Chapter: Some Basic Concept of Chemistry

Examples

1. Calculate the molecular mass of glucose $\left(C_6H_{12}O_6\right)$ molecule.

Answer: Here, we have given the molecular mass of glucose i.e. $\left(C_6H_{12}O_6\right)$

So, from that we get

$$
\begin{aligned}
&= 6(12.011\ u)+12(1.008\ u)+6(16.00\ u) \\
&= (72.066\ u)+(12.096\ u)+(96.00\ u) \\
&= 180.162\ u
\end{aligned}
$$

2. A compound contains 4.07% hydrogen, 24.27% carbon and 71.65% chlorine. Its molar mass is 98.96 g. What are its empirical and molecular formulas?

Answer: We have total 100 g of compound masses. So, it divides as $4.07\ g$ hydrogen, $24.27\ g$ carbon and $71.65\ g$ chlorine

As we know atomic mass of hydrogen, carbon, chlorine are $1.008g,\ 12.01g,\ 35.453g$ respectively.

Now, divide given masses by their atomic mass. We get:

$$\text{Moles of hydrogen} = \frac{4.07\ g}{1.008\ g} = 4.04$$

$$\text{Moles of carbon} = \frac{24.27\ g}{12.01\ g} = 2.021$$

$$\text{Moles of chlorine} = \frac{71.65\ g}{35.453\ g} = 2.021$$

As there is smallest value i.e. 2.021 So, it gives ratio of $2{:}1{:}1$ for $H{:}C{:}CL$ after dividing moles with smallest value

Thus, empirical formula is CH_2CL

Now, we have to find empirical formula mass for CH_2CL We get:

$$
\begin{aligned}
&= 12.01+(2\times 1.008)+35.453 \\
&= 49.48\ g
\end{aligned}
$$

$$\text{So, } \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96\ g}{49.48\ g} = 2 = n$$

Now, finally multiply empirical formula with $n = 2$ we get molecular formula.

Hence molecular formula $= C_2H_4Cl_2$

3. Calculate the amount of water g produced by the combustion of 16 g of methane.

Answer: Here, methane is CH_4

So, balanced equation for its combustion

$CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(g)$

$16\ g$ of methane implies to one mole

From this, one mole of $CH_4(g)$ gives two mole of $H_2O(g)$

Two mole of $H_2O(g) = 2\times(2+16) = 2\times18 = 36\ g$

One mole of $H_2O(g) = 18\ g\ H_2O$

$$\Rightarrow \frac{18\ gH_2O}{1\ molH_2O} = 1$$

Thus, $2\ mol\ H_2O \times \frac{18\ g\ H_2O}{1\ mol\ H_2O}$

$$= 2\times18\ g\ H_2O$$
$$= 36\ g\ H_2O$$

4. How many moles of methane are required to produce $22\ g\ CO_2(g)$ after combustion?

Answer: As chemical equation is $CH_4(g)+2O_2(g) \rightarrow CO_2(g)+2H_2O(g)$

We get $44\ gCO_2(g)$ from $16\ gCH_4(g)$

No. of moles of $CO_2(g) = 22\ g\ CO_2(g) \times \frac{1\ mol\ CO_2(g)}{44\ g\ CO_2(g)}$

$$= 0.5\ mol\ CO_2(g)$$

Thus, $0.5\ mol\ CO_2(g)$ we get this from $0.5\ mol\ CH_4(g)$ or $0.5\ mol$ of $CH_4(g)$ required to produce $22\ g\ CO_2(g)$

5. 50.0 g of $N_2(g)$ and 10.0 kg of $H_2(g)$ are mixed to produce $NH_3(g)$. Calculate the amount of $NH_3(g)$ formed. Identify the limiting reagent in the production of $NH_3(g)$ in this situation.

