

Unit 3: Stoichiometry

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Class:10

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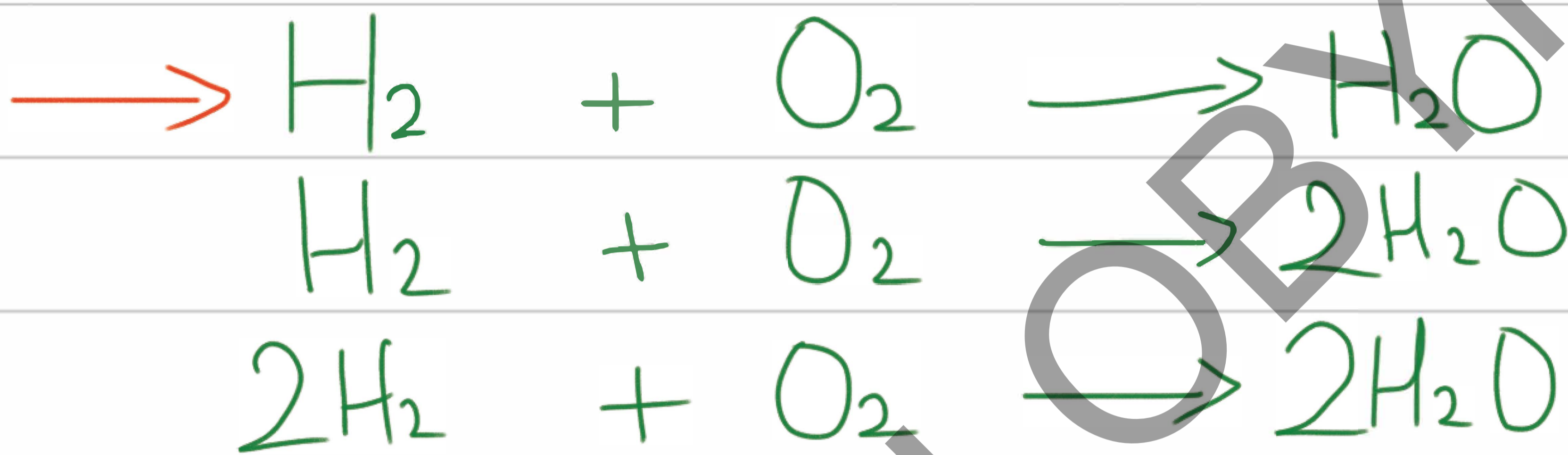
"The calculation of reactant and products in a balanced chemical reaction".

or

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

Example:

Stoichen=Element
Metron=Measure



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: 6.022×10^{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)

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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:



2 mol of hydrogen + 1 mol of oxygen \longrightarrow 2 mol of water

2kg chicken + 1kg rice \longrightarrow 2kg biryani

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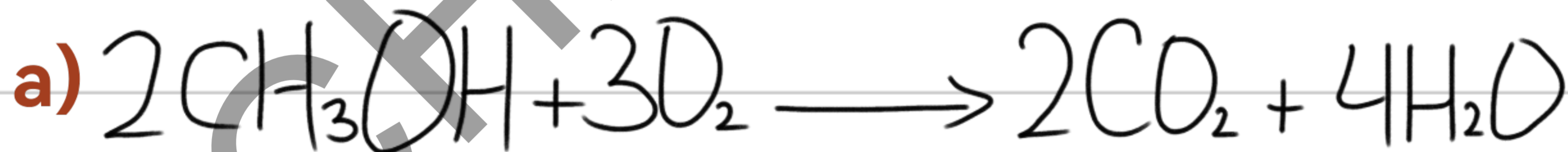
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Explanation of mole ratio.

SI amount of substance: mole
Symbolic representation: mol

Example 3.1: (Methanol burns according to the following)



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



$$C.F = \frac{F_{ind}}{G_{iven}} = \frac{3}{2} = 1.5 \text{ moles}$$

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

$$\begin{aligned} &= \text{Desired Quantity} = \text{Given} \times \text{C.F} \\ &= 3.5 \text{ mol} \times 1.5 \\ &= 5.25 \text{ moles} \end{aligned} \quad \text{answer}$$

b) How many moles of water are produced?

step 1 Given : Find
 $2\text{CH}_3\text{OH} : 4\text{H}_2\text{O}$

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

Step 2 :- Desired Quantity = C.F \times Given

$$\begin{aligned} &= 2 \text{ mol} \times 3.50 \\ &= 7 \text{ moles} \end{aligned} \quad \text{answer}$$

homework

Part a)

Concept assessment 3.1

Step 1

Given : Find
2 : 1

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

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step 2

$$\begin{aligned} \text{Desired Quantity} &= \text{Given} \times \text{C.F} \\ &= 8.0 \times 0.5 \\ &= 4 \text{ moles of } \text{N}_2 \end{aligned} \quad \text{Answer}$$

Part b)

step 1:

Given : Find
2 : 3

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

step 2:

$$\begin{aligned} \text{Derived Quantity} &= \text{Given} \times C.F \\ &= 8.0 \times 1.5 \\ &= 12 \text{ moles of } H_2 \text{ Answer-} \end{aligned}$$

Example 3.2

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SOLUTION:

Given : Find

Fe_2O_3 : Fe

1 mol : 2 mol

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

rough

$$\begin{aligned} Fe_2O_3 \\ &= 55.8 \times 2 + 16 \times 3 \\ &= 111.6 + 48 \\ &= 159.6 \end{aligned}$$

Conversion of Mass in g to moles-

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

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

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

$$= 5.32 \times 55.9$$

$$= 297.3 \text{ grams}$$

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Final answer

Important Formulas

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

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

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

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

Concept assessment 3.2

Solution:

Given : Find

H_2 : O_2
2 : 1

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

$$\begin{aligned} \text{Desire Quantity} &= C.F \times \text{Given} \\ &= 0.5 \times 50 \\ &= 25 \text{ moles} \end{aligned}$$

$$\begin{aligned} n &= \frac{100}{2} \\ &= 50 \end{aligned}$$

$$\begin{aligned} \text{Mass in g} &= n \times \text{Molar mass} \\ &= 25 \times 32 \\ &= 800 \text{ grams} \end{aligned}$$

$$\begin{aligned} \text{O}_2 &= 16 \times 2 \\ &= 32 \end{aligned}$$

Molar volume

1)

$$1 \text{ mol} = 24 \text{ dm}^3$$

$$1 \text{ mol of } \text{H}_2 \text{ RTP} = 24 \text{ dm}^3$$

$$1 \text{ mol of } \text{NH}_3 \text{ RTP} = 24 \text{ dm}^3$$

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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

$$1 \text{ mol of } \text{Cl}_2 = 24 \text{ dm}^3$$

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

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

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Concept assessment 3.3

a) $1 \text{ mol of } \text{O}_2 = 24 \text{ dm}^3$

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

$$x$$

$$x = \frac{50}{24}$$

$$= 2.083 \text{ mol}$$

Answer

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b) $1 \text{ mol of } \text{N}_2 = 24 \text{ dm}^3$

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

$$x = 0.80 \times 24$$

$$= 19.2 \text{ dm}^3$$

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:

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

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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 :

$$\% \text{ of H} = \frac{2}{18} \times 100 = 11.11\%$$

$$\% \text{ of O} = \frac{16}{18} \times 100 = 88.88\%$$

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

$$\begin{aligned}\text{Molar mass} &= 1 \times 2 + 16 \\ &= 2 + 16 \\ &= 18 \text{ g/mol}\end{aligned}$$

2) H_2SO_4 :

$$\% \text{ of H} = \frac{2}{98} \times 100 = 2.04\%$$

$$\% \text{ of S} = \frac{32}{98} \times 100 = 32.65\%$$

$$\% \text{ of O} = \frac{64}{98} \times 100 = 65.30\%$$

$$65.30 + 2.04 + 32.65 = 100\% \checkmark$$

$$\begin{aligned}\text{Molar Mass} &= 1 \times 2 + 32 + 16 \times 4 \\ &= 2 + 32 + 64 \\ &= 34 + 64 \\ &= 98 \text{ g/mol-}\end{aligned}$$

3) $\text{C}_6\text{H}_{12}\text{O}_6$:

