Unit 3: Stoichiometry

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"The calculation of reactant and products in a balanced chemical reaction".

or

"The study of relative amount of substance involved in a chemical reaction".

Stoichen=Element
Metron=Measure

Example:

$$\rightarrow H_2 + O_2 \rightarrow H_2O$$

$$H_2 + O_2 \rightarrow 2H_2O$$

$$2H_2 + O_2 \rightarrow 2H_2O$$

Left = Right

2 mol of hydrogen + 1 mol of oxygen → 2 mol of water

Mole

"Mole is a measuring unit of amount of substance"

→ "Atomic mass,
Formula mass and molecule mass of a substance expressed in grams"

one mol atom: 6.022x10 23 atom=Mass in g

Example:

Term	Particle Type	Molar Mass	Example (1 mole = ? grams)
O	Oxygen atoms	16 g/mol	16 g of O atoms
O_2	Oxygen molecules	32 g/mol	32 g of O_2 gas
NaCl	Ionic compound	58.5 g/mol	58.5 g of table salt

Mole :
⏭ Amount of a substance.
⏭ atomic mass in grams of a substance.
⏭ formula mass in grams of a substance.
⏭ molecular mass in grams of a substance.
⏭ unit is necessary.
⏭ unit is mol.
⏭ Example:
✴ one mole of oxygen is 16g .
(this is for atom)
✴ one mole of NaCl is 58.5g.(this is formula unit)
12:23 AM

Mole ratio/Conversion factor:

"Mole ratio tell us how many mole of each substance react or are produced in a chemical reaction "

Or

"A balanced chemical equation will tell you the ratio of amount reactant produce to the mole".

Example: $H_2 + O_2 \longrightarrow 2H_2O$

2 mol of hydrogen + 1 mol of oxygen ⟶ 2 mol of water

2kg chicken + 1kg rice ⟶ 2kg biryani

Explanation of mole ratio.

- **SI amount of substance: mole**
- **Symbolic representation: mol**

Example 3.1: (Methanol burns according to the following)

a) $2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$

3.50 mol of CH_3OH are burnt in oxygen,Calculate:

Given:

Step 1: Mole ratio /conversion factor between given and find from balance chemical equation.

Given : Find

$2CH_3OH : 3O_2$

$$C.F = \frac{Find}{Given} = \frac{3}{2} = 1.5 \text{ moles}$$

Step 2: Find desired quantity by multiplying given value in numerical.

$= \text{Desired Quantity} = \text{Given} \times C.F$

$= 3.5\,mol \times 1.5$

$= 5.25$ moles answer

b) How many moles of water are produced?

step 1

Given : Find

$2CH_3OH : 4H_2O$

$C.F = \frac{\text{Find}}{\text{Given}} = \frac{4}{2} = 2$ moles

Step 2 :- Desired Quantity = C.F × Given

$= 2\,mol \times 3.50$

$= 7$ moles answer

homework

Part a)

Concept assessment 3.1

Step 1

Given : Find

2 : 1

$C.F = \frac{1}{2} = 0.5$ moles

step 2 Desired Quantity = Given × C.F

$= 8.0 \times 0.5$

$= 4$ moles of N_2 Answer

Part b)

step 1: Given : Find

2 : 3

$$C.F = \frac{Find}{Given} = \frac{3}{2} = 1.5 \text{ moles}$$

step 2: Derived Quantity = Given × C.F

$$= 8.0 \times 1.5$$

$$= 12 \text{ moles of } H_2$$ Answer-

Example 3.2

SOLUTION:

Given : Find

Fe_2O_3 : Fe

1 mol : 2 mol

$$C.F = \frac{Find}{Given} = \frac{2}{1} = 2 \text{ mol}$$

rough

Fe_2O_3

$= 55.8 \times 2 + 16 \times 3$

$= 111.6 + 48$

$= 159.6$

Conversion of Mass in g to moles-

$$n = \frac{\text{Mass in g}}{\text{Molar mass}} = \frac{425}{159.6} = 2.66 \text{ mol}$$

$$\text{Derived Quantity} = C.F \times \text{Given}$$
$$= 2 \times 2.66$$
$$= 5.32 \text{ moles}$$

$$\text{Mass in gram} = n \times \text{molar mass}$$
$$= 5.32 \times 55.9$$

$$= 297.3 \text{ grams}$$

Final answer

Important Formulas

1) $n = \dfrac{\text{Mass in gram}}{\text{Molar mass}}$

2) $\text{Mass in gram} = n \times \text{Molar mass}$

3) $n = \dfrac{N\,(\text{Particle})}{N_A \text{ avogadro's}}$

4) $N\,(\text{Particle}) = n \times N_A$

Concept assessment 3.2

Solution:

Given : Find
H_2 : O_2
2 : 1

$$C.F = \frac{Find}{Given} = \frac{1}{2} = 0.5 \text{ moles}$$

Desire Quantity = C.F x Given
= 0.5 x 50
= 25 moles

$$n = \frac{100}{2} = 50$$

Mass in g = n x Molar mass
= 25 x 32
= 800 grams

O_2 = 16x2
= 32

Molar volume

1)

1 mol = 24 dm3

1 mol of H_2 RTP = 24 dm3

1 mol of NH_3 RTP = 24 dm3

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RTP
(Room Temperature pressure)
T = 25 deg C
P = 1 atm

2) 1 mol of H_2 STP = 24.414 dm3/litre

Exercise 3.3

1 mol of Cl_2 = $24 dm^3$

2.5 mol = $\frac{1}{x}$

$x = \frac{2.5 \times 24}{1} = 60 dm^3$ Answer

Concept assessment 3.3

a) 1 mol of O_2 = $24 dm^3$

$\frac{50 dm^3}{x} = 24 dm^3$

$x = \frac{50}{24} = 2.083 mol$ **Answer**

b) 1 mol of N_2 = $24 dm^3$

$0.80 = \frac{1}{x}$

$x = 0.80 \times 24$

$= 19.2 dm^3$ **Answer**

Percentage composition

Definition:

"Percentage composition is the ratio of mass of each element to the total mass of compound multiplied by 100"