Answer: Balanced equation is $N_2(g)+3H_2(g) \rightleftharpoons 2NH_3(g)$

No. of moles of N_2 $= 50.0\text{ kg } N_2 \times \frac{1000\text{ g } N_2}{1\text{ kg } N_2} \times \frac{1\text{ mol } N_2}{28.0\text{ g } N_2} = 17.86\times10^2$ mol

No. of moles of H_2 $= 10.00\text{ kg}H_2 \times \frac{1000\text{ g}H_2}{1\text{ kg}H_2} \times \frac{1\text{ mol}H_2}{2.016\text{ g}H_2} = 4.96\times10^3$ mol

From above equations, $1\text{ mol } N_2(g)$ requires $3\text{ mol } H_2(g)$

Thus, 17.86×10^2 mol of N_2 requires moles of $H_2(g)$ i.e.

$= 17.86\times10^2\text{ mol } N_2 \times \frac{3\text{ mol}H_2(g)}{1\text{ mol } N_2(g)} = 5.36\times10^3\text{ mol } H_2$

Since we have only $4.96\times10^3\text{ mol } H_2$. Thus, dihydrogen is limiting reagent here.

Thus, $NH_3(g)$ formed from 4.96×10^3 mol

As $3\text{ mol } H_2(g)$ gives $2\text{ mol } NH_3(g)$

$4.96\times10^3\text{ mol } H_2(g) \times \frac{2\text{ mol } NH_3(g)}{3\text{ mol } H_2(g)}$

$= 3.30\times10^3\text{ mol } NH_3(g)$

After converting in grams it looks like

$1\text{ mol } NH_3(g) = 17.0\text{ g } NH_3(g)$

$3.30\times10^3\text{ mol } NH_3(g) \times \frac{17.0\text{ g } NH_3(g)}{1\text{ mol } NH_3(g)}$

$= 3.30\times10^3\times17\text{ g } NH_3(g)$

$= 56.1\times10^3\text{ g } NH_3$

$= 56.1\text{ kg } NH_3$

6. A solution is prepared by adding 2 g of a substance A to 18 g of water. Calculate the mass per cent of the solute

Answer: Mass percent of A $= \frac{\text{Mass of A}}{\text{Mass of solution}} \times 100$

$$= \frac{2g}{2g \text{ of A} + 18g \text{ of water}} \times 100$$

$$= \frac{2g}{20g} \times 100$$

$$= 10\%$$

7. Calculate the molarity of $NaOH$ in the solution prepared by dissolving its 4 g in enough water to form 250 ml of the solution.

Answer: Molarity $= \dfrac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$

$$= \frac{\dfrac{\text{Mass of NaOH}}{\text{Molar mass of NaOH}}}{0.250 \text{ L}}$$

$$= \frac{\dfrac{4 \text{ g}}{40 \text{ g}}}{0.250 \text{ L}}$$

$$= \frac{0.1 \text{ mol}}{0.250 \text{ L}}$$

$$= 0.4 \text{ mol}^{-1}$$

$$= 0.4 \text{ M}$$

8. The density of $3\,M$ solution of $NaCl$ is $1.25\,g\,ml^{-1}$. Calculate the molarity of the solution.

Answer: Here molarity denotes as M i.e. $M = 3 \text{ mol L}^{-1}$

So, mass of $NaCL$ in one litre solution $= 3 \times 58.5 = 175.5$ g

Then, mass of one litre solution $= 1000 \times 1.25 = 1250$ g

Thus, mass of water in this solution $= 1250 - 75.5 = 1074.5$ g

$$M = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}} = \frac{3 \text{ mol}}{1.0745 \text{ kg}} = 2.79 \text{ m}$$

9. A piece of metal is $3\,\text{inch}$ (represented by in) long. What is its length in cm.

Answer: As we know that $1 \text{ in} = 2.54 \text{ cm}$

So, $\dfrac{1 \text{ in}}{2.54 \text{ cm}} = 1 = \dfrac{2.54 \text{ cm}}{1 \text{in}}$

