Introduction

The purpose of this book is to provide relevant material for each subject in O-level education here in Tanzania. The first edition contains civics, history, geography, biology, chemistry and physics. The content is ordered by syllabus topic and contains relevant definitions and solved problems as they have appeared on NECTA examinations. Though it is impossible to predict NECTA topics and questions, I feel that a student who knows all of the information provided here can get a B in the subject if they are also able to understand English and have competency in the subject material. This is not meant to be a primary resource, but rather it is intended to help guide students and teachers towards relevant topics and questions for study and discussion. This book is for students taking form 4 examinations. Some form 1 and 2 topics are not covered, since they have not appeared on the examinations.

The expectation of this book is that it will provide a base of knowledge that each student will have by the time they come to take their national examinations. In class and in further study, topics and questions can be expanded upon to provide the student with the competency he requires to be successful on his national examinations. Students are encouraged to look at future topics before they are taught in class, so that the teacher can spend class time explaining difficult material, rather than writing definitions or notes on the board.

This work could not have been done without the help of my fellow teachers and staff here at Abbey Secondary School. I am grateful for their contributions to this project. I hope that each year we can update and improve these study guides so that our school can continue to grow academically.

- Jeff Rodwell
jeff.rodwell@gmail.com

Legal Statement

Additional Credits

Civics - Juma Seif
History - Ramadhani Mndeme
Geography - Field JK Osera
Biology - Gastone Ndunguru
Chemistry - Gastone Ndunguru

Chemistry

Reference Materials

0.1.0 Equations and oxidation numbers
0.2.0 Periodic table

Form 1

1.1.0 Introduction to chemistry
  1.1.1 The concept of chemistry
  1.1.2 Importance of chemistry in life

1.2.0 Lab techniques and safety
  1.2.1 Rules and safety precautions in a chemistry lab
  1.2.2 First aid and first aid kit
  1.2.3 Basic chemistry laboratory apparatus and their uses
  1.2.4 Warning signs

1.3.1 Heat sources
  1.3.2 Types of flames

1.4.0 The scientific procedure
  1.4.1 Significance of the scientific procedure
  1.4.2 The main steps of the scientific procedure
  1.4.3 Application of the scientific procedure

1.5.0 Matter
  1.5.1 The concept of matter
  1.5.2 States of matter
  1.5.3 Physical and chemical changes
  1.5.4 Elements and symbols
  1.5.5 Compounds and mixtures
  1.5.6 Separation of mixtures

1.6.0 Air, combustion, rusting and fire fighting
  1.6.1 Composition of air
  1.6.2 Combustion
  1.6.3 Fire fighting
  1.6.4 Rusting

Form 2

2.1.0 Oxygen
  2.1.1 Preparation and properties of oxygen
  2.1.2 Uses of oxygen

2.2.0 Hydrogen
  2.2.1 Preparation and properties of hydrogen
  2.2.2 Uses of hydrogen

2.3.0 Water
  2.3.1 Occurrence and nature of water
  2.3.2 Properties of water
  2.3.3 Treatment and purification of water
  2.3.4 Uses of water

2.4.0 Fuels and energy
  2.4.1 Fuels sources
  2.4.2 Categories of fuels
  2.4.3 Uses of fuel
  2.4.4 Conservation of energy
  2.4.5 Renewable energy (biogas)

2.5.0 Atomic structure
  2.5.1 The atom
  2.5.2 Subatomic particles
  2.5.3 Electronic arrangements
  2.5.4 Atomic number, mass number and isotopy

2.6.0 Periodic classification
  2.6.1 Periodicity
  2.6.2 General trends

2.7.0 Formula bonding and nomenclature
  2.7.1 Valence and chemical formulae
  2.7.2 Oxidation state
  2.7.3 Radicals
  2.7.4 Covalent bonding
  2.7.5 Electrovalent bonding

Form 3

3.1.0 Chemical equations
  3.1.1 Molecular equations
  3.1.2 Ionic equations
  3.1.3 Mass problems

3.2.0 Hardness of water
  3.2.1 The concept of hardness of water
  3.2.2 Types of hardness of water
  3.2.3 Treatment and purification of hard water

3.3.0 Acids, bases and salts
  3.3.1 Acids and bases
  3.3.2 Indicators
  3.3.3 Salts

3.4.0 The mole concept and related calculations
  3.4.1 The mole as a unit of measurement
  3.4.2 Application of the mole concept

3.5.0 Volumetric analysis
  3.5.1 Standard volumetric apparatus
  3.5.2 Standard solutions
  3.5.3 Volumetric calculations
  3.5.4 Application of volumetric analysis

3.6.0 Ionic theory and electrolysis
  3.6.1 Ionic theory
  3.6.2 The mechanism of electrolysis
  3.6.3 Laws of electrolysis
  3.6.4 Application of electrolysis

3.7.0 Chemical kinetics, equilibrium and energetics
  3.7.1 The rate of chemical reactions
  3.7.2 Factors affecting the rate of chemical reactions

0.1.0 Equations and Oxidation Numbers

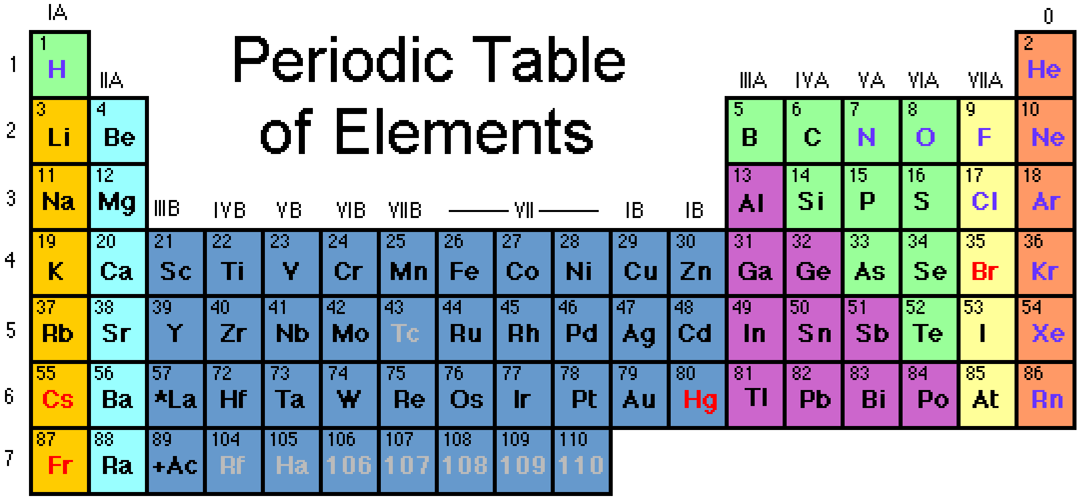
Avogadro's Constant = 6.02×10^{23}
1 Faraday (F) = 96500 Coulombs (C)
Gas molar volume (GMV) at STP = $22.4dm^3$
Atomic masses: H = 1, C = 12, N = 14, O = 16, Na = 23, Mg = 24, S = 32, Cl = 35.5, Ca = 40, Mn = 55, Fe = 56, Cu = 63.5, Zn = 65, Ag = 108, I = 127, Pb = 207

Name	Formula	Oxidation State
Ammonia	NH_4^+	+1
Sulphate	SO_4^{2-}	-2
Carbonate	CO_3^{2-}	-2
Hydroxide	OH^-	-1
Nitrate	NO_3^-	-1
Phosphate	PO_4^{3-}	-3
Chlorate	ClO_3^-	-1
Hydrogen Sulphate	HSO_4	-1

Molarity Equation $Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$	**Concentration Equation** $Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$
Dilution Law $\frac{M_A V_A}{M_B V_B} = \frac{N_A}{N_B}$	**Quantity of Electricity Equation** $Quantity\ of\ \ Electricity(C) = Current(I) \times Time(t)$ $Q = IT$
Moles Equation	**Relative Molecular Mass Equation**

$$Moles = \frac{Mass}{Molar\ Mass}$$	$$Relative\ Molecular\ Mass = 2 \times Vapour\ Density$$
Electrochemical Equivalent Equation $$m = ZQ$$	

0.2.0 Periodic Table



Periodic Table of Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 106	107 107	108 108	109 109	110 110								

Form 1 Topics

1.5.0 Matter

1.5.2 States of matter

Condensation - The change in the phase of matter from gaseous to liquid droplets

Dry Ice - Solid form of carbon dioxide (CO_2) which sublimes into the air

Evaporation - Causes the vapourization of a liquid, but occurs only on the surface of a liquid

Freezing - The process of water or any other liquid changing from a liquid to a solid

Froth - Foam consisting of bubbles on a liquid

Melting - The process of a substance changing from a solid to a liquid

Sublimation - The change of a substance from a solid to a gas without becoming a liquid (ex. Ammonia salts sublime when heated)

Avogadro's Law - States that equal volumes of all gases at the same temperature and pressure contain the same number of molecules

Ideal Gas - A theoretical gas composed of a set of randomly moving particles which obeys they ideal gas law

Ideal Gas Law - The equation of a state of a hypothetical ideal gas which approximates the behaviour of many gases under varying conditions combining Boyle's Law and Charles's Law

Boyle's Law - Describes the inversely proportional relationship between pressure and volume of a gas (as volume increases, pressure decreases)

Charles's Law - A law which describes out gases tend to expand when heated, showing the direct relationship between temperature and volume (as temperature increases, volume increases)

Kinetic Theory of Gases - Explains the behaviour of gases based on the movement of their molecules

Avagadro's Hypothesis - Requires that equal volumes of all ideal gases have the same number of molecules at STP

Differentiate between solids, liquids and gases -

Liquids - Particles are loosely packed and are bound to each other with weaker forces than those of solids giving them definite volume but not definite shape. This causes liquids to take the shape of the container in which they are placed

Solids - Particles are closely packed and bound by strong inter-particle attraction which makes solids rigid and geometrical giving them definite shape and volume

- **Gases -** The particles are separated by greater distances and forces of attraction are virtually non-existent which results in particles which are free to move in any direction. This causes gases to posses neither a definite volume nor shape and they occupy the whole volume of the vessel in which they are placed

1.5.3 Physical and chemical changes

Types of changes - Physical, chemical

- **Physical Change -** Any change that does not involve a chemical change of the substance
  - **Examples of physical change -** Paper burning, glass breaking, changing of a liquid to a solid, dissolving of salt in water, heating iodine
- **Chemical Change -** Bonds are broken and new bonds are formed in a compound
  - **Examples of chemical change -** Rusting of iron, sodium metal reacting with water, heating magnesium in air or heating ammonium chloride, adding sulphuric acid to water, ripening of fruits, burning paper

Why rusting of iron is a chemical change - Rusted iron cannot be converted back to its original form, heat is produced during rust formation

Rusting is best prevented <u>by painting</u>

1.5.5 Compounds and mixtures

Compound - Substances formed by the combination of two or more elements in a way that causes them to change chemically. The smallest particle of a compound is a molecule

Molecule - The smallest amount of a substance which can exist independently which consists of two or more atoms combined together (ex. Cl_2)

Chromatography - Used to describe the laboratory techniques used for the separation of mixtures

Decanting - A process used for the separation of mixtures by pouring a solution from a container in order to leave behind precipitate at the bottom of the original container

Distillation - The process of separating mixtures based on different boiling points. It is a physical process which involves evaporation and condensation

Filtration - A process used for the separation of solids from liquids by passing them through a medium in which only liquid can pass, leaving behind the solid material

Destructive Distillation - A set of chemical reactions that cracks or breaks down macromolecules into smaller parts (ex destructive distillation of wood makes charcoal)

Fractional Distillation - The process used in the separation of liquids based on different boiling points

Solvent Extraction - The method used to separate two solvents

Crystallization - The process of the formation of solid crystals from precipitate in a solution

Dehydration - Removal of water from a compound

Desiccation - The state of extreme dryness or the process of extreme drying surrounding environment. Desiccants are hygroscopic substances

Desiccators - Keeps substances constantly dry

Effervescence - The escape of gas from an aqueous solution and the foaming or fizzing which results from the release of this gas

Efflorescent - The loss of water or a solvent of crystallization from a hydrated or solvated salt to the atmosphere on exposure to air

Hygroscopic Substance - One which absorbs water from the atmosphere when exposed to air without forming a solution

Precipitation - The formation of a solid in a solution

Saturated Solution - One which cannot dissolve any more solute at a particular temperature

Saturation Point - The point when a solution cannot dissolve any more of a substance and the excess will appear as precipitate

Suspension - A heterogeneous fluid containing solid particles that have not yet fallen to the bottom as precipitate

Miscibility - Liquids ability to mix with each other forming a homologous solution

- **Miscible Liquids -** Pairs of liquids which mix together uniformly in any proportion (ex. water and alcohol)
- **Immiscible Liquids -** Pairs of liquids which do not mix completely and form two separate layers (ex. water and oils)

Solubility - Is the ability of a solute to dissolve in a liquid solvent to form a homogeneous solution of the solute in the solvent

- **Solute -** A substance which dissolves in a solvent (e.g. common salt, NaCl is a solute, water is a solvent)
- **Solvent -** A substance which dissolves a solute (e.g. water)

Mixture - A material system of two or more different substances which are mixed together but not combined chemically

- **Heterogeneous Mixture -** A mixture where the composition is not uniform
- **Homogeneous Mixture -** A mixture where the composition is uniform
- **Why air is a mixture -** Can be separated by physical means, composition is variable, no energy is given out when air is formed, its properties are the average of its constituent gases

1.6.0 Air, combustion rusting and fire fighting

1.6.1 Fire basics

Kindling Point - The temperature at which a fuel will catch fire

Requirements for a fire - Presence of combustible material, presence of oxygen, material is at or above the kindling point

Products of combustion - Heat, light, CO_2, H_2O, sometimes carbon (soot)

Classes of Fires - A, B, C, D, E

- **A -** A fire which is produced by burning solids (ex. wood, coal, paper, furniture, plastics, clothes)
  - **How to put out a class A fire -** Water
- **B -** A fire which is produced by burning a flammable liquid (ex. petrol, kerosene, alcohol, fats)
  - **How to put out a class B fire -** Fire blanket, sand, foam, carbon dioxide
- **C -** A fire which is produced by burning a liquefied gas (ex. acetylene, hydrogen gas, coal gas)
  - **How to put out a class C fire -** Foam or carbon dioxide extinguisher
- **D -** A fire which is produced by burning an alkali metal (ex sodium, potassium)
- **E -** A fire which is produced in electrical equipment
  - **How to put out a class E fire -** Unplug the device, use carbon dioxide extinguisher, never use water because water increases electrical conductivity and can cause a larger fire

Why water is not used to put out oil (class B) fires - Water is denser than oil, so it will sink below the oil and the oil will continue to burn

Function of a damp blanket to put out a fire on your clothes - The damp blanket will limit the oxygen which is necessary for combustion to occur. This will also lead to an accumulation of CO_2 gas which is not used in combustion

Form 2 Topics

2.2.0 Hydrogen/preparation/properties

2.2.1 Preparation of hydrogen gas

Methods of collecting gases in the laboratory - Downward delivery, downward displacement of water, upward delivery

- **Downward delivery -**
  - **Used to collect -** Chlorine, chloride gas
    - **Chlorine -** This method is best because chlorine is denser than air and will sink into the collection container. Since chlorine is soluble in water you cannot use displacement
- **Downward displacement of water -**
  - **Used to collect -** Hydrogen
    - **Hydrogen -** Since hydrogen gas is less dense than water, it will rise up into the collection container. Water is also useful because hydrogen gas is colourless and can be seen as bubbles in the water
- **Upward delivery -**
  - **Used to collect -** Ammonia
    - **Ammonia -** This method is best because ammonia has a lower density than air which means that it will rise up into the collection container. Since it is very soluble in water you cannot use displacement

2.2.2 Uses of hydrogen gas

Uses of hydrogen gas - Used in fuels, conversion of liquid oils to solid oils, making petrol, used in the Haber process to manufacture ammonia

2.3.0 Water

2.3.1 Properties of water

Properties of water - Effective solvent, high specific heat capacity, liquid on earth surface, expands on freezing, high surface tension, neutral pH

2.4.0 Fuels

2.4.1 Fuel sources

Fuel - Is a substance that produces energy when burnt

Examples of liquid fuel - Petrol, diesel

Advantages of liquid fuel over solid fuel -

Examples of gaseous fuel - Methane, propane, coal gas, water, gas

Gaseous fuels used in industry - Coal gas (hydrogen, methane, ethane, CO), liquid gas (hydrogen, CO)

Examples of solid fuel - Charcoal, coal

Characteristics of good fuel - Naturally occurs in large amount, non-toxic, easy to ignite, widely used, high heat output, does not produce too much waste when burnt

Examples of Fossil Fuels - Oil, coal, natural gas

2.5.0 Atomic Structure

2.5.1 The atom

Dalton's atomic theory -

- All elements are made up of tiny indivisible particles called atoms
- Atoms cannot be created, destroyed or sub divided (and are the smallest particles which can undergo a chemical change)
- Atoms of different elements have different properties and different masses
- When atoms combine to form a molecule they do so in simple whole numbers
- All molecules of one compound are identical and are different from the molecules of all other compounds

2.5.2 Subatomic particles

Electron - A particle which carries a negative charge, it is smaller than protons and neutrons

Proton - Positively charged particle of an atom which has a mass equal to that of a hydrogen atom

Neutron - A subatomic particle with no electric charge and the same mass as a proton

Radioactivity - The process by which an unstable nucleus of an atom loses energy by emitting ionizing particles or radiation

Element - A substance which cannot be split into two or more simpler substances by any chemical means (ex. carbon, oxygen)

Remember: *All chemical reactions involve moving electrons from one element to another. Protons and neutrons never transfer (this only happens in nuclear reactions, not chemical reactions)*

2.5.3 Electronic arrangements

1. What is the electronic arrangement for an element with an atomic number of 14?

Step 1: Fill the first shell. *The first shell has 2 spaces* 2 filled, 12 remaining **Step 2: Fill the second shell.** *If there are no remaining atoms then you are finished.* *The second shell has 8 spaces* 2:8 filled, 4 remaining	**Step 3: Fill the last shell with the remaining atoms.** *If there are no remaining atoms then you are finished. The last shell has 8 spaces* *2:8:4, 0 remaining, therefore the electron arrangement is* ***2:8:4***

2.5.4 Isotope, atomic mass and atomic number

Differentiate between mass number and atomic number -

Mass Number - The sum of the number of protons and neutrons in a nucleus

Atomic Number - The number of protons in the nucleus

Isotopes - Atoms of the same element (same number of protons) and atomic number but a different number of neutrons

1. Boron is known to have two stable isotopes of atomic masses 10.016g and 11.013g. These isotopes normally exist in nature in proportions of 20% and 80% respectively. From this data, what is the average atomic mass of boron?