→ It shows how much element contributes to total mass of compound

Formula:

$$\text{Percentage of Element} = \frac{\text{Mass of element in 1 mol compound}}{\text{Molar mass whole compound}} \times 100$$

Example: (MgO)

$$\% \text{ of } Mg = \frac{\text{Mass of } Mg}{\text{Molar mass of } MgO} \times 100$$

$$\% \text{ of } Mg = \frac{24}{40} \times 100 = 60.3\% \text{ by Mass}$$

$$\% \text{ of } O = \frac{16}{40} \times 100 = 39.7\% \text{ of Mass}$$

By mass 60.3% Mg and 39.7% O is present in MgO.Sum of % of all element in compound should be 100.

concept 3.4

1) H_2O:

Molar mass
$= 1\times2+16$
$= 2+16$
$= 18\ g/mol$

$\%\ of\ H = \frac{2}{18}\times100 = 11.11\%$

$\%\ of\ O = \frac{16}{18}\times100 = 88.88\%$

$11.11\% + 88.88\% = 99.99\% \sim 100\%$ ✓

2) H_2SO_4:

Molar Mass
$= 1\times2+32+16\times4$
$= 2+32+64$
$= 34+64$
$= 98\ g/mol$

$\%\ of\ H = \frac{2}{98}\times100 = 2.04\%$

$\%\ of\ S = \frac{32}{98}\times100 = 32.65\%$

$\%\ of\ O = \frac{64}{98}\times100 = 65.30\%$

$65.30 + 2.04 + 32.65 = 100\%$ ✓

3) $C_6H_{12}O_6$:

Molar Mass
$= 12\times6+12+16\times6$
$= 72+12+96$
$= 180\ g/mol$

$\%\ of\ C = \frac{72}{180}\times100 = 40\%$

$\%\ of\ H = \frac{12}{180}\times100 = 6.66\%$

$\%\ of\ O = \frac{96}{180}\times100 = 53.33\%$

$40 + 6.66 + 53.33 = 100\%$ ✓

4) NH_4NO_3 :

Molar mass
$= 14+4\times1+14+16$
$= 14+4+14+16\times3$
$= 14+4+14+48$
$= 80\,g/mol$

$$\%\ of\ N = \frac{28}{80} \times 100 = 35\%$$

$$\%\ of\ H = \frac{4}{80} \times 100 = 5\%$$

$$\%\ of\ O = \frac{48}{80} \times 100 = 60\%$$

$$35\% + 5\% + 60\% = 100\%$$

Limiting and non limiting reactant:

Limiting Reactant:

"A limiting reactant is the one that produce least number of product in reaction ".

or

"The reactant that is completely consumed in a chemical reaction".

→ The amount of product form will depend upon amount of limiting reactant.

Non- Limiting Reactant:

The reactant in excess/left utilized/unreacted after reaction is completed is non-limiting.

For example:

According to balanced chemical equation

1) 2 mol of H_2 : 2 mol of H_2O
10 mol of H_2 : x

$$x = \frac{20^{10} \text{ mol}}{2 \text{ mol}} = 10 \text{ mols of } H_2O$$

answer

→ limiting reactant

2) 1 mol of O_2 : 2 mol of H_2O
10 mol of O_2 : x

$$x = \frac{10 \times 2}{1} = 20 \text{ mols of } H_2O$$

answer

→ excess reactant

Example 3.4

Zn :-

$$n = \frac{\text{mass in g}}{\text{Molar mass}} = \frac{200}{65.3} = 3.06 \text{ moles}$$

H_2SO_4 :- $n = \frac{200 \text{ g}}{98} = 2.04$ moles.

a) Zn :- Given : Find
1 mol of Zn : 1 mol of H_2

$$C.F = \frac{Find}{Given} = \frac{1}{1} = 1\ mol$$

$$Desired\ Quantity = C.F \times Given$$
$$= 1 \times 3.06$$
$$= 3.06\ moles\ of\ Zn$$

H_2SO_4 :- Given : Find

1 mol of H_2SO_4 : 1 mol of H_2

$$C.F = \frac{1}{1} = 1\ mol$$

$$Desired\ Quantity = 1 \times 2.04$$
$$= 2.04\ mols\ of\ H_2SO_4$$

b) Conversion of moles into mass

$$Mass\ of\ Hydrogen\ Produced = n \times Molar\ mass$$

$$Mass = 2.04 \times 2$$
$$= 4.08\ g$$

Concept assessment exercise 3.5:

Question number 01:

Conversion of mass into moles:

Mg:- $n = \frac{1.50}{24} = 0.06$ moles

S:- $n = \frac{1.50}{32} = 0.04$ moles

Mg:- Given : Find

1 mol of Mg : 1 mol of MgS

$C.F = \frac{1}{1} = 1$ mol

$D.Q = n \times Given$

$= 1 \times 0.06$

$= 0.06$ mol of Mg

S:- Given : Find

1 mol of S : 1 mol of MgS

$C.F = \frac{1}{1} = 1$ mole

$D.Q = n \times Given$

$= 1 \times 0.04$

$= 0.04$ moles of S

Conversion of moles into mass

Mass in g of MgS produced = n × Molar mass of MgS

$= 0.04 \times 56$

$= 2.24\,g$ Answer

Question number 02

Conversion of mass into moles

Zn:- $n = \frac{10}{65.3} = 0.15$ moles

S:- $n = \frac{10}{32} = 0.31$ moles

a) Zn:- Given : Find

1 mol of Zn : 1 mol of ZnS

$C.F = \frac{1}{1} = 1\,mol$

$D.Q = 1 \times 0.15$

$= 0.15$ moles of Zn

S:- Given : Find

1 mol of S : 1 mol of ZnS

$C.F = \frac{1}{1} = 1\ mol$

$D.Q = n \times Given$

$= 1 \times 0.31$

$= 0.31$ moles of S

Conversion of mole into mass:

Mass in g of ZnS produced $= n \times$ Molar mass

$= 0.15 \times 95.5$

$= 14.32\ g$

Question number 03

Conversion of mass into mole:

Al :- $n = \frac{54}{27} = 2\ mol$

Br_2 :- $n = \frac{200}{160} = 1.25\ mol$

a) Al :- Given : Find

2 mol of Al : 2 mol of $AlBr_3$

$C.F = \frac{2}{2} = 1$ mol

$D.Q = 1 \times 2$

$= 2$ mole of Al

Br_2:-

Given : Find

3 mole of Br_2 : 2 mole of $AlBr_3$

$C.F = \frac{Find}{Given} = \frac{2}{3} = 0.66$

$D.Q = 0.66 \times 1.25$

$= 0.82$ moles of Br

b) conversion of mole into mass

mass in g of $AlBr_3$ produced: 0.82×267

$= 218.9$ g Answer

Theoretical yield

"The amount of product as calculated from a balanced chemical equation is known as theoretical yield".

