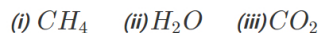


Q1. Calculate the molar mass of the following:

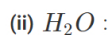


Ans.



Molecular weight of methane, CH_4

$$\begin{aligned} &= (1 \times \text{Atomic weight of carbon}) + (4 \times \text{Atomic weight of hydrogen}) \\ &= [1(12.011 \text{ u}) + 4(1.008 \text{ u})] \\ &= 12.011 \text{ u} + 4.032 \text{ u} \\ &= 16.043 \text{ u} \end{aligned}$$



Molecular weight of water, H_2O

$$\begin{aligned} &= (2 \times \text{Atomic weight of hydrogen}) + (1 \times \text{Atomic weight of oxygen}) \\ &= [2(1.0084) + 1(16.00 \text{ u})] \\ &= 2.016 \text{ u} + 16.00 \text{ u} \\ &= 18.016 \text{ u} \\ &\text{So approximately} \\ &= 18.02 \text{ u} \end{aligned}$$



= Molecular weight of carbon dioxide, CO_2

$$\begin{aligned} &= (1 \times \text{Atomic weight of carbon}) + (2 \times \text{Atomic weight of oxygen}) \\ &= [1(12.011 \text{ u}) + 2(16.00 \text{ u})] \\ &= 12.011 \text{ u} + 32.00 \text{ u} \\ &= 44.011 \text{ u} \\ &\text{So approximately} \\ &= 44.01 \text{ u} \end{aligned}$$

Q2. Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

Ans.

Now for Na_2SO_4 .

Molar mass of Na_2SO_4

$$\begin{aligned} &= [(2 \times 23.0) + (32.066) + 4(16.00)] \\ &= 142.066 \text{ g} \end{aligned}$$

$$\text{Formula to calculate mass percent of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$$

Therefore, Mass percent of the sodium element:

$$= \frac{46.0g}{142.066g} \times 100$$

$$= 32.379$$

$$= 32.4\%$$

Mass percent of the sulphur element:

$$= \frac{32.066g}{142.066g} \times 100$$

$$= 22.57$$

$$= 22.6\%$$

Mass percent of the oxygen element:

$$= \frac{64.0g}{142.066g} \times 100$$

$$= 45.049$$

$$= 45.05\%$$

Q3. Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.

Ans.

Percent of Fe by mass = 69.9 % [As given above]

Percent of O₂ by mass = 30.1 % [As given above]

Relative moles of Fe in iron oxide:

$$= \frac{\text{percent of iron by mass}}{\text{Atomic mass of iron}}$$

$$= \frac{69.9}{55.85}$$

$$= 1.25$$

Relative moles of O in iron oxide:

$$= \frac{\text{percent of oxygen by mass}}{\text{Atomic mass of oxygen}}$$

$$= \frac{30.1}{16.00}$$

$$= 1.88$$

Simplest molar ratio of Fe to O:

$$= 1.25: 1.88$$

$$= 1: 1.5$$

$$\approx 2: 3$$

Therefore, empirical formula of iron oxide is Fe_2O_3 .

Q4. Calculate the amount of carbon dioxide that could be produced when

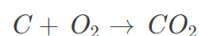
(i) 1 mole of carbon is burnt in air.

(ii) 1 mole of carbon is burnt in 16 g of dioxygen.

(iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Ans.

(i) 1 mole of carbon is burnt in air.



1 mole of carbon reacts with 1 mole of O_2 to form one mole of CO_2 .

Amount of CO_2 produced = 44 g

(ii) 1 mole of carbon is burnt in 16 g of O_2 .

1 mole of carbon burnt in 32 grams of O_2 it forms 44 grams of CO_2 .

Therefore, 16 grams of O_2 will form $\frac{44 \times 16}{32}$

= 22 grams of CO_2

(iii) 2 moles of carbon are burnt in 16 g of O_2 .

Since oxygen is the limiting reactant here, the 16g (0.5 mol) of O_2 will react with 6g of carbon (0.5 mol) to form 22 g of carbon dioxide. The remaining 18g of carbon (1.5 mol) will not undergo combustion.

Q5. 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 $82.0245 \text{ g mol}^{-1}$.

.

Ans.

