



Beaconhouse School System

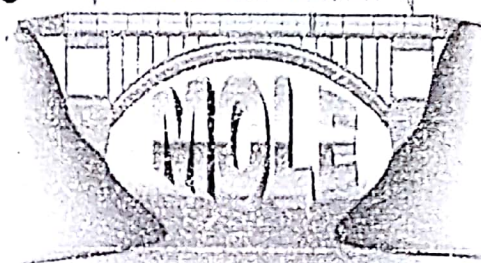
GARDEN TOWN CAMPUS

LEVEL
6
CHEMISTRY
5070

TOPIC 3

MOLES

Grams $\rightarrow$ Bridge between mass and number $\leftarrow$ Count



AND

STOICHIOMETRY

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Class :

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TOPIC: 3

MOLES AND STOICHIOMETRY

syllabus

(f)	define relative molecular mass, M_r , and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses
(g)	calculate the percentage mass of an element in a compound when given appropriate information
(h)	calculate empirical and molecular formulae from relevant data
(i)	*calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24 dm^3 at room temperature and pressure); calculations involving the idea of limiting reactants may be set (questions on the gas laws and the calculations of gaseous volumes at different temperatures and pressures will not be set)
(j)	*apply the concept of solution concentration (in mol/dm^3 or g/dm^3) to process the results of volumetric experiments and to solve simple problems (appropriate guidance will be provided where unfamiliar reactions are involved)
(k)	calculate % yield and % purity

RELATIVE MOLECULAR MASS

Relative Atomic Mass (elements) – A_r

The relative atomic mass (A_r) of an element is the average mass of its one atom, compared to $1/12$ the mass of one atom of carbon-12.

- A_r of element = Nucleon number of element (e.g. A_r of Na = 23)
- A_r is a ratio and has no units.
- A_r values may not always be whole numbers as some elements like chlorine occur as mixtures of isotopes.

Relative Molecular Mass / Relative Formula Mass – M_r

The relative molecular mass (M_r) of an element or compound is the average mass of its one molecule, compared to $1/12$ the mass of one atom of carbon-12.

- Relative molecular mass is used for molecules (e.g. O_2 , H_2O);
- Relative formula mass is used for ionic compounds (e.g. CuSO_4 , CaCO_3).
- $M_r = (A_r \text{ of 1st element} \times \text{No. of atoms}) + \dots$
- e.g. M_r of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = [(64 \times 1) + (32 \times 1) + (16 \times 4)] + 5[(1 \times 2) + (16 \times 1)] = 250$

MOLE

Mole is actually a counting unit

In daily life, we encounter an use several counting units, to make counting easier. For example:

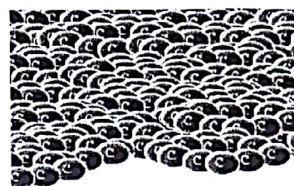
Similarly, we got to believe that:

Counting unit	units	Example
Pair	2	Shoes
Dozen	12	Eggs
Pack	52	Playing cards
Ream	1000	Sheets of printing paper
Gross	144	Drawing pins

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"Mole is also a counting unit for atoms, molecules, ions, electrons etc."

Counting Unit	Number of Units	Examples
MOLE	6.02×10^{23}	A mole of carbon, a mole of water, a mole of glucose, a mole of electrons



Mass of one mole of substance:

Mole differs from other counting units in the following two sense

- The number of units in a mole is extremely large.
- The mass of one mole of a particular substance is always a fixed and constant quantity.

$$\text{MASS OF 1 MOLE} = A_r \text{ or } M_r$$

e.g. 1 mole of carbon = 12 g 1 mole of water = 18 g
 1 mole of glucose = 180 g 1 mole of sodium chloride = 58.50 g
 but 1 pair of shoes, 1 dozen of eggs does not always necessary have same mass.

The number of units in a mole – Avogadro's Number (or Avogadro Constant)

The number of atoms in one mole of an element, the number of molecules in one mole of a molecular compound or number ions in one mole of an ionic species is termed as Avogadro's number, N_A . Its value is 6.02×10^{23}

Counting moles by weighing

There is no direct way of measuring moles nor its possible to count such large number of particles. However, by weighing correct masses of these substances serves both the above purposes. Following simple principles help you to count the enormous number of molecules just in a matter of seconds.

$$\text{One mole of an element} = \text{Relative atomic mass } (A_r) \text{ of element in grams}$$

Examples

1 mole of hydrogen (H) atoms = 1 g = 6.02×10^{23} H atoms
 1 mole of Carbon (C) atoms = 12 g = 6.02×10^{23} C atoms
 1 mole of Sulphur (S) atoms = 32 g = 6.02×10^{23} S atoms

$$\text{One mole of diatomic element} = \text{Relative molecular mass } (M_r) \text{ of element in grams}$$

1 mole of Hydrogen (H_2) gas = 2 g = 6.02×10^{23} H_2 molecules
 1 mole of Oxygen (O_2) gas = 32 g = 6.02×10^{23} O_2 molecules
 1 mole of Nitrogen (N_2) gas = 28 g = 6.02×10^{23} N_2 molecules

$$\text{One mole of compound} = \text{Relative Molecular mass } (M_r) \text{ of Compound in grams}$$

Examples

1 mole of Water (H_2O) atoms = 18 g = 6.02×10^{23} H_2O molecules
 1 mole of Carbon dioxide (CO_2) = 44 g = 6.02×10^{23} CO_2 molecules
 1 mole of Glucose ($C_6H_{12}O_6$) = 180 g = 6.02×10^{23} $C_6H_{12}O_6$ molecules

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Mole – defined

A mole is defined as the amount of substance (element, compound or ion etc.), which contains Avogadro number of particles (6.02×10^{23} atoms/ions/molecules etc.)

Mole – Examples

Element	Symbol	Relative atomic mass	Mass of 1 mole	Number of atoms in 1 mole
Hydrogen	H	1	1 g	6.02×10^{23}
Helium	He	4	4 g	6.02×10^{23}
Carbon	C	12	12 g	6.02×10^{23}
Oxygen	O	16	16 g	6.02×10^{23}

Element	Formula	Relative molecular Mass	Mass of 1 mole	Number of molecules in 1 mole
Hydrogen	H ₂	2	2 g	6.02×10^{23}
Chlorine	Cl ₂	71	71 g	6.02×10^{23}
Oxygen	O ₂	32	32 g	6.02×10^{23}
Iodine	I ₂	254	254 g	6.02×10^{23}

Compound	Formula	Relative molecular / formula Mass	Mass of 1 mole	Number of molecules / ions in 1 mole
Water	H ₂ O	18	2 g	6.02×10^{23}
Ammonia	NH ₃	17	17 g	6.02×10^{23}
Carbon dioxide	CO ₂	44	44 g	6.02×10^{23}
Glucose	C ₆ H ₁₂ O ₆	180	180 g	6.02×10^{23}
Sodium Chloride	NaCl	58.50	58.50 g	6.02×10^{23}

Remember!

- 1 mole = M_r or A_r
- Same moles of different element have same atoms but different mass
- Same moles of different compounds have same molecules but different mass

Calculation based on moles

$$1 \text{ moles} = \frac{\text{Mass in g}}{A_r \text{ or } M_r}$$

$$2 \text{ Mass} = \text{moles} \times A_r \text{ or } M_r$$

$$3 \text{ } A_r \text{ or } M_r = \frac{\text{Mass}}{\text{moles}}$$

Sample 1 Example

Calculate the number of moles of an aluminum saucepan with a mass of 270 g

$$\text{Number of moles of Al} = \frac{270 \text{ g}}{27} = 10 \text{ mol. Al}$$

Remember...! Don't forget the unit of mass g

Sample 2 Example

Carbon tetrachloride, CCl₄, is colourless liquid used in manufacture of Fluorocarbons and an industrial solvent. What is the mass of 0.0758 moles of carbon tetrachloride.

$$\begin{aligned} \text{Molecular Mass of Carbon tetrachloride} &= 154 \\ \text{Number of moles of CCl}_4 &= 0.0758 \text{ mol} \\ \text{Number of moles} &= 0.0758 \times 154 = 11.6732 \text{ g} \end{aligned}$$

MOLES / MASS OF A ELEMENT/ION FROM ITS COMPOUND



1

moles of an element from mole of compound = Number of atoms of element in the compound $\times$ moles of compound

2

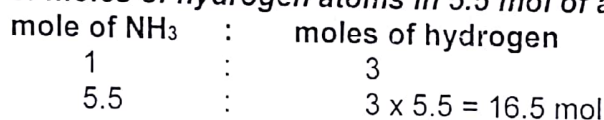
moles of an element from mass of compound = $\frac{\text{Number of atoms of element in the compound} \times \text{mass of compound}}{M_r \text{ of compound}}$

3

mass of an element from mass of compound = $\frac{\text{Number of atoms of element in the formula of compound} \times A_r \times \text{mass of compound}}{M_r \text{ of compound}}$

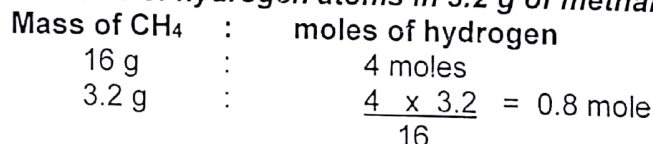
Sample Example 1

What is the number of moles of hydrogen atoms in 5.5 mol of ammonia?



Sample Example 2

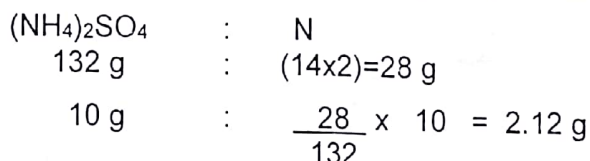
What is the number of moles of hydrogen atoms in 3.2 g of methane?



Sample Example 3

What is the mass of nitrogen present in 10 of ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$?

$$M_r \text{ of } (\text{NH}_4)_2\text{SO}_4 = [14 + (1 \times 4)] \times 2 + 32 + (16 \times 4) = 132$$



Sample Example 4

In a given 75 g sample of Calcium Phosphate, $\text{Ca}_3(\text{PO}_4)_2$, calculate

(a) the mass of Calcium

(b) the mass of oxygen

(c) the percentage of phosphate ions, PO_4^{3-}

(a) $M_r \text{ of calcium phosphate, } \text{Ca}_3(\text{PO}_4)_2 = (40 \times 3) + 2[31 + (16 \times 4)] = 310$

$$\text{Mass of calcium} = \frac{3 \times 40}{310} \times 75 = 29.03 \text{ g Ca}$$

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$$(b) \text{ Mass of oxygen} = \frac{8 \times 16}{310} \times 75 = 30.97 \text{ g Ca}$$

$$(c) \text{ Mass of phosphate ions} = \frac{3 \times [31 + (16 \times 4)]}{310} \times 75 = 49.97 \text{ g PO}_4^{3-} \text{ ions}$$

$$\text{Percentage of phosphate ions} = \frac{45.97}{75} \times 100 = 61.293 \% \text{ PO}_4^{3-}$$

Sample Example 5

Calculate the percentage of water of crystallization in 35.7 g sample of washing soda, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

$$M_r \text{ of washing soda, } \text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = (23 \times 2) + 12 + (16 \times 3) + 10[(1 \times 2) + 16] = 286$$

$$\text{Mass of water of crystallization} = \frac{10 \times 18}{286} \times 35.7 = 22.47 \text{ g H}_2\text{O}$$

$$\text{Percentage of water of crystallization} = \frac{22.47}{35.7} \times 100 = 62.94 \% \text{ H}_2\text{O}$$

MOLAR VOLUME

Principle:

A fixed volume of any gas under the same conditions of temperature and pressure always contains the same number of molecules. This means that 10 cm^3 of hydrogen contains same number of molecules as 10 cm^3 of carbon dioxide if the gases are measured at the same temperature and pressure.

H_2

$V = 10 \text{ cm}^3$
 $T = 25^\circ\text{C}$
 $P = 1 \text{ atm}$

CO_2

$V = 10 \text{ cm}^3$
 $T = 25^\circ\text{C}$
 $P = 1 \text{ atm}$

Same moles = Same molecules

1 mole = 24 dm^3 or 24000 cm^3

At room conditions (i.e. 25°C and standard atmospheric pressure),
one mole of gas has a volume of 24 dm^3 or 24000 cm^3 .

This is called the molar volume at r.t.p. (room temperature and pressure).

Gas	Molecular Formula	Mass of gas that has a volume of 24 dm^3 (molar mass)	Volume at r.t.p
Hydrogen	H_2	2 g	24 dm^3
Ammonia	NH_3	17 g	
Nitrogen	N_2	28 g	
Oxygen	O_2	32 g	
Carbon dioxide	CO_2	44 g	
Propane	C_3H_8	44 g	

1

Volume of gas r.t.p = moles $\times$ 24

2

Volume of gas r.t.p = $\frac{\text{Mass}}{M_r} \times 24$

3

moles at r.t.p = $\frac{\text{Volume at rtp}}{24}$

4

Mass of gas = $\frac{\text{Volume of gas r.t.p}}{24} \times M_r$

Sample Example 1

In an experiment, 0.04 mol of nitrogen gas was produced (measured at rtp). Calculate the volume of gas.

$$\text{Volume of N}_2 \text{ gas} = 0.04 \text{ mol} \times 24 \text{ dm}^3 = \boxed{0.96 \text{ dm}^3}$$

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Sample Example 2

What is the volume of 7 g of nitrogen, N_2 , at room conditions ?

$$\text{Volume of } N_2 \text{ gas} = \frac{7g}{28} \times 24 \text{ dm}^3 = 6 \text{ dm}^3$$

Sample Example 3

In a particular reaction, at room temperature and pressure, 350 cm^3 of carbon dioxide gas was evolved. What would be number of moles of gas? (M_r of $CO_2 = 44$)

$$\text{Volume of gas} = 350 \text{ cm}^3 = 350/1000 = 0.350 \text{ dm}^3$$

$$\text{Number of moles of } CO_2 = \frac{0.350 \text{ dm}^3}{24 \text{ dm}^3 \text{ mol}^{-1}} = 0.015 \text{ mol}$$

or alternatively

$$\text{Number of moles of } CO_2 = \frac{350 \text{ cm}^3}{24000 \text{ cm}^3 \text{ mol}^{-1}} = 0.015 \text{ mol}$$

Sample Example 4

What mass of 200 cm^3 of methane gas will be required for complete combustion into carbon dioxide and water? (M_r of $CH_4 = 16$)

$$\text{Mass of methane} = \frac{200 \text{ cm}^3}{24000 \text{ cm}^3} \times 16 \text{ g mol}^{-1} = 0.13 \text{ g } CH_4$$

PERCENTAGE COMPOSITION

Definition:

The mass of a particular element, ion or unit expressed as a percentage of total mass in a compound is called Percentage Composition.

$$\text{Percentage of an element in a compound} = \frac{\text{number of atoms of that element in the formula} \times A_r \text{ of element}}{M_r \text{ of the compound}} \times 100$$

Sample Example

Calculate the percentage of

- (a) Copper ions (b) Sulphate atoms (c) Sulphate ions
(d) Oxygen atoms (e) water (f) anhydrous copper sulphate

in hydrated copper sulphate crystals?

$$M_r \text{ of hydrated copper sulphate, } CuSO_4 \cdot 5H_2O = 64 + 32 + (16 \times 4) + 5[(1 \times 2) + 16] = 250$$

$$(a) \text{ Percentage of copper ions} = \frac{1 \times 64}{250} \times 100 = 26\% \text{ Cu}^{2+} \text{ ions}$$

$$(b) \text{ Percentage of sulphate ions} = \frac{1 \times 96}{250} \times 100 = 38.4\% \text{ SO}_4^{2-} \text{ ions}$$

$$(c) \text{ Percentage of sulphur atoms} = \frac{1 \times 32}{250} \times 100 = 12.8\% \text{ S atoms}$$

$$(d) \text{ Percentage of oxygen atoms} = \frac{4 \times 16}{250} \times 100 = 25.6\% \text{ O atoms}$$

$$(e) \text{ Percentage of water molecules} = \frac{5 \times 18}{250} \times 100 = 36\% \text{ H}_2\text{O molecules}$$

$$(f) \text{ Percentage of anhydrous copper sulphate} = \frac{1 \times 60}{250} \times 100 = 24\% \text{ CuSO}_4 \text{ atoms}$$

EMPIRICAL FORMULA

Definition:

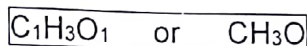
The formula which gives simplest whole number ratio of atoms of each element present in one molecule or formula unit of a compound.

Sample Example 1

Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7% hydrogen and 51.6% oxygen. Its molar mass is 62.1 grams mol. Determine its empirical formula.

	C	H	O
Percentage Composition	38.7	9.7	51.6
Mole Ratio =	$\frac{38.7}{12} = 3.225$	$\frac{9.7}{1} = 9.7$	$\frac{51.6}{16} = 3.225$
<u>%age composition</u> Ar			
Simplest atomic ratio = <u>Mole Ratio</u>	$\frac{3.225}{3.225} = 1$	$\frac{9.7}{3.225} = 3$	$\frac{3.225}{3.225} = 1$
Least mole ratio			

Empirical Formula of Ethylene Glycol =



Sample Example 2

12.0 g of anhydrous magnesium sulphate combines with 12.6 g of water to form hydrated magnesium sulphate. What is the formula of the hydrated magnesium sulphate? [$M_r(\text{MgSO}_4) = 120$]

	MgSO_4	:	H_2O
moles	$\frac{12.0}{120} = 0.1$	:	$\frac{12.6}{18} = 0.7$
	1	:	7

Formula of the hydrated magnesium sulphate = $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

Sample Example 3

Analysis of an organic acid isolated from red ants shows that it contains 0.060 g of carbon, 0.010 g of hydrogen and 0.16 g of oxygen. Calculate the empirical formula for this acid.

	C	H	O
Mole Ratio	$\frac{0.060}{12} = 0.005$	$\frac{0.010}{1} = 0.010$	$\frac{0.16}{16} = 0.010$
Simplest atomic ratio	$\frac{0.005}{0.005} = 1$	$\frac{0.010}{0.005} = 2$	$\frac{0.010}{0.005} = 2$

Empirical Formula: CH_2O_2

MOLECULAR FORMULA DETERMINATION

Step I

Determine the empirical formula or is already given.

Step II

Determine the empirical formula mass.

e.g. Empirical formula mass of $\text{CH}_2\text{O} = 12 + (1 \times 2) + 16 = 30$

Step III

Determine the multiple factor, n , by the formula

$$n = \frac{\text{Molecular Mass (given)}}{\text{Empirical formula mass}}$$

e.g. Multiple factor, $n = \frac{180}{30} = 6$

Step IV

Determine the molecular formula as follows:

$$\text{Molecular formula} = n (\text{Empirical Formula})$$

e.g. Molecular formula of a compound having molecular mass of 180 and empirical formula, CH_2O , is given as

$$\text{Molecular Formula} = 6 (\text{CH}_2\text{O}) = \text{C}_{(1 \times 6)} \text{H}_{(2 \times 6)} \text{O}_{(1 \times 6)} = \text{C}_6\text{H}_{12}\text{O}_6$$

STOICHIOMETRY

pronunciation : Stoy—key—ahma—tree.

Definition:

The branch of chemistry that deals with the quantitative relationships between reactants and products based on balanced chemical equation is known as Stoichiometry.

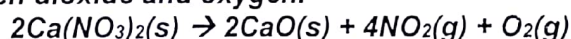
Significance:

The knowledge of stoichiometry is important because it helps us to calculate theoretically:

- The amount of products that can be produced as a result of given quantities of reactant(s) — Theoretical Yield of a product.
- The amount of reactant(s) required to produce a definite amount of product.
- The quantity of one reactant that reacts with the other(s).
- The amount of product produced along with other products.
- The percentage purity of a substance.
- The percentage yield of a product.
- The composition and identity of an unknown substance, etc.
- The volume of gases required or produced during a reaction.

mole ratio

The carbonates of many metallic elements decompose when heated. The nitrates of metallic elements also decompose when heated. Calcium nitrate decomposes to form calcium oxide, nitrogen dioxide and oxygen.

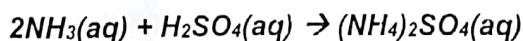


A 0.010 mol sample of calcium nitrate is heated. Calculate the number of moles of gas produced when this sample is completely decomposed.

mol ratio	$\text{Ca}(\text{NO}_3)_2$	:	$\text{NO}_2 + \text{O}_2$
	2 mol	:	(4+1) 5 mol
	0.010 mol	:	$0.010 \times \frac{5}{2} = 0.025 \text{ mol}$

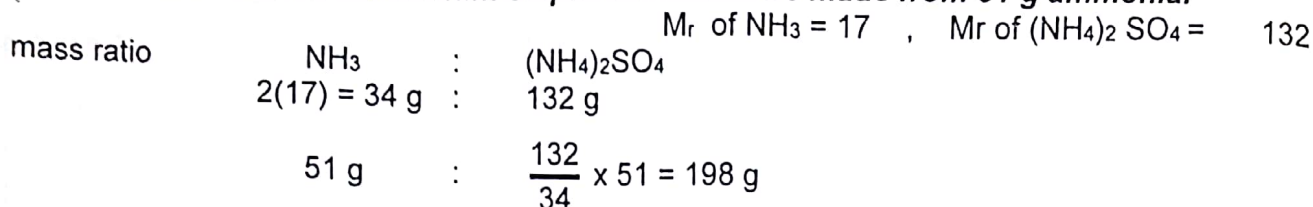
mass ratio

Ammonium sulphate can be made by reacting aqueous ammonia with dilute sulphuric acid.



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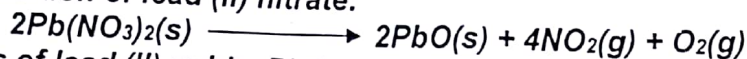
Calculate the mass of ammonium sulphate that can be made from 51 g ammonia.



DETERMINATION OF VOLUME OF A GAS REQUIRED / PRODUCED

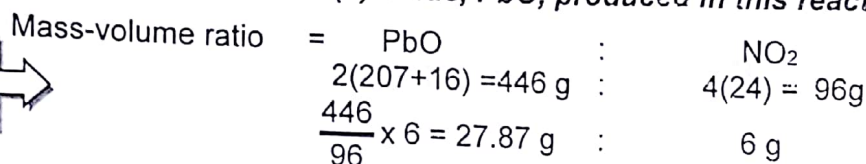
Sample Example 1

In an experiment, 6 dm³ of NO_2 gas measured at r.t.p. was produced by decomposition of lead (II) nitrate:

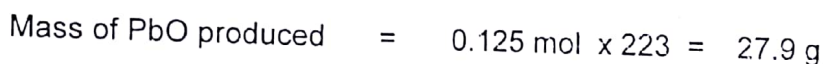
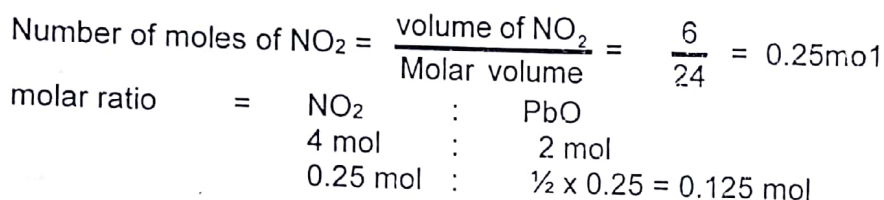


Calculate the mass of lead (II) oxide, PbO, produced in this reaction.

1st method

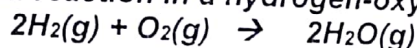


2nd method



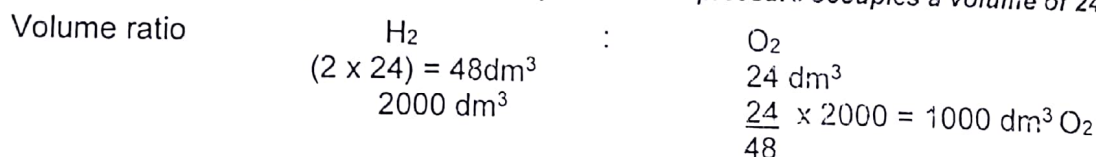
Sample Example 2

Hydrogen-oxygen fuel cells are used to generate electricity. The overall reaction in a hydrogen-oxygen fuel cell is shown below.



A hydrogen-oxygen fuel cell uses 2000 dm³ of hydrogen measured at room temperature and pressure.

Calculate the volume of oxygen, measured at room temperature and pressure, used by the fuel cell. [One mole of any gas at room temperature and pressure occupies a volume of 24 dm³.]



CONCENTRATION OF SOLUTIONS

Solution: A solution is a homogeneous mixture (having a uniform composition) of two or more substances)

Binary Solution: A solution consisting of only two components is termed as a binary solution.

Components of a binary solution:

A binary solution consists of two components:

- Solute:** The component of a solution in relatively smaller quantity is called a solute.
- Solvent:** The component of a solution in relatively larger quantity is called a solvent.

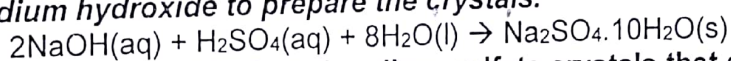
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Concentration of a solution:

- 1 Concentration in moles per dm^3 = $\frac{\text{moles}}{\text{Volume of solution in } \text{dm}^3}$
- 2 Concentration in moles per dm^3 = $\frac{\text{mass in grams}}{M_r \times \text{Volume of solution in } \text{dm}^3}$
- 3 Concentration in moles per dm^3 = $\frac{\text{moles}}{\text{Volume of solution in } \text{cm}^3} \times 1000$
- 4 Concentration in moles per dm^3 = $\frac{\text{mass in grams}}{M_r \times \text{Volume of solution in } \text{cm}^3} \times 1000$
- 5 Moles = Concentration (mol/dm^3) $\times$ Volume of solution in dm^3
- 6 Moles = Concentration (mol/dm^3) $\times$ M_r $\times$ Volume of solution in dm^3
- 7 Moles = Concentration (mol/dm^3) $\times$ $\frac{\text{Volume of solution in } \text{cm}^3}{1000}$
- 8 Moles = Concentration (mol/dm^3) $\times$ $\frac{M_r \times \text{Volume of solution in } \text{dm}^3}{1000}$
- 9 Concentration in g/dm^3 = Concentration in $\text{mol/dm}^3 \times M_r$ of solute

Sample Example 1

A student prepares some crystals of hydrated sodium sulfate by titrating aqueous sodium hydroxide with sulfuric acid. The student uses 25.0 cm^3 of 1.60 mol/dm^3 sodium hydroxide to prepare the crystals.



Calculate the maximum mass of hydrated sodium sulfate crystals that can be formed.

$$\text{Moles of NaOH} = 1.60 \times \frac{25}{1000} = 0.04 \text{ mol}$$

Mol ratio	NaOH	:	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
	2 mol	:	1 mol
	0.04 mol	:	$\frac{1}{2} \times 0.04 = 0.02 \text{ mol}$

$$\text{Maximum mass of hydrated sodium sulfate} = 0.02 \text{ mol} \times 322 (M_r) = 6.44 \text{ g } \text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$$

Sample Example 2

15.0 cm^3 of 1.0 mol/dm^3 potassium hydroxide just neutralise 20.0 cm^3 of a solution of nitric acid. What is the concentration of the acid?

$$\text{Moles of KOH} = \frac{1.0 \times 15}{1000} = 0.015 \text{ mol KOH}$$

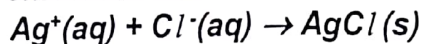
Mole ratio	KOH	:	HNO_3
	1 mol	:	1 mol
	0.015 mol	:	0.015 mol

$$\text{Concentration of acid} = \frac{n}{V} = \frac{0.015}{20} \times 1000 = 0.75 \text{ mol dm}^{-3}$$

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Sampled Example 1

Silver ions react with chloride ions.



It is found that 5 cm^3 of a 0.1 mol/dm^3 solution of the chloride of metal X needs 10 cm^3 of 0.1 mol/dm^3 silver nitrate for complete reaction. What is the formula of the chloride?

$$\text{Moles of chloride} = 0.1 \times \frac{5}{1000} = 0.0005 \text{ mol}$$

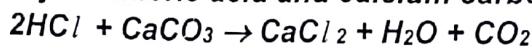
$$\text{Moles of AgNO}_3 = 0.1 \times \frac{10}{1000} = 0.0010 \text{ mol}$$

$$0.0010 = 2 \times 0.0005$$

So the formula XCl_2

Sampled Example 2

The reaction between hydrochloric acid and calcium carbonate is shown.



Which volume of 1.0 mol/dm^3 hydrochloric acid is needed to react completely with 1.0 g of calcium carbonate ($M_r = 100$)?

Mass mole ratio	CaCO_3	:	HCl
	100 g	:	2 mol
	1 g	:	$\frac{2}{100} \times 1 = 0.02 \text{ mol}$

$$\text{concentration of HCl} = \frac{\text{number of moles}}{\text{volume}}$$

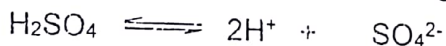
$$1.0 \text{ mol dm}^{-3} = \frac{0.02}{V}$$

$$V = 0.02 \text{ dm}^3 \text{ or } 20 \text{ cm}^3$$

CALCULATION OF IONS FROM A COMPOUND

Sampled Example 1What is the concentration of hydrogen ions in 0.05 mol/dm^3 sulfuric acid?

A 0.025 g/dm^3 B 0.05 g/dm^3 C 0.10 g/dm^3 D 2.0 g/dm^3



$$1 \text{ mol} \quad 2 \text{ mol}$$

$$0.05 \text{ mol}$$

$$2 \times 0.05 = 0.1 \text{ mol}$$

$$A_r \text{ of } \text{H}^+ = 1$$

$$\text{conc(g/dm}^3\text{)} = 0.1 \text{ mol/dm}^3 \times 1 = 0.1 \text{ g dm}^{-3}$$

Correct Option: C

Sampled Example 2

A student titrates 20.0 cm^3 of a metal hydroxide, M(OH)_2 , of concentration 0.060 mol/dm^3 with a strong acid of concentration 0.050 mol/dm^3 . It requires 24.0 cm^3 of acid to neutralise the metal hydroxide.