The average mass of boron is the sum of the proportions of each isotope

Step 1: Calculate the proportions

Proportion of isotope 1 = 10.016 * 0.2 = 2.0032

Proportion of isotope 2 = 11.013 * 0.8 = 8.8104

Step 2: Add the proportions together

Average = Proportion of isotope 1 + proportion of isotope 2

Average = (10.016 * 0.2) + (11.013 * 0.8)

Average = 2.0032 + 8.8104

Average = **10.81g**

2. ^{204}J, ^{206}K, ^{207}L and ^{A}M are isotopes of element H whose abundances are 2%, 24%, 22% and X% respectively. Calculate the abundance of X% and the mass number A of isotope M given that the relative atomic mass of element H is 207

This question is asking for two things. First is the % abundance of ^{A}M and second for the mass number of A. To find the % abundance of ^{A}M you simply calculate the remaining percent by subtracting the percents given from 100%:

100% - 2% - 24% - 22% = 52%

So the percent abundance is 52%

The following equation represents how the mass number of an element is the average of all isotopes

$$\frac{(204\times 2)+(206\times 2)+(207\times 22)+(A\times 52)}{100}=207$$

$$\frac{9906+52A}{100}=207$$

$$9906+52A=20700$$

$$52A=17904$$

$$A=208$$

*So the mass number is **208***

2.6.0 Periodic Classification

2.6.1 Periodicity and general trends

Periodicity - Refers to the pattern seen in the periodic table resulting from the chemical properties of each element and how they change going down or to the right of the table

Ionization Energy - Is the minimum amount of energy required to remove one mole of electrons from one mole of gaseous atoms or molecules. It decreases down the group in the periodic table

Ionization - The process of gaining or losing electrons from a neutral atom or molecule

Alkali Earth Metals - Elements found in group II in the periodic table, they are soft compared to other metals, have low melting points and are so reactive that they are never found in nature uncombined with other elements. They are powerful reducing agents (they give up an electron easily)

Examples of alkali earth metals - Potassium (K), sodium (Na), lithium (Li)

Halogens - A series of nonmetal elements (group VII and VIIA), which contain all three states of matter at standard temperature and pressure and are highly reactive. Halogens include fluorine, chlorine, bromine, and iodine

Noble Gases - A group of elements which are odorless, colourless and do not react easily with other substances. This is due to their outer shell being full, so electrons are not shared

2.6.2 Periodicity questions

1. (a) T and K are elements found in the periodic table. The atomic number of T is 16 and K is 19.

(i) In which group and period of the periodic table does the element T appear?

The period and group of an element are determined by filling their valency shells with electrons. The first shell can hold 2 electrons and the 2nd one holds 8, the 3rd one holds 8. The remainder represents which group (column) the element belongs to. The period represents how many shells the atom has.

Since T has 16 electrons, we start by filling the first two valency shells with the remainder going into the third shell → 2:8:6

Since the remainder is 6, the element T is found in group 6. We filled electrons up until the 3rd shell, so this means it is in the 3rd period. The 2, 8 and 6 represent the three shells, with the 1st and 2nd ones being full and the remainder in the outer shell.

Group of T is 6 and the period is 3

(ii) Is element T a metal or non-metal?
It is a non metal because it is found in group 6.
(iii) Write a molecular formula of a compound formed between T and K

To find this we must find the group of element K. Element K fills its valency shells like 2:8:8:1, therefore it is in group 1. Group 1 has a +1 charge, while element T which is in group 6 has a -2 charge. Therefore it requires two element K's per element T.

Molecular formula: K_2T

(b) (i) Which particles are atoms of the same element in the list given? $^{40}_{18}A, ^{38}_{20}B, ^{38}_{18}C, ^{40}_{19}D$

$^{40}_{18}A, ^{38}_{18}C$ *because they both have an atomic number of 18. The bottom number represents the atomic number (number of protons) and the top number represents its mass (neutrons + protons). A and C are isotopes of the same element*
(ii) Give the electronic configuration of sodium and neon if the atomic number of sodium is 11 and neon is 10
Sodium - 2:8:1
Neon - 2:8
(iii) Why can't neon react with sodium?
This is because neon's outer shell is full making it a noble gas, so it cannot form covalent bonds. The first shell can hold 2 electrons, the 2nd holds 8 and the 3rd holds 8. All noble gases have their outer shells full.

2.7.0 Formula bonding and nomenclature

2.7.1 Valence and chemical formulae

Chemical Equation - Representation of reactants and products in a chemical reaction
Empirical Formula - A formula of a compound which shows the simplest whole number ratio between each element in a compound (ex. H_2O)
Molecular Formula - Is a formula which shows the actual number of atoms present in a compound
Valency - The ability of an atom to combine with other atoms. It represents the number of electrons an atom can gain, lose or share by forming bonds with another element
Electronegativity - The ability of an atom to attract electrons towards it
Ion - An atom or molecule where the total number of electrons is not equal to the total number of protons resulting in a positive or negative electrical charge
Ionic Compound - A type of compound formed by two or more atoms by electron transfer from one atom to another (ex. $Na^+ + Cl^- \rightarrow NaCl$)

2.7.2 Oxidation state

Oxidation Number - Indicates the number of electrons that a particular element has lost gained or shared when forming a compound
Variable Valency - Represents elements that have more than one valency (ex. Fe^{2+} or Fe^{3+})

2.7.3 Radicals

Radical - Is a group of atoms that exist in several compounds but cannot exist on their own (ex. SO_4^{2-})

2.7.4 Covalent bonding

Types of bonds - Covalent, ionic (electrovalent), metallic, dative (dipolar)
- **Covalent Bond -** Characterized by the sharing of pairs of electrons between atoms
  - **Examples of covalent bonds -** Bonding between hydrogen and bromine
- **Ionic (Electrovalent) Bond -** A bond formed by the attraction between two oppositely charged ions. This usually involves a metal and a nonmetal ion, when a metal loses an electron and the nonmetal accepts them
  - **Examples of ionic bonds -** NaCl
- **Metallic Bond -** The electromagnetic interaction between metals and electrons
  - **Examples of metallic bonds -** Metals like zinc

Dative (Coordinate, Dipolar) Bonding - A type of bonding where both shared electrons come from the same atom (ex. CO)

Why covalent compounds do not conduct electricity - Because covalent compounds have no free ions which are necessary for carrying electric current in molten or solution form

2.7.5 Electrovalent bonding

Electrovalent Bond - Formed between two oppositely charged ions, usually between a metal and nonmetal

Differentiate between electrovalent and covalent compounds -

Electrovalent Compound	Covalent Compound
• Usually crystalline solid • Have high melting points • Generally soluble in water • Generally insoluble in organic solvents • Conduct electricity when in molten or solution form	• Often liquids or gases • Have low melting points • Generally insoluble in water • Generally soluble in organic solvents • Do not conduct electricity

2.7.6 Valency questions

1. What is the IUPAC name of H_2SO_3?

The total charge of the molecule must be 0. This means that the sum of all parts must equal zero. By looking at the periodic table we see that H is +1 and O_3 is -2. We are looking for the charge of S, since it is sulphuric acid

Given: Charge of H_2 = +1 * 2 = +2
Charge O_3 = -2 * 3 = -6
Charge of S = ?

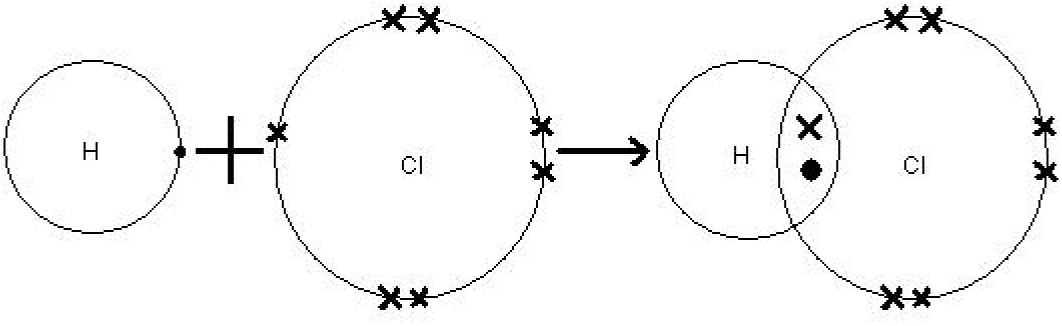
2 - 6 + S = 0
S = 4

Therefore the answer is **Sulphuric (iv) acid**

2. (a) Draw an electron diagram to show the covalent bonding between hydrogen and chlorine in a hydrogen chloride gas molecule

Electron diagrams show how two atoms form a covalent bond in a molecule. First you must think about how many outer electrons there will be for each of the atoms. This is done by looking at the periodic table. Hydrogen is found in the first column; therefore it has 1 electron in its outer shell. Chlorine is in the 7th column; therefore it has 7 electrons in its outer shell. The left side of the equation shows both of these atoms. Two different symbols are used to show the electrons of the two atoms. Circles represent hydrogen and crosses represent chlorine.

Since chlorine has one open place to take an electron and hydrogen has one to share, they will form a covalent bond.



(ii) **Explain why covalent compounds do not conduct electricity -** *Covalent compounds have no free ions which are necessary for carrying electric current in molten or solution form*

(b) What is the oxidation state of a Fe particle in a $FeCl_3$ molecule?

We know that chlorine lies in the 7th column. This indicates that it carries a charge of -1, meaning that it is able to share one electron from another atom. Since there are 3 chlorine atoms attached to the Fe atom, this means that the charge must be +3. Mathematically it is represented like this:

$$x + (-1 + 3) = 0$$
$$x - 3 = 0$$
$$x = +3$$

Form 3 Topics

3.1.0 Molecular and Ionic Equations

3.1.1 Molecular Equations

1. What is the formula of the compound formed by element M of group III and element X of group VI?

To solve this problem you must know the valency of elements in each group. Group I always has a +1 valency while group VII has a -1 valency. Group VIII has no valency since they are noble gases and do not react with other elements. Since it is asking about group III and VII, we need to think about how the valencies change. If it is +1 in group I, then you have to add 2 more to get to group III, making element M +3. Since element X is in group VI which is one less than group VII, we subtract 1 to get a -2 valency

Step 1: Find valencies Element M (Group III) = +3 Element X (Group VI) = -2 **Step 2: Find number of atoms in the compound.** *We need the valencies to add to zero, so we multiply valencies of the individual elements by a positive whole number to get +6 and -6*	Element M: +3 * 2 = +6 Element X: -2 * 3 = -6 Therefore, there are 2 atoms of element M and 3 atoms of element X in the compound. This gives the formula $\mathbf{M_2X_3}$

2. 2F was required to deposit one mole of metallic element M from an aqueous solution of its salt. If element M has no variable valency, what is the empirical formula of its phosphate?

Given: 1F = 1 mole of e^-
Therefore M requires 2 moles of e^- (2F = 2 moles of e^-)
So M is M^{2+}, *note that it is + because it is a metallic element*
PO_4 is always -3

Step 1: Find valencies Element M = +2 PO_4 = -3 **Step 2: Find number of atoms in the compound.** *We need the valencies to add to zero, so we multiply valencies of the individual elements by a positive whole number to get +6 and -6*	Element M: +2 * 3 = +6 PO_4: -3 * 2 = -6 *Therefore, there are 3 atoms of element M and 2 atoms of PO_4 in the compound. This gives the formula* $\mathbf{M_3(PO_4)_2}$

3. An oxide of iron, 4.5g by mass, was completely reduced by heating it in a certain reducing agent and 3.15g of iron was produced. Calculate the empirical formula of the compound

To calculate the empirical formula of a compound, you must first find the ratio between the atomic mass and mass given in the problem for each element. Then you take the ratio and divide both by the smallest value to get the numbers used to form the equation

Step 1: Calculate the mass of oxygen. *Subtract the mass of the iron from the iron oxide* 4.5g - 3.15g = 1.35g of oxygen **Step 2: Find the ratio between the mass of**	$O = \frac{0.084375}{0.05625} = 1.5 \text{ or } \frac{3}{2}$ **Step 4: Write the chemical equation using these numbers**

oxygen and the atomic mass
Relative atomic masses: Fe = 56 O = 16

$$Fe = \frac{3.15}{56} = 0.05625$$

$$O = \frac{1.35}{16} = 0.084375$$

Step 3: Divide both ratios by the smallest ratio
Smallest ratio: Fe = 0.05625

$$Fe = \frac{0.05625}{0.05625} = 1$$

$$Fe_1O_{\frac{3}{2}}$$

Step 5: Since we cannot have fractions in a chemical equation, you must multiply by 2 to get whole numbers

$$Fe_{1*2} = Fe_2$$

$$O_{\frac{3}{2}*2} = O_3$$

Step 6: Combine these to form the final formula
Fe_2O_3

4. Balance the equation $Zn(NO_3)_{2(s)} \xrightarrow{Heat} ZnO_{(s)} + NO_{2(g)} + O_{2(g)}$ and then say what type of chemical reaction is represented by this equation.