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$$\% \text{ of C} = \frac{72}{180} \times 100 = 40\%$$

$$\% \text{ of H} = \frac{12}{180} \times 100 = 6.66\%$$

$$\% \text{ of O} = \frac{96}{180} \times 100 = 53.33\%$$

$$40 + 6.66 + 53.33 = 100\% \checkmark$$

$$\begin{aligned}\text{Molar Mass} &= 12 \times 6 + 12 + 16 \times 6 \\ &= 72 + 12 + 96 \\ &= 180 \text{ g/mol-}\end{aligned}$$

4) NH_4NO_3 :

$$\begin{aligned}\text{Molar mass} &= 14 + 4 \times 1 + 14 + 16 \\ &= 14 + 4 + 14 + 16 \times 3 \\ &= 14 + 4 + 14 + 48 \\ &= 80 \text{ g/mol}\end{aligned}$$

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

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

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

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

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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:

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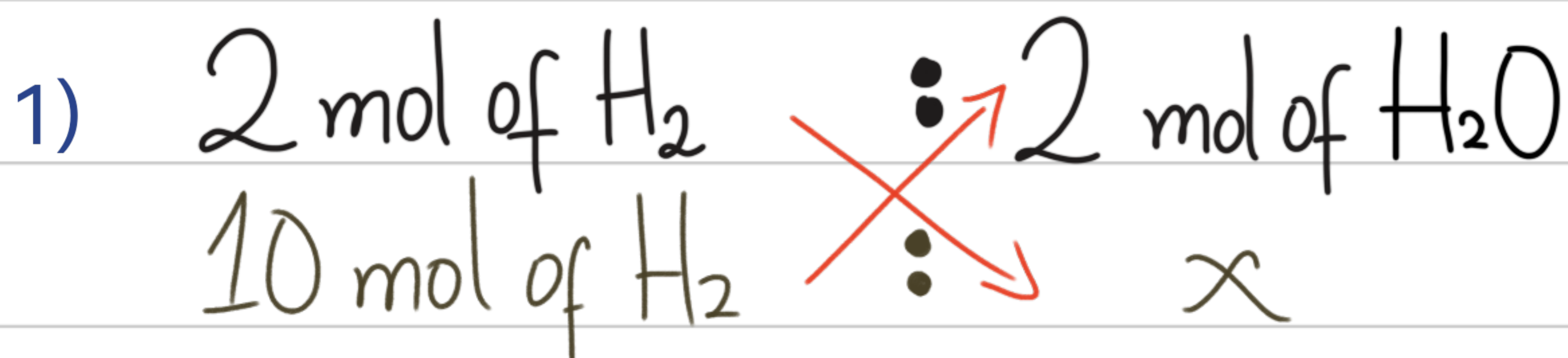
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The reactant in excess/left utilized/unreacted after reaction is completed is non-limiting.

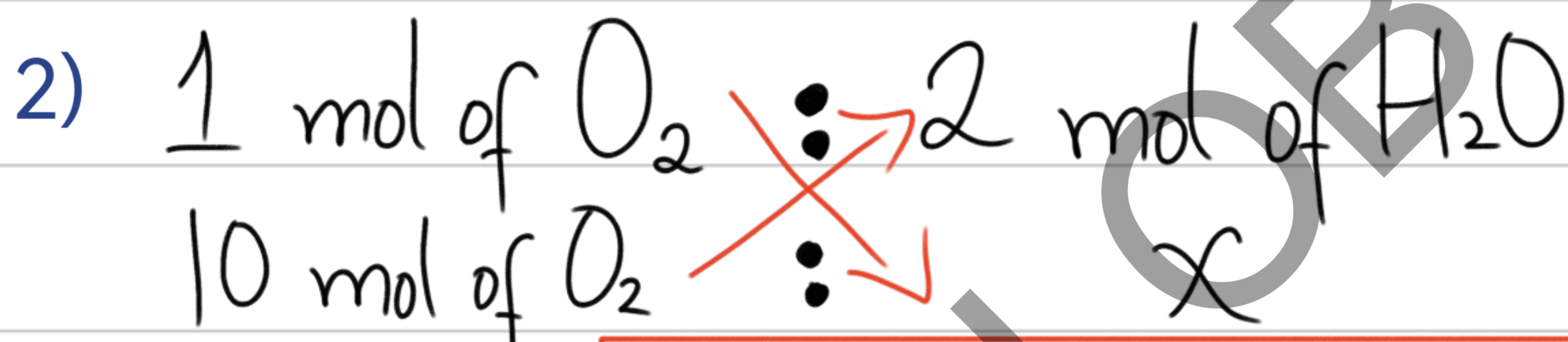
For example:

According to balanced chemical equation



$$x = \frac{20^{10} \text{ mol}}{2 \text{ mol}} = 10 \text{ mols of H}_2\text{O} \quad \text{answer}$$

limiting reactant



$$x = \frac{10 \times 2}{1} = 20 \text{ mols of H}_2\text{O} \quad \text{answer}$$

excess reactant

Example 3.4

Zn:-

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

H₂SO₄:- $n = \frac{200 \text{ g}}{98} = 2.04 \text{ moles}$

a) **Zn:-** Given : Find



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

$$\begin{aligned}\text{Desired Quantity} &= C.F \times \text{Given} \\ &= 1 \times 3.06 \\ &= 3.06 \text{ moles of Zn}\end{aligned}$$

H_2SO_4 :- Given : Find
1 mol of H_2SO_4 : 1 mol of H_2

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

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$$\begin{aligned}\text{Desired Quantity} &= 1 \times 2.04 \\ &= 2.04 \text{ mols of } \text{H}_2\text{SO}_4\end{aligned}$$

b) Conversion of moles into mass

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

$$\begin{aligned}\text{Mass} &= 2.04 \times 2 \\ &= 4.08 \text{ g}\end{aligned}$$

Concept assessment exercise 3.5:

Question number 01:

Conversion of mass into moles:

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

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

Mg:- Given : Find

1 mol of Mg : 1 mol of MgS

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

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$$\begin{aligned} \text{D.Q} &= n \times \text{Given} \\ &= 1 \times 0.06 \\ &= 0.06 \text{ mol of Mg} \end{aligned}$$

S:-

Given : Find

1 mol of S : 1 mol of MgS

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

$$\begin{aligned} \text{D.Q} &= n \times \text{Given} \\ &= 1 \times 0.04 \\ &= 0.04 \text{ moles of S} \end{aligned}$$

Conversion of moles into mass

$$\begin{aligned} \text{Mass in g of MgS produced} &= n \times \text{Molar mass of MgS} \\ &= 0.04 \times 56 \\ &= 2.24 \text{ g} \end{aligned}$$

Answer

Question number 02

Conversion of mass into moles

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

S:-

$$n = \frac{10}{32} = 0.31 \text{ moles}$$

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

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

$$\begin{aligned} \text{D.O} &= 1 \times 0.15 \\ &= 0.15 \text{ moles of Zn} \end{aligned}$$

S:- Given : Find
1 mol of S : 1 mol of ZnS

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

$$\begin{aligned} D.Q &= n \times \text{Given} \\ &= 1 \times 0.31 \\ &= \underline{0.31 \text{ moles of S}} \end{aligned}$$

Conversion of mole into mass:

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

$$= 0.15 \times 95.5$$

$$= \underline{14.32 \text{ g}}$$

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Question number 03

Conversion of mass into mole:

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

$$\text{Br}_2:- n = \frac{200}{160} = 1.25 \text{ mol}$$

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a) **Al**:- Given : Find

2 mol of Al : 2 mol of AlBr_3

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

$$D.Q = 1 \times 2 \\ = 2 \text{ mole of Al}$$

Br_2 :- Given : Find

3 mole of Br_2 : 2 mole of $AlBr_3$

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

$$D.Q = 0.66 \times 1.25 \\ = 0.82 \text{ moles of Br}$$

b) conversion of mole into mass

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

$$= 218.9 \text{ g}$$

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."

