NCERT SOLUTIONS

CLASS - 11th



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Class : 11th
Subject : Chemistry
Chapter : 1
Chapter Name : Some Basic Concepts Of Chemistry

Q1.1 Calculate the molar mass of the following:
(i) H_2O
(ii) CO_2
(iii) CH_4

Answer. (i) H_2O:
The molecular mass of water, H_2O
$= (2 \times \text{ Atomic mass of hydrogen}) + (1 \times \text{ Atomic mass of oxygen })$
$= [2(1.0084) + 1(16.00u)]$
$= 2.016u + 16.00u$
$= 18.016$
$= 18.02u$

(ii) CO_2
The molecular mass of carbon dioxide, CO_2
$= (1 \times \text{ Atomic mass of carbon }) + (2 \times \text{ Atomic mass of oxygen })$
$= [1(12.011u) + 2(16.00u)]$
$= 12.011u + 32.00u$
$= 44.01u$

(iii) CH_4:
The molecular mass of methane, CH_4
$= (1 \times \text{ Atomic mass of carbon }) + (4 \times \text{ Atomic mass of hydrogen })$
$= [1(12.011u) + 4(1.008u)]$
$= 12.011u + 4.032u$
$= 16.043u$

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Q1.2 Calculate the mass percent of different elements present in sodium sulphate (Na_2SO_4).

Answer. The molecular formula of sodium sulphate is Na_2SO_4
Molar mass of $Na_2SO_4 = [(2 \times 23.0) + (32.066) + 4(16.00)]$
$= 142.066g$
$\text{Mass percent of an element} = \frac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$
$\therefore$ Mass percent of sodium:

$= \frac{46.0\text{g}}{142.066\text{g}} \times 100$
$= 32.379$
$= 32.4\%$
Mass percent of sulphur:
$= \frac{32.066\text{g}}{142.066\text{g}} \times 100$
$= 22.57$
$= 22.6\%$
Mass percent of oxygen:
$= \frac{64.0\text{g}}{142.066\text{g}} \times 100$
$= 45.049$
$= 45.05\%$

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Q1.3 Determine the empirical formula of an oxide of iron, which has 69.9% iron and 30.1% dioxygen by mass.

Answer. $\%$ of iron by mass $= 69.9\%[\text{ Given }]$
$\%$ of oxygen by mass $= 30.1\%[\text{ Given }]$
Relative moles of iron in iron oxide:
$= \frac{\%\text{ of iron by mass}}{\text{Atomic mass of iron}}$
$= \frac{69.9}{55.85}$
$= 1.25$
Relative moles of oxygen in iron oxide:
$= \frac{\%\text{ of oxygen by mass}}{\text{Atomic mass of oxygen}}$
$= \frac{30.1}{16.00}$
$= 1.88$
Simplest molar ratio of iron to oxygen:
$= 1.25 : 1.88$
$= 1 : 1.5$
$= 2 : 3$
$\therefore$ The empirical formula of the iron oxide is Fe_2O_3

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Q1.4 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. The balanced reaction of combustion of carbon can be written as:
(i) As per the balanced equation, 1 mole of carbon burns in 1 mole of dioxygen (air) to Produced 1 mole of carbon dioxide.

(ii) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mole of carbon to give 22 g of carbon dioxide. Hence, it is a limiting reactant.
(iii) According to the question, only 16 g of dioxygen is available. It is a limiting reactant. Thus, 16 g of dioxygen can combine with only 0.5 mole of carbon to give 22 g of carbon dioxide.

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Q1.5 Calculate the mass of sodium acetate $\left(\mathrm{CH}_3 \mathrm{COONa}\right)$ required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is 82.0245 gmol^{-1}.

Answer. $0.375 \mathrm{M}$ aqueous solution of sodium acetate.
$\equiv 1000 \mathrm{mL}$ of solution containing 0.375 moles of sodium acetate.
$\therefore$ Number of moles of sodium acetate in 500 mL.
$=\frac{0.375}{1000} \times 500$
$=0.1875$ mole
Molar mass of sodium acetate = $82.0245 \mathrm{gmole}^{-1}$(Given)
$\therefore$ Required mass of sodium acetate $=\left(82.0245 \mathrm{gmol}^{-1}\right)(0.1875 \text { mole})$
$=15.38 \mathrm{g}$

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Q1.6 Calculate the concentration of nitric acid in moles per litre in a sample which has a density, $1.41 \mathrm{gmL}^{-1}$ and the mass percent of nitric acid in it being 69%.