This is an example of a decomposition reaction because the original substance has been broken down into its constituent (simpler) substances. Since atoms are not consumed in chemical reactions, the total of each element on each side of the equation must be the same

Step 1: See if the equation is balanced

L Side			R Side	
Zn	1		Zn	1
N	2		N	1
O	6		O	5

They are not balanced

Step 2: Choose a number to multiply the left side by to begin trying to balance the equation. *Two is a good number to begin with. In this case we chose 2, but if two does not work, we can choose 3 or 4 and so on until the equation balances. If there is more than one molecule on the left side, start with one, then the other, then both and so on*

$$2Zn(NO_3)_2$$

This gives us the following table (note that the right side has not changed yet):

L Side			R Side	
Zn	2		Zn	1
N	4		N	1
O	12		O	5

Step 3: Now think of what we can multiply the molecules on the right side by to equal the values on the left side. *In this case we can multiply zinc by 2 and nitrogen by 4. Since you cannot multiply 5 by a whole number to get 12, we will look at oxygen last. Hopefully by balancing zinc and nitrogen, oxygen will balance as well*

$$2ZnO + 4NO_2 + O_2$$

Gives us this table:

L Side			R Side	
Zn	2		Zn	2
N	4		N	4
O	12		O	12

Step 4: Now we see that we have balanced the equation, so we can now just combine the two sides to get the final formula

$$2Zn(NO_3)_{2(s)} \rightarrow 2ZnO_s + 4NO_{2(g)} + O_{2(g)}$$

5. Determine the empirical formula of a substance that has the following composition by mass: 49.5% manganese and 50.5% oxygen

Given: Percentage composition Mn - 49.5%
Percentage composition O - 50.5%
Atomic mass Mn - 55
Atomic mass O - 16

Step 1: Calculate ratio between % composition and atomic mass for each element. *To find the empirical formula, you have to find the ratio between each component. This is found by dividing the percentage*

oxygen is the lowest, we will use it as the base value for calculating the empirical formula. The value for Mn will be divided by the value for oxygen to see how many more atoms there are *composition by the atomic mass.*

$$\frac{\%\ Composition\, Mn}{Atomic\ Mass\ Mn} = \frac{49.50}{55} = 0.9$$

$$\frac{\%\ Composition\, O}{Atomic\ Mass\ O} = \frac{50.50}{16} = 3.16$$

Step 2: Find the ratio between Mn and O atoms in the compound. *Since the value for*

in the compound

$$\frac{0.9}{0.9} = 1Mn \quad \frac{3.16}{0.9} = 3.5\ O$$

Step 3: Calculate the final formula. *Since there is a decimal value for oxygen, we must multiply it by 2 to get a whole number. In the process we must multiply the Mn by 2 also. This gives us 2 Mn atoms for every 7 O atoms, therefore the empirical formula will be* $\boldsymbol{Mn_2O_7}$

3.1.2 Ionic Equations

1. Write the ionic equation for $AgNO_3 + HCl \longrightarrow AgCl + HNO_3$

Step 1: Write all ions of the equation separately
$AG^+ + NO_3^- + H^+ + Cl^- \rightarrow AgCl + H^+ + NO_3^-$

Step 2: Cancel spectator ions. *If you see an ion on both sides of the equation, cancel them*

$AG^+ + \cancel{NO_3^-} + \cancel{H^+} + Cl^- \rightarrow AgCl + \cancel{H^+} + \cancel{NO_3^-}$

Step 3: Write the ionic equation with the remaining reactants and products
$Ag^+ + Cl^- \rightarrow AgCl$

3.1.3 Mass problems

1. The mass of anhydrous Copper (II) Sulphate produced by heating 50g Copper (II) Sulphate hydrated ($CuSO_4{\cdot}5H_2O$) is -

This question is asking about the weight of the Copper (II) sulphate without the water. Once it is heated, the water will evaporate, leaving only $CuSO_4$. If we find the ratio between anhydrous and hydrated $CuSO_4$ and then multiply it by the grams of hydrated $CuSO_4$, we will get the grams of $CuSO_4$. The ratio represents the percentage of $CuSO_4$ in $CuSO_4{\cdot}5H_2O$

Atomic masses: H = 1, C = 12, O = 16, S = 32, Cu = 63.5

Step 1: Find weights of each compound
$CuSO_4 = (63.5 + 32 + (16 * 4)) = 159.5g$
$CuSO_4{\cdot}5H_2O = (159.5) + 5 * ((1 * 2) + 16) = 249.5$
Step 2: Find the ratio between the two compounds

$$\frac{159.5}{249.5} = 0.639$$

Step 3: Multiply the ratio by the original mass
0.639 * 50g = **32g**

Another simpler example of this type of question is as follows: What is the weight of oxygen in 180g of water (H_2O)?

*Since the weight of H_2O is 18 (16 + 1 * 2) and oxygen on its own is 16, we can get a simple ratio of oxygen to water*

$\frac{16}{18} = 0.88$*, we can take this and multiply it by the original grams of water (180g) to get* ***160g of oxygen***

2. What is the loss in mass when 100g of calcium carbonate is strongly heated to a constant mass?

This question is asking what the remaining mass will be after burning has released CO_2 gas and left behind a solid precipitate.

$$CaCO_3 \xrightarrow{Heat} CaO + CO_2$$

By looking at the equation we see that there is just one mole of CO_2 created in this reaction, so the answer is simply the molar mass of CO_2

C = 12, O = 16
$CO_2 = 12 + (16 * 2) =$ **44g**

3. A certain type of soil requires 40kg of nitrogen per litre in order to fulfill its requirements for nitrogen. Calculate in kg the quantity of ammonium sulphate $(NH_4)_2SO_4$ fertilizer required to meet this demand

Step 1: Find the molar mass of $(NH_4)_2SO_4$	**Step 2: Find the grams of nitrogen required**
N = 14 x 2 = 28	40 kg = 40,000 g of N required
H = 1 x 4 x 2 = 8	$\frac{40,000g \times 132g}{28g} = 188,571.7g$
S = 32 x 1 = 32	
O = 16 x 4 = 64	
Molar mass = 28 + 8 + 32 + 64 = 132	***188.57 kg*** *of $(NH_4)_2SO_4$ will be required*

3.2.0 Hardness of Water

3.2.1 Concept, types and treatment of hard water

Temporary hardness of water - Is due to dissolved calcium hydrogen carbonate $(Ca(HCO)_3)_2$ and can be softened by boiling the water. Hard water requires more soap for washing because the soap reacts with the minerals to form scum. Chemical equation of this is:

$$Ca(HCO_3)_2 \xrightarrow{boiling} CaCO_3 + H_2O + CO_2$$

Permanent hardness of water - Is due to dissolved sulphates of calcium and magnesium and cannot be softened by boiling. Chemical equation of this is: $Ca^{2+} + CO_3^{2-} \longrightarrow CaCO_3$

Methods of water treatment - Filtration, precipitation of mineral salts calcium hydroxide, chlorination, fluoridation

Detergent can form lather even in hard water

3.3.0 Acids, Bases, and Salts

3.3.1 Acids and bases

pH - The measure of the acidity or basicity of a solution, where 7.0 is neutral. Anything over 7.0 is considered a base and anything below 7.0 is considered an acid. The scale is from 1-14

Base/Alkali - A substance which forms hydroxyl ions (OH^-) when dissolved in water. Concentrated solutions are caustic. Have pH of 7.1 - 14

- **Strong Base -** A base which completely ionizes in an aqueous solution (ex potassium hydroxide (KOH), sodium hydroxide (NaOH))
- **Weak Base -** A base which slightly ionizes in a dilute aqueous solution (ex ammonia (NH_3))
- **Basicity of an base -** The measure of how basic a base is (ex. how high the pH is, 8 is slightly basic, 14 is highly basic)
- **Basic Solution -** Occurs when a solution has a surplus of hydroxyl (OH^-) ions

Acid - A compound that when dissolved in water gives a pH of less than 7.0. Turns blue litmus paper red. The give up hydrogen ions (H^+)

- **Strong Acid -** An acid which ionizes completely in an aqueous solution (ex hydrochloric acid (HCl), sulfuric acid (H_2SO_4), hydrobromic acid (HBr))
- **Weak Acid -** An acid which slightly ionizes in a dilute aqueous solution to give hydrogen ions (H+) (ex acetic acid (CH_3COOH))
- **Acidity of an acid -** The measure of how acidic an acid is (ex how low the pH is, 6 is slightly acidic, 1 is highly acidic)
- **Acidic Solution -** Occurs when a solution has a surplus of hydrogen ions (H^+)

Neutralization - A type of chemical reaction where an acid and a base react to form water and a salt

Acids turn blue litmus red and bases turn red litmus blue

3.3.2 Indicators

**Found in the volumetric analysis section 3.5.0*

3.3.3 Salts

Acidic Salt - Salts which contain replaceable hydrogen because not all of their hydrogen was replaced by metal (e.g. $NaHSO_4$)

Salt - A chemical substance formed when either all or part of the ionizable hydrogen of an acid is replaced by a metallic ion or ammonium ion

- **Soluble Salts -** Salts which dissolve in water
- **Insoluble Salts -** Salts which do not dissolve in water
- **Sparing Soluble Salts -** Salts which do not dissolve in water, but can with an increase in temperature

Comparison of deliquescence and efflorescence -

Deliquescence - The tendency of a salt to absorb water from the atmosphere and form a solution (e.g. calcium chloride ($CaCl_2$)

Efflorescence - The tendency of a salt to loose its water to the atmosphere (e.g. sodium carbonate (Na_2CO_3)

Normal salt is a compound where all ionizable hydrogen has been replaced

3.4.0 Mole Concept

3.4.1 Mole as a unit of measurement

Stoichiometry - Is the quantitative relationship of reacting substances

Mole - Is the measurement of the amount of substance containing Avogadro's number of particles (6.02 x 10^{23} particles). This is the same number of particles as the number of atoms contained in 12g of carbon-12

Avogadro's Number - Number of atoms in one mole of any gas at STP (6.02×10^{23})

Atomicity - The total number of atoms present in one molecule of substance (ex. 2 hydrogen atoms in H_2O)

Molar Mass - Represents the mass of something per amount of substance (ex. the mass of C in CO_2)

Molar Solution - Is a solution that contains one mole of solute in one dm^3 or one litre

Molarity (Molar Concentration) - Is the amount of solute per volume of solution

Molar Volume of Gases - The volume occupied by one mole of the gas at STP

Vapour Density - The mass of a certain volume of the vapour gas compared with the mass of the same volume of hydrogen at the same temperature

$$Amount\ of\ substance = n = \frac{m}{M} = \frac{volume\ of\ substance}{molar\ volume}$$

$$2 \times Vapour\ density = \frac{mass\ of\ 1\ molecule\ of\ gas}{mass\ of\ 1\ atomof\ H_2}$$ *(Note it is 2 x vapour density because hydrogen gas is diatomic (there are two atoms of hydrogen in H_2)*

3.4.2 Application of the mole concept

1. What is the molarity of a solution containing 26.5g of anhydrous sodium carbonate in $5dm^3$ of solution?

Given: 26.5g anhydrous sodium carbonate (Na_2CO_3)
$5dm^3$ solution
Atomic masses: O = 16, Na = 23, C = 12

The equation for molarity is: $Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$

The equation for concentration is: $Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$

Step 1: Find the concentration

$$Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$$

$$Concentration(g/dm^3) = \frac{26.5}{5} = 5.3$$

Step 2: Find the molar mass of Na_2CO_3

Na_2CO_3 = (2 * 23) + 12 + (16 * 3) = 106g

Step 3: Calculate molarity

$$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$$

$$Molarity = \frac{5.3}{106} = 0.05M$$

The question can also be asked like this: What mass of pure sulphuric acid is found in $400cm^3$ of its 0.1M aqueous solution?

Given: Volume = $400cm^3$ or $0.4dm^3$ sulphuric acid (H_2SO_4)
Molarity = 0.1M

Step 1: Calculate molar mass of H_2SO_4

H_2SO_4 = (1 * 2) + 32 + (16 * 4) = 98g

Step 2: Calculate concentration

Step 3: Calculate mass

$$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$$

$$0.1M = \frac{C}{98}$$

$$C = 9.8g/dm^3$$

$$Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$$

$$9.8 = \frac{Mass}{0.4}$$

$$Mass = 3.92g$$

2. The sample containing the greatest number of atoms is:
a. 6g of hydrogen atoms b. 6g of sulphur atoms c. 6g of carbon atoms d. 6g of oxygen atoms e. 6g of helium atoms

The answer is a) because hydrogen is the lightest substance, therefore it takes more hydrogen atoms to weigh 6g than it does for heavier elements like sulphur or even helium

3. What is the mass of 11.2dm^3 of ammonia gas at STP?

Given: 11.2dm^3 of ammonia gas (NH_3)
Atomic masses: H = 1, N = 14

Formula to calculate moles of a gas at STP: $Moles = \frac{volume\ given(dm^3)}{22.4(dm^3mol^{-1})}$ *(Only works at STP)*

Step 1: Calculate number of moles of ammonia

$$Moles = \frac{volume\ given(dm^3)}{22.4(dm^3mol^{-1})}$$

$$Moles = \frac{11.2}{22.4} = 0.5\ moles$$

Step 2: Calculate molecular mass of NH_3

NH_3 = (14 + (1 * 3)) = 17

Step 3: Calculate mass by multiplying by the molecular mass of NH_3 by the number of moles

17 * 0.5 = **8.5g**

4. Ethyl alcohol has a molecular formula of C_2H_6O. If 9.2g of ethyl alcohol is available, calculate the
(i) Number of molecules of ethyl alcohol in 9.2g of ethyl alcohol

This question is about molar conversions. To solve these types of problems you must first find the weight of the molecule. To do this you first find the atomic weight of each element then multiply that by the number of atoms of that element which are present in the molecule. Then you add all of these together to get the molar mass. Then you divide the mass given in the problem by the molar mass to get the number of moles. To find the number of molecules, you must multiply the number of moles by Avogadro's constant

Step 1: Find the molar mass of C_2H_6O. First find the masses of each element in the molecule

C = 12	H = 1	O = 16
C = 12 x 2	H = 1 x 6	O = 16x1
C = 24	H = 6	O = 16

Step 2: Sum each element to find the molar mass

24 + 6 + 16 = 46 g/mol

Step 3: Divide the mass given in the problem by the molar mass to find the number of moles

Num of Moles = $\frac{9.2}{46} = 0.2$ moles

Step 4: Calculate the number of molecules by multiplying the number of moles from step 3 by Avogadro's constant

Avogadro's constant = 6.02 x 10^{23}
N = 0.2 x (6.02 x 10^{23})
N = 1.204 x 10^{23} molecules

(ii) Percentage by weight of oxygen in 9.2g of ethyl alcohol

Since we calculated the molar mass of the molecule earlier, we can use it to solve this problem. Out of the 46 g/mol, we know that 16g of it is oxygen, therefore 16/46 of the molecule is oxygen. Then we simply multiply this fraction (which represents the percentage of oxygen in ethyl alcohol) by the total grams of ethyl alcohol to get the weight of oxygen used. Now we can find the percentage by weight by dividing this result by the mass of the compound

Step 1: Write out the variables

O = 16 x 1　Molar Mass of C_2H_6O = 46
O = 16　Mass given 9.2g of C_2H_6O

Step 2: Find the fraction of oxygen present in C_2H_6O. *Divide the mass of oxygen by the molar mass of the compound. Multiply this by 9.2g to get the mass of oxygen by weight of C_2H_6O*

$$\frac{16}{46}\times 9.2 = 3.2g$$ of oxygen in C_2H_6

Step 3: Calculate the percent weight. *Divide the mass of oxygen by the total mass given in the problem*

$$\frac{3.2}{9.2}\times 100 = 34.78\%$$

5. (a) How many hydrogen atoms are there in $2.57\text{x}10^{-6}$g of hydrogen?

$$number\ of\ \ atoms\ of\ \ x = mass\ of\ \ x \times Avogadro's\ number$$

$$number\ of\ \ atoms\ of\ \ hydrogen = 2.57\times 10^{-6}\times 6.02\times 10^{23}$$

$$number\ of\ \ atoms\ of\ \ hydrogen = 1.54714\times 10^{-18}$$

This solution was simple because we are dealing with molecular hydrogen (H_2), if it was a compound, we would have to account for each element separately

(b) Ammonium sulphate combines with sodium hydroxide solution to produce sodium sulphate. With the aid of a balanced chemical equation calculate the mass of sodium hydroxide that would combine with 52g of ammonium sulphate

$$(NH_4)_2SO_{4(aq)} + 2NaOH_{(aq)} \longrightarrow NaSO_{4(aq)} + 2H_2O_{(l)} + 2NH_{3(g)}$$

Atomic masses: N = 14, H = 1, S = 32, O = 16, Na = 23

Step 1: Find the atomic masses of the two compounds

Atomic mass of $(NH_4)_2SO_4$ = ((14+(1*4)) * 2)+(32+(16*4)) = 132g
Atomic mass of 2NaOH = 2 * (23+16+1) = 80g
Note that we multiply by 2 because in the chemical equation there are 2 NaOH molecules for every $(NH_4)_2SO_4$ molecule

Step 2: Find the ratio between the atomic mass of sodium hydroxide and the atomic mass of ammonium sulphate

$$\frac{Atomic\ mass\ of\ \ 2NaOH}{Atomic\ mass\ of\ \ (NH_4)_2SO_4} = \frac{80}{132} = 0.606$$

Step 3: Multiply the ratio by the grams of ammonium sulphate to get the grams of sodium hydroxide

$$0.606\times 52 = 31.52g$$

The amount of sodium hydroxide is ***31.52g***

6. 20 cm^3 of a solution containing 7g/dm^3 of metal hydroxide, XOH, were exactly neutralized by 25cm^3 of 0.10M hydrochloric acid

(a) Write a balanced chemical equation for the neutralization of the metal hydroxide XOH and hydrochloric acid

$$XOH + HCl \longrightarrow XCL + H_2O$$

(b) Calculate the concentration of the metal hydroxide XOH in moles per dm^3

Data Given: V_B = 20 cm^3
Concentration of B = 7g/cm^3
V_A = 25cm^3
M_A = 0.1M

Use $M_BV_B = M_AV_A$ instead of the dilution law because N_A and N_B equal 1

$$M_B\times 20 = 0.1\times 25$$

$$M_B = 0.1\times\frac{25}{20}$$

$$M_B = 0.125M$$

(c) Calculate the molar mass of X

Since $Molarity = \frac{Concentration\ g/dm^3}{Molar\ Mass\ g/mol}$, we can say that

$$Molar\ Mass\ g/mol = \frac{Concentration\ g/dm^3}{Molarity}$$

$$Molar\ Mass = \frac{7}{0.125} = 56\ g/mol$$

(d) Identify the element X

Mass of X = (Mass of XOH) - (mass of O + mass of H)
Mass of X = 56 - (16 + 1)
Mass of X = 39

Potassium has an atomic number of 39, therefore X is ***potassium***

7. (a) What would be the molarity of a solution if 46g of sodium hydroxide (NaOH) were dissolved in 2000cm^3 of solution?

