Question 1. Calculate the molecular mass of the following:
(i) H_2O (ii) CO_2 (iii) CH_4

Answer:
(i) Molecular mass of H_2O = 2(1.008 amu) + 16.00 amu=18.016 amu
(ii) Molecular mass of CO_2= 12.01 amu + 2 x 16.00 amu = 44.01 amu
(iii) Molecular mass of CH_4= 12.01 amu + 4 (1.008 amu) = 16.042 amu

Question 2. Calculate the mass percent of different elements present in sodium sulphate ($Na_2\,SO_4$).
Answer:

$$\text{Mass \% of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$$

Now, $\text{Molar mass of } Na_2SO_4 = 2\,(23.0) + 32.0 + 4 \times 16.0$
$= 142 \text{ g mol}^{-1}$,

$$\text{Mass percent of sodium} = \frac{46}{142} \times 100 = 32.39\ \%$$

$$\text{Mass percent of sulphur} = \frac{32}{142} \times 100 = 22.54\ \%$$

$$\text{Mass percent of oxygen} = \frac{64}{142} \times 100 = 45.07\ \%$$

Question 3. Determine the empirical formula of an oxide of Iron which has 69.9 % iron and 30.1 % dioxygen by mass.
Answer:

Element	*Symbol*	*% by mass*	*Atomic mass*	*Moles of the element (Relative no. of moles)*	*Simplest molar ratio*	*Simplest whole number molar ratio*
Iron	Fe	69.9	55.85	$\frac{69.9}{55.85} = 1.25$	$\frac{1.25}{1.25} = 1$	2
Oxygen	O	30.1	16.00	$\frac{30.1}{16.00} = 1.88$	$\frac{1.88}{1.25} = 1.5$	3

$\therefore$ Empirical formula = Fe_2O_3.

Question 4. Calculate the amount of carbon dioxide that could be produced when
(i) 1 mole of carbon is burnt in air.
(ii) 1 mole of carbon is burnt in 16 g of dioxygen.
(iii) 2 moles of carbon are burnt in 16 g of dioxygen.
Answer: The balanced equation for the combustion of carbon in dioxygen/air is

$$\underset{\text{1 mole}}{C\,(s)} + \underset{\substack{\text{1 mole}\\(32\text{ g})}}{O_2\,(g)} \longrightarrow \underset{\substack{\text{1 mole}\\(44\text{ g})}}{CO_2\,(g)}$$

(i) In air, combustion is complete. Therefore,CO_2 produced from the combustion of 1 mole of carbon = 44 g.
(ii) As only 16 g of dioxygen is available, it can combine only with 0.5 mole of carbon, i.e., dioxygen is the limiting reactant.

Hence, CO_2 produced = 22 g.
(iii) Here again, dioxygen is the limiting reactant. 16 g of dioxygen can combine only with 0.5 mole of carbon. CO_2 produced again is equal to 22 g.

Question 5. Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 g mol^{-1}

Answer: 0.375 M aqueous solution means that 1000 mL of the solution contain sodium acetate = 0.375 mole

$\therefore$ 500 mL of the solution should contain sodium acetate = $\frac{0.375}{2}$ mole

Molar mass of sodium acetate = 82.0245 g mol^{-1}

$\therefore$ Mass of sodium acetate required = $\frac{0.375}{2}$ mole, $\times$ 82.0245 g mol^{-1} = **15.380 g.**

Question 6. Calculate the concentration of nitric acid in moles per litre in a sample which has a density 1.41 g mL^{-1} and the mass percent of nitric acid in it is being 69%.
Answer: Mass percent of 69% means that 100 g of nitric acid solution contain 69 g of nitric acid by mass.
Molar mass of nitric acid HNO_3= 1 + 14 + 48 = 63 g mol^{-1}

$$\therefore \text{ Moles in 69 g } HNO_3 = \frac{69 \text{ g}}{63 \text{ g mol}^{-1}} = 1.095 \text{ mole}$$

$$\text{Volume of 100 g nitric acid solution} = \frac{100 \text{ g}}{1.41 \text{ g mL}^{-1}} = 70.92 \text{ mL} = 0.07092 \text{ L}$$

$$\therefore \text{ Conc. of } HNO_3 \text{ in moles per litre} = \frac{1.095 \text{ mole}}{0.07092 \text{ L}} = \mathbf{15.44 \text{ M.}}$$

Question 7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)? (Atomic mass of Cu= 63.5 amu)
Answer: 1 mole of $CuSO_4$ contains 1 mole (1 g atom) of Cu
Molar mass of $CuSO_4$= 63.5 + 32 + 4 x 16 = 159.5 g mol^{-1}
Thus, Cu that can be obtained from 159.5 g of $CuSO_4$ = 63.5 g

$$\therefore \text{ Cu that can be obtained from 100 g of } CuSO_4 = \frac{63.5}{159.5} \times 100 \text{ g} = \mathbf{39.81 \text{ g.}}$$

Question 8. Determine the molecular formula of an oxide of iron in which the mass percent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.8 g mol^{-1}(Atomic mass: Fe = 55.85, O = 16.00 amu)Calculation of Empirical Formula. See Q3.
Answer: Empirical formula mass of Fe_2O_3 = 2 x 55.85 + 3 x 16.00 = 159.7 g mol^{-1}

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.8}{159.7} = 1$$

Hence, molecular formula is same as empirical formula, viz., Fe_2O_3.

Question 9. Calculate the atomic mass (average) of chlorine using the following data:

	% Natural Abundance	Molar Mass
^{35}Cl	75.77	34.9689
^{37}Cl	24.23	36.9659

Answer:

Fractional abundance of ^{35}Cl = 0.7577, Molar mass = 34.9689
Fractional abundance of ^{37}Cl = 0.2423, Molar mass = 36.9659
$\therefore$ Average atomic mass = (0.7577) (34.9689 amu) + (0.2423) (36.9659 amu)
= 26.4959 + 8.9568 = **35.4527**

Question 10. In three moles of ethane (C_2H_6), calculate the following:
(i) Number of moles of carbon atoms
(ii) Number of moles of hydrogen atoms
(iii) Number of molecules of ethane
Answer:
(i) 1 mole of C_2H_6 contains 2 moles of carbon atoms
3 moles of C_2H_6 will C-atoms = 6 moles
(ii) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms
3 moles of C_2H_6 will contain H-atoms = 18 moles
(*iii*) 1 mole of C_2H_6 contains Avogadro's no., *i.e.*, 6.02×10^{23} molecules
$\therefore$ 3 moles of C_2H_6 will contain ethane molecules = $3 \times 6.02 \times 10^{23}$
= **18.06×10^{23} molecules**

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