Answer. Mass percent of nitric acid in the sample $=69 \%[$ Given $]$
Thus, 100g of nitric acid contains 69g of nitric acid by mass. Molar mass of nitric acid $\left(\mathrm{HNO}_3\right)$
$=\{1+14+3(16)\} \mathrm{gmol}^{-1}$
$=1+14+48$
$=63 \mathrm{gmol}^{-1}$
$\therefore$ Number of moles in 69 g of HNO_3
$=\frac{69 \mathrm{g}}{63 \mathrm{gmol}^{-1}}$
$=1.095 \mathrm{mol}$
Volume of 100g nitric acid solution
$=\frac{\text { Mass of solution }}{\text { density of solution }}$
$=\frac{100 \mathrm{g}}{1.41 \mathrm{gmL}^{-1}}$
$=\frac{100 \mathrm{g}}{1.4 \lg \mathrm{mL}^{-1}}$
$=70.92 \mathrm{mL} \equiv 70.92 \times 10^{-3} \mathrm{L}$
Concentration of nitric acid
$=\frac{1.095 \mathrm{mole}}{70.92 \times 10^{-3} \mathrm{L}}$
$=15.44 \mathrm{mol} / \mathrm{L}$
$\therefore$ Concentration of nitric acid $=15.44 \mathrm{mol} / \mathrm{L}$

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Q1.7 How much copper can be obtained from 100 g of copper sulphate $(\mathrm{CuSO_4})$?

Answer. 1 mole of $\mathrm{CuSO_4}$ contains 1 mole of copper.
Molar mass of cus $O_4 = (63.5) + (32.00) + 4(16.00)$
$= 63.5 + 32.00 + 64.00$
$= 159.5\mathrm{g}$
$159.5\mathrm{g}$ of $\mathrm{CuSO_4}$ contains 53.5 g of copper.
$\Rightarrow$ 100g of cus$\mathrm{O_4}$ will contain $\frac{63.5 \times 100\mathrm{g}}{159.5}$ of copper.
$\therefore$ AMount of copper that can be obtained from 100 g $\mathrm{CuSO_4} = \frac{63.5 \times 100}{159.5}$
$= 39.81\mathrm{g}$

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Q1.8 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.

Answer. Mass percent of iron (Fe) = $= 69.9\%$ (Given)
Mass percent of oxygen (O) = $30.1\%(\text{ Given })$
Number of moles of iron present in the oxide = $\frac{69.90}{55.85}$
$= 1.25$
Number of moles of oxygen present in the oxide $= \frac{30.1}{16.0}$
$= 1.88$
$Ratio of iron to oxygen in the oxide,$
$= 1.25 : 1.88$
$= \frac{1.25}{1.25} : \frac{1.88}{1.25}$
$= 1 : 1.5$
$= 2 : 3$
$\therefore$ The empirical formula of the oxide is $\mathrm{Fe_2O_3}$
Empirical formula mass of $\mathrm{Fe_2O_3} = [2(55.85) + 3(16.00)]\mathrm{g}$
Molar mass of $\mathrm{Fe_2O_3} = 159.69\mathrm{g}$
$\therefore n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{159.69\mathrm{g}}{159.7\mathrm{g}}$
$= 0.999$
$= 1(\text{ approx })$
Molecular formula of a compound is obtained by multiplying the empirical formula with n.
Thus, the empirical formula of the given oxide is $\mathrm{Fe_2O_3}$ and n is 1 .
Hence, the molecular formula of the oxide is $\mathrm{Fe_2O_3}$.

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Q1.9 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. The average atomic mass of chlorine.

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

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

$= 26.4959 + 8.9568$

$= 35.4527\text{u}$

$\therefore$ The average atomic mass of chlorine $= 35.4527\text{u}$

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Q1.10 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) 1 mole of C_2H_6 contains 2 moles of carbon atoms.

$\therefore$ Number of moles of carbon atoms in 3 moles of C_2H_6

$= 2 \times 3 = 6$

(ii) 1 mole of C_2H_6 contains 6 moles of hydrogen atoms.