Molarity Formula: $Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$

Concentration Formula: $Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$

Data Given: Volume: 2000cm^3 or 2dm^3
Mass: 46g
Atomic masses: Na = 23, O = 16, H = 1
Mass of NaOH: 23 + 16 + 1 = 40g

Step 1: Find the concentration	**Step 2: Find the molarity**
$Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$	$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$
$Concentration(g/dm^3) = \frac{46}{2} = 23g/dm^3$	$Molarity = \frac{23}{40} = 0.575M$

(b) 8.50g of a sample of iron required just 75cm^3 of 3.00 M HCl to dissolve it to give a neutral solution. Calculate the percentage purity of the iron sample

Given: 8.5g impure iron
Volume: 75cm^3 of 3.00M HCl
Molar mass of Fe = 56

$$Fe + 2HCl \longrightarrow FeCl_2 + H_2$$

Purity Formula: $\%Purity = \frac{mass\ of\ pure}{mass\ of\ impure} \times 100\%$

The mass of impure iron was given, so we need to find the mass of pure iron. We get the mass of pure iron by calculating the number of moles of iron used and multiply this by the molar mass.

$$Mass\ of\ iron = moles \times molar\ mass$$

We know the molar mass of Fe is 56, so we now need to find the number of moles of Fe present. It is going to be half of the number of HCl moles used because the formula shows that for every two molecules of HCl used in the reaction, one molecule of Fe is used.

$$Moles\ of\ \ HCl\ used = Molarity \times Volume = 3 \times \frac{75}{1000} = 0.225 moles$$

$$\frac{0.225}{2} = 0.1125 moles\ of\ \ Fe$$

Now we have all of the information needed to solve the problem

$$Mass\ of\ \ iron = moles \times molar\ \ mass = 0.1125 \times 56 = 6.3g$$

$$\%Purity = \frac{mass\ of\ \ pure}{mass\ of\ \ impure} \times 100\% = \frac{6.3}{8.5} \times 100\% = 74.12\%$$

3.5.0 Volumetric Analysis (Titration)

3.5.1 Standard volumetric apparatus

Methyl Orange (MO) - An indicator used in titrations. Acid - red, neutral - orange, base - yellow. pH range is 4.0-6.0

Phenolphthalein (POP) - An indicator used in titrations. Base - pink, neutral - colourless. pH range is 8.3-10.0

Suitable indicators for titration -
Strong acid against a strong base - Methyl orange (MO) or Phenolphthalein (POP)
Strong base against a weak acid - Phenolphthalein (POP)
Weak base against a strong acid - Methyl orange (MO)
Weak acid against weak base - None

3.5.2 Standard solutions

Standard Solution - Describes a solution with a known concentration expressed in molarity (M). They are used in titration to determine the concentration of a substance in a solution

3.5.3 Volumetric calculations

Titration (Volumetric Analysis) - A method used in a laboratory to determine the concentration of a reactant

1. A solution of sodium carbonate was prepared in order to get a 2M solution. 200cm^3 of this solution was used in a titration experiment. What will be the number of moles present in 200cm^3 of 2M solution used in the titration?

The equation for molarity is: $Molarity = \frac{Concentration(g/dm^3)}{Molar\ \ Mass(g/mol)}$

The equation for concentration is: $Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$

Given Data: Molarity = 2M solution of sodium carbonate (Na_2CO_3)
Volume = 200 cm^3 or 0.2 dm^3 solution
Atomic masses: O = 16, C = 12, Na = 23

Step 1: Calculate the molar mass of Na_2CO_3 Na_2CO_3 = (23 * 2) + 12 + (16 * 3) = 106g **Step 2: Calculate the concentration**	$212 = \frac{Mass}{0.2}$ $Mass = 42.4g$ **Step 4: Find number of moles**

$$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$$

$$2 = \frac{C}{106}$$

$$C = 212$$

Step 3: Calculate the mass

$$Concentration(g/dm^3) = \frac{Mass(g)}{Volume(dm^3)}$$

$$Moles = \frac{Mass\ of\ Sample}{Molar\ Mass}$$

$$Moles = \frac{42.4}{106} = 0.4$$

2. Washing soda has a formula $Na_2\ CO_3 \cdot H_2O$. 7.15g of hydrated sodium Carbonate was dissolved to make 250cm^3 of solution 25.0 cm^3 of this solution was titrated with 0.25M Hydrochloric acid (0.25 mol per litre or 1000cm^3) using screened methyl orange indicator. The end point was reached when exactly 20.0cm^3 of the acid was added.

$$Na_2CO_3 + 2HCl_{(aq)} \longrightarrow 2NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

(2a) How do you know that the end point of the titration has been reached?

It is indicated by the colour change of the methyl orange from yellow to orange. Yellow - base, orange - neutral, red/pink - acid

(2b) Find the number of moles of hydrochloric acid used in this titration

Given: 0.25 moles of HCl initially
20cm^3 volume of acid used
1000 cm^3 volume of the whole solution

We need to find what percentage of the HCl was used in this titration. Since we know that 20cm^3 of the volume was used and there was 1000 cm^3 volume initially, the percentage will be 20/1000. We then multiply this percentage by the original moles of HCl present to get the amount of HCl used

$$Moles\ of\ HCl\ used = 0.25 \times \frac{20}{1000}$$

$$Moles\ of\ HCl\ used = 0.005$$

(2c) Find the concentration of sodium carbonate in moles per litre

To find the concentration of sodium carbonate we simply use the dilution law and solve for M_B, which is the molarity of the base

Dilution Law: $\dfrac{M_AV_A}{M_BV_B} = \dfrac{N_A}{N_B}$

Given: M_A = 0.25M
V_A = 20 cm^3
N_B = 1

V_B = 25 cm^3
N_A = 2
M_B = ?

Step 1: Rearrange the dilution law to solve for M_B

$$M_B = \frac{M_AV_AN_B}{V_BN_A}$$

Step 2: Solve for M_B

$$M_B = \frac{0.25 \times 20 \times 1}{25 \times 2}$$

$M_B = 0.1$, *therefore the concentration of sodium carbonate is 0.1 moles/litre because M represents molarity*

(2d) Find the formula mass of $Na_2CO_3nH_2O$

$$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$$

Given: Molarity = 0.1M
Concentration = 28.6 g/dm^3

Simply calculate the molar mass from the Molarity equation above

$$0.1 = \frac{28.6}{Molar\ Mass}$$

$$Molar\ Mass = \frac{28.6}{0.1} = 286g$$

(2e) Find the number of molecules of water of crystallization in hydrated sodium carbonate

Given: H = 1, C = 12, O = 16, Na = 23

Mass of $Na_2CO_3nH_2O$ = 286g

Simply solve for n algebraically

$$Na_2CO_3nH_2O = 286$$

$$(23*2)+12+(16*3)+n((1*2)+16) = 286$$

$$106+18n = 286$$

$$18n = 180$$

$$n = 10$$

3. Samples of sodium hydroxide solution of unknown concentration were titrated against 0.2M ethanoic acid, and the results are as tabulated below:

Volume of pipette used = 10cm^3

Burette Readings (cm^3)	Pilot	1	2	3
Final Reading	21.50	45.05	30.00	40.00
Initial Reading	1.00	25.00	10.05	20.00

(3a) Find the average volume of the acid used

You do not include the pilot when finding the average. Only use 1, 2 and 3

$$\frac{20.05+19.95+20}{3} = 20$$

(3b) If the chemical equation of the reaction is

$$NaOH_{(aq)} + CH_3COOH_{(aq)} \longleftrightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$$

Calculate the Molarity of the sodium hydroxide solution and its concentration in g/dm3

Dilution Law: $$\frac{M_AV_A}{M_BV_B} = \frac{N_A}{N_B}$$

M_A = 0.2M V_A = 20 cm^3 N_B = 1	V_B = 10 cm^3 N_A = 1 M_B = ?

Step 1: Rearrange the dilution law for M_B	**Step 2: Solve for M_B**
$M_B = \frac{M_AV_AN_B}{V_BN_A}$	$M_B = \frac{0.2\times 20\times 1}{10\times 1} = 0.4M$

(3c) (i) Name the most suitable indicator for this reaction

POP (Phenolphthalein) because it is a titration of a weak acid against a strong base

(ii) If the indicator you named in (III) above was used, what would be the colour change at the end point of the titration

The colour change will be from pink to colourless. POP starts pink in a base solution and becomes colourless when in a neutral solution

4. A solution of HCl contains 3.6g of the acid in 1dm^3 of solution. 20cm^3 of this solution is neutralized by sodium carbonate solution in which 1.1g of the carbonate is dissolved in 100cm^3 of solution.

(4a) Calculate the mole ratio for the reaction

$$2HCl + Na_2CO_3 \longrightarrow 2NaCl + CO_2 + H_2O$$

$$Mole\ Ratio = \frac{N_A}{N_B} = \frac{2}{1}$$

There are two HCl molecules for every 1 Na_2CO_3 molecule, therefore the ratio is 2:1

(4b) Calculate the molarities of the acid and carbonate solution

$$Molarity = \frac{Concentration(g/dm^3)}{Molar\ Mass(g/mol)}$$

For the acid:

Given: Concentration = $3.6g/dm^3$
Atomic Masses = Cl = 35.5, H = 1
Molar Mass = 35.5 + 1 = 36.5g

$$Molarity = \frac{3.6}{36.5} = 0.1M$$

For the carbonate:

Given: Concentration = $11g/dm^3$ because the concentration was given in $100cm^3$ not $1000cm^3$
Atomic Masses = Na = 23, C = 12, O = 16
Molar Mass = Na_2CO_3 = (2 * 23) + 12 + (16 * 3) = 106g

$$Molarity = \frac{11}{106} = 0.1M$$

(4c) Calculate the volume of the carbonate solution used

Given: M_A = 0.1M
$V_A = 20cm^3$
N_B = 1
N_A = 2 (*Note that the number is 2 because the ratio between N_A and N_B as shown in part a is 2:1*)
M_B = 0.1M
V_B = ?

Dilution Law: $\frac{M_A V_A}{M_B V_B} = \frac{N_A}{N_B}$

Step 1: Rearrange the dilution law for V_B	**Step 2: Solve for V_B**
$V_B = \frac{M_A V_A N_B}{M_B N_A}$	$V_B = \frac{0.1 \times 20 \times 1}{0.1 \times 2} = 10cm^3$

3.6.0 Electrolysis

3.6.1 Ionic theory

Electrolyte - A substance which when in molten or solution state conducts an electric current and is decomposed by it

Examples of electrolytes - Salts, acids, alkali
Strong Electrolyte - Dissociates completely in water to form positive or negative ions
Weak Electrolyte - Dissociates partially in water so that most of the molecules remain undissociated

Conductor - Any substance in solid form that allows the passage of an electric current

Coulomb - Is a unit of electric charge. One coulomb is the amount of electric charge transported in one second by a steady current of one ampere

3.6.2 Mechanism of electrolysis

Electrolysis - Is the decomposition of an electrolyte due to the passage of an electric current through it

Electrochemical Equivalent - The mass of a substance or element liberated during electrolysis by one coulomb (or 1 ampere in 1 second) of electricity

Electrode - A piece of conductor metal or graphite through which electrons enter or leave an electrolyte

Oxidizing Agent - Substance that causes oxidation but itself is said to be reduced at the end of the reaction. Oxidizing agents are generally nonmetals since they accept an electron

Reducing Agent - An element or compound in a redox reaction that reduces another compound or element. Reducing agents are generally metals since they donate an electron when forming an ionic bond

Cathode - Is the negative electrode used in electrolysis. This is where you place the item that you want to cover with a metal during electroplating. This is the electrode where electric current enters the electrolyte because current moves from positive to negative (electrons move from negative to positive)

Cation - A positively charged ion that gets discharged at the cathode during electrolysis, formed when metals lose electrons. The discharge of cations from the cathode during electrolysis is reduction

Examples of cations - Na^+, NH_4^+, Ca^{2+}, Fe^{2+}, Fe^{3+}, Cu^{2+}, Zn^{2+}, Pb^{2+}

Ionic equation of cathode - $X + e^- \longrightarrow X^-$ **, X is reduced**

Anode - Is the positive electrode used in electrolysis

Anion - A negatively charged ion that gets discharged at the anode during electrolysis

Examples of anions - Rust, CO_3^{2-}, HCO_3^-, NO_3^-, SO_4^{2-}, Cl^-

Ionic equation of anode - $X^- \longrightarrow X + e^-$ **, X is oxidized**

Factors affecting the selection of ion discharge at electrodes - Concentration of the competing ions, position of the ion on the electrochemical series, nature of electrode used

Example of oxidizing - $Cu^+ \rightarrow Cu^{2+}$, electron is lost, therefore the charge increases by +1

Example of reduction - Cu+ → Cu, electron is gained, therefore the charge decreases by -1

3.6.3 Laws of electrolysis

Faraday's 1st law of electrolysis - The mass of a substance dissolved from an electrode is proportional to the quantity of electricity passing through the electrolyte (M = ZQ)

Faraday's 2nd law of electrolysis - When the same quantity of electricity is passed through solutions of different electrolytes, the amount of the elements deposited is proportional to the chemical equivalents of the elements

The amount of an element liberated by electrolysis depends upon current and time taken to complete the process

3.6.4 Application of electrolysis & questions

Electroplating - Is the process of coating metallic objects with a layer of another metal by using electrolysis

Uses of electroplating - Covering a metal so that it becomes abrasion and wear resistant, resists corrosion (rusting), becomes more lubricated, gains aesthetic qualities

1. When a current of 5A is passed through molten aluminum chloride for 3.5 hours, what will be the number of coulombs?

We use the formula for calculating coulombs (quantity of electricity)

Coulombs (C) = Current (I) * Time (T)

C = IT

Note that time must be in ***seconds***

Given: Time (T) = 3.5 hours

Current (I) = 5A

Step 1: Calculate time in seconds	**Step 2: Calculate coulombs**
3.5 hours * 60 minutes * 60 seconds = 12600 secs	C = IT
	C = 5 * 12600
	C = 63000 Coulombs (C)

2. A light bulb uses a current of 0.6 amps. How many Faradays will be used by this light each hour?

This question is asking for the quantity of electricity in Faradays. Quantity of electricity can be calculated in Coulombs first using Q = It, where I is current and t is time in seconds. Then we can use the definition of a Faraday which is 1F = 96500C to solve for the number of Faradays.

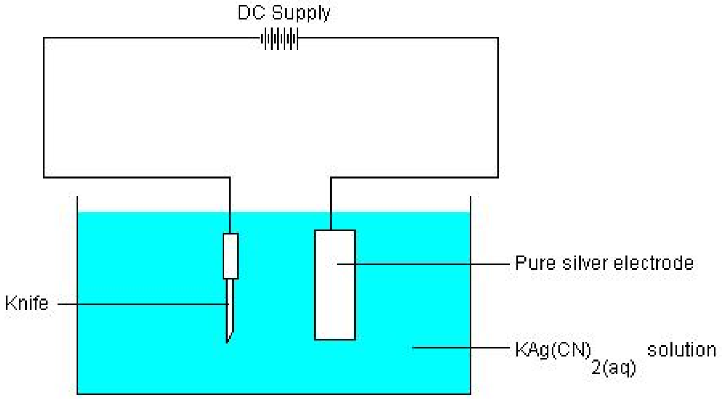
Given: I = 0.6 A

t = 1 hour * 60 mins * 60 secs = 3600 secs

1F = 96500C

Step 1: Calculate Coulombs used	**Step 2: Calculate Faradays used**
Q = It	1F = 96500C
Q = 0.6 * 3600	F = 2160 / 96500
Q = 2160 Coulombs	F = **0.2238F**

3. The apparatus shown below was used in an electroplating experiment with an iron know with silver



(i) Which electrode is the cathode?
The cathode is the iron knife

(ii) Write ionic equations to represent the reactions taking place at the electrodes

At the cathode: $Ag^{+}{}_{(aq)} + \overset{-}{e} \longrightarrow Ag_{(s)}$

At the anode: $Ag_{(s)} \longrightarrow Ag^{+}{}_{(aq)} + \overset{-}{e}$

(iii) A which electrode was the reduction taking place?
The reduction took place at the cathode (iron knife electrode)

4. A sample of impure silver with a mass of 3.45g was used as the anode during an electrolysis purifying process. The cathode was made up of pure gold with a mass of 6.45g. After the electrolysis the cathode was found to weight 9.66g

(i) Calculate the number of coulombs of electricity passed.
First we need to think about what is given in the problem and what information will be useful to us. We first write out the masses given in the problem. Since it is a problem calculating coulombs in an electrolysis process we will have to know the amount of Ag required for 1F and the conversion of F to C

Mass at anode (impure Ag) M_a = 3.45g
After electrolysis, mass of cathode = 9.66g
1F = 96500C
Mass at cathode (gold) M_c = 6.45g
108g of Ag = 1F
Now we can begin to solve the problem

Step 1: Calculate the amount of Ag deposited at the cathode. *Subtract the original mass from the new mass of the cathode*
9.66 - 6.45 = 3.21g

Step 2: Calculate the amount of F required. *Divide the grams of Ag deposited at the cathode by the number of grams of Ag required for 1F*

$$\frac{3.21}{108} = 0.0297F$$

Step 3: To calculate the required number of coulombs. *Multiply the result of step 2 by 96500*

$$C = 0.0297F \times 96500F/C$$

$$C = 2868.19\ Coulombs$$

(ii) What is the percentage purity of the impure silver?
To solve this problem we need to find the amount of Ag transferred which represents pure Ag. This was done in Step 1 above by subtracting the original mass from the new mass at the cathode.