(i) Calculate the number of moles of acid in 24.0 cm^3 of the acid.

$$0.050 \times \frac{24.0}{1000} = 0.0012 \text{ mol}$$

1 mol of base give two mole OH^-

(ii) Calculate the number of moles of OH^- ions in 20.0 cm^3 of the metal hydroxide.

$$\text{Mole of } \text{OH}^- = \left(0.060 \times \frac{20.0}{1000}\right) 2 = 0.0024 \text{ mol}$$

VOLUMETRIC EXPERIMENTS

- ✓ In such questions an acid and base is given.
- ✓ Volume and concentration of one of them, acid or base, is known
- ✓ Concentration of the other solutions is required to be calculated whose volume is given.

How to solve?

1. Calculate the number of moles of solution whose concentration and volume is given

$$\text{moles} = \frac{C \times V}{1000}$$
2. Make the mole ratio of acid and base and calculate the moles of solution whose concentration is unknown. Usually there is one of the category will be there:

Acid : base

Acid : base

Acid : base

1 : 1

1 : 2

2 : 1

e.g

HCl : NaOH

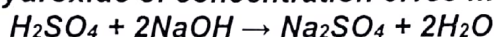
H₂SO₄ : NaOHHCl : Ca(OH)₂

3. calculate the concentration of solution whose volume is given.

$$\text{Conc.} = \frac{\text{moles}}{\text{Volume}} \times 1000$$

**Sample 1
Example**

12.0 cm³ of an aqueous solution of sulphuric acid exactly neutralised 20.0 cm³ of a solution of sodium hydroxide of concentration 0.150 mol/dm³.



Calculate the concentration, in mol/dm³ of the aqueous sulphuric acid.

$$n(\text{NaOH}) = 0.150 \times \frac{20.0}{1000} = 0.003 \text{ mol}$$

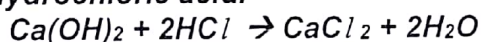
$$\text{mole ratio} = \text{H}_2\text{SO}_4 : \text{NaOH} = 1 : 2$$

$$n(\text{H}_2\text{SO}_4) = \frac{1}{2} \times 0.003 = 0.0015 \text{ mol}$$

$$\text{Conc.} = \frac{0.0015}{12.0} \times 1000 = 0.125 \text{ mol dm}^{-3}$$

**Sample 2
Example**

25.0 cm³ of an aqueous solution of calcium hydroxide is exactly neutralised by 18.0 cm³ of 0.040 mol/dm³ hydrochloric acid.



Calculate the concentration, in mol/dm³, of the aqueous calcium hydroxide.

$$\text{moles of HCl} = 0.040 \text{ mol dm}^{-3} \times \frac{18.0}{1000} = 0.00072 \text{ mol HCl}$$

$$\text{mol ratio} \quad \text{HCl} : \text{Ca(OH)}_2$$

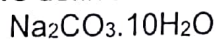
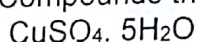
$$2 \text{ mol} : 1 \text{ mol}$$

$$0.00072 \text{ mol} : \frac{1}{2} \times 0.00072 = 0.00036 \text{ mol}$$

$$\text{Concentration} = \frac{0.00036 \text{ mol}}{25} \times 1000 = 0.0144 \text{ mol dm}^{-3}$$

CALCULATING WATER OF CRYSTALLISATION

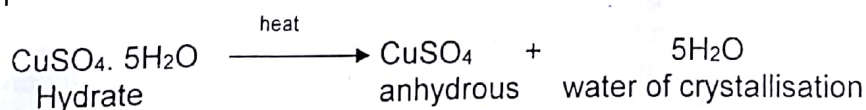
Hydrate: Compounds that have definite number of water molecules attached chemically.



Water of Crystallisation: The definite number of water molecules attached to the crystal structure.

Anhydrous Compound:

Compound formed when water of crystallization is lost by heating.



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Sample Example

12.0 g of anhydrous magnesium sulphate combines with 12.6 g of water to form hydrated magnesium sulphate. What is the formula of the hydrated magnesium sulphate? [$M_r(\text{MgSO}_4) = 120$]

	MgSO_4	:	H_2O
moles	$\frac{12.0}{120}$	:	$\frac{12.6}{18}$
	= 0.1		= 0.7
	1	:	7

Formula: $\text{MgSO}_4 \cdot 5\text{H}_2\text{O}$

PERCENTAGE PURITY

Sometimes the chemicals used in the preparation of other substances are not pure and contains one or more than one type of impurities in them. It is assumed in such problems that the impurities form no part of the reaction; therefore their mass should be excluded. In this way, through stoichiometric problems can help us to determine how much is the mass of pure substance in a given sample.

Some points to remember:

Students should keep in mind the following points in mind while dealing with such problems. However, these points pertain in general, and should not be considered absolute in-all cases.

- Usually, it is the mass of one of the reactant, which is provided in the question.
- Never use mass of impure (or natural sample) in calculating number of moles, or in any other direct calculations.
- Calculate the mass of the reactant from the given data by normal stoichiometric procedure.
- Calculate the percentage purity of the sample by the following formula:

$$\text{percentage purity} = \frac{\text{mass of pure sample present} \times 100}{\text{mass of (impure) sample}}$$

Sample Example

An impure sample of calcium carbonate (say limestone), CaCO_3 , contains calcium sulphate as an impurity. When hydrochloric acid was added to 8 g of the sample, 1200 cm^3 of gas was produced (measured at rtp). Calculate the percentage purity of the calcium carbonate sample.



$$\begin{array}{lcl} \text{Mass Volume ratio} & = & \text{CO}_2 : \text{CaCO}_3 \\ & & 24000\text{cm}^3 : 100\text{ g} \\ & & 1200\text{ cm}^3 : \frac{100}{24000} = 5.0\text{ g CaCO}_3 \end{array}$$

$$\text{Percentage Purity} = \frac{5.0}{8} \times 100 = 62.50\% \text{ pure CaCO}_3$$

PERCENTAGE YIELD

Yield: The amount of product obtained as a result of a chemical reaction is called as Yield.

Theoretical (Calculated Yield):

The amount of product obtained as a result of calculations based on a balanced chemical equation, without reference to the conditions affecting that amount is called as theoretical or calculated yield.

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Practical (Actual/Observed Yield):

The amount of product obtained by actual performance of reaction carried out in a laboratory in an industrial undertaking under the conditions affecting that amount is called as Practical (Actual/Observed) Yield.

Percentage Yield:

The actual yield expressed as a percentage of theoretical yield is called percentage (or practical) yield. It can be calculated as:

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

Why actual yield is usually less than theoretical yield:

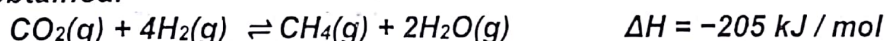
- 1) Inaccurate methods of measurement or observations.
- 2) Careless working or mishandling of the materials or apparatuses by the user.
- 3) Occurrence of side reactions.
- 4) Reversibility of the reaction.
- 5) Personal limitations of the observer.
- 6) Inability of the user to recover the product completely from the reaction mixture.
- 7) Spillage or vaporization of reactants/ products.
- 8) Limitations of purification techniques like crystallization, filtration, distillation etc.

Points to remember:

- 1) Mass/moles of one of the reactant and that of one of products is also given in the problem.
- 2) The term yield pertains only to the product and not to the reactant.
- 3) The amount of product given in the problem is always a actual/experimental yield.

Sample Example

When carbon dioxide reacts with hydrogen in a sealed container, an equilibrium mixture is obtained.



This reaction is exothermic. In an experiment, 220 g of carbon dioxide and an excess of hydrogen are reacted in a sealed container until an equilibrium is established. A mass of 46 g of methane is produced.

- (i) Calculate the mass of methane that should have been made if the percentage yield was 100%.

Mass ratio	CO ₂	:	CH ₄
	44g	:	16g
	220	:	$\frac{16}{44} \times 220 = 80 \text{ g}$

- (ii) Calculate the percentage yield of methane in this experiment.

$$\frac{46}{80} \times 100 = 57.5\%$$

LIMITING REACTANT

What is a Limiting Reactant:

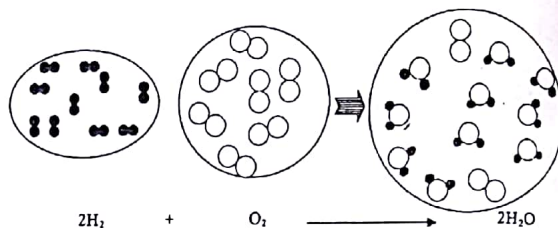
The reactant which is used earlier in a chemical reaction due to non-stoichiometric amounts of reactants taken, and is responsible for the amount of product produced is called Limiting reactant.

Non-Stoichiometric amounts:

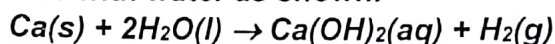
The amounts of reactants not taken according to the balanced chemical equation.

Illustration

According to equation 2 moles of Hydrogen gas react with one mole oxygen to form two moles of water. But if we take 10 moles of hydrogen and 7 moles of O_2 , then only 5 moles of water are produced, as 2 moles of O_2 are taken in excess. Therefore, H_2 is the limiting reactant.

**Sample Example 1**

Calcium reacts with water as shown.

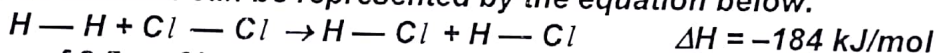


What is the total mass of the solution that remains when 40 g of calcium reacts with 100 g of water? ($A_r : Ca = 40, M_r = 18$)

Mass ratio	=	Ca	:	H_2O
		40 g	:	$(18 \times 2) = 36$ g
Mass unreactive H_2O	=	$100 - 36 = 64$ g		
Mass ratio	=	Ca	:	$Ca(OH)_2$
		40 g	:	74 g
Mass of the solution	=	$74 + 64 = 138$ g		

Sample Example 2

Hydrogen reacts with chlorine to make hydrogen chloride. The reaction is exothermic. The reaction can be represented by the equation below.



A mixture of 2.5 g of hydrogen and 142 g of chlorine is allowed to react.

(i) Which gas, hydrogen or chlorine, is in excess?

Explain your answer.

(ii) Calculate the mass of hydrogen chloride formed.

(i) Moles of hydrogen = $\frac{2.5 \text{ g}}{2} = 1.25 \text{ mol}$, Moles of chlorine = $\frac{142 \text{ g}}{71} = 2 \text{ mol}$

Mol ratio	H_2	:	Cl_2
	1 mol	:	1 mol
For 1.25 mol	:	1.25 mol required	but 2 mol are available
So, chlorine is in excess.			

(ii) Mol mass ratio H_2 : HCl

1 mol	:	$2(36.5) = 73$ g
1.25 mol	:	?

$$\text{Mass of HCl} = \frac{73}{1} \times 1.25 = 91.25 \text{ g HCl}$$

FINDING FORMULA OF A HYDROCARBON

A hydrocarbon has a molecular formula of C_xH_y . When it burns in a plentiful supply of air it will form carbon dioxide and steam. The general equation for burning is: -



It means one mole of hydrocarbon C_xH_y will produce x moles of carbon dioxide and $y/2$ moles of steam.

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**Sample 1
Example**

0.5 moles of a hydrocarbon on complete combustion, in a plentiful supply of air, produced 24 dm³ of carbon dioxide and 18 g of water. What is the formula of the hydrocarbon?

$$\text{moles of carbon dioxide} = \frac{24}{24} = 1 \text{ mol CO}_2$$

$$\text{moles of water} = \frac{18}{18} = 1 \text{ mol H}_2\text{O}$$

Mole ratio:	Hydrocarbon	:	CO ₂	:	H ₂ O.
	0.5 mol	:	1	:	1
	1 mol	:	2	:	2

To find the value of x and y:

$$\text{From the equation } x \text{ CO}_2 = 2 \text{ CO}_2$$

$$\text{Therefore } x = 2$$

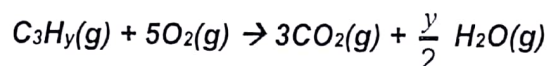
$$\text{And } \frac{y}{2} \text{ H}_2\text{O} = 2 \text{ H}_2\text{O}$$

$$y = 4$$

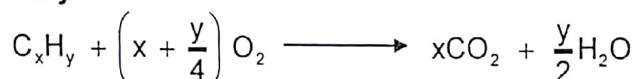
Hence, formula of hydrocarbon is C₂H₄.

**Sample 2
Example**

A hydrocarbon, C₃H_y, burns in air to form carbon dioxide and water.



What is the value of y?



$$x + \frac{y}{4} = 5$$

$$3 + \frac{y}{4} = 5$$

$$\frac{y}{4} = 5 - 3$$

$$y = 2 \times 4 = 8$$

$$\text{so } \text{C}_3\text{H}_y = \text{C}_3\text{H}_8$$

Group

* 58-71 Lanthanoid series
† 90-103 Actinoid series

a = relative atomic mass
 X = atomic symbol

The volume of one mole of any gas is 24 dm^3 at room temperature and pressure (r.t.p.).

EXAMINATION QUESTIONS

MOLES

MCQ'S

- Which quantity is the same for one mole of ethanol and one mole of ethane?
- A mass B number of atoms C number of molecules/ D volume at r.t.p.
- Which gas contains the same number of molecules as 9 g of water?
- A 2 g of hydrogen B 14 g of nitrogen C 32 g of oxygen D 44 g of carbon dioxide
- What is the mass of one mole of carbon-12?
- A 0.012 g B 0.024 g C 1 g D 12 g
4. What is the relative molecular mass M_r of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$?
- A 160 B 178 C 186 D 250
- What is the ratio of the number of molecules in 71 g of gaseous chlorine to the number of molecules in 2 g of gaseous hydrogen? [Relative atomic masses A_r (atomic weights): H, 1; Cl, 35.5]
- A 1:1 B 1:2 C 2:1 D 71:2
- What has the same mass as 0.25 mol of copper atoms?
- A 0.5 mol of oxygen molecules B 1 mol of sulfur dioxide molecules
C 1.5 mol of water molecules D 2 mol of oxygen atoms
- The relative atomic mass of chlorine is 35.5. What is the mass of 2 moles of chlorine gas?
- A 17.75 g B 35.5 g C 71 g D 142 g
- Using the Periodic Table for the relative atomic masses, which has the least mass?
- A 0.1 moles of silicon dioxide, SiO_2 B 0.5 moles of oxygen, O_2
C 0.5 moles of lithium, Li D 1.0 moles of ammonia, NH_3

CALCULATING MASS/MOLES OF A SUBSTANCE FOR A COMPOUND

MCQ'S

- All ammonium salts on heating with sodium hydroxide produce ammonia gas. From which ammonium salt can the greatest mass of ammonia be obtained?
- A 0.5 mol $(\text{NH}_4)_3\text{PO}_4$ B 0.5 mol $(\text{NH}_4)_2\text{SO}_4$ C 1.0 mol NH_4Cl D 1.0 mol NH_4NO_3
- What is the mass of aluminium in 204 g of aluminium oxide, Al_2O_3 ?
- A 26 g B 27 g C 54 g D 108 g
- What is the mass of oxygen contained in 72 g of pure water? [Relative atomic masses: H = 1; O = 16]
- A 16 g B 32 g C 64 g D 70 g
- What is the number of moles of hydrogen atoms in 3.2 g of methane?
- A 0.02 B 0.2 C 0.4 D 0.8

STRUCTURED QUESTIONS

1. What mass of molybdenum can be made from 125 g of molybdenum(VI) oxide? [A_r : Mo, 96]

2. What mass of titanium can be made from 125 g of titanium(IV) chloride?

MOLAR VOLUME

MCQ'S

- What is the ratio of the volume of 2 g of hydrogen to the volume of 16 g of methane, both volumes at r.t.p.
 A 1 to 1 B 1 to 2 C 1 to 8 D 2 to 1
- A volume of ethane, C_2H_6 , at r.t.p. has a mass of 20 g. What is the mass of an equal volume of propene, C_3H_6 , at r.t.p.
 A 20 g B 21 g C 28 g D 42 g

PERCENTAGE COMPOSITION

MCQ'S

- The relative molecular mass, M_r , of copper(II) sulphate, $CuSO_4$, is 160. The relative molecular mass, M_r , of water is 18. What is the percentage by mass of water in copper(II) sulphate crystals, $CuSO_4 \cdot 5H_2O$?
 A $\frac{18 \times 100}{160}$ B $\frac{5 \times 18 \times 100}{160 + 18}$ C $\frac{18 \times 100}{160 + 18}$ D $\frac{5 \times 18 \times 100}{160 + (5 \times 18)}$
- Which substance has the highest percentage by mass of nitrogen?
 A NH_4NO_3 $M_r = 80$ B $(NH_4)_2SO_4$ $M_r = 132$
 C $CO(NH_2)_2$ $M_r = 60$ D $(NH_4)_3PO_4$ $M_r = 149$
- Which fertiliser contains the greatest percentage by mass of nitrogen?
 A $(NH_4)_2HPO_4$ $M_r = 132$ B $(NH_4)_2SO_4$ $M_r = 132$
 C NH_4NO_3 $M_r = 80$ D $CO(NH_2)_2$ $M_r = 60$

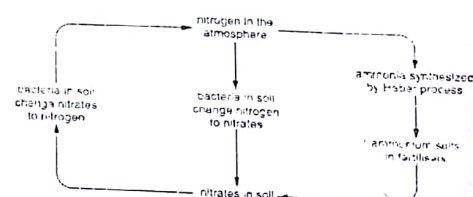
- The table below shows some information about two copper ores, tenorite and cuprite. Both contain copper oxide.

ore	formula of copper oxide in ore	oxidation number of copper	percentage of copper by mass
tenorite		+2	80.0%
cuprite	Cu_2O		

STRUCTURED QUESTIONS

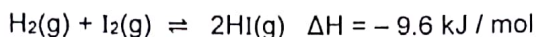
- What is the formula of the copper compound in tenorite? CuO
- What is the oxidation number of copper in cuprite, Cu_2O ? $+1$
- Calculate the percentage of copper by mass in Cu_2O . 88.9%

- A simplified diagram of the nitrogen cycle is shown below. Fertilisers are added to the soil to improve crop yields. A farmer has the choice of two fertilisers, ammonium nitrate, NH_4NO_3 , or Diammonium hydrogen phosphate, $(NH_4)_2HPO_4$. Show by calculation which of these fertilisers contains the greater percentage of nitrogen by mass. You must show your working.



- Fertilisers supply the essential elements, nitrogen, phosphorus and potassium for plant growth. A bag of fertiliser contains 500 g of ammonium sulfate, $(NH_4)_2SO_4$, and 500 g of potassium nitrate, KNO_3 . Calculate the percentage by mass of nitrogen in the bag of fertiliser.
 $\frac{14 \times 2 \times 500}{132 \times 500 + 101 \times 500} \times 100 = 17.54\%$
- Fertilisers are used to promote plant growth and increase crop yield. Three fertilisers are potassium chloride, potassium nitrate and ammonium phosphate. Ammonium phosphate is an ionic compound containing the phosphate ion, PO_4^{3-} .
 (i) Write the formula for ammonium phosphate. $(NH_4)_3PO_4$
 (ii) Calculate the percentage by mass of nitrogen in ammonium phosphate. 29.1%

5. When hydrogen and iodine are heated in a sealed container an equilibrium is reached with the product, hydrogen iodide.

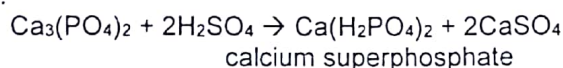


At 400 °C the equilibrium mixture contains 0.4000 moles of hydrogen, 0.07560 moles of iodine and 1.344 moles of hydrogen iodide. Calculate the percentage of iodine molecules, I_2 , by mass in this equilibrium mixture.

6. The percentage of methane by volume in the air is 0.00014%. Calculate the mass of methane in 1 000 dm³ of air.

7. Ammonia is used to make fertilisers such as ammonium phosphate, $(\text{NH}_4)_3\text{PO}_4$. Calculate the percentage by mass of nitrogen in ammonium phosphate.

8. Sulfuric acid is used to make superphosphate fertilisers. A mixture of the fertiliser and calcium sulfate is formed. This mixture is used by farmers.



Calculate the percentage by mass of calcium sulfate in the mixture of calcium superphosphate and calcium sulfate. (The relative formula mass of calcium superphosphate is 234.)

9. Iron is extracted in a blast furnace. The raw materials required are

• iron ore, which contains iron(III) oxide, Fe_2O_3 , • limestone • coke (carbon) • air.
Another ore of iron contains an oxide with the formula Fe_3O_4 . Calculate the percentage by mass of iron in Fe_3O_4 .

10. Ammonia is used to make fertilisers. The table gives some information about two fertilisers made from ammonia.

fertiliser	formula	relative formula mass (M_r)
ammonium nitrate	NH_4NO_3	80
Urea	$(\text{NH}_2)_2\text{CO}$	60

Use the data in the table to show that urea contains a greater percentage by mass of nitrogen than ammonium nitrate.

11. Ammonia is used to make fertilisers.

The table gives some information about two fertilisers made from ammonia.

fertiliser	formula	relative formula mass (M_r)
ammonium nitrate	NH_4NO_3	80
urea	$(\text{NH}_2)_2\text{CO}$	60

Use the data in the table to show that urea contains a greater percentage by mass of nitrogen than ammonium nitrate.

[2]

12. Sulfuric acid is used to make the fertiliser potassium sulfate, K_2SO_4 . Calculate the percentage by mass of potassium in this fertiliser.

EMPIRICAL FORMULA

MCQ'S

- The empirical formula of a liquid compound is C_2H_4O . To find the empirical formula, it is necessary to know the

A	density of the compound.	B	percentage composition of the compound.
C	relative molecular mass of the compound.	D	volume occupied by 1 mole of the compound
- Analysis of a sample of an oxide of nitrogen gave the following data.

•	percentage by mass of nitrogen	47%
•	percentage by mass of oxygen	53%

 What is the empirical formula of this oxide? [Ar: N, 14; O, 16]

A	NO	B	NO ₂	C	N ₂ O	D	N ₂ O ₃
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- What is the empirical formula of a compound containing 12 g of carbon, 2 g of hydrogen and 16 g of oxygen only?

A	CHO	B	CHO ₂	C	CH ₂ O	D	C ₂ H ₂ O
---	-----	---	------------------	---	-------------------	---	---------------------------------
- An organic compound has the molecular formula $C_8H_{16}O_4$. What is the empirical formula of the compound?

A	C_2H_4O	B	$C_4H_8O_2$	C	$C_6H_{12}O_3$	D	$C_8H_{16}O_4$
---	-----------	---	-------------	---	----------------	---	----------------
- A compound containing only the elements carbon and hydrogen has 80.0% by mass of carbon. What is its empirical formula?

A	C_3H	B	CH_3	C	CH_4	D	C_2H_6
---	--------	---	--------	---	--------	---	----------
- A compound contains 70% by mass of iron and 30% by mass of oxygen. What is its empirical formula? [Ar: O, 16; Fe, 56]

A	FeO	B	Fe ₂ O ₃	C	Fe ₃ O ₂	D	Fe ₃ O ₄
---	-----	---	--------------------------------	---	--------------------------------	---	--------------------------------

STRUCTURED QUESTIONS

- Ethene can also be converted into a compound that contains carbon, hydrogen and oxygen. A sample of the compound was analysed and found to contain 0.72 g of carbon, 0.18 g of hydrogen and 0.96 g of oxygen. Show that the empirical formula of the compound is CH_3O .

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- A sample of a compound of iron is analysed. The sample contains 0.547 g of potassium, 0.195 g of iron, 0.252 g of carbon and 0.294 g of nitrogen. Calculate the empirical formula of this compound.

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- Dry air contains mainly nitrogen and oxygen together with small amounts of argon and carbon dioxide. A small amount of xenon is present in the air. Several compounds of xenon have been made in recent years. A compound of xenon contained 9.825 g of xenon, 1.200 g of oxygen and 5.700 g of fluorine. Determine the empirical formula of this compound.

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- Many inks contain salts of the metals potassium, iron, cobalt and nickel in addition to ethanoic acid and gallic acid. Analysis of 21.25 g of gallic acid showed that it contained 10.50 g of carbon, 0.75 g of hydrogen and 10.00 g of oxygen. Show that the empirical formula of gallic acid is $C_7H_6O_5$.

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Carbon and graphite are two forms of carbon.

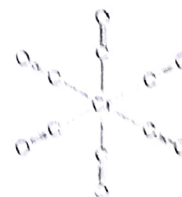
Carbon monoxide can be formed by the reduction of carbon dioxide with red-hot carbon. Carbon monoxide reacts with chromium to form chromium carbonyl. The structure of chromium carbonyl is shown below. Write the empirical formula for chromium carbonyl.



diamond



graphite



Analysis of V showed that it contained 79.9 % copper and 20.1 % oxygen by mass. Calculate the empirical formula of the black solid V.

Iron(II) sulfate crystals decompose when heated to give three gases U, V and W and an orange-brown solid T.

- Gas U was tested with filter paper soaked with acidified potassium dichromate(VI). The filter paper changed colour from orange to green.
- Analysis of gas V showed it contained 40.0% sulfur and 60.0% oxygen by mass.
- When gas W was condensed it formed a colourless liquid that turned anhydrous copper(II) sulfate from white to blue.
- Solid T was dissolved in dilute nitric acid. Aqueous ammonia was added drop by drop and a red-brown precipitate was obtained. Calculate the empirical formula of gas V.

Analysis of compound X shows it has the following composition.

element	percentage by mass
nitrogen	11.1
hydrogen	3.20
chromium	41.3
oxygen	44.4

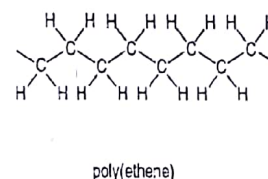
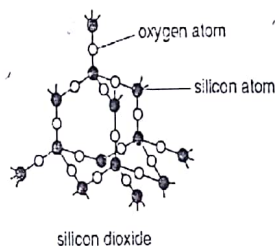
Show that X has the formula $N_2H_8Cr_2O_7$.

Parts of the structures of silicon dioxide and poly(ethene) are shown below.

The compound used to make the monomer of the silicone fluid has the following composition by mass.

C = 18.6 g, Cl = 55.0 g, H = 4.65 g, Si = 21.7 g

Deduce the empirical formula of this compound.
empirical formula



- 10 The farmer uses another fertiliser. This fertiliser has the following percentage composition by mass.

Ca, 17.1%; H, 1.7%; P, 26.5%; O, 54.7%

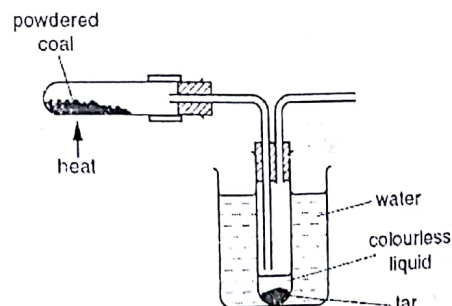
Calculate the empirical formula of this compound and suggest the formula of the anion present in the fertiliser.

- 11 Coal is a mixture of carbon compounds with a small amount of sulfur. A sample of coal is heated in the absence of air using the apparatus shown. The tar also contains a compound with the following composition.

element	percentage by mass
carbon 76.60	
hydrogen 6.38	
oxygen 17.02	

- 12 Plants contain many different types of carbon compounds. When wood chips are heated in the absence of air, a colourless distillate is formed. The distillate also contains a compound with the following composition.

element	percentage by mass
carbon	37.5
hydrogen	12.5
oxygen	50.0



empirical formula

- 13 A 11.09 g sample of an oxide of copper contains 9.86 g of copper. Deduce the empirical formula of this oxide of copper.

empirical formula

STRUCTURED QUESTIONS

1. Both ethanoic acid and butanoic acid are found in some plants and bacteria. Butanoic acid can be converted into an ester by heating it with an alcohol and a few drops of concentrated sulphuric acid. A sample of an ester contains 0.18 g of carbon, 0.03 g of hydrogen and 0.08 g of oxygen. The relative molecular mass of the ester is 116. Calculate both the empirical and molecular formulae of this ester.

mol. ratio $\frac{C}{12} = \frac{H}{1} = \frac{O}{16}$ $\frac{0.039}{12} = \frac{0.009}{1} = \frac{0.009}{16}$

2. Sorrel is a small green plant. Analysis of 10.0 g of carboxylic acid X shows that it contains 2.67 g carbon, 0.220 g hydrogen and 7.11 g oxygen.

- (i) Deduce the empirical formula of X. $\frac{2.67}{12} = 0.2225$ $\frac{0.2225}{0.2225} = 1$ $\frac{2.11}{18} = 0.1172$ $\frac{0.1172}{0.1172} = 1$
- (ii) The relative molecular mass of X is 90. Deduce the molecular formula of X.

Empirical formula mass = $12 + (16 \times 2) = 44$
Molecular formula = $3 \times (CH_2O) = C_3H_6O_3$

3. Alkane G contains 84% carbon by mass. Calculate the molecular formula for G. molecular formula?

- 4 (a) Construct the equation for the complete combustion of cyclohexene.

- (b) Another cycloalkene has the following percentage composition by mass.
C, 88.2%; H, 11.8%

- (i) Use the percentage composition by mass to show that the empirical formula of this cycloalkene is C_5H_8 .

$$\frac{C}{H} = \frac{0.826}{0.102} = \frac{8.09}{1} \quad \frac{O}{H} = \frac{0.175}{0.102} = \frac{1.72}{1}$$

- (ii) The cycloalkene has a relative molecular mass, M_r , of 68.

Draw the structure of the cycloalkene, showing all of the atoms and all of the bonds.

Methane, ethane and propane are all gases at room temperature.

Ethane reacts with chlorine in the presence of ultraviolet light to give a number of different compounds.

A 1.00 g sample of one of these compounds contains 0.040 g of hydrogen, 0.242 g of carbon and 0.718 g of chlorine.

(i) Calculate the empirical formula of this compound.

empirical formula[2]

(ii) The relative molecular mass of the compound is 99.