$\therefore$ Number of moles of carbon atoms in 3 moles of C_2H_6

$= 3 \times 6 = 18$

(iii) 1 mole of C_2H_6 contains 6.023×10^{23} molecules of ethane.

$\therefore$ Number of molecules in 3 moles of C_2H_6

$= 3 \times 6.023 \times 10^{23} = 18.069 \times 10^{23}$

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Q1.11 What is the concentration of sugar $(C_{12}H_{22}O_{11})$ in $\text{mol}L^{-1}$ if its 20g are dissolved in enough water to make a final volume up to 2L?

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

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

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

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

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

$= \frac{0.0585\text{mol}}{2\text{L}}$
$= 0.02925\text{molL}^{-1}$
$\therefore$ Molar concentration of sugar $= 0.02925\text{molL}^{-1}$

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Q1.12 If the density of methanol is 0.793kgL^{-1}. what is its volume needed for making 2.5 L of its 0.25 M solution?

Answer. Molar mass of methanol $(\text{CH}_3\text{OH}) = (1 \times 12) + (4 \times 1) + (1 \times 16)$
$= 32\text{gmol}^{-1}$
$= 0.032\text{kgmol}^{-1}$
Molarity of methanol solution $= \frac{0.793\text{kgL}^{-1}}{0.032\text{kgmol}^{-1}}$
$= 24.78\text{molL}^{-1}$
Since density is mass per unit volume
Applying,
$M_1V_1 = M_2V_2$
Given solution) (Solution to be prepared
$\left(24.78\text{molL}^{-1}\right)\text{V}_1 = (2.5\text{L})\left(0.25\text{molL}^{-1}\right)$
$\text{V}_1 = 0.0252\text{L}$
$\text{V}_1 = 25.22\text{mL}$

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Q1.13 Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below:
$1\text{Pa} = 1\text{Nm}^{-2}$
If mass of air at sea level is 1034 g cm^{-2}. calculate the pressure in pascal.

Answer. Pressure is defined as force acting per unit area of the surface.
$P = \frac{F}{A}$
$= \frac{1034\text{g} \times 9.8\text{ms}^{-2}}{\text{cm}^2} \times \frac{1\text{kg}}{1000\text{g}} \times \frac{(100)^2\text{cm}^2}{1\text{m}^2}$
$= 1.01332 \times 10^5\text{kgm}^{-1}\text{s}^{-2}$
We know,
$1\text{N} = 1\text{kgms}^{-2}$
Then,
$1\text{Pa} = 1\text{Nm}^{-2} = 1\text{kgm}^{-2}\text{s}^{-2}$
$1\text{Pa} = 1\text{kgm}^{-1}\text{s}^{-2}$
$\therefore$ Pressure $= 1.01332 \times 10^5\text{Pa}$

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Q1.14 What is the SI unit of mass? How is it defined?

Answer. The SI unit of mass is kilogram (kg). 1 kilogram is defined as the mass equal to the mass of the international prototype of kilogram.

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Q1.15 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.

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

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Q1.16 What do you mean by significant figures?

Answer. Significant figures are those meaningful digits that are known with certainty. They indicate uncertainty in an experiment or calculated value. For example, if 15.6 mL is the result of an experiment, then 15 is certain while 6 is uncertain, and the total number of significant figures are 3. Hence, significant figures are defined as the total number of digits in a number including the last digit that represents the uncertainty of the result.

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Q1.17 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 equivalent to 1 part out of 1 million $\left(10^{6}\right)$ parts.
$\therefore$ Mass percent of 15ppm chloroform in water.
$=\frac{15}{10^{6}} \times 100$
$\approx 1.5 \times 10^{-3} \%$

(ii) 100g of the sample contains 1.5×10^{-3}g of CHCl_{3}.
$\Rightarrow$ 1000g of the sample contains 1.5×10^{-2}g of CHCl_{3}
$\therefore$Molality of chloroform in water.
$=\frac{1.5 \times 10^{-2} \mathrm{g}}{\text{Molar mass of } \mathrm{CHCl}_{3}}$
Molar mass of $\mathrm{CHCl}_{3}=12.00+1.00+3(35.5)$
$=119.5 \mathrm{gmol}^{-1}$
$\therefore$ Molality of chloroform in water $=0.0125 \times 10^{-2} \mathrm{m}$
$=1.25 \times 10^{-4} \mathrm{m}$