Actual yield:

" The amount of product actually produced in a chemical reaction is known as actual yield".

Percentage yield:

"Percentage Yield is a measure of how efficient a chemical reaction is. It compares the actual amount of product obtained (actual yield) to the maximum possible amount(theoretical yield) and expresses it as a percentage."

$$= \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

→ Actual yield is always less than theoretical yield due to:

1. Mechanical losses (spills, transfer errors)
2. Side reactions (unwanted byproducts)
3. Reversible reactions (incomplete conversion)
4. Impurities (in reactants)
5. Human/measurement errors

Example 3.5:

Given : Find

$CuSO_4$: Cu

1 mol of $CuSO_4$: 1 mol of Cu

$$C.F = \frac{1}{1} = 1 \text{ mol}$$

$$D.Q = 1 \times 7.9 \times 10^{-3}$$
$$= 7.9 \times 10^{-3} \text{ mol}$$

Conversion of mole into gram

$$\text{Mass in gram} = n \times \text{Molar mass}$$
$$= 7.9 \times 10^{-3} \times 63.5$$

$$\text{Theoretical Yield} = 0.5072 \text{ g}$$

$$\% \text{ Yield} = \frac{0.392}{0.5072} \times 100$$

$$= 77.28\%$$ answer

Concept exercise 3.6:

Question number 01

Given: 2 mol : CH
Actual yield : 177 g

Required: theoretical : ?
Actual : ?
% yield: ?

Given : Find

CH_4 : CCl_4

1 mol of CH_4 : 1 mol of CCl_4

$$C.F = \frac{1}{1} = 1 \text{ mol}$$

$$D.Q = 1 \times 2 = 2 \text{ mol}$$

conversion of mole to gram

$$\text{Mass in g} = 2 \times 154$$
$$= 308 \text{ g}$$

$$\% \text{ Yield} = \frac{\text{Actual}}{\text{Theoretical}} \times 100$$

$$= \frac{177}{308} \times 100$$

$$= 57.46 \%$$ Answer

Question number 02

Conversion of mass into mole

$$n = \frac{\text{mass in g}}{\text{Molar mass}} = \frac{27.9}{270.6} = 0.103$$

solution Given : Find

PBr_2 : Br_2

2 mol : 3 mol

$$C.F = \frac{3}{2} = 1.5 \text{ mol}$$

$$D.Q = 1.5 \times 0.103$$
$$= 0.15 \text{ moles}$$

Conversion of moles into mass

$$\text{Mass in g} = 0.15 \times 270.6$$
$$= 40.59 \text{ g}$$

$$\% \text{ Yield} = \frac{27.9}{40.5} \times 100$$

$$= 68.8\%$$ Answer

Determining Emperical and molecular formula:

definitions ←

"Shows simple ratio of elements in a compound"	"It shows actual number of atoms in a compound"
6 : 12 : 6 1 : 2 : 1 $= CH_6O$	

Example 3.6:

Step 1:- Calculate moles of each substance by given percentage

$$C = \frac{40\%}{12} = 3.3\,mol$$

$$H = \frac{6.7\%}{1} = 6.7\,mol$$

$$O = \frac{53.3\%}{16} = 3.3\,mol$$

Step 2:- calculate mole ratio of each element

$$\text{mole ratio} = \frac{\text{moles of Element}}{\text{Lowest value of mol}}$$

$$\text{Mole ratio of C} = \frac{3.3}{3.3} = 1$$

$$\text{Mole ratio of H} = \frac{6.7}{3.3} = 2$$

$$\text{Mole ratio of O} = \frac{3.3}{3.3} = 1$$

$= C_1H_2O_1$ Emperical Formula

Molecular formula = n x Empirical formula

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

$$= \frac{180}{30} = 6$$

Molecular formula = n x Empirical Formula

$$= 6 \times (CH_2O)$$

Molecular Formula = $C_6H_{12}O_6$ glucose.

Question 01

concept assessment 3.7:-

Step 1: Calculating moles of each element.

$$C = \frac{26.4}{12} = 2.2 \text{ mol (least value)}$$

$$H = \frac{3.3}{1} = 3.3 \text{mol}$$

$$O = \frac{70.3}{16} = 4.3 \text{mol}$$

step 2: calculating mole ratio of each element.

C : H : O

$$\frac{2.2}{2.2} : \frac{3.3}{2.2} : \frac{4.39}{2.2}$$

$$(1 : 1.5 : 2) \times 2$$ → to make whole digit

$$2 : 3 : 4$$

$C_2H_3O_2$ Empirical formula

Molecular formula = n x Empirical formula

$$n = \frac{\text{Molar mass}}{\text{Empirical formula}}$$

$$n = \frac{91}{91} = 1 \text{ mol}$$

Molecular formula $= 1 \times (C_2H_3O_2)$

$= C_2H_3O_2$ molecular formula

Question 02

Step 1: Calculating moles of each element.

$$C = \frac{18.24}{12} = 1.52 \text{ mol}$$

$$H = \frac{0.51}{1} = 0.51 \text{ mol}$$

$$F = \frac{16.9}{18.9} = 0.89 \text{ mol}$$

step 2: calculating mole ratio of each element.

$$C : H : F$$

$$\frac{1.52}{0.51} : \frac{0.51}{0.51} : \frac{0.89}{0.51}$$

$$(3 : 1 : 1.74) \times 4$$

$$12 : 4 : 7$$

$= C_{12}H_4F_7$ Answer

Molecular formula = n x empirical formula

molar mass of empirical formula $= 12 \times 12 + 4 + 19 \times 7$

$= 281$

$$n = \frac{562}{281} = 2$$

Molecular Formula $= 2 \times (C_{12}H_4F_7)$

$= C_{24}H_8F_{14}$ answer

Percentage purity:

"Amount of pure substance in a sample represented as percentage of sample's total mass"

formula: Percentage purity = $\frac{\text{actual mass of compound}}{\text{Total n}}$

Applications:

1. Medicines – Ensures correct drug dosage.
2. Chemicals – Checks quality in labs & industries.
3. Food – Tests additives & nutrients.
4. Fertilizers – Measures active ingredients.
5. Metals – Determines gold/silver purity.
6. Pollution Control – Analyzes impurities in air/water.