0.375 M aqueous solution of CH_3COONa

= 1000 mL of solution containing 0.375 moles of CH_3COONa

Therefore, no. of moles of CH_3COONa in 500 mL

$$= \frac{0.375}{1000} \times 1000$$

$$= 0.1875 \text{ mole}$$

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

Therefore, mass that is required of CH_3COONa

$$= (82.0245 \text{ g mol}^{-1})(0.1875 \text{ mole})$$

$$= 15.38 \text{ gram}$$

Q6. 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 per cent of nitric acid in it being 69%

Ans.

Mass percent of HNO_3 in sample is 69 %

Thus, 100 g of HNO_3 contains 69 g of HNO_3 by mass.

Molar mass of HNO_3

$$= \{1 + 14 + 3(16)\} \text{ g mol}^{-1}$$

$$= 1 + 14 + 48$$

$$= 63 \text{ g mol}^{-1}$$

Now, No. of moles in 69 g of HNO_3 :

$$= \frac{69 \text{ g}}{63 \text{ g mol}^{-1}}$$

$$= 1.095 \text{ mol}$$

Volume of 100g HNO_3 solution

$$= \frac{\text{Mass of solution}}{\text{density of solution}}$$

$$= \frac{100 \text{ g}}{1.41 \text{ g mL}^{-1}}$$

$$= 70.92 \text{ mL}$$

$$= 70.92 \times 10^{-3} \text{ L}$$

Concentration of HNO_3

$$= \frac{1.095 \text{ mole}}{70.92 \times 10^{-3} \text{ L}}$$

$$= 15.44 \text{ mol/L}$$

Therefore, Concentration of HNO_3 = 15.44 mol/L

Q7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

Ans.

1 mole of $CuSO_4$ contains 1 mole of Cu.

Molar mass of $CuSO_4$

$$= (63.5) + (32.00) + 4(16.00)$$

$$= 63.5 + 32.00 + 64.00$$

$$= 159.5 \text{ gram}$$

159.5 gram of $CuSO_4$ contains 63.5 gram of Cu.

Therefore, 100 gram of $CuSO_4$ will contain $\frac{63.5 \times 100 \text{ g}}{159.5}$ of Cu.

$$= \frac{63.5 \times 100}{159.5}$$

$$= 39.81 \text{ gram}$$

Q8. Determine the molecular formula of an oxide of iron, in which the mass per cent of iron and oxygen are 69.9 and 30.1, respectively.

Ans.

Here,

Mass percent of Fe = 69.9%

Mass percent of O = 30.1%

No. of moles of Fe present in oxide

$$= \frac{69.90}{55.85}$$

$$= 1.25$$

No. of moles of O present in oxide

$$= \frac{30.1}{16.0}$$

$$= 1.88$$

Ratio of Fe to O in oxide,

$$= 1.25 : 1.88$$

$$= \frac{1.25}{1.25} : \frac{1.88}{1.25}$$

$$= 1 : 1.5$$

$$= 2 : 3$$

Therefore, the empirical formula of oxide is Fe_2O_3

Empirical formula mass of Fe_2O_3

$$= [2(55.85) + 3(16.00)] \text{ gr}$$

$$= 159.69 \text{ g}$$

$$\text{Therefore } n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.69 \text{ g}}{159.7 \text{ g}}$$

$$= 0.999$$

$$= 1 (\text{approx})$$

The molecular formula of a compound can be obtained by multiplying n and the empirical formula.

Thus, the empirical of the given oxide is Fe_2O_3 and n is 1.

Q9. Calculate the atomic mass (average) of chlorine using the following data:

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

Ans.

Average atomic mass of Cl.

$$= [(\text{Fractional abundance of } ^{35}\text{Cl})(\text{molar mass of } ^{35}\text{Cl}) + (\text{fractional abundance of } ^{37}\text{Cl})(\text{Molar mass of } ^{37}\text{Cl})]$$

$$= \left[\left(\frac{75.77}{100} (34.9689 \text{ u}) \right) + \left(\frac{24.23}{100} (36.9659 \text{ u}) \right) \right]$$

$$= 26.4959 + 8.9568$$

$$= 35.4527 \text{ u}$$

Therefore, the average atomic mass of Cl = 35.4527 u

Q10. In three moles of ethane (C₂H₆), calculate the following:

(i) Number of moles of carbon atoms.