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

→ 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

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

$$\begin{aligned} D.O &= 1 \times 7.9 \times 10^{-3} \\ &= 7.9 \times 10^{-3} \text{ mol} \end{aligned}$$

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Conversion of mole into gram

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

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

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

$$= 77.28\% \text{ answer}$$

Concept exercise 3.6:

Question number 01

Given: 2 mol : CH₄
Actual yield : 177 g

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

Given : Find

CH₄ : CCl₄

1 mol of CH₄ : 1 mol of CCl₄

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

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

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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
PBr₂ : Br₂
2 mol : 3 mol

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

$$\begin{aligned} \text{D.Q} &= 1.5 \times 0.103 \\ &= 0.15 \text{ moles} \end{aligned}$$

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Conversion of moles into mass

$$\begin{aligned} \text{Mass in g} &= 0.15 \times 270.6 \\ &= 40.59 \text{ g} \end{aligned}$$

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

$$= 68.8\%$$

Answer

Determining Empirical and molecular formula:

definitions



"Shows simple ratio of elements in a compound"

6 : 12 : 6

1 : 2 : 1

=CH₂O

"It shows actual number of atoms in a compound"

Example 3.6:

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

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

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

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

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Step 2:- calculate mole ratio of each element

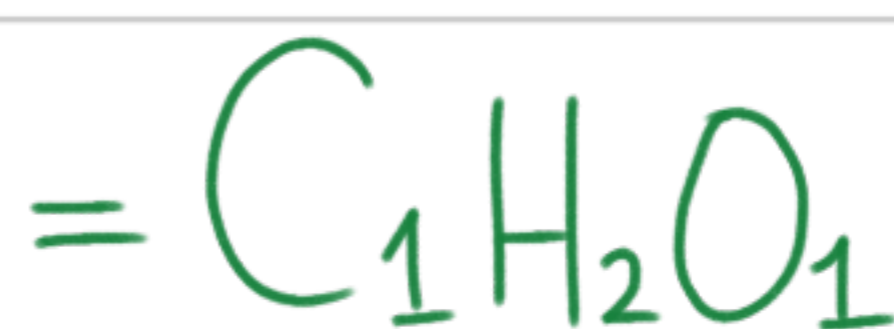
mole ratio = moles of Element

Lowest value of mol

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

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

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



Empirical
Formula

Molecular formula = $n \times$ Empirical formula

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

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

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$$\begin{aligned}\text{Molecular formula} &= n \times \text{Empirical Formula} \\ &= 6 \times (\text{CH}_2\text{O})\end{aligned}$$

Molecular Formula = $\text{C}_6\text{H}_{12}\text{O}_6$

glucose.

Question 01

concept assessment 3.7:-

Step 1: Calculating moles of each element.

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

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

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

step 2: calculating mole ratio of each element.

C : H : O

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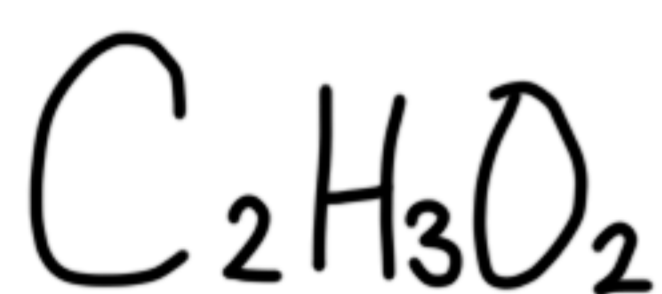
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$$\frac{2.2}{2.2} : \frac{3.3}{2.2} : \frac{4.39}{2.2}$$

$$(1 : 1.5 : 2) \times 2 \rightarrow \text{to make whole digit}$$

$$2 : 3 : 4$$



Empirical formula

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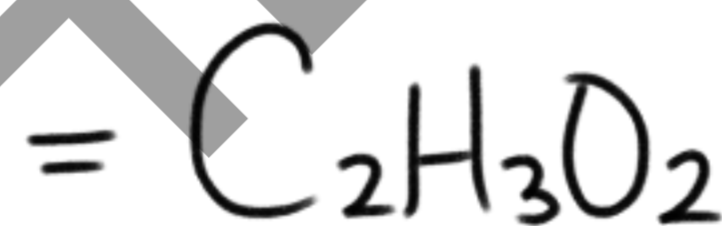
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Molecular formula = n x Empirical formula

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

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

$$\text{Molecular formula} = 1 \times (\text{C}_2\text{H}_3\text{O}_2)$$



molecular formula

Question 02

Step 1: Calculating moles of each element.

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

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

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

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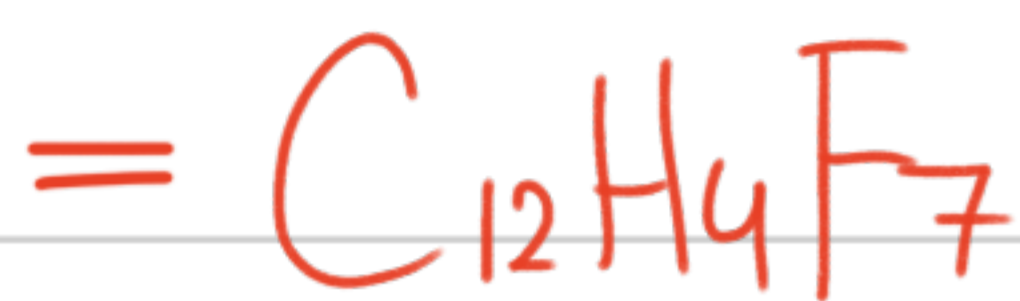
step 2: calculating mole ratio of each element.



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

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

$$12 : 4 : 7$$



Answer

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Molecular formula = n x empirical formula

$$\begin{aligned} \text{molar mass of empirical formula} &= 12 \times 12 + 4 + 19 \times 7 \\ &= 281 \end{aligned}$$

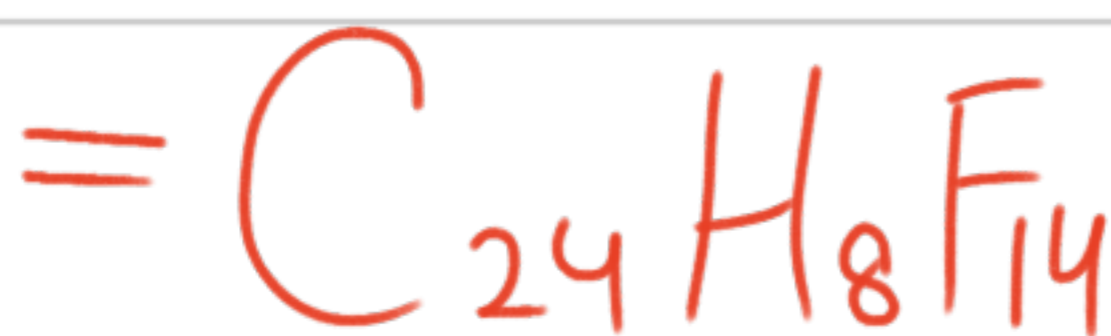
$$n = \frac{562}{281} = 2$$

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



answer

Percentage purity:

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

formula:
$$\text{Percentage purity} = \frac{\text{actual mass of compound}}{\text{Total mass of sample}} \times 100$$

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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 = $\frac{\text{Actual mass of compound} \times 100}{\text{Total mass of sample}}$
= $\frac{4.2 \times 100}{6.3}$
= 66.67 %

Concept assessment 3.8:

Total mass: 70g
Actual mass: 56g

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

$$= 80\%$$

answer

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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:

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

Example 3.8:

Given data:

Mass in g of Urea = 40g

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

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Required:

Molarity = ?

Moles = ?

Formula:

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

$$M = \frac{\text{no of moles}}{\text{Volume in dm}^3}$$

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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 $\text{KMnO}_4 = 0.05$ moles
- Volume of solution $= \frac{600}{1000} = 0.60 \text{ dm}^3$

Required:

Molarity=?

Formula:

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

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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 dm³ is 0.083 M.

Concept assessment 3.9:

a) Given:

moles of KClO₃ = 1.5 moles

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

Required:

Molarity = ?