Purpose: Ensures safety, quality, and effectiveness in scientific & industrial uses.

Example 3.7

A student synthesized a compound and obtained a sample weighing 6.3g. Then he performed chemical analysis and found that the sample contains 4.2g of pure compound. Calculate the percentage purity of the compound.

Solution

Actual mass of compound = 4.2g

The total mass of the sample = 6.3g

Percentage purity of compound = Actual mass of compound x 100/ Total mass of sample

= 4.2 x100/6.3

= 66.67 %

Concept assesssment 3.8:

Total mass: 70g
Actual mass:56g

$$= \frac{56 \times 100}{70}$$

$= 80\%$ answer

Concentration unit:

Definition:

"The quantity of solute present in given amount of solvent/solution."

Formula:

$$\text{Concentration} = \frac{\text{Quantity of solute}}{\text{Total amount of solvent}}$$

Example: A) 5g of NaCl in 100g water → Highest concentration
B) 10g NaCl in 100g water
C) 5g of NaCl in 200g water

1) Molarity:

Definition:

"Number of moles of solute dissolved per dm^3 of solution."

Formula:

$$M = \frac{\text{no. of moles of solute}}{(\text{Volume})dm^3 \text{ of solution}}$$

Example 3.8:

Given data:

Mass in g of Urea = 40g

$$\text{Volume of Solution} = \frac{500cm^3}{1000} = 0.5dm^3.$$

Required:

Molarity = ?

Moles = ?

Formula:

$$n = \frac{\text{mass in gram}}{\text{Molar mass}}$$

$$M = \frac{\text{no of moles}}{\text{Volume in dm3}}$$

Solution:

conversion of mass into moles:

$$n = \frac{40}{60} = 0.667\text{mol}$$

putting value in formula of molarity:

$$M = \frac{0.667}{0.5} = 1.334\text{mol}$$

result:

Example 3.9:

Given Data:

- Number of moles KMno4= 0.05 moles
- Volume of solution $= \frac{600}{1000} = 0.60$ dm3

Required:

Molarity=?

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Formula:

$$M = \frac{\text{No of moles}}{\text{Volume}}$$

Solution:

$$M = \frac{0.05}{0.6} = 0.083\,M$$

Result:

The Molarity of the solution (containing 0.05 moles) with a volume of 0.60 dm^3 is 0.083 M.

Concept assessment 3.9:

a) Given:

moles of $KClO_3$ = 1.5 moles

Volume of solution = $\frac{250\ cm^3}{1000}$ = 0.25 dm^3

Required:

Molarity = ?

Formula:

$$M = \frac{\text{No of moles}}{\text{Volume}}$$

Solution:

Putting values in formula:

$$M = \frac{1.5\,mol}{0.25\,dm^3}$$

$$\boxed{M = 6M}$$

Result:

The molarity of solution containing 1.5 moles of KClO and 250cm is 6M.

b) Given data:

Mass in g = 75g

Volume of Solution = $1.25\ dm^3$

Required:

Molarity = ?

Moles = ?

Formula:

$$n = \frac{\text{mass in gram}}{\text{Molar mass}}$$

$$M = \frac{\text{no of moles}}{\text{Volume in dm3}}$$

Solution:

convert mass into moles:

$$n = \frac{75}{122.5} = 0.612\ mol$$

Molar mass

$= 39 + 35.5 + 16 \times 3$

$= 122.5\ g/mol$

calculating molarity now:

$$M = \frac{0.612}{1.25} = 0.49M$$

Result:

The Molarity of solution containing 75g of KClO and 1.25dm solution is 0.49M.

c) **Given data:**

Mass of $KClO_3$ = 0.25 g

Volume= $\frac{50\ cm^3}{1000}$ = $0.05 dm^3$

Required:

no. of moles= ?

Molarity = ?

Formulas:

$$n = \frac{\text{mass in gram}}{\text{Molar mass}}$$

$$M = \frac{\text{no of moles}}{\text{Volume in dm3}}$$

Solution:

Convert mass into moles:

$$n = \frac{0.25g}{122.5} = 0.00204\ mol$$

Calculate Molarity:

$$M = \frac{0.00204}{0.05} = 0.0408M$$

Result:

Molarity of solution containing 0.25g of $KClO_3$ and 50 cm^3 of solution is 0.0408M.

Length of solution:

Definition:

"Length of solution refers to concentration of solute in a given volume of solution".

Mathematical formula:

$$\text{Strength} = \frac{\text{Mass of Solute (g)}}{\text{Volume of Solution } (dm^3 \text{ or } cm^3)}$$

→ g/dm^3
→ g/cm^3

$1\ dm^3 = 1000\ cm^3$

♡ $\frac{cm^3}{1000} = dm^3$

♡ $dm^3 \times 1000 = cm^3$

Example 3.10:

given:

Mass of Salt = 20 gram

Volume = 2 dm^3

Required:

Strength = ?

Formula:

$$\text{Strength} = \frac{\text{Mass of Solute}}{\text{Volume}}$$

Solution:

putting value in solution:

$$\text{Strength} = \frac{20g}{2\ dm^3} = 10\ g/dm^3$$

Result:

The strength of solution is 10 g/dm^3.