(ii) Number of moles of hydrogen atom

(iii) Number of molecules of ethane

Ans.

(a) 1 mole C₂H₆ contains two moles of C- atoms.

∴ No. of moles of C- atoms in 3 moles of C₂H₆.

$$= 2 * 3$$

$$= 6$$

(b) 1 mole C₂H₆ contains six moles of H- atoms.

∴ No. of moles of C- atoms in 3 moles of C₂H₆.

$$= 3 * 6$$

$$= 18$$

(c) 1 mole C₂H₆ contains six moles of H- atoms.

∴ No. of molecules in 3 moles of C₂H₆.

$$= 3 * 6.023 * 10^{23}$$

$$= 18.069 * 10^{23}$$

Q11. 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 2L?

Ans.

Molarity (M) is as given by,

$$= \frac{\text{Number of moles of solute}}{\text{Volume of solution in Litres}}$$

$$= \frac{\text{Mass of sugar}}{\text{Molar mass of sugar} \times 2 L}$$

$$= \frac{20 g}{\frac{[(12 \times 12) + (1 \times 22) + (11 \times 16)]g}{2 L}}$$

$$= \frac{20 g}{342 g}$$

$$= \frac{0.0585 mol}{2 L}$$

$$= 0.02925 \text{ mol } L^{-1}$$

Therefore, Molar concentration = 0.02925 mol L⁻¹

Q12. If the density of methanol is 0.793 kg L⁻¹, what is its volume needed for making 2.5 L of its 0.25 M solution?

Ans.)

Molar mass of CH_3OH

$$= (1 * 12) + (4 * 1) + (1 * 16)$$

$$= 32 \text{ g mol}^{-1}$$

$$= 0.032 \text{ kg mol}^{-1}$$

Molarity of the solution

$$= \frac{0.793 \text{ kg L}^{-1}}{0.032 \text{ kg mol}^{-1}}$$

$$= 24.78 \text{ mol L}^{-1}$$

(From the definition of density)

$$M_1 V_1 = M_2 V_2 \therefore (24.78 \text{ mol L}^{-1}) V_1 = (2.5 \text{ L}) (0.25 \text{ mol L}^{-1})$$

$$V_1 = 0.0252 \text{ Litre}$$

$$V_1 = 25.22 \text{ Millilitre}$$

Q13. Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:

$$1 \text{ Pa} = 1 \text{ N m}^{-2}$$

If mass of air at sea level is 1034 g cm^{-2} , calculate the pressure in pascal

Ans.

As per definition, pressure is force per unit area of the surface.

$$P = \frac{F}{A}$$

$$= \frac{1034 \text{ g} \times 9.8 \text{ ms}^{-2}}{\text{cm}^2} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{(100)^2 \text{ cm}^2}{1 \text{ m}^2}$$

$$= 1.01332 \times 10^5 \text{ kg m}^{-1} \text{ s}^{-2}$$

Now,

$$1 \text{ N} = 1 \text{ kg m s}^{-2}$$

Then,

$$1 \text{ Pa} = 1 \text{ Nm}^{-2}$$

$$= 1 \text{ kgm}^{-2} \text{ s}^{-2}$$

$$\text{Pa} = 1 \text{ kgm}^{-1} \text{ s}^{-2} \therefore \text{Pressure (P)} = 1.01332 \times 10^5 \text{ Pa}$$

Q14. What is the SI unit of mass? How is it defined?

Ans.

Si Unit: Kilogram (kg)

Mass:

"The mass equal to the mass of the international prototype of kilogram is known as mass."

Q15. Match the following prefixes with their multiples:

	Prefixes	Multiples
(a)	femto	10
(b)	giga	10^{-15}
(c)	mega	10^{-6}
(d)	deca	10^9
(e)	micro	10^6

Ans.

	Prefixes	Multiples
(a)	femto	10^{-15}
(b)	giga	10^9
(c)	mega	10^6
(d)	deca	10
(e)	micro	10^{-6}

Q16. What do you mean by significant figures?

Ans.

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.