Formula:

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

Solution:

Putting values in formula:

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

$$M = 6M$$

Result:

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

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b) Given data:

$$\text{Mass in g} = 75\text{g}$$

$$\text{Volume of Solution} = 1.25 \text{ dm}^3$$

Required:

$$\text{Molarity} = ?$$

$$\text{Moles} = ?$$

Formula:

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

$$M = \frac{\text{no of moles}}{\text{Volume in dm}^3}$$

Solution:

convert mass into moles:

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

$$\begin{aligned} \text{Molar mass} \\ &= 39 + 35.5 + 16 \times 3 \\ &= 122.5 \text{ g/mol} \end{aligned}$$

calculating molarity now:

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

Result:

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

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c) Given data:

Mass of $\text{KClO}_3 = 0.25 \text{ g}$

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

Required:

no. of moles = ?

Molarity = ?

Formulas:

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

$$M = \frac{\text{no of moles}}{\text{Volume in dm}^3}$$

Solution:

Convert mass into moles:

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

Calculate Molarity:

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

Result:

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

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Length of solution:

Definition:

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

Mathematical formula:

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

→ g/dm^3
→ g/cm^3

$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

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

♥ $\text{dm}^3 \times 1000 = \text{cm}^3$

Example 3.10:

given:

$$\text{Mass of Salt} = 20 \text{ gram}$$

$$\text{Volume} = 2 \text{ dm}^3$$

Required:

$$\text{Strength} = ?$$

Formula:

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

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Solution:

putting value in solution:

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

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Result:

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

Example 3.11:

Given:

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

$$\text{Volume of Sol} = 500 \text{ cm}^3$$

required:

$$\text{Strength} = ?$$

formula:

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

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solution:

putting value in solution:

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

Result:

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

Example 3.12:

given:

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

Required:

$$\text{Strength in g/cm}^3 = ?$$

Formula:

$$1 \text{ dm}^3 = 1000 \text{ cm}^3.$$

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

$$\text{dm}^3 \times 1000 = \text{cm}^3$$

solution:

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

$$1 \text{ g/dm}^3 = 0.001 \text{ g/cm}^3$$

$$50 \text{ g/dm}^3 = x$$

$$x = 50 \times 0.001$$

$$= 0.05 \text{ g/cm}^3$$

result:

The strength of solution in g/cm³ is 0.05 g/cm³.

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Problem involving interconversion of Molarity and Strength:

Molarity \longleftrightarrow Strength
↓
interconversion

Example 3.13:

Given:

$$\text{moles}_{\text{NaOH}} = 0.25\text{M}$$

$$\text{no. of moles} = 0.25\text{M}$$

$$\text{Volume} = 1 \text{ dm}^3$$

Required:

$$\text{Mass in grams} = ?$$

Formula:

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

Solution:

$$\text{Mass} = 0.25 \times 40$$

$$= 10 \text{ gram/dm}^3$$

$$\begin{aligned} \text{Molar mass of NaOH:} \\ &= 23 + 16 \\ &= 40 \end{aligned}$$

Result:

10 g/dm³ mass of NaOH is present per dm³ of solution.

Now converting g/dm into g/ cm .

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$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

$$10 \text{ g NaOH} = 1000 \text{ cm}^3$$

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

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

Example 3.14:

Given:

$$\text{Mass of KOH} = 5.8 \text{ g}$$

$$\text{Volume} = 1 \text{ dm}^3$$

Required:

$$\text{Molarity} = M = ?$$

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:

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

$$n = 0.1 \text{ mol}$$

$$\begin{aligned} & \star \text{ KOH Molar Mass :-} \\ & = 39 + 16 + 1 \\ & = 56 \text{ g/mol} \end{aligned}$$

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Putting value in formula:

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

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Result: Molarity of solution is 0.10 M.

Concept 3.10:

question#01

Given data:

$$\text{Mass in g of NaOH} = 25 \text{ g}$$

$$\text{Volume of Solution} = 1 \text{ dm}^3$$

required:

$$\text{Molarity} = ?$$

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:

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

$$= 0.625 \text{ mol}$$

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calculate molarity now :

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

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

question#02:

given data:

$$\text{Molarity} = 1.2 \text{ M}$$

$$\text{Volume in dm}^3 = 1 \text{ dm}^3 -$$

$$\text{no. of moles} = 1.2 \text{ mol} -$$

required:

$$\text{Mass in g of NaOH in g/cm}^3 \text{ and g/dm}^3 = ?$$

Formula:

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

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

solution:

Calculate mass in g:

$$\begin{aligned} \text{Mass in g} &= 1.2 \text{ mol} \times 40 \text{ g/mol} \\ &= 48 \text{ g/dm}^3 - \end{aligned}$$

convert into cm^3 :

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

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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:

$$\text{Mass in g} = 10\text{g}$$

$$\text{Volume of Solution} = 1\text{dm}^3 -$$

$$\text{Molar mass} = 6.57 \times 10^4 -$$

Required:

$$\text{no. of moles} = ?$$

$$\text{Molarity} = ?$$

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:

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

$$= 0.000153 \text{ mol}$$

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Calculate molarity

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

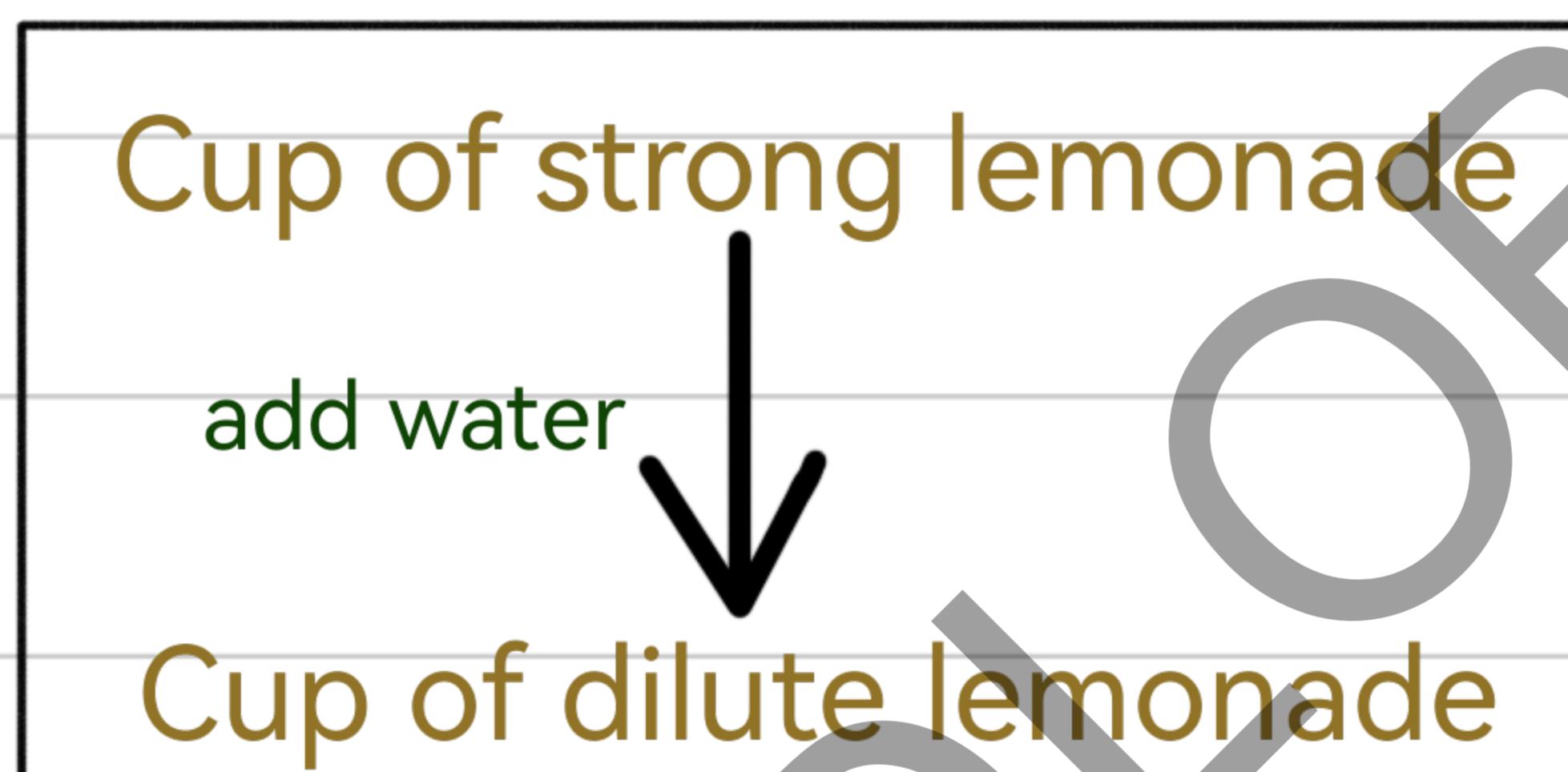
result: molarity of solution containing 10g of haemoglobin dissolved in water to make up 1 dm³ of solution is $1.536 \times 10^{-4} \text{ M}$.