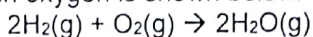
Deduce the molecular formula of the compound.

.....[1]

STOICHIOMETRY

MCQ'S

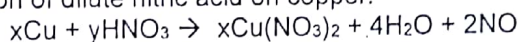
The equation for the burning of hydrogen in oxygen is shown below.



Which information does this equation give about the reaction?

- A 36 g of steam can be obtained from 16 g of oxygen.
- B 2 g of hydrogen combine with 1 g of oxygen.
- ☒ C 2 mol of steam can be obtained from 1 mol of oxygen.
- D 2 atoms of hydrogen combine with 2 atoms of oxygen.

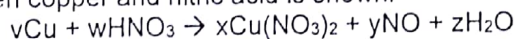
The equation represents the action of dilute nitric acid on copper.



What are the values of x and y?

- A x = 1, y = 4
- B x = 1, y = 8
- C x = 3, y = 4
- ☒ D x = 3, y = 8

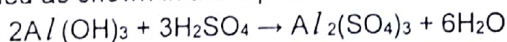
The equation for the reaction between copper and nitric acid is shown.



v, w, x, y and z are whole numbers. Which values of v, w, x, y and z balance the equation?

	V	W	X	Y	Z
A	1	2	1	1	1
B	1	4	1	2	2
C	3	4	3	2	2
D	3	8	3	2	4

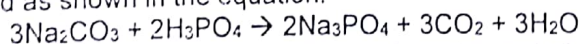
Aluminium sulphate can be obtained as shown in the equation.



How many moles of sulphuric acid are needed to produce 0.5 mol of aluminium sulphate?

- A 0.5
- B 1.0
- C 1.5
- D 3.0

Carbon dioxide can be obtained as shown in the equation.



How many moles of phosphoric acid, H_3PO_4 , are needed to produce 1.5 mol of carbon dioxide?

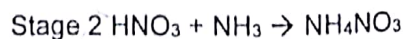
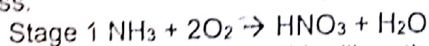
- A 0.5
- B 1.0
- C 1.5
- D 2.0

One mole of magnesium is dissolved in excess aqueous ethanoic acid, CH_3COOH .

How many moles of hydrogen, H_2 , will be produced?

- A 0.5
- ☒ B 1
- C 2
- D 4

The fertiliser ammonium nitrate (NH_4NO_3 , $M_r = 80$) is manufactured from ammonia (NH_3 , $M_r = 17$) by a two-stage process.

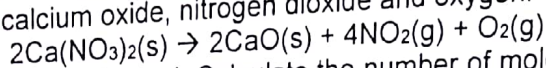


What is the maximum mass of fertiliser that can be made if only 17 tonnes of ammonia is available?

- A 34 tonnes
- ☒ B 40 tonnes
- C 80 tonnes
- D 97 tonnes

STRUCTURED QUESTIONS

1. The carbonates of many metallic elements decompose when heated. The nitrates of metallic elements also decompose when heated. Calcium nitrate decomposes to form calcium oxide, nitrogen dioxide and oxygen.

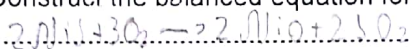


A 0.010 mol sample of calcium nitrate is heated. Calculate the number of moles of gas produced when this sample is completely decomposed.

mol rate $\text{Ca}(\text{NO}_3)_2$ 0.010 mol 0.010 mol 0.020 mol
 2 mol (1.1) 5 mol 0.020 mol

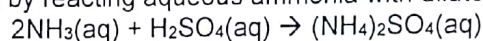
2. Nickel is a transition element. It is manufactured in a four-stage process from nickel(II)sulphide, NiS.
- Stage 1 – nickel(II) sulphide is heated in air to form nickel(II) oxide and sulphur dioxide.
 - Stage 2 – nickel(II) oxide is heated with carbon to give impure nickel.
 - Stage 3 – impure nickel is reacted with carbon monoxide to make nickel tetracarbonyl, $\text{Ni}(\text{CO})_4$.
 - Stage 4 – nickel tetracarbonyl is decomposed to give pure nickel.

- (i) Construct the balanced equation for the reaction in stage 1.



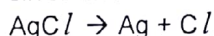
- (ii) Calculate the mass of sulphur dioxide that is formed when 182 kg of nickel sulphide is heated in air.

3. Ammonium sulphate can be made by reacting aqueous ammonia with dilute sulphuric acid.

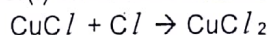


Calculate the mass of ammonium sulphate that can be made from 51 g ammonia.

4. Sunglasses can be made from photochromic glass. When bright light strikes photochromic glass it darkens. Photochromic glass contains small amounts of silver chloride, AgCl , and copper(I) chloride, CuCl . In the presence of bright light, silver chloride decomposes into silver atoms which make the glass go dark, and into chlorine atoms.

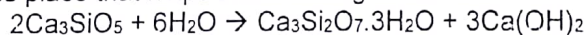


Chlorine atoms immediately react with copper(I) chloride to make copper(II) chloride.



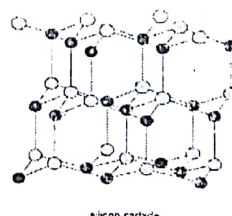
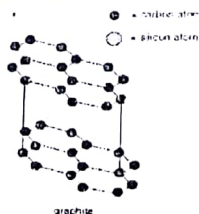
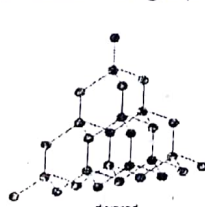
When the exposure to bright light ends, silver atoms reduce copper(II) chloride back into copper(I) chloride and silver chloride. Calculate the maximum mass of silver that can be formed when 0.287 g of silver chloride decomposes.

5. This question is about calcium compounds. Cement is made by heating calcium carbonate and clay together at a very high temperature. One of the compounds produced is a form of calcium silicate, Ca_3SiO_5 . In the presence of water a chemical reaction takes place that helps in the setting of cement.



Calculate the mass of calcium hydroxide formed from 912 g of Ca_3SiO_5 .

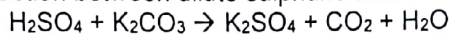
3. The structures of diamond, graphite and silicon carbide are shown below.



When a 1.20 g sample of graphite is completely burnt in oxygen, 4.40 g of carbon dioxide are produced. What mass of carbon dioxide is made when a 1.20 g sample of diamond is completely burnt in oxygen?

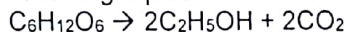
mass of carbon dioxide:

-) Fertilisers are soluble salts containing one or more of the essential elements required for plant growth. Potassium sulphate can be prepared by the reaction between dilute sulphuric acid and potassium carbonate.



Calculate the mass of potassium sulphate that can be prepared from 3.45 g of potassium carbonate.

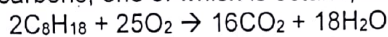
The table shows the formula of the first three members of the alcohol homologous series. Ethanol can be manufactured from either ethene or glucose. The fermentation of glucose can be represented by the following equation.



Calculate the maximum mass of ethanol that could be made from 36 tonnes of glucose.

alcohol	formula
methanol	CH_3OH
ethanol	$\text{C}_2\text{H}_5\text{OH}$
propanol	$\text{C}_3\text{H}_7\text{OH}$

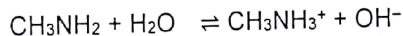
Petrol (gasoline) is a mixture of hydrocarbons, one of which is octane, C_8H_{18} . Octane burns in air.



A petrol-powered motor car travels at a constant speed of 80 km/h. For every kilometer travelled 108 g of carbon dioxide are formed. When the motor car travels 100 km calculate the mass of carbon dioxide emitted by the car,

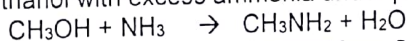
- ii) the mass of petrol burned by the car assuming that petrol is 100% octane.

Methylamine, CH_3NH_2 , is a base which has similar properties to ammonia. When methylamine dissolves in water, the following equilibrium is set up.



- a) Methylamine is a gas. Calculate the volume occupied by 6.2 g of methylamine at room temperature and pressure.

- b) Methylamine is made by reacting methanol with excess ammonia under pressure in the presence of a catalyst.



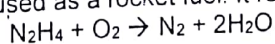
Calculate the theoretical yield of methylamine that can be obtained from 240 kg of methanol.

10 Paraffin (kerosene) is a mixture of hydrocarbons. It is used as a fuel for the jet engines of an aircraft.

- a) When paraffin burns in a jet engine some nitrogen monoxide, NO , is formed. This is because the high temperature of the engine allows nitrogen to react with oxygen. Write an equation to describe how nitrogen monoxide is formed in this reaction. Calculate the mass of nitrogen monoxide formed from 55 kg of nitrogen.

- b) Using the information that one mole contains 6.02×10^{23} particles, calculate the number of electrons in one mole of NO molecules.

11. Hydrazine, N_2H_4 , is a liquid that has been used as a rocket fuel. It reacts with oxygen as shown in the equation.

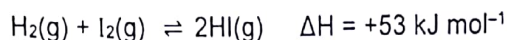


This reaction is highly exothermic.

- i) Calculate the volume of oxygen, measured at room temperature and pressure, needed to completely combust 1.00 tonne of hydrazine. [One tonne is 10^6 grams. One mole of any gas at room temperature and pressure occupies a volume of 24 dm^3 .]

- (ii) A rocket burns hydrazine in an atmosphere of oxygen. Both hydrazine and oxygen are stored in the rocket as liquids. Suggest why oxygen is stored as a liquid rather than as a gas.
-

- 12 Hydrogen and iodine react together to form hydrogen iodide in a reversible redox reaction. The forward reaction is endothermic.



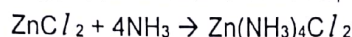
Hydrogen and hydrogen iodide are colourless gases whereas iodine gas is purple.

Calculate the maximum mass of hydrogen iodide that can be made from 45.3 g of hydrogen.

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13. Solid zinc chloride absorbs ammonia to form tetrammine zinc chloride, $\text{Zn}(\text{NH}_3)_4\text{Cl}_2$.

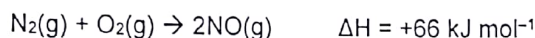


Calculate the maximum yield, in grams, of tetrammine zinc chloride formed when 3.4 g of zinc chloride reacts with excess ammonia.

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- 14 Uranium is a radioactive metal. It has two main isotopes, uranium-235 with a nucleon number of 235 and uranium-238 with a nucleon number of 238. Uranium is manufactured from uranium(IV) oxide, UO_2 , in a two-step process.
- Step 1 – uranium(IV) oxide is heated with hydrogen fluoride to make uranium(IV) fluoride, UF_4 , and water.
- Step 2 – uranium(IV) fluoride is reduced by magnesium to give uranium and one other product.
- Calculate the mass of uranium that can be made from 1.00 tonne of uranium(IV) oxide. [One tonne is one million grams.]
-

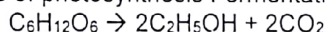
- 15 Nitric oxide, NO , is an atmospheric pollutant formed inside car engines by the reaction between nitrogen and oxygen.



This reaction is endothermic. Calculate the mass of nitric oxide formed when 100 g of nitrogen reacts completely with oxygen.

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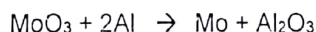
- 16 Glucose, $\text{C}_6\text{H}_{12}\text{O}_6$, is one of the products of photosynthesis. Fermentation converts glucose into ethanol, a biofuel.



Calculate the maximum mass of ethanol that can be made from 1 tonne of glucose. [One tonne is one million grams.]

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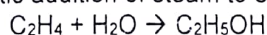
- 17 Molybdenum, atomic number 42, is manufactured by the displacement reaction between molybdenum(VI) oxide and aluminium.



Calculate the mass of aluminium needed to make 1 tonne of molybdenum. [1 tonne is one million grams.]

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- 18 Ethanol can be manufactured by the catalytic addition of steam to ethene.



If the reactants are not recycled, only 5% of the ethene is converted to ethanol. Calculate the mass of ethanol formed from 0.4 tonnes of ethene when only 5% of the ethene is converted to ethanol. [1 tonne is 1 000 000 grams]

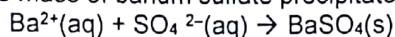
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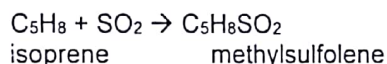
- 9 A sample containing 0.233 g of an unknown Group I element is added to excess ethanol. The volume of hydrogen gas formed at room temperature and pressure is 400 cm³. Calculate the relative atomic mass, A_r , of the Group I element and suggest the identity of the element.

relative atomic mass = identity of the element =

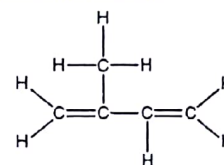
- 10 The concentration of sulfate ions in seawater is 1.24 g / dm³. Excess aqueous barium chloride is added to a 50.0 cm³ sample of seawater. Calculate the mass of barium sulfate precipitated in this reaction.



- 11 When rubber is heated in the absence of air, a small amount of isoprene is formed. The structure of isoprene is shown. Isoprene reacts with sulfur dioxide to form methylsulfolene.

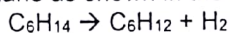


Calculate the maximum mass of methylsulfolene that can be formed from 100 g of isoprene.



mass of methylsulfolene = g

- 12 Cyclohexane can be manufactured from hexane as shown in the equation.



Calculate the mass of cyclohexane that can be made from 258 g of hexane.
[M_r of cyclohexane = 84]

mass of cyclohexane = g

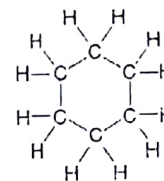
- 13 Cyclohexane, C_6H_{12} , is a cycloalkane. Cycloalkanes react in a similar way to alkanes.

- (a) Cyclohexane can be manufactured from hexane as shown in the equation.



Calculate the mass of cyclohexane that can be made from 258 g of hexane.
[M_r of cyclohexane = 84]

Molar mass of C₆H₁₂ = 84



cyclohexane

mass of cyclohexane = g [2]

- (b) Another cycloalkane has the following percentage composition by mass.
C, 85.7%; H, 14.3%

- (i) Use the percentage composition by mass to show that the empirical formula of the cycloalkane is CH_2 .

- (ii) Draw the structure of the cycloalkane with an M_r of 56, showing all of the atoms and all of the bonds.

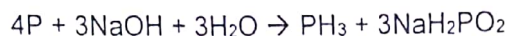
Calculate the mass of zinc in the sample of brass.

$$\begin{array}{l} \text{H}_2 \quad \text{Zn} \\ 24 \text{ dm}^3 \quad (24) \\ 0.075 \text{ dm}^3 \quad (0.075) \\ \frac{0.075}{24} \times 0.07 = 0.19 \text{ g Zn} \end{array}$$

Calculate the percentage of zinc in the sample of brass.

$$\% \text{ of Zn} = \frac{0.19}{1.12} \times 100 = 16.96\%$$

Phosphine, PH_3 , is a gas which has a smell of garlic. It is formed when white phosphorus is warmed with aqueous sodium hydroxide.



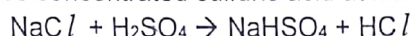
Calculate the maximum mass of phosphine formed when 1.86 g of phosphorus reacts with excess aqueous sodium hydroxide.

$$\begin{array}{l} \text{Ar of P} = 31 \quad \text{Ar of PH}_3 = 34 \quad \text{Mass ratio} \quad \text{P} \quad \text{PH}_3 \\ 1.86 \text{ g} \quad 1.86 \text{ g} \quad 4(31) = 124 \quad 134 \times 0.075 = 0.114 \end{array}$$

Calculate the volume of phosphine formed from 1.86 g of phosphorus at r.t.p.

$$\text{Moles of PH}_3 = 0.11 - 0.015 \text{ mol} \quad \text{Ar of PH}_3 = 0.015 \times 24 = 0.36 \text{ dm}^3$$

Hydrogen fluoride, hydrogen chloride and hydrogen iodide are all acidic gases. A student makes hydrogen chloride by reacting sodium chloride with excess concentrated sulfuric acid at room temperature and pressure.



Calculate the maximum volume of hydrogen chloride that can be made from 0.2 moles of sodium chloride at room temperature and pressure.

$$\text{Mol-Volume ratio} \quad 1 \text{ mol} \quad 24 \text{ dm}^3 \quad 0.2 \times 24 = 4.8 \text{ dm}^3$$

The overall reaction for the electrolysis of aqueous sodium hydroxide is shown below.



Some submarines use this reaction to provide oxygen for the occupants to breathe. Calculate the mass of water which must be electrolysed to make 2500 dm³ of oxygen at room temperature and pressure. [One mole of any gas at room temperature and pressure occupies a volume of 24 dm³.]

$$\text{Vol-mass ratio} \quad 24 \text{ dm}^3 \quad 32 \text{ g} \quad 2500 \text{ dm}^3 \quad 2500 \times \frac{32}{24} = 3333.33 \text{ g}$$

Hydrogen-oxygen fuel cells are used to generate electricity. The overall reaction in a hydrogen-oxygen fuel cell is shown below.



A hydrogen-oxygen fuel cell uses 2000 dm³ of hydrogen measured at room temperature and pressure. Calculate the volume of oxygen, measured at room temperature and pressure, used by the fuel cell. [One mole of any gas at room temperature and pressure occupies a volume of 24 dm³.]

$$\text{Vol ratio} \quad 2 \text{ H}_2 \quad 1 \text{ O}_2 \quad 2000 \text{ dm}^3 \quad 2000 \div 2 = 1000 \text{ dm}^3$$

The relative atomic mass of magnesium can be determined in the laboratory by finding the volume of hydrogen given off when magnesium reacts with hydrochloric acid.



0.036 g of magnesium reacts at room temperature and pressure with excess hydrochloric acid to produce 36 cm³ of hydrogen. 1 mole of any gas at room temperature and pressure occupies 24 dm³. Show by calculation that the relative atomic mass of magnesium is 24.

$$\text{H}_2 \quad 1 \text{ mol} \quad 24 \text{ dm}^3 \quad 0.036 \text{ g} \quad 0.036 \times 24000 = 0.864 \text{ g}$$

Potassium chlorate, KClO_3 , decomposes to form potassium chloride and oxygen.



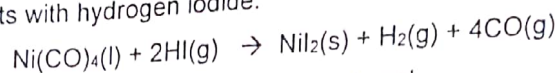
(a) Calculate the percentage by mass of oxygen in potassium chlorate.

$$39\%$$

(b) Calculate the maximum volume of oxygen formed at room temperature and pressure when 12.25 g of potassium chlorate is completely decomposed.

$$\text{KClO}_3 = 122.5 \quad 12.25$$

- 9 Nickel carbonyl, $\text{Ni}(\text{CO})_4$, reacts with hydrogen iodide.



- (a) Calculate the percentage by mass of nickel in nickel carbonyl.

54.5 %

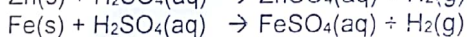
- (b) Calculate the maximum volume of gas formed at room temperature and pressure when 1.71 g of nickel carbonyl reacts completely with hydrogen iodide.

mol nickel
carbonyl
= $\frac{1.71}{171}$

mol gases = $(0.01) \times 5$

maximum volume of gas formed

- 10 Both zinc and iron react with dilute sulfuric acid.



A 0.65 g sample of zinc is reacted with excess sulfuric acid.

Calculate the volume of hydrogen, measured at room temperature and pressure, formed in this reaction.

Moles of Zn and $\text{H}_2 = 0.01$
Vol of $\text{H}_2 = 0.24 \text{ dm}^3$

volume of hydrogen = dm^3

CONCENTRATION OF SOLUTIONS

- 2 dm^3 of aqueous sodium hydroxide of concentration 5 mol / dm^3 were required for an experiment. How many moles of sodium hydroxide were needed to make up this solution?
A 2.5 B 5 C 7 D 10
- What is the concentration of iodine, I_2 , molecules in a solution containing 2.54 g of iodine in 250 cm^3 of solution?
A 0.01 mol / dm^3 B 0.02 mol / dm^3 C 0.04 mol / dm^3 D 0.08 mol / dm^3
- How many moles per dm^3 of gaseous carbon dioxide are there if 4.4 g occupies 500 cm^3 ?
A 0.1 mol / dm^3 B 0.2 mol / dm^3 C 2.2 mol / dm^3 D 8.8 mol / dm^3
- What is the concentration of a solution containing 1.0 g of sodium hydroxide in 250 cm^3 of solution?
A 0.025 mol / dm^3 B 0.10 mol / dm^3 C 0.25 mol / dm^3 D 1.0 mol / dm^3

MCQ'S

STRUCTURED QUESTIONS

- A student prepares some crystals of hydrated sodium sulfate by titrating aqueous sodium hydroxide with sulfuric acid. The student uses 25.0 cm^3 of 1.60 mol / dm^3 sodium hydroxide to prepare the crystals.
$$2\text{NaOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) + 8\text{H}_2\text{O}(\text{l}) \rightarrow \text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}(\text{s})$$

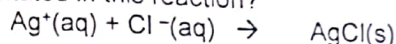
Calculate the maximum mass of hydrated sodium sulfate crystals that can be formed.
.....
.....

Seawater contains many dissolved ions. The table shows the concentration of some of these ions in a typical sample of seawater.

Calculate the concentration, in mol / dm³, of sulfate ions in seawater.

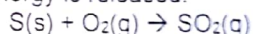
ion	formula	concentration, g / dm ³
chloride	Cl ⁻	18.00
sodium	Na ⁺	10.56
sulfate	SO ₄ ²⁻	2.65
magnesium	Mg ²⁺	1.26
calcium	Ca ²⁺	0.40
potassium	K ⁺	0.38
hydrogencarbonate	HCO ₃ ⁻	0.14

Excess aqueous silver nitrate is added to a 25.0 cm³ sample of seawater. What mass of silver chloride is precipitated in this reaction?



Magnesium chloride, MgCl₂, is present in seawater at a concentration of 1.26 g / dm³. Calculate the concentration of chloride ions, in mol / dm³, arising from the magnesium chloride in seawater.

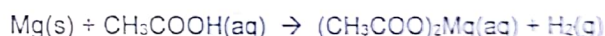
When one mole of sulfur burns, 247 kJ of energy is released.



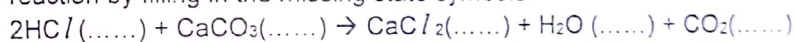
Calculate the volume of 0.10 mol / dm³ sodium hydroxide which contains 3.2 g of sodium hydroxide.

volume of aqueous sodium hydroxide = cm³

Magnesium reacts with ethanoic acid to make magnesium ethanoate and hydrogen. Write the equation for this reaction. Use the equation to calculate the mass of magnesium needed to react completely with 50 cm³ of 1.0 mol/dm³ of ethanoic acid.



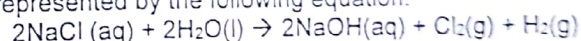
An experiment was carried out to measure the rate of reaction between excess powdered calcium carbonate and dilute acids. In Experiment 1, 25 cm³ of 1.5 mol/dm³ hydrochloric acid was used. Complete the equation for the reaction by filling in the missing state symbols.



i) Calculate the total volume of carbon dioxide that is made from this reaction at r.t.p.

Chlorine can be made by the electrolysis of concentrated aqueous sodium chloride.

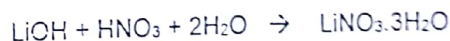
The overall process can be represented by the following equation.



55 dm³ of 3.5 mol / dm³ aqueous sodium chloride is electrolysed.

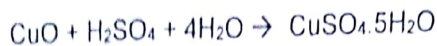
What is the maximum volume of chlorine that can be formed, measured at room temperature and pressure?

3 A student prepared some crystals of hydrated lithium nitrate by reacting aqueous lithium hydroxide with dilute nitric acid.



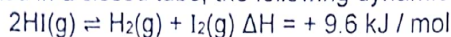
The student used 20.0 cm³ of 0.500 mol / dm³ of lithium hydroxide to prepare the crystals. Calculate the maximum mass of hydrated lithium nitrate crystals that could be made.

- 9 A student prepared some crystals of hydrated copper(II) sulfate by reacting excess insoluble copper(II) oxide with dilute sulfuric acid.



The student used 15.0 cm^3 of 2.00 mol / dm^3 sulfuric acid to prepare the crystals.
Calculate the maximum mass of hydrated copper(II) sulfate crystals that could be made.

- 10 When hydrogen iodide, HI, is heated in a closed tube, the following dynamic equilibrium is established.



The table shows the concentrations of HI(g), H₂(g) and I₂(g) in the equilibrium mixture at 25 °C and 450 °C

substance	concentration at 25 °C / mol / dm ³	concentration at 450 °C / mol / dm ³
HI(g)	0.94	0.79
H ₂ (g)	0.033	0.11
I ₂ (g)	0.033	0.11

The tube has a volume of 50 cm^3 .

Calculate the mass of hydrogen iodide in the equilibrium mixture at 25 °C.

CALCULATION OF IONS FROM A COMPOUND

- What is the concentration of hydrogen ions in 0.05 mol / dm^3 sulfuric acid?
A 0.025 g / dm^3 B 0.05 g / dm^3 C 0.10 g / dm^3 D 2.0 g / dm^3
- A student titrates 20.0 cm^3 of a metal hydroxide, $\text{M}(\text{OH})_2$, of concentration 0.060 mol / dm^3 with a strong acid of concentration 0.050 mol / dm^3 . It requires 24.0 cm^3 of acid to neutralise the metal hydroxide.
 - Calculate the number of moles of acid in 24.0 cm^3 of the acid.
 - Calculate the number of moles of OH^- ions in 20.0 cm^3 of the metal hydroxide.

VOLUMETRIC EXPERIMENTS

- When added to 20 cm^3 of 0.5 M sulphuric acid, which substance would give a neutral solution?
A 20 cm^3 of 0.5 M sodium hydroxide B 10 cm^3 of 0.5 M sodium hydroxide
C 40 cm^3 of 1.0 M sodium hydroxide D 20 cm^3 of 1.0 M sodium hydroxide
- 15.0 cm^3 of 1.0 mol / dm^3 potassium hydroxide just neutralise 20.0 cm^3 of a solution of nitric acid. What is the concentration of the acid?
A 0.75 mol / dm^3 B 1.0 mol / dm^3 C 1.5 mol / dm^3 D 7.5 mol / dm^3
- The equation shown represents the neutralisation of aqueous sodium hydroxide with dilute sulfuric acid.
 $2\text{NaOH}(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
How much sulfuric acid is required to neutralise 100 cm^3 of 1.0 mol / dm^3 NaOH?
A 50 cm^3 of 2.0 mol / dm^3 sulfuric acid B 100 cm^3 of 1.0 mol / dm^3 sulfuric acid
C 25 cm^3 of 0.5 mol / dm^3 sulfuric acid D 50 cm^3 of 1.0 mol / dm^3 sulfuric acid

- Potassium reacts with water as shown in the equation.
 $2\text{K}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{K}^+(\text{aq}) + 2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g})$

STRUCTURED QUESTIONS

A sample of 0.195 g of potassium was added to 500 cm³ of cold water. When the reaction was finished, 100 cm³ of 0.100 mol/dm³ hydrochloric acid was added to form solution X.

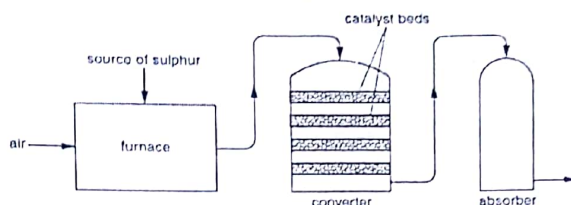
Calculate the number of moles of hydroxide ions formed when the potassium was added to water.

i) Calculate the number of moles of hydrogen ions in 100 cm³ of 0.100 mol/dm³ hydrochloric acid.

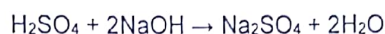
ii) Give an ionic equation to represent the neutralisation reaction.

iii) Suggest a pH value for solution X. Explain your answer.

The diagram shows the stages in the manufacture of sulphuric acid.



2.0 cm³ of an aqueous solution of sulphuric acid exactly neutralised 20.0 cm³ of a solution of sodium hydroxide of concentration 0.150 mol/dm³.



Calculate the concentration, in mol/dm³ of the aqueous sulphuric acid.

A fertiliser contains three compounds:
ammonium sulphate, (NH₄)₂SO₄,
iron(II) sulphate, FeSO₄,
sand, SiO₂.

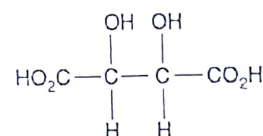
a) Calculate the percentage by mass of nitrogen in ammonium sulphate.

b) The mass of iron(II) ions in a sample of fertiliser can be determined by the reaction between iron(II) ions and acidified potassium manganate(VII), KMnO₄. A student analysed a sample of the fertiliser. He dissolved the sample in 25.0 cm³ of dilute sulphuric acid and titrated the solution formed with 0.0200 mol / dm³ potassium manganate(VII). The student used 22.5 cm³ of potassium manganate(VII) to reach the end-point.

i) Calculate the number of moles of potassium manganate(VII) used in the titration.

ii) One mole of potassium manganate(VII) reacts with five moles of iron(II) ions. Calculate the mass, in grams, of iron(II) ions in the sample analysed.

4 Red grapes contain a number of coloured pigments. Some red grapes are crushed and the pigments extracted with a solvent. The deep red solution contains a mixture of pigments. Tartaric acid can also be extracted from grape juice. The structure of tartaric acid is shown below.



i) Deduce the empirical formula of tartaric acid.

ii) A solution of tartaric acid was titrated with 0.100 mol / dm³ potassium hydroxide.
 $\text{C}_2\text{H}_2(\text{OH})_2(\text{CO}_2\text{H})_2 + 2\text{KOH} \rightarrow \text{C}_2\text{H}_2(\text{OH})_2(\text{CO}_2\text{K})_2 + 2\text{H}_2\text{O}$
tartaric acid

It required 6.00 cm³ of the potassium hydroxide solution to neutralise 20.0 cm³ of tartaric acid. Calculate the concentration, in mol / dm³, of the tartaric acid solution.