Mass of pure Ag = 3.21
Mass of impure Ag = 3.45

Now you can divide them, the numerator is the pure Ag and the denominator is the impure Ag.

$$\frac{3.21}{3.45}\times 100 = 93.04\%$$

5. (i) Write a chemical equation for the discharging process at the anode and cathode when dilute sulphuric acid is electrolyzed using platinum electrodes

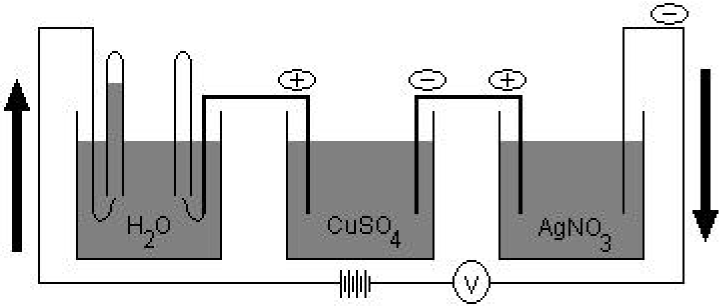
Dilute sulphuric acid H_2SO_4

Cathode: $4H^{+}{}_{(aq)} + 4\overline{e} \rightarrow 2H_{2(g)}$

Anode: $4OH^{-}{}_{(aq)} \rightarrow 2H_2O + O_2 + 4\overline{e}$

6. This figure represents an experiment where Faraday’s second law was illustrated by connecting in series three cells containing water where dilute sulphuric acid, copper sulphate solution and silver nitrate solution were added. A current of 1 A was passed through the solution for 2 seconds.

The volumes and hence weights of hydrogen and oxygen liberated were calculated. The weight of copper and silver formed by the electrolytes of copper and silver nitrate solutions were also measured.



The results of the experiment are shown below:

Element	Current	Time (s)	Mass of element deposited in grams (g)	Quantity of electricity	Electrochemical equivalent
Hydrogen	1.0	2	$2.0892\text{x}10^{-5}$		
Oxygen	1.0	2	$1.658\text{x}10^{-4}$		
Copper	1.0	2	$6.587\text{x}10^{-4}$		
Silver	1.0	2	$2.236\text{x}10^{-3}$		

(6a) Calculate the quantity of electricity passed and the electrochemical equivalent of the elements in the table above

*The quantity of electricity is simply found by multiplying the current by the time in seconds. This is represented by Q = IT (Quantity = Current * Time)*

Quantity of electricity = Current (I) * Time (T)
Quantity of electricity = 1.0 * 2
Quantity of electricity = **2 Coulombs**

To find the electrochemical equivalent you use the formula m = ZQ

m = ZQ
m = mass of element
Z = electrochemical equivalent
Q = Quantity of electricity

Since m and Q are given, we can rearrange the equation to Z = m/Q to solve for Z. Only hydrogen is shown, but the process is the same for all
For Hydrogen:

$$Z = \frac{m}{Q}$$

$$Z = \frac{2.0892 \times 10^{-5}}{2}$$

$$Z = 1.0446 \times 10^{-5}\, g / C$$

Element	Quantity of electricity (C)	Electrochemical equivalent
Hydrogen	2	1.0446×10^{-5}
Oxygen	2	8.29×10^{-5}
Copper	2	3.2935×10^{-4}
Silver	2	1.118×10^{-3}

(6b) If the Faraday constant is 96500 C, calculate the chemical equivalent of the four elements

$$Chemical\ Equivalent\ (C) = Electrochemical\ Equivalent \times Faradays\ Constant$$

Ex. Hydrogen

$$Chemical\ Equivalent\ (C) = Electrochemical\ Equivalent \times Faradays\ Constant$$

$$Chemical\ Equivalent\ (C) = (1.0446 \times 10^{-5}) \times 96500$$

$$Chemical\ Equivalent\ (C) = 1.008C$$

Element	Chemical Equivalent (C)
Hydrogen	1.008
Oxygen	8.000
Copper	31.782
Silver	107.887

(6c) What relationship is there between an electrochemical equivalent of an element and its chemical equivalent?

*Chemical Equivalent = Electrochemical Equivalent * Faraday's Constant*

3.7.0 Chemical Reactions and Chemical Equilibrium

3.7.1 Rate of reactions

Chemical Kinetics - The field of study which deals with the rate of chemical reactions

Indicators of reaction - Evolution of gas, colour change, temperature change, precipitation, change in mass

Equilibrium - The state of reaction at which the rate of forward reaction is equal to the rate of backward reaction (ex. the rate of formation of products equals the rate of formation of reactants)

Types of reactions - Decomposition, substitution, reversible, irreversible, catalysis, combination (synthesis), double displacement, thermal decomposition, oxidation, neutralization, redox

- **Decomposition -** Separation of a compound into simpler compounds or elements (often by heating)
- **Substitution (Displacement) -** A reaction where one reactive element replaces a less reactive one (ex activation series)
- **Reversible -** One which can proceed in either direction due to varying conditions such as temperature, pressure and concentration
- **Irreversible -** Reactions which proceed in one direction (forward) only no matter what the conditions are
- **Precipitation -** A reaction where an insoluble compound is formed when two solutions are mixed
- **Catalysis -** A reaction where a catalyst is present to speed up the reaction
- **Combination (Synthesis) -** A reaction where two or more elements form a single molecule as a product
- **Double Displacement -** A type of reaction where two reactants exchange bonds resulting in products with similar bonding between them (ex. $AgNO_3 + HCl \rightarrow HNO_3 + AgCl$)
- **Thermal Decomposition -** Chemical decomposition of a compound by heat (e.g. decomposition of water into H_2 and O_2 at high temperatures)

Oxidation - This reaction is important to animals because it makes chemical energy available to animals through the Kreb's cycle (formation of ATP). It involves the transfer of electrons

Neutralization Reaction - Is a reaction between an acid and a base resulting in the formation of a salt and water

Redox (Reduction-oxidation Reaction) - A reaction where one substance is oxidized and the other is reduced (ex combustion, respiration, rusting, corrosion)

3.7.2 Factors affecting rate of chemical reactions

Factors affecting rate of a chemical reaction - Concentration, pressure, temperature, catalyst, surface area of reactant, light (sometimes), nature of the substance (surface area)

Rate of a chemical reaction - The amount of reactions taken per unit of time

Activation Energy - The energy which must be overcome for a chemical reaction to occur. Catalysts lower the activation energy required for a reaction to take place, meaning less energy is needed for the reaction to occur. As the activation energy is lowered, the rate of reaction increases

Catalysts - A chemical substance which changes the speed of a chemical reaction, but itself is not consumed and remains unchanged chemically and unchanged in mass at the end of the reaction. Small proportions of catalysts are needed to increase the rate of reaction

Catalysis - The change in rate of a reaction due to the presence of a catalyst

Homogeneous Catalysis Reaction - A sequence of reactions involving a catalyst in the same phase as the reactants (i.e. they are all gases/solids/liquids)

Heterogeneous Catalysis Reaction - A sequence of reactions involving a catalyst of a different phase than the reactants (i.e. reactants are liquid, catalyst is solid)

Substances with a higher surface area undergo reactions at a faster rate (i.e. powder reacts faster than granules, and granules react faster than large pieces)

3.7.4 Equilibrium reaction

Le Chatelier's Principle - If a chemical system at equilibrium experience a change in concentration, temperature, volume or partial pressure, then the equilibrium shifts to counteract the imposed change and a new equilibrium is established

Factors affecting the position of equilibrium in a reversible chemical reaction - Temperature, pressure, concentration, catalyst

3.7.5 Endo/exothermic reactions

Endothermic Reaction - A reaction which absorbs energy in the form of heat from the surrounding environment. Products are favoured by an increase in temperature. Products have higher energy than the reactants, therefore the products are unstable. The total energy required to break the bonds is lower than the energy required to form new bonds. Represented by $+\Delta H$

Examples of endothermic processes - Evaporation of water, photosynthesis

Exothermic Reaction - A reaction which releases energy in the form of heat into the surrounding environment. The reactants have more energy than the products; therefore the products are more stable. The total energy required to break the bonds is higher than the energy required to form a new bond. Represented by $-\Delta H$

Examples of exothermic reactions - Explosions, condensation of rain from water vapour, mixing alkalis and acids, mixing water and strong acids, combustion of fuels (coal/wood/charcoal)

3.7.6 Chemical reactions questions

1. Some marble chips (calcium carbonate) were placed in a beaker containing an excess of dilute hydrochloric acid standing on a reading balance. The mass of the beaker and its contents were recorded every two minutes as shown in the table below

Time (Minutes)	0	2	4	6	8	10	12
Mass (g)	126.44	126.31	126.19	126.09	126.03	126.00	126.00

(a) Why was there a loss of mass?

There was a loss of mass because CO_2 gas was formed during the reaction and released into the air

(b) Write the equation for the reaction

$$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$$

(c) State different ways that the reaction can be sped up

- *You could have used powdered carbonate instead of chips because powdered material has a larger surface area and will react faster than chips which have a smaller surface area*

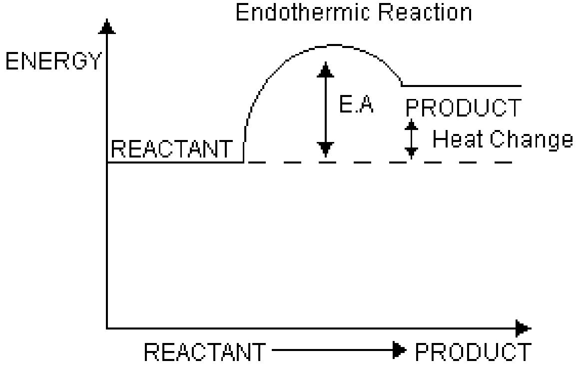
- *Use concentrated HCl instead of dilute HCl*
- *Increase temperature*

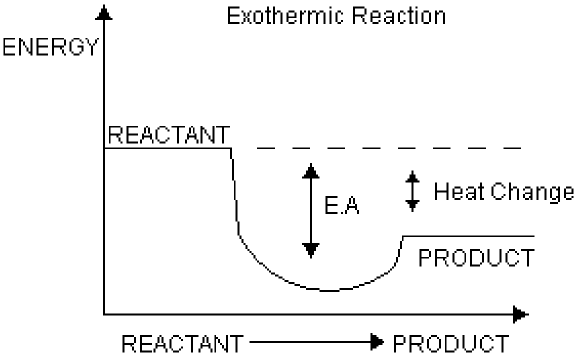
(d) Why did the mass remain constant after 10 minutes?
The reaction ended because all of the marble was used up

(e) Write the name and formula of the two ions remaining in the final solution
The ions remaining are calcium ions (CA^{2+}) and chloride ions (Cl^-)

2. (a) Distinguish between an endothermic and exothermic reaction
Endothermic Reaction - A reaction in which heat is taken or absorbed from the surroundings
Exothermic Reaction - A reaction where heat is given out or released into the surroundings

(2b) Draw a simple graph of the energy profile for an endothermic and exothermic reaction





For an endothermic reaction the heat change is positive. For an exothermic reaction the heat change is negative

(2c) Name the types of reaction represented by each of the following chemical reactions

(i) $2KClO_{3(s)} \xrightarrow{\Delta,MnO_2} 2KCl + 3O_{2(s)}$

Catalysis reaction

(ii) $Fe_{(s)} + S_{(s)} \xrightarrow{\Delta} FeS_{(s)}$

Combination (synthesis) reaction

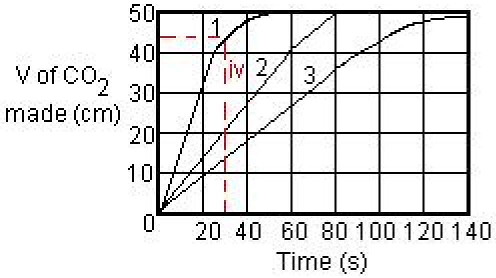
(iii) $AgNO_{3(aq)} + NaCl_{(aq)} \longrightarrow AgCl_{(s)} + NaNO_{3(aq)}$

Double decomposition reaction

(iv) $NH_4Cl_{(s)} \overset{\Delta}{\longleftrightarrow} NH_{3(g)} + HCl_{(g)}$

Thermal decomposition reaction

3. A pupil has carried out three experiments at 20°C to study the rate of reaction between marble chips and hydrochloric acid by observing the rate of creation of carbon dioxide. In each case marble chips of equal masses but of different sizes were used with excess 2.00 M hydrochloric acid. Finally the student showed the results graphically as seen below



(a) Use the graph to answer the following questions
(i) Which experiment gave the fastest reaction?

Experiment 1 gave the fastest reaction since it has the largest slope

(ii) In which experiment did the pupil use the largest size marble chips?
In experiment 3 because the reaction was the slowest due to a lower surface area

(iii) At what time did experiment 2 produce the largest volume of carbon dioxide?
In the first 20 seconds because its slope is the greatest during that period

(iv) What was the volume of carbon dioxide produce at the 30th second in experiment 1?
Roughly 45 cm^3 was produced. To find this we look at the intersection between 30 seconds and line 1, then read the y axis which is between 40 and 50 cm^3. This is seen by the dotted line in the graph

(v) Why did the volume of carbon dioxide not increase any further after 130 seconds in experiment 3?
All of the reactants were used up; therefore there could no longer be a reaction

(b) What will happen to the rate of production of carbon dioxide in experiments 1,2 and 3 if excess 3.000M hydrochloric acid is used instead of excess 2.00M hydrochloric acid? What will happen to the rate of production of carbonic acid?
The rate of the reaction will increase because increasing the concentration of a participating reactant always increases the rate of reaction

3.8.0 Extraction of metals

3.8.1 Chemical properties of metals

Metal - An element which can ionize by electron loss

Physical properties of metals - They are malleable, ductile, lustrous, sonorous (can produce sound), good conductors of heat and electricity, high tensile strength (can support heavy loads), high melting and boiling points, have high density

Chemical properties of metals - They are reducing agents, are involved in displacement reactions, most can react with an acid to produce a salt, few form compounds with hydrogen, they form electrovalent chlorides

Order of activity series from most reactive to least reactive - K, Na, Ca, Mg, Al, Zn, Fe, Pb, Cu

Reduction - The process of electron loss or a decrease in oxidation number

Reducing Agent - A substance which donates its electrons during the reaction and is said to be oxidized by the end of the reaction

Reducing Power - The easiness of electron donation of an element (metals at the top of the activity series have high reducing power and those at the bottom have lower reducing power)

Metal Extraction - The process of obtaining a pure form of metal from their natural ores

Ore - The natural occurrence of an impure metal compound

Activity Series - A series of metals from highest to lowest which refers to the metals ability to react with acids and water

Reactivity Series - The series of elements in order of increasing electrode potential

Chemically active metals - K, Ca, Na, Mg, Al

Base Metals - Zn, Fe, Pb, Cu

Noble/inert Metals - Ag, Hg, Au

Metals found in nature as true elements - Cu, Hg, Ag, Au *(this is because they are the least reactive)*

Solvay Process - The major industrial process for the production of sodium carbonate (soda ash)

Coke - A solid carbonaceous residue derived from the destructive distillation of coal

Producer Gas - A fuel gas made from coke or other carbonaceous material

Alloy - A solid solution of two or more metals. The alloy has different properties than the parent metals

Amalgam - An allow of mercury with another metal

Malleability - The degree to which a metal can be deformed without fracturing

Ductility - A property which describes the extent to which a solid can be deformed without fracture by tensile stress

Brittle - Used to describe materials which when subjected to stress will break easily

1. The following elements are arranged in order of an activity series: Na, Ca, Mg, Zn, Fe, H, Cu, Hg
From these elements, name:

(i) A metal which react with cold water
Na because it is at the beginning of the activity series and is therefore more reactive

(ii) A metal which burns in steam but does not react with cold water
Ca because they are less reactive than Na so they cannot react with cold water, but they can react with steam since the temperature is higher

