CHAPTER

1

Some Basic Concepts of Chemistry

CLASS 11TH

NCERT EXERCISE AND SOLUTIONS - CHEMISTRY

Q. 1. **Calculate the molar mass of the following:**

(i) H_2O **(ii)** CO_2 **(iii)** CH_4

ANSWER:-

(i) Molar mass of the water, H_2O

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

= [2(1.0084) + 1(16.00u)]

= 2.016u + 16.00u

= 18.016

= 18.02u

(ii) The molecular mass of carbon dioxide, CO_2

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

= [1(12.011u) + 2(16.00u)]

= 12.011u + 32.00u

= 44.01u

(iii) The molecular mass of methane, CH_4

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

= [1(12.011u) + 4(1.008u)]

= 12.011u + 4.032u

= 16.043u

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

ANSWER:-

The molecular formula of sodium sulphate is Na_2SO_4

Molar mass of Na_2SO_4 = [(2 × 23.0) + (32.066) + 4(16.00)] = 142.066 g

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

Mass percent of sodium :-

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

$= 32.379$

$= 32.4\%$

Q. 3. **Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.**

ANSWER:-

% of iron by mass = 69.9% [Given]

of oxygen by mass = 30.1% [Given]

Relative moles of iron in iron oxide:

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

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

$= 1.25$

Relative moles of oxygen in iron oxide

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

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

$= 1.88$

Simplest molar ratio of iron to oxygen:

$= 1.25:1.88$

$= 1:1.5$

$= 2:3$

The empirical formula of the iron oxide is $Fe_2 O_3$

Q. 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:-

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

$C + O_2 \rightarrow CO_2$

1 mole of C produces 1 mole of which corresponds to 44 g of CO_2.

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

16 g of dioxygen corresponds to $\frac{16}{32} = 0.5$ moles.

Here, dioxygen is the limiting reagent.

It will produce 0.5 moles of which corresponds to $0.5 \times 44 = 22$ g of CO_2

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

16 g of dioxygen corresponds to $\frac{16}{32} = 0.5$ moles.

Here, dioxygen is the limiting reagent.

It will produce 0.5 moles of CO_2 which corresponds to $0.5 \times 44 = 22$ g of CO_2

Q. 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 solution of CH_3COONa = 0.375 moles of CH_3COONa/1000ml

No. of moles of CH_3COONa in $500ml = \frac{0.375}{1000} \times 500 = 0.1875$ moles

Molar mass of sodium acetate = 82.0245 g/mole

Mass of CH_3COONa required = $82.0245 \times 0.1875 = 15.38$

Q. 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 per cent of nitric acid in it being 69%.**

ANSWER:-

Mass percent of nitric acid in the sample = 69% [Given]

Thus, 100 g of nitric acid contains 69 g of nitric acid by mass.

molar mass of nitric acid (HCO_3)

$= \{1 + 14 + 3(16)\} gmol^{-1}$

$= 1 + 14 + 48$

$= 63\ gmol^{-1}$

Number of moles in 69 g of HNO_3

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

$= 1.095$ mol

Volume of 100 g of nitric acid solution

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

$= \frac{100g}{1.41gm\ L^{-1}}$

$= 70.92mL$

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

Concentration of nitric acid

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

= 15.44 mol/L

∴ Concentration of nitric acid = 15.4 mol/L

Q. 7. How much copper can be obtained from 100 g of copper sulphate ($CuSO_4$)?

ANSWER:-

1 mole of $CuSO_4$ contain 1 mole (1 g atom) of Cu

Molar mass of $CuSO_4$ = 63.5 + 32 + 4 × 16

= 159.5g mol^{-1}

Thus, Cu that can be obtained from 159.5 g of $CuSO_4$ = 63.5g

∴ Cu that can be obtained from 100 g of $CuSO_4$

$$= \frac{63.5}{159.5} \times 100\text{g} = 39.81\text{g}$$

Q. 8. 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.

ANSWER:-

The iron oxide has iron and 30.1% dioxygen by mass.

Thus, 100 g of iron oxide contains 69.9 g iron and 30.1 g dioxygen. The number of moles of iron present in 100 g of iron oxide are $\frac{69.9}{55.8} = 1.25$ The number of moles of dioxygen present in 100 g of iron oxide are $\frac{30.1}{32} = 0.94$

The ratio of the number of oxygen atoms to the number of carbon atoms present in one formula unit of iron oxide is $\frac{2 \times 0.94}{1.25} = 1.5 : 1 = 3 : 2$. Hence, the formula of the iron oxide is Fe_2O_3.

