NCERT Solutions for Class-XI Chemistry

Chapter-1
NCERT Chemistry Class 11

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

1. (i) H_2O :
The molecular mass of water, H_2O
= (2 × Atomic mass of hydrogen) + (1 × 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 :
The molecular mass of carbon dioxide, CO_2
= (1 × Atomic mass of carbon) + (2 × Atomic mass of oxygen)
= [1(12.011 u) + 2 (16.00 u)]
= 12.011 u + 32.00 u
= 44.01 u

(iii)CO_4 :
The molecular mass of methane, CH_4
= (1 × Atomic mass of carbon) + (4 × 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).

2. Atomic Mass of Sodium = 23g
Atomic Mass of Sulphur = 32g
Atomic Mass of Oxygen = 16g
From the molecular formula of sodium sulphate, it is understood that in 1 mole of sodium sulphate, 1 mole of sulphur, 4 moles of oxygen and 2 moles of sodium is present.
So Molecular Mass of Sodium Sulphate = [23×2] + [32×1] + [16×4]
= 46 + 32 + 64
= 142g
Percentage of Sodium in sodium sulphate = [46×100]/142

= [4600/142]

= 32.394%

Percentage of Sulphur in sodium sulphate = [32×100]/142

= 3200/142

= 22.535%

Percentage of Oxygen in sodium sulphate = [64×100]/142

= 6400/142

= 45.070%

Therefore the percentage of different elements in sodium sulphate is given below:

Sr. No.	Element	Percentage
1.	Sodium	32.39
2.	Sulphur	22.54
3.	Oxygen	45.07

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

3. % 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:

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

$= \frac{30.1}{16.00}$

= 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 $Fe2O_3$.

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.

4. The reaction of carbon with oxygen is given as follows:

$$C + O_2 \rightarrow CO_2$$

Number of mole 1 1 1

(i) By observing the above equation, it is understood that when 1 mole of carbon is burnt in air, then the amount of carbon dioxide produced is 44g [Molecular Mass of carbon dioxide].
(ii) In the above equation 32g of oxygen react with 1 mole of carbon, then 44g of carbon dioxide is produced.
32g of Oxygen → 44g of Carbon Dioxide

16g of Oxygen → X g of Carbon Dioxide
X = [44×16]/32
= 704/32
= 22g
Therefore 16g of oxygen when reacted with 1 mole of carbon we get 22g of CO_2.
(iii) Even though 2 moles of carbon is available for reaction but <u>16g of oxygen can react with only 0.5 moles of carbon</u> to give 22g of CO_2. In this oxygen acts as limiting reagent and carbon acts as an excess reagent. <u>So 1.5 moles carbon would be left unreacted.</u>

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}

5. 0.375 M aqueous solution of sodium acetate
≡ 1000 mL of solution containing 0.375 moles of sodium acetate
∴ Number of moles of sodium acetate in 500 mL
$= \frac{0.375}{1000} \times 500$
= 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 per cent of nitric acid in it being 69%.

6. Given:
Density of nitric acid = 1.41 g/mol
Percentage of Nitric acid = 69%
Solution:
69% of nitric acid signifies 69g of nitric acid dissolved in 100g of solvent.
Volume of nitric acid solution = [mass of solution]/ density of solution
= 100/1.41
= 70.922 ml
Number of moles of Nitric Acid = Mass of Nitric acid/ Molecular Mass of nitric acid
= 69/63
= 1.095 moles

Molarity, $M_o = \frac{\text{Number of moles of solute}}{\text{Volume of solution in ml}} \times 1000$
$= \frac{1.0925}{70.922} \times 1000$
= 1092.5/70.92
= 15.404 M
≈ 15.4 M

Therefore concentration of nitric acid is 15.4 mole/L.

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

7. 1 mole of $CuSO_4$ contains 1 mole of copper.

Molar mass of $CuSO_4$ = (63.5) + (32.00) + 4(16.00)

= 63.5 + 32.00 + 64.00

= 159.5 g 159.5 g of $CuSO_4$ contains 63.5 g of copper.

⇒ 100 g of $CuSO_4$ will contain $\frac{63.5 \times 100\text{ g}}{159.5}$ of copper.

Amount of copper that can be obtained from 100 g $CuSO_4$ $= \frac{63.5 \times 100}{159.5}$

= 39.81 g

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. Given that the molar mass of the oxide is 159.69 g mol^{-1}.

8. Finding Empirical Formula:

To determine the empirical formula of iron oxide, we need to find the simplest whole number ratio.

