Exercise:
Q 1. Calculate the molar mass of the following:
(i) CH_4 (ii) H_2O (iii) CO_2

Ans.

(i) CH_4:
Molecular mass of CH_4 = Atomic mass of C + 4 x Atomic mass of H
= 12 + 4 x 1
= 16 u

(ii) H_2O:
Molar mass of water H_2O
Atomic mass of H = 1
Atomic mass of O = 16
H_2O = 2xH+1xO
Molar mass of water = 2x1+16 = 18g/mol

(iii) CO_2:
Molecular mass of CO_2 = Atomic mass of C + 2 x Atomic mass of O
= 12 + 2 x 16
= 44 u

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

= [(2 x 23.0) + (32.066) + 4(16.00)]

= 142.066 g

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

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.

Given there is an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass:

Relative moles of iron in iron oxide:

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

$= \frac{69.9}{55.85}$

= 1.25

Relative moles of oxygen in iron oxide:

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

$= \frac{30.1}{16.00}$

= 1.88

The simplest molar ratio of iron to oxygen:

⇒ 1.25: 1.88 ⇒ 1: 1.5 ⇒ 2: 3

Therefore, the empirical formula of the 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.

$$C + O_2 \rightarrow CO_2$$

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.

Here again, dioxygen is the limiting reactant. 16g of dioxygen can combine only with 0.5mol of carbon. CO_2 produced again is equal to 22g.

***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 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

= $\frac{0.375}{1000} \times 500$

= 0.1875 mole

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

Therefore, the mass of CH_3COONa

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

= 15.38 grams

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)} $g.mol^{-1}$

= 1 + 14 + 48

= 63g mol^{-1}

Now, no. of moles in 69 g of HNO_3

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

= 1.095 mol

Volume of 100g HNO_3 solution

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

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

= 70.92mL

= 70.92 x 10^{-3} L

Concentration of HNO_3

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

= 15.44mol/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 grams

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

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

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

= 39.81 grams

Q8. 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.

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)] g

= 159. 7g

The molar mass of Fe_2O_3 = 159.69g

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

= 0.999

= 1(approx)

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

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

Therefore, the molecular formula of the oxide is Fe_2O_3

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.

Fractional Abundance of ^{35}Cl = 0.7577 and Molar mass = 34.9689

Fractional Abundance of ^{37}Cl = 0.2423 and Molar mass = 36.9659

Average Atomic mass = (0.7577 x 34.9689)amu + (0.2423 x 36.9659)

= 26.4959 + 8.9568 = 35.4527

Q10. In three moles of ethane (C_2H_6), calculate the following:
(i) Number of moles of carbon atoms.
(ii) Number of moles of hydrogen atom

(iii) Number of molecules of ethane

Ans.

(i) 1 mole of C_2H_6 contains two moles of C- atoms.

∴ No. of moles of C- atoms in 3 moles of C_2H_6.

= 2 x 3

= 6

(ii) 1 mole of C_2H_6 contains six moles of H- atoms.

∴ No. of moles of H- atoms in 3 moles of C_2H_6.

= 3 x 6

= 18

(iii) 1 mole of C_2H_6 contains 1 mole of ethane- atoms.

∴ No. of molecules in 3 moles of C_2H_6

= 3 x 6.023 x 10^{23}

= 18.069 x 10^{23}

Q11. 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?

Ans.

Molarity (M) is as given by,

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

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

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

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

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

= 0.02925 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 methanol (CH_3OH)

= 32 $gmol^{-1}$= 0.032 $kgmol^{-1}$

molarity of the given solution

$$= \frac{W_2 in kg}{M_{w_2} \times V_{(sol)} L} = \frac{d_{sol}(kgL^{-1})}{Mw_2(kg)}$$

$$= \frac{0.793 kgL^{-1}}{0.032 kgmol^{-1}} = 24.78M$$

$$Applying \underset{(Given solution)}{M_1 \times V_1} = \underset{(solution to be prepared)}{M_2 V_2}$$

24.78 x V_1 = 0.25 x 2.5 L

or V_1 = 0.02522L = 25.22mL

Q13. Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:
1Pa = 1N m^{-2}
If mass of air at sea level is 1034 g cm^{-2}, calculate the pressure in pascal

Ans.