Q. 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:-

Average atomic mass of chlorine

$$= \frac{75.77 \times 34.9689 + 24.23 \times 36.9659}{100} = 35.45\text{g / mol}$$

Q. 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:-

We are given 3 moles of ethane.

(i) Number of moles of carbon atoms = 3 × 2 = 6 mol

(ii) Number of moles of hydrogen atoms = 3 × 6 = 18 mol

(iii) Number of molecules of ethane = $3 \times 6.023 \times 10^{23} = 1.8069 \times 10^{24}$

Q. 11. **What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L?**

ANSWER:-

"Molarity (M) of a solution is given by,"

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

$$= \frac{\text{Mass of sugar/molar mass of sugar}}{2L}$$

$$= \frac{20g / [(12\times12)+(1\times22)+(11\times16)]g}{2L}$$

$$= \frac{20g / 342g}{2L}$$

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

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

Molar concentration of sugar = 0.02925 $molL^{-1}$

Q. 12. **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?**

ANSWER:-

Molar mass of methanol (CH_3OH) = (1 × 12) + (4 × 1) + (1 × 16)

= 32 g mol^{-1}

= 0.032 kg mol^{-1}

$$\text{Molarity of methanol solution} = \frac{0.793kgL^{-1}}{0.032kgmol^{-1}}$$

= 24.78 mol L^{-1}

(Since density is mass per unit volume)

Applying,

$M_1 V_1 = M_2 V_2$

(Given solution) (Solution to be prepared)

$(24.78 \text{ mol } L_{-1}) V_1 = (2.5 \text{ L})(0.25 \text{ mol}^{-1})$

$V_1 = 0.0252 \text{ L}$

$V_1 = 25.22 \text{ mL}$

Q. 13. Pressure is determined as force per unit area of the surface. The SI unit of pressure, Pascal is as shown below:-

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

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

ANSWER:-

Pressure is defined as force acting per unit area of the surface

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

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

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

We know,

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

Then,

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

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

$\therefore$ Pressure $= 1.01332 \times 10^5$ Pa.

Q .14. What is the SI unit of mass? How is it defined?

ANSWER:-

The SI unit of mass is kilogram Kilogram is defined as the mass equal to the mass of the international prototype of kilogram.

Q. 15. Match the following prefixes with their multiples:

	Prefixes	Multiples
(i)	micro	10^6
(ii)	deca	10^9
(iii)	mega	10^{-6}
(iv)	giga	10^{-15}
(v)	femto	10

ANSWER:-

	Prefixes	Multiples
(i)	micro	10^{-6}
(ii)	deca	10
(iii)	mega	10^{6}
(iv)	giga	10^{9}
(v)	femto	10^{-15}

Q. 16. What do you mean by significant figures?

ANSWER:-

Significant figures:

Significant figures in the measured value of a physical quantity is the sum of reliable digits and the first uncertain digit.

OR

The number of digits in a measurement about which we are certain, plus one additional digit, the first one about which we are not certain is known as significant figures or significant digits.

Q. 17. 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.

ANSWER:-

Contamination level: 15 ppm (parts per million)

This means 15 parts of chloroform in 1,000,000 parts of water by mass.

Percent by mass is calculated as parts per 100.

We know 1 ppm = 1 part per 1,000,000 parts.

Therefore, 15 ppm = 15 parts per 1000,000 parts.

Calculate percent by mass:

$$\text{Percent by mass} = \left(\frac{15 \text{ parts}}{1,000,000 \text{ parts}}\right) \times 100$$

$$\text{Percent by mass} = \left(\frac{15}{1,000,000}\right) \times 100$$

"Percent by mass = $15 \times 10^{-6} \times 100$

"Percent by mass = 15×10^{-4} %

So, the contamination level in percent by mass is 1.5×10^{-3}%.

Given mass of chloroform: 15 grams

Molar mass of chloroform: 119.5 g/mol

Number of moles of chloroform:

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{15\text{g}}{119.5\text{g/mol}} = 0.25 \text{ moles}$$

$$\text{Molality} = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kg}}$$

$$= \frac{0.125 \text{ moles}}{1,000\text{kg}}$$

$$\text{Molality} = 0.125 \times 10^{-3} \text{ mol/kg} = 1.25 \times 10^{-4} \text{ mol/kg}$$

So, the molality of chloroform in the water sample is 1.25×10^{-4} **mol**/kg.