Element	Percentage	Atomic Mass	Relative Number of Moles	Simplest Mole Ratio	Whole Number Ratio
Iron[Fe]	69.9	55.84	69.9/55.84 = 1.252	1.252/1.252 = 1	2
Oxygen [O]	30.1	16	30.1/16 = 1.88	1.88/1.252 = 1.5	3

Therefore the empirical formula for iron oxide is Fe_2O_3.

Finding Molecular Formula:

Empirical Formula Mass = [55.84×2] + [16×3]

= 111.68 + 48

= 159.68g

Molecular Mass [given] = 159.69g

n = [Molecular Mass]/ [Empirical Formula Mass]

= 159.69/159.68

= 1

Molecular Formula = n × Empirical Formula

= 1 × Fe_2O_3.

= Fe_2O_3

Therefore the Molecular Formula is Fe_2O_3.

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

9. The average atomic mass of chlorine

$$= \left[\begin{pmatrix}\text{Fractional abudance} \\ \text{of } ^{35}\text{Cl}\end{pmatrix}\begin{pmatrix}\text{Molar mass} \\ \text{of } ^{35}\text{Cl}\end{pmatrix} + \begin{pmatrix}\text{Fractional} \\ \text{abundance} \\ \text{of } ^{37}\text{Cl}\end{pmatrix}\begin{pmatrix}\text{Molar mass} \\ \text{of } ^{37}\text{Cl}\end{pmatrix}\right]$$

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

$= 26.4959 + 8.9568$

$= 35.4527$ u

$\therefore$ The average atomic mass of chlorine = 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.

10. In 1 mole of ethane (C_2H_6), it is seen that 2 moles of carbon and 6 moles of hydrogen is present.

(i) Since three moles of ethane is present,

So Number of moles of carbon present $= 3\times2$

$= 6$

Therefore 6 moles of carbon is present in 3 moles of ethane.

(ii) Since three moles of ethane is present,

So Number of moles of hydrogen present $= 3\times6$

$= 18$

Therefore 18 moles of hydrogen is present in 3 moles of ethane.

(iii) 1 mole of ethane contains 6.023×10^{23} molecules of ethane [By Avogadro's Law]

Therefore Number of molecules if ethane in 3 moles of ethane

$= 3\times6.023\times10^{23}$

$= 18.069\times10^{23}$

Therefore Number of molecules of ethane present in 3 moles of ethane is 18.069×10^{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?

11. Molarity (M) of a solution is given by,

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

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

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

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

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

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

$\therefore$ Molar concentration of sugar = $0.02925\ mol\ L^{-1}$

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?

12. Molecular Formula of Methanol = [CH_3OH]
Atomic Mass of Hydrogen = 1g
Atomic Mass of Oxygen = 16g
Atomic Mass of Carbon = 12g
So Molecular Mass of sugar = [12×1] + [4×1] + [16×1]
= 12 + 4 + 16
= 32g
= 0.032kg
Molarity of methanol solution = [0.793]/0.032
= 24.781 mol/L
Applying the relation, $M_1V_1 = M_2V_2$
24.781×X = 2.5×0.25
X = [2.5×0.25]/24.781
= 0.025 L
= 25.22 ml
Therefore volume needed for making 2.5 L of its 0.25 M solution is 25.22 ml.

13. 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.

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

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

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

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

We know, 1 N = 1 kg ms–2

Then,

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

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

$\therefore$ Pressure = 1.01332×10^5 Pa

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

14. The SI system was internationally accepted in 1960. SI is the abbreviation of "Systeme International" Units which is French equivalent of "International System of Units".

The SI unit of Mass is kilogram.
1 kilogram is defined as the mass of a platinum-iridium cylinder kept in Paris. In practice, the mass of 1litre of water at 4°C is 1 kilogram. On atomic scale, 1 kilogram is equivalent to the mass of 5.0188× 10^{25} atoms of $_6C^{12}$ [isotope of carbon].

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

15.

	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?

16. The total number of digits in a measured physical quantity is called the number of significant figures. The number of significant figures refers to the precision of a measured quantity and may be defined as follows:
The number of significant figures in a measured physical quantity refers to number of digits written, including the last one whose value is uncertain.
Rules for Determination of Significant Figures:
[i] All non-zero digits are significant.
For example: 146 cm has three significant figures.
[ii] The zeros placed to the left of the first non-zero digit in the given physical quantity are not significant.
For example: 0.56cm has two significant figures.
[iii] The zeros placed to the right of the decimal point are significant.
For example: 15,0cm has three significant figures.
[iv] The zeros placed between two non-zero digits are significant.
For example: 6.08cm has three significant figures.