Therefore, "the total number of digits in a number with the Last digit the shows the uncertainty of the result is known as significant figures."

Q17. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

(i) Express this in per cent by mass.

(ii) Determine the molality of chloroform in the water sample.

Ans.

(a) 1 ppm = 1 part out of 1 million parts.

Mass percent of 15 ppm chloroform in H_2O

$$= \frac{15}{10^6} \times 100$$

$$= \approx 1.5 \times 10^{-3} \%$$

(b) 100 gram of the sample is having 1.5×10^{-3} g of $CHCl_3$.

1000 gram of the sample is having 1.5×10^{-2} g of $CHCl_3$.

\therefore Molality of $CHCl_3$ in water

$$= \frac{1.5 \times 10^{-2} \text{ g}}{\text{Molar mass of } CHCl_3}$$

Molar mass ($CHCl_3$)

$$= 12 + 1 + 3 (35.5)$$

$$= 119.5 \text{ gram mol}^{-1}$$

Therefore, molality of $CHCl_3$ in water

$$= 1.25 \times 10^{-4} \text{ m}$$

Q18. Express the following in the scientific notation:

(i) 0.0048

(ii) 234,000

(iii) 8008

(iv) 500.0

(v) 6.0012

Ans.

(a) $0.0048 = 4.8 \times 10^{-3}$

(b) $234,000 = 2.34 \times 10^5$

(c) $8008 = 8.008 \times 10^3$

(d) $500.0 = 5.000 \times 10^2$

(e) $6.0012 = 6.0012$

Q19. How many significant figures are present in the following?

(a) 0.0027

(b) 209

(c) 6005

(d) 136,000

(e) 900.0

(f) 2.0035

Ans.

(i) 0.0027: 2 significant numbers.

(ii) 209: 3 significant numbers.

(iii) 6005: 4 significant numbers.

(iv) 136,000: 3 significant numbers.

(v) 900.0: 4 significant numbers.

(vi) 2.0035: 5 significant numbers.

Q20. Round up the following upto three significant figures:

- (a) 35.217
- (b) 11.4108
- (c) 0.05577
- (d) 2806

Ans.

- (a) The number after round up is: 35.2
- (b) The number after round up is: 11.4
- (c) The number after round up is: 0.0560
- (d) The number after round up is: 2810

Q21. 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

(a) 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 = dm³

Ans.

(1) If we fix the mass of N₂ at 28 g, the masses of N₂ that will combine with the fixed mass of N₂ are 32 gram, 64 gram, 32 gram and 80 gram.

The mass of O₂ bear whole no. ratio of 1: 2: 2: 5. Therefore, the given information obeys the law of multiple proportions.

The law of multiple proportions states, "If 2 elements combine to form more than 1 compound, then the masses of one element that combines with the fixed mass of another element are in the ratio of small whole numbers."

(2) Convert:

(a) 1 km = ___ mm = ___ pm

$$\bullet 1 \text{ km} = 1 \text{ km} * \frac{1000 \text{ m}}{1 \text{ km}} * \frac{100 \text{ cm}}{1 \text{ m}} * \frac{10 \text{ mm}}{1 \text{ cm}}$$

$$\therefore 1 \text{ km} = 10^6 \text{ mm}$$

$$\bullet 1 \text{ km} = 1 \text{ km} * \frac{1000 \text{ m}}{1 \text{ km}} * \frac{1 \text{ pm}}{10^{-12} \text{ m}}$$

$$\therefore 1 \text{ km} = 10^{15} \text{ pm}$$

Therefore, 1 km = 10⁶ mm = 10¹⁵ pm

(b) 1 mg = ___ kg = ___ ng

$$\bullet 1 \text{ mg} = 1 \text{ mg} * \frac{1 \text{ g}}{1000 \text{ mg}} * \frac{1 \text{ kg}}{1000 \text{ g}}$$

$$1 \text{ mg} = 10^{-6} \text{ kg}$$

$$\bullet 1 \text{ mg} = 1 \text{ mg} * \frac{1 \text{ g}}{1000 \text{ mg}} * \frac{1 \text{ mg}}{10^{-3} \text{ g}}$$