(iii) Any other element which reacts when heated in steam
Mg because it is less reactive than Na, but it can react with steam due to the higher temperature

(iv) An element which has oxide which decomposes on heating
Hg because it is the least reactive in the series. It is found very low on the reactivity series
(v) Those elements which do not displace hydrogen from dilute hydrochloric acid
Hg and Cu cannot displace hydrogen because they are found at the bottom of the series and cannot displace those which are more reactive and higher up in the series

2. (a) Sodium, magnesium, zinc, copper and silver are five metals which appear in this order in the activity series: Sodium being the most reactive and silver being the least reactive. Which one of these metals is
(i) Likely to tarnish most rapidly when exposed to the air
Sodium is most likely to tarnish because it is highly reactive
(ii) Most likely to be found in nature
Silver is most likely to be found in nature because it is the least reactive
(iii) Least likely to react with steam
Silver or copper because they are the least reactive

(b) Two of the metals above are usually extracted by electrolysis of their molten chlorides. Name the two metals and the reason this method is used
Magnesium and sodium are extracted by electrolysis of their molten chlorides. They are extracted by electrolysis of their molten chlorides because they are highly reactive. If aqueous solutions are used, the H^+ and OH^- ions would interfere with the discharge in the electrodes, so that process cannot be used

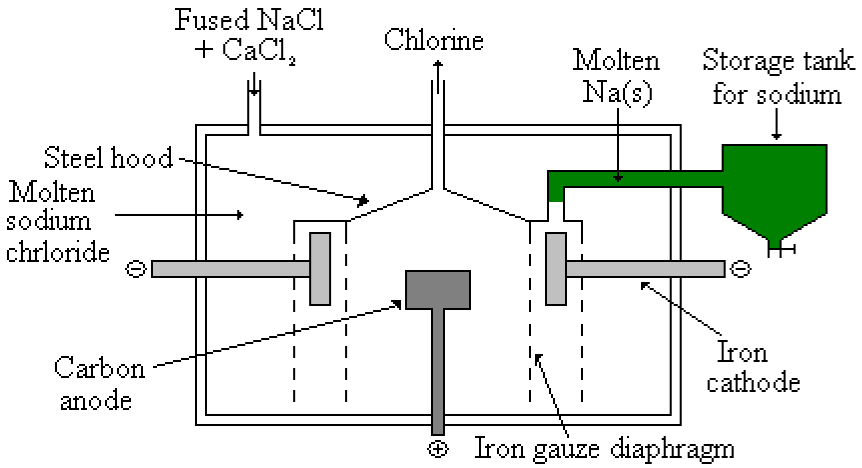
(c) (i) To which electrode will the sodium ions in an aqueous dilute solution of sodium chloride migrate during electrolysis
Since the sodium ions are positive, they will migrate to the cathode
(ii) What other ions will migrate to the cathode?
H^+ ions since they are positively charged
(iii) Which ions will be discharged at the anode? Give reasons for your answer
H^+ ions will be discharged due to their position in the electrochemical series

3.8.3 Extraction of metals by electrolytic reduction

Metals extracted by electrolytic reduction - K, Na, Ca, Al, Mg
Why these metals are extracted by electrolysis - This is because they are the most reactive metals, so they are never found as free elements in nature, they are only found in ores
Ways of extracting metals through electrolysis - Reduction of their oxides by carbon (coke) or carbon monoxide, oxidation of their sulphates to form a metal or metal oxide

Extraction of sodium (Na)

Occurrence of sodium - Occurs mainly has a chloride (common salt, NaCl), sodium carbonate, sodium nitrate (saltpeter), sodium chloride (NaCl)
Process of extraction of sodium -
Apparatus - Sodium is extracted by electrolysis in a special cell called Down's electrolytic cell
Ore Used - NaCl mixed with $CaCl_2$
Down's Process - Sodium chloride (NaCl) is electrolyzed in a molten state with calcium chloride ($CaCl_2$) to lower the melting point of NaCl. The outer shell of the cell is made of iron lined with firebrick and there is a diaphragm of iron gauze screens in the carbon anode from the ring shaped iron cathode that surrounds it. Chlorine escapes via the hood while the sodium collects in the inverted trough over the cathode, rises up the pipe and is tapped off through an iron vessel. The gauze steel wire prevents the mixing of the products of electrolysis
Down's electrolytic Cell - Cell designed for the extraction of sodium metal from NaCl



Reaction at the cathode - Sodium ions gain an electron and sodium metal is deposited and collected

$$Na^{+} + e - \longrightarrow Na_{(s)}$$

Reaction at the cathode - Each chloride ion loses an electron to combine with another chloride ion to form chlorine gas which is collected as a useful byproduct

$$Cl^{-} - e^{-} \longrightarrow Cl$$

Properties of sodium - Soft metal with a silvery appearance, highly electropositive, tarnishes rapidly in air forming the monoxide (Na_2O), reacts with water to form the hydroxide (NaOH), the hydroxide reacts with CO_2 in air to form sodium carbonate (Na_2CO_3), it is kept in liquid paraffin (kerosene) in the lab

Uses of sodium - As an alloy with lead for use as an antiknock agent in petrol engines, used in the isolation of gold from the metallic ore, to prepare sodium peroxide, as a heat conductor in nuclear reactors, as a vapour in sodium lamps, in making various alloys

Extraction of aluminum

Ores used in extraction - Bauxite, cryolite, corundum, kaolin

Process of aluminum extraction -

Bayer Process - Bauxite is dissolved in sodium hydroxide solution under pressure to form sodium aluminate ($NaAl(OH)_4$). This solution is filtered and then seeded with pure aluminum oxide to cause the aluminum hydroxide to crystallize in a pure state. Then it is heated in rotary kilns to produce pure aluminum oxide. The pure aluminum oxide is dissolved in molten cryolite and the mixture acts as the electrolyte in the cell

At the cathode - Aluminum is then deposited at the cathode which is the carbon lining of the container

$$Al^{3+} + 3e^{-} \longrightarrow Al(s)$$

At the anode - Oxygen is evolved at the carbon anode

$$2O^{2-} \longrightarrow O_{2(g)} + 4e^{-}$$

Uses of aluminum - To make light weight roofing, as household items like saucepans and kettles, in the manufacture of alloys, as an overhead cable, as foil for wrapping food or other products, as a reducing agent in the production of molten iron for welding

3.8.4 Extraction of metals by chemical reduction

Metals extracted by chemical reduction - Zn, Fe, Pb

Why these metals are extracted by chemical reduction - These metals are moderately reactive so they are found as oxides, carbonates or sulphides which are then reduced by carbon or carbon monoxide

Four stages of extraction by chemical reduction - Concentration of the ore, roasting of the ore, reduction, refining

Concentration - Some impurities are separated from the ore by physical means (by hand or washing with water, or by magnetic means)

Roasting - The concentrated ore is roasted in air at high temperatures to produce the metal oxide which is easier to reduce to the pure metal than the sulphide or carbonate

Reduction - The oxide is then heated in a furnace in the presence of a reducing agent (usually coke or carbon) to the metal which then must be refined

Refining - The metal with impurities is then purified by electrolysis or other chemical means

Extraction of iron

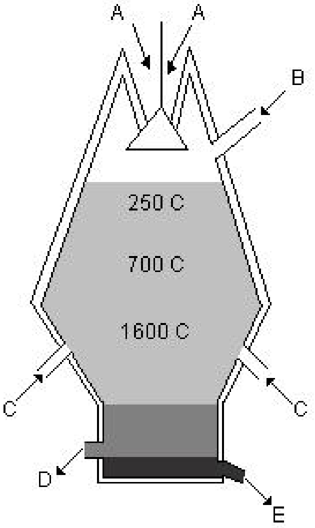
Ores used in extraction - Hematite (Fe_2O_3), magnetite (Fe_3O_4), siderite ($FeCO_3$), limonite ($FeO(OH)$)

Functions of coke - Acts as a reducing agent, produces reducing agent, produces heat energy required for the process

Functions limestone - To remove impurities of silicon dioxide, to produce a reducing agent, formation of slag which is used for road making or fertilizers

Functions of waste gases - Used for preheating the blast furnace, used for steam raising, used for heating coke ovens

Why it is named a blast furnace - Since air is forced in at a high pressure it is being blasted



2. This figure represents a blast furnace for the extraction of iron from its ore.

(a) (i) Name the ore used for the extraction of iron

Hematite (Fe_2O_3) is the ore commonly used in the extraction of iron

(ii) What substances are fed into the furnace at point A?

Hematite (Fe_2O_3), limestone ($CaCO_3$), coke (C)

(iii) Name the substance introduced at point C

Hot air

(b) Write a balanced chemical equation for the reactions taking place at (i) 1600 C (ii) 700 C (iii) 250 C

(i) $C + O_2 \longrightarrow CO_2$ (ii) $2CO_2 + C \longrightarrow 2CO$

(iii) $Fe_3O_4 + 4CO \longleftrightarrow 3Fe + 4CO$

(c) (i)What are the products produced at D and E?

D - Molten slag *E - Molten iron*

(ii) Why was $CaCO_3$ added to the blast furnace?

It is added to remove silica (SiO_2) impurities after undergoing decomposition

(iii) What is the function of CO formed in the blast furnace?

Its function is to reduce Fe_2O_3 to Fe

(d) Uses of impure iron metal - *To manufacture cooking utensils, used to manufacture roofing materials*

(e) If 80kg of iron ore (Fe_2O_3) were allowed to react with carbon monoxide (CO) during the extraction process, how many kilograms of iron (Fe), would be obtained?

Given: 80kg if iron ore (Fe_2O_3)

Step 1: Find the equation representing the reaction. *The left side is just the iron ore + carbon monoxide. The right side will be the products which are iron and the remaining CO. Then balance the equation*

$$Fe_2O_3 + 3CO \longleftrightarrow 2Fe + 3CO$$

Step 2: Calculate the number of moles of Fe_2O_3. *Since the mass was given in kg, we need to multiply it by 1000 to get it in grams*

$$moles\ of\ Fe_2O_3 = \frac{mass\ of\ Fe_2O_3}{molar\ mass\ of\ Fe_2O_3}$$

$$\frac{80 \times 1000}{(56 \times 2) + (16 \times 3)} = 500\ moles$$

Step 3: Calculate the number of moles of Fe. *Since there are two atoms of Fe in every molecule of Fe_2O_3, you will multiply the moles of Fe_2O_3 by two get the number of moles of Fe produced by one mole of Fe_2O_3*

500 * 2 = 1000 moles of Fe produced

Step 4: Multiply the moles of Fe by the atomic weight of Fe

*1000 * 56 =* ***56000 g or 56kg***

Extraction of copper

Ores used in extraction - Copper pyrite ($CuFeS_2$), malachite ($CuCO_3$)

Why silicon dioxide is added - It is used to remove the impurities of iron (II) oxide by converting it into slag

1. Write the process for the extraction of copper metal from its ore with balanced chemical equations for each step

(i) Roasting of concentrated ore in the air

$$2CuFeS_{2(s)} + 4O_{2(g)} \longrightarrow Cu_{2(s)}S_{(s)} + 3SO_{2(g)} + 2FeO_{(s)}$$

(ii) Heating of the roasted ore with silica in the absence of air

$$FeO_{(s)} + SiO_2 \longrightarrow FeSiO_{3(l)}$$

(iii) Burning copper sulphide ore Cu_2S in with a regulated supply of air

$$Cu_2S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)} + 2Cu_{(l)}$$

(iv) Purification of copper by electrolysis using copper sulphate solution electrolyte, pure copper cathode and impure copper obtained from the extraction anode

At Anode: $Cu \longrightarrow Cu^{2+} + 2e$

At Cathode: $Cu^{2+} + 2e \longrightarrow Cu$

Uses of copper - As an electrical conductor in copper wires and cables, to make water pipes and boilers, used as a primary component of alloys (brass or bronze), for ornamental work

3.8.5 Environmental consideration

(Not found in exams)

3.9.0 Compounds of metals

3.9.1 Oxides

Oxide - A chemical compound containing at least one oxygen atom as well as at least one other element

Methods of oxide production - Direct, indirect

Direct - The heating of metal directly in air

Indirect - Done by thermal decomposition of carbonates, hydroxides and nitrates (assuming the metal is below in the series

Amphoteric Oxides - React with acids to form normal salts and with bases to form complex salts

Examples of amphoteric oxides - PbO, ZnO, Al_2O_3

Uses of metallic oxides - Preparation calcium carbide (CaC_2), lining of furnaces, formation of slag, drying agent, manufacturing of mortar, manufacturing of cement and plasters

3.9.2 Hydroxides

Hydroxide - A chemical compound containing a hydroxide (-OH) group

Methods of hydroxide preparation -

Direct - Used for soluble hydroxides by the direct combination of metal and water

Indirect - Used for insoluble hydroxides by using a solution of soluble salts with a solution of alkalis

Uses of calcium hydroxide - Cheap laboratory base for preparation of bleaching powder by passing chlorine over moist calcium hydroxide, used in agricultural operations as a liming material, used to manufacture mortar, treatment for hard water, manufacturing of paper

Properties of metallic hydroxides - When heated they decompose (except potassium hydroxide and sodium hydroxide), they form salts and water when reacting with acids

Uses of metal hydroxides - In agriculture it is used to correct soil pH as a liming material, used as a fertilizer, domestic cleaners, softening of water, used in qualitative analysis, making artificial silk, soap, paper and as a bleaching agent

3.9.3 Carbonates and hydrogen carbonates

Carbonate - A chemical compound containing a carbonate (CO_3) group

Natural occurrences of calcium carbonate - Chalk, limestone, marble, shells of eggs and sea animals

Preparation of metallic carbonates -

Insoluble Carbonates - Prepared by a precipitation reaction in a salt solution or a solution of alkali

Soluble Carbonates - Prepared by passing excess carbon dioxide gas through a solution of carbonate

Properties of carbonates - Some are soluble some are insoluble, they react with acids to form salt, water and carbon dioxide, they decompose when heated to give a metallic oxide and carbon dioxide (except for potassium carbonate and sodium carbonate), in cold conditions carbonate reacts with magnesium

Uses of carbonates - Softening of water (ex Na_2CO_3), manufacturing of glass, used in qualitative and qualitative analysis, $NaHCO_3$ is used in baking and the removal of grease

3.9.4 Nitrates

Nitrate - A chemical compound containing a nitrate (NO_3^-) group

Preparation of nitrates - Usually they are prepared by methods involving crystallization by dissolving a metal, carbonate, oxide or alkali in dilute nitric acid

Properties of nitrates - All nitrates decompose on heating to give products according the reactivity series, all are soluble in water, they are hydrated (except sodium and potassium nitrates)

Test for nitrate - Brown ring test

Uses of nitrates - Used as food preservatives, used as fertilizer ($NaNO_3$), used to make photographic film

3.9.5 Chlorides

Chloride - A chemical compound containing a Cl^- ion

Preparation of chlorides - By the action of metals with dilute HCl, from lead and silver salts, by the reaction of HCl and oxides/carbonates/hydroxides, by direct combination of constituent elements

Use of chlorides - Used as a bleaching agent, kills diseases (cholera, typhoid), used to manufacture chemicals, combines with metals to form salts used in home cleaning, salt used in the petroleum industry, used in the manufacturing of dry batteries

3.9.6 Sulphates

Sulphate - A chemical compound containing a sulphate (SO_4^{2-}) group

Preparation of sulphates -

Soluble - By the action of a metal, carbonate, hydroxide, or oxide with dilute sulphuric acid

Insoluble - By double decomposition

Uses of sulphates - $CuSO_4$ is used in plaster casts, $FeSO_4$ is used in the manufacturing of ink, used in the manufacturing of pigments and medicines, used in dyes and making leather

Form 4 Topics

4.1.0 Nonmetals and their compounds

4.1.1 General chemical properties of nonmetals

Physical properties of nonmetals - Usually brittle, bad conductors of heat and electricity (except graphite), physically weak, have low melting and boiling points (except diamonds), they are gases, volatile liquids, or solids at room temperature, usually have low density

Chemical properties of nonmetals - Are electron acceptors (electronegative elements, ex Cl^-, O^{2-}, S^{2-}), they are oxidizing agents (react by gaining electrons to form negative ions), form acidic or neutral oxides, hydrides and chlorides are covalent (CH_4, NH_3, HCl, H_2S), cannot replace hydrogen in an acid to form a salt (like metals do) because nonmetals are electron acceptors

Electronegativity - The ability of an element to attract electrons towards itself

Electrovalent Series - List of elements which go from the highest to lowest electronegativity ($F \rightarrow Cl \rightarrow Br \rightarrow I$)

Effect of an element with higher electronegativity on a compound with an element with lower electronegativity - The element with higher electronegativity will displace the element with a lower electronegativity (the opposite is never true). The higher the electronegativity of a substance, the stronger an oxidant it becomes

Physical properties differentiating metals and non metals -

Metal	Non-metal
• Good conductors of heat and electricity • Can be cut sharp • Shiny surface when cut • Malleable	• Poor conductors of heat and electricity • Generally dull when cut • Generally not shiny when cut • Non-malleable

4.1.2 Chlorine

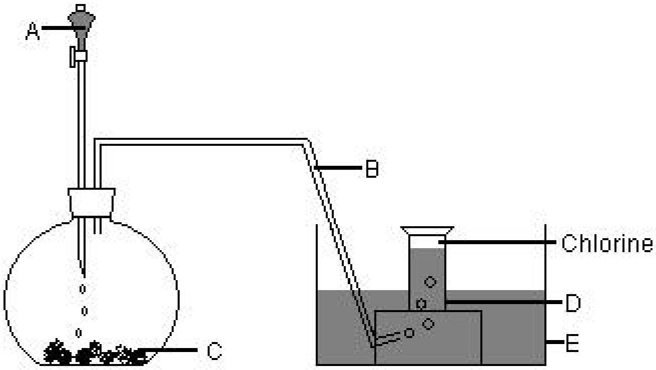
Preparation of chlorine gas - An oxidizing agent like potassium permanganate ($KMnO_4$) with concentrated hydrochloric acid (HCl)

Physical properties of chlorine - Greenish yellow gas, unpleasant choking smell, denser than air, moderately soluble in water, poisonous gas

Chemical properties of chlorine - Bleaching agent, no effect on dry litmus paper but wet litmus paper turns colourless

Uses of chlorine gas - Bleaching agent, kills dangerous germs in water like cholera and typhoid, used to manufacture organic solvents like chloroform and carbon tetrachloride, used as a precursor for insecticides and pesticides

1. The diagram below shows the preparation of chlorine gas in a laboratory fume chamber. Study the diagram and answer the questions which follow



(a) What do the letters A, B, C, D, and E represent?