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.

17. (i) 1 ppm is equivalent to 1 part out of 1 million (106) parts.

Mass percent of 15 ppm chloroform in water $= \frac{15}{10^6} \times 100$

$\simeq 1.5 \times 10^{-3}$ %

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

⇒ 1000 g of the sample contains 1.5×10^{-2} g of $CHCl_3$.

∴ Molality of chloroform in water

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

Molar mass of $CHCl_3$ = 12.00 + 1.00 + 3(35.5) = 119.5 g mol^{-1}

∴ Molality of chloroform in water = 0.0125×10^{-2} m

= 1.25×10^{-4} m

18. Express the following in the scientific notation:

(i) 0.0048

(ii) 234,000

(iii)8008

(iv)500.0

(v) 6.0012

18. (i) 0.0048

The scientific notation of 0.0048 is 4.8×10^{-3}

(ii) 234,000

The scientific notation of 234,000 is 2.34×10^{-5}

(iii) 8008

The scientific notation of 8008 is 8.01×10^{3}

(iv) 500.0

The scientific notation of 500 is 5×10^{2}

(v) 6.0012

The scientific notation of 6.0012 is 6×10^{0}

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

(i) 0.0025

(ii) 208

(iii)5005

(iv)1,26,000

(v) 2.0034

19. (i) 0.0025

There are 2 significant figures.

(ii) 208

There are 3 significant figures.

(iii)5005

There are 4 significant figures.

(iv)1,26,000

There are 3 significant figures.

(v) 500.0

There are 4 significant figures.

(vi)2.0034

There are 4 significant figures.

20. Round up the following upto three significant figures:

(i) 34.216

(ii) 10.4107

(iii)0.04597

(iv)2808

20. (i) 34.216 – 34.2

(ii) 10.4107-10.4

(iii) 0.04597-0.046

(iv) 2808 – 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 km = ... mm = ... pm

(ii) 1 mg = ... kg = ... ng

(iii)1 mL =...................... L = ... dm^3

21. (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, 64 g, 32 g, 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.

(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 km = 10^6 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 km = 10^{15} pm

Hence, 1 km = 10^6 mm = 10^{15} pm

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

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

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

$\Rightarrow$ 1 mg = 10^6 ng

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

(iii) $1\text{ mL} = 1\text{ mL} \times \dfrac{1\text{ L}}{1000\text{ mL}}$

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

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

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

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

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

22. Speed = [Distance]/ Time

Distance = $[3.0 \times 10^8] \times [2\times10^{-9}]$

= 0.6 m

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

23. A limiting reagent determines the extent of a reaction. It is the reactant which is the first to get consumed during a reaction, thereby causing the reaction to stop and limiting the amount of products formed.

(i) According to the given reaction, 1 atom of A reacts with 1 molecule of B. Thus, 200 molecules of B will react with 200 atoms of A, thereby leaving 100 atoms of A unused. Hence, B is the limiting reagent.

(ii) According to the reaction, 1 mole of A reacts with 1 mole of B. Thus, 2 mole of A will react with only 2 mole of B. As a result, 1 mole of A will not be consumed. Hence, A is the limiting reagent.

(iii) According to the given reaction, 1 atom of A combines with 1 molecule of B. Thus, all 100 atoms of A will combine with all 100 molecules of B. Hence, the mixture is stoichiometric where no limiting reagent is present.

(iv) 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of B will combine with only 2.5 mole of A. As a result, 2.5 mole of A will be left as such. Hence, B is the limiting reagent.

(v) According to the reaction, 1 mole of atom A combines with 1 mole of molecule B. Thus, 2.5 mole of A will combine with only 2.5 mole of B and the remaining 2.5 mole of B will be left as such. Hence, A is the limiting reagent.

24. Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

(i) Calculate the mass of ammonia produced if 2.00 × 103 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?

24. The reaction given in the question is not balanced.

Balancing the above chemical reaction:

$N_2 + 3H_2 \rightarrow 2NH_3$

(i) Finding mass of NH_3 produced:

Atomic Mass of Hydrogen = 1g

Atomic Mass of Nitrogen = 14g

Molecular Mass of Nitrogen = 14 + 3

= 17 g

From the chemical equation we observe that 28g of nitrogen react with 6g of hydrogen to give 34g of ammonia

Therefore 1 mole of dinitrogen reacts with 3 moles dihydrogen to give 2 moles of ammonia.