$$1 \text{ mg} = 10^6 \text{ ng}$$

$$\text{Therefore, } 1 \text{ mg} = 10^{-6} \text{ kg} = 10^6 \text{ ng}$$

$$(c) 1 \text{ mL} = \text{---} \text{ L} = \text{---} \text{ dm}^3$$

$$\bullet 1 \text{ mL} = 1 \text{ mL} * \frac{1 \text{ L}}{1000 \text{ mL}}$$

$$1 \text{ mL} = 10^{-3} \text{ L}$$

$$\bullet 1 \text{ mL} = 1 \text{ cm}^3 = 1 * \frac{1 \text{ dm} \times 1 \text{ dm} \times 1 \text{ dm}}{10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm}} \text{ cm}^3$$

$$1 \text{ mL} = 10^{-3} \text{ dm}^3$$

$$\text{Therefore, } 1 \text{ mL} = 10^{-3} \text{ L} = 10^{-3} \text{ dm}^3$$

Q22. If the speed of light is $3.0 \times 10^8 \text{ m s}^{-1}$, calculate the distance covered by light in 2.00 ns

Ans.

Time taken = 2 ns

$$= 2 \times 10^{-9} \text{ s}$$

Now,

$$\text{Speed of light} = 3 \times 10^8 \text{ m s}^{-1}$$

So,

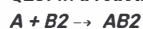
Distance travelled in 2 ns = speed of light * time taken

$$= (3 \times 10^8)(2 \times 10^{-9})$$

$$= 6 \times 10^{-1} \text{ m}$$

$$= 0.6 \text{ m}$$

Q23. In a reaction



Identify the limiting reagent, if any, in the following reaction mixtures.

(a) 2 mol X + 3 mol Y

(b) 100 atoms of X + 100 molecules of Y

(c) 300 atoms of X + 200 molecules of Y

(d) 2.5 mol X + 5 mol Y

(e) 5 mol X + 2.5 mol Y

Ans.

Limiting reagent:

It determines the extent of a reaction. It is the first to get consumed during a reaction, thus causes the reaction to stop and limiting the amt. of products formed.

(a) 2 mol X + 3 mol Y

1 mole of X reacts with 1 mole of Y. Similarly, 2 moles of X reacts with 2 moles of Y, so 1 mole of Y is unused. Hence, X is limiting agent.

(b) 100 atoms of X + 100 molecules of Y

1 atom of X reacts with 1 molecule of Y. Similarly, 100 atoms of X reacts with 100 molecules of Y. Hence, it is a stoichiometric mixture where there is no limiting agent.

(c) 300 atoms of X + 200 molecules of Y

1 atom of X reacts with 1 molecule of Y. Similarly, 200 atoms of X reacts with 200 molecules of Y, so 100 atoms of X are unused. Hence, Y is limiting agent.

(d) 2.5 mol X + 5 mol Y

1 mole of X reacts with 1 mole of Y. Similarly, 2.5 moles of X reacts with 2.5 moles of Y, so 2.5 mole of Y is unused. Hence, X is limiting agent.

(e) 5 mol X + 2.5 mol Y

1 mole of X reacts with 1 mole of Y. Similarly 2.5 moles of X reacts with 2.5 moles of Y, so 2.5 mole of X is unused. Hence, Y is limiting agent.

Q24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation: $N_2(g) + H_2(g) \rightarrow 2NH_3(g)$

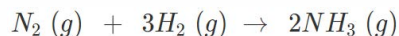
(a) What is the mass of NH_3 produced if 2×10^3 g N_2 reacts with 1×10^3 g of H_2 ?

(b) Will the reactants N_2 or H_2 remain unreacted?

(c) If any, then which one and give its mass.

Ans.

(a) Balance the given equation:



Thus, 1 mole (28 g) of N_2 reacts with 3 mole (6 g) of H_2 to give 2 mole (34 g) of NH_3 .

2×10^3 g of N_2 will react with $\frac{6}{28} \times 2 \times 10^3$ g NH_3

2×10^3 g of N_2 will react with 428.6 g of H_2 .