- (iii) Tartaric acid is purified by recrystallisation. On analysis, 8.00 g of impure tartaric acid was found to contain 7.40 g of pure tartaric acid. Calculate the percentage purity of the impure tartaric acid.

5 Iron(II) sulphate, FeSO_4 , is easily oxidised to iron(III) sulphate.

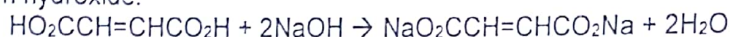
- (a) Calculate the percentage by mass of iron in iron(II) sulphate.

(b) An impure sample of iron(II) sulphate was analysed by titration. The sample was dissolved in 25.0 cm^3 of dilute sulphuric acid and then titrated against 0.0400 mol/dm^3 potassium dichromate(VI) solution. 19.0 cm^3 of potassium dichromate(VI) solution was required to reach the end-point.

- (i) Calculate the number of moles of potassium dichromate(VI) used in the titration.

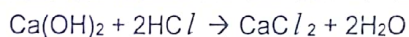
(ii) One mole of potassium dichromate(VI) reacts with six moles of iron(II) ions. Calculate the mass, in grams, of iron(II) ions in the sample analysed.

6 Fumaric acid is a colourless solid which can be extracted from plants. A solution of fumaric acid was titrated against aqueous sodium hydroxide.



18.0 cm^3 of 0.200 mol/dm^3 sodium hydroxide were required to neutralise 60.0 cm^3 of fumaric acid solution. Calculate the concentration, in mol/dm^3 , of the fumaric acid solution.

7 Chlorine, bromine and iodine are non-metals in Group VII of the Periodic Table. Their molecules are diatomic. An aqueous solution of calcium hydroxide was titrated with 0.0150 mol/dm^3 hydrochloric acid.

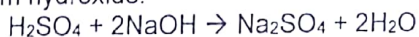


It required 6.00 cm^3 of this aqueous hydrochloric acid to neutralise 20.0 cm^3 of the calcium hydroxide solution. Calculate the concentration, in mol/dm^3 , of the calcium hydroxide solution.

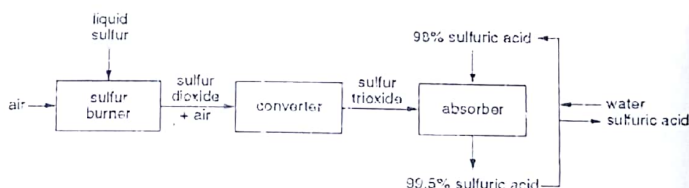
8 Sulfamic acid, SO_3NH_2 , is a weak acid used to remove limescale from kettles. A 0.105 g sample of sulfamic acid is dissolved in 25.0 cm^3 of water. The sulfamic acid solution requires 10.8 cm^3 of 0.100 mol/dm^3 potassium hydroxide for complete neutralisation. Calculate the number of moles of sulfamic acid that react with one mole of potassium hydroxide.

9 A flow diagram for the manufacture of sulfuric acid is shown below.

Aqueous sulfuric acid is titrated with aqueous sodium hydroxide.

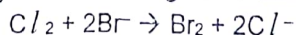


It requires 28.0 cm^3 of 0.100 mol/dm^3 aqueous sodium hydroxide to neutralise 9.50 cm^3 of sulfuric acid. Calculate the concentration, in mol/dm^3 , of the aqueous sulfuric acid. Give your answer to 3 significant figures.

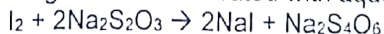


10. Seawater contains chloride, bromide and iodide ions.

Bromine can be manufactured by bubbling chlorine through seawater.



The concentration of sodium chlorate(I) in a solution can be found by reacting sodium chlorate(I) with excess acidified potassium iodide and then titrating the iodine liberated with aqueous sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.

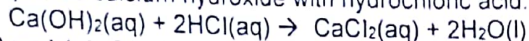


A solution of sodium thiosulfate contains 12.4 g of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, in 1.00 dm^3 of solution.

- (i) Calculate the concentration of the sodium thiosulfate solution in mol/dm^3 .

- 1) 23.6 cm³ of this sodium thiosulfate solution reacts with exactly 12.5 cm³ of aqueous iodine. Calculate the concentration, in mol / dm³, of the aqueous iodine.

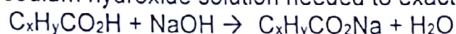
2. A student titrated 10.0 cm³ of aqueous calcium hydroxide with hydrochloric acid.



It required 4.00 cm³ of 0.0100 mol / dm³ hydrochloric acid to neutralise 10.0 cm³ of aqueous calcium hydroxide. Calculate the concentration of the calcium hydroxide.

3. Calculate the mass of Fe₃O₄ formed when 2.80 g of iron completely reacts with excess steam.

4. A solution containing 0.172 g of an unknown carboxylic acid, C_xH_yCO₂H, is titrated with 0.100 mol / dm³ aqueous sodium hydroxide. The volume of sodium hydroxide solution needed to exactly neutralise the acid is 23.2 cm³.



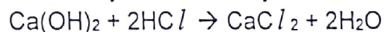
Calculate the relative formula mass, M_r, of the carboxylic acid and suggest its identity.

5. A sample of 25.0 cm³ of 0.500 mol / dm³ aqueous ammonium nitrite is heated. Calculate the volume of nitrogen formed, measured at room temperature and pressure.

6. Cement is made by heating clay with crushed calcium carbonate. During this process, the calcium carbonate is first converted to calcium oxide.

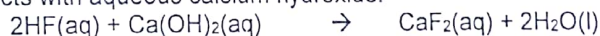


Concrete is made from cement, sand and water. When set, concrete is slightly porous. When rain water soaks through concrete, some of the uncombined calcium oxide dissolves to form calcium hydroxide. 25.0 cm³ of an aqueous solution of calcium hydroxide is exactly neutralised by 18.0 cm³ of 0.040 mol/dm³ hydrochloric acid.



Calculate the concentration, in mol/dm³, of the aqueous calcium hydroxide.

16. Dilute hydrofluoric acid reacts with aqueous calcium hydroxide.



What is the minimum volume, in cm³, of 0.150 mol / dm³ Ca(OH)₂ required to react completely with a solution containing 0.200 g of HF?

volume of Ca(OH)₂(aq) = cm³

17. Copper reacts with hot concentrated aqueous sulfuric acid.



An excess of copper is added to 25.0 cm³ of hot 14.0 mol / dm³ H₂SO₄.

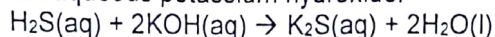
Use this information, together with the equation, to calculate the maximum volume of SO₂ formed.

The gas volume is measured at room temperature and pressure.

volume of SO₂ =

18. Hydrogen sulfide, H_2S , has a simple molecular structure. It is soluble in water.

Aqueous hydrogen sulfide reacts with aqueous potassium hydroxide.



What is the minimum volume, in cm^3 , of 0.150 mol / dm^3 KOH required to completely react with a solution containing 0.170 g of H_2S ?

volume of $\text{KOH}(\text{aq}) = \dots\dots\dots \text{cm}^3$

CALCULATING WATER OF CRYSTALLIZATION

1. One mole of a sample of hydrated sodium sulphide contains 162 g of water of crystallisation. What is the correct formula of this compound?

A $\text{Na}_2\text{S} \cdot 3\text{H}_2\text{O}$ B $\text{Na}_2\text{S} \cdot 5\text{H}_2\text{O}$ C $\text{Na}_2\text{S} \cdot 7\text{H}_2\text{O}$ D $\text{Na}_2\text{S} \cdot 9\text{H}_2\text{O}$

2. The formula for hydrated copper(II) nitrate is $\text{Cu}(\text{NO}_3)_2 \cdot x\text{H}_2\text{O}$. It contains 36.5% water of crystallisation by mass. What is the value of x ? [Ar: H, 1; N, 14; O, 16; Cu, 64]

A 4 B 5 C 6 D 7

MCQ'S

STRUCTURED QUESTIONS

Between the 13th and the 19th Century artists used a green pigment called verdigris. They made the pigment by hanging copper foil over boiling vinegar. Verdigris has the formula $[\text{Cu}(\text{CH}_3\text{CO}_2)_2]_2 \cdot \text{Cu}(\text{OH})_2 \cdot x\text{H}_2\text{O}$. It has a relative formula mass of 552 . Calculate the value of x in the formula.

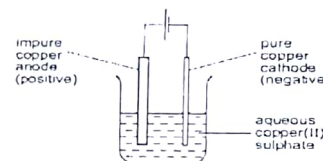
PERCENTAGE PURITY

1. A sample of copper contains a metal impurity which is below copper in the reactivity series. The diagram shows the apparatus used for refining the sample. The loss in mass of the anode (positive electrode) is 50 g and the gain in mass of the cathode (negative electrode) is 45 g . What is the percentage purity of this sample of copper?

A 10.0% B 11.1% C 90.0% D 95.0%

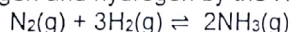
2. A piece of chalk has a mass of 23.0 g . Chalk is impure calcium carbonate. When analysed, the chalk is found to contain 0.226 moles of pure calcium carbonate. [M_r : CaCO_3 , 100]. What is the percentage purity of the piece of chalk?

A 0.983% B 1.02% C 77.0% D 98.3%



PERCENTAGE YIELD

Ammonia is manufactured from nitrogen and hydrogen by the Haber process.



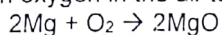
What is the percentage yield when 60 kg of ammonia is produced from 60 kg of hydrogen?

A 5.9% B 17.6% C 35.3% D 50.0%

MCQ'S

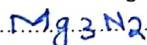
STRUCTURED QUESTIONS

1. Magnesium reacts with oxygen in the air to form magnesium oxide.

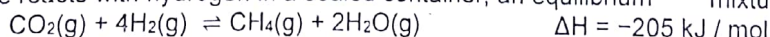


- (i) If the yield of the reaction is 75% calculate the mass of magnesium oxide formed when 12 kg of magnesium burns in excess air.

- (ii) Magnesium nitride is also formed when magnesium burns in air. Magnesium nitride is an ionic compound. Deduce the formula for magnesium nitride.



2. When carbon dioxide reacts with hydrogen in a sealed container, an equilibrium mixture is obtained.



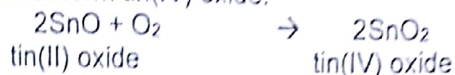
This reaction is exothermic. In an experiment, 220 g of carbon dioxide and an excess of hydrogen are reacted in a sealed container until an equilibrium is established. A mass of 46 g of methane is produced.

- (i) Calculate the mass of methane that should have been made if the percentage yield was 100% .

- 7 (a) A 9.50 g sample of a chloride of tin contains 5.95 g of tin.
Deduce the empirical formula of this chloride of tin.

empirical formula

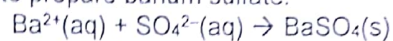
- (b) Tin(II) oxide reacts with oxygen to form tin(IV) oxide.



When a sample of 13.5 g of tin(II) oxide is reacted with oxygen, 12.7 g of tin(IV) oxide is formed.
Calculate the percentage yield of tin(IV) oxide.

- 8 Acids are neutralised by alkalis.

Aqueous sodium sulfate can be used to prepare barium sulfate.



In an experiment, 20.0 cm³ of 0.550 mol / dm³ of barium nitrate was added to excess aqueous sodium sulfate.

- (i) Calculate the maximum mass of barium sulfate that could be made. [The relative formula mass of BaSO₄ is 233]

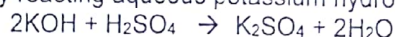
maximum mass of barium sulfate = g

- (ii) A mass of 1.92 g of dry barium sulfate was obtained. Calculate the percentage yield of barium sulfate.

percentage yield of barium sulfate = %

- 9 Acids are neutralised by insoluble bases.

Potassium sulfate can be prepared by reacting aqueous potassium hydroxide with dilute sulfuric acid.



In an experiment, 20.0 cm³ of 0.650 mol / dm³ sulfuric acid is just neutralised by aqueous potassium hydroxide.

- (i) Calculate the maximum mass of potassium sulfate, K₂SO₄, that could be prepared.
[The relative formula mass of K₂SO₄ is 174.]

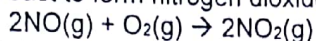
maximum mass of potassium sulfate = g [2]

- (ii) After crystallisation, 1.72 g of dry potassium sulfate was obtained. Calculate the percentage yield of potassium sulfate.

percentage yield of potassium sulfate = %

LIMITING REACTANT

Nitrogen monoxide and oxygen react to form nitrogen dioxide.



What is the maximum volume of nitrogen dioxide that could be obtained when 1 dm³ of nitrogen monoxide reacts with 2 dm³ of oxygen?

- (A) 1 dm³ B 2 dm³ C 3 dm³ D 4 dm³

Dilute ethanoic acid and dilute hydrochloric acid both react with magnesium ribbon to form hydrogen. Magnesium ribbon reacts with hydrochloric acid as shown in the equation.



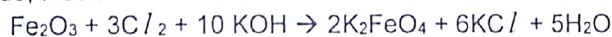
A 0.24 g sample of magnesium ribbon is added to 5.0 cm³ of 2.0 mol/dm³ hydrochloric acid.

- i) Which reactant, magnesium or hydrochloric acid, is in excess? Use calculations to explain your answer.

- ii) Calculate the maximum mass of magnesium chloride that can be formed in this reaction.

- iii) A 0.24 g sample of magnesium ribbon is added to 5.0 cm³ of 2.0 mol/dm³ ethanoic acid. Explain why this reaction forms the same volume of hydrogen but takes place much more slowly than the reaction of the same mass of magnesium with 5.0 cm³ of 2.0 mol/dm³ hydrochloric acid.

Analysis of a compound Z obtained from the planet Mars showed Z has the following composition. K₂FeO₄ can be prepared in the laboratory by the reaction between iron(III) oxide, Fe₂O₃, chlorine, Cl₂, and potassium hydroxide, KOH.



A 2.00 g sample of Fe₂O₃ is added to 20.0 cm³ of 4.00 mol dm⁻³ KOH.

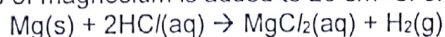
- i) Calculate the amount, in moles, of Fe₂O₃ used.

- ii) Calculate the amount, in moles, of KOH used.

- iii) Which reagent, Fe₂O₃ or KOH, is in excess in this reaction?

element	percentage by mass
potassium	39.4
iron	28.3
oxygen	32.3

3. A sample of 0.030 g of small pieces of magnesium is added to 20 cm³ of 0.10 mol / dm³ hydrochloric acid.



Show by calculation which reactant is in excess.

4. Chlorine reacts with sodium hydroxide.



A volume of 144 cm³ of chlorine gas, measured at room temperature and pressure, is passed into 38.0 cm³ of 0.250 mol / dm³ sodium hydroxide. Show by calculation which reactant is in excess.

PREPARED BY : KASHIF HASNAT MALIK

FINDING FORMULA OF A HYDROCARBON

MCQ'S

- A 10 cm³ sample of a gaseous hydrocarbon is completely burnt in oxygen. The total volume of the products is 70 cm³. Which equation represents the combustion of the hydrocarbon?
 A $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ B $\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
 C $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$ D $2\text{C}_2\text{H}_6(\text{g}) + 7\text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$
- Which hydrocarbon will burn completely in oxygen to give equal numbers of moles of carbon dioxide and water?
 A C_2H_6 B C_3H_8 C C_4H_{10} D C_5H_{12}
- Two different hydrocarbons each contain the same percentage by mass of hydrogen. It follows that they have the same
 A empirical formula. B number of isomers.
 C relative molecular mass. D structural formula.
- When butanol, represented by $\text{C}_4\text{H}_w\text{OH}$, burns in air, carbon dioxide and water are formed.

$$\text{C}_4\text{H}_w\text{OH} + x\text{O}_2 \rightarrow 4\text{CO}_2 + y\text{H}_2\text{O}$$
 Which values of w, x and y balance the equation?

	w	x	y
A	8	6	4
B	9	6	4
C	9	6	5
D	10	7	5
- A hydrocarbon, C_3H_y , burns in air to form carbon dioxide and water.

$$\text{C}_3\text{H}_y(\text{g}) + 5\text{O}_2(\text{g}) \rightarrow 3\text{CO}_2(\text{g}) + \frac{y}{2}\text{H}_2\text{O}(\text{g})$$
 What is the value of y?
 A 4 B 6 C 7 D 8
- In an experiment, 1 cm³ of a gaseous hydrocarbon X required 4 cm³ of oxygen for complete combustion to give 3 cm³ of carbon dioxide. All gas volumes are measured at r.t.p. Which formula represents X?
 A C_2H_2 B C_2H_4 C C_3H_4 D C_3H_8
- When 20 cm³ of a gaseous alkene burns in an excess of oxygen, 60 cm³ of carbon dioxide are formed. Both volumes are measured at r.t.p. What is the formula of the alkene?
 A C_3H_6 B C_3H_8 C C_6H_{12} D C_6H_{14}
- In an experiment, 1 cm³ of a gaseous hydrocarbon, Z, requires 4 cm³ of oxygen for complete combustion to give 3 cm³ of carbon dioxide. All gas volumes are measured at r.t.p. Which formula represents Z?
 A C_2H_2 B C_2H_4 C C_3H_4 D C_3H_8

FINDING FORMULA OF A COMPOUND

MCQ'S

- The mass of one mole of a chloride formed by a metal Y is 74.5 g. What is the formula of the chloride?
 A Y_3Cl B Y_2Cl C YCl D YCl_2
- The formula of china clay (aluminium silicate) was shown in an old book as $\text{Al}_2\text{O}_3 \cdot 2\text{SiO}_2 \cdot 2\text{H}_2\text{O}$. This formula is shown in a modern book as $\text{Al}_2(\text{OH})_x\text{Si}_2\text{O}_y$. What are the values of x and y in the modern formula?

	x	y
A	2	4
B	2	5
C	4	3
D	4	5
- In an experiment 264 g of strontium reacts with 213 g of chlorine. What is the formula of strontium chloride?
 A SrCl B SrCl_2 C SrCl_3 D Sr_2Cl
- The element X forms a gaseous molecule X_2 . One volume of X_2 combines with one volume of hydrogen to form two volumes of a gaseous hydride. What is the formula for the hydride of X?
 A HX B HX_2 C H_2X D H_2X_2
- What is the ratio of the number of molecules in 71 g of gaseous chlorine to the number of molecules in 2 g of gaseous hydrogen? [Relative atomic masses A_r : H, 1; Cl, 35.5]
 A 1:1 B 1:2 C 2:1 D 71:2

Nickel can be refined by reacting the impure metal with carbon monoxide. The impurities do not react with carbon monoxide. A volatile compound called nickel carbonyl is formed. This is decomposed to give pure nickel and carbon monoxide. Nickel carbonyl has the formula $\text{Ni}(\text{CO})_x$. The relative molecular mass of nickel carbonyl is 171. Calculate the value of x.

STRUCTURED QUESTIONS

FINDING FORMULA OF A HYDROCARBON

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MCQ'S

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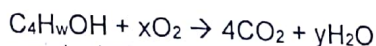
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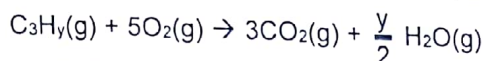
- A empirical formula.
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When butanol, represented by $\text{C}_4\text{H}_9\text{OH}$, burns in air, carbon dioxide and water are formed.



Which values of w, x and y balance the equation?

A hydrocarbon, C_3H_y , burns in air to form carbon dioxide and water.



What is the value of y?

- A 4 B 6 C 7 D 8

In an experiment, 1 cm³ of a gaseous hydrocarbon X required 4 cm³ of oxygen for complete combustion to give 3 cm³ of carbon dioxide. All gas volumes are measured at r.t.p. Which formula represents X?

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MCQ'S

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STRUCTURED QUESTIONS

MARK SCHEME

MOLES

MCQ'S

MCQ	Key	Solution & Explanation
1	C	1 mole of any compound contains same number of molecules
2	B	Same moles = same molecules. moles (H_2O) = $\frac{9}{18} \text{ g} = 0.5 \text{ mol}$ A $\frac{2}{2} = 1 \text{ mol}$ B $\frac{14}{28} = 0.5 \text{ mol}$ C $\frac{32}{32} = 1 \text{ mol}$ D $\frac{44}{44} = 1 \text{ mol}$
3	D	1 mol of any element = A_r of element
4	D	$64 + 32 + (16 \times 4) + 5[(1 \times 2) + 16] = 250$
5	A	No. of molecules ratio Cl_2 : H_2 $\frac{71}{71}$: $\frac{2}{2}$ 1 : 1
6	A	Mass of Cu = $0.25 \times 64 = 16 \text{ g}$ A : O_2 $0.5 \times 32 = 16 \text{ g}$ B : SO_2 $1 \times 64 = 64 \text{ g}$ C : H_2O $1.5 \times 18 = 27 \text{ g}$ D : O $2 \times 16 = 32 \text{ g}$
7	D	Mass : $2 \times 71 = 142 \text{ g}$
8	C	$0.1 \text{ mol SiO}_2 = 28 + (16 \times 2) = 60 \times 0.1 = 6 \text{ g}$ $0.5 \text{ mol O}_2 = (16 \times 2) = 32 \times 0.5 = 16 \text{ g}$ $0.5 \text{ mol Li} = 7 \times 0.5 = 3.5 \text{ g}$ $1.0 \text{ mol NH}_3 = 14 + (1 \times 3) = 17 \times 10 = 17 \text{ g}$

CALCULATING MASS OF ELEMENT/ION OF COMPOUND

MCQ'S

MCQ	Key	Solution & Explanation
1	A	Ammonium salt A 1 mol ∴ 0.5 B 1 mol ∴ 0.5 C 1 mol D 1 mol Ammonia released 3 mol $0.5 \times 3 = 1.5 \text{ mol}$ 2 mol $0.5 \times 2 = 1 \text{ mol}$ 1 mol 1 mol
2	D	A_r of Al = 27, M_r of $\text{Al}_2\text{O}_3 = 102$ Mass ratio Al_2O_3 : Al $102 \text{ g} : 27 \times 2 = 54 \text{ g}$ $204 \text{ g} : \frac{54}{102} \times 204 = 108 \text{ g}$
5	C	Mass ratio H_2O : O $18 \text{ g} : 16 \text{ g}$ $72 \text{ g} : \frac{16}{18} \times 72 = 64 \text{ g}$
4	D	M_r of $\text{CH}_4 = 16 = 1 \text{ mol}$ CH_4 : H $16 \text{ g} : 4 \text{ mol}$ $3.2 \text{ g} : \frac{4}{16} \times 3.2 = 0.8 \text{ mol}$

STRUCTURED QUESTIONS

- Molybdenum (VI) oxide MoO_3
 $96 + (16 \times 3) = 144 \text{ g}$
 125 g

$\text{Mo}^{6+}\text{O}^{2-}$ or MoO_3
 Mo
 96 g

$\frac{96}{144} \times 125 = 83.33 \text{ g}$
- Titration (IV) chloride TiCl_4
 Mass ratio TiCl_4 : Ti
 190 g : 48 g
 125 g : $\frac{48}{190} \times 125 = 31.58 \text{ g}$

M_r of $\text{TiCl}_4 = 48 + (35.5 \times 4) = 190$

MOLAR VOLUME

MCQ'S

MCQ	Key	Solution & Explanation
1	A	H_2 : CH_4 $\frac{2 \text{ g}}{2} = 1 \text{ mol}$: $\frac{16 \text{ g}}{16} = 1$
2	C	M_r of $\text{C}_2\text{H}_6 = 30$ moles (C_2H_6) = $\frac{20}{30} = 0.67 \text{ mol}$ M_r of $\text{C}_3\text{H}_8 = 42$ $\text{Mol} = \frac{\text{mass}}{M_r}$ or mass = $0.67 \times 42 = 28 \text{ g}$

PERCENTAGE COMPOSITION

MCQ'S

MCQ	Key	Solution & Explanation
1	D	
2	C	A_r of N = 14 A $\frac{14 \times 2}{80} \times 100 = 35\%$ B $\frac{14 \times 2}{132} \times 100 = 21.2\%$ C $\frac{14 \times 2}{60} \times 100 = 47.7\%$ D $\frac{14 \times 2}{149} \times 100 = 28.2\%$
3	D	$A.$ $\frac{14 \times 2}{132} \times 100 = 21.21\%$ $B.$ $\frac{14 \times 2}{132} \times 100 = 21.21\%$ $C.$ $\frac{14 \times 2}{80} \times 100 = 35\%$ $D.$ $\frac{14 \times 2}{60} \times 100 = 46.7\%$

EXAMINATION STRUCTURED QUESTIONS

- (i) CuO (ii) +1
 (iii) M_r of $\text{Cu}_2\text{O} = 144$
 $\% \text{ of Cu} = \frac{64 \times 2}{144} \times 100 = 88.9\%$
- M_r of $\text{NH}_4\text{NO}_3 = 80$ M_r of $(\text{NH}_4)_2\text{HPO}_4 = 132$
 $\% \text{ of N in } \text{NH}_4\text{NO}_3 = \frac{14 \times 2}{132} \times 100 = 21.2\%$
- M_r of $(\text{NH}_4)_2\text{SO}_4 = 132$ M_r of $\text{KNO}_3 = 101$
 $\text{Mass of N in } (\text{NH}_4)_2\text{SO}_4 = \frac{14 \times 2}{132} \times 500 = 106.06 \text{ g}$
 $\text{Mass of N in } \text{KNO}_3 = \frac{14}{101} \times 500 = 69.31 \text{ g}$
 $\text{Total mass of N} = 106.06 + 69.31 = 175.37$
 $\% \text{ of N} = \frac{175.37}{500 + 500} = 17.54\%$

4. (i) $(\text{NH}_4)_3\text{PO}_4$
 (ii) M_r of $(\text{NH}_4)_3\text{PO}_4 = 149$; $\% \text{ of N} = \frac{143 \times 3}{149} \times 100 = 28.19\%$
 $= 0.8000 \text{ g}$
 $= 19.2024 \text{ g}$
5. Mass (H_2) = 0.4000×2
 Mass (I_2) = 0.07560×254
 Mass (HI) = 1.344×128
 Total mass of $\text{H}_2 + \text{I}_2 + \text{HI}$ = 192.0344 g
 $\% \text{ of I}_2 = \frac{19.2024}{192.0344} \times 100 = 9.99\% \approx 10\%$
6. Vol % = $\frac{\text{Vol of Methane}}{\text{total volume}} \times 100$
 $0.0014 = \frac{\text{Vol of Methane}}{1000} \times 100$
 Vol. of Methane = $\frac{0.0014}{100} \times 1000 = 0.0014 \text{ dm}^3$
 Volume = moles $\times 24$
 $0.00014 = \text{moles} \times 24$
 Mass = $5.83 \times 10^{-5} \times 16 = 9.33 \times 10^{-4} \text{ g}$
7. molar mass of $(\text{NH}_4)_3\text{PO}_4 = 149$ $\frac{42}{149} \times 100 = 28.19\% / 28.2\%$
8. M_r of $\text{CaSO}_4 = 272$ M_r of $\text{Ca}(\text{H}_2\text{PO}_4)_2 = 234$
 Total mass of $\text{CaSO}_4 + \text{Ca}(\text{H}_2\text{PO}_4)_2$
 $2(272) + 234 = 506$
 $\% \text{ of CaSO}_4 = \frac{272}{506} \times 100 = 53.76\%$
9. M_r of $\text{Fe}_3\text{O}_4 = (56 \times 3) + (16 \times 4) = 232$ $\% \text{ age of Fe} = \frac{56 \times 3}{232} \times 100 = 72.4\%$
10. $\% \text{ of N in } \text{NH}_4\text{NO}_3 = \frac{14 \times 2}{80} \times 100 = 35\%$ $\% \text{ of N in } (\text{NH}_2)_2\text{CO} = \frac{14 \times 2}{60} \times 100 = 46.7\%$
11. Percentage of N in ammonium nitrate = 35% (1)
 Percentage of N in urea = 47% (1) OR Both formulae contain two nitrogen atoms (1)
 Urea has a lower relative formula mass (1)
12. Relative formula mass = 174 (1) Percentage of K = 44.8%

EMPIRICAL FORMULA

MCQ'S

MCQ 3

MCQ	Key	Solution & Explanation												
1	B													
MCQ	Key	Solution & Explanation												
2	A	<table> <tr> <th></th><th>N</th><th>O</th></tr> <tr> <td>Mole</td><td>$\frac{47}{14} = 3.357$</td><td>$\frac{53}{16} = 3.125$</td></tr> <tr> <td>Atoms</td><td>$\frac{3.357}{3.125} = 1.1$</td><td>$\frac{3.125}{3.125} = 1$</td></tr> <tr> <td></td><td>1</td><td>1</td></tr> </table> Empirical formula NO		N	O	Mole	$\frac{47}{14} = 3.357$	$\frac{53}{16} = 3.125$	Atoms	$\frac{3.357}{3.125} = 1.1$	$\frac{3.125}{3.125} = 1$		1	1
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	C	H	O											
Mole	$\frac{12}{12} = 1$	$\frac{2}{1} = 2$	$\frac{16}{16} = 1$											
	1	2	1											
4	A													

PREPARED BY KASHIF HASNAT MALIK

5	B	C 80% $\frac{80}{12}$ $\frac{6.67}{6.67} = 1$	H 20% $\frac{20}{1}$ $\frac{20}{6.67} = 3$	CH ₃
6	B	Fe $\frac{70}{56} = 1.25$ Multiply by 2 $\frac{1.25}{1.25} = 1$ 2	O $\frac{30}{16} = 1.875$ $\frac{1.875}{1.25} = 1.5$ 3	Fe ₂ O ₃

STRUCTURED QUESTIONS

1.

	C	H	O
mol ratio	$\frac{0.72}{12} = 0.06$	$\frac{0.18}{1} = 0.18$	$\frac{0.96}{16} = 0.06$
Atomic ratio	$\frac{0.06}{0.06} = 1$	$\frac{0.18}{0.06} = 3$	$\frac{0.06}{0.06} = 1$
Empirical formula = CH ₃ O			

2.