Pressure is the force (i.e., weight) acting per unit area

But weight = mg

∴ Pressure = Weight per unit area

$$= \frac{1034g \times 9.8ms^{-2}}{cm^2}$$

$$= \frac{1034g \times 9.8ms^{-2}}{cm^2} \times \frac{1kg}{1000g} \times \frac{100cm \times 100cm}{1m \times 1m} \times \frac{1N}{kgms^{-2}} \times \frac{1Pa}{1Nm^{-2}}$$

= 1.01332 x 10^5 Pa

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

Ans.

The S.I unit of mass is kilogram (kg). A kilogram is equal to the mass of a platinum-iridium cylinder kept at the International Bureau of Weights and Measures at Service, France.

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 that 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.

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

Mass percent of 15 ppm chloroform in H_2O

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

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

(ii) $Molarity = \frac{15/119.5}{10^6 \times 10^{-3}} = 1.25 \times 10^{-4}$

Q18. Express the following in the scientific notation:
(i) 0.0048
(ii) 234,000
(iii) 8008
(iv) 500.0
(v) 6.0012

Ans.

(i) 0.0048= 4.8×10^{-3}

(ii) 234,000 = 2.34×10^5

(iii) 8008 = 8.008×10^3

(iv) 500.0 = 5.000×10^2

(v) 6.0012 = 6.0012×10^0

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

(a) 0.0025

(b) 208

(c) 5005

(d) 126,000

(e) 500.0

(f) 2.0034

Ans.

(a) 0.0025: 2 significant numbers.

(b) 208: 3 significant numbers.

(c) 5005: 4 significant numbers.

(d) 126,000:3 significant numbers.

(e) 500.0: 4 significant numbers.

(f) 2.0034: 5 significant numbers.

Q20. Round up the following upto three significant figures:

(a) 34.216

(b) 10.4107

(c)0.04597

(d)2808

Ans.

(a) The number after round up is: 34.2

(b) The number after round up is: 10.4

(c)The number after round up is: 0.0460

(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 = …………………. dm^3

Ans.

(a)

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.

(b)

i. $1km = 1km \times \frac{1000m}{1km} \times \frac{100cm}{1m} \times \frac{10mm}{1cm} = 10^6mm$

$1km = 1km \times \frac{1000m}{1km} \times \frac{1pm}{10^{-12}m} = 10^{15}pm$

ii. $1mg = 1mg \times \frac{1g}{1000mg} \times \frac{1kg}{1000g} = 10^{-6}kg$

$1mg = 1mg \times \frac{1g}{1000mg} \times \frac{1ng}{10^{-9}g} = 10^6ng$

iii. $1mL = 1mL \times \frac{1L}{1000mL} = 10^{-3}L$

1 mL = 1 cm^3

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

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

Ans.

Time taken = 2 ns

= 2 x 10^{-9} s

Now,

Speed of light = 3 × 10^8 ms^{-1}

We know that,

Distance = Speed x Time

So,

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

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

= 6×10^{-1} m

= 0.6 m

Q23. 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

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 limits the amount of product formed.

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

1 atom of A reacts with 1 molecule of B. Similarly, 200 atoms of A reacts with 200 molecules of B, so 100 atoms of A are unused. Hence, B is the limiting reagent.

(ii) 2 mol A + 3 mol B

1 mole of A reacts with 1 mole of B. Similarly, 2 moles of A reacts with 2 moles of B, so 1 mole of B is unused. Hence, A is the limiting reagent.

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

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

(iv) 5 mol A + 2.5 mol B

1 mole of A reacts with 1 mole of B. Similarly 2.5 moles of A reacts with 2.5 moles of B, so 2.5 moles of A is unused. Hence, B is the limiting reagent.

(v) 2.5 mol A + 5 mol B

1 mole of A reacts with 1 mole of B. Similarly, 2.5 moles of A reacts with 2.5 moles of B, so 2.5 moles of B is unused. Hence, A is the limiting reagent.

Q24. 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 NH_3 produced if 2 x 10^3 g N_2 reacts with 1 x 10^3 g of H_2?

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass.

Ans.

(i) 1 mol of N_2 i.e., 28 g reacts with 3 moles of H_2 i.e., 6 g of H_2

∴ 2000 g of N_2 will react with H_2 = $\frac{6}{28} \times 200g = 428.6g$

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

= 34 g

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

= 2428.57 g

(ii) H_2 will remain unreacted

(iii) Mass left unreacted = 1000g – 428.6g = 571.4g

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 × 23) + 12 + (3 × 16)
= 106 g mol^{-1}

1 mole of Na_2CO_3 means 106 g of Na_2CO_3
Therefore, 0.5 mol of Na_2CO_3

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

= 53 g of Na_2CO_3
0.5 M of Na_2CO_3 = 0.5 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:

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

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

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

Q27. Convert the following into basic units:

(i) 28.7 pm

(ii) 15.15 pm

(iii) 25365 mg

Ans.

