sodium chloride (NaCl) and water (H2O). If the actual yield of NaCl obtained is 17.5 g, calculate the percentage yield.

Answer:

Step	Description	Calculation	Result
1	Calculate the molar mass of NaOH	Molar mass of NaOH = 23.00 g/mol + 16.00 g/mol + 1.01 g/mol = 40.01 g/mol	40.01 g/mol
2	Calculate moles of NaOH	Moles of NaOH = Mass (g) / Molar mass (g/mol)	20.0 g / 40.01 g/mol ≈ 0.499 moles
3	Use stoichiometry to find moles of NaCl and H2O	Balanced equation: NaOH + HCl → NaCl + H2O Moles of NaCl = Moles of NaOH	1 mole of NaOH produces 1 mole of NaCl 0.499 moles
4	Calculate the theoretical yield of NaCl	Theoretical yield (g) = × Molar mass (g/mol)	0.499 moles × (58.44 g/mol) = 29.20 g (NaCl)
5	Calculate the percentage yield	Percentage yield = (Actual yield/ Theoretical yield)	(17.5 g / 29.20 g) × 100% ≈ 60.05%

		×100%	

PERCENTAGE PURITY

- Percentage purity is a measure used to express the purity of a substance within a sample or compound
- $percentage\ purity\ =\ percentage\ yield\ \times 100\%$

Example: When zinc granules are added to copper (II) chloride solution, copper metal and zinc chloride are formed. 200 g of impure zinc granules were added to an excess of copper (II) chloride solution. 171 g of copper metal was obtained. Find the percentage purity of the zinc granules.

Answer:

Step	Description	Calculation	Result
1	Calculate the theoretical mass of copper	Molar mass of copper (Cu) = 63.55 g/mol Moles of copper = Mass (g) / Molar mass (g/mol) = 171 g / 63.55 g/mol	2.691 moles
2	Calculate the moles of zinc that reacted.	Balanced equation: Zn + CuCl2 → Cu + ZnCl2 Moles of copper = Moles of zinc	2.691 moles
3	Calculate the theoretical mass of pure zinc.	Molar mass of zinc (Zn) = 65.38 g/mol Theoretical mass of pure zinc = Moles of zinc × Molar mass of zinc = (2.691 moles × 65.38 g/mol)	175.938 g

4	Calculate the percentage purity of the zinc granules.	Percentage purity = (Actual mass of pure zinc / Theoretical mass of pure zinc) × 100	88.0%

Summary

Key Points	Definition	Example
Relative Atomic Mass (Ar)	Ar is the average mass of one atom of an element compared to 1/12 of the mass of a carbon-12 atom.	Ar = Sum of (% abundance × isotopic mass)
Relative Molecular Mass (Mr)	Mr is the average mass of one molecule compared to 1/12 of the mass of a carbon-12 atom.	Mr of compound = Total Ar of all atoms
Empirical and Molecular Formulae	Empirical formula is the simplest ratio of atoms in a compound; Molecular formula shows the actual number of atoms in a molecule.	
The Mole	A mole represents a quantity of 6.02×10^{23} particles (Avogadro's number).	1 mole = 6.02×10^{23} particles
Mole Formulas	$Mole = \frac{no.of\ particles}{6 \times 10^{23}}$ $Mole = \frac{volume\ of\ gas\ (dm^3)}{24cm^3/mol}$	

	$Mole = \frac{mass\ of\ substance\ (g)}{M_r}$ $conc.(mol/dm^3) = \frac{no.moles\ of\ solute\ (mol)}{volume\ of\ solvent\ (dm^3)}$ $conc.(g/dm^3) = \frac{mass\ of\ solute\ (g)}{volume\ of\ solvent\ (dm^3)}$ $1dm^3=1000cm^3$	
Limiting reagent	Limiting reagent is the reagent that limits the amount of product formed	
Percentage Yield and Purity	Percentage yield measures how much product is actually obtained, while percentage purity measures the purity of a substance.	% Yield = (Actual yield / Theoretical yield) × 100%

What are some common pitfalls?

1. Molar Mass Calculation:

Pitfall: Incorrectly calculating molar mass by summing atomic masses without considering subscripts.

Elaboration: Ensure that you understand how to calculate molar mass by multiplying the atomic mass of each element by its subscript in the chemical formula.

Example: Calculate the molar mass of calcium carbonate ($CaCO_3$).