Example 3.11:

Given:

Mass of Solute = 25 g

Volume of Sol = 500 cm^3

required:

Strength = ?

formula:

$$\text{Strength} = \frac{\text{Mass of Solute}}{\text{Volume}}$$

solution:

putting value in solution:

$$\text{Strength} = \frac{25g}{500cm^3} = 0.05\ g/cm^3$$

Result: The strength of solution is 0.05g/cm^3.

Example 3.12:

given:

Strength in $g/dm^3 = 50\ g/dm^3$.

Required:

Strength in $g/cm^3 = ?$

Formula:

$$1\ dm^3 = 1000\ cm^3$$

$$\frac{cm^3}{1000} = dm^3$$

$$dm^3 \times 1000 = cm^3$$

solution:

$$\frac{1g}{dm^3} = \frac{1g}{1000 \times 1} = 0.001 g/cm^3.$$

$$1\ g/dm^3 = 0.001\ g/cm^3$$

$$50\ g/dm^3 = x$$

$$x = 50 \times 0.001$$

$$= 0.05\ g/cm^3$$

result:

The strength ofsolution in g/cm3 is 0.05 g/cm^3.

Problem involving interconversion of Molarity and Strength:

Molarity ⇄ Strength
↓
interconversion

Example 3.13:

Given:

$moles_{NaOH} = 0.25M$

no. of moles = 0.25M

Volume = 1 dm^3.

Required:

Mass in grams = ?

Formula:

1) Mass in g = n x Molar mass

Molar mass of NaOH :
= 23 + 16
= 40

Solution:

Mass = 0.25 x 40

= 10 gram/dm^3 .

Result:

10 g/dm^3 mass of NaOH is present per dm^3 of solution.

Now converting g/dm into g/ cm .

$$1\,dm^3 = 1000\,cm^3$$

$$10\,g\ NaOH = 1000\,cm^3$$

$$x = 1\,cm^3$$

$$x = \frac{10 \times 1}{1000} = 0.01\ g/cm^3$$

Example 3.14:

Given:

Mass of KOH = 5.8 g

Volume = $1\,dm^3$

Required:

Molarity = M = ?

Formula:

1) $$\text{Molarity} = \frac{\text{no. of moles}}{\text{Volume in } dm^3}$$

2) $$n = \frac{\text{Mass in g}}{\text{Molar Mass}}$$

Solution:

calculating moles:

$$n = \frac{5.8\,g}{56\,g/mol}$$

$$n = 0.1\,mol$$

* KOH Molar Mass :-
= 30 + 16 + 1
= 56 g/mol.

Putting value in formula:

$$M = \frac{0.1 \text{ moles}}{1 \text{ dm}^3} = 0.1M$$

Result: Molarity of solution is 0.10 M.

Concept 3.10:

question#01

Given data:

Mass in g of NaOH = 25g

Volume of Solution = 1 dm^3

required:

Molarity = ?

formula:

$$\text{Molarity} = \frac{\text{Moles of Sol}}{\text{Volume of Sol}}$$

$$n = \frac{\text{Mass in g}}{\text{Molar mass}}$$

solution:

convert mass into moles:

$$n = \frac{25g}{40g/mol}$$

$$= 0.625 \text{ mol}$$

calculate molarity now :

$$M = \frac{0.625}{1\ dm^3} = 0.625\ g/dm^3$$

result: the molarity of solution is $0.625 g/dm^3$.

question#02:

given data:

Molarity = 1.2M

Volume in dm^3 = $1\ dm^3$.

no. of moles = 1.2 mol.

required:

Mass in g of NaOH in g/cm^3 and g/dm^3 = ?

Formula:

1) Mass in g = n × Molar mass

2) $cm^3 = \frac{\text{Mass in g}}{1000\ cm^3}$

solution:

Calculate mass in g:

$$\text{Mass in g} = 1.2\ mol \times 40\ g/mol$$
$$= 48\ g/dm^3.$$

convert into cm^3 :

$$cm^3 = \frac{48\ g}{1000\ cm^3} = 0.048\ g/cm^3.$$

result:

The mass of $NaOH_3$ having molarity is 48g/dm^3 in dm^3 and 0.048g/cm^3 in cm^3.

question # 03:

given:

Mass in g = 10g

Volume of Solution = 1dm^3.

Molar mass = 6.57×10^4.

Required:

no. of moles = ?

Molarity = ?

formulas:

1) $n = \frac{\text{Mass in g}}{\text{Molar mass}}$

2) $\text{Molarity} = \frac{\text{moles of solute}}{\text{Volume of solution}}$

solution:

convert mass into mole:

$n = \frac{10}{65100}$

$= 0.000153$ mol

Calculate molarity

$$M = \frac{0.0001538\ mol}{1\ dm^3} = 1.536 \times 10^{-4}\ M$$

result:molarity of solution containing 10g of haemoglobin dissolvedin water to make up 1dm3 of solution is 1.536×10^{-4}M.

Dilution of solutions:

Definition:

"Dilution is a process by which you make a solution less concentrated by adding more liquid(water)"

Example:

Cup of strong lemonade

add water ↓

Cup of dilute lemonade

- Only volume changes not the solute
- Dilution does not change amount of solute only spread it out in more solvent

Formula:

$$M1V2 = M2V2$$

M1= concentration of original solution
M2= concentration you want after dilution
V1= Volume of original solution
V2= Total volume of solution

Activity # 01:

Prepare 250cm of solution of 0.1M from concentrated hydrochloric acid

Given data:

M_1 = Concentration of Original sol = HCL

M_2 = Find/after dilution concentration

V_2 = Total volume of solution

required:

V_1 = Volume of Original solution.

formula:

M1V2 = M2V2

solution:

$$V_1 = \frac{M_2 \times V_2}{M_1}$$

putting values in formula:

$$V_1 = \frac{0.1M \times 250}{12 M}$$

$$V_1 = 2.08 cm^3.$$

result:

volume of original solution is 2.08 cm^3.

Example 3.15:

Given data: M_2 = Molarity needed to prepare KNO_3 =

$V_2 = 500 cm^3$

M_1 = 4M given solution.