(i) 28.7 pm

1 pm = 10^{-12} m

28.7 pm = 28.7 x 10^{-12} m

= 2.87 x 10^{-11} m

(ii) 15.15 pm

1 pm = 10^{-12} m

15.15 pm = 15.15 × 10^{-12} m

= 1.515 × 10^{-11} m

(iii) 25365 mg

1 mg = 10^{-3} g

1 mg = 10^{-6} kg

25365 mg = 25365 x 10^{-6} kg

25365 mg = 2.5365 × 10^{-2} 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 of Au (s)

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

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

= 3.06×10^{21} atoms of Au (s)

(ii) 1 g of Na (s)

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

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

= 0.262×10^{23} atoms of Na (s)

= 26.2×10^{21} atoms of Na (s)

(iii) 1 g of Li (s)

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

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

= 0.86×10^{23} atoms of Li (s)

= 86.0×10^{21} atoms of Li (s)

(iv)1 g of Cl_2 (g)

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

(Molar mass of Cl_2 molecule = 35.5 × 2 = 71 g mol^{-1})

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

= 0.0848 x 10^{23} atoms of Cl_2 (g)

= 8.48 x 10^{21} atoms of Cl_2 (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

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

0.040 = $\frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + n_{H_2O}}$ ––(1)

No. of moles present in 1 L water:

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

Substituting the value of nH_2O in eqn (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

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

= 2.314 M

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

Ans.

1 mole of carbon atoms

= 6.023 x 10^{23} atoms of carbon

= 12 g of carbon

Therefore, mass of 1 atom of ^{12}C

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

= 1.993 x 10^{-23} 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 × 5.364

(iii) 0.0125 + 0.7864 + 0.0215

Ans.

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

Least precise number = 0.112

Therefore, no. of significant numbers in the answer

= No. of significant numbers in 0.112

= 3

(ii) 5 × 5.364

Least precise number = 5.364

Therefore, no. of significant numbers in the answer

= No. of significant numbers in 5.364

= 4

(iii) 0.0125 + 0.7864 + 0.0215

As the least no. of decimal place in each term is 4. Hence, 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* $g\ mol^{-1}$	*0.337 %*
^{38}Ar	*37.96272* $g\ mol^{-1}$	*0.063 %*
^{40}Ar	*39.9624* $g\ mol^{-1}$	*99.600 %*

Ans.

Molar mass of Argon:

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

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

= 39.947 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 mole of Ar = 6.023 x 10^{23} atoms of Ar

Therefore, 52 moles of Ar = 52 x 6.023 x 10^{23} atoms of Ar

= 3.131 x 10^{25} 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 x 10^{23} atoms of He

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

= 7.829 x 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

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

= 0.9218 g

18 g of water contains 2 g of hydrogen

Therefore, 0.690 g of water will contain hydrogen

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

= 0.0767 g

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

= 0.9217 g + 0.0767 g

= 0.9984 g

Therefore, % of C in the compound

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

= 92.32 %

% of H in the compound

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

= 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

Weight of 10 L of gas at STP = 11.6 g

Therefore, weight of 22.4 L of gas at STP

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

= 25.984 g

$\approx$ 26 g

(iii) Molecular formula

Empirical formula mass:

CH = 12 + 1

= 13 g

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

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

= 2

Therefore, molecular formula = 2 x CH = C_2H_2.

Q35. 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?

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

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

= 0.6844 g

Given chemical reaction,

$$CaCO_3\ (s)\ +\ 2\ HCl\ (aq)\ \rightarrow\ CaCl_2\ (aq)\ +\ CO_2\ (g)\ +\ H_2O\ (l)$$

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

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

= 0.9375 g

Q36. 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?

Ans.

1 mole of MnO_2 = 55 + 2 × 16 = 87 g

4 mole of HCl = 4 × 36.5 = 146 g

1 mole of MnO_2 reacts with 4 mol of HCl

Hence,

5 g of MnO_2 will react with:

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

= 8.4 g HCl

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