Hence, $\frac{1\text{ in}}{2.54\text{ cm}} = \frac{2.54\text{ cm}}{1\text{in}} = 1$

Accordingly, $3\text{ in} = 3\text{ in} \times \frac{2.54\text{ cm}}{1\text{ in}} = 3\times 2.54\text{ cm} = 7.62\text{ cm}$

10. A jug contains 2 L of milk. Calculate the volume of the milk in m^3.

Answer: As we know $1\text{ L} = 1000\text{ cm}^3$, $1\text{ m} = 100\text{ cm}$

We get $\frac{1\text{ m}}{100\text{ cm}} = 1 = \frac{100\text{ cm}}{1\text{ m}}$

So, for m^3 , we cubed this unit, we get:

$$\left(\frac{1\text{ m}}{100\text{ cm}}\right)^3 \Rightarrow \frac{1\text{ m}^3}{10^6\text{ cm}^3} = (1)^3 = 1$$

Thus, $2\text{ L} = 2\times 1000\text{ cm}^3$

Now multiply it with unit factor we get

$$2\times 1000\text{ cm}^3 \times \frac{1\text{ m}^3}{10^6\text{ cm}^3} = \frac{2\text{ m}^3}{10^3} = 2\times 10^{-3}\text{ m}^3$$

11. How many seconds are there in 2 days.

Answer: As we know that $1\text{ day} = 24\text{ hours (h)}$

$$\frac{1\text{ day}}{24\text{ h}} = 1 = \frac{24\text{ h}}{1\text{ day}}$$

And we also know that $1\text{ h} = 60\text{ min}$

$$\frac{1\text{ h}}{60\text{ min}} = 1 = \frac{60\text{ min}}{1\text{ h}}$$

Now, for calculating seconds in two days, we have

$$= 2\text{ day} \times \frac{24\text{ h}}{1\text{ day}} \times \frac{60\text{ min}}{1\text{ h}} \times \frac{60\text{ s}}{1\text{ min}}$$

$$= 2\times 24\times 60\times 60\text{ s}$$

$$= 172800\text{ s}$$

Exercise

1. Calculate the molecular mass of the following

(i) H_2O

(ii) CO_2

(iii) CH_4

Answer: (i) H_2O represents hydrogen. So, molecular mass of H_2O is

$= (2\times$ atomic mass of hydrogen) + $(1\times$ atomic mass of oxygen)

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

= 2.016 u+16.00 u

= 18.016

= 18.02 u

(ii) CO_2 represents Carbon dioxide. So, molecular mass of CO_2 is

$= (1\times$ atomic mass of carbon) + $(2\times$ atomic mass of oxygen)

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

= 12.011 u+32.00 u

= 44.01 u

(iii) CH_4 represents methane, So, molecular mass of CH_4 is

$= (1\times$ atomic mass of carbon) + $(4\times$ atomic mass of hydrogen)

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

= 12.011 u+4.032 u

= 16.043 u

2. Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Answer: As we know that molecular formula of sodium sulphate is (Na_2SO_4)

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

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

Thus, using this formula, we find mass percent of sodium, sulphur, oxygen

Mass percent of sodium $= \dfrac{46.0\text{ g}}{142.066\text{ g}} \times 100 = 32.379 = 32.4\%$

Mass percent of sulphur $= \dfrac{32.066\text{ g}}{142.066\text{ g}} \times 100 = 22.57 = 22.6\%$

Mass percent of oxygen $= \dfrac{64.0\text{ g}}{142.066\text{ g}} \times 100 = 45.049 = 45.05\%$

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

Answer: We have given that

Mass percent of iron $= 69.9\%$

Mass percent of oxygen $= 30.1\%$

So, relative mole of iron in iron oxide $= \dfrac{\%\text{ of iron by mass}}{\text{Atomic mass of iron}}$

$$= \frac{69.9}{55.85} = 1.25$$

And relative mole of oxygen in iron oxide $= \dfrac{\%\text{ of oxygen by mass}}{\text{Atomic mass of oxygen}}$

$$= \frac{30.1}{16.00} = 1.88$$

Thus, ratio $= 1.25 : 1.88 = 1 : 1.5 = 2 : 3$

Hence empirical formula of iron oxide $= Fe_2O_3$

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) According to balance equation, one mole of carbon burns in one mole of dioxygen for producing one mole of carbon dioxide

(ii) According to given condition, here only 16 g of dioxygen is available. Thus, it react with 0.5 mole of carbon to give 22 g of carbon dioxide. Hence, it shows limiting reactant.