Therefore amount of dihydrogen required to react with 2000g of dinitrogen = [6×2000]/28

= 12000/28

= 428.57g

Therefore 2000g of dinitrogen will produce = [34×2000]/28

= 68000/28

= 2428.57g of Ammonia

Therefore 2428.57g of ammonia is produce when 2000g of dinitrogen is used.

(ii) Since dinitrogen is present in lesser amount, it is the limiting reagent and dihydrogen is the excess reagent and hence H_2 will be left unreacted.

(iii) Mass of dihydrogen left unreacted = 1000 – 428.57

= 571.43g

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

25. Molar mass of Na_2CO_3 = (2 × 23) + 12.00 + (3× 16) = 106 g mol^{-1}

Now, 1 mole of Na_2CO_3 means 106 g of Na_2CO_3.

$\therefore$ 0.5 mol of Na_2CO_3 $= \frac{106}{1\text{ mole}} \times 0.5 \text{ mol } Na_2CO_3$

= 53 g Na_2CO_3

⇒ 0.50 M of Na_2CO_3 = 0.50 mol/L Na_2CO_3

Hence, 0.50 mol of Na_2CO_3 is present in 1 L of water or 53 g of Na_2CO_3 is present in 1 L of water.

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

26. 1 pm = 10^{-12} m

28.7 pm = 28.7×10^{-12}m

$= 2.87\times10^{-12}$m

(ii) 15.15 pm

1 pm = 10^{-12} m

15.15 pm = 15.15×10^{-12}m

$= 1.515\times10^{-12}$m

(iii) 25365 mg

1 mg = 10^{-3} g

1 mg = 10^{-6} kg

∴ 25365 mg = 25365×10^{-6} kg

$= 2.5365\times10^{-2}$ kg

27. Convert the following into basic units:

(i) 28.7 pm

(ii) 15.15 pm

(iii)25365 mg

27. (i) 28.7 pm:

1 pm = 10^{-12} m

∴ 28.7 pm = 28.7×10^{-12} m

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

(ii) 15.15 pm:

1 pm = 10^{-12} m

∴ 15.15 pm = 15.15×10^{-12} m

$= 1.515 \times 10^{-12}$ m

(iii)25365 mg:

1 mg = 10^{-3} g

25365 mg = $2.5365 \times 10^{4} \times 10^{-3}$ g

Since,

1 g = 10^{-3} kg

2.5365×10^{1} g = $2.5365 \times 10^{-1} \times 10^{-3}$ kg

∴ 25365 mg = 2.5365×10^{-2} kg

28. Which one of the following will have 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)

28. (i) Atomic Mass of Gold [Au] = 197g

197g of Gold [Au] contains 6.023×10^{23} atoms

Therefore 1 g of Gold will contain = $[6.023\times10^{23}]/197$

$= 3.057\times10^{21}$ atoms

(ii) Atomic Mass of Sodium [Na] = 23g

23g of Sodium [Na] contains 6.023×10^{23} atoms

Therefore 1 g of Sodium [Na] will contain = $[6.023\times10^{23}]/23$

= 2.619×10^{22} atoms

(iii) Atomic Mass of Lithium [Li] = 7g

7g of Lithium [Li] contains 6.023×10^{23} atoms

Therefore 1 g of Lithium [Li] will contain = [6.023×10^{23}]/7

= 8.604×10^{22} atoms

(iv) Atomic Mass of Chlorine [Cl_2] = 71g

71g of Chlorine [Cl_2] contains 6.023×10^{23} atoms

Therefore 1 g of Chlorine [Cl_2] will contain = [6.023×10^{23}]/71

= 8.483×10^{21} atoms

Therefore 1g of Lithium [Li] will contains the maximum number 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).

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

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

Number of moles present in 1 L water:

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

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

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

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

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

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

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

$$n_{C_2H_5OH} = 2.314\text{ mol}$$

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

= 2.314 M

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

30. By Avogadro's Law

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

6.023×10^{23} atoms of carbon → 12g of Carbon

1 atom of Carbon → X g of Carbon

X = [12]/ 6.023×10^{23}

= 1.992×10^{-23}g of Carbon

Therefore mass of 1 atom of Carbon is 1.992×10^{-23}g.

31. 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

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

Least precise number of calculation = 0.112

∴ Number of significant figures in the answer

= Number of significant figures in the least precise number = 3

(ii) 5 × 5.364

Least precise number of calculation = 5.364

∴ Number of significant figures in the answer = Number of significant figures in 5.364 = 4

(iii)0.0125 + 0.7864 + 0.0215

Since the least number of decimal places in each term is four, the number of significant figures in the answer is also 4.