Dilution of solutions:

Definition:

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

Example:



→ Only volume changes not the solute

→ Dilution does not change amount of solute only spread it out in more solvent

Formula:

$$M_1V_1 = M_2V_2$$

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M₁ = concentration of original solution

M₂ = concentration you want after dilution

V₁ = Volume of original solution

V₂ = Total volume of solution

Activity # 01:

Prepare 250 cm³ of solution of 0.1 M 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:

$$M_1 V_1 = M_2 V_2$$

solution:

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

putting values in formula:

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

$$V_1 = 2.08 \text{ cm}^3$$

result:

volume of original solution is 2.08 cm^3

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Example 3.15:

Given data: M_2 = Molarity needed to prepare KNO_3 =

$$V_2 = 500 \text{ cm}^3$$

$M_1 = 4 \text{ M}$ given solution-

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Required:

V_1 = Volume of solution needed to Prepare KNO_3 .

formula:

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

solution:

putting values in formula:

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

$$= 125 \text{ cm}^3$$

Result:

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

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Activity 3.2

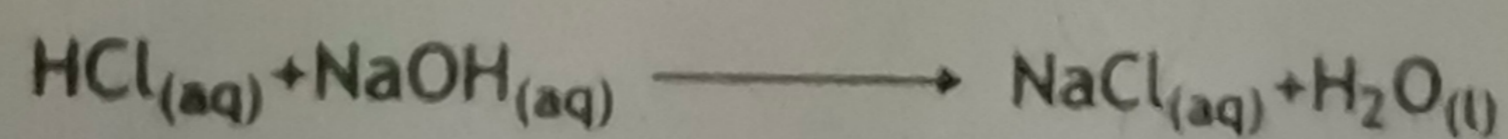
To standardise the given solution of hydrochloric acid.

Note: Perform this activity in the chemistry laboratory.

You will need:

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

Chemical equation



$n_1 = 1 \text{ mole}$ $n_2 = 1 \text{ mole}$

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Concept 3.11:

question 01

Given:

$$M_1 = 0.1\text{M} , M_2 = ?$$

$$V_1 = 9.8 \text{ cm}^3 , V_2 = 10 \text{ cm}^3$$

$$n_1 = 1 , n_2 = 1$$

formulas:

$$\begin{array}{ccc} \text{Acid} & & \text{Base} \\ \frac{M_1 V_1}{n_1} & = & \frac{M_2 V_2}{n_2} \end{array}$$

solution:

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1)

By substituting values:

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

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

$$M_2 = 0.098 \text{ M}$$

2)

We know that $1 \text{ dm}^3 = 1000 \text{ cm}^3$ amount per NaOH per cm^3 :

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

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

question 02:

given:

$$M_1 V_1 = M_2 V_2$$

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

$$V_1 = 200\text{cm}^3, V_2 = ?$$

solution:

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

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

$$V_2 = 600\text{cm}^3$$

amount of water added = final volume - initial volume

$$= 600\text{cm}^3 - 200\text{cm}^3$$

$$= 400\text{cm}^3$$

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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.

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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 ₂ O(for glucose)	Example:C ₆ H 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.

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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:

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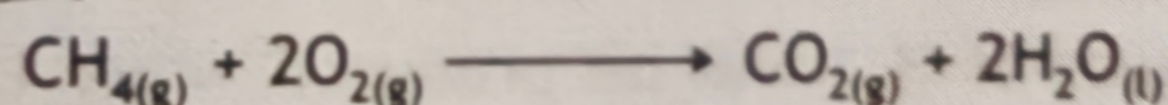
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EXERCISE

1. Multiple Choice Questions (MCQs)

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



- (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 $\text{C}_{20}\text{H}_{30}\text{O}$. 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} ?



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

short answers:

question # 01

Solution:

Calculating moles of elements:

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

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

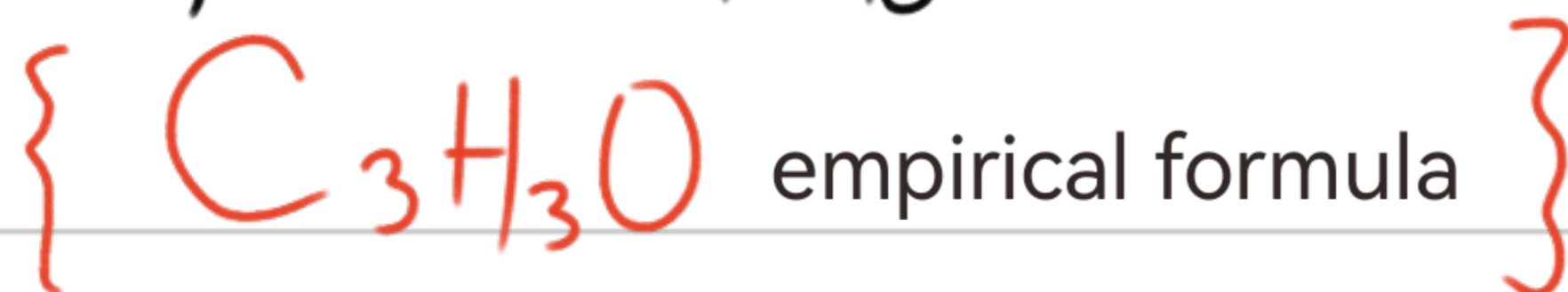
$$O = \frac{29.09}{16} = 1.818 \text{ moles-}$$

calculating mole ratio of each elements:

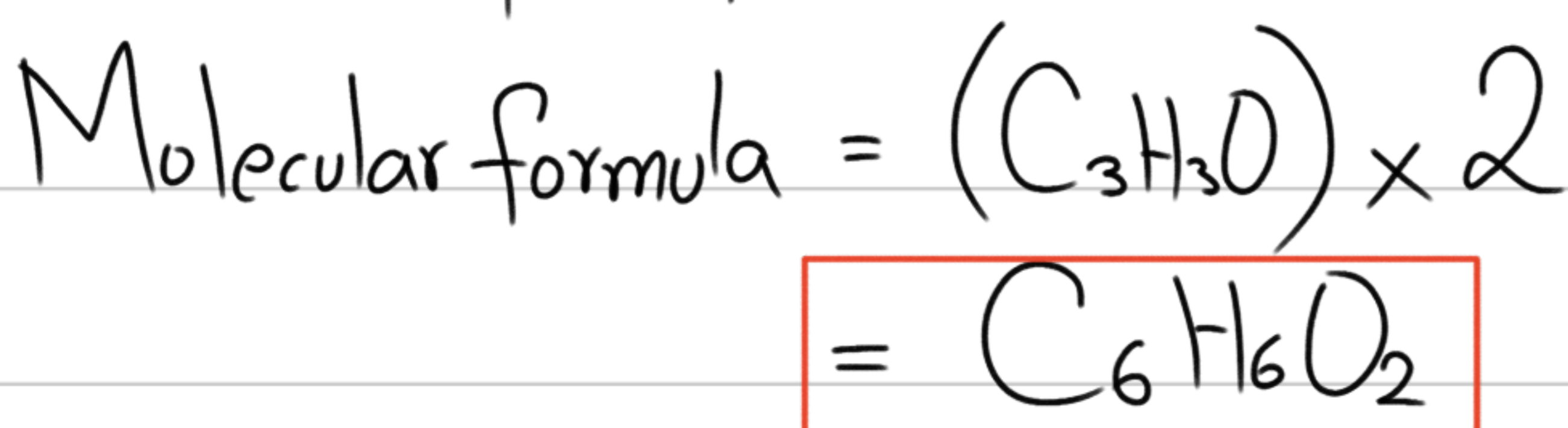
$$\text{mol ratio of C} = \frac{5.454}{1.818} = 3$$