	K	Fe	C	N
mol ratio	$\frac{0.547}{39} = 0.0140$	$\frac{0.195}{56} = 0.00348$	$\frac{0.252}{12} = 0.0210$	$\frac{0.294}{14} = 0.0210$
atomic ratio	$\frac{0.0140}{0.00348} = 4$	$\frac{0.00348}{0.00348} = 1$	$\frac{0.0210}{0.00348} = 6$	$\frac{0.0210}{0.00348} = 6$

Empirical formula = K₄FeC₆N₆ or K₄Fe(CN)₆

3.

	Xe	O	F
mol ratio	$\frac{9.825}{131} = 0.075$	$\frac{1.200}{16} = 0.075$	$\frac{5.700}{19} = 0.300$
atomic ratio	$\frac{0.075}{0.075} = 1$	$\frac{0.075}{0.075} = 1$	$\frac{0.300}{0.075} = 4$

Empirical formula = XeOF₄

4.

	C	H	O
mol ratio	$\frac{10.50}{12} = 0.875$	$\frac{0.75}{1} = 0.75$	$\frac{10.00}{16} = 0.625$
atomic ratio	$\frac{0.875}{0.075} = 11.67$ 5 (1.4) 7	$\frac{0.75}{0.075} = 10$ 5 (1.2) 6	$\frac{0.625}{0.075} = 8.33$ 5 (1) 5s

Multiply by 5
or divided by 0.2Empirical formula C₇O₅H₆5. CrC₆O₅ ALLOW: Cr(CO)₆

6.	Cu $\frac{79.9}{64}$	O $\frac{20.1}{16}$	
Mol ratio	$\frac{79.9}{64} = 1.25$	$\frac{20.1}{16} = 1.25$	
Atomic ratio	$\frac{1.25}{1.25} = 1$	$\frac{1.25}{1.25} = 1$	Empirical formula : CuO

7.

	S	O
mol ratio	$\frac{40.0}{32} = 1.25$	$\frac{60.0}{16} = 3.75$
atomic ratio	$\frac{1.25}{1.25} = 1$	$\frac{3.75}{1.25} = 3$

Empirical formula SO_3

8.

	N	H	Cr	O
Mole ratio	$\frac{11.1}{14} / 0.793$	$\frac{3.2}{1} / 3.2$	$\frac{41.3}{52} / 0.794$	$\frac{44.4}{16} / 2.78$
simplified ratio	$\frac{0.793}{0.793} / 1$	$\frac{3.2}{0.793} / 4$	$\frac{0.794}{0.793} / 1$	$\frac{2.78}{0.793} / 3.5$
$\times 2$	2	8	2	7

Mole ratio line (1) Simplified ratio line (1)
Idea of the $\times 2$ (1)

9.

	C	Cl	H	Si
Mole ratio	$\frac{18.6}{12} = 1.5$	$\frac{55.0}{35.5} = 1.55$	$\frac{4.65}{1} = 4.65$	$\frac{21.7}{28} = 0.775$
Atomic	$\frac{1.55}{0.775} = 2$	$\frac{1.55}{0.775} = 2$	$\frac{4.65}{0.775} = 6$	$\frac{0.775}{0.775} = 1$

Empirical formula $\text{C}_2\text{Cl}_2\text{H}_6\text{Si}$

10.

	Ca	H	P	O
Moles	$\frac{17.1}{40} = 0.4275$	$\frac{1.7}{1} = 1.7$	$\frac{26.5}{31} = 0.8548$	$\frac{54.7}{16} = 3.418$
Simple ratio	$\frac{0.4275}{0.4275} = 1$	$\frac{1.7}{0.4275} = 3.98 \approx 4$	$\frac{0.8548}{0.4275} = 1.99 \approx 2$	$\frac{3.4188}{0.4275} = 7.997 \approx 8$

11.

	C	H	O
Moles :	$\frac{76.60}{12} = 6.38$	$\frac{6.38}{1} = 6.38$	$\frac{17.02}{16} = 1.06$
Atom :	$\frac{6.38}{1.06} = 6.02$	$\frac{6.38}{1.06} = 6.02$	$\frac{1.06}{1.06} = 1$
	≈ 6	≈ 6	1
Empirical formula :	$\text{C}_6\text{H}_6\text{O}$		

12

	C	H	O
Moles :	$\frac{37.5}{12} = 3.125$	$\frac{12.5}{1} = 12.5$	$\frac{50.0}{16} = 3.125$
Atom :	$\frac{3.125}{3.125} = 1$	$\frac{12.5}{3.125} = 4$	$\frac{3.125}{3.125} = 1$
Empirical formula :	CH_4O		

13 Cu_2O (2)If two marks not scored, mol Cu = $\frac{9.86}{64}$ AND mol O = $\frac{1.23}{16}$ scores 1 mark

MOLECULAR FORMULA DETERMINATION

STRUCTURED QUESTIONS

1.

	C	H	O
mol ratio	$\frac{0.18\text{g}}{12} = 0.015$	$\frac{0.03\text{g}}{1} = 0.03$	$\frac{0.08\text{g}}{16} = 0.005$
atomic ratio	$\frac{0.015}{0.005} = 3$	$\frac{0.03}{0.005} = 6$	$\frac{0.005}{0.005} = 1$

Empirical formula = $\text{C}_3\text{H}_6\text{O}$

$n = \frac{116}{(12 \times 3) + (1 \times 6) + 16} = \frac{116}{58} = 2$ Molecular Formula = $2(\text{C}_3\text{H}_6\text{O}) = \text{C}_6\text{H}_{12}\text{O}_2$

2. (a) (i)

	C	H	O
mol ratio	$\frac{2.67}{12} = 0.2225$	$\frac{0.220}{1} = 0.220$	$\frac{7.11}{16} = 0.4443755$
atomic ratio	$\frac{0.2225}{0.220} = 1.01 \approx 1$	$\frac{0.220}{0.220} = 1$	$\frac{0.444375}{0.220} = 2.02 \approx 2$

Empirical formula = CHO_2

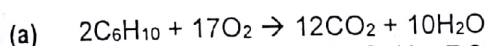
(ii) Empirical formula mass = $12 + 1 + (16 \times 2) = 45$

$n = \frac{90}{45} = 2$

Molecular formula = $2(\text{CHO}_2) = \text{C}_2\text{H}_2\text{O}_4$

3. (% H is) 16 (%) (1)

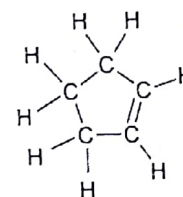
Moles C: $\frac{84}{12}$ Moles H: $\frac{16}{1}$ or mole ratio 7:16 (1)

Molecular formula is C_7H_{16} 

(b) (i) Mole ratio C: H = 7.35 : 11.8 (1)

Idea of dividing by smallest / simplest ratio is 1 : 1.6 AND $\times 5$ (1)

(ii)



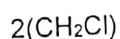
5 (i)

C	H	Cl
$\frac{0.242}{12}$	$\frac{0.040}{1}$	$\frac{0.178}{35.5}$
= 0.020	0.040	0.020
1	2	1



(ii) Empirical formula mass of $\text{CH}_2\text{Cl} = 49.5$

$n = \frac{99}{49.5} = 2$

molecular formula $\text{C}_2\text{H}_4\text{Cl}_2$

STOICHIOMETRY

MCQ'S

MCQ	Key	Solution & Explanation	
1	C	$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ 2mol 1 mol 2 mol 2(2)=4g 32g 2(18) = 36g	
2	D		
3	D		
4	C	mol ratio $\text{Al}_2(\text{SO}_4)_3$: H_2SO_4 1 mol : 3 mol 0.5 mol : $3 \times 0.5 = 1.5$ mol	
5	B	mol ratio CO_2 : H_3PO_4 3mol : 2 mol 1.5 mol : $\frac{2}{3} \times 1.5 = 1$ mol	

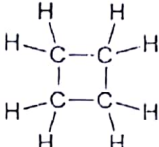
PREPARED BY : KASHIF HASNAT MALIK

6	B	$\text{Mg} + 2\text{CH}_3\text{COOH} \rightarrow (\text{CH}_3\text{COO})_2\text{Mg} + \text{H}_2$ $\text{1mol} \qquad \qquad \qquad \qquad \qquad \qquad \qquad \qquad \qquad \text{1mol}$							
7	B	NH_3	+	2O_2	$\longrightarrow$	HNO_3	+	H_2O	
		HNO_3	+	NH_3	$\longrightarrow$	NH_4NO_3			
		2NH_3	+	2O_2	$\longrightarrow$	NH_4NO_3	+	H_2O	
		Mass ratio	:	NH_3	:	NH_4NO_3			
				$2(17)34\text{g}$	:	80g			
				17 tonne	:	$\frac{80}{34} \times 17 = 40 \text{ tonnes}$			

STRUCTURED QUESTIONS

- mol ratio $\text{Ca}(\text{NO}_3)_2$: $\text{NO}_2 + \text{O}_2$
 2 mol : $(4+1) 5 \text{ mol}$
 0.010 mol : $0.010 \times \frac{5}{2} = 0.025 \text{ mol}$
- (i) equation $2 \text{NiS} + 3 \text{O}_2 \rightarrow 2 \text{NiO} + 2 \text{SO}_2$
(ii) $\text{Mr of SO}_2 = 64$, $\text{Mr of NiS} = 91$
Mass ratio NiS : SO_2
 $2(91)182\text{g}$: $2(64) = 128\text{g}$
- $\text{Mr of NH}_3 = 17$, $\text{Mr of (NH}_4)_2\text{SO}_4 = 132$
Moles ratio NH_3 : $(\text{NH}_4)_2\text{SO}_4$
 $2(17) = 34 \text{ g}$: 132 g
 51 g : $\frac{132}{34} \times 51 = 198 \text{ g}$
- $\text{Mr of Ag} = 108$, $\text{Mr of AgCl} = 143.5$
Mass ratio AgCl : Ag
 143.5 g : 108 g
 0.287 g : $\frac{108}{143.5} \times 0.287 = 0.216 \text{ g}$
- $\text{Mr of Ca}_3\text{SiO}_5 = 228$, $\text{Mr of Ca(OH)}_2 = 74$
mass ratio Ca_3SiO_5 : Ca(OH)_2
 $2(228) = 456\text{g}$: $3(74) = 222 \text{ g}$
 912 g : $\frac{222}{456} \times 912 = 444 \text{ g}$
- (i) answer 4.40 g
(ii) $\text{Mr of K}_2\text{CO}_3 = 138$, $\text{Mr of K}_2\text{SO}_4 = 174$
mass ratio K_2CO_3 : K_2SO_4
 138 g : 174 g
 3.45 g : $\frac{174}{138} \times 3.45 = 4.35 \text{ g}$
- Mass ratio $\text{C}_6\text{H}_{12}\text{O}_6$: $\text{C}_2\text{H}_5\text{OH}$
 180 g : $2(46) = 92 \text{ g}$
 36 tonnes : $\frac{92}{180} \times 36 = 18.4 \text{ tonnes}$
- (i) $108 \text{ g} \times 100 \text{ km} = 10800 \text{ g}$ or 10.8 kg
(ii) mass ratio CO_2 : C_6H_6
 $16(44) = 704 \text{ g}$: $2(144) = 288 \text{ g}$
 10800 g : $\frac{288}{704} \times 10800 = 3497.73 \text{ g} = 3.498 \text{ kg}$
- (a) $\text{Mr of CH}_3\text{NH}_2 = 31$, Moles of $\text{CH}_3\text{NH}_2 = \frac{6.2}{31} = 0.2 \text{ mol}$
Vol. of $\text{CH}_3\text{NH}_2 = 0.2 \times 24 = 4.8 \text{ dm}^3$
(b) $\text{Mr of CH}_3\text{OH} = 32$
Mass ratio CH_3OH : CH_3NH_2
 32g : 31g
 240 kg : $\frac{31}{32} \times 240 = 232.5 \text{ kg}$

- 10 (a) mass ratio N_2 : NO
 28 g : 2(30) = 60 g
 55 kg : $\frac{60}{28} \times 55 = 117.86 \text{ kg}$
- (b) 1 molecule of NO = 7+8 = 15 electrons
 1 mol of NO = $15 \times 6.02 \times 10^{23} = 9.03 \times 10^{24}$ electrons
11. (i) moles $\text{N}_2\text{H}_4 = \frac{1\,000\,000}{32} = 31\,250$ mole
 moles $\text{O}_2 = \text{moles } \text{N}_2\text{H}_4$ Volume of O_2 ($31\,250 \times 24$) = $750\,000 \text{ dm}^3 / 7.5 \times 10^5 \text{ dm}^3$;
 Alternative for 1st two stages:
 $32 \text{ g } \text{N}_2\text{H}_4 \rightarrow 32 \text{ g } \text{O}_2$ (1 mark) moles $\text{O}_2 = 1\,000\,000 / 32 = 31\,250$ (allow ecf) (1 mark)
- (ii) it / liquid oxygen takes up less space / room ; ALLOW: able to store more in liquid form / gaseous volume too high / maximum storage capacity. IGNORE: less easily spread out/no gas can escape / less possibility of an explosion / to prevent reaction with other substances
- 12 Mr of $\text{H}_2 = 2 \text{ g}$: Mr of HI = 128 g
 Mass ratio : H_2 : HI
 2g : 2(128) = 256g
 45.3 g : $\frac{256}{2} \times 45.3 = 5798.4 \text{ g}$
13. Mr of $\text{ZnCl}_2 = 136$, Mr of $\text{Zn}(\text{NH}_3)_4\text{Cl}_2 = 204$
 Mass ratio ZnCl_2 : $\text{Zn}(\text{NH}_3)_4\text{Cl}_2$
 136 g : 204 g
 3.4 g : $\frac{204}{136} \times 3.4 = 5.1 \text{ g}$
- 14 $\text{UO}_2 + 4\text{HF} \rightarrow \text{UF}_4 + 2\text{H}_2\text{O}$
 $\text{UF}_4 + 2\text{Mg} \rightarrow \text{U} + 2\text{MgF}_2$
 $\text{UO}_2 + 4\text{HF} + 2\text{Mg} \rightarrow \text{U} + 2\text{MF}_2 + 2\text{H}_2\text{O}$ (combined equation)
 Mr of $\text{UO}_2 = 238 + (16 \times 2) = 270$
 Mass ratio UO_2 : U
 270 g : 238
 1 tonne : $\frac{238}{270} \times 1 = 0.881 \text{ tonne}$
- 15 Mass ratio N_2 : NO
 28g : 2(30)60g
 100g : $\frac{60}{28} \times 100 = 214.29 \text{ g}$
- 16 Mass ratio $\text{C}_5\text{H}_{12}\text{O}_5$: $\text{C}_2\text{H}_5\text{OH}$
 180 g : 2(46) = 92g
 1 Tonne : $\frac{92}{180} \times 1 = 0.511 \text{ tonne}$
- 17 Ar of Mo = 96
 Mass ratio Mo : Al
 96g : 2(27) = 54 g
 1 tonne : $\frac{54}{96} \times 1 = 0.5625 \text{ tonne}$
- 18 Mass ratio C_2H_4 : $\text{C}_2\text{H}_5\text{OH}$
 28g : 46g
 0.4 tonne : $\frac{46}{28} \times 0.4 = 0.657 \text{ tonne} = 0.657 \times 5/100 = 0.0329 \text{ tonne}$
- 19 Moles of hydrogen = $\frac{400}{24000} = 0.01667$ (1) (mark for working or correct answer)
 Moles of alkali metal = $0.01667 \times 2 = 0.03334$ (1) (mark for working or correct answer)
 Ar = 7 (1) Lithium / Li (1) ALLOW: appropriate answer based on incorrect Ar in step 3

- 20 M_r of SO_4^{2-} = 96 M_r of BaSO_4 = 233
 mass ratio SO_4^{2-} : BaSO_4
 96 g : 233 g
 1.24 : $\frac{233}{96} \times 1.24 = 3.0096$ g
 Mass in 1000 dm^3 : mass in 50 cm^3
 3.0096 : $3.0096 \times \frac{50}{1000} = 0.151$ g
21. M_r of C_5H_8 = $(5 \times 12) + (1 \times 8) = 68$ M_r of $\text{C}_5\text{H}_8\text{SO}_2$ = $68 + 32 + (16 \times 2) = 132$
 Mass ratio C_5H_8 : $\text{C}_5\text{H}_8\text{SO}_2$
 68 g : 132 g
 100 g : $\frac{132}{68} \times 100 = 194.12$ g
22. M_r of C_6H_{14} = $(6 \times 12) + (1 \times 14) = 86$ M_r of C_6H_{12} = $(6 \times 12) + (1 \times 12) = 84$
 mass ratio C_6H_{14} : C_6H_{12}
 86 g : 84 g
 258 g : $\frac{84}{86} \times 258 = 252$ g
- 23 (a) Moles of C_6H_{14} = 3.0 (1) Mass of C_6H_{12} = 252 (g)
- (b) (i) Mole ratio C : H = 7.14 : 14.3 or 85.7 / 12 and 14.3 / 1(1) Divide by 7.14 to get empirical formula
 (ii)
- 
24. (a) ANY TWO FROM
 Dissolves
 Blue / green solution
 Fizzes / bubbles / effervescence
- (b) $\text{CuCO}_3 \cdot \text{Cu}(\text{OH})_2 + 4\text{HCl} \rightarrow 2\text{CuCl}_2 + \text{CO}_2 + 3\text{H}_2\text{O}$
 Correct formulae
 Balancing
- (c) Mass volume ratio CO_3^{2-} : CO_2
 60g : 24000 cm^3
 $\frac{60}{24000} \times 96 = 0.24$ g

DETERMINATION OF VOLUME OF A GAS REQUIRED / PRODUCED

MCQ'S

MCQ	Key	Solution & Explanation	
1	A	Mol ratio = volume ratio Mol/vol ratio H_2 : O_2 2 : 1	
2	D	mol – vol ratio CO_2 : CaCO_3 24000 cm^3 : 1 mol 24 cm^3 : $\frac{1}{24000} \times 24 = 0.001$ mol	
3	B	vol. ratio NH_3 : N_2 2(24) dm^3 : 24 $\text{dm}^3 = 48$ dm^3 1 : $\frac{24}{48} \times 1 = 0.5$ dm^3	
4	D	mol – vol ratio CO_2 : CaCO_3 24000 cm^3 : 1 mol 24 cm^3 : $\frac{1}{24000} \times 24 = 0.001$ mol	

STRUCTURED QUESTIONS

1. Mr of $\text{MgCO}_3 = 84$ and 1 mol of $\text{CO}_2 = 24 \text{ dm}^3$
 Mass – vol ratio MgCO_3 : CO_2
 84 g : 24 dm^3
 10.5 g : $\frac{24}{84} \times 10.5 = 3 \text{ dm}^3$
2. (i) reaction is faster ALLOW: larger surface area for reaction NOT: reaction is fast (comparison needed)
 (ii) Vol – mass ratio H_2 : Zn
 24 dm^3 : 65 g
 0.075 dm^3 : $\frac{65}{24} \times 0.075 = 0.195 \text{ g Zn}$
 (iii) %Zn = $\frac{0.195}{1.2} \times 100 = 16.25\%$
3. (i) Ar of P = 31 Mr of $\text{PH}_3 = 34$
 Mass ratio P : PH_3
 4(31) = 124 g : 34 g
 1.86 g : $\frac{34}{124} \times 1.86 = 0.51 \text{ g}$
 (ii) Moles of $\text{PH}_3 = \frac{0.51}{34} = 0.015 \text{ mol}$ Vol of $\text{PH}_3 = 0.015 \times 24 = 0.36 \text{ dm}^3$
4. Mol – volume ratio NaCl : HCl
 1 mol : 24 dm^3
 0.2 : $\frac{24}{1} \times 0.2 = 4.8 \text{ dm}^3$
5. Vol – mass ratio O_2 : H_2O
 24 dm^3 : 2(18) = 36 g
 2500 dm^3 : $\frac{36}{24} \times 2500 = 3750 \text{ g}$
6. Vol ratio H_2 : O_2
 2(24) = 48 dm^3 : 24 dm^3
 2000 dm^3 : $\frac{24}{48} \times 2000 = 1000 \text{ dm}^3$
7. Mass – vol ratio H_2 : Mg
 36 cm^3 : 0.036 g
 24000 cm^3 : $\frac{0.036}{36} \times 24000 = 24 \text{ g (Ar of Mg)}$
- 8 (a) 39% (2)
 If 2 marks not scored, molar mass of $\text{KCIO}_3 = 122.5$ scores 1 mark
 (b) $\text{KCIO}_3 = \frac{12.25}{122.5}$ or 0.10 (mol) (1)
 moles $\text{O}_2 = 1.5 \times 0.01 = 0.15$ moles / idea of multiplying moles by 1.5 (1)
 volume of $\text{O}_2 = 3.6 \text{ dm}^3 / 3600 \text{ cm}^3$, correct unit must be included (1)
- 9 (a) 34.5% (2)
 If two marks not scored, 171 (for molar mass of nickel carbonyl) scores 1 mark
 (b) mol nickel carbonyl = $\frac{1.71}{171}$ OR 0.01 (mol) (1)
 mol gases = (0.01) $\times 5$ / idea of multiplying mol $\times 5$ (1)
 volume of gases = 1.2 $\text{dm}^3 / 1200 \text{ cm}^3$ (units must be correct)
- 10 Moles of Zn and $\text{H}_2 = 0.01$ (1) Volume of $\text{H}_2 = 0.24 \text{ dm}^3$ (1)

CONCENTRATION OF SOLUTIONS

MCQ'S

Solution & Explanation		
MCQ	Key	
1	D	$5 \text{ mol dm}^{-3} \times 2 \text{ dm}^3 = 10 \text{ mol}$ $A_r \text{ of I} = 127$, $M_r \text{ of I}_2 = 254$
2	C	$n = \frac{2.54}{254} = 0.01 \text{ mol}$ $C = \frac{0.01}{250} \times 1000 = 0.04 \text{ mol dm}^{-3}$
3	B	$C = \frac{n}{V}$ or $C = \frac{m}{M_r \times V}$ $C = \frac{4.4}{44 \times 500} \times 1000 = 0.2 \text{ mol dm}^{-3}$
4	B	$\text{Conc.} = \frac{\text{mass}}{M_r \times \text{Volume}} = \frac{1.0}{40 \times 250} \times 1000 = 0.10 \text{ mol dm}^{-3}$

1 $M_r \text{ of Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O} = (23 \times 2) + 32 + (16 \times 4) + 10(18) = 32$

Moles (NaOH) $= 1.60 \times \frac{25.0}{1000} = 0.04 \text{ mol}$

Moles – mass ratio

NaOH	:	$\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$
2 mol	:	322g
0.04 mol	:	$\frac{322}{2} \times 0.04 = 6.44 \text{ g}$

2 (a) $M_r \text{ of SO}_4^{2-} = 96$
 $\text{Conc. (mol dm}^{-3}) = \frac{\text{Conc. (g dm}^{-3})}{M_r} = \frac{2.65}{96} = 0.027 \text{ mol dm}^{-3}$

(b) moles of Cl^- in $\text{dm}^3 = \frac{19.00}{35.5} = 0.535 \text{ mol dm}^{-3}$

moles of Cl^- in $25 \text{ dm}^3 = 0.535 \times \frac{25}{1000} = 0.0134 \text{ mol}$

mol ratio Cl^- : AgCl
 1 mol : $(108 + 35.5) = 143.5 \text{ g}$
 $0.0134 \text{ mol} : \frac{14.35}{1} \times 0.0134 = 1.92 \text{ g}$

3 conc. of $\text{MgCl}_2 = 1.26 \text{ g dm}^{-3}$ $M_r \text{ of MgCl}_2 = 95$

conc. MgCl_2 in $\text{mol dm}^{-3} = \frac{1.26}{95} = 0.0133 \text{ mol dm}^{-3}$

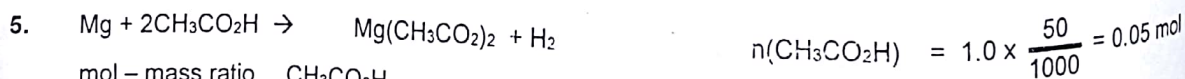
conc. of Cl^- in $\text{mol dm}^{-3} = 0.0133 \times 2 = 0.0266 \text{ mol dm}^{-3}$

4. $M_r \text{ of NaOH} = 40$ Mole $= \frac{3.2}{40} = 0.08$

Conc. $= \frac{\text{mol}}{\text{vol}}$

0.10 $= \frac{0.08}{\text{Vol}}$

Vol $= \frac{0.08}{0.01} = 0.8 \text{ dm}^3 = 800 \text{ cm}^3$



mol – mass ratio

$\text{CH}_3\text{CO}_2\text{H}$	:	Mg
2 mol	:	24 g
0.05 mol	:	$\frac{24}{2} \times 0.05 = 0.60 \text{ g}$

6. (i) $\text{moles(HCl)} = 1.5 \times \frac{25}{1000} = 0.0375 \text{ mol}$
 mol-vol ratio $\text{HCl} : \text{CO}_2$
 $2 \text{ mol} : 24 \text{ dm}^3$
 $0.0375 \text{ mol} : \frac{24}{2} \times 0.0375 = 0.45 \text{ dm}^3 = 450 \text{ cm}^3$

(ii) $25/1000 \times 1.5 \times 24 = 0.9 \text{ dm}^3 / 900 \text{ cm}^3$

7. $\text{Moles of NaCl} = 3.5 \times 55 = 192.5 \text{ mol}$
 Mol - vol. ratio $\text{NaCl} : \text{Cl}_2$
 $2 \text{ mol} : 24 \text{ dm}^3$
 $192.5 \text{ mol} : \frac{24}{2} \times 192.5 = 2310 \text{ dm}^3$

$\text{mol LiOH} = 0.500 \times \frac{20.0}{1000} = 0.01 \text{ mol}$

$\text{mass} = 123 \times 0.01 = 1.23 \text{ g (1)}$

molar mass of hydrated lithium nitrate = 123 (1)

$\text{mol sulfuric acid} = 2.0 \times \frac{15}{1000} = 0.03 \text{ (1)}$

$\text{mass of hydrated copper sulfate} = 7.5 \text{ g (1)}$

molar mass of hydrated copper sulfate = 250 (1)

$\text{mol (HI)} = 0.94 \times \frac{50}{1000} = 0.047 \text{ mol}$

$\text{mass (HI)} = 0.047 \times 128 = 6.016 \text{ g}$

VOLUMETRIC EXPERIMENTS

MCQ'S

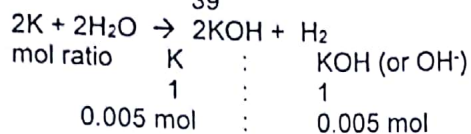
MCQ	Key	Solution & Explanation
1	D	$\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$ mole ratio $1 : 2$ moles of NaOH should be double than that of H_2SO_4 $\text{mole of H}_2\text{SO}_4 = 0.5 \times \frac{20}{1000} = 0.01 \text{ mol}$ so moles of NaOH = 0.02 mol A $\frac{20}{1000} \times 0.5 = 0.01 \text{ mol}$ B $\frac{10}{1000} \times 0.5 = 0.05 \text{ mol}$ C $\frac{40}{1000} \times 1.0 = 0.04 \text{ mol}$ D $\frac{20}{1000} \times 1.0 = 0.02 \text{ mol}$
2	A	$\text{KOH} + \text{HNO}_3 \rightarrow \text{KNO}_2 + \text{H}_2\text{O}$ $\text{Moles (KOH)} = 1.0 \times \frac{15.0}{1000} = 0.015 \text{ mol}$ Mol ratio KOH : $\text{HNO}_3 = 1 : 1$ $\text{Concentration (HNO}_3) = \frac{0.015}{20.0} \times 1000 = 0.75 \text{ mol dm}^{-3}$
3	D	$\text{Moles of NaOH} = 1.0 \times \frac{100}{100} = 0.1 \text{ mol}$ Mole ratio $\text{NaOH} : \text{H}_2\text{SO}_4 = 2 : 1$ So mole of H_2SO_4 required is $\frac{1}{2} \times 0.1 = 0.05 \text{ mol}$ A $2.0 \times 50/100 = 0.1 \text{ mol}$ B $1.0 \times \frac{100}{100} = 0.1 \text{ mol}$ C $0.5 \times \frac{25}{100} = 0.0125 \text{ mol}$ D $1.0 \times \frac{50}{100} = 0.05 \text{ mol}$

PREPARED BY : KASHIF HASNAT MALIK

STRUCTURED QUESTIONS

1. 2 dm^3 of aqueous sodium hydroxide of concentration 5 mol / dm^3 were

$$(i) \quad n(K) = \frac{0.195}{39} = 0.005 \text{ mol}$$



$$(ii) \quad n(H^+) = C \times V = 0.100 \times \frac{100}{1000} = 0.010 \text{ mol}$$



- (iv) pH is 1 to 4
because an excess of HCl present
or an extra 0.005 mol acid present

$$2. \quad n(NaOH) = 0.150 \times \frac{20.0}{1000} = 0.003 \text{ mol}$$

$$n(H_2SO_4) = \frac{1}{2} \times 0.003 = 0.0015 \text{ mol}$$

$$\text{Conc.} = \frac{0.0015}{12.0} \times 1000 = 0.125 \text{ mol dm}^{-3}$$

$$3. \quad (a) \quad M_r \text{ of } (NH_4)_2SO_4 = 132$$

$$\% \text{ of N} = \frac{14 \times 2}{132} \times 100 = 21.2\%$$

$$(d) \quad (i) \quad 0.0200 \times \frac{22.5}{1000} = 4.5 \times 10^{-4} \text{ mol KMnO}_4$$

$$(ii) \quad n(Fe^{2+}) = 4.5 \times 10^{-4} \times 5 = 2.25 \times 10^{-3} \text{ mol}$$

$$\text{mass } (Fe^{2+}) = 2.25 \times 10^{-3} \times 56 = 0.126 \text{ g}$$

$$4 \quad (i) \quad C_2H_3O_3$$

$$(ii) \quad \text{Moles (KOH)} = 0.100 \times \frac{6.00}{1000} = 0.0006 \text{ mol}$$