(iii) According to given condition, here only $16\ g$ of dioxygen is available then it shows limiting reactant. Hence, $16\ g$ of dioxygen is combined with only 0.5 mole of carbon to give $22\ g$ of carbon dioxide

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: Here molar denotes as M

We have given $0.375\ M$ aqueous solution of sodium acetate which is equal to $1000\ ml$ of solution which contains 0375 moles of sodium acetate

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

Given that Molar mass $= 82.0245\ g\ mole^{-1}$

Thus, required mass $= \left(82.0245\ g\ mol^{-1}\right)(0.1875\ mole) = 15.38\ g$

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 being 69%.

Answer: We have given that mass percent of nitric acid $= 69\%$

According to mass, $100\ g$ of nitric acid contains $69\ g$ of nitric acid

So, molar mass $= \{1+14+3(16)\}g\ mol^{-1}$

$$= 1+14+48$$

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

No. of moles in $69\ g$ $= \frac{69\ g}{63\ g\ mol^{-1}} = 1.095\ mol$

Volume of $100\ g$ nitric acid solution $= \frac{\text{Mass of solution}}{\text{density of solution}}$

$$= \frac{100\ g}{1.41\ g\ mL^{-1}} = 70.92\ mL$$

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

Now, concentration $= \frac{1.095\ mole}{70.92 \times 10^{-3}\ L} = 15.44\ mol/L$

7. How much copper can be obtained from 100 g of copper sulphate $CuSO_4$.

Answer: Here, one mole of $CuSO_4$ contains one mole of copper

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

$= 63.5+32.00+64.00$
$= 159.5\text{ g}$

Therefore, 159.5 g of $CuSO_4$ contains 63.5 g of copper

This implies that 100 g of $CuSO_4$ contains $\frac{63.5\times100\text{ g}}{159.5}$ of copper

Thus, amount of copper from 100 g of $CuSO_4$ $= \frac{63.5\times100}{159.5} = 39.81$

8. Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9, 30.1 respectively. Given that the molar mass of the oxide is 159.69 g mol^{-1}.

Answer: We have given that mass percent of iron Fe and oxygen O are $69.9, 30.1$ respectively.

No. of moles of iron in oxide $= \frac{69.90}{55.85} = 1.25$

No. of moles of oxygen in oxide $= \frac{30.1}{16.0} = 1.88$

Ratio of iron to oxygen in oxide is given by

$1.25 : 1.88$

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

$1 : 1.5$

$2 : 3$

So, empirical formula $= Fe_2O_3$

Empirical formula mass $= [2(55.85)+3(16.00)]\text{g} = 159.69\text{ g}$

Thus, $n = \frac{\text{Molar mass}}{\text{Emprical formula mass}}$

$= \frac{159.69\text{ g}}{159.7\text{ g}}$

$= 0.999 = 1\text{ (approx)}$

As $n = 1$, so molecular formula $= Fe_2O_3$ after multiplying n with empirical formula

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: So, average atomic mass of chlorine is

$=$ [(Fractional abundance of ^{15}Cl)(Molar mass of ^{15}Cl) + (Fractional abundance of ^{37}Cl)(Molar mass of ^{37}Cl)

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

$= 26.4959+8.9568$

$= 35.4527$ u

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) One mole of C_2H_6 have two moles of carbon atoms

So, no. of moles $= 2\times3 = 6$

(ii) One mole of C_2H_6 have six moles of hydrogen atoms

So, no. of moles $= 3\times6 = 18$

(iii) One mole of C_2H_6 have 6.023×10^{23} molecules of ethane

So, no. of molecules $= 3\times6.023\times10^{23} = 18.069\times10^{23}$

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 2 L.