Q. 18. **Express the following in the scientific notation:**

(i) **0.0048** **(ii)** **234,000**
(iii) **8008** **(iv)** **500.0**
(v) **6.0012**

ANSWER:-

(i) $0.0048 * 10^{-3}$ **(ii)** $234,000 = 2.34 * 10^{5}$
(iii) $8008 = 8.008 * 10^{3}$ **(iv)** $500.0 = 5.000 * 10^{2}$
(v) $6.0012 = 6.0012$

Q. 19. **How many significant figures are present in the following?**

(i) **0.0025** **(ii)** **208** **(iii)** **5005** **(iv)** **126,000**
(v) **500.0** **(vi)** **2.0034**

ANSWER:-

(i) 0.0025 contains 2 significant figures.
(ii) 208 contains 3 significant figures.
(iii) 5005 contains 4 significant figures.
(iv) 126,00 contains 3 significant figures.
(v) 500.0 contains 4 significant figures.
(vi) 2.0034 contains 5 significant figures.

Q. 20. **Round up the following upto three significant figures:**

(i) **34.216** **(ii)** **10.4107** **(iii)** **0.04597** **(iv)** **2808**

ANSWER:-

(i) The number 34.216 is rounded to three significant figures as 34.2 .
(ii) The number 10.4107 is rounded to three significant figures as 10.4 .
(iii) The number 0.04597 is rounded to three significant figures as 0.046 .
(iv) The number 2808 is rounded to three significant figures as 2810

Q. 21. **The following data are obtained when dinitrogen and dioxygen react together to form different compounds:**

	Mass of dinitrogen	Mass of dioxygen
(i)	14 g	16 g
(ii)	14 g	32 g
(iii)	28 g	32 g
(iv)	28 g	80 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) 1mg = kg = ng**

(iii) 1 mL = L = dm^3

ANSWER:-

(a) If we fix the mass of dinitrogen at 28 g , then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 32 g, 64g, 32g, and 80 g .

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. The law states that if two elements combine to form more than one compound, then the masses of one element that combines with the fixed mass of another element are in the ratio of small whole numbers.

(b) (i) $1\text{km} = 1\text{km} \times \frac{1000\text{m}}{1\text{km}} \times \frac{100\text{cm}}{1\text{m}} \times \frac{10\text{mm}}{1\text{cm}}$

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

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

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

(ii) $1\text{mg} = 1\text{mg} \times \frac{1\text{g}}{1000\text{mg}} \times \frac{1\text{kg}}{100\text{g}}$

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

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

(iii) $1\text{mL} = 1\text{mL} \times \frac{1\text{L}}{100\text{mI}}$

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

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

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

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

Q. 22. **If the speed of light is 3.0×10^8 m s^{-1}, calculate the distance covered by light in 2.00 ns.**

ANSWER:-

According to the question:

Time taken to cover the distance = 2.00ns

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

Speed of light = 3.0×10^8 ms^{-1}

Distance travelled by light in 2.00 ns

= Speed of light × Time taken

$= (3.0 \times 10^8\ ms^{-1})\ (2.00 \times 10^{-9}\ s)$

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

= 0.600m

Q. 23. **In a reaction $A + B_2 \rightarrow AB_2$ Identify the limiting reagent, if any, in the following reaction mixtures.**

(i) 300 atoms of A + 200 molecules of B

(ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B

(iv) 5 mol A + 2.5 mol B

(v) 2.5 mol A + 5 mol B

ANSWER:-

(i) 200 molecules of *B* will react with 200 atoms of *A* and 100 atoms of A will remain unreacted. Hence, B_2 is the limiting reagent.

(ii) 2 moles of *A* will react with 2 moles of B_2 and 1 mole of B_2 will remain unreacted. Thus, A is the limiting reagent.

(iii) 100 atoms of will react with 100 molecules of B_2. Thus, both reagents are present in stoichiometric amounts.

(iv) 2.5 mole of B_2 will react with 2.5 mole of and 2.5 mole of *A* will remain unreacted. Hence, B_2 is the limiting reagent.

(v) 2.5 moles of *A* will react with 2.5 moles of B_2 and 2.5 moles of B_2 will remain unreacted. Thus, is the limiting reagent.

Q. 24. **Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation: $N_2(g) + H_2(g)\ 2NH_3(g)$ (i) Calculate the mass of ammonia produced if 2.00×10^3 g dinitrogen reacts with 1.00×10^3 g of dihydrogen. (ii) Will any of the two reactants remain unreacted? (iii) If yes, which one and what would be its mass?**

ANSWER:-

1 mol of, i.e., 28 g reacts with 3 mole of H_2, i.e., 6 g of H_2.