Required:

V_1 = Volume of solution needed to Prepare KNO_3.

formula:

$$V_1 = \frac{M_2V_2}{M_1}$$

solution:

putting values in formula:

$$V_1 = \frac{1M \times 500}{4M}$$

$$= 125 cm^3.$$

Result:

Volume of solution needed to prepare KNO_3 is 125 cm^3.

Activity 3.2

To standardise the given solution of hydrochloric acid.

Note: Perform this activity in the chemistry laboratory.

You will need:

- Burette, Pipette, burette stand, beakers, conical flash, glass rod.
- Standard 0.1M NaOH solution and phenolphthalein.

Chemical equation

$$HCl_{(aq)} + NaOH_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$

n_1 = 1 mole n_2 = 1 mole

Concept 3.11:

question 01

Given:

$M_1 = 0.1M$, $M_2 = ?$

$V_1 = 9.8 cm^3$, $V_2 = 10 cm^3$

$n_1 = 1$, $n_2 = 1$

formulas:

Acid Base

$$\frac{M_1V_1}{n_1} = \frac{M_2V_2}{n_2}$$

solution:

1)

By substituting values:

$$\frac{0.1 \times 9.8}{1} = \frac{M_2 \times 10}{1}$$

$$\frac{0.1 \times 9.8}{10} = M_2$$

$$M_2 = 0.098\,M$$

2)

We know that 1dm3 = 1000cm3 amount per NaOH per cm3:

$$= \frac{0.098}{1000}$$

$$= 9.8 \times 10^{5} \text{ mole per cm}^3$$

question 02:

given:

$$M_1V_1 = M_2V_2$$

$$M_1 = 6M \quad , \quad M_2 = 2M$$

$$V_1 = 200cm^3 , V_2 = ?$$

solution:

$$6M \times 200cm^3 = 2 \times V_2$$

$$\frac{6 \times 200}{2} = V_2$$

$$V_2 = 600cm^3$$

amount of water added = final volume – initial volume

$$= 600cm^3 - 200cm^3$$

$$= 400\ cm^3$$

Result:
The chemical should min 400 cm3 of water with the original solution to achieve the desired concentration of 2M

SLO questions:

1. Define Stoichiometry.
Answer:
The calculation of reactants and products in chemical reactions using balanced equations, based on the law of conservation of mass.

2. What is a Mole?
Answer:
A mole is the amount of substance containing 6.022×10^{23} particles (Avogadro's number).

3. State the Law of Conservation of Mass.
Answer
Mass is neither created nor destroyed in a chemical reaction (total mass of reactants = total mass of products).

4. What is a Limiting Reagent?
Answer:
The reactant that gets completely consumed first, stopping the reaction and limiting the product formed.

5. What is Percentage Yield?
Answer:

$$\text{Percentage Yield} = \left(\frac{\text{Actual Yield}}{\text{Theoretical Yield}}\right) \times 100$$

Measures reaction efficiency.

6. What is STP? Give values.
Answer:
Standard Temperature and Pressure (STP):
- Temperature = 273 K (0°C)
- Pressure = 1 atm (760 mmHg)
- 1 mole of gas occupies 22.4 L at STP.

7. What is Avogadro's Law?
Answer
Equal volumes of gases at same temperature and pressure contain equal number of molecules.

8. Why Must Chemical Equations Be Balanced?
Answer:
To obey the law of conservation of mass (same number of atoms of each element on both sides).

11. What is Molar Mass?
Answer:
Mass of 1 mole of a substance (units: g/mol). Example: Molar mass of H_2O = 18 g/mol.

12. Differentiate between empirical formula and molecular formula.

Empirical Formula	Molecular Formula
Simplest whole-number ratio of atoms in a compound.	Exact number of atoms of each element in a molecule.
Does not give the actual number of atoms.	Represents the true molecular structure.
Example: CH_2O (for glucose)	Example: $C_6H_{12}O_6$ (actual formula of glucose)
Can be the same for different compounds (e.g., CH_2 for ethene and cyclopropane).	Unique to a specific compound.

13. Differentiate between actual yield and theoretical yield.

Actual Yield	Theoretical Yield
Amount of product obtained in a real experiment.	Maximum possible amount of product calculated from stoichiometry.
Always less than or equal to theoretical yield.	Ideal yield assuming 100%efficiency.
Affected by experimental errors,side reactions,and impurities.	Based on balanced chemical equations and limiting reactants.
Measured in grams or moles.	Calculated in grams or moles.

14. Differentiate between dilute solutions and concentrated solutions.

Dilute Solution	Concentrated Solution
Contains a small amount of solute relative to solvent.	Contains a large amount of solute relative to solvent.
Low concentration(e.g.,0.1 M HCl).	High concentration(e.g.,12 M HCl).
Less solute particles per unit volume.	More solute particles per unit volume.
Often used in titrations or sensitive reactions.	Often used as stock solutions in labs.

Exercise questions:

EXERCISE

1. Multiple Choice Questions (MCQs)

i. What mass of CO_2 is produced by the combustion of each mol of CH_4?

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(l)}$$

(a) 16g (b) 44g
(c) 32g (d) none

ii. A flask contains 500 cm^3 of SO_2 at RTP. The flask contains;

(a) 32 g (b) 2.4 g
(c) 64 g (d) 1.33g

iii. A necklace has 6g of diamond in it. How many carbon atoms are present in it?

(a) 1 (b) 0.5
(c) 1.5 (d) 0.25

iv. What is the mass of aluminium in 204 g of the aluminium oxide, Al_2O_3?

(a) 26 g (b) 27 g
(c) 54 g (d) 108 g

v. The reactant which is consumed earlier and gives least quantity of product is called;

(a) Reactant in excess (b) Stoichiometry
(c) Limiting reactant (d) Stoichiometric amount

vi. Which one of the following compounds contains the highest percentage by mass of nitrogen?

(a) NH_3 (b) N_2H_4
(c) NO (d) NH_4OH

vii. Vitamin A has a molecular formula of $C_{20}H_{30}O$. The number of moles of vitamin A in 500 mg of its capsule will be;

(a) 1.7 (b) 1.7×10^{-3}
(c) 1.05 (d) 3.01×10^{-3}

viii. When one mole of each of the following is completely burnt in oxygen, which will give the largest mass of CO_2?