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Q1.18 Express the following in the scientific notation:
(i) 0.0048
(ii) 234,000
(iii) 8008
(iv) 500.0
(v) 6.0012

Answer. (i) $0.0048=4.8 \times 10^{-3}$
(ii) $234,000=2.34 \times 10^{5}$
(iii) $8008=8.008 \times 10^{3}$
(iv) $500.0=5.000 \times 10^{2}$
(v) $6.0012=6.0012$

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Q1.19 How many significant figures are present in the following?
(i) 0.0025
(ii) 208
(iii) 5005
(iv) 126,000
(v) 500.0
(vi) 2.0034

Answer. (i) There are 2 significant figures.
(ii) There are 3 significant figures.
(iii) There are 4 significant figures.
(iv) There are 3 significant figures.
(v) There are 4 significant figures.

(vi) There are 5 significant figures.

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Q1.20 Round up the following upto 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

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Q1.21 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

Answer. (a) If we fix the mass of dinitrogen at 28 g, then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 32 g, 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.
(b) (i) $1\mathrm{km} = 1\mathrm{km} \times \frac{1000\mathrm{m}}{1\mathrm{km}} \times \frac{100\mathrm{cm}}{1\mathrm{m}} \times \frac{10\mathrm{mm}}{1\mathrm{cm}}$
$\therefore 1\mathrm{km} = 10^6\mathrm{mm}$
$1\mathrm{km} = 1\mathrm{km} \times \frac{1000\mathrm{m}}{1\mathrm{km}} \times \frac{1\mathrm{pm}}{10^{-12}\mathrm{m}}$
$\therefore 1\mathrm{km} = 10^{15}\mathrm{pm}$
Hence , $1\mathrm{km} = 10^6\mathrm{mm} = 10^{15}\mathrm{pm}$

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

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

(iii) $1\mathrm{mL}=1\mathrm{mL}\times\frac{1\mathrm{L}}{1000\mathrm{mL}}$
$\Rightarrow 1\mathrm{mL}=10^{-3}\mathrm{L}$
$1\mathrm{mL}=1\mathrm{cm}^3=1\mathrm{cm}^3\frac{1\mathrm{dm}\times1\mathrm{dm}\times1\mathrm{dm}}{10\mathrm{cm}\times10\mathrm{cm}\times10\mathrm{cm}}$
$\Rightarrow 1\mathrm{mL}=10^{-3}\mathrm{dm}^3$
$\therefore 1\mathrm{mL}=10^{-3}\mathrm{L}=10^{-3}\mathrm{dm}^3$

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Q1.22 If the speed of light is $3.0\times10^{8}\mathrm{ms}^{-1}$. calculate the distance covered by light in 2.00 ns.

Answer. According to the question:
Time taken to cover the distance = 2.00 ns
$=2.00\times10^{-9}\mathrm{s}$
Speed of light $=3.0\times10^{8}\mathrm{ms}^{-1}$
Distance travelled by light in 2.00 ns
= Speed of light x Time taken
$=\left(3.0\times10^{8}\mathrm{ms}^{-1}\right)\left(2.00\times10^{-9}\mathrm{s}\right)$
$=6.00\times10^{-1}\mathrm{m}$
$=0.600\mathrm{m}$

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Q1.23 In a reaction
$\mathrm{A}+\mathrm{B}_2\rightarrow\mathrm{AB}_2$
Identify the limiting reagent, if any, in the following reaction mixtures.
(i) 300 atoms of A + 200 molecules of B
(ii) 2 mol A + 3 mol B
(iii) 100 atoms of A + 100 molecules of B
(iv) 5 mol A + 2.5 mol B
(v) 2.5 mol A + 5 mol B

Answer. A 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 Iimiting reagent.
(ii) According to the reaction, 1 mol of A reacts with 1 mol of 3. Thus, 2 mol of A will react with only 2 mol of B. As a result, 1 mol 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 mol of atom A combines with 1 mol of molecule B. Thus, 2.5 mol of B will combine with only 2.5 mol of A. As a result, 2.5 mol of A will be left as such. Hence, 3 is the limiting reagent.
(v) According to the reaction, 1 mol of atom A combines with 1 mol of molecule B. Thus, 2.5 mol of A will combine with only 2.5 mol of B and the remaining 2.5 mol of B will be left as such. Hence, A is the limiting reagent.