Incorrect Calculation: Molar Mass = Ca (40.08 g/mol) + C (12.01 g/mol) + O (16.00 g/mol) = 68.09 g/mol

Correct Calculation: Molar Mass = Ca (1 × 40.08 g/mol) + C (1 × 12.01 g/mol) + O (**3 ×** 16.00 g/mol) = 100.09 g/mol

2. Balanced Equations:

Pitfall: Using unbalanced chemical equations for stoichiometric calculations.

Elaboration: Remember to balance chemical equations before performing any stoichiometric calculations. Make it a regular practice

Example: Consider the reaction: $N_2 + H_2 \rightarrow NH_3$

Unbalanced Equation: $N_2 + H_2 \rightarrow NH_3$

Correct Balanced Equation: $N_2 + \underline{\mathbf{3}}H_2 \rightarrow \underline{\mathbf{2}}NH_3$

3. Stoichiometric Calculations:

Pitfall: Misinterpreting coefficients as mole ratios.

Elaboration: Remember that the coefficients in a balanced equation represent the mole ratios of reactants and products. Use these ratios correctly in stoichiometric calculations.

Example: In the balanced equation $2H_2 + O_2 \rightarrow 2H_2O$, find the number of moles of O_2 needed to react with 4 moles of H_2.

Incorrect Calculation: 4 moles of H_2 = 4 moles of O_2 (misinterpreting coefficients)

Correct Calculation: 4 moles of H_2 × (**<u>1 mole of O_2 / 2 moles of H_2</u>**) = 2 moles of O_2

4. Limiting Reactants:

Pitfall: Failing to identify the limiting reactant, leading to incorrect yield calculations.

Elaboration: Determine the limiting reactant by comparing the stoichiometric coefficients and the initial amounts of reactants. Limiting reactants determine the maximum amount of product.

Example: Consider the reaction of 3 moles of hydrogen (H_2) and 2 moles of oxygen (O_2). Determine the limiting reactant and the moles of water formed.

Incorrect Calculation: 3 moles of H_2 react with 2 moles of O_2, so O_2 is limiting, and 2 moles of water are formed.

Correct Calculation: To find the limiting reactant, calculate the moles of water formed by each reactant.

Moles of H_2 produce 3 moles of water, and moles of O_2 produce 1 mole of water. Since only 1 mole of water can be formed from O_2, O_2 is the limiting reactant.

5. Concentration Units:

Pitfall: Mixing up units when working with different concentration units (mol/dm^3, g/dm^3).

Elaboration: Clearly define the units and provide examples of converting between them. Use consistent units in calculations.

Example: Calculate the concentration of glucose ($C_6H_{12}O_6$) in a solution containing 20 g of glucose in 500 cm^3 of water.

Incorrect Calculation: Concentration = 20 g / 500 cm^3 = 0.04 g/dm^3

Correct Calculation: First, convert the volume to dm^3 (500 cm^3 = 0.5 dm^3), then calculate the concentration in g/dm^3: Concentration = 20 g / 0.5 dm^3 = 40 g/dm^3

Study tips

Studying mole concepts and stoichiometry can be challenging, but with the right strategies, students can master these topics effectively. Here are some specific studying tips:

1. Understand the Mole Concept Thoroughly:

Begin by grasping the concept of a mole as a unit of measurement for the quantity of substances. Understand how it relates to Avogadro's number and the number of particles in one mole.

2. Stoichiometry Practice Sets:

Build your stoichiometry skills through practice sets that include various types of stoichiometry problems, including mole-to-mole, mass-to-mole, and volume-to-mole conversions.

3. Work on Balancing Chemical Equations:

Master the skill of balancing chemical equations. This is crucial for stoichiometric calculations. Practice balancing equations until it becomes second nature.

4. Identify the Limiting Reactant:

Focus on recognizing the limiting reactant in a reaction. Practice problems that involve identifying the reactant that limits the amount of product formed.

5. Understand Units and Conversions:

Pay close attention to units in stoichiometry problems. Ensure consistency in units and practice converting between different units (e.g., grams to moles, moles to liters, dm^3 to cm^3).

6. Work on Stoichiometry Story Problems:

Practice solving stoichiometry "story problems" where you are given a scenario and asked to calculate reactant quantities, product yields, or other related quantities.

Remember that mastering mole concept and stoichiometry may take time and practice. Be patient with yourself and keep working on these topics regularly to build your confidence and proficiency.

Study hard and you can do it! ☺