(a) Carbon monoxide (b) Diamond
(c) Ethane (d) Methane

ix. One mole of ethanol and one mole of ethane have an equal;

(a) Mass (b) Number of atoms
(c) Number of electrons (d) Number of molecules

x. How many moles of oxygen are needed for the complete combustion of two moles of butane, C_4H_{10}?

$$2C_4H_{10(g)} + 13O_{2(g)} \longrightarrow 8CO_{2(g)} + 10H_2O_{(l)}$$

(a) 12 mol (b) 13 mol
(c) 4 mol (d) 10 mol

short answers:

question # 01

Solution:

Calculating moles of elements:

$$C = \frac{65.45}{12} = 5.454 \text{ moles}$$

$$H = \frac{5.45}{1} = 5.45 \text{ moles}$$

$$O = \frac{29.09}{16} = 1.818 \text{ moles}$$

calculating mole ratio of each elements:

$$\text{mol ratio of } C = \frac{5.454}{1.818} = 3$$

$$\text{mol ratio of } H = \frac{5.45}{1.818} = 3$$

$$\text{mol ratio of } O = \frac{1.818}{1.818} = 1$$

$\{ C_3H_3O$ empirical formula $\}$

$$n = \frac{\text{molar mass}}{\text{Empirical formula}} = \frac{110}{55} = 2$$

$$\text{Molecular formula} = (C_3H_3O) \times 2$$

$$= C_6H_6O_2$$

Question#02

Solution:

Mass of $AlCl_3 = 125g$

Molar mass $= 27 + (3 \times 35.5) = 27 + 106.5$

$= 133.5 \, g/mol$

Mols of $AlCl_3 = \frac{125}{133.5} = 0.937$ moles

According to balanced chemical equation:

2 moles of $AlCl_3$ = 6 mol of HCl

1 mol of $AlCl_3 = \frac{6}{2}$ mol of HCl

0.937 mol = x

$$x = \frac{0.937 \times 6}{2}$$

x = 2.81 moles of HCl

Result:

number of moles of HCl prodeced is 2.81 moles

question # 03

solution:

Mass of $Ca(NO_3)_2$ = 1 g

Molar mass = $40.08 + (14.01 \times 2) + (16 \times 6)$

$= 40.08 + 28.02 + 96$

$= 164.10$ g/mol

moles of $Ca(NO_3)_2 = \frac{Mass}{Molar\ mass} = \frac{1.0}{164.10} = 0.00609$ mol

According to balanced chemical equation:

1 mol of $Ca(NO_3)_2$ = 3 moles of O_2

0.00609 mol of $Ca(NO_3)_2$ = x

$x = 0.00609 \times 3$

$x = 0.0183$ moles of O_2

Result:

number of moles of oxygen needed is 0.0183 moles.

question # 04

Solution:

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Mass of Al = 25g

Molar mass of Al = 27

$$\text{Moles of Al} = \frac{\text{mass}}{\text{Molar}} = \frac{25}{27} = 0.926 \text{ moles}$$

$\text{Mass of S} = 25g$

$\text{Molar mass (S)} = 32.06$

$$\text{Moles of S} = \frac{25}{32} = 0.78125 \text{ moles}$$

according to balanced chemical equations:

For Al:

$2 \text{ mol of Al} = 1 \text{ mol of } Al_2S_3$

$1 \text{ mol of Al} = \frac{1}{2} \text{ mol of } Al_2S_3$

$0.926 \text{ moles of Al} = x$

$$x = \frac{0.926}{2} = 0.463 \text{ moles of } Al_2S_3$$

For S=

$3 \text{ mol of S} = 1 \text{ mol of } Al_2S_3$

$1 \text{ mol of S} = \frac{1}{3} \text{ mol of } Al_2S_3$

$0.78125 \text{ mol of S} = x$

$$x = \frac{0.78125}{3} = 0.260 \text{ moles of } Al_2S_3$$

result:

As,S is producing lesser moles of product,so,S is limiting reactant.

Question #05

solution:

Calculating moles of elements:

$$Na = \frac{20.2}{23} = 0.878 \text{ mol}$$

$$S = \frac{37.6}{32} = 1.175 \text{ mol}$$

$$O = \frac{42.2}{16} = 2.638 \text{ mol}$$

calculating mole ratio of each elements:

$$\text{Mole ratio of Na} = \frac{0.878}{0.878} = 1$$

$$\text{Mole ratio of S} = \frac{1.175}{0.878} = 1.338$$

$$\text{Mole ratio of O} = \frac{2.638}{0.878} = 3.005$$

$\times 3$ to make whole digit $= 3, 4, 9$

{ $Na_3S_4O_9$ Empirical formula }

Molar mass of $Na_3O_9S_4 = 341$

$$n = \frac{\text{Molar mass}}{\text{Empirical formula}} = \frac{682.8}{341} = 2$$

therefore ,molecular formula: $= 2(Na_3S_4O_9)$

$$= Na_6S_8O_{18}$$

Question #06

Solution:

Mass of $CaCO_3$ = 68.1 g

Molar mass = 40 + 12 + (3×16) = 100 g/mol

Moles of $CaCO_3$ = $\frac{\text{Mass}}{\text{Molar}}$ = $\frac{68.1}{100}$ = 0.681 mol

Mass of HCl = 57.6 g

Molar mass = 1 + 35 = 36

Moles of HCl = $\frac{57.6}{36}$ = 1.433 mol

According to balanced chemical equation:

1 mol of $CaCO_3$ = 2 mol of HCl

0.681 mol of $CaCO_3$ = x

$x = 0.681 \times 2$

= 1.362 moles of HCl

so, HCl is in excess

limiting reactant is CaCO3.