Given:

Amt of $H_2 = 1 \times 10^3$

28 g of N_2 produces 34 g of NH_3

Therefore, mass of NH_3 produced by 2000 g of N_2

$$= \frac{34 \text{ g}}{28 \text{ g}} \times 2000 \text{ g}$$

$$= 2430 \text{ g of } NH_3$$

(b) H_2 is the excess reagent. Therefore, H_2 will not react.

(c) Mass of H_2 unreacted

$$= 1 \times 10^3 - 428.6 \text{ g}$$

$$= 571.4 \text{ g}$$

Q25. How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Ans.

Molar mass of Na_2CO_3 :

$$= (2 \times 23) + 12 + (3 \times 16)$$

$$= 106 \text{ g mol}^{-1}$$

1 mole of Na_2CO_3 means 106 g of Na_2CO_3

Therefore, 0.5 mol of Na_2CO_3

$$= \frac{106 \text{ g}}{1 \text{ mol}} \times 0.5 \text{ mol } Na_2CO_3$$

$$= 53 \text{ g of } Na_2CO_3$$

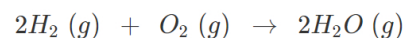
$$0.5 \text{ M of } Na_2CO_3 = 0.5 \text{ mol/L } Na_2CO_3$$

Hence, 0.5 mol of Na_2CO_3 is in 1 L of water or 53 g of Na_2CO_3 is in 1 L of water.

Q26. If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Ans.

Reaction:



2 volumes of dihydrogen react with 1 volume of dioxygen to produce two volumes of vapour.

Hence, 10 volumes of dihydrogen will react with five volumes of dioxygen to produce 10 volumes of vapour.

Q27. Convert the following into basic units:

(i) 29.7 pm

(ii) 16.15 pm

(iii) 25366 mg

Ans.

(i) 29.7 pm

$$1 \text{ pm} = 10^{-12} \text{ m}$$

$$29.7 \text{ pm} = 29.7 \times 10^{-12} \text{ m}$$

$$= 2.97 \times 10^{-11} \text{ m}$$

(ii) 16.15 pm

$$1 \text{ pm} = 10^{-12} \text{ m}$$

$$16.15 \text{ pm} = 16.15 \times 10^{-12} \text{ m}$$

$$= 1.615 \times 10^{-11} \text{ m}$$

(iii) 25366 mg

$$1 \text{ mg} = 10^{-3} \text{ g}$$

$$25366 \text{ mg} = 2.5366 \times 10^{-1} \times 10^{-3} \text{ kg}$$

$$25366 \text{ mg} = 2.5366 \times 10^{-2} \text{ kg}$$

Q28. Which one of the following will have the largest number of atoms?

(i) 1 g Au (s)

(ii) 1 g Na (s)

(iii) 1 g Li (s)

(iv) 1 g of Cl_2 (g)

Ans.

(i) 1 g Au (s)

$$= \frac{1}{197} \text{ mol of Au (s)}$$

$$= \frac{6.022 \times 10^{23}}{197} \text{ atoms of Au (s)}$$

$$= 3.06 \times 10^{21} \text{ atoms of Au (s)}$$

(ii) 1 g Na (s)

$$= \frac{1}{23} \text{ mol of Na (s)}$$

$$= \frac{6.022 \times 10^{23}}{23} \text{ atoms of Na (s)}$$

$$= 0.262 \times 10^{23} \text{ atoms of Na (s)}$$

$$= 26.2 \times 10^{21} \text{ atoms of Na (s)}$$

(iii) 1 g Li (s)

$$= \frac{1}{7} \text{ mol of Li (s)}$$

$$= \frac{6.022 \times 10^{23}}{7} \text{ atoms of Li (s)}$$

$$= 0.86 \times 10^{23} \text{ atoms of Li (s)}$$

$$= 86.0 \times 10^{21} \text{ atoms of Li (s)}$$

(iv) 1 g of Cl_2 (g)

$$= \frac{1}{71} \text{ mol of } Cl_2 \text{ (g)}$$

(Molar mass of Cl_2 molecule = $35.5 \times 2 = 71 \text{ g mol}^{-1}$)

$$= \frac{6.022 \times 10^{23}}{71} \text{ atoms of } Cl_2 \text{ (g)}$$

$$= 0.0848 \times 10^{23} \text{ atoms of } Cl_2 \text{ (g)}$$

$$= 8.48 \times 10^{21} \text{ atoms of } Cl_2 \text{ (g)}$$

Therefore, 1 g of Li (s) will have the largest no. of atoms.