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

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



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



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Question#02

Solution:

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$$\text{Mass of AlCl}_3 = 125 \text{ g}$$

$$\begin{aligned}\text{Molar mass} &= 27 + (3 \times 35.5) = 27 + 106.5 \\ &= 133.5 \text{ g/mol.}\end{aligned}$$

$$\text{Mols of AlCl}_3 = \frac{125}{133.5} = 0.937 \text{ moles}$$

According to balanced chemical equation:

$$2 \text{ moles of AlCl}_3 = 6 \text{ mol of HCl}$$

$$1 \text{ mol of AlCl}_3 = \frac{6}{2} \text{ mol of HCl}$$

$$0.937 \text{ mol} = x$$

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

$$x = 2.81 \text{ moles of HCl}$$

Result:

number of moles of HCl produced is 2.81 moles

question # 03

solution:

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$$\text{Mass of Ca(NO}_3)_2 = 1 \text{ g}$$

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

$$= 40.08 + 28.02 + 96$$

$$= 164.10 \text{ g/mol}$$

$$\text{moles of Ca(NO}_3)_2 = \frac{\text{Mass}}{\text{Molar mass}} = \frac{1.0}{164.10} = 0.00609 \text{ mol}$$

According to balanced chemical equation:

$$1 \text{ mol of Ca(NO}_3)_2 = 3 \text{ moles of O}_2$$

$$0.00609 \text{ mol of Ca(NO}_3)_2 = x$$

$$x = 0.00609 \times 3$$

$$x = 0.0183 \text{ moles of O}_2$$

Result:

number of moles of oxygen needed is 0.0183 moles.

question # 04

Solution:

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

$$\text{Molar mass of Al} = 27$$

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

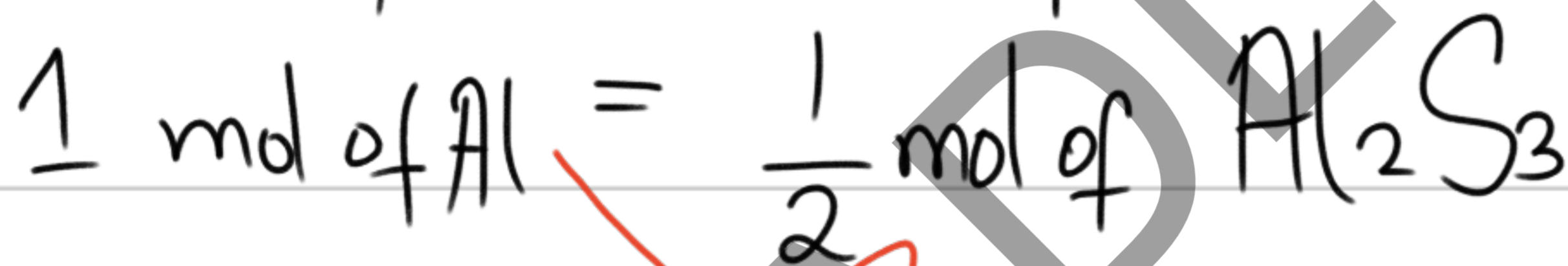
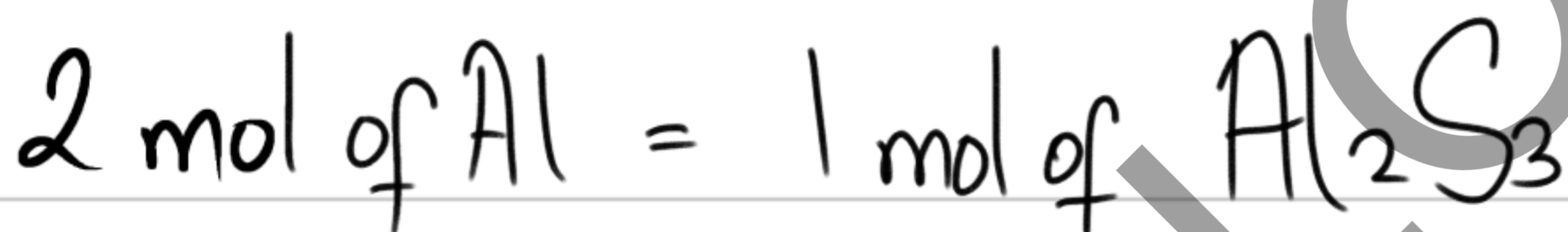
$$\text{Mass of S} = 25 \text{ g}$$

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

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

according to balanced chemical equations:

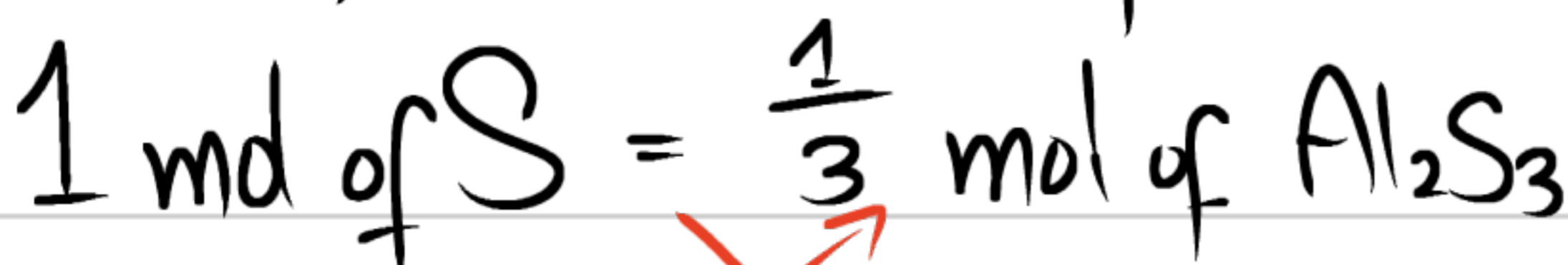
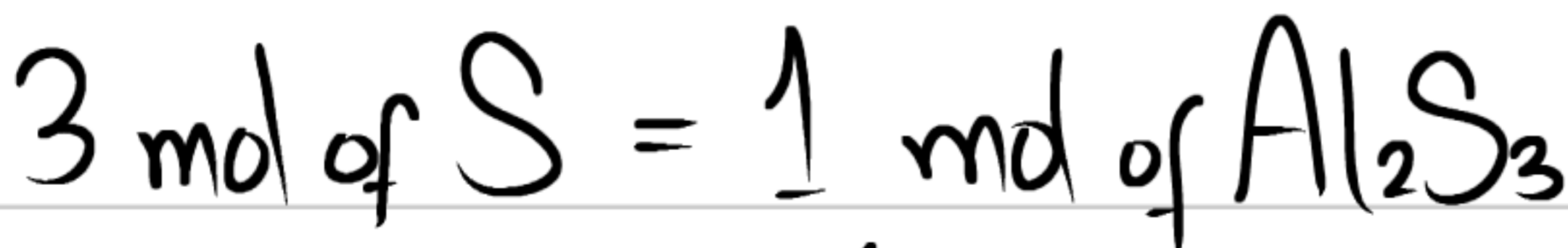
For Al:



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

$$x = \frac{0.926}{2} = 0.463 \text{ moles of Al}_2\text{S}_3$$

For S=



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

$$x = \frac{0.78125}{3} = 0.260 \text{ moles of Al}_2\text{S}_3$$

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result:

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

Question #05

solution:

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Calculating moles of elements:

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

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

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

calculating mole ratio of each elements:

$$\text{Mole ratio of Na} = \frac{0.878}{0.878} = 1 = 3$$

$$\text{Mole ratio of S} = \frac{1.175}{0.878} = 1.338 \times 3 = 4$$

$$\text{Mole ratio of O} = \frac{2.638}{0.878} = 3.005 \text{ to make whole digit} = 9$$