Answer: Here, we denoted molarity as M

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

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

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

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

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

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

12. If the density of methanol is 0.793 kg L^{-1} what is the volume needed for making 2.5 L of its 0.25 M solution.

Answer: Methanol represents as CH_3OH

Molar mass $= (1\times12)+(4\times1)+(1\times16) = 32\text{ g mol}^{-1} = 0.032\text{ kg mol}^{-1}$

Molarity $= \dfrac{0.793\text{ kg L}^{-1}}{0.032\text{ kg mol}^{-1}} = 24.78\text{ mol L}^{-1}$

Now, applying $M_1V_1 = M_2V_2$ we get:

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

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

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

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

Answer: $P = \dfrac{F}{A}$

$$= \frac{1034\text{ g}\times9.8\text{ ms}^{-2}}{\text{cm}^2}\times\frac{1\text{ kg}}{1000\text{ g}}\times\frac{(100)^2\text{ cm}^2}{1\text{ m}^2}$$

$$= 1.01332\times10^5\text{ kg m}^{-1}\text{ s}^{-2}$$

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

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

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

So, pressure in pascal $= 1.01332 \times 10^5$ Pa

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

Answer: The SI unit of mass is kilogram i.e. kg and it is defined as mass which is equal to international prototype of kilogram

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}

16. What do you mean by significant figures?

Answer: Those meaningful digits which known with creativity is known as significant figures. They indicate uncertainty or calculated value in figures. In other words, we can say that significant figures defined as total no. of digits in no. which include last digits that shows uncertainty of the result.

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 percent by mass.

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

Answer: (i) 1 ppm is equal to one part out of one million parts

So, mass percent of 15 ppm $= \dfrac{15}{10^6} \times 100 = 1.5 \times 10^{-3}\%$

(ii) 100 g contains 1.5×10^{-3} g of $CHCl_3$

$1000\ g$ contains $1.5\times10^{-2}\ g$ of $CHCl_3$

$$\text{Molality} = \frac{1.5\times10^{-2}\ g}{\text{Molar mass of } CHCl_3}$$

$$\text{Molar mass} = 12.00+1.00+3(35.5) = 119.5\ g\ mol^{-1}$$

$$\text{Molality} = 0.0125\times10^{-2}\ m = 1.25\times10^{-4}\ m$$

18. Express the following in the scientific notation:

(i) 0.0048

(ii) 234000

(iii) 8008

(iv) 500.0

(v) 6.0012

Answer: (i) $0.0048 = 4.8\times10^{-3}$

(ii) $234000 = 2.34\times10^{5}$

(iii) $8008 = 8.008\times10^{3}$

(iv) $500.0 = 5.000\times10^{2}$

(v) $6.0012 = 6.0012$

19. How many significant figures are present in the following

(i) 0.0025

(ii) 208

(iii) 5005

(iv) 126000

(v) 500.0

(vi) 2.0034

Answer: (i) Here, two significant figures are present

(ii) Here, three significant figures are present

(iii) Here, four significant figures are present

(iv) Here, three significant figures are present

(v) Here, four significant figures are present

(vi) Here, five significant figures are present

20. Round up the following up to three significant figures:

(i) 34.216

(ii) 10.4107

(iii) 0.04597

(iv) 2808

Answer: (i) 34.2

(ii) 10.4

(iii) 0.0460

(iv) 2810

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) $1\,\mathrm{km} = ...\,\mathrm{mm} = ...\,\mathrm{pm}$

(ii) $1\,\mathrm{mg} = ...\,\mathrm{kg} = ...\,\mathrm{ng}$

(iii) $1\,\mathrm{mL} = ...\,\mathrm{L} = ...\,\mathrm{dm}^3$

Answer: (a) Let us fix mass of nitrogen at $28\ \mathrm{g}$ then mass of dioxygen combine with mass of dinitrogen i.e. $32\mathrm{g}, 64\mathrm{g}, 32\mathrm{g}, 80\mathrm{g}$. So, dioxygen have ratio $1:2:2:5$ which shows law of multiple proportions. And that law states that if two elements combine to form more than one then mass of one combine with fixed mass of another in that ratio which we obtained.

