

NCERT Solutions for Class 11 Chemistry Chapter 1
Some Basic Concepts of Chemistry

Exercise:

Q 1. Calculate the molar mass of the following:

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

Ans.

(i) CH_4 :

$$\begin{aligned}\text{Molecular mass of CH}_4 &= \text{Atomic mass of C} + 4 \times \text{Atomic mass of H} \\ &= 12 + 4 \times 1 \\ &= 16 \text{ u}\end{aligned}$$

(ii) H_2O :

$$\begin{aligned}\text{Molar mass of water H}_2\text{O} \\ \text{Atomic mass of H} &= 1 \\ \text{Atomic mass of O} &= 16 \\ \text{H}_2\text{O} &= 2 \times \text{H} + 1 \times \text{O} \\ \text{Molar mass of water} &= 2 \times 1 + 16 = 18 \text{ g/mol}\end{aligned}$$

(iii) CO_2 :

$$\begin{aligned}\text{Molecular mass of CO}_2 &= \text{Atomic mass of C} + 2 \times \text{Atomic mass of O} \\ &= 12 + 2 \times 16 \\ &= 44 \text{ u}\end{aligned}$$

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:

$$\begin{aligned}&= \frac{46.0 \text{ g}}{142.066 \text{ g}} \times 100 \\ &= 32.379 \\ &= 32.4\%\end{aligned}$$

Mass percent of the sulphur element:

$$\begin{aligned}&= \frac{32.066 \text{ g}}{142.066 \text{ g}} \times 100 \\ &= 22.57 \\ &= 22.6\%\end{aligned}$$

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Mass percent of the oxygen element:

$$\begin{aligned} &= \frac{64.0g}{142.066g} \times 100 \\ &= 45.049 \\ &= 45.05\% \end{aligned}$$

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:

$$\begin{aligned} &= \frac{\text{percent of iron by mass}}{\text{Atomic mass of iron}} \\ &= \frac{69.9}{55.85} \\ &= 1.25 \end{aligned}$$

Relative moles of oxygen in iron oxide:

$$\begin{aligned} &= \frac{\text{percent of oxygen by mass}}{\text{Atomic mass of oxygen}} \\ &= \frac{30.1}{16.00} \\ &= 1.88 \end{aligned}$$

The simplest molar ratio of iron to oxygen:

$$\Rightarrow 1.25 : 1.88 \Rightarrow 1 : 1.5 \Rightarrow 2 : 3$$

Therefore, the empirical formula of the iron oxide is Fe_2O_3 .

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

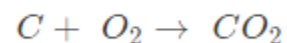
(i) 1 mole of carbon is burnt in air.

(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.



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

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

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

= 0.1875 mole

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

Therefore, the mass of CH_3COONa

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

= 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.

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Molar mass of HNO_3

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

$$= 1 + 14 + 48$$

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

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,

$$\text{Concentration of } \text{HNO}_3 = 15.44 \text{ mol/L}$$

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

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

$$= 63.5 + 32.00 + 64.00$$

$$= 159.5 \text{ grams}$$

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159.5 grams of CuSO_4 contains 63.5 grams of Cu.

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

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

$$= 39.81 \text{ grams}$$

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

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

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$$= 159.7\text{g}$$

The molar mass of $\text{Fe}_2\text{O}_3 = 159.69\text{g}$

$$\text{Therefore } n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.69\text{ g}}{159.7\text{ g}}$$
$$= 0.999$$
$$= 1(\text{approx})$$

The molecular formula of a compound can be obtained by multiplying n 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 $^{35}\text{Cl} = 0.7577$ and Molar mass = 34.9689

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

Average Atomic mass = $(0.7577 \times 34.9689)\text{amu} + (0.2423 \times 36.9659)$

$$= 26.4959 + 8.9568 = 35.4527$$

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.

\therefore No. of moles of C- atoms in 3 moles of C_2H_6 .

$$= 2 \times 3$$

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$$= 6$$

(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 .

$$= 3 \times 6$$

$$= 18$$

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

∴ No. of molecules in 3 moles of C_2H_6

$$= 3 \times 6.023 \times 10^{23}$$

$$= 18.069 \times 10^{23}$$

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,

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

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

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

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

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

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

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.

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Molar mass of methanol (CH₃OH)

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

molarity of the given solution

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

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

$$\text{Applying } M_1 \times V_1 = M_2 V_2$$

(Given solution) (solution to be prepared)

$$24.78 \times V_1 = 0.25 \times 2.5 \text{ L}$$

$$\text{or } V_1 = 0.02522 \text{ L} = 25.22 \text{ mL}$$

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.

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

But weight = mg

∴ Pressure = Weight per unit area

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

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

$$= 1.01332 \times 10^5 \text{ Pa}$$

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:

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

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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

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

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

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

Q18. Express the following in the scientific notation:

(i) 0.0048

(ii) 234,000

(iii) 8008

(iv) 500.0

(v) 6.0012

Ans.