A - Concentrated hydrochloric acid *B - Delivery tube* *C - Potassium manganate VI*
D - Gas jar *E - Water trough*

(i) Why is the gas prepared in the fume chamber?

The gas prepared in the fume chamber because chlorine is a poisonous gas which is dangerous to inhale

(ii) Can the gas be collected over water? Why?

The chlorine gas can be collected over water because it is not soluble in water. The water in this experiment removes hydrogen chloride gas which is an impurity and not water

(iii) What will happen to a damp blue litmus paper if it is introduced into a gas jar full of chlorine gas?

It will turn the blue litmus paper red and then bleaches it

(iv) What will happen if a gas jar of hydrogen sulphide is inverted over a gas jar of chlorine such as the two gases get mixed? Write a balanced equation for the reaction which takes place between hydrogen sulphide gas and chlorine gas

Hydrogen sulphide gas decolorizes a solution of yellowish-green chlorine. A yellow precipitate of sulphur forms on the jar which liberates hydrogen chloride gas

$$H_2S_{(aq)} + Cl_{2(g)} \longrightarrow S_{(s)} + 2HCl_{(aq)}$$

(b) Give a balanced chemical equation for the method of preparation of chlorine used in this question

$$2KMnO_{4(s)} + 16HCl_{(aq)} \longrightarrow 5Cl_{2(g)} + 2KCl_{(aq)} + 2MnCl_{2(aq)} + 8H_2O_{(l)}$$

4.1.3 Hydrogen chloride

Preparing a sample of dry hydrogen chloride gas - It is prepared by the reaction between concentrated sulphuric acid (H_2SO_4) and sodium chloride (NaCl)

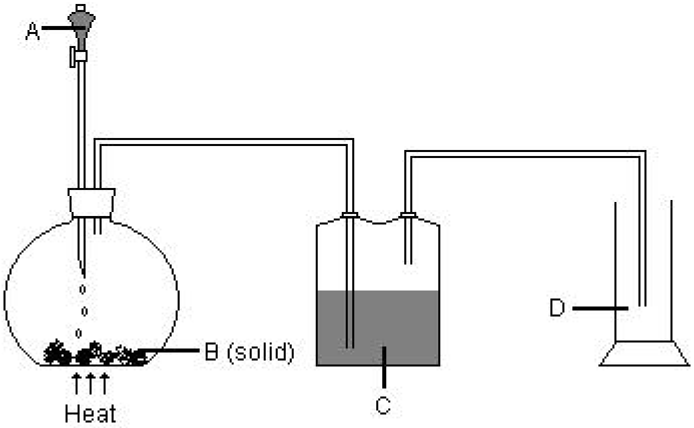
Physical properties of hydrogen chloride gas - Colourless gas, has a choking smell, acidic (turns damp litmus paper red), denser than air, fumes in moist air, very soluble in water

Chemical properties of hydrogen chloride gas - Turns blue litmus paper red, does not support combustion, gives white fumes with ammonia gas, decomposes to liberate hydrogen gas when reacted with certain metals like Zn or Fe, with oxidizing agents it decomposes to give salt, water & chlorine gas

Test for HCl(g) - Reaction with silver nitrate which forms a white precipitate of AgCl

Uses of hydrogen chloride gas - Synthesis of chlorine gas, preparation of hydrogen gas, production of hydrochloric acid, used in qualitative and quantitative analysis

1. (a) This picture represents the laboratory preparation of hydrogen chloride gas. Name the parts labeled A, B, C and D



A - Concentrated sulphuric acid B - Sodium chloride
C - Concentrated sulphuric acid D - Hydrogen chloride gas

(b) (i) Do you think the gas can be collected over water? Give reasons for your answer
The gas cannot be collected over water because it is soluble in water

(ii) Explain the test for this gas
The test for hydrogen chloride gas is the formation of dense white fumes with ammonia gas due to the formation of NH_4Cl

(iii) What is the function of C?
The function of sulphuric acid is to dry the gas since it has a high affinity to water

(iv) Name the method used in the collection of this gas
Downward delivery

(v) Write a balanced chemical equation for the reaction taking place during the preparation of the hydrogen chloride gas

$$NaCl + H_2SO_4 \longrightarrow NaHSO_4 + HCl$$

(c) Write chemical equations for the reaction between
(i) Ammonia gas and hydrogen chloride

$$NH_{3(g)} + HCl_{(g)} \longrightarrow NH_4Cl$$

(ii) Hydrogen chloride gas and water

$$HCl_{(g)} + H_2O_{(l)} \longrightarrow H_3O_{(aq)}{}^{+} + Cl^{-}{}_{(aq)}$$

This is showing the mixing an acid with water forms a diluted aqueous liquid

4.1.3.2 Hydrochloric acid

Properties of $HCl_{(aq)}$ - Is acidic to litmus changing it from blue to red
With carbonates - Salt, water and carbon dioxide are produced
With bases - It forms salt and water only
With metals - It reacts to form salts and hydrogen gas is liberated
With silver nitrate - A white precipitate of silver chloride is formed
With an oxidizing agent - Chlorine gas is produced

Uses of $HCl_{(aq)}$ - Manufacturing of chlorine, cleaning metal surfaces before galvanizing, preparation of soluble chlorides, used in qualitative and quantitative analysis

4.1.4 Sulphur

Allotropy - Different structures of the same element. For example, diamond and graphite, diamonds are bonded tetrahedrally and graphite is bonded in sheets

Allotropes - Different forms of the same element which exist in the same physical state (e.g. diamonds and graphite are allotropes of carbon)

Sulphur- Is a yellow, nonmetallic, naturally occurring element found mainly in the upper surfaces of the earth in Sicily and deep in the earth in the USA

Frasch Process - The industrial process used in the extraction of sulphur based on the use of sulphurs low melting point

Allotropes of sulphur - Rhombic, monoclinic, amorphous (flower of sulphur), colloidal, plastic

Rhombic Sulphur - Made of yellow translucent crystals which is octahedral in structure and changes to monoclinic form after it is heated above 96°C

Monoclinic Sulphur - Consists of needle shaped amber transparent crystals which reverts back to rhombic sulphur below 96°C

Amorphous Sulphur - Produced when a saturated solution of hydrogen sulphide gas in water is left exposed to the air bring about a white powder of amorphous sulphur

Colloidal Sulphur - A yellow suspension formed when dilute hydrochloric acid is added to dilute sodium thiosulphate

Plastic Sulphur - A rubber like form of sulphur formed when boiling sulphur is poured into water, after a few days the plastic sulphur reverts back o the hard rhombic type

Physical properties of sulphur - It is a yellow solid existing in allotropic form, when heated it melts to an amber (reddish brown) liquid, sublimes easily on cooling and when condensed from a vapour forms a deposit called flower sulphur, insoluble in water but soluble in organic solvent, poor conductor of heat and electricity

Chemical properties of sulphur - Combines directly with metals to give metallic sulphide, can react with metals without heating, reacts with nonmetals (ex w/oxygen to form sulphur dioxide or with carbon to form carbon disulphide), burns easily with a blue flame in the presence of air to give sulphur dioxide gas which is acidic, used as a reducing agent

Uses of sulphur - Manufacturing of sulphuric acid, to vulcanize (harden) rubber, the manufacture of matches and gunpowder, manufacturing of chemicals, insecticides, in medicine for ointments and drugs

4.1.5 Sulphur dioxide

Preparation of sulphur dioxide - Heating copper and concentrated sulphuric acid (H_2SO_4) or by reacting sodium sulphate and dilute hydrochloric acid then heating

Physical properties of sulphur dioxide - A colourless gas, has an irritating smell, is readily soluble in water, can be liquefied at ordinary room temperature at 3 atm pressure, it is poisonous to vegetation, it is denser than air so it is observed by downward delivery

Chemical properties of sulphur dioxide - An acidic gas reacting with water to form sulphurous acid which turns blue litmus red, it neutralizes alkalis

Uses of sulphur dioxide - Bleaching of delicate materials, manufacture of sulphuric acid, used as a germicide and fumigant, used for the preservation of juicy fruits, used in refrigerators

Hazards of sulphur dioxide - It is dangerous because it can react with oxygen and water in the body to form sulphurous acid which will destroy tissues in the body

4.1.6 Sulphuric acid

Contact Process - The method of producing sulfuric acid in high concentrations needed for industrial processes

Commercial uses of sulphuric acid - Manufacture of inorganic fertilizers, extraction of metals, to clean metal surfaces, manufacturing synthetic fibers/paints/dyes/explosives, creation of fungicide, manufacture other chemicals like HCl/sulphates/alums/nitric acid, used as an electrolyte in car batteries, used to refine petroleum, use to manufacture detergents and soaps

Properties of sulphuric acid - A dense oily liquid, able to react with bases to form salts, reacts with metals to produce salt and hydrogen gas, reacts with carbonates to form salt, water and carbon dioxide, works as an oxidizing agent, has a high affinity to water (its better to add the acid to the water than the water to the acid), it is used as a drying agent to remove water from various compounds

Uses of sulphuric acid - Used in the manufacturing of fertilizers, manufacturing of paints and pigments, manufacturing of fibres, production of chemicals, used in accumulators, refining of petroleum and extraction of metals, manufacturing of detergents and soap

4.1.7 Nitrogen

Preparation of nitrogen in the lab - Reacting ammonium chloride and sodium nitrate while heating

Importance of nitrogen - Essential to plants and animals for the production of proteins, plants cannot absorb nitrogen directly (only in the form of nitrates)

Industrial preparation of nitrogen - By fractional distillation of liquid air from the atmosphere

Physical properties of nitrogen - Slightly soluble in water, colourless and odorless gas, slightly less dense than air, not poisonous, boils at 195°C

Uses of nitrogen - Used to manufacture ammonia by Haber's process, manufacturing of fertilizer, used as a coolant, manufacturing of nitric acid, manufacturing of plastics

Chemical test for nitrogen - Does not burn and does not support combustion and has no effect with lime water

Nitrogen Fixing - The process whereby free atmospheric nitrogen is made useable by changing it into nitrate and it is absorbed by plants

Ways to fix nitrogen naturally - Bacteria living at the roots of leguminous plants (like beans) have the ability to convert atmospheric nitrogen into nitrates, some bacteria living freely in the soil can also fix atmospheric nitrogen, lightning can cause atmospheric nitrogen to combine with oxygen to form nitrogen oxide which then dissolves in rain water to form nitric acid which is converted to nitrates in soil

Nitrifying Bacteria - Oxidize ammonia and convert it into nitrates

Denitrifying Bacteria - Convert ammonia into free nitrogen (N_2)

Nitrogen gas is the most abundant gas in the air and is inert and will not react with metals under normal conditions

4.1.8 Ammonia

Haber Process - Used in the industrial preparation of ammonia and is represented by the equation

$$N_{2(g)} + 3H_{2(g)} \overset{catalyst}{\leftrightarrow} NH_{3(g)} \quad \Delta H = -46.2kJ/mol$$

Catalyst used in the production of ammonia - Iron oxide from magnetite

Properties of ammonia - A colourless, tasteless gas with a pungent choking smell, less dense than air (collected by upward delivery), extinguishes a burning splint, it is the only alkaline gas which turns red litmus to blue

Uses of ammonia - Fertilizer, softening of temporary hard water, cleaner, precursor for explosives, refrigeration, antimicrobial, source of nitrogen for living systems

4.1.9 Carbon

Allotropes of carbon - Diamond, graphite

Diamond - Very hard, used in drilling machines and glass cutting

Graphite - Very soft, black and greasy, used as lubricant, non-metal which conducts electricity

Similarities between diamond and graphite - Forms crystalline in nature, both are pure carbon (formed entirely of carbon atoms), chemically similar (e.g. when burnt, they release CO_2)

Differences between diamond and graphite -

Property	Diamond	Graphite
Appearance	Colourless, transparent, glittering	Black with metallic luster
Density	3.51 g/cm^3	2.25 g/cm^3
Hardness	Hardest known element	Not very hard
Conductivity	Does not conduct electricity	Conducts electricity like metals
Bonding	Strong C-C covalent bonds arranged octahedrally to form a giant molecular crystal	Strong C-C bonds in hexagonal rings in sheets with weak forces between the layers
Cleavage	Difficult, along octahedral boundaries	Easy, along sheets
Method of synthesis	Prepared from graphite at very high pressure and temperature	Prepared from coke and silica mixture at a high temperature
Uses	Drills, precious stone, cutting glass, grinding	Lubricant, pencils, electrode in electrolytic reactions, moderator in nuclear reactors, converted under pressure to synthetic diamonds

Effect of carbon monoxide in the blood - Carbon monoxide binds to the iron in hemoglobin in red blood cells to form carboxyl which prevents the hemoglobin from taking oxygen, eventually leading to death

4.1.10 Carbon dioxide

Test for carbon dioxide (CO_2) - When it is bubbled through calcium hydroxide (lime water) it causes small chalk particles to suspend in the water causing a milky appearance

4.1.11 Nonmetals questions

1. The industrial preparation of ammonia in the Haber process is represented by:

$$N_{2(g)} + 3H_{2(g)} \xleftrightarrow{Catalyst} NH_{3(g)} \quad \Delta H = -46.2kJ/mol$$

What will happen if?

(i) The temperature of the equilibrium mixture is increased?