(b) (i)

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

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

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

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

$$\therefore\ 1\text{ km} = 10^{6}\text{ mm} = 10^{15}\text{pm}$$

(b) (ii)

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

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

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

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

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

(b)(iii)

$$1\text{ mL} = 1\text{ mL}\times\frac{1\text{ L}}{1000\text{ mL}}$$

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

$$1\text{ mL} = 1\text{ cm}^{3} = 1\frac{1\text{dm}\times1\text{dm}\times1\text{dm}}{10\text{ cm}\times10\text{ cm}\times10\text{ cm}}\text{cm}^{3}$$

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

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

22. If the speed of light is $3.0\times10^{8}\text{ms}^{-1}$ calculate the distance covered by light in 2.00 ns.

Answer: We have given that time $= 2.00\text{ ns} = 2.00\times10^{-9}\text{s}$

And speed of light $= 3.0\times10^{8}\text{ms}^{-1}$

As we know that Distance $=$ Speed $\times$ Time

$$= \left(3.0\times10^{8}\text{ ms}^{-1}\right)\left(2.00\times10^{-9}\text{s}\right)$$
$$= 6.00\times10^{-1}\text{ m}$$
$$= 0.600\text{ m}$$

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

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

(ii) $2\text{ mol A} + 3\text{ mol B}$

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

(iv) $5\text{ mol A} + 2.5\text{ mol B}$

(v) $2.5\text{ mol A} + 5\text{ mol B}$

Answer: (i) As we have given the reaction so accordingly, one atom of A reacts with one molecule of B. Hence, 200 molecules of B react with 200 atoms of A. So, remaining 100 atoms of A are unused. Thus, B is limiting reagent.

(ii) As we have given the reaction so accordingly, one mole of A reacts with one mole of B. Hence, two moles of A react with two moles of B. So, remaining one mole of A are not consumed. Thus, A is limiting reagent.

(iii) As we have given the reaction so accordingly, one atom of A combines with one molecule of B. Hence, all 100 atoms of A combine with all 100 molecules of B. So, this is stoichiometric. Thus, no limiting reagent is there.

(iv) As we have given the reaction so accordingly, one atom of A combines with one mole of molecule of B. Hence, 2.5 mole of B combines with 2.5 mole of A. So, remaining 2.5 mole of A is unused. Thus, B is limiting reagent.

(v) As we have given the reaction so accordingly, one mole of A combines with one mole of molecule of B. Hence, 2.5 mole of A combines with 2.5 mole of B. So, remaining 2.5 mole of B will left. Thus, A is limiting reagent.

24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation $N_{2(g)} + H_{2(g)} \rightarrow 2NH_{3(g)}$

(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: (i) First, balance chemical equation which is given $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

Here, one mole of dinitrogen reacts with three mole of dihydrogen to give two mole of ammonia

2.00×10^3 g dinitrogen reacts with $\frac{6\text{ g}}{28\text{ g}} \times 2.00 \times 10^3$ g dihydrogen

i.e. 2.00×10^3 g dinitrogen reacts with 428.6 g dihydrogen

So, amount of dihydrogen = 1.00×10^3 g

Thus, N_2 is limiting reagent

Hence, $28\ g$ of N_2 produces $34\ g$ of NH_3

Thus, Mass of ammonia $= \frac{34\ g}{28\ g} \times 2000\ g = 2428.57\ g$

(ii) Here, N_2 is limiting reagent and H_2 is excess reagent. So, H_2 remain unreacted