Mol ratio	KOH	:	tartaric acid
	2	:	1
	0.0006	:	$\frac{1}{2} \times 0.006 = 0.003 \text{ mol}$

$$\text{Concentration (tartaric acid)} = \frac{0.003}{20.0} \times 1000 = 0.15 \text{ mol dm}^{-3}$$

$$(iii) \quad \frac{7.40}{8.00} \times 100 = 92.5\%$$

$$5 \quad (a) \quad M_r \text{ of } FeSO_4 = 152$$

$$\% \text{ mass of Fe} = \frac{56}{152} \times 100 = 36.8\%$$

$$(b) \quad (i) \quad \text{moles } (K_2Cr_2O_7) = 0.0400 \times \frac{19.0}{1000} = 0.00076 \text{ mol}$$

$$(ii) \quad \text{mol } (Fe^{2+}) = 0.00076 \times 6 = 0.00456 \text{ mol}$$

$$\text{mass } (Fe^{2+}) = 0.00456 \times 56 = 0.255 \text{ g}$$

$$6 \quad \text{moles (NaOH)} = 0.200 \times \frac{18.0}{1000} = 0.0036 \text{ mol}$$

mol ratio	NaOH	:	fumaric acid
	2 mol	:	1 mol
	0.0036 mol	:	$\frac{1}{2} \times 0.0036 = 0.0018 \text{ mol}$

$$\text{conc. (fumaric acid)} = \frac{0.0018}{60.0} \times 1000 = 0.003 \text{ mol dm}^{-3}$$

$$7 \quad \text{Mol (HCl)} = 0.0150 \times \frac{6.00}{1000} = 0.0009 \text{ mol}$$

$$\begin{array}{lcl} \text{Mol ratio} & \text{HCl} & : \quad \text{Ca(OH)}_2 \\ & 2 & : \quad 1 \\ & 0.0009 & : \quad \frac{1}{2} \times 0.0009 = 0.000045 \text{ mol} \end{array}$$

$$\text{concentration (Ca(OH)}_2) = \frac{0.000045}{20} \times 1000 = 0.00225 \text{ mol dm}^{-3}$$

$$8 \quad \text{Moles of sulfamic acid} = \frac{0.105}{97} / 0.00107 (1)$$

$$\text{Moles of KOH} = \frac{10.8}{1000} \times 0.100 / 0.00108 (1)$$

so reacts with one mole (1)

$$9 \quad \text{moles NaOH} = 0.100 \times \frac{28.0}{1000} = 0.0028 \text{ mol}$$

$$\begin{array}{lcl} \text{mol ratio} & = & \text{NaOH} : \quad \text{H}_2\text{SO}_4 \\ & & 2 \quad : \quad 1 \\ & 0.0028 & : \quad 0.0028 \times \frac{1}{2} = 0.0014 \text{ mol} \end{array}$$

$$\text{conc. (H}_2\text{SO}_4) = \frac{0.00147}{9.50} \times 1000 = 0.147 \text{ mol dm}^{-3}$$

$$10. \quad (i) \quad \begin{array}{lcl} \text{Mr of Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O} & = & (23 \times 2) + (32 \times 2) + (16 \times 3) + 5(18) = 248 \\ \text{conc. in mol dm}^{-3} & \neq & 12.4 \div 248 = 0.05 \text{ mol dm}^{-3} \end{array}$$

$$(ii) \quad \text{mol (Na}_2\text{S}_2\text{O}_3) = 0.05 \times \frac{23.6}{1000} = 0.00118 \text{ mol}$$

$$\begin{array}{lcl} \text{Mol ratio} & \text{Na}_2\text{S}_2\text{O}_3 & : \quad \text{I}_2 \\ & 2 & : \quad 1 \\ & 0.00118 & : \quad \frac{1}{2} \times 0.0018 = 0.00059 \text{ mol} \end{array}$$

$$\text{Conc. (I}_2) = \frac{0.00059}{12.5} \times 1000 = 0.0472 \text{ mol dm}^{-3}$$

$$11. \quad n(\text{HCl}) = 0.0100 \times \frac{4.00}{1000} = 4 \times 10^{-5} \text{ mol}$$

$$\begin{array}{lcl} \text{mol ratio} & \text{HCl} & : \quad \text{Ca(OH)}_2 \\ & 2 & : \quad 1 \\ & 4 \times 10^{-5} & : \quad \frac{1}{2} \times 4 \times 10^{-5} = 2 \times 10^{-5} \text{ mol} \end{array}$$

$$\text{conc. (Ca(OH)}_2) = \frac{2 \times 10^{-5}}{10.0} \times 1000 = 2 \times 10^{-3} \text{ mol dm}^{-3}$$

$$12. \quad 3\text{Fe} + 4\text{H}_2\text{O(g)} \rightleftharpoons 4\text{H}_2\text{(g)} + \text{Fe}_3\text{O}_4 \quad \Delta H = +35 \text{ kJ / mol}$$

A_r of Fe = 56,

A_r of Fe_3O_4 = 232

$$\begin{array}{lcl} \text{Mass ratio} & \text{Fe} & : \quad \text{Fe}_3\text{SO}_4 \\ & 3(56) = 168\text{g} & : \quad 232 \text{ g} \\ & 2.80 \text{ g} & : \quad \frac{232}{168} \times 2.80 = 3.867 \text{ g} \end{array}$$

13. moles (NaOH) = $0.100 \times \frac{23.2}{1000} = 0.00232 \text{ mol}$

mole ratio NaOH : acid
 $\frac{1}{0.00232 \text{ mol}} : \frac{1}{0.00232 \text{ mol}}$

moles = $\frac{\text{mass}}{M_r}$

$0.00232 = \frac{0.172}{M_r}$

$M_r = \frac{0.172}{0.00232} = 74.14 \approx 74$

$M_r \text{ of } C_xH_yCO_2H = 74$

$C_xH_y + 12 + 32 + 1 = 74$

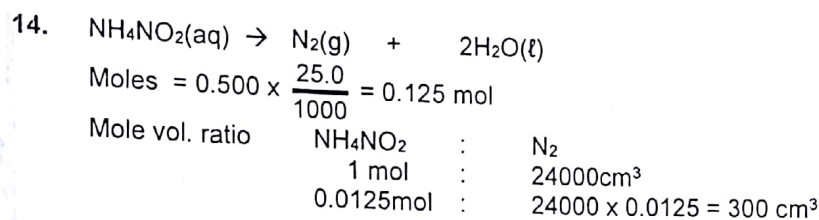
$C_xH_y + 45 = 74$

$C_xH_y = 74 - 45$

$C_xH_y = 29$

$12 \times 2 + 5 = 29$ $x - 2 \text{ and } H = 5$ if $x = 2$ then $y = 5$

Identity of acid = $C_2H_5CO_2H$ propanoic acid



15. moles (HCl) = $0.040 \times \frac{18.0}{1000} = 7.2 \times 10^{-4} \text{ mol}$

mole ratio HCl : $Ca(OH)_2$
 $2 \text{ mol} : 1 \text{ mol}$

$7.2 \times 10^{-4} \text{ mol} : \frac{1}{2} \times 7.2 \times 10^{-4}$

conc. $(Ca(OH)_2) = \frac{3.6 \times 10^{-4}}{25.0} \times 1000 = 0.0144 \text{ mol dm}^{-3}$

16. Moles of HF = 0.01 (1)
 Moles of $Ca(OH)_2 = 0.005$ / moles of $Ca(OH)_2 = 0.5 \times$ moles of HF (1)
 Volume = 33.3 cm^3 (1)

17. moles (H_2SO_4) = $14.0 \times \frac{25.0}{1000} = 0.35 \text{ mol}$

Mol ratio = $H_2SO_4 : SO_2 = 2:1$

Moles of $SO_2 = 0.35 \times \frac{1}{2} = 0.175 \text{ mol}$

Vol. of $SO_2 = 0.175 \times 24 = 4.2 \text{ dm}^3$ or 4200 cm^3

18. $M_r \text{ of } H_2S = (1 \times 2) + 32 = 34$

Mass - mol ratio $H_2S : KOH$
 $34 \text{ g} : 2 \text{ mol}$

$0.170 \text{ g} : \frac{2}{34} \times 0.170 = 0.01 \text{ mol}$

$C = \frac{n}{V} \Rightarrow 0.150 = \frac{0.01}{V}$

$V = \frac{0.01}{0.150} = 0.0667 \text{ dm}^3 = 66.67 \text{ cm}^3$

CALCULATING WATER OF CRYSTALLIZATION

MCQ'S

MCQ	Key	Solution & Explanation
1	D	$\frac{162}{18} = 9$ $\text{Na}_2\text{S} \cdot 9\text{H}_2\text{O}$
2	C	$\text{Mr of Cu(NO}_3)_2 = 64 + 2 [14 + (16 \times 3)] = 188$ $\text{Mr of H}_2\text{O} = 18$ $\text{Cu(NO}_3)_2 : \text{H}_2\text{O}$ $(100 - 36.5) = \frac{63.5}{188} : \frac{36.5}{18}$ $0.338 : 2.028$ $\frac{0.338}{0.338} = 1 : \frac{0.028}{0.338} = 6$

$$\begin{aligned} \text{Mr of [Cu(CH}_3\text{CO}_2)_2]_2 \text{ Cu(OH)}_2 \cdot x\text{H}_2\text{O} &= 552 \\ 2[64 + 2 \{(12 \times 2) + (1 \times 3) + (16 \times 2)\}] + 64 + \{16 + 1\}2 + x(2 + 16) &= 552 \\ 462 + 18x &= 552 \\ 18x &= 90 \\ x &= 5 \end{aligned}$$

STRUCTURED QUESTIONS

PERCENTAGE PURITY

MCQ'S

MCQ	Key	Solution & Explanation
1	C	$\frac{45}{50} \times 100 = 90\%$
2	D	$\text{Mass of CaCO}_3 = 0.226 \times 100 = 22.6 \text{ g}$ $\% \text{ purity} = \frac{22.6}{23.0} \times 100 = 98.3\%$

PERCENTAGE YIELD

MCQ'S

MCQ	Key	Solution & Explanation
1	B	$\text{Mass ratio H}_2 : \text{NH}_3$ $3(2) = 6\text{g} : (17 \times 2) 34\text{g}$ $60 \text{ kg} : \frac{34}{6} \times 60 = 340 \text{ kg}$ $\text{Expected held} = 340 \text{ kg}$ $\% \text{ yield} = \frac{60}{340} \times 100 = 17.6\%$

STRUCTURED QUESTIONS

1. (i) Mass ratio $\text{Mg} : \text{MgO}$
 $2(24) = 48 \text{ g} : 2(40) 80$
 $12 \text{ kg} : \frac{80}{48} \times 12 = 20 \text{ kg}$

$\% \text{ yield} = \frac{\text{Actual Yield}}{\text{Calculated Yield}} \times 100$
 $75 = \frac{\text{Actual Yield}}{20} \times 100$
 $\text{Actual Yield} = 75 \times \frac{20}{100} = 15 \text{ kg}$

(ii) Mg_3N_2 2. (i) Mass ratio CO_2 : CH_4

44g : 16g

220 g : $\frac{16}{44} \times 220 = 80 \text{ g}$ (ii) $\frac{46}{80} \times 100 = 57.5\%$ 3. (a) Mr of $(\text{NH}_4)_3\text{PO}_4 = 149$ % of N = $\frac{14 \times 3}{149} \times 100 = 28.19\% \approx 28.2\%$ (b) (i) moles $(\text{H}_3\text{PO}_4) = 1.25 \times \frac{25.0}{1000} = 0.03125 \text{ mol}$ mol ratio : H_3PO_4 : NH_3

1 : 3

0.03125 mol : $3 \times 0.03125 = 0.09375 \text{ mol}$ conc. $(\text{NH}_3) = \frac{0.09375}{45.3} \times 1000 = 2.07 \text{ mol dm}^{-3}$ (ii) Mr of $(\text{NH}_4)_3\text{PO}_4 = 149$ Mass of $(\text{NH}_4)_3\text{PO}_4 = 0.03125 \times 149 = 4.66 \text{ g}$ (iii) % yield = $\frac{2.93}{4.66} \times 100 = 62.9\%$

4. Mr of HCN = 27

Mr of $\text{CH}_4 = 16$

Mass ratio

 CH_4 : HCN $2(16) = 32 \text{ g}$: $2(27) = 54 \text{ g}$ 500 g : $\frac{54}{32} \times 500 = 843.75$ $843.75 \times 65\% = 548.44 \text{ g}$ 5. Mr of glycolic acid, $\text{CH}_2(\text{OH})\text{COOH} = (12 \times 2) + (16 \times 3) + (1 \times 4) = 76$ Mr of methanol, $\text{CH}_2\text{O} = 12 + (1 \times 2) + 16 = 30$

Mass ratio

Methanol : glycolic acid

30 g : $\frac{76}{30} \times 1800 = 4560 \text{ g}$ % yield = $\frac{\text{actual yield}}{\text{calculated yield}} \times 100$ $\Delta 5 = \frac{\text{actual yield}}{4560} \times 100$ Actual Yield = $\frac{45 \times 4560}{100} = 2052 \text{ g}$

6. (i) Mol ratio

NO : N_2 (ii) $\frac{1.0}{1.2} \times 100 = 83.3\%$

2 : 1

 2.4 dm^3 : $\frac{1}{2} \times 2.4 = 1.2 \text{ dm}^3$ 7 (a) mol Sn = $\frac{5.95}{119}$ mol Cl = $\frac{3.55}{35.5}$

dividing masses by correct atomic masses (1)

(mol Sn = 0.05 and mol Cl = 0.1) formula is SnCl_2 (b) mol tin(II) oxide = $\frac{13.5}{135}$ OR 0.10 molmass tin(IV) oxide expected = $0.10 \times 151 = 15.1 \text{ g}$ (1)% yield = $\frac{12.7}{15.1} \times 100 = 84\%$ (1)

(i) Moles $(\text{Ba}(\text{NO}_3)_2) = \frac{20.0 \times 0.550}{1000} = 0.011 \text{ mol}$
 Moles ratio $\text{Ba}^{2+} : \text{BaSO}_4 = 1:1$
 Mass $(\text{BaSO}_4) = 0.011 \times 233 = 2.563 \text{ g}$

(ii) $\frac{1.92}{2.563} \times 100 = 74.91\%$

(i) moles $(\text{H}_2\text{SO}_4) = \frac{20.0 \times 0.650}{1000} = 0.013 \text{ mol}$
 mole ratio = $\text{H}_2\text{SO}_4 : \text{K}_2\text{SO}_4 = 1:1$
 mass $(\text{K}_2\text{SO}_4) = 0.013 \times 174 = 2.62 \text{ g}$

(ii) $\frac{1.72}{2.262} \times 100 = 76.01\%$

LIMITING REACTANT

MCQ

	NO	:	O ₂	:	NO ₂
	2		1		1
Vol ratio = mol ratio	1		2		1

NO is limiting. 1 dm³ of NO require $\frac{1}{2}$ dm³ O₂ to form 1 dm³ of NO₂. O₂ is an excess.

Correct Option : A

STRUCTURED QUESTIONS

1. (i) mol (HCl) = $2.0 \times \frac{5.0}{1000} = 0.01 \text{ mol}$

mol (Mg) = $\frac{0.24}{24} = 0.01$

moles of HCl should be double than of Mg i.e. 0.02 mol of HCl is expected. So Mg is in excess.

(ii) mol ratio

HCl	:	MgCl ₂
2 mol	:	1 mol
0.01 mol	:	$\frac{1}{2} \times 0.01 = 0.005 \text{ mol}$

Mass of MgCl₂ = $0.005 \times 95 = 0.475 \text{ g}$

(iii) ANY 3 of: [3]

- same number of moles of each acid / same amount of replaceable hydrogen in each acid / same number of hydrogen ions which react in each acid;

ALLOW: same concentration of each acid at the same volume

- hydrochloric acid is a strong acid and ethanoic acid is a weak acid / hydrochloric acid is stronger than ethanoic acid ORA;

- hydrochloric acid fully ionised and ethanoic acid partially ionised

ALLOW: hydrochloric acid more ionised than ethanoic acid ORA

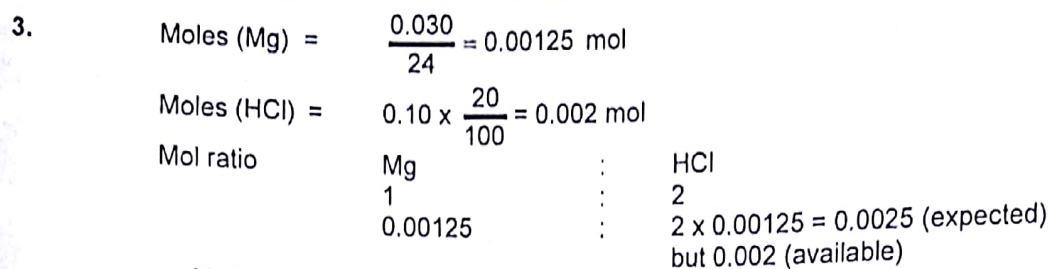
- higher concentration of hydrogen ions in hydrochloric acid / lower concentration of hydrogen ions in ethanoic acid; more collisions per unit time / collision rate higher with hydrochloric than with ethanoic acid ORA

2. (i) M_r of Fe₂O₃ = 160

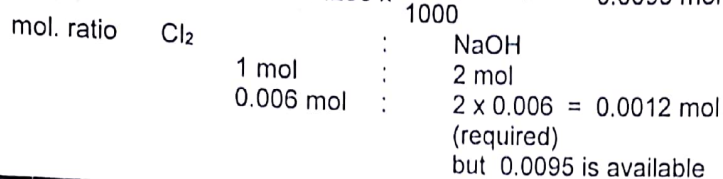
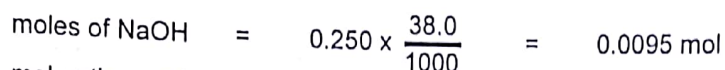
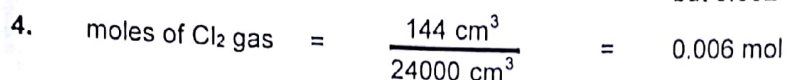
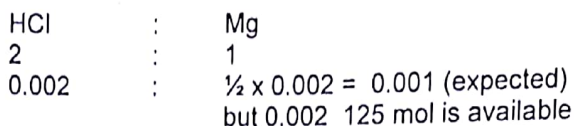
Moles = $\frac{2.00}{160} = 0.0125 \text{ mol}$

(ii) moles (KOH) = $4.00 \times \frac{20.0}{1000} = 0.08 \text{ mol}$

(iii) 1 mol of Fe₂O₃ react with 10 mol KOH
 For 0.08 mol of KOH, 0.008 mol Fe₂O₃ required but available is 0.0125 mol
 Fe₂O₃ is an excess.



so Mg is in excess
or alternatively



so Cl_2 is in excess

FINDING FORMULA OF A HYDROCARBON

MCQ'S

MCQ	Key	Solution & Explanation	MCQ'S																																								
1	C																																										
2	B	$\text{C}_x\text{H}_y + (x + \frac{y}{4}) \text{O}_2 \rightarrow x\text{CO}_2 + \frac{y}{2} \text{H}_2\text{O}$ No. of atoms of C = moles of CO_2 And no. of atom of H = $\frac{1}{2}$ moles of H_2O																																									
3	A																																										
4	C																																										
5	D	$\text{C}_x\text{H}_y + (x + \frac{y}{4}) \text{O}_2 \rightarrow x\text{CO}_2 + \frac{y}{2} \text{H}_2\text{O}$ $x + \frac{y}{4} = 5$ $3 + \frac{y}{4} = 5$ $\frac{y}{4} = 5 - 3 \qquad \frac{y}{4} = 2 \qquad y = 8$																																									
6	C	$\text{C}_x\text{H}_y + (x + \frac{y}{4}) \text{O}_2 \rightarrow x\text{CO}_2 + \frac{y}{2} \text{H}_2\text{O}$ $1\text{cm}^3 \quad 4\text{cm}^3 \rightarrow 3\text{cm}^3 + \frac{y}{2} \text{H}_2\text{O}$ Mol ratio <table> <tr> <td>X</td> <td>:</td> <td>O_2</td> <td>:</td> <td>CO_2</td> </tr> <tr> <td>1</td> <td>:</td> <td>4</td> <td>:</td> <td>3</td> </tr> </table> So <table> <tr> <td>x</td> <td>:</td> <td></td> <td>:</td> <td></td> </tr> <tr> <td>$x + \frac{y}{4}$</td> <td>:</td> <td></td> <td>:</td> <td></td> </tr> <tr> <td>$3 + \frac{y}{4}$</td> <td>:</td> <td></td> <td>:</td> <td></td> </tr> <tr> <td>$\frac{y}{4}$</td> <td>:</td> <td>$4 - 3$</td> <td>:</td> <td>$\frac{y}{4}$</td> </tr> <tr> <td>y</td> <td>:</td> <td>4</td> <td>:</td> <td></td> </tr> <tr> <td>x</td> <td>:</td> <td></td> <td>:</td> <td>C_3H_4</td> </tr> </table>	X	:	O_2	:	CO_2	1	:	4	:	3	x	:		:		$x + \frac{y}{4}$	:		:		$3 + \frac{y}{4}$	:		:		$\frac{y}{4}$	:	$4 - 3$	:	$\frac{y}{4}$	y	:	4	:		x	:		:	C_3H_4	
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x	:		:	C_3H_4																																							

PREPARED BY : KASHIF HASNAT MALIK

MCQ	Key	Solution & Explanation
7	A	$\text{alkene} + \text{O}_2 \longrightarrow \text{CO}_2 + \text{H}_2\text{O}$ $20\text{cm}^3 \quad \quad \quad 60\text{cm}^3$ $1 \quad \quad \quad 3$ 1 mol of alkene produces 3 moles CO_2 which means there are 3 carbon atoms In alkene no. of 'H' are double than 'C' atom.
8	C	using $\text{C}_x\text{H}_y + (x + y/4)\text{O}_2 \rightarrow \text{CO}_2 + y/2 \text{H}_2\text{O}$ $\begin{array}{ccccccc} & & \text{O}_2 & : & \text{CO}_2 & & \\ \text{Vol ratio} & : & 1\text{cm}^3 & : & 4\text{cm}^3 & : & 3\text{cm}^3 \\ \text{Vol ratio} = \text{mol ratio} & 1 & : & 4 & : & 3 \end{array}$ so $x = 3$ and $x + y/4 = 4$ $3 + y/4 = 4$ $y/4 = 1$ $y = 4$ so $\text{C}_x\text{H}_y = \text{C}_3\text{H}_4$

FINDING FORMULA OF A COMPOUND

MCQ'S

MCQ	Key	Solution & Explanation
1	C	Cl shows -1 oxidation state in binary compound so A and B cannot be option in which Cl show -3 and -2 oxidation state. For 'C' A_r of Cl = 35.5, so $75.5 - 35.5 = 39$ i.e. potassium For D $74.5 - (35.5 \times 2) = 3.5$, no metal have A_r of 3.5
2	D	or $\text{Al}_2\text{O}_3 \cdot \text{SiO}_2 \cdot 2\text{H}_2\text{O}$ or $\text{Al}_2\text{Si}_2\text{O}_4\text{H}_4 \text{O}_5$ or $\text{Al}_2(\text{OH})_4\text{Si}_2\text{O}_5$
3	B	mol $\frac{264\text{g}}{88} = 3$ $\frac{213\text{g}}{35.5} = 6$ atom $\frac{3}{3} : \frac{6}{3}$ 1 : 2 Formula : SrCl_2

MCQ	Key	Solution & Explanation
4	A	$\text{X}_2 + \text{H}_2 \rightarrow 2\text{HX}$ $1 \text{ vol} \quad 1 \text{ vol} \quad 2 \text{ vol}$
5	A	No. of molecules ratio $\text{Cl}_2 : \text{H}_2$ $\frac{71}{71} : \frac{2}{2}$ $1 : 1$

STRUCTURED QUESTIONS

$$\begin{array}{lcl} M_r \text{ of Ni (CO)}_x & = & 171 \\ 59 + x(12 + 16) & = & 171 \\ 28x & = & 171 - 59 \end{array} \quad \quad \quad \begin{array}{lcl} 28x & = & 112 \\ x & = & 4 \end{array}$$

CALCULATION OF IONS FROM A COMPOUND

- C

$$\text{H}_2\text{SO}_4 \rightleftharpoons 2\text{H}^+ + \text{SO}_4^{2-}$$

$$\begin{array}{ccc} 1 \text{ mol} & 2 \text{ mol} & \\ 0.05 \text{ mol} & 2 \times 0.05 = 0.1 \text{ mol} & \end{array} \quad \quad \quad 0.1 \times 1 = 0.1 \text{ g dm}^{-3}$$
- (i) $0.050 \times \frac{240}{1000} = 0.0012 \text{ mol}$ (ii) Mole of $\text{OH}^- = \left(0.060 \times \frac{20.0}{1000}\right) 2 = 0.0024 \text{ mol}$

1 mol of base give two mole OH^-

PREPARED BY : KASHIF HASNAT MALIK

ATP QUESTIONS (Year-wise)

June 1999

- 3 A student used zinc to prepare zinc nitrate. The zinc nitrate was then heated to produce zinc oxide. Some zinc was put in a weighed evaporating dish. The dish was reweighed.

Mass of dish + zinc = 22.20 g

Mass of dish = 15.70 g

- (a) Calculate the mass of zinc.

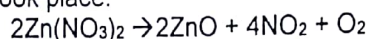
The mass of zinc = g. [1]

An acid was slowly added until all the zinc had dissolved.

Zinc nitrate solution was produced.

- (b) Name the acid.....[1]

The solution was evaporated to dryness and the resultant solid was heated in a fume cupboard. The following reaction took place.



- (c) Why was the heating done in a fume cupboard?

.....[1]

After cooling, the dish was weighed. It was then heated again, cooled and reweighed.

- (d) Why was the dish heated again?

.....[1]

Final mass of dish + zinc oxide = 23.40 g

- (e) Calculate the mass of zinc oxide. g [1]

- (f) Using your answers to (a) and (e), calculate the percentage yield of zinc oxide from the initial mass of zinc. Ar : Zn, 65; O, 16.

.....[2]

November 1999

- 2 A student added 100 cm³ of 1.00 mol / dm³ barium nitrate to 100 cm³ of 2.00 mol / dm³ sodium sulphate. Barium sulphate was produced.

- (a) (i) Give the formula of barium sulphate.....
(ii) Describe the appearance of barium sulphate in the resulting mixture.

- (iii) How could the barium sulphate be removed from the mixture?

- (b) Which of the reagents, barium nitrate or sodium sulphate, was in excess? Explain your answer

explanation

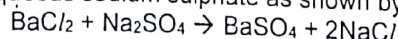
- (c) (i) The student expected a yield of 0.100 mol of barium sulphate. Calculate the expected mass of barium sulphate, using your answer in (a) (i). [Ar: Ba = 56, S = 32, O = 16]

- (ii) The mass of dry barium sulphate was 12.5 g.
Calculate the percentage yield of barium sulphate.

.....[2]

June 2000

- 5 Aqueous barium chloride reacts with aqueous sodium sulphate as shown by the equation



When 0.20 moles of barium chloride was added to 0.40 moles sodium sulphate the number of moles of barium sulphate produced was

- (a) 0.20, (b) 0.40, (c) 0.60, (d) 0.80.

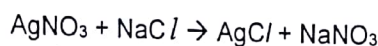
NOVEMBER 2000

A student added 20 cm³ of 2.0 mol/dm³ aqueous silver nitrate to a beaker containing aqueous sodium chloride. A precipitate of silver chloride was produced.

50 cm³ of 1.0 mol/dm³

- (a) (i) What colour was the precipitate?
- (ii) How could the precipitate be separated from the solution?.....[2]
- (b) (i) How many moles of silver nitrate are there in 20 cm³ of 2.0 mol/dm³ aqueous silver nitrate?
- (ii) How many moles of sodium chloride are there in 50 cm³ of 1.0 mol/dm³ aqueous sodium chloride?

Sodium chloride reacts with silver nitrate according to the equation below.[2]



- (c) Using this equation and your answers in (b), calculate the mass of silver chloride produced in the reaction. (Ar: Ag, 108; Cl, 35.5.)

- (d) The experiment was repeated using 50 cm³ of 1.0 mol/dm³ aqueous silver nitrate with 20 cm³ of 2.0 mol/dm³ sodium chloride. How does this mass of silver chloride produced compare with the mass obtained in (c)? Explain your answer.[2]

A student made magnesium oxide by heating magnesium in air. What mass of magnesium oxide, MgO, was made from 6.0 g of magnesium? (Ar: Mg, 24; O, 16.)

- (a) 8.0g (b) 10.0g (c) 12.0g (d) 16.0g

NOVEMBER 2001

A student was given some hydrated copper(II) sulphate crystals, CuSO₄·xH₂O. They were placed in a previously weighed test-tube which was then reweighed.

- (a) What colour are hydrated copper(II) sulphate crystals?
- Mass of test-tube + hydrated copper(II) sulphate crystals = 9.25 g
- Mass of test-tube = 5.40 g

- (b) Calculate the mass of hydrated copper(II) sulphate used in the experiment.[1]

The crystals were gently heated until they became anhydrous, i.e. no more water vapour was given off. The crystals changed colour and became powdery.

- (c) What colour was the copper(II) sulphate after heating?

Mass of test-tube + copper(II) sulphate after heating = 7.90 g

- (d) (i) Calculate the mass of copper(II) sulphate which remained after heating.[2]

- (ii) Calculate the mass of water lost from the crystals.

- (e) Calculate
- (i) the relative formula mass of anhydrous copper(II) sulphate, (Ar: Cu, 64; S, 32; O, 16)

- (ii) the relative molecular mass of water.