$\therefore$ 2000g of N_2 will react with $H_2 = \frac{6}{28} \times 2000\text{g} = 42.85\text{g}$. Thus, N_2 is the limiting reagent while H_2 is the excess reagent.

2 mol of N_2, i.e., 28 g of N_2 produces NH_3 = 2 mol = 34g

Therefore, 2000 g will produces $NH_3 = \frac{34}{28} \times 2000g$

= 2428.57 g

b. H_2 will remain unreacted.

c. Mass left unreacted = 1000g − 42.85g = 571.4 g

Q. 25. How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

ANSWER:-

0.5 mol $NaNO_3$ means the quantity of weight of Na_2CO_3 present whereas 0.5 M Na_2CO_3 means 0.5 mol.Na_2CO_3 present in 1 L solution. it is a measure of quantity of Na_2CO_3 in the solution.

Q. 26. If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

ANSWER:-

The balanced chemical equation for the reaction is,

$2H_2 + O_2 \rightarrow 2H_2O$

The volume of a gas is directly proportional to the number of moles.

Thus, 2 volumes of dihydrogen react with 1 volume of oxygen to form 2 volumes of water.

Thus, 10 volumes of dihydrogen react with 5 volume of oxygen to form 10 volumes of water.

Q. 27. Convert the following into basic units:

(i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

ANSWER:-

SI unit of length is metre and that of mass is Kg . The conversions into SI unit are shown below:

(i) $28.7pm = 28.7pm \times \frac{1m}{10^{12}pm} = 2.87 \times 10^{-11}m$

(ii) $15.15pm = 15.15pm \times \frac{1m}{10^{12}pm} = 1.515 \times 10^{-11}m$

(iii) $25365mg = 25365mg \times \frac{1kg}{10^{6}mg} = 2.5365 \times 10^{-2}kg$

Q. 28. 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)

ANSWER:-

(i) 1 g of Au $(s) = \frac{1}{197}$ mol of Au (s) $= \frac{6.022 \times 10^{23}}{197}$ atoms of Au (s)

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

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

$= \frac{6.022\times10^{23}}{23}\text{ atoms of Na (s)}$

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

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

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

$= \frac{6.022\times10^{23}}{7}\text{ atoms of Li(s)}$

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

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

(iv) $1\text{g of Cl}_2(g) = \frac{1}{71}\text{mol of Cl}_2(g)$

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

$= \frac{6.022\times10^{23}}{71}\text{ molecules of Cl}_2(g)$

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

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

As one molecule of Cl_2 contains two atoms of Cl.

Number of atoms of Cl $= 2 \times 8.48 \times 10^{21} = 16.96 \times 10^{21}$ atoms of Cl

Hence, 1 g of Li will have the largest number of atoms.

Q. 29. 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)

ANSWER:-

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

Number of moles present in 1 L water:-

$$n_{H_2O} = \frac{1000g}{18gmol^{-1}}$$

$$n_{H_2O} = 55.55mol$$

Substituting the value of nH_2 in equation (1),

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

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

$0.96 n_{C_2H_5OH} = 2.222 mol$

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

$n_{C_2H_5OH} = 2.314 mol$

$\therefore$ Molarity of solution $= \frac{2.314 mol}{1L}$

$= 2.314 M$

Q. 30. **What will be the mass of one ^{12}C atom in g?**

ANSWER:-

1 mole of carbon atoms = 6.023×10^{23} atoms of carbon = 12g of carbon

$\therefore$ Mass of one ^{12}C atom $= \frac{12g}{6.022 \times 10^{23}}$

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

Q. 31. **How many significant figures should be present in the answer of the following calculations?**

(i) **0.02856 298.15 × 0.112 0.5785**

(ii) **5 × 5.364**

(iii) **0.0125 + 0.7864 + 0.0215**

ANSWER:-

(i) The answer of the calculation $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$ will contain 3 significant figures as 0.112 contains 3 significant figures which is the lowest number of significant figures.

(ii) The answer to the calculation 5 × 5.364 will contain 4 significant figures as 5.364 contains 4 significant figures which are the lowest number of significant figures. Here, the exact figure 5 is not considered.