Q29. Calculate the molarity of a solution of ethanol in water, in which the mole fraction of ethanol is 0.040 (assume the density of water to be one).

Ans.

Mole fraction of C_2H_5OH

$$= \frac{\text{Number of moles of } C_2H_5OH}{\text{Number of moles of solution}}$$

$$0.040 = \frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + n_{H_2O}} \quad \text{---(1)}$$

No. of moles present in 1 L water:

$$n_{H_2O} = \frac{1000 \text{ g}}{18 \text{ g mol}^{-1}} \quad n_{H_2O} = 55.55 \text{ mol}$$

Substituting the value of n_{H_2O} in eq (1),

$$\frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + 55.55} = 0.040$$

$$n_{C_2H_5OH} = 0.040 n_{C_2H_5OH} + (0.040)(55.55)$$

$$0.96 n_{C_2H_5OH} = 2.222 \text{ mol}$$

$$n_{C_2H_5OH} = \frac{2.222}{0.96} \text{ mol} \quad n_{C_2H_5OH} = 2.314 \text{ mol}$$

Therefore, molarity of solution

$$= \frac{2.314 \text{ mol}}{1 \text{ L}}$$

$$= 2.314 \text{ M}$$

Q30. What will be the mass of one ^{12}C atom in g?

Ans.

1 mole of carbon atoms

$$= 6.023 \times 10^{23} \text{ atoms of carbon}$$

$$= 12 \text{ g of carbon}$$

Therefore, mass of 1 ^{12}C atom

$$= \frac{12 \text{ g}}{6.022 \times 10^{23}}$$

$$= 1.993 \times 10^{-23} \text{ g}$$

Q31. How many significant figures should be present in the answer of the following calculations?

(i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$

(ii) 5×5.365

(iii) $0.012 + 0.7864 + 0.0215$

Ans.

(i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$

Least precise no. of calculation = 0.112

Therefore, no. of significant numbers in the answer

= No. of significant numbers in the least precise no.

= 3

(ii) 5×5.365

Least precise no. of calculation = 5.365

Therefore, no. of significant numbers in the answer

= No. of significant numbers in 5.365

= 4

(iii) $0.012 + 0.7864 + 0.0215$

As the least no. of decimal place in each term is 4, the no. of significant numbers in the answer is also 4.

Q32. Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	Molar mass	Abundance
^{36}Ar	$35.96755 \text{ g mol}^{-1}$	0.337 %
^{38}Ar	$37.96272 \text{ g mol}^{-1}$	0.063 %
^{40}Ar	$39.9624 \text{ g mol}^{-1}$	99.600 %

Ans.

Molar mass of Argon:

$$= \left[(35.96755 \times \frac{0.337}{100}) + (37.96272 \times \frac{0.063}{100}) + (39.9624 \times \frac{99.600}{100}) \right]$$

$$= [0.121 + 0.024 + 39.802] \text{ g mol}^{-1}$$

$$= 39.947 \text{ g mol}^{-1}$$

Q33. Calculate the number of atoms in each of the following

(i) 52 moles of Ar

(ii) 52 u of He

(iii) 52 g of He

Ans.

(i) 52 moles of Ar

$$1 \text{ mole of Ar} = 6.023 \times 10^{23} \text{ atoms of Ar}$$

$$\text{Therefore, } 52 \text{ mol of Ar} = 52 \times 6.023 \times 10^{23} \text{ atoms of Ar}$$

$$= 3.131 \times 10^{25} \text{ atoms of Ar}$$

(ii) 52 u of He

1 atom of He = 4 u of He

OR

4 u of He = 1 atom of He

1 u of He = $\frac{1}{4}$ atom of He

52 u of He = $\frac{52}{4}$ atom of He

= 13 atoms of He

(iii) 52 g of He

4 g of He = 6.023×10^{23} atoms of He

52 g of He = $\frac{6.023 \times 10^{23} \times 52}{4}$ atoms of He

= 7.8286×10^{24} atoms of He

Q34. A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Find:

(i) Empirical formula

(ii) Molar mass of the gas, and

(iii) Molecular formula

Ans.