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

$$\begin{aligned} \text{Molar mass} \\ \text{of } \text{Na}_3\text{O}_9\text{S}_4 &= \\ &= 341 \end{aligned}$$

therefore, molecular formula: $= 2 (\text{Na}_3\text{S}_4\text{O}_9)$



Question #06

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Solution:

$$\text{Mass of CaCO}_3 = 68.1 \text{ g}$$

$$\text{Molar mass} = 40 + 12 + (3 \times 16) = 100 \text{ g/mol}$$

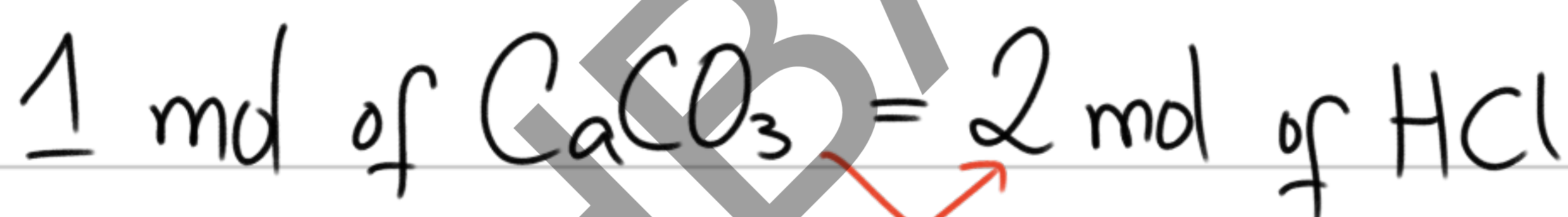
$$\text{Moles of CaCO}_3 = \frac{\text{Mass}}{\text{Molar}} = \frac{68.1}{100} = 0.681 \text{ mol}$$

$$\text{Mass of HCl} = 57.6 \text{ g}$$

$$\text{Molar mass} = 1 + 35 = 36$$

$$\text{Moles of HCl} = \frac{57.6}{36} = 1.6 \text{ mol}$$

According to balanced chemical equation:



$$0.681 \text{ mol of CaCO}_3 = x$$

$$x = 0.681 \times 2$$

$$= 1.362 \text{ moles of HCl}$$

so, HCl is in excess

limiting reactant is CaCO₃.

Question#07

given:

actual yield = 1.85 g

Theoretical yield = 2.5 g

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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

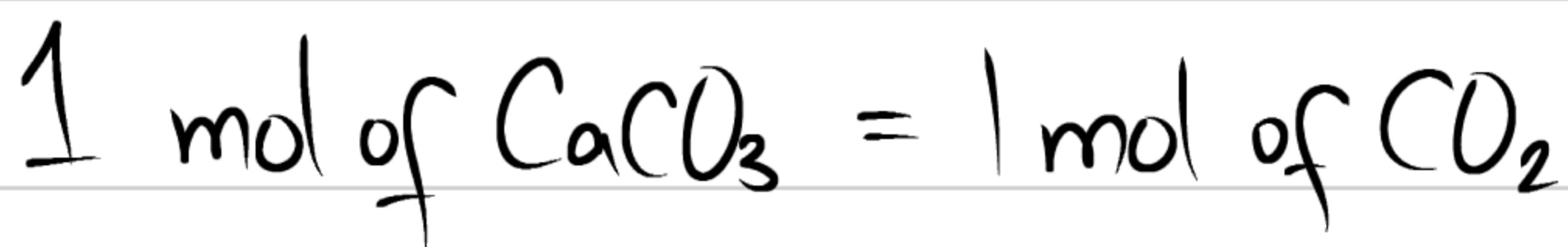
solution:

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:



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

$$\begin{aligned} \text{Mass of CaCO}_3 &= \text{Moles} \times \text{Molar mass} \\ &= 1 \times 100 \\ &= 100 \text{ g} \end{aligned}$$

So, 100 g of limestone was heated.

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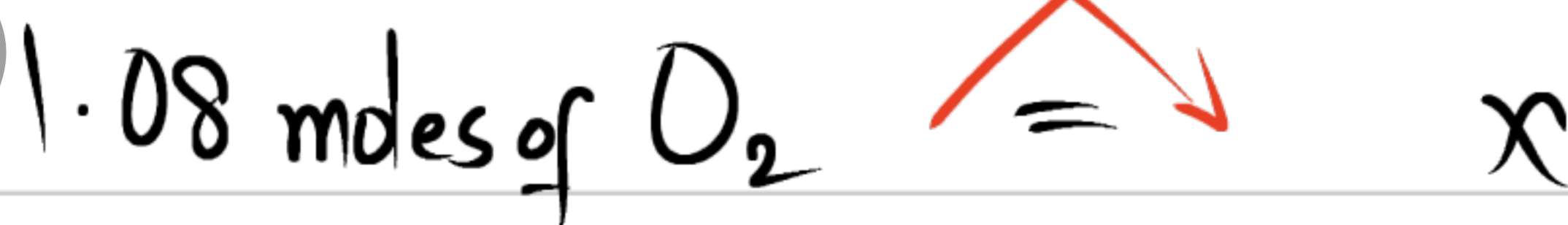
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Question#09

Solution: a) moles of CO₂ produced:-

According to balanced chemical equation:



$$x = \frac{1.08 \times 2}{3}$$

$$= 0.72 \text{ Moles of CO}_2$$

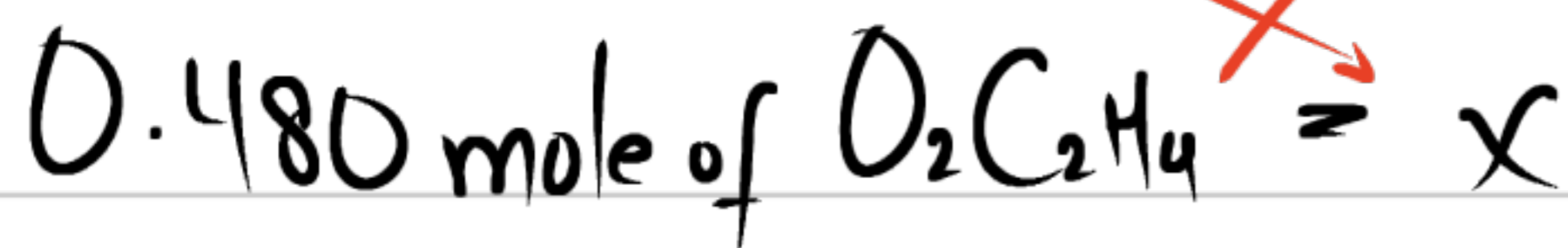
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Solution:-b) reactants left:

According to balanced chemical equation:



$$x = 0.480 \times 3$$

$$= 1.44 \text{ moles of } O_2$$

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but, we only have 1.08 moles of O_2 , which is less than required.

so, O_2 is limiting reactant and it will be completely used.

Since, C_2H_4 is in excess, so some will remain

Hence, C_2H_4 is the reactant left unreacted.

Question # 10

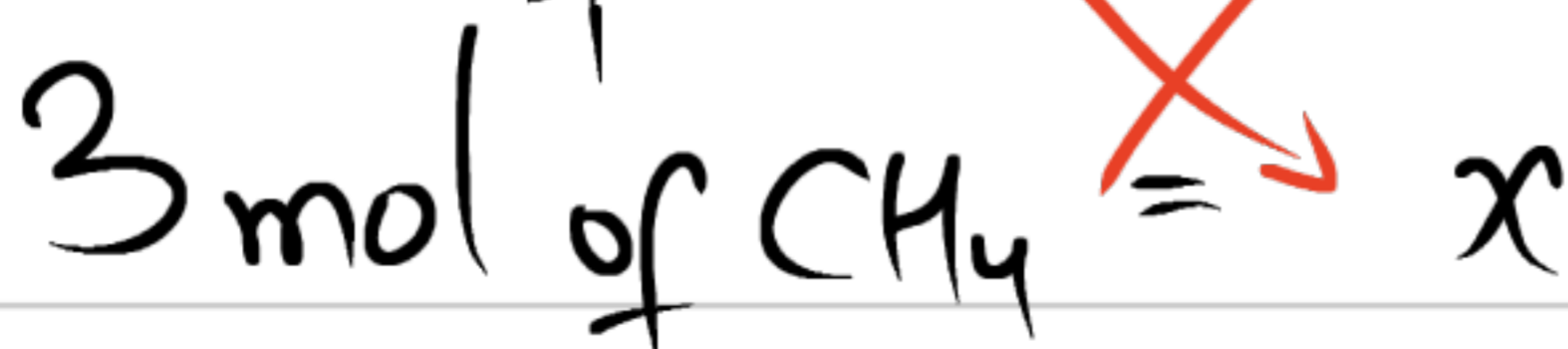
Solution:

$$\text{Mass of } CH_4 = 48g$$

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

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

according to balanced chemical equation:



$$x = 3 \text{ moles of } CO_2$$

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

$$\begin{aligned}\text{Volume of CO}_2 &= 3 \text{ mole} \times 24 \text{ dm}^3 \text{ mole} \\ &= 72 \text{ dm}^3\end{aligned}$$