(iii) The answer to the calculation 0.0125 + 0.7864 + 0.0215 will contain 4 significant figures as the result will be reported to 4 decimal places

Q. 32. **Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:-**

Isotope	Isotopic molar mass	Abundance
^{36}Ar	$35.96755 gmol^{-1}$	0.337%
^{38}Ar	$37.96272 gmol^{-1}$	0.063%
^{40}Ar	$39.9624 gmol^{-1}$	99.600%

ANSWER:-

Molar mass of argon

$$= \left[\left(35.96755 \times \frac{0.337}{100}\right) + \left(37.96272 \times \frac{0.063}{100}\right) + \left(39.9624 \times \frac{90.60}{100}\right)\right] \text{gmol}^{-1}$$

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

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

Q. 33. 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**

ANSWER:-

(i) 1 mole of Ar = 6.022×10^{23} atoms of Ar

52 mol of Ar = $52 \times 6.022 \times 10^{23}$ atoms of Ar

= 3.131×10^{25} atoms of Ar

(ii) Atomic mass of He = 4amu

Therefore, number of He atoms in $52\text{u} = \frac{52}{4} = 13$ atoms of He

(iii) Gram atomic mass of He = 4 g

Or 4 g of He contains = 6.022×10^{23} atoms

Therefore, 52 g of He contains $= \frac{6.022 \times 10^{23}}{4 \times .52} = 7.83 \times 10^{24}$ atom

Q. 34. 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. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

ANSWER:-

(i) 1 mole (44 g) of CO_2 contains 12 g of carbon.

$\therefore$ 3.338 g of CO_2 will contain carbon $= \frac{12\text{g}}{44\text{g}} \times 3.38\text{g} = 0.9217$ g

18 g of water contains 2 g of hydrogen.

$\therefore$ 0.690 g of water will contain hydrogen $= \frac{2g}{18g} \times 0.690 = 0.0767$ g

Since carbon and hydrogen are the only constituents of the compound, the total mass of the compoun

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

$= 0.9984\text{g}$

$\therefore$ Percent of C in the compound $\frac{0.9217\text{g}}{0.9984\text{g}} \times 100$

$= 92.32\%$

Percent of H in the compound $= \frac{0.0767g}{0.9984g} \times 100$

$= 7.68\%$

Moles of carbon in the compound $= \frac{92.32}{12.00}$

$= 7.69$

Moles of hydrogen in the compound $= \frac{7.68}{1}$

$= 7.68$

$\therefore$ Ratio of carbon to hydrogen in the compound = 7.69:7.68

= 1:1

Hence, the empirical formula of the gas is CH .

(ii) Given,

Weight of 10.0 L of the gas at S.T.P) = 11.6 g

$\therefore$ Weight of22.4L of gas at STP $= \frac{11.6g}{10.0L} \times 22.4L$

= 25.984 g

≈26 g

Hence, the molar mass of the gas is 26 g .

(iii) Empirical formula mass of CH = 12 + 1 = 13g

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

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

N = 21

Molecular formula of gas = $(CH)_n$

= C_2H_2

Q. 35. **5 Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction, $CaCO_3$ (s) + 2 HCl (aq) → $CaCl_2$ (aq) + CO_2(g) + H_2O(l) What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl?**

ANSWER:-

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 wates

≡ 27.375 g of HC1 is present in 1 L of water

Thus, 1000 mL of solution contains 27.375 g of HC1.

$\therefore$ Amount of HCl present in 25 mL of solution

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

$= 0.6844g$

From the given chemical equation,

$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

$2\text{mol of HCl}(2\times 36.5 = 73\text{g})$ react with $1\text{mol of } CaCO_3(100\text{g})$

$\therefore$ Amount of $CaCO_3$ that will react with $0.6844\text{g} = \frac{100}{73}\times 0.6844$

$= 0.9375\text{g}$

Q. 36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction 4 HCl (aq) + MnO_2(s) → $2H_2O$ (l) + $MnCl_2$(aq) + Cl_2 (g) How many grams of HCl react with 5.0 g of manganese dioxide?

ANSWER:-

1 mol [55 + 2 × 16 = 87g]MnO_2 reacts completely with 4 mol [4 × 36.5 = 146 g] of HCl.

$\therefore$ 5.0 g of MnO_2 will react with $= \frac{146\text{g}}{87\text{g}}\times 5.0\text{g of HCl}$

$= \frac{(146\text{ g})}{(87\text{ g})} \times 5.0\text{ g of HCl}$

= 8.4 g of HCl)

Hence, 8.4 g of HCl will react completely with 5.0 g of manganese dioxide.