(i) Empirical formula

1 mole of CO_2 contains 12 g of carbon

Therefore, 3.38 g of CO_2 will contain carbon

$$= \frac{12 \text{ g}}{44 \text{ g}} \times 3.38 \text{ g}$$

$$= 0.9217 \text{ g}$$

18 g of water contains 2 g of hydrogen

Therefore, 0.690 g of water will contain hydrogen

$$= \frac{2 \text{ g}}{18 \text{ g}} \times 0.690$$

$$= 0.0767 \text{ g}$$

As hydrogen and carbon are the only elements of the compound. Now, the total mass is:

$$= 0.9217 \text{ g} + 0.0767 \text{ g}$$

$$= 0.9984 \text{ g}$$

Therefore, % of C in the compound

$$= \frac{0.9217 \text{ g}}{0.9984 \text{ g}} \times 100$$

$$= 92.32 \%$$

% of H in the compound

$$= \frac{0.0767 \text{ g}}{0.9984 \text{ g}} \times 100$$

$$= 7.68 \%$$

Moles of C in the compound,

$$= \frac{92.32}{12.00}$$

$$= 7.69$$

Moles of H in the compound,

$$= \frac{7.68}{1}$$

$$= 7.68$$

Therefore, the ratio of carbon to hydrogen is,

7.69: 7.68

1: 1

Therefore, the empirical formula is CH.

(ii) Molar mass of the gas, and

Weight of 10 L of gas at STP = 11.6 g

Therefore, weight of 22.4 L of gas at STP

$$= \frac{11.6 \text{ g}}{10 \text{ L}} \times 22.4 \text{ L}$$

$$= 25.984 \text{ g}$$

$$\approx 26 \text{ g}$$

(iii) Molecular formula

Empirical formula mass:

$$\text{CH} = 12 + 1$$

$$= 13 \text{ g}$$

$$n = \frac{\text{Molar mass of gas}}{\text{Empirical formula mass of gas}}$$

$$= \frac{26 \text{ g}}{13 \text{ g}}$$

$$= 2$$

Therefore, molecular formula is $(\text{CH})_n$ that is C_2H_2 .

Q35. Calcium carbonate reacts with aqueous HCl to give CaCl₂ and CO₂ according to the reaction, $\text{CaCO}_3 (\text{s}) + 2 \text{HCl} (\text{aq}) \rightarrow \text{CaCl}_2 (\text{aq}) + \text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{l})$

What mass of CaCO₃ is required to react completely with 25 mL of 0.75 M HCl?

Ans.

0.75 M of HCl

≡ 0.75 mol of HCl are present in 1 L of water

≡ [(0.75 mol) × (36.5 g mol⁻¹)] HCl is present in 1 L of water

≡ 27.375 g of HCl is present in 1 L of water

Thus, 1000 mL of solution contains 27.375 g of HCl

Therefore, amt of HCl present in 25 mL of solution

$$= \frac{27.375 \text{ g}}{1000 \text{ mL}} \times 25 \text{ mL}$$

$$= 0.6844 \text{ g}$$

Given chemical reaction,



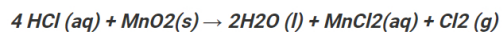
2 mol of HCl (2 × 36.5 = 73 g) react with 1 mol of CaCO_3 (100 g)

Therefore, amt of CaCO_3 that will react with 0.6844 g

$$= \frac{100}{73} \times 0.6844 \text{ g}$$

$$= 0.9375 \text{ g}$$

Q36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction:



How many grams of HCl react with 5.0 g of manganese dioxide?

Ans.

$$1 \text{ mol of } \text{MnO}_2 = 55 + 2 \times 16 = 87 \text{ g}$$

$$4 \text{ mol of HCl} = 4 \times 36.5 = 146 \text{ g}$$

1 mol of MnO_2 reacts with 4 mol of HCl

5 g of MnO_2 will react with:

$$= \frac{146 \text{ g}}{87 \text{ g}} \times 5 \text{ g HCl}$$

$$= 8.4 \text{ g HCl}$$

Therefore, 8.4 g of HCl will react with 5 g of MnO_2 .