- (f) Using your answers to (d) and (e), calculate
- (i) how many moles of anhydrous copper(II) sulphate remained after heating,
- (iii) how many moles of water were lost on heating.[2]
- (g) The value of x in the formula $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$ can be found as follows.
- $$x = \frac{\text{answer to (f)(ii)}}{\text{answer to (f)(i)}}$$

Calculate the value of x . Hence write the formula of hydrated copper(II) sulphate crystals.[2]

The formula is

November 2002

- 2 Silver iodide may be made by the reaction between aqueous potassium iodide and aqueous silver nitrate. A student added 50 cm^3 of 1.0 mol/dm^3 potassium iodide to 30 cm^3 of 2.0 mol/dm^3 silver nitrate.



- (a) (i) Describe what was seen during the reaction.
- (ii) How could the silver iodide be removed from the mixture? [3]
- (b) (i) Which of the reagents potassium iodide or silver nitrate was in excess? Explain your answer.
- answer [3]
- explanation [5]
- (ii) Calculate the mass of silver iodide formed (Ar: Ag, 108; I, 127.)
- (c) The student did another experiment to make silver chloride by adding 50 cm^3 of 1.0 mol/dm^3 potassium chloride to 30 cm^3 of 2.0 mol/dm^3 silver nitrate,
- (i) Describe the appearance of the silver chloride on forming,
- on standing for a few minutes.
- (ii) Was the mass of silver chloride more than, the same or less than the mass of silver iodide in (b)(ii)? Explain your answer. (Ar: Ag, 108; Cl, 35.5.)
- answer [5]
- explanation

JUNE 2003

- 3 A student used magnesium oxide to prepare magnesium nitrate. The magnesium nitrate was then heated to re-form magnesium oxide. The object of the experiment was to determine whether the mass of magnesium oxide produced was the same as the amount used initially. Some magnesium oxide was put into a weighed evaporating dish. The dish was reweighed.

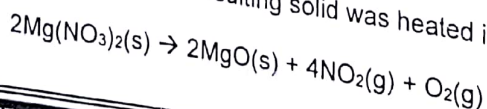
Mass of dish + magnesium oxide = 14.70 g
Mass of dish = 8.90 g

- (a) Calculate the mass of magnesium oxide used.

The mass of magnesium oxide = g [1]

- (b) An acid was slowly added until all the magnesium oxide had dissolved. Magnesium nitrate was produced. Name the acid.

- (c) The solution was evaporated to dryness and the resulting solid was heated in a fume cupboard. The following reaction took place.



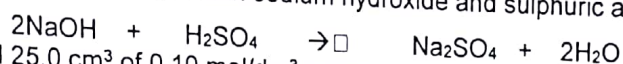
PREPARED BY : KASHIF HASNAT MALIK

After cooling, the dish was weighed. It was then heated again, cooled and reweighed.
The final mass of the dish and contents was 14.40 g.

- Why was the heating done in a fume cupboard?
- Why was the dish reweighed?
- Calculate the mass of magnesium oxide obtained. [1]
- Using your answers to (a) and (c)(iii) calculate the percentage yield of magnesium oxide. g [1]
- Suggest one reason why the experiment did not produce the same amount of magnesium oxide as was used at the beginning of the experiment. % [1]

NOVEMBER 2006

The equation for the reaction between sodium hydroxide and sulphuric acid is shown below.



A student placed 25.0 cm³ of 0.10 mol/dm³ aqueous sodium hydroxide in a flask. A few drops of an indicator were added. Aqueous sulphuric acid was added from a burette until the end-point was reached. Which of the following amounts of sulphuric acid would exactly neutralise the aqueous solution of sodium hydroxide in the flask?

- 25.0 cm³ of 0.050 mol/dm³
- 25.0 cm³ of 0.10 mol/dm³
- 50.0 cm³ of 0.050 mol/dm³
- 50.0 cm³ of 0.10 mol/dm³

JUNE 2007

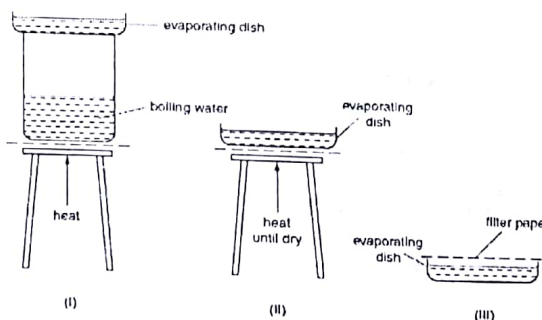
A student made some zinc sulphate by reacting zinc carbonate with dilute sulphuric acid. The powdered zinc carbonate was added to a beaker half-filled with dilute sulphuric acid. Carbon dioxide was produced.

- Give a test for carbon dioxide.
test
observation [1]
- The acid was in excess. How did the student know that the reaction had stopped? [1]
 - If the zinc carbonate was in excess, what additional observation would the student make? [1]

To ensure that all the acid had been neutralised, the student added excess zinc carbonate. The mixture was well stirred.

- How was the unreacted zinc carbonate removed from the mixture?

The solution produced in (c) was transferred to an evaporating dish. The student was told to select one of the following techniques to produce large crystals of zinc sulphate.



- Which technique, (i), (ii) or (iii), would produce the largest crystals? [1]

The student then repeated the experiment, this time adding 100 cm³ of 0.25 mol/dm³ sulphuric acid to an excess of zinc carbonate.

The equation for the reaction is



- Calculate the number of moles of sulphuric acid used in this experiment. moles [1]
- Use your answer to (e) and the equation to calculate the mass of zinc sulphate produced. [Ar: Zn, 65; S, 32; O, 16] g [1]
- Calculate the volume of carbon dioxide produced during the reaction. [One mole of a gas occupies 24 dm³ at room temperature and pressure.] dm³ [1]

PREPARED BY : KASHIF HASNAT MALIK

NOVEMBER 2007

- 4 A student was asked to make a sample of barium sulphate, BaSO_4 . She added 100 cm^3 of 0.20 mol / dm^3 sulphuric acid to 60 cm^3 of 0.25 mol / dm^3 barium nitrate. The equation for the reaction is
- $$\text{Ba}(\text{NO}_3)_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{HNO}_3$$

- (a) Describe the appearance of barium sulphate in the resulting mixture. [1]
- (b) Calculate
- (i) the number of moles of sulphuric acid used in the experiment, moles [1]
- (ii) the number of moles of barium nitrate used in the experiment, moles [1]
- (c) Using your answers to (b)(i) and (ii) calculate the maximum mass of barium sulphate that could be produced in the reaction. [Ar: Ba, 137; S, 32; O, 16] g [1]

The barium sulphate was removed from the solution by filtration. It was dried and weighed.

- (d) The mass of barium sulphate obtained was 3.35 g. Calculate the percentage yield of barium sulphate. % [1]

Another student, doing the same experiment and using the same quantities of barium nitrate and sulphuric acid, obtained 3.60 g of product.

- (e) Suggest a reason for this increased mass of product. [1]
- (f) Suggest a different barium salt that could have been used instead of barium nitrate to produce barium sulphate. [1]

JUNE 2008

- 2 A student was given some hydrated sodium carbonate crystals, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, where x is a whole number. They were placed in a previously weighed container, which was reweighed.

mass of container + sodium carbonate crystals = 9.87 g

mass of container = 5.83 g

- (a) Calculate the mass of sodium carbonate crystals used in the experiment. g [1]

The container and crystals were heated to remove the water of crystallisation and then reweighed. This process was repeated until there was no further change in mass.

- (b) Describe the appearance of the sodium carbonate crystals after heating. [2]

mass of container + sodium carbonate after heating = 7.35 g

- (c) (i) Calculate the mass of sodium carbonate which remained after heating. g
- (ii) Calculate the mass of water which was lost from the crystals. g

- (d) (i) Calculate the relative formula mass of sodium carbonate, Na_2CO_3 , and the relative formula mass of water. [Ar: Na, 23; C, 12; O, 16; H, 1]

relative formula mass of sodium carbonate

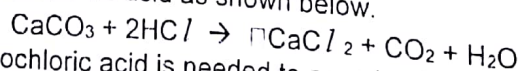
relative formula mass of water

- (e) Using your answers to (c) and (d), calculate

- (i) the number of moles of sodium carbonate which remained after heating, [1]
- (ii) the number of moles of water which were lost on heating. [1]

- (f) Using your answers to (e) calculate the value of x in the formula $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$ [1]

- 8 Calcium carbonate reacts with hydrochloric acid as shown below.



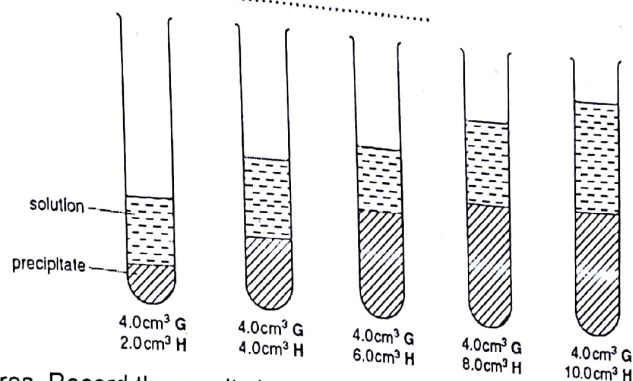
$x =$ [2]

What volume of 1.0 mol/dm^3 hydrochloric acid is needed to completely react with 2.0 g of calcium carbonate? [Mr: CaCO_3 , 100]

- (a) 20 cm^3 (b) 40 cm^3 (c) 200 cm^3 (d) 400 cm^3

The reaction between lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2$, and potassium iodide, KI , produces a precipitate of lead(II) iodide.
 Write the equation for the reaction between lead(II) nitrate and potassium iodide.

(a) Solution G is a solution of potassium iodide.
 Solution H is 1.0 mol/dm^3 lead(II) nitrate.
 4.0 cm^3 of G was placed in each of five test-tubes.
 The diagrams below show the results of adding different volumes of H to each of the tubes.
 Each test-tube was left for a while to allow the precipitate to settle.

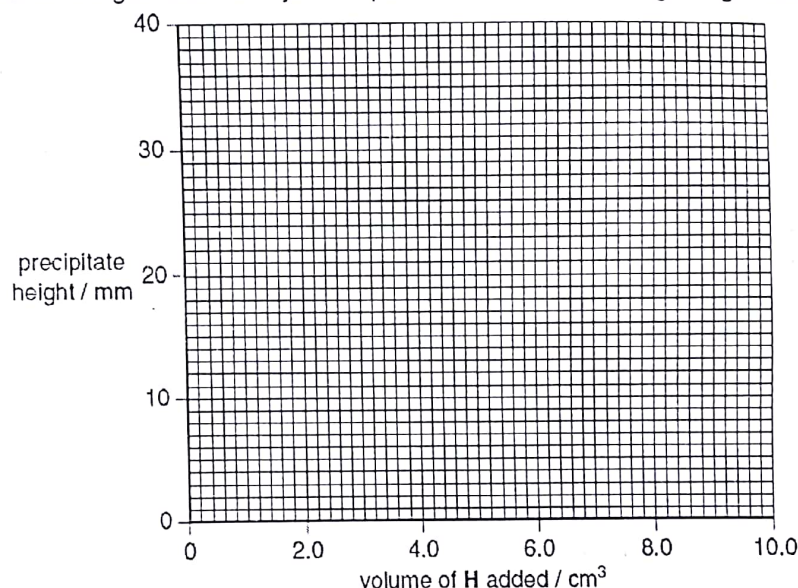


(b) Measure the height of each precipitate in millimetres. Record the results in the table below.

test-tube	1	2	3	4	5
volume of H added / cm^3	2.0	4.0	6.0	8.0	10.0
precipitate height / mm					

[2]

(c) Plot these results on the grid below and join the points with two intersecting straight lines.



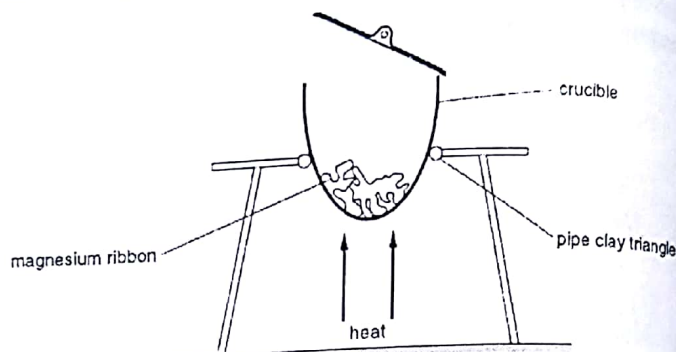
- (d) (i) What was the minimum volume of H required to completely react with 4.0 cm^3 G? cm^3 [1]
 (ii) Using your answers to (a) and (d)(i), calculate the concentration of potassium iodide in G. mol/dm^3 [2]
 (e) The experiment was repeated using 2.0 mol/dm^3 lead(II) nitrate and solution G.
 (i) Calculate the minimum volume of 2.0 mol/dm^3 lead(II) nitrate which was required to completely react with 4.0 cm^3 of G. cm^3 [1]
 (ii) What was the height of the precipitate in the test-tube when this reaction occurred? [1]

- 1 A student did the following experiment to find the formula of magnesium oxide.
A 10 cm length of magnesium ribbon was loosely coiled and placed in a previously weighed crucible which was then reweighed.
- mass of crucible + magnesium = 13.08 g
mass of crucible = 12.72 g

(a) Calculate the mass of magnesium..... g [1]

The crucible was placed on a pipe clay triangle and heated strongly for several minutes.

During the heating the crucible lid was lifted and replaced several times. The magnesium was converted into magnesium oxide.



(b) (i) Why was it necessary for a lid to be placed on the crucible during heating?

(ii) Why was the lid lifted during heating?

(c) Describe the appearance of

(i) magnesium..... [1]

(ii) magnesium oxide.....

After cooling, the crucible was weighed. It was then reheated, cooled and reweighed. final mass of crucible + magnesium oxide = 13.32 g

(d) Why was the crucible re-heated?

(e) (i) Calculate the mass of magnesium oxide. g [1]

(ii) Calculate the mass of oxygen that reacted with the magnesium. g [1]

(f) Using your answers to (a) and (e)(ii), calculate the formula of magnesium oxide.
[Ar: Mg, 24; O, 16]

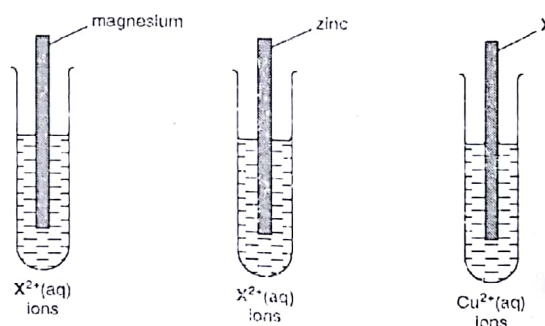
.....[2]

NOVEMBER, 2010/41

- 5 A student did an experiment to compare the reactivities of three different metals. Three test-tubes were arranged as in the diagrams below. There was a deposit formed in all three tubes. X could be

(a) calcium
(c) silver

(b) iron
(d) sodium



- 6 A compound, P, contains 40.0% carbon, 6.7% hydrogen and 53.3% oxygen. What is its empirical formula? [Ar: C, 12; H, 1; O, 16]

(a) CHO

(b) C₂H₂O

(c) CH₂O

(d) CHO₂

NOVEMBER, 2010/42

1. The apparatus below contains 0.100 mol/dm³ sulfuric acid.

(a) (iii) Using your answer to (a)(ii), calculate the number of moles of 0.100 mol/dm³ sulfuric acid.

..... moles [1]

(b) (i) The sulfuric acid was poured into a beaker and 0.12 g of magnesium was added. The magnesium reacted with the sulfuric acid and hydrogen was produced.

How many moles of magnesium were added?
[Ar: Mg, 24]

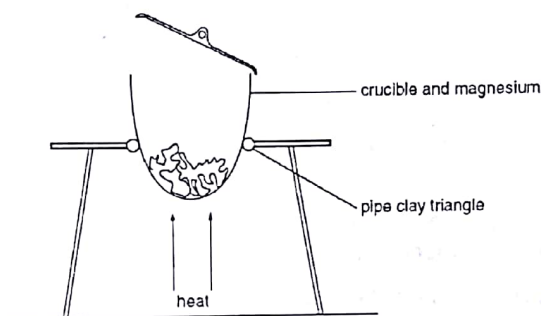
..... moles [1]



- (ii) Write the equation for the reaction between magnesium and sulfuric acid.[1]
- (iii) Using your answers to (a)(iii), (b)(i) and (b)(ii), suggest which reagent was in excess, magnesium or sulfuric acid? Explain your answer.[1]
- (c) (ii) Calculate the volume of hydrogen gas produced in this reaction. [1 mol of a gas measured at 25 °C occupies a volume of 24 dm³.]
- In an experiment to find the formula of the oxide of element M, 2.0 g of the element was burnt in oxygen. The mass of metal oxide obtained was 2.8 g. [Ar: M, 40; O, 16] What was the formula of the metal oxide?
- (a) M₂O (b) MO (c) MO₂ (d) MO₃

NOVEMBER, 2011/41

A student does an experiment to find the formula of magnesium oxide. A 10 cm length of magnesium is loosely coiled and placed in a weighed crucible. The crucible is heated for several minutes during which the crucible lid is raised from time to time. The magnesium changes to magnesium oxide.



- (a) mass of crucible + magnesium = 14.33 g
mass of crucible = 13.85 g

Calculate the mass of magnesium. g [1]

- (b) Describe the appearance of

- (i) magnesium, [1]
- (ii) magnesium oxide. [2]

- (c) After cooling, the crucible is weighed. It is reheated, cooled and reweighed. Why is this done? [1]

- (d) Final mass of crucible + magnesium oxide = 14.65 g
Calculate

- (i) the mass of magnesium oxide, g
- (ii) the mass of oxygen which reacts with the magnesium. g [2]

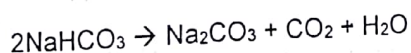
- (e) Using your answers to (a) and (d)(ii), calculate the formula of magnesium oxide. [Ar: O, 16; Mg, 24] [2]

- (f) (i) Give an equation for a reaction in which magnesium oxide reacts with an acid. [1]

- (ii) By referring to your equation, state whether magnesium oxide can be classed as an acidic, basic or amphoteric oxide. [1]

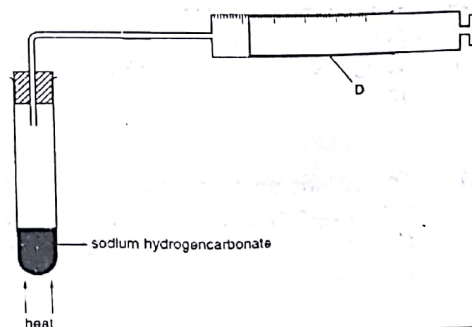
NOVEMBER, 2011/42

- 4 A student heats solid sodium hydrogencarbonate in the apparatus shown below. The carbon dioxide gas produced is collected in apparatus D.



- (a) Name apparatus D. [1]

- (b) Give a test for carbon dioxide. test observation



(c) On heating the sample of sodium hydrogencarbonate, 120 cm³ of carbon dioxide is evolved. The gas is measured at room temperature and pressure.

(i) Calculate the number of moles of carbon dioxide evolved.
[1 mole of a gas occupies 24 dm³ at room temperature and pressure.]

..... moles [1]

(ii) Using your answer to (c)(i) and the equation for the reaction, deduce the number of moles of sodium hydrogencarbonate decomposed.

..... moles [1]

(iii) Using your answer to (c)(ii), calculate the mass of sodium hydrogencarbonate decomposed. [Ar: H, 1; C, 12; O, 16; Na, 23]

..... g [1]

June 2012/41

2 When copper is heated in air it reacts with oxygen to form an oxide.
A student does an experiment to find the formula of copper oxide.

(a) Describe the appearance of copper.

[1]

(b) Some copper is placed in a previously weighed crucible and reweighed.

mass of crucible + copper = 5.92 g

mass of crucible = 4.65 g

Calculate the mass of copper used in the experiment.

..... g [1]

(c) The crucible containing the copper is heated and copper oxide is produced. The crucible with copper oxide is weighed.

mass of crucible + copper oxide = 6.24 g

Calculate the mass of copper oxide produced.

..... g [1]

(d) Using your answers to (b) and (c) calculate the mass of oxygen that combines with the copper.

..... g [1]

(e) Using your answers to (b) and (d) calculate the formula of copper oxide. Show your working. [Ar: O, 16; Cu, 64]

[2]

June 2012/42

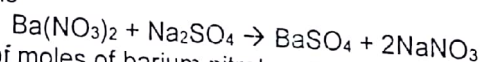
3 A student adds 50 cm³ of 1.50 mol / dm³ barium nitrate to 100 cm³ of 1.00 mol / dm³ sodium sulfate. Barium sulfate is produced.

(a) (i) Describe the appearance of barium sulfate in the resulting mixture.

(ii) How does the student produce a pure sample of barium sulfate from the original mixture?

[1]

(b) The equation for the reaction is



(i) Calculate the number of moles of barium nitrate present in 50 cm³ of 1.50 mol / dm³ solution.

..... moles [1]

(ii) Calculate the number of moles of sodium sulfate in 100 cm³ of 1.00 mol / dm³ solution.

..... moles [1]

(iii) Deduce the number of moles of barium sulfate produced.

..... moles [1]

(iv) Calculate the mass of barium sulfate produced.
[Ar: O, 16; S, 32; Ba, 137]

..... g [2]

[Total: 9]

7 The organic compound T contains 85.7% carbon and 14.3% hydrogen. What is its empirical formula? [Ar: H, 1; C, 12]

(a) CH

(b) CH₂

(c) CH₃

(d) CH₄

PREPARED BY: KASHIF HASNAT MALIK

November 2012/41

A student makes zinc sulfate by adding powdered zinc carbonate to a beaker half-filled with dilute sulfuric acid. The equation for the reaction is



- (d) The student repeats the experiment, this time adding excess zinc carbonate to 100 cm³ of 0.5 mol / dm³ sulfuric acid. Calculate the number of moles of sulfuric acid used in the experiment.

..... moles [1]

- (e) Using your answer to (d) and the equation, calculate

(i) the mass of zinc sulfate which is produced, [Ar: O, 16; S, 32; Zn, 65]

..... g [2]

November 2012/42

- (d) The student dries and weighs the silver iodide.

mass of silver iodide = 4.70 g

Calculate the number of moles of silver iodide in this mass. [Ar: Ag, 108; I, 127]

..... moles [1]

- (e) The concentration of aqueous potassium iodide used is 1.00 mol / dm³. It reacts with aqueous silver nitrate according to the following equation.



Using your answer from (d), deduce the number of moles of silver nitrate used in the reaction.

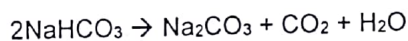
..... moles [1]

- (f) Using your answers to (a) and (e) calculate the concentration of the aqueous silver nitrate.

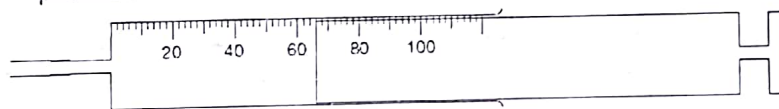
..... mol / dm³ [1]

June 2013/41

A student heats some sodium hydrogencarbonate in the apparatus shown below. The reaction produces carbon dioxide.

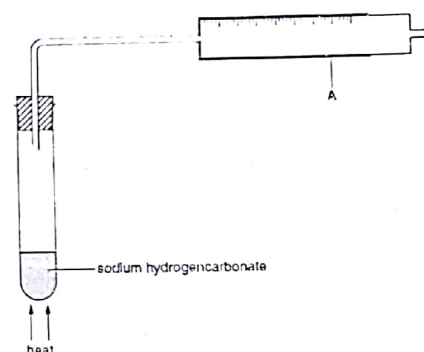


- (c) The diagram below shows apparatus A at the completion of the reaction. The carbon dioxide collected is at room temperature and pressure.



What volume of carbon dioxide is collected?

..... cm³ [1]



- (d) Using your answer to (c), calculate the number of moles of carbon dioxide collected, measured at room temperature and pressure.

(One mole of a gas occupies 24 000 cm³ at room temperature and pressure.)

..... moles [1]

- (e) (i) Using the equation for the reaction and your answer to (d) calculate the number of moles of sodium hydrogencarbonate used in the experiment.

..... moles [1]

- (ii) Calculate the relative formula mass of sodium hydrogencarbonate. [Ar: H, 1; C, 12; O, 16; Na, 23]

..... [1]

- (iii) Using your answers to (e)(i) and (e)(ii) calculate the mass of sodium hydrogencarbonate used in the experiment.

..... g [1]

- 6 A student finds that a hydrocarbon contains 88.9% by mass of carbon. What is its empirical formula?

[Ar: H, 1; C, 12]

(a) CH₂

(b) CH₃

(c) C₂H₃

(d) C₂H₅

- 7 Aqueous sodium hydroxide reacts with sulfuric acid.
 $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
 Which of the following aqueous solutions of sodium hydroxide will produce 1.42 g of sodium sulfate when reacting with excess sulfuric acid? [Mr: Na_2SO_4 , 142]

- (a) 100 cm³ of 0.100 mol / dm³ of sodium hydroxide
 (b) 50 cm³ of 0.200 mol / dm³ of sodium hydroxide
 (c) 50 cm³ of 0.400 mol / dm³ of sodium hydroxide
 (d) 100 cm³ of 0.050 mol / dm³ of sodium hydroxide

June 2013/42

- 1 Iron(II) sulfate crystals have the formula $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$, where x is a whole number. A student is asked to find the value of x . The crystals are placed in a previously weighed crucible which is then reweighed. [1]

- (a) What colour are iron(II) sulfate crystals? [1]
 (b) Mass of crucible + iron(II) sulfate crystals = 9.01 g
 Mass of crucible = 5.97 g
 Calculate the mass of iron(II) sulfate crystals used in the experiment.g [1]

- (c) The crystals are gently heated until no more water is given off.
 The crucible and contents are cooled and reweighed.

Mass of crucible and iron(II) sulfate after heating = 7.66 g

- (i) Calculate the mass of iron(II) sulfate which remains after heating.g [1]

- (ii) Calculate the mass of water lost from the crystals.g [1]

- (iii) Calculate the number of moles of iron(II) sulfate that remain after heating.

[Mr: FeSO_4 , 152]

..... moles [1]

- (iv) Calculate the number of moles of water which are lost on heating.

[Mr: H_2O , 18]

.....moles [1]

- (d) (i) Using your answers to (c)(iii) and (c)(iv) calculate the number of moles of water combined with one mole of iron(II) sulfate.

..... moles [1]

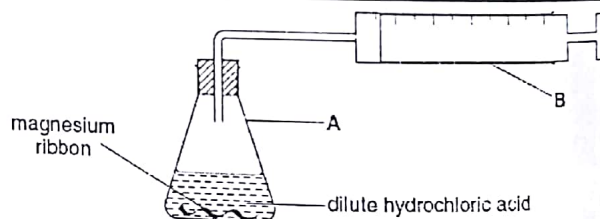
- (ii) What is the value of x in the formula $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$? $x =$ [1]

- 7 A student adds 10.0 cm³ of 0.200 mol / dm³ sulfuric acid to an excess of magnesium. Hydrogen gas is produced. The experiment is repeated with a different acid solution, again using an excess of magnesium. Which acid solution will give twice the volume of hydrogen?

- (a) 20 cm³ of 0.200 mol / dm³ hydrochloric acid
 (b) 20 cm³ of 0.100 mol / dm³ sulfuric acid
 (c) 40 cm³ of 0.200 mol / dm³ hydrochloric acid
 (d) 40 cm³ of 0.050 mol / dm³ sulfuric acid

November 2013/41

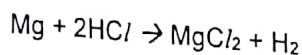
- 1 A student adds a known mass of magnesium ribbon to 100 cm³ of dilute hydrochloric acid (an excess) in the apparatus shown below. Hydrogen gas is evolved. The time taken to collect 48 cm³ of gas at room temperature and pressure is measured.



- (a) (i) Name apparatus A. [1]
 (ii) Name apparatus B. [1]

- (b) (i) Calculate the number of moles of hydrogen in the 48 cm³ of gas. [1]
 [1 mole of any gas occupies 24 000 cm³ at room temperature and pressure.]

- (ii) Use the equation below to deduce the mass of magnesium used in the experiment to produce 48 cm³ of hydrogen at room temperature and pressure.moles [1]
 [Ar : Mg, 24]



.....g [1]

PREPARED BY : KASHIF HASNAT MALIK

A compound contains 0.48 g of carbon, 0.08 g of hydrogen and 0.64 g of oxygen. [Ar : H, 1; C, 12; O, 16]
What is the empirical formula of the compound?

- (a) CHO (b) CHO₂ (c) CH₂O (d) CH₂O₂

November 2013/42

The fertiliser ammonium nitrate is a source of nitrogen. It has the formula NH₄NO₃. It can be made by adding an acid to aqueous ammonia.

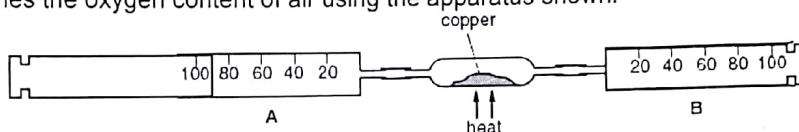
(a) Name and give the formula of this acid. name formula

(c) (i) Calculate the mass of nitrogen contained in 1000 g of ammonium nitrate.
[Ar: H, 1; N, 14; O, 16]g [2]

(ii) What volume would the mass of nitrogen calculated in (i) occupy in the gaseous state at room temperature and pressure?
[One mole of a gas occupies 24 dm³ at room temperature and pressure.] dm³ [1]

June 2014/41

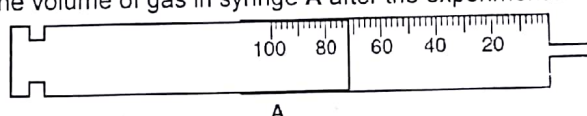
A student determines the oxygen content of air using the apparatus shown.



Syringe A contains 90 cm³ of air. The air is forced over heated copper into syringe B. The air is then forced back over the heated copper into syringe A.

The process is repeated several times until the volume of gas forced back into syringe A is constant. The apparatus is allowed to cool to room temperature.