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Q1.24 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:
$\mathrm{N_2(g) + H_2(g) \rightarrow 2NH_3(g)}$
(i) Calculate the mass of ammonia produced if 2.00×10^3g dinitrogen reacts with $1.00 \times 10^3 g$ of dihydrogen.
(ii) Will any of the two reactants remain unreacted?
(iii) If yes, which one and what would be its mass?

Answer. (i) Balancing the given chemical equation,
$\mathrm{N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}}$
From equation, 1 mole (28 g) of nitrogen reacts with 3 mole (6 g) of dihydrogen to give 2 mole (34 g) of ammonia.
$\Rightarrow 2.00 \times 10^3$g of di nitrogen will react with $\frac{6\mathrm{g}}{28\mathrm{g}} \times 2.00 \times 10^3$g dihydrogen i.e., 2.00×10^3g of nitrogen will react with 428.6 of dihydrogen.
Given,
Amount of dihydrogen = 1.00×10^3g
Hence, $\mathrm{N_2}$ is the limiting reagent.
$\therefore 28\mathrm{g}$ of $\mathrm{N_2}$ produces $34\mathrm{g}$ of $\mathrm{NH_3}$
Hence, mass of ammonia produced by 2000 g of $\mathrm{N_2} = \frac{34\mathrm{g}}{28\mathrm{g}} \times 2000\mathrm{g}$
$= 2428.57\mathrm{g}$
(ii) $\mathrm{N_2}$ is the limiting reagent and $\mathrm{H_2}$ is the excess reagent. Hence, $\mathrm{H_2}$ will remain unreacted.
(iii) Mass of dihydrogen left unreacted = $1.00 \times 10^3\mathrm{g} - 428.6\mathrm{g}$
$= 571.4\mathrm{g}$

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Q1.25 How are 0.50 mol $\mathrm{Na_2CO_3}$ and $0.50\mathrm{M}$ $\mathrm{Na_2CO_3}$ different?

Answer. Molar mass of $\mathrm{Na_2CO_3} = (2 \times 23) + 12.00 + (3 \times 16)$
$= 106\mathrm{gmol^{-1}}$
Now, 1 mole of $\mathrm{Na_2CO_3}$ means 106g of $\mathrm{Na_2CO_3}$
$\therefore 0.5\mathrm{mol}$ of $\mathrm{Na_2CO_3} = \frac{106\mathrm{g}}{1\mathrm{mole}} \times 0.5\mathrm{mol}\mathrm{Na_2CO_3}$
$= 53\mathrm{g}\mathrm{Na_2CO_3}$
$\Rightarrow 0.50\mathrm{M}$ of $\mathrm{Na_2CO_3} = 0.50\mathrm{mol/L}\mathrm{Na_2CO_3}$
Hence, $0.50\mathrm{mol}$ of $\mathrm{Na_2CO_3}$ is present in 1 L of water or 53 g of $\mathrm{Na_2CO_3}$ is present in 1 L of water.

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Q1.26 If 10 volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Answer. Reaction of dihydrogen with dioxygen can be written as:
$2\mathrm{H}_{2(\mathrm{s})}+\mathrm{O}_{2(x)} \longrightarrow 2\mathrm{H}_2\mathrm{O}_{(x)}$
Now, two volumes of dihydrogen react with one volume of dihydrogen to produce two volumes of water vapour. Hence, ten volumes of dihydrogen will react with five volumes of dioxygen to produce ten volumes of water vapour.

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Q1.27 Convert the following into basic units:
(i) 28.7 pm
(ii) 15.15 pm
(iii) 25365 mg

Answer. (i) $28.7\mathrm{pm}$:
$1\mathrm{pm}=10^{-12}\mathrm{m}$
$\therefore 28.7\mathrm{pm}=28.7 \times 10^{-12}\mathrm{m}$
$=2.87 \times 10^{-11}\mathrm{m}$
(ii) $15.15\mathrm{pm}$:
$1\mathrm{pm}=10^{-12}\mathrm{m}$
$\therefore 15.15\mathrm{pm}=15.15 \times 10^{-12}\mathrm{m}$
$=1.515 \times 10^{-12}\mathrm{m}$
(iii) $25365\mathrm{mg}$:
$1\mathrm{mg}=10^{-3}\mathrm{g}$
$25365\mathrm{mg}=2.5365 \times 10^4 \times 10^{-3}\mathrm{g}$
Since,
$1\mathrm{g}=10^{-3}\mathrm{kg}$
$2.5365 \times 10^1\mathrm{g}=2.5365 \times 10^{-1} \times 10^{-3}\mathrm{kg}$
$\therefore 25365\mathrm{mg}=2.5365 \times 10^{-2}\mathrm{kg}$