(iii) Maa of dihydrogen which left unreacted $= 1.00 \times 10^3 g - 428.6 g = 571.4\ g$

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

Answer: firstly, we find molar mass of $Na_2CO_3 = (2\times23)+12.00+(3\times16) = 106\ g\ mol^{-1}$

Here, one mole of Na_2CO_3 implies 106 g of Na_2CO_3

So, $0.50\ mol\ Na_2CO_3 = \frac{106\ g}{1\ mole} \times 0.5\ mol\ Na_2CO_3$

$$= 53\ g\ Na_2CO_3$$
$$= 0.50\ M\ of\ Na_2CO_3$$
$$= 0.50\ mol/L\ Na_2CO_3$$

26. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Answer: Here is the reaction of dihydrogen gas which can be written as

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

So, two volumes of dihydrogen react with one of dihydrogen to obtain two volumes of water vapour

Thus, ten volumes react with five to produce ten volumes of water vapour.

27. Convert the following into basic units:

(i) 28.7 pm

(ii) 15.15 pm

(iii) 25365 mg

Answer: (i)

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

$28.7\ pm = 28.7\times10^{-12}\ m = 2.87\times10^{-11}\ m$

(ii)

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

$15.15\ \text{pm} = 15.15\times10^{-12}\ \text{m} = 1.515\times10^{-12}\ \text{m}$

(iii)

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

$25365\ \text{mg} = 2.5365\times10^{4}\times10^{-3}\ \text{g}$

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

$2.5365\times10^{1}\ \text{g} = 2.5365\times10^{-1}\times10^{-3}\ \text{kg}$

$25365\ \text{mg} = 2.5365\times10^{-2}\ \text{kg}$

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

(i) 1 g Au(s)

(ii) 1 g Na(s)

(iii) 1g Li(s)

(iv) 1 g of Cl_2(g)

Answer: (i)

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

$$= \frac{6.022\times10^{23}}{197}\text{ atoms of Au(s)}$$

$$= 3.06\times10^{21}\text{ atoms of Au(s)}$$

(ii)

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

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

$$= 0.262\times10^{23}\text{ atoms of Na(s)}$$

$$= 26.2\times10^{21}\text{ atoms of Na(s)}$$

(iii)

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

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

$= 0.86\times10^{23}$ atoms of Li(s)

$= 86.0\times10^{21}$ atoms of Li(s)

(iv)

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

$= \frac{6.022\times10^{23}}{71}$ atoms of $Cl_2(g)$

$= 0.0848\times10^{23}$ atoms of $Cl_2(g)$

$= 8.48\times10^{21}$ atoms of Cl_2 (g)

Thus, 1 g of Li (s) this have largest no. of atoms.

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: Ethanol denotes as C_2H_5OH

Mole fraction $= \frac{\text{Number of moles of } C_2H_3OH}{\text{Number of moles of solution}}$

$$0.040 = \frac{n_{C_2H_5OH}}{n_{C_2H_5OH} + n_{H_2O}} \quad \text{......(i)}$$

No. of moles in one litre water

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

$$n_{H_2O} = 55.55 \text{ mol}$$

Substitute values in (i)

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

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

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

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

$n_{C_2H_5OH} = 2.314$ mol

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

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

Answer: Here, One mole of carbon atom $= 6.023\times10^{23} = 12$ g

So, mass of one ^{12}C atom in g $= \dfrac{12 \text{ g}}{6.022\times10^{23}} = 1.993\times10^{-23}$ g

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

(i) $\dfrac{0.02856\times298.15\times0.112}{0.5785}$

(ii) 5×5.364

(iii) $0.0125+0.7864+0.0215$

Answer: (i) Least precise no. $= 0.112$

No. of significant figures $=$ no. of significant figures in least precise no. $= 3$

(ii) Least precise no. $= 5.364$

No. of significant figures $=$ no. of significant figures in least precise no. $= 4$