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

Question # 11

solution:

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

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

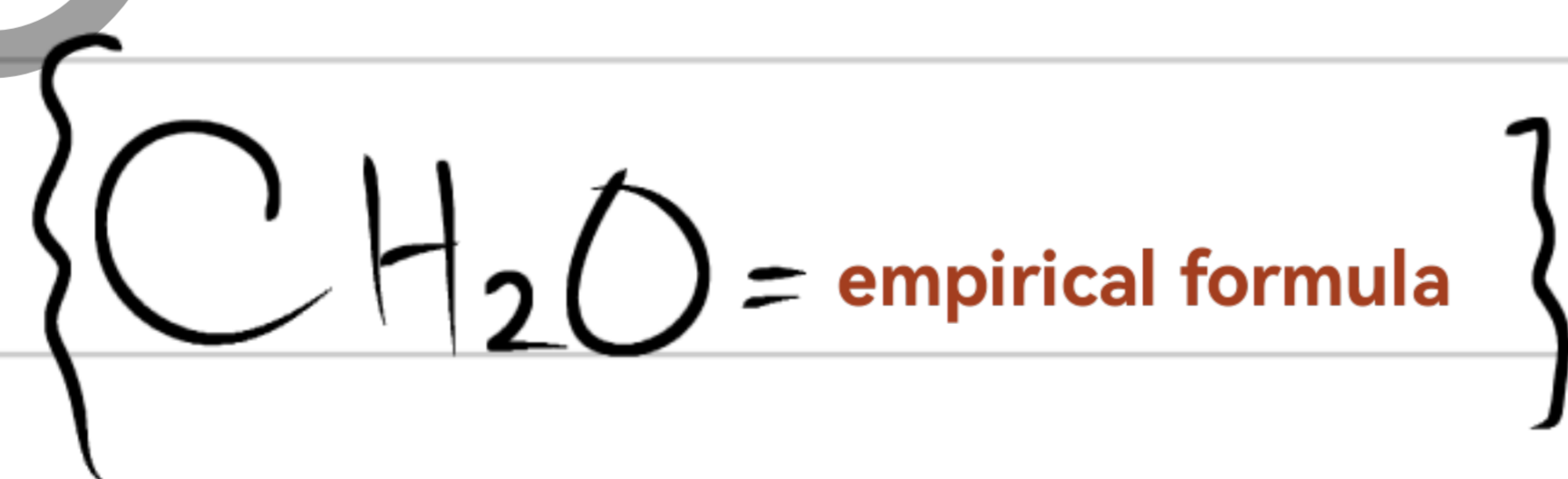
$$O = \frac{53.3}{10} = 5.33 \text{ mol}$$

According to balanced chemical equation:

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

$$\text{Mole ratio of H} = \frac{6.7}{3.33} = 2$$

$$\text{Mole ratio of O} = \frac{5.33}{3.33} = 1$$



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Question 12

Solution: a) Concentration in g/cm³

Mass of Sodium Chloride (NaCl) = 10.0 g

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

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

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

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b) Concentration in mol/dm^3 :-

Mass of Sodium Chloride = 10 g

Molar mass = $23 + 35 = 58 \text{ g/mol}$

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

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

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

question no. 13

Solution:

Volume of NaOH = $25 \text{ cm}^3 = \frac{25}{1000} = 0.025 \text{ dm}^3$

Concentration of NaOH = 0.10 mol/dm^3

Volume of HCl = $\frac{20}{1000} = 0.020 \text{ dm}^3$

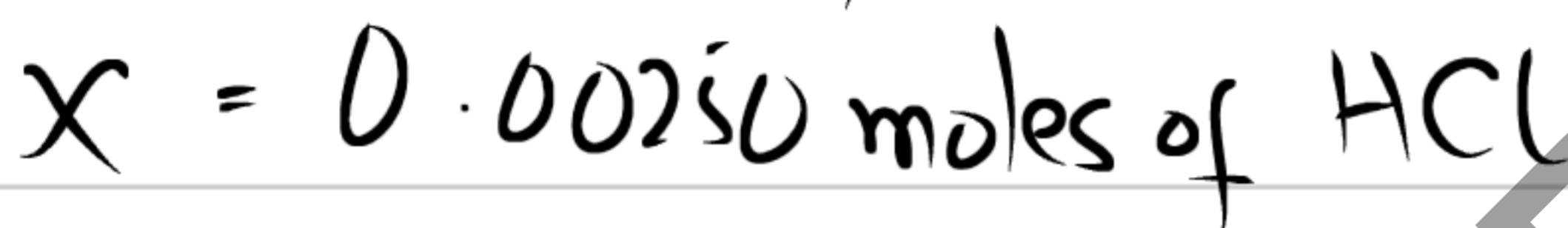
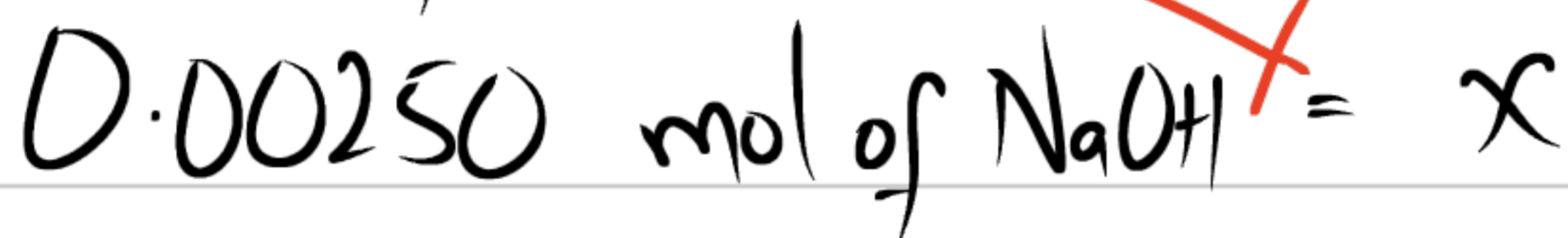
$$\begin{aligned}\text{Moles of NaOH} &= \text{Concentration} \times \text{Volume} \\ &= 0.10 \times 0.0250 \\ &= 0.00250 \text{ mol.}\end{aligned}$$

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According to balanced chemical equation:



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

Hence, concentration of HCl is 0.125 mol/dm³.

Question # 14

a)

Concentration is the amount of solution (in this case sodium chloride – NaCl) dissolved in a specific volume of solvent (water), typically expressed as mol.

Yes, cubic decimales (mol/dm³).

In this case, 0.5 mol/dm³ NaCl solution means 0.5 moles of NaCl are dissolved in 1 dm³ (1000 cm³) of water.

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b) solution:

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

$$\text{Concentration of NaCl} = 0.5 \text{ mol/dm}^3$$

$$\text{Moles of NaCl} = \text{Concentration} \times \text{Volume}$$

$$= 0.5 \times 0.250$$

$$= 0.125 \text{ mole}$$

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$$\text{Mass of NaCl} = \text{Moles} \times \text{Molar mass}$$

$$= 0.125 \times 58$$

$$= 7.25 \text{ g}$$

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c) solution:

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

$$\text{Molar mass of NaCl} = 23 + 35 = 58$$

$$\text{Concentration in (g/cm}^3\text{)} = \text{Concentration} \times \text{Molar mass}$$

$$= 0.5 \times 58$$

$$= 29 \text{ g/dm}^3$$

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 cm³ volumetric flask (ex bracket first).
 - Add about 100 – 150 cm³ 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 cm³ calibration line.
5. Mix thoroughly: Stopper the flask and invert several times to ensure uniform mixing.