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Q1.28 Which one of the following will have the largest number of atoms?
(i) 1 g Au (s)
(ii) 1 g Na (s)
(iii) 1 g Li (s)
(iv) $1\mathrm{g}$ of $\mathrm{Cl}_2(\mathrm{g})$

Answer. $1\mathrm{g}$ of $\mathrm{Au(s)}=\frac{1}{197}$ mol of $\mathrm{Au(s)}$

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

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

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

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

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

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

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

$= \frac{6.022 \times 10^{33}}{7}$ $atomsofLi(s)$

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

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

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

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

$= \frac{6.022 \times 10^{33}}{71}$ atoms of $Cl_2(g)$

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

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

$Hence, 1gofLi(s)willhavethelargestnumberofatoms.$

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Q1.29 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. Mole fraction of $C_2H_5OH = \frac{\text{Number of moles of } C_2H_5OH}{\text{Number of moles of solution}}$

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

Number of moles present in 1 L water:

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

$n_{H_2O} = 55.55 mol$

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

$\frac{n_{C_2H_3OH}}{n_{C_2H_3OH} + 55.55} = 0.040$

$n_{C_2H_3OH} = 0.040 n_{C_2H_3OH} + (0.040)(55.55)$

$0.96 n_{C_2H_3OH} = 2.222 mol$

$h_{C_2H_3OH} = \frac{2.222}{0.96} mol$

$h_{C_2H_3OH} \equiv 2.314 mol$

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

$= 2.314 M$

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Q1.30 What will be the mass of one ^{12}C atom in g?

Answer. 1 mole of carbon atoms $= 6.023 \times 10^{23}$ atoms of carbon
$= 12$g of carbon
$\therefore$ Mass of one $^{12}C = \frac{12g}{6.022 \times 10^{23}}$
$= 1.993 \times 10^{-23}$g

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Q1.31 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) $0.0125 + 0.7864 + 0.0215$

Answer. (i) $\frac{0.02856 \times 298.15 \times 0.112}{0.5785}$
Least precise number of calculation = 0.112
$\therefore$ 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 $= 5.364$
$\therefore$ 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.

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Q1.32 Use the data given in the following table to calculate the molar mass of naturally occuring argon isotopes:

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

Answer. Molar mass of argon
$= \left[\left(35.96755 \times \frac{0.337}{100}\right) + \left(37.96272 \times \frac{0.063}{100}\right) + \left(39.9624 \times \frac{90.60}{100}\right)\right] \text{gmol}^{-1}$
$= [0.121 + 0.024 + 39.802]\text{gmol}^{-1}$
$= 39.947\text{gmol}^{-1}$

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Q1.33 Calculate the number of atoms in each of the following (i) 52 moles of Ar (ii) 52 u of He (iii) 52 g of He.

Answer. (i) 1 mole of $\mathrm{Ar} = 6.022 \times 10^{23}$ atoms of Ar
$\therefore$ 52 mol of $\mathrm{Ar} = 52 \times 6.022 \times 10^{23}$ atoms of Ar
$= 3.131 \times 10^{25}$ atoms of Ar
(ii) 1 atom of He $= 4$ u of He
Or,
4u of $\mathrm{He} = 1$ atom of He
1u of $\mathrm{He} = \frac{1}{4}$ atom of He
52u of $\mathrm{He} = \frac{52}{4}$ atom of He
$= 13$ atoms of He
(iii) 4g of $\mathrm{He} = 6.022 \times 10^{23}$ atoms of He
$\therefore$ 52 g of He = $\frac{6.022 \times 10^{23} \times 52}{4}$ atoms of He
$= 7.8286 \times 10^{24}$ atoms of He

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Q1.34 A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate
(i) empirical formula,
(ii) molar mass of the gas, and (iii) molecular formula.