(iii) here, least no. in each term is 4. So, significant figures in answer $= 4$

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

Isotopes	Isotopic 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%

Answer: Here, molar mass of argon is

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

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

$= 39.947$ g mol^{-1}

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)

$1 \text{ mole of Ar} = 6.022 \times 10^{23} \text{ atoms of Ar}$

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

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

(ii)

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

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

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

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

$= 13 \text{ atom of He}$

(iii)

$4 \text{ g of He} = 6.022 \times 10^{23} \text{ atoms of He}$

$52 \text{ g of He} = \frac{6.022 \times 10^{23} \times 52}{4} \text{ atoms of He}$

$= 7.8286 \times 10^{24} \text{ atoms of He}$

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)

One mole of carbon dioxide contains $12\ \text{g}$ of carbon.

Therefore, $3.38\ \text{g}$ of carbon dioxide contain carbon $= \dfrac{12\ \text{g}}{44\ \text{g}} \times 3.38\ \text{g} = 0.9217\ \text{g}$

$18\ \text{g}$ water contains $2\ \text{g}$ hydrogen.

Therefore, $0.690\ \text{g}$ of water contain hydrogen $= \dfrac{2\ \text{g}}{18\ \text{g}} \times 0.690 = 0.0767\ \text{g}$

As carbon and hydrogen are only constituents of compound

So, total mass $=\ 0.9217\ \text{g} + 0.0767\ \text{g} = 0.9984\ \text{g}$

Percent of carbon $= \dfrac{0.9217\ \text{g}}{0.9984\ \text{g}} \times 100 = 92.32\%$

Percent of hydrogen $= \dfrac{0.0767\ \text{g}}{0.9984\ \text{g}} \times 100 = 7.68\%$

Moles of carbon $= \dfrac{92.32}{12.00} = 7.69$

Moles of hydrogen $= \dfrac{7.68}{1} = 7.68$

Ration of carbon to hydrogen $=\ 7.69 : 7.68 = 1 : 1$

Thus, empirical formula $=\ CH$

(ii) As it is given that

Weight of $10.0\ \text{L}$ of gas $=\ 11.6\ \text{g}$

So, weight of $22.4\ \text{L}$ of gas $= \dfrac{11.6\ \text{g}}{10.0\ \text{L}} \times 22.4\ \text{L} = 25.984\ \text{g} \approx 26\ \text{g}$

(iii) Empirical formula of $CH\ =\ 12+1 = 13\ \text{g}$

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

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

$= 2$

So, molecular formula after putting value of n in empirical formula we get

$= (CH)_n = C_2H_2$

35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to reaction $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(1)}$ 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 present in one litre of water.

$= \left[(0.75\text{ mol}) \times \left(36.5\text{ g mol}^{-1}\right)\right]$ HCl present in one litre of water.

$= 27.375$ g HCl present in one litre of water.

So, amount of HCl in 25 mL $= \dfrac{27.375\text{ g}}{1000\text{ mL}} \times 25\text{ mL} = 0.6844\text{ g}$

According to given chemical reaction we get:

Two mol of HCl $(2 \times 36.5 = 71\text{ g})$ react with one mol of $CaCO_3$

Amount of $CaCO_3$ with 0.6844 g $= \dfrac{100}{71} \times 0.6844\text{ g} = 0.9639\text{ g}$

36. Chlorine is prepared in the laboratory by treating manganese dioxide with aqueous hydrochloric acid according to the reaction $4HCl_{(aq)} + MnO_{2(s)} \rightarrow 2H_2O_{(l)} + MnCl_{2(aq)} + Cl_{2(g)}$

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

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

So, 5.0 g of MnO_2 reacts with $= \dfrac{146\text{ g}}{87\text{ g}} \times 5.0\text{ g of HCl} = 8.4\text{ g of HCl}$

Thus, 8.4 g of HCl react with 5.0 g of manganese dioxide.