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%

32. Atomic Mass of First Isotope of Argon = 35.96755g

Natural Abundance of First Isotope of Argon, P_{a1}= 0.337% = 0.00337

Total Atomic Mass of 1st Isotope, M_{a1} = 35.96755×0.00337

= 0.12121g

Atomic Mass of Second Isotope of Argon = 37.96272g

Natural Abundance of Second Isotope of Argon, P_{a2} = 0.063%

= 0.00063

Total Atomic Mass of 2nd Isotope, M_{a2} = 37.96272×0.00063

= 0.02392g

Atomic Mass of Third Isotope of Argon = 39.9624g

Natural Abundance of Third Isotope of Argon, P_{a3} = 99.6%

Total Atomic Mass of 3rd Isotope, M_{a3} = 39.9624×0.996

= 39.8026g

So Average Atomic Mass of Argon $= \frac{M_{a1} + M_{a2} + M_{a3}}{P_{a1} + P_{a2} + P_{a3}}$

$\rightarrow \frac{0.12121 + 0.02392 + 39.8026}{0.996 + 0.00063 + 0.0037}$

→ 39.9477g

Therefore the average atomic mass of Argon is 39.95g

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.

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

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

= 3.131 × 1025 atoms of Ar

(ii) 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)4 g of He = 6.022×10^{23} atoms of He

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

= 7.8286×10^{24} 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) 52 u of He

(iii)52 g of He.

34. Finding Percentage of Carbon:

44 parts of CO_2→12 parts of C

OR

44g of CO_2→12 g of C

Therefore according to the question

3.38 g of CO_2 → X g of C

X = [12× 3.38] /44

=0.921g

18 g of water → 2g of hydrogen

So 0.690 g of water → X g of hydrogen

X = [2× 0.690]/18

=0.0767g

Since carbon and hydrogen are the only constituents of the compound, the total mass of the compound is obtained as follows:

= 0.9217 g + 0.0767 g = 0.9984 g

Percentage of carbon = [100×weight of carbon]/weight of compound

= [0.921× 100]/0.998

= 92.32%

Also percentage of hydrogen =[100×weight of hydrogen]/weight of compound

= [0.0766 × 100]/0.998

= 7.68%

(i) Finding Empirical Formula:

Element	Atomic Mass	Percentage	Relative Number of Moles	Simplest Mole Ratio	Whole Number Ratio
Carbon [C]	12	92.32	92.32/12 = 7.69	7.69/7.68= 1.001	1
Hydrogen [H]	1	7.68	7.68/1	7.68/7.68 = 1	1

Empirical formula of the compound = CH

Therefore the Empirical Formula of the compound is CH

(ii) Weight of 22.4 L of gas of STP = [11.6g×22.4L]/10.0L

= 25.985g

≈ 26 g

Therefore the molecular mass of the gas is 26 g.

(iii) Finding Molecular formula:

Empirical formula mass = 12+1 = 13g

Also molecular mass =26 g [as obtained in step (ii)]

n = [Molecular Mass] / [Empirical Formula Mass]

= 26/13

= 2

Now molecular formula = n ×Empirical Formula = 2 × CH = C_2H_2

Therefore the molecular formula of the compound is C_2H_2.

35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction, $CaCO_{3(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$

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

35. 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.

∴ Amount of HCl present in 25 mL of solution

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

= 0.6844 g

From the given chemical equation,

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

2 mol of HCl (2 × 36.5 = 71 g) react with 1 mol of CaCO3 (100 g).

∴ Amount of CaCO3 that will react with 0.6844 g $= \frac{100}{71} \times 0.6844\ g$

= 0.9639 g

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

$4HCl_{(aq)} + MnO_{2(s)} \rightarrow 2H2O_{(l)} + MnCl_{2(aq)} + Cl_{2(g)}$

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

36. Atomic Mass of Manganese = 55g

Atomic Mass of Oxygen = 16g

Atomic Mass of Hydrogen = 1g

Molecular Mass of MnO_2 = 55 + [16×2]

= 55 + 32

= 87g

Molecular Mass of HCl = 1 + 35.5

= 36.5g

So by observing the chemical reaction we see that,

1moles of MnO_2 reacts with 4moles of HCl

87g of MnO_2 → 4×36.5g of HCl

So 5g of MnO_2 → X g of HCl

X = [146×5]/87

X = 730/87

X = 8.39g

So 8.4g of HCl will react completely with 5g of MnO_2.