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

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

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

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

(v) $6.0012 = 6.0012 \times 10^0$

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

(a) 0.0025

(b) 208

(c) 5005

(d) 126,000

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(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:

	<i>Mass of dioxygen</i>	<i>Mass of dinitrogen</i>
(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:

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- (i) $1 \text{ km} = \dots\dots\dots \text{ mm} = \dots\dots\dots \text{ pm}$
(ii) $1 \text{ mg} = \dots\dots\dots \text{ kg} = \dots\dots\dots \text{ ng}$
(iii) $1 \text{ mL} = \dots\dots\dots \text{ L} = \dots\dots\dots \text{ dm}^3$

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)

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

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

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

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

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

$$1 \text{ mL} = 1 \text{ cm}^3$$

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

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,

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

We know that,

$$\text{Distance} = \text{Speed} \times \text{Time}$$

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So,

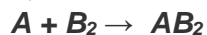
Distance travelled in 2 ns = speed of light \times time taken

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

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

$$= 0.6 \text{ m}$$

Q23. In a reaction



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

(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

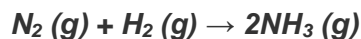
1 mole of A reacts with 1 mole of B. Similarly 2.5 moles of A reacts with 2.5 moles of B, so 2.5 moles of A is unused. Hence, B is the limiting reagent.

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(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:



(i) Calculate the mass of NH_3 produced if $2 \times 10^3 \text{ g}$ N_2 reacts with $1 \times 10^3 \text{ g}$ 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

$$\therefore 2000 \text{ g of } \text{N}_2 \text{ will react with } \text{H}_2 = \frac{6}{28} \times 2000 \text{ g} = 428.6 \text{ g}$$

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 $\text{NH}_3 = 2 \text{ mol}$

= 34 g

Therefore, 2000 g will produce $\text{NH}_3 = \frac{34}{28} \times 2000 \text{ g}$

= 2428.57 g

(ii) H_2 will remain unreacted

(iii) Mass left unreacted = $1000 \text{ g} - 428.6 \text{ g} = 571.4 \text{ g}$

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

$$= \frac{106 \text{ g}}{1 \text{ mol}} \times 0.5 \text{ mol } \text{Na}_2\text{CO}_3$$

= 53 g of Na_2CO_3

0.5 M of $\text{Na}_2\text{CO}_3 = 0.5 \text{ mol/L } \text{Na}_2\text{CO}_3$

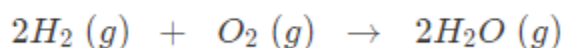
Hence, 0.5 mol of Na_2CO_3 is in 1 L of water or 53 g of Na_2CO_3 is in 1 L of water.

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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 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

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

$$28.7 \text{ pm} = 28.7 \times 10^{-12} \text{ m}$$

$$= 2.87 \times 10^{-11} \text{ m}$$

(ii) 15.15 pm

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

$$15.15 \text{ pm} = 15.15 \times 10^{-12} \text{ m}$$

$$= 1.515 \times 10^{-11} \text{ m}$$

(iii) 25365 mg

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

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

$$25365 \text{ mg} = 25365 \times 10^{-6} \text{ kg}$$

$$25365 \text{ mg} = 2.5365 \times 10^{-2} \text{ kg}$$

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 Cl₂ (g)

Ans.

(i) 1 g of 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 of 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 of 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 Cl₂ (g)

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

(Molar mass of Cl₂ molecule = 35.5 × 2 = 71 g mol⁻¹)

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

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

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

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 eqn (1),

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

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

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

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

Therefore, molarity of solution

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

$$= 2.314 \text{ M}$$

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

Ans.

1 mole of carbon atoms

= 6.023×10^{23} atoms of carbon

= 12 g of carbon

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

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

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

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) 5×5.364

(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) $0.0125 + 0.7864 + 0.0215$

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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	$35.96755 \text{ g mol}^{-1}$	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:

$$\begin{aligned} &= \left[(35.96755 \times \frac{0.337}{100}) + (37.96272 \times \frac{0.063}{100}) + (39.9624 \times \frac{99.600}{100}) \right] \\ &= [0.121 + 0.024 + 39.802] \text{ g mol}^{-1} \\ &= 39.947 \text{ g mol}^{-1} \end{aligned}$$

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 = 6.023×10^{23} atoms of Ar

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

= 3.131×10^{25} atoms of Ar

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(ii) 52 u of He

1 atom of He = 4 u of He

OR

4 u of He = 1 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 atoms of He

(iii) 52 g of He

4 g of He = 6.023×10^{23} atoms of He

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

= 7.829×10^{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

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

= 0.9218 g

18 g of water contains 2 g of hydrogen

Therefore, 0.690 g of water will contain hydrogen

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

$$= 0.0767 \text{ g}$$

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.

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(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

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

$$= 25.984 \text{ g}$$

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

(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 = 2 x CH = C₂H₂.

Q35. Calcium carbonate reacts with aqueous HCl to give CaCl₂ and CO₂ according to the reaction, CaCO₃ (s) + 2 HCl (aq) → CaCl₂(aq) + CO₂ (g) + H₂O(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

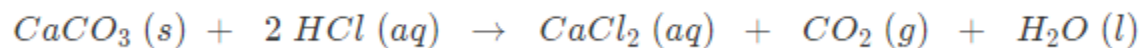
Therefore, amt of HCl present in 25 mL of solution

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

$$= 0.6844 \text{ g}$$

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Given chemical reaction,



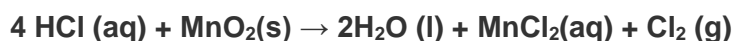
2 mol of HCl ($2 \times 36.5 = 73$ g) react with 1 mol of CaCO_3 (100 g)

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.

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

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

1 mole of MnO_2 reacts with 4 mol of HCl

Hence,

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 .