Question#07

given:

actual yield = 1.85 g
Theoretical yield = 2.5 g

required:

percentage yield= ?

formula:

$$\text{Percentage Yield} = \frac{\text{Actual Y}}{\text{Theoretical Y}} \times 100$$

solution:

Putting values in formula:

$$P.Y = \frac{1.85}{2.5} \times 100$$

$$= 74\%$$

Result: percentage yield is 74%.

question#08

solution:

At room temp and pressure ,1 mole of any gas occupies 24dm3

so,

$$\text{Moles of } CO_2 = \frac{\text{Volume at RTP}}{24} = \frac{24}{24} = 1 \text{ mol}$$

according to balanced chemical equation:

$$1 \text{ mol of } CaCO_3 = 1 \text{ mol of } CO_2$$

so, if 1 mol of CO2 is produced than 1 mol of CaCO3 was heated.

$$\text{Mass of } CaCO_3 = \text{Mass} \times \text{Molar mass}$$
$$= 1 \times 100$$
$$= 100 \text{ g}$$

So, 100 g of limestone was heated.

Question#09

Solution: a) moles of CO2 produced:-

According to balanced chemical equation:

$$3 \text{ mol of } CO_2 \text{ produced} = 2 \text{ moles of } O_2$$
$$1 \text{ mol of } O_2 = \frac{2}{3} \text{ mols of } CO_2$$
$$1.08 \text{ moles of } O_2 = x$$
$$x = \frac{1.08 \times 2}{3}$$

$$= 0.72 \text{ Moles of } CO_2$$

Solution:-b) reactants left:

According to balanced chemical equation:

1 mole of C_2H_4 = 3 moles of O_2

0.480 mole of $O_2C_2H_4$ = x

$x = 0.480 \times 3$

$= 1.44$ moles of O_2

but,we only have 1.08 moles of O2 ,which is less than required.

so, O2 is limiting reactant and it will be completely used.

Since, C2H4 is in excess,so some will remain

Hence, C2H4 is the reactant left unreacted.

Question # 10

Solution:

Mass of CH_4 = 48g

Molar mass of CH_4 = $12 + (1 \times 4) = 16$ g/mol

Moles of CH_4 = $\frac{\text{mass}}{\text{molar}} = \frac{48}{16} = 3$ mol

according to balanced chemical equation:

1 mol of CH_4 = 1 mol of CO_2

3 mol of CH_4 = x

$x = 3$ moles of CO_2.

At RTP ,1 mole of any gas occupies 24dm3.

Volume of CO_2 = 3 mole × 24 dm^3 mole
= 72 dm^3

result:- Hence, 72dm3 carbon dioxide gas is produced.

Question # 11

solution:

$C = \frac{40}{12} = 3.33 mol$

$H = \frac{6.7}{10} = 6.7 mol$

$O = \frac{53.3}{10} = 3.33 mol$

According to balanced chemical equation:

Mole ratio of C = $\frac{3.33}{3.33} = 1$

Mole ratio of H = $\frac{6.7}{3.3} = 2$

Mole ratio of O = $\frac{3.33}{3.33} = 1$

{ CH_2O = empirical formula }

Question 12

Solution: a) Concentration in g/cm^3

Mass of Sodium Chloride (NaCl) = 10.0 g

Volume of Solution = $500 cm^3 = \frac{500}{1000} = 0.5 dm^3$

Concentration of solution $g/dm^3 = \frac{\text{mass of solute}}{\text{Volume of Solution}}$

$= \frac{10.0}{0.5} = 20 g/dm^3$

b) Concentration in mol/dm^3 :-

Mass of Sodium Chloride = 10 g

Molar mass = 23 + 35 = 58 g/mol.

Moles of NaCl = $\frac{10.0}{58} = 0.172$ mol

Concentration of solution $mol/dm^3 = \frac{\text{moles}}{\text{Volume}}$

$= \frac{0.172}{0.5} = 0.344 \ mol/dm^3$

question no. 13

Solution:

Volume of NaOH = $25 cm^3 = \frac{25}{1000} = 0.0250 dm^3$.

Concentration of NaOH = $0.10 \ mol/dm^3$.

Volume of HCl = $\frac{20}{1000} = 0.200 dm^3$.

Moles of NaOH = Concentration × Volume

$$= 0.10 \times 0.0250$$

$$= 0.00250 \text{ mol}$$

According to balanced chemical equation:

1 mol of NaOH = 1 mol of HCl

0.00250 mol of NaOH = x

x = 0.00250 moles of HCl

$$\text{Concentration of HCl} = \frac{\text{moles}}{\text{Volume}} = \frac{0.00250}{0.0200} = 0.125 \text{ mol/dm}^3$$

Hence, concentration of HCl is 0.125 mol/dm3。

Question # 14

a)

Concentration is the amount of solution (in this case s sodium chloride – NaCl) dissolved in a specific volume of solvent (water), typically expressed as mol.
Yes, cubic decimates (mol/dm^3).

In this case, 0.5 mol/dm^3 NaCl solution means 0.5 moles of NaCl are dissolved in 1 dm^3 (1000 cm^3) of water.

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b) solution:

Volume of NaCl $= \frac{250}{100} = 0.250\ dm^3$

Concentration of NaCl $= 0.5\ mol/dm^3$.

Moles of NaCl = Concentration × Volume

$= 0.5 \times 0.250$

$= 0.125$ mole

mass of NaCl = Moles × Molar mass

$= 0.125 \times 58$

$= 7.25\ g$

c) solution:

Concentration in $(mol/dm^3) = 0.5\ mol/dm^3$

Molar mass of NaCl $= 23 + 35 = 58$

Concentration in (g/dm^3) = Concentration × Molar mass

$= 0.5 \times 58$

$= 29\ g/dm^3$

1. Calculate required mass: As shown above, 7.25 g of NaCl is needed.

2. Weigh the solute: Accurately weigh 1.25 g of NaCl using an electronic balance.

3. Dissolve the solute:
   - Transfer the NaCl into a 250 cm^3 volumetric flask (ex bracket first).
   - Add about 100 – 150 cm^3 of distilled water and swirl until fully dissolved.
   - Transfer to volumetric flask (if wet already). Rinse any remaining solute from the beaker into the flask using distilled water.

4. Make up volume:
   - Add distilled water until the bottom of the meniscus reaches the 250 cm^3 calibration line.

5. Mix thoroughly: Stopper the flask and invert several times to ensure uniform mixing.