Answer. (i) 1 mole $(44\ \mathrm{g})$ of $\mathrm{CO_2}$ contains 12g of carbon.
$\therefore$ 3.38g of $\mathrm{CO_2}$ will contain carbon $= \frac{12\mathrm{g}}{44\mathrm{g}} \times 3.38\mathrm{g}$
$= 0.9217g$
18 g of water contains 2 g of hydrogen.
$\therefore$ 0.690 g of water will contain hydrogen = $\frac{2\mathrm{g}}{18\mathrm{g}} \times 0.690$
$= 0.0767\mathrm{g}$
Since carbon and hydrogen are the only constituents of the compound, the total mass of the compound is:
$= 0.9217\mathrm{g} + 0.0767\mathrm{g}$
$= 0.9984\mathrm{g}$
$\therefore$ Percent of C in the compound $= \frac{0.9217\mathrm{g}}{0.9984\mathrm{g}} \times 100$
$= 92.32\%$
Percent of H in the compound $= \frac{0.0767\mathrm{g}}{0.9984\mathrm{g}} \times 100$
$= 7.68\%$
Moles of carbon in the compound $= \frac{92.32}{12.00}$
$= 7.69$
Moles of hydrogen in the compound $= \frac{7.68}{1}$
$= 7.68$

$\therefore$ Ratio of carbon to hydrogen in the compound $= 7.69 : 7.68$
$= 1 : 1$
Hence, the empirical formula of the gas is CH.

(ii) Given,
Weight of 10.0L of the gas (at S.T.P) $= 11.6\text{g}$
$\therefore$ Weight of 22.4 L of gas at STP = $\frac{11.6\text{g}}{10.0\text{L}} \times 22.4\text{L}$
$= 25.984\text{g}$
$\approx 26\text{g}$
Hence, the molar mass of the gas is 26 g.

(iii) Empirical formula mass of CH = 12 + 1 = 13 g
$n = \frac{\text{Molar mass of gas}}{\text{Empirical formula mass of gas}}$
$= \frac{26\text{g}}{13\text{g}}$
$n = 2$
$\therefore$ Molecular formula of gas $= (\text{CH})_n$
$= \text{C}_2\text{H}_2$

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Q1.35 Calcium carbonate reacts with aqueous HCl to give CaCl_2 and CO_2 according to the reaction, $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$.
What mass of CaCO_3 is required to react completely with 25 mL of 0.75 M HCl?

Answer. 0.75M of $\text{HCl} \equiv 0.75\text{mol}$ of HCl are present in 1L of water
$\equiv \left[\left(0.75 \text{ mol}\right) \times \left(36.5\text{gmol}^{-1}\right)\right]$ HCl is present in 1L of water
$\equiv 27.375\text{g}$ of HCl is present in 1L of water
Thus, 1000 mL of solution contains 27.375 g of HCL.
$\therefore$ Amount of HCl present in 25mL
$= \frac{27.375\text{g}}{1000\text{mL}} \times 25\text{mL}$
$= 0.6844\text{g}$
From the given chemical equation,
$\text{CaCO}_{3(s)} + 2\text{HCl}_{(\text{aq})} \longrightarrow \text{CaCl}_{2(aq)} + \text{CO}_{2(\text{g})} + \text{H}_2\text{O}_{(g)}$
2 mol of $\text{HCl}(2 \times 36.5 = 71\text{g})$ react with 1 mol of $\text{CaCO}_3(100\text{g})$
$\therefore$ Amount of CaCO_3 that will react with $0.6844\text{g} = \frac{100}{71} \times 0.6844\text{g}$
$= 0.9639\text{g}$

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Q1.36 Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction
$4\text{HCl}(\text{aq}) + \text{MnO}_2(\text{s}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{MnCl}_2(\text{aq}) + \text{Cl}_2(\text{g})$
How many grams of HCl react with 5.0 g of manganese dioxide?

Answer. 1 mol $[55+2\times16=87\mathrm{g}]\mathrm{MnO_2}$ reacts completely with 4 mol $[4\times36.5=146\mathrm{g}]$ of HCL.
$\therefore 5.0\mathrm{g}$ of MnO $_2$ will react with
$=\frac{146\mathrm{g}}{87\mathrm{g}}\times5.0\mathrm{g}$ of HCL
$=8.4\mathrm{g}$ of HCl
Hence, $8.4\mathrm{g}$ of HCl will react completely with $5.0\mathrm{g}$ of manganese dioxide.

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