The diagram below shows the volume of gas in syringe A after the experiment is finished.



(a) Copper reacts with oxygen in the air to produce copper(II) oxide.

(i) Construct the equation for this reaction. [1]

(ii) What colour is copper(II) oxide?
..... [1]

(b) (i) What is the volume of gas remaining in syringe A? cm³ [1]

(ii) Name the major component of the gas remaining in syringe A. [1]

(iii) Calculate the volume of oxygen that reacts with the copper. cm³ [1]

(iv) Using your answer to (b)(iii), calculate the number of moles of oxygen that react with the copper.
[One mole of a gas occupies 24 000 cm³ at room temperature and pressure.] moles

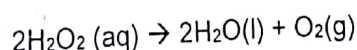
(v) Using your equation in (a)(i) and your answer to (b)(iv) calculate the mass of copper that reacts with the oxygen. [Ar: Cu, 64] g [1]

(c) In another experiment 60 cm³ of oxygen is required to react with all the copper.

Calculate the volume of air required to provide this volume of oxygen. cm³

6 A student decomposes aqueous hydrogen peroxide using manganese(IV) oxide, MnO₂, as a catalyst.

The equation for the reaction is



100 cm³ of hydrogen peroxide is allowed to completely decompose and 120 cm³ of oxygen is produced.
[One mole of a gas occupies 24 000 cm³ at room temperature and pressure.]

The concentration of the hydrogen peroxide is

- (a) 0.01 mol / dm³ (b) 0.10 mol / dm³ (c) 0.05 mol / dm³ (d) 0.50 mol / dm³

PREPARED BY : KASHIF HASNAT MALIK

- 2 A student adds 50.0 cm^3 of 1.0 mol / dm^3 aqueous silver nitrate to a beaker containing 40.0 cm^3 of 1.5 mol / dm^3 aqueous sodium bromide.
- (a) (i) A precipitate of silver bromide is produced. [1]
Suggest the colour of the precipitate. [1]
- (ii) How is the precipitate separated from the reaction mixture? [1]
- (b) (i) Calculate the number of moles of silver nitrate in 50.0 cm^3 of 1.0 mol / dm^3 silver nitrate.moles [1]
- (ii) Calculate the number of moles of sodium bromide in 40.0 cm^3 of 1.5 mol / dm^3 sodium bromide.moles [1]
- (c) The equation for the reaction is $\text{AgNO}_3 + \text{NaBr} \rightarrow \text{NaNO}_3 + \text{AgBr}$
Using your answers to (b)(i), (b)(ii) and the equation, calculate the mass of silver bromide produced in the experiment. [Ar: Ag, 108; Br, 80]g [2]
- (d) The student repeats the experiment using 50.0 cm^3 of 1.0 mol / dm^3 silver nitrate with 60.0 cm^3 of 0.5 mol / dm^3 sodium bromide. Calculate the mass of silver bromide produced in this experiment.g

- 4 What mass of magnesium oxide is produced when 3.0 g of magnesium is burned in excess oxygen?
[Ar: Mg, 24; O, 16]

(a) 3.0 g (b) 4.0 g (c) 5.0 g (d) 6.0 g

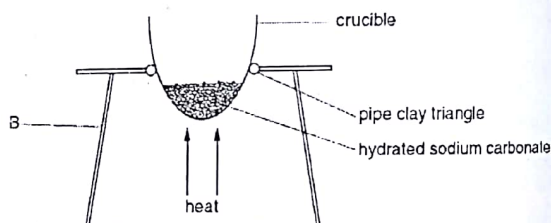
November 2014/41

- 3 A student heats a sample of hydrated sodium carbonate, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, in a crucible until all the water of crystallisation is removed. The purpose of the experiment is to determine the value of x in the formula $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$.

(a) Name the apparatus

B.....
.....

(b) How can the student be sure that all the water of crystallisation has been removed?
.....
.....



(c) After heating 0.715 g of the hydrated salt, $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$, 0.265 g of the anhydrous salt, Na_2CO_3 , remained.

- (i) Calculate the mass of water of crystallisation removed. g [1]
- (ii) Calculate the relative formula mass of sodium carbonate, Na_2CO_3 .
[Ar: Na, 23; C, 12; O, 16]

Calculate the relative formula mass of water, H_2O . [Ar: H, 1; O, 16]

(iii) Calculate the number of moles of Na_2CO_3 in 0.265 gmoles of Na_2CO_3

Calculate the number of moles of H_2O in your answer to (c)(i).moles of H_2O

(d) Using your answers to (c)(iii), calculate the value of x in the formula $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O}$.
.....

$x =$

- 7 A compound contains 0.92 g of sodium, 1.42 g of chlorine and 2.56 g of oxygen.
[Ar: Na, 23; Cl, 35.5; O, 16] Its empirical formula is

(a) NaC/O (b) NaC/O_2 (c) NaC/O_3 (d) NaC/O_4

November 2014/42

A student heats 0.336 g of a metal carbonate, MCO_3 . The sample decomposes according to the equation shown

$$MCO_3 \rightarrow MO + CO_2$$

0.176 g of carbon dioxide is produced.

- (a) Describe a test for carbon dioxide gas.
..... [1]
- (b) How can the student be sure that all the MCO_3 decomposes?
..... [1]
- (c) (i) Calculate the mass of MO produced.
..... [1]
- (ii) Calculate the number of moles in 0.176 g of carbon dioxide.
[Ar: C, 12; O, 16] g [1]
- (iii) Use the equation
$$MCO_3 \rightarrow MO + CO_2$$

and your answer to (c)(i) to deduce the number of moles of MO produced
..... moles [1]
- (iv) Using your answers to (c)(i) and (c)(iii), calculate the relative formula mass of MO .
..... moles [1]
- (v) Calculate the relative atomic mass of M .
[Ar: O, 16] relative formula mass [1]
- [1]

A compound Q contains 0.69 g of sodium, 1.92 g of sulfur and 1.44 g of oxygen.
[Ar: Na, 23; S, 32; O, 16] Its empirical formula is

- (a) Na_2SO_3 (b) Na_2SO_4 (c) $Na_2S_2O_3$ (d) $Na_3S_2O_3$

June 2015/41

- (a) (i) Describe the appearance of zinc.
..... [1]
- (ii) Zinc oxide can be made from zinc by heating in air.
Construct the equation for this reaction.
..... [1]
- (iii) Which compound may be used to convert zinc oxide into zinc nitrate?
..... [1]
- (b) When zinc nitrate is heated in a fume cupboard the following reaction takes place.
$$2Zn(NO_3)_2(s) \rightarrow 2ZnO(s) + 4NO_2(g) + O_2(g)$$
- (i) Suggest why the heating is done in a fume cupboard.
..... [1]
- (ii) Calculate the number of moles of zinc nitrate in 3.78 g of zinc nitrate.
[Ar: Zn, 65; N, 14; O, 16] moles [1]
- (iii) Using the equation for the reaction and your answer to (b)(ii) calculate the volume of each gas produced when 3.78 g of zinc nitrate is heated.
[1 mole of a gas occupies a volume of 24 000 cm^3 at room temperature and pressure.]

volume of NO_2 cm^3
volume of O_2 cm^3

A student is asked to make 250 cm^3 of 0.20 mol / dm^3 sodium hydroxide.
[Ar: Na, 23; O, 16; H, 1] What is the mass of sodium hydroxide required?

- (a) 1.0 g (b) 2.0 g (c) 4.0 g (d) 8.0 g

- 2 A student uses the apparatus shown to find the energy released by the combustion of each of three alcohols, X, Y and Z. In each experiment 0.1 mol of alcohol is burnt and the initial temperature of the water is 20 °C.

(b) When analysed, alcohol Y is found to contain 60.0% carbon, 13.3% hydrogen and 26.7% oxygen by mass. Its relative formula mass is 60. [Ar: C, 12; H, 1; O, 16]

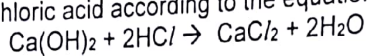
(i) Calculate the empirical formula and hence the molecular formula of Y.

empirical formulamolecular formula

- 4 Water containing a little dilute sulfuric acid is electrolysed using carbon electrodes. If 20 cm³ of oxygen is produced at the positive electrode, what volume of hydrogen is produced at the negative electrode?

(a) 10 cm³ (b) 20 cm³ (c) 30 cm³ (d) 40 cm³

- 7 Calcium hydroxide reacts with hydrochloric acid according to the equation shown.



Which of the following solutions of hydrochloric acid will produce 1.11g of calcium chloride when reacted with excess calcium hydroxide? [Mr: CaCl₂, 111]

(a) 25 cm³ of 0.200 mol / dm³ of hydrochloric acid (b) 50 cm³ of 0.400 mol / dm³ of hydrochloric acid
(c) 75 cm³ of 0.050 mol / dm³ of hydrochloric acid (d) 100 cm³ of 0.100 mol / dm³ of hydrochloric acid

November 2015/41

- 4 Compound Q contains 40.00% carbon, 6.67% hydrogen and 53.33% oxygen by mass. What is the empirical formula of Q? [Ar: C, 12; H, 1; O, 16]

(a) CHO (b) CH₂O (c) CH₂O₂ (d) C₄H₂O

November 2015/42

- 1 A student determines the mass of copper in mixture A which is composed of only copper and zinc. Zinc reacts with dilute sulfuric acid. Copper does not. Dilute sulfuric acid is mixed with A in the apparatus below. A gas is given off which is collected in B.

(a) Name apparatus B.

.....

(b) (i) Name the gas collected in B. Give a test and observation to identify the gas.

name of gas

test and observation

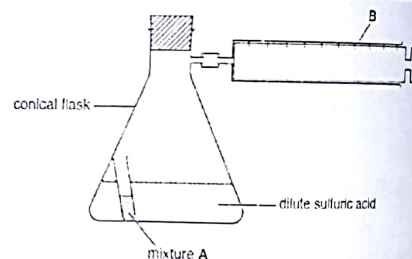
(ii) Construct an equation for the reaction between zinc and dilute sulfuric acid.

(c) When all the zinc has reacted, the volume of gas collected in B is 96.0 cm³ when measured at room temperature and pressure. [1 mole of any gas occupies 24 000 cm³ at room temperature and pressure.]

(i) Calculate the number of moles of gas in 96.0 cm³.

(ii) Using your answers to (b)(ii) and (c)(i) calculate the mass of zinc in A. [Ar: Zn, 65]

(iii) The mass of mixture A is 1.20 g. Calculate the mass of copper in mixture A.

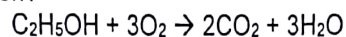


compound, T, contains 41.0 % potassium, 33.7 % sulfur and 25.3 % oxygen by mass. What is the empirical formula of T? [Ar: K, 39; S, 32; O, 16]

- (a) $K_4S_3O_3$ (b) K_2SO_3 (c) K_2SO_4 (d) $K_2S_2O_3$

June 2016/41

What volume of carbon dioxide gas, measured at room temperature and pressure, is produced when 9.2 g of ethanol is burned in an excess of oxygen?



Ar: H, 1; C, 12; O, 16]

[1 mole of a gas occupies a volume of 24 000 cm³ at room temperature and pressure.]

- (a) 240 cm³ (b) 480 cm³ (c) 960 cm³ (d) 9 600 cm³

June 2016/42

- (a) When zinc is heated in air it reacts with oxygen to form an oxide.
A student does an experiment to find the formula of zinc oxide.
Some zinc is placed in a previously weighed crucible and reweighed.

mass of crucible + zinc = 7.04 g
mass of crucible = 5.74 g

- (i) Calculate the mass of zinc used in the experiment. g [1]

The crucible containing the zinc is heated and zinc oxide is produced. The crucible with zinc oxide is weighed.

mass of crucible + zinc oxide = 7.36 g

- (ii) Calculate the mass of zinc oxide produced. g [1]

- (iii) Using your answers to (i) and (ii), calculate the mass of oxygen that combines with the zinc. g [1]

- (iv) Using your answers to (i) and (iii), calculate the formula of zinc oxide. Show all your working.
[Ar : Zn, 65; O, 16]

formula [2]

- (b) When zinc reacts with dilute hydrochloric acid, a gas is produced.
Name the gas. Give a test and observation to identify the gas.

gas
test and observation

W-2016-41

- 3 A student finds that a compound contains 4.8 g of carbon, 0.8 g of hydrogen and 6.4 g of oxygen.
[Ar: C, 12; H, 1; O, 16]

What is the empirical formula of the compound?

- (a) CH_2O (b) C_2H_2O (c) C_2HO_2 (d) $C_4H_8O_4$

W-2016-42

- 3 A student burns 4.8 g of an element X in excess oxygen. The mass of the oxide produced is 8.0 g.
What is the empirical formula of the oxide of X? [Ar: X, 24; O, 16]

- (a) XO (b) X_2O (c) X_2O_2 (d) X_3O_5

Mark Scheme ATP QUESTIONS

MCQs					
Session	Q. No.	Answer Key	Session	Q. No.	Answer Key
November 2010/41	5	B	June 2016/41	4	D
November 2010/42	6	B	November 2015/42	3	D
June 2012/42	7	B	November 2015/41	4	B
November 2013/41	3	C	June 2015/42	7	B
June 2014/41	6	B	June 2015/42	4	D
June 2014/42	4	C	June 2015/41	7	B
November 2014/41	7	A	November 2014/42	4	D
November 2016/41	3	A	November 2016/42	3	A

June 1999

- 3 (a) $22.20 - 15.70 = 6.50 \text{ g}$ 1 (b) Nitric acid 1 (c) Toxic gas, NO_2 is provided 1
 (d) To obtain a constant mass 1 (e) $23.40 - 15.70 = 7.70 \text{ g}$ 1
 (f) M_r of $\text{Zn}(\text{NO}_3)_2 = 81$
 $\text{Zn} : \text{ZnO}$
 $65 : 81$
 $6.50 : \frac{81}{65} \times 6.50 = 8.1 \text{ g (expected) \% of Zn} = \frac{7.70}{8.1} \times 100 = 95\%$

November 1999

- 2 (a) (i) BaSO_4 (ii) White ppt (iii) Filtration 1
 (b) Moles of $\text{Ba}(\text{NO}_3)_2 = 1.00 \times \frac{100}{1000} = 0.1 \text{ mol}$
 Mole of $\text{Na}_2\text{SO}_4 = 2.00 \times \frac{100}{1000} = 0.2 \text{ mol}$
 Mol ratio : $\text{Ba}(\text{NO}_3)_2 : \text{Na}_2\text{SO}_4 = 1 : 1$
 But Na_2SO_4 is more than expectation
 (c) (i) M_r of $\text{BaSO}_4 = 56 + 32 + (16 \times 4) = 152$ (ii) $\frac{12.5}{15.2} \times 100 = 82.2\%$
 Mass of $\text{BaSO}_4 = 152 \times 0.100 = 15.2 \text{ g}$

June 2000

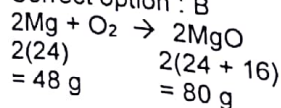
5. Correct option : A
 Mol ratio $\text{BaCl}_2 : \text{Na}_2\text{SO}_4$
 $1 : 1$
 $0.2 : 0.2 \text{ (expected)}$
 0.4 (available)
 Na_2SO_4 is excess, BaCl_2 is limiting
 Mol ratio $\text{BaCl}_2 : \text{BaSO}_4 = 1 : 1$ Moles of $\text{BaSO}_4 = 0.2 \text{ mol}$

NOVEMBER 2000

- 3 (a) (i) White 1 (ii) Filtration (b) (i) $2.0 \times \frac{20}{1000} = 0.04 \text{ mol}$
 (ii) $1.0 \times \frac{50}{1000} = 0.05 \text{ mol}$
 (c) mol ratio $\text{AgNO}_3 : \text{NaCl}$
 $1 : 1$
 $0.04 : 0.04 \text{ (expected)}$ 0.05 (available)
 NaCl is in excess
 Mol ratio $\text{AgNO}_3 : \text{AgCl} = 1 : 1$
 So 0.04 will produce 0.04 will of AgCl
 M_r of $\text{AgCl} = 0.04 \times 143.5 = 5.74 \text{ g}$
 (d) Now NaCl is limiting
 Mol ratio $\text{NaCl} : \text{AgCl} = 1 : 1$

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Mass of AgCl = $0.04 \times 143.5 = 5.74\text{g}$
Correct option : B



$$6.0\text{ g} \quad \frac{80}{48} \times 6.0 = 10.0\text{ g}$$

NOVEMBER 2001

- 2 (a) Blue (b) $9.25 - 5.40 = 3.85\text{ g}$ (c) White
(d) (i) $7.90 - 5.40 = 2.50\text{ g}$ (ii) $3.85 - 2.50 = 1.35\text{ g}$
(e) (i) $64 + 32 + (16 \times 4) = 160$ (ii) $(1 \times 2) + 16 = 18$
(f) (i) $\frac{2.50}{160} = 0.0156$ (ii) $\frac{1.35}{18} = 0.075$ (g) $\frac{0.075}{0.0156} = 4.81 \approx \text{CuSO}_4 \cdot 5\text{H}_2\text{O} (1)$

November 2002

2. (a) (i) yellow precipitate (ii) filtration
(b) (i) silver nitrate
Moles of KI = $1.0 \times \frac{50}{1000} = 0.05\text{ mol}$
Mol ratio KI : AgNO₃ is 1 : 1
Moles of AgNO₃ = $2.0 \times \frac{30}{1000} = 0.06\text{ mol}$
(c) (i) Mr of Ag I = $108 + 127 = 235$
white precipitate
Mass of Ag I = $235 \times 0.05 = 11.75\text{ g}$
(ii) Darkens on standing
mass of Ag I is more than that of AgCl
Molar mass of Ag I is greater than that of AgCl

JUNE 2003

- 3 (a) $14.70 - 8.90 = 5.80\text{ g}$ (b) Nitric acid
(iii) $14.40 - 8.90 = 5.50\text{ g}$ (d) $\frac{5.50}{5.80} \times 100 = 94.8\% \approx 95\%$ (e) small amount of powder was lost.
(c) (i) NO₂, a toxic gas is produced
(ii) to ensure the decomposition is complete

November 2006

Correct option : A
JUNE 2007

- 2 (a) Test : pass the gas through lime water
(b) Observation : Lime water goes from colourless to milky
(i) effervescence stopped
(ii) solid remained (answers must be different in (i) and (ii)) (no precipitates or solubility in answers) [1]
(c) filtration (d) (iii) (e) $0.25 \times \frac{100}{1000} = 0.025\text{ moles}$
(f) mole ratio = H₂SO₄ : ZnSO₄ = 1 : 1
Mr of ZnSO₄ = 161 mass of ZnSO₄ = $0.25 \times 161 = 4.025\text{ g}$
(g) mole of H₂SO₄ : Vol. of CO₂
1 mol : 24 dm³
0.025 mol : $24 \times 0.025 = 0.6\text{ dm}^3$ [1]

[1]

NOVEMBER 2007

- 4 (a) white precipitate (1). (b) (i) $0.20 \times \frac{100}{1000} = 0.02\text{ moles}$ (ii) $0.25 \times \frac{60}{1000} = 0.015\text{ moles}$
(c) Mol ratio H₂SO₄ : Ba(NO₃)₂
1 : 1
0.02 : 0.02 (expected)
0.015 (available)
Mol ratio Ba(NO₃)₂ : BaSO₄
1 : 1
Mr of BaSO₄ = $137 + 32 + (16 \times 4) = 233$
Ba(NO₃)₂ is limiting reactant
(d) $\frac{3.35}{3.495} = 95.85\%$ (e) product was not dried etc (1) (f) barium chloride, carbonate, bromide, or iodide (1)
Mass of BaSO₄ = $233 \times 0.015 = 3.495\text{ g}$
(not phosphate);

[7]

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JUNE 2008

- 2 (a) $9.87 - 5.83 = 4.04 \text{ g}$ (b) white (1) powder (1)
 (c) (i) $7.35 - 5.83 = 1.52 \text{ g}$ (ii) $4.04 - 1.52 = 2.52 \text{ g}$
 (d) (i) $(23 \times 2) + 12 + (16 \times 3) = 106$ (ii) $(1 \times 2) + 16 = 18$
 (e) (i) $\frac{1.52}{106} = 0.0143$ (ii) $\frac{2.52}{18} = 0.140$ (f) $\frac{0.140}{0.0143} = 9.8 \approx 10 \times 10$

8. Correct Option : B
 CaCO_3 : HCl
 100 g : 2 mol
 2 g : $\frac{2}{100} \times 2 = 0.04 \text{ mol}$

$$C = \frac{n}{V} \times 1000$$

$$1.0 = \frac{0.04 \times 1000}{V}$$

$$V = 40 \text{ cm}^3$$

- 11 (a) $\text{Pb}(\text{NO}_3)_2 + 2\text{KI} \rightarrow \text{PbI}_2 + 2\text{KNO}_3$ correct formulae (1) balanced (1)
 (b) 8.0 (+/- 0.5), 16., 24.0, 28.0, 28.0 all correct (2), one error (1), more than 1 error, 0 marks.
 (c) All (candidate's) points correctly plotted (1) two intersecting lines (2)
 (Points joined by a curve, 1 mark, points joined by a series of st. lines 0 marks)
 (d) (i) 7.0 cm^3 (1)
 (ii) moles $\text{Pb}(\text{NO}_3)_2$: moles $\text{KI} = 1:2$
 concentration of $\text{KI} = 3.5 \text{ mol/dm}^3$ Correct answer with working (2)
 (evidence of some correct working (1)) Mole ratio of 1:1 gives 1.75 mol/dm^3
 (e) (i) $3.5 (1) \text{ cm}^3$ (half of answer (d) (i)) (ii) $28 \text{ mm} (1)$ [12]

June, 2009

- 1 (a) $13.08 - 12.72 = 0.36 \text{ g}$ [1]
 (b) (i) to prevent product escaping (ii) to allow air (oxygen) to enter crucible [1]
 (c) (i) shiny, silver or grey (ii) white/yellow/grey powder [1]
 (d) to ensure that all Mg had reacted, etc.
 (e) (i) $13.32 - 12.72 = 0.6 \text{ g}$ (ii) $0.60 - 0.36 = 0.24 \text{ g}$ [1]
 (f) Mg : O
 $\frac{0.36}{24} : \frac{2.24}{16}$
 $0.015 : 0.015$
 1 : 1

November 2010/41

6. C H O
 $\frac{40.0}{12} = 3.33$ $\frac{6.7}{1} = 6.7$ $\frac{53.3}{16} = 3.33$
 $\frac{3.33}{3.33} = 1$ $\frac{6.7}{3.33} = 2$ $\frac{3.33}{3.33} = 1$
 CH_2O

Correct Option : C

November 2010/42

- 1 (a) (iii) $0.100 \times \frac{44}{1000} = 0.0044 \text{ moles}$ (b) (i) $\frac{0.12}{24} = 0.005 \text{ moles}$
 (ii) $\text{Mg} + \text{H}_2\text{SO}_4 \rightarrow \text{MgSO}_4 + \text{H}_2$
 (c) (ii) Mol ratio = Mg : $\text{H}_2\text{SO}_4 = 1 : 1$
 Mg is in excess : 0.044 g of Mg was expected but available is 0.05 mol
 $0.0044 \times 24 = 0.106 \text{ dm}^3 (1)$

November 2011/41

- 3 (a) $0.48 (1) \text{ g}$ (b) (i) silver/grey/shiny metal/solid (1) (ii) white solid/powder (1)
 (c) to ensure constant weight or that reaction was complete (1)
 (d) (i) $0.8 (1) \text{ g}$ (ii) $0.32 (1) \text{ g}$ (e) $0.48/24 = 0.02$ $0.32/16 = 0.02 (1)$ $\text{MgO} (1)$
 (f) (i) $\text{MgO} + 2\text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2\text{O} (1)$ or $+ \text{H}_2\text{SO}_4 + \text{HNO}_3$ (ii) basic (1) [10]

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- 4 (a) syringe (1) (b) lime water turns milky (1) (c) (i) 0.005 (1) (ii) 0.01 (1)
 (iii) $0.01 \times 84 = 0.84 \text{ g}$ (1) [5]

June 2012/41

- 2 (a) brown/red-brown/pink/orange + metal/solid (1) (b) 1.27 (1) g (c) 1.59 (1) g
 (d) 0.32 (1) g (e) $1.27/64 = 0.02$; $0.32/16 = 0.02$ (1) $1:1 = \text{CuO}$ (1) [6]

June 2012/42

- 3 (a) (i) white ppt./solid (1) (ii) filter precipitate (1), wash with water (1) dry the solid (1)
 (b) (i) 0.075 (1) (ii) 0.1 (1) (iii) 0.075 (1) (iv) $233 (1) \times 0.075 = 17.48 \text{ g}$ (1)

November 2012/41

4. (d) 0.05 (1) (e) (i) molar mass = 161 (1); mass = $161 \times 0.05 = 8.05 \text{ g}$ (1)

November 2012/42

1. (d) 0.02 (1) moles (e) 0.02 (1) moles (f) 0.80 (1) moles (ecf for (e) and (f) from (d))

June 2013/41

- 1 (c) 66 cm^3 (d) $\frac{66}{24000} = 0.00275 \text{ moles}$
 (e) (i) moles ratio = $\text{NaHCO}_3 = 1:2$
 Moles (NaHCO_3) = $0.00275 \times 2 = 0.0055 \text{ moles}$
 (ii) $M_r = 23 + 1 + 12 + (16 \times 3) = 84$ (iii) $84 \times 0.0055 = 0.462 \text{ g}$

6. % of H = $100 - 88.9 = 11.1$

C : $\frac{88.9}{12} = 7.41$ H = $\frac{11.1}{1} = 11.1$

C = $7.41 / 7.41$ H = $\frac{11.1}{7.41} = 1.49$



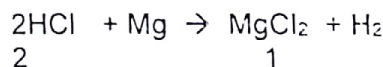
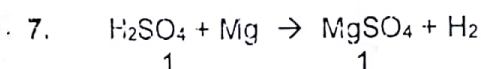
Correct option C

7. Moles (NaSO_4) = $\frac{1.42}{142} = 0.01 \text{ mol}$

Correct option C

June 2013/42

- 1 (a) green (1) (b) $(9.01 - 5.97) = 3.04 \text{ (g)}$ (c) (i) $7.66 - 5.97 = 1.69 \text{ (g)}$
 (ii) $3.04 - 1.69 = 1.35 \text{ (g)}$ (iii) $\frac{1.69}{152} = 0.011 \text{ (moles)}$ (iv) $\frac{1.35}{18} = 0.075 \text{ (moles)}$
 (d) (i) $\frac{0.075}{0.0111} = 6.76 \text{ moles}$ (ii) $x = 7$



	Moles of acid	Moles of H_2
Moles of H_2SO_4	$0.200 \times \frac{10.0}{1000} = 0.002$	0.002

A	$0.200 \times \frac{20}{1000} = 0.004$	0.002
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B	$0.100 \times \frac{20}{1000} = 0.002$	0.002
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C	$0.200 \times \frac{40}{1000} = 0.008$	0.004
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D	$0.050 \times \frac{40}{1000} = 0.002$	0.002
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Correct option C

November 2013/4

- 1 (a) (i) conical flask (ii) gas syringe/syringe (b) (i) 0.002 (ii) 0.048 (g)

November 2013/42

- 2 (a) nitric (acid), HNO_3 (1) (both) (c) (i) molar mass of $\text{NH}_4\text{NO}_3 = 80$ (1)
 $28/80 \times 1000 = 350$ (g) (1) [2]
 (ii) $(350/28) \times 24 = 300$ (dm³) (1)

June 2014/41

- 1 (a) (i) $2\text{Cu NO}_2 \rightarrow 2\text{CuO}$ (ii) black (iii) 18 (1) cm³
 (b) (i) 72 (1) cm³ (ii) nitrogen (c) 300 (1) cm³
 (iv) 0.00075 (1) moles (v) 0.096 (1) g

June 2014/42

2. (a) (i) off white or cream or pale yellow (ii) filtration / filter (b) (i) 0.05 (1) moles
 (c) $0.05 (1) \times 188 = 9.4$ g (d) $0.03 (1) \times 188 = 5.64$ g (ii) 0.06 (1) moles

November 2014/41

- 3 (a) tripod (b) heat to constant mass (c) (i) 0.45 g (ii) 106, 18 (iii) 0.0025, 0.025
 (d) 10

November 2014/42

- 3 (a) limewater turns milky (b) heat to constant mass (c) (i) 0.16 (1) g (ii) 0.004 (1) moles
 (iii) 0.004 (1) moles (iv) 40 (v) $(40 - 16) = 24$

June 2015/41

- 1 (a) (i) Silver / grey solid (ii) $2\text{Zn} + \text{O}_2 = 2\text{ZnO}$ (iii) Nitric acid / HNO_3
 (b) (i) Toxic / poisonous gas evolved (ii) $3.78 / 189 = 0.02$ (iii) $0.02 \times 24000 \times 2 = 960$ cm³
 $0.02 \times 24000 \times 0.5 = 240$ cm³

June 2015/42

2. (b) (i) $(60 / 12 = 5 \quad 13.3 / 1 = 13.3 \quad 26.7 / 16 = 1.67)$
 $\quad \quad \quad 3 \quad \quad \quad 3 \quad \quad \quad 1$
 Empirical Formula = $\text{C}_5\text{H}_{13}\text{O}$ (1) Reject $\text{C}_5\text{H}_7\text{OH}$
 Molecular formula = $\text{C}_5\text{H}_{13}\text{O}$ (1)

November 2015/42

- 1 (a) (gas) syringe (ii) $\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 + \text{H}_2$
 (b) (i) hydrogen
 burning splint pops or pops in a flame (iii) 0.94 (g)
 (c) (i) 0.004 (mol) (ii) 0.26 (g)

June 2015/42

- 3 (a) (i) 1.3(0) g 1 (ii) 1.62 g 1 (iii) 0.32 g 1 (iv) M1 1.30 / 65 and 0.32 / 16
 or 0.02 and 0.02
 or both 1 / 50 (1)
 M2 ZnO (1)
 (b) Hydrogen
 Pops in a flame / lighted splint pops / burning splint pops