The equilibrium will shift to the left because reactants are favoured by an increase in temperature for exothermic reactions

(ii) More nitrogen gas is added to the equilibrium mixture?
The equilibrium will shift to the right and more NH_3 will be produced, since an increase in reactants will create more products

(iii) The formed ammonia gas is removed from the equilibrium mixture?
The equilibrium will shift to the right and compensate for the NH_3 that was removed in following Le Chatelier's principle

2. (a) Write down the balanced equations for manufacturing sulphur dioxide (SO_2) from
(i) Sulphur

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)}$$

(ii) Hydrogen sulphide

$$H_2S_{(g)} + H_2SO_{4(l)} \longrightarrow 2H_2O_{(l)} + S_{(s)} + SO_{2(g)}$$

(iii) A sulphide ore

$$4FeS_{2(g)} + 11O_2 \longrightarrow 2FeO_{3(s)} + 8SO_{2(g)}$$

(b) Write down a balanced chemical equation for the combustion of sulphur dioxide

$$2SO_{2(g)} + O_{2(g)} \xleftrightarrow{PtCatalyst,500C} 2SO_{3(g)}$$

4.2.0 Organic Chemistry

4.2.1 Introduction to organic chemistry

Isomer - Compounds with the same molecular formula but different structural formulas
Isomerism - The occurrence of two or more compounds with the same molecular formula but having different structural formulas
Monomer - An atom or small molecule that can form longer chains called polymers
Polymerization - The process of combining monomer molecules together in a chemical reaction to form 3 dimensional chains called polymers
Aldehydes - An organic compound containing a formyl group with the structure R-CHO
Amine - Organic compounds and functional groups that contain a basic nitrogen atom with a lone pair which are derivatives of ammonia where one or more hydrogen atoms has been replaced by an alkyl or aryl group
Cracking - The decomposition of complex organic compounds into simpler ones
Functional Group - Specific groups of atoms within molecules that are responsible for the characteristic chemical reactions of those molecules
Lubricating Oils - Oils which reduce the friction between two moving surfaces to improve efficiency and reduce wear (breakdown)
Petroleum Products - Asphalt, diesel, lubricating oils, paraffin wax, gasoline, kerosene, petrochemicals, tar, fuel
Viscosity - The measure of the resistance of a fluid which is thought about in terms of thickness of thinness (ex water is 'thin', honey is 'thick')

4.2.2 Hydrocarbons

Hydrocarbon - Is an organic compound consisting entirely of hydrogen and carbon atoms. Hydrocarbons produce CO_2 (carbon dioxide), H_2O (water) and energy when burned
Homologous Series - Is a series of compounds which conform to a general formula, undergo chemical reactions, physical properties change gradually in the same direction in the series, general methods of preparation can be applied to any member in the series and differ in composition by CH_2
Hydrogenation - Is a chemical reaction between molecular hydrogen (H_2) and another compound or element. Commonly used to reduce or saturate organic compounds like alkanes or alkynes (ex ethyne (C_2H_2) reacts with H_2 to form C_2H_6 since two of the three triple bonds are broken, freeing up four spaces for hydrogen atoms, resulting in the formation of an alkane from an alkyne through hydrogenation)

$KMnO_4$ (potassium permanganate) is used to distinguish saturated and unsaturated hydrocarbons
Nickel is used as a catalyst in the hydrogenation of oils
Methane is found in natural gas

4.2.3 Properties of hydrocarbons

Types of hydrocarbons - Saturated, unsaturated
- **Saturated Hydrocarbons -** Organic compounds containing no double or triple bonds (alkanes)
- **Types of saturated hydrocarbons -** Alkanes
  - **Alkanes -** Are hydrocarbons whose atoms are linked together exclusively by single bonds. They are saturated hydrocarbons. Methane (CH_4) is the simplest alkane. Chain alkanes follow the homologous series C_nH_{2n+2}
- **Unsaturated Hydrocarbons -** Are a type of hydrocarbon with double or triple bonds which means that they can undergo additional reactions (alkenes/alkynes)
- **Types of unsaturated hydrocarbons -** Alkenes, alkynes
  - **Alkenes -** Are hydrocarbons which contain at least one double carbon-carbon bond. They are unsaturated hydrocarbons. Ethylene (C_2H_4) is the simplest alkene. Follow the homologous series C_nH_{2n}
  - **Alkynes -** Are hydrocarbons which contain a triple carbon-carbon bond. They are unsaturated hydrocarbons. Ethyne (acetylene C_2H_2) is the simplest alkyne. Follows the homologous series C_nH_{2n-2}

4.2.4 Alcohols

Alcohols - Any organic compound in which a hydroxyl function group (-OH) is bound to a carbon atom. They follow the general formula $C_nH_{2n+1}OH$. The simplest alcohol is methanol (CH_3OH)

Fermentation - An anaerobic process of respiration which involves the breakdown of carbohydrates into alcohols, carbon dioxide and energy

Ester - Chemical compounds formed by condensing an alcohol with an acid
- **Esterification -** A chemical reaction where an alcohol and an acid form an ester as a product
- **Uses of esters -** Used as fragrances or oils, backbone of DNA, creation of explosives, plastics

Converting an alkane to an alcohol - One of the hydrogen atoms is replaced by an OH group (ex CH_4 becomes CH_3OH)

When beer or wine is exposed to the air, alcohol is converted into acetic acid

4.2.5 Carboxylic Acids

Carboxylic Acid - Are organic compounds with at least one carboxyl group (-COOH) with a general formula of R-COOH, where R is a functional group. The general formula is $C_nH_{2n+1}COOH$

Soap - The salt of fatty acids

4.2.6 Organic chemistry questions

1. The empirical formula of a certain compound is CH_3. Its vapour density is 30. What is its molecular formula?

Vapour density = 30

$$Relative\ Molecular\ Mass = 2 \times Vapour\ Density$$

Step 1: Calculate relative molecular mass of CH_3 $Relative\ Molecular\ Mass = 2 \times 30$ $Relative\ Molecular\ Mass = 60$ **Step 2: Calculate n** C = 12, H = 1	n(CH_3) = RAM n(CH_3) = 60 n(15) = 60 n = 4 To find the final molecule you need to multiply the CH_3 by 4, 4(CH_3),which gives **C_4H_{12}**

2. By naming the reagents, stating the conditions whenever possible and writing a balanced equation for each, describe how ethene could be converted into:

(i) Ethane - **By hydrogenation**

$$CH_2 = CH_2 + H_2 \xrightarrow{Ni,Catalyst} CH_3CH_3$$

(ii) Chloroethane - **By reacting it with HCl gas**

$$CH_2 = CH_2 + HCl \longrightarrow CH_3CH_2Cl$$

(iii) 1, 2-dibromoethane - **By bromination**

$$CH_2 = CH_2 + Br_2 \longrightarrow CH_2BrCH_2Br$$

(iii) Ethanol - **By reaction with HCl gas and reaction with water**

$$CH_2 = CH_2 + HCl \longrightarrow CH_3CH_2Cl$$

$$CH_3CH_2Cl + H_2O \longrightarrow CH_3CH_2OH + HCl$$

3. Briefly explain how you would carry out the following conversions
Ethene → Ethane → Ethanol → Acetaldehyde → Acetic acid

Ethene → Ethane - **By hydrogenation**

$$CH_2 = CH_2 + H_2 \xrightarrow{Ni,Catalyst} CH_3CH_3$$

Ethane → Ethanol - **By chlorination**

$$CH_3CH_3 + Cl_2 \xrightarrow{uv\ light} CH_3CH_2Cl + HCl$$

Ethanol → Acetaldehyde - **By adding water**

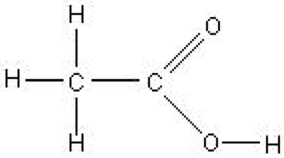
$$CH_3CH_2Cl + H_2O \longrightarrow CH_3CH_2OH + HCl$$

Acetaldehyde → Acetic acid - **Using oxidizing agent like potassium dichromate**

Step 1: $CH_3CH_2OH + K_2Cr_2O_7 \longrightarrow CH_3CHO$

Step 2: $CH_3CHO + K_2Cr_2O_7 \longrightarrow CH_3COOH$

4. (a) Name the compound formed and write its open structure formula
This is an acid because it follows the general formal of acids which is $C_nH_{2n}O_2$



(b) Write balanced chemical equations of the reactions between CH_2O and the following substances

(i) Sodium metal

$$2CH_3COOH + 2Na \longrightarrow 2CH_3COONa + H_2$$

(ii) Ethanol

$$CH_3COOH + CH_3CH_2OH \longrightarrow CH_3COOCH_2CH_3 + H_2O$$

(iii) Sodium hydroxide

$$CH_3COOH + NaOH \longrightarrow CH_3COONa + H_2O$$

4.3.0 Soil Chemistry

4.3.1 Soil Formation

**Found in the Geography study guide*

Soil Formation - Creation of soil through the weathering of parent material
Soil Structure - The arrangement of soil particles within the soil
Flocculation - Binding together of soil particles

4.3.2 Soil reaction

Soil Reaction (pH) - The degree of acidity or alkalinity of the soil

- **Causes of soil reaction -** Thunderstorms and lightning, decomposition of organic matter, amount of mineral element present in soil solution, weathering, too much rainfall causing leaching, addition of fertilizers, liming, bush burning, overcropping, irrigation
- **Importance of soil reaction -** Solubility of different elements depends on the pH of the soil, determines the availability of plant nutrients, costs money to return to normal, influences enzyme activities in the soil, can discourage diseases, enables farmers to select acid or base tolerant crops

Liming - The process of adding basic compounds in acidic soil to reduce its acidity

States of agricultural soil - Active, potential

- **Active Acidity -** A state of agricultural soil which has a higher concentration of H^+ ions in the soil solution than in the soil particles
- **Potential Acidity -** A state of agricultural soil which has a lower concentration of H^+ ions in the soil solution than in the soil particles

Causes of soil acidity - Leaching of bases by irrigation or heavy rainfall, dissolving of carbon dioxide to form carbonic acid, application of acid forming fertilizers like sulphate of ammonia

- **Disadvantages of soil acidity -** Affects plant growth by harming the roots, prevents the availability of essential elements, increases the solubility of ions to toxic levels, prevents activity of microorganisms
- **Advantages of soil acidity -** Enables farmers to select acid tolerant crops, may discourage disease

Causes of soil basicity - Availability of cations, overliming, low amount of microbes in the soil

- **Advantages of soil alkalinity -** Enables formers to select suitable crops for basic soil (ex coffee, tea), favors growth of most crops due to high content of available cations
- **Disadvantages of soil alkalinity -** Availability of too many bases may lead to toxic levels in the plant

4.3.3 Plant nutrients in soil

Ways atmospheric nitrogen is transformed into usable forms for plants -

- **Nitrogen Fixation -** A process where atmospheric nitrogen is converted into nitrates or ammonium ions in the soil by bacteria in the nodules of leguminous plants
- **Mineralization -** Conversion of organic nitrogenous compounds into inorganic forms by microorganisms
- **Nitrification -** The chemical way that nitrifying bacteria use ammonium as a source of energy which oxidizes it into nitrates which are useable by plants
- **Lightning -** Lightning causes atmospheric nitrogen to combine with oxygen to form nitrogen oxide, which is then dissolved by rain water and added to the soil

Functions of various elements in plants - N, P, Zn

- **Nitrogen -** Synthesis of proteins/nucleic acids/chlorophyll, root growth and cation exchange, promotes plant growth and use of carbohydrates
- **Phosphorus -** Synthesis of nucleic acids/ATP, disease resistance, root growth, increases grain yield for cereals, speeds up crop maturation
- **Zinc -** Component of enzymes responsible for anaerobic respiration

Types of elements needed by plants - Macronutrients, micronutrients (trace elements)

- **Macronutrients** - Elements in the soil needed by plants in a large amount
  - **Examples of macronutrients in plants -** Nitrogen, phosphorus, potassium, sulphur
- **Micronutrients (Trace Elements) -** Elements in the soil needed by plants in a small amount
  - **Examples of trace nutrients in plants -** Zinc, chlorine, boron, copper, iron

Denitrification - A microbially facilitated process of nitrate reduction that ultimately produces molecular nitrogen (N_2) which is released into the atmosphere

Denitrifying Bacteria - Add nitrogen to the atmosphere by converting nitrates in the soil into free nitrogen

Symbiotic Bacteria - Bacteria which live in a host and they depend on each other for nutrition (e.g. nitrogen fixing bacteria living symbiotically with leguminous plants)

4.3.4 Manures and fertilizers

Nitrogenous straight fertilizers - Sulphate of ammonia ($(NH_4)_2SO_4$ or SA), ammonium sulphate nitrate ($(NH_4)_2SO_4$ + NH_4NO_3 or A.S.N), urea ($CO(NH_2)_2$), Calcium ammonium nitrate ($CaCO_3$ + NH_4NO_3 or C.A.N)

- **Sulphate of Ammonia -** Consists of small white crystals, contains about 21% nitrogen and 23.5% sulphur, is a hygroscopic and deliquescent compound, if applied repeatedly to soil it will become acidic due to hydrolysis
- **Ammonium Sulphate Nitrate -** Mixture of two salts (ammonia sulphate and ammonium nitrate), consists of yellowish granules, about 26% sulphur, not strongly acidic like ammonium sulphate but is more hygroscopic
- **Urea -** Consists of white crystals, contains about 46% nitrogen, soluble and hygroscopic, does not damage plant leaves so it can be applied to plants, easily taken up by plant roots through leaves, causes slight acidity to the soil if applied repeatedly
- **Calcium Ammonium Nitrate -** Mixture of two chemicals (Ammonium nitrate and calcium carbonate), consists of gray granules, contains about 20% nitrogen, is hygroscopic, does not cause acidity in the soil

Advantages of artificial fertilizers - Chemical composition is constant, easy to apply, nutrient elements are balanced in relation to each other, easy to store, easy to absorb, contain nutrients in concentrated form

Disadvantages of artificial fertilizers - Do not assist in conservation and building of organic matter, do not improve soil structure, do not conserve soil nitrogen, do not encourage microbial activity, do not increase soil temperature, alters pH easily, does not contain all elements in a single fertilizer

Effects of artificial fertilizers - Encourage rapid growth of algae in water bodies causing a lack of oxygen in the water for fish, can enter the water supply from water runoff and poison humans

Manure - A fertilizer which contains the decayed organic matter from animal wastes or plant remains

Types of organic manure - Kraal manure (waste product from animals), farm yard manure (mixture of grasses and animal wastes), poultry manure, composite manure, green manure

Advantages of manure - Conserves nitrogen in the soil, improves soil structure through humus formation, almost all kinds of nutrients are available, encourages microbial activity, increases soil temperature, assists in the conservation of organic matter, makes the soil fertile and porous

Disadvantages of manure - Has high moisture content making them heavy for transportation, have indefinite chemical compositions, nutrients are supplied in a low amount, nutrients are unbalanced in relation to each other, sometimes manure must be supplemented by artificial fertilizers,

Advantages of organic manure over artificial fertilizers - Improves soil structure through humus formation, improves soil porosity/aeration/drainage, encourages conservation of soil nitrogen, supplies all kinds of plant nutrients, encourages microbial activities which assist in the mineralization of nutrients, prevents leaching

Methods of fertilizer application - Broadcasting, top dressing, side dressing, banding, in rows, plough-sole

Broadcasting - Used in closely spaced crops where nutrients are required in the narrow top soil zone

Top Dressing - The application of fertilizers uniformly over the soil after crop emergence

Side Dressing - Application of fertilizer beside the row of crops

Banding - The placement of fertilizer below or to the side of the seed or plant

In rows - Seeds are placed directly to on to the fertilizer

Plough-sole - Fertilizer is allowed to fall into the ground when the plough is ploughing the field

4.3.5 Soil fertility and productivity

Soil Fertility - The status of a soil with respect to the availability of nutrients

Why fertile soil is not always productive - Climate of the area (rainfall/temperature), crop types, soil texture, technology used, soil structure, poor farming methods

Ways to improve soil fertility - Crop rotation, reafforestation, renewal of humus, use of compost/manure, proper use of artificial fertilizer

Effects of excessive nitrogen to plants - Delays plant maturity, weakens plants leading to lower quality and yield, reduces sugar content

Factors affecting soil fertility - Soil erosion, burning, water lodging causing leaching, monoculture, harvesting

Ways to improve soil fertility - Adding of manure, adding sand or clay, tilling the soil, crop rotation, weed control, mulching, good farming/harvesting practices, soil erosion control

Soil Erosion - Is the shifting of topsoil from one place to another

Ways to control soil erosion - Use of organic manures, intercropping, crop rotation, mulching, afforestation, destocking

Why fertile soil may not be productive - All of the following can negatively effect soil: climate of the area, type of crop grown, soil structure of the soil, soil texture, technology employed, pests and diseases

Most commonly used nitrates by plants - Ammonium (NH_4^+) or nitrate ions (NO_3), since they are easily taken up by roots

Why nitrogen gas is not directly available to plants - Very few organisms are capable of assimilating free nitrogen to make it available to plants as nitrate ions. Leguminous plants use a species called rhizobium to fix atmospheric nitrogen

Leaching - The process where nutrients are removed from the soil by being dissolved in water

Mulching - Practice of covering the soil with a layer of plant remains to conserve soil moisture

Terracing - The creation of step-like landforms on the sides of mountains to capture water and increase land size for farming

Chemicals which neutralize soil acid - Lime (calcium carbonate), magnesium carbonate, calcium oxide, ashes from charcoal and wood, silicates of calcium and magnesium, pulp from paper manufacturing

Chemicals which neutralize basic soil - Sulphur (S_2), calcium sulphate, gypsum

4.3.6 Soil questions

1. Calculate the percentage composition of nitrogen in ammonium sulphate fertilizer $(NH_4)_2SO_4$

To find the percentage composition of an element in a given compound, you need to first find the masses of each element, and then you find the total mass of the compound and divide that total mass by the mass of the element in the question. This will give you the percent composition

Step 1: Find masses of each element in the compound

N = 14, H = 1, S = 32, O = 16

Step 2: Multiply their masses by the number of

28 + 8 + 32 + 64 = 132 g/mol

Step 4: Take the mass from step 2 of the element in the question and divide it by the total mass of the compound and multiply by 100 to get a

atoms present in the compound	**percentage**
N = 14 x 2 = 28 H = 4 x 1 x 2 = 8 S = 32 O = 16 x 4 = 64	$\frac{28}{132}\times 100 = 21.21\%$
Step 3: Sum these to get the total mass of the compound	*The percentage composition of N is* ***21.21%***

4.4.0 Pollution

4.4.1 Concept of pollution

**The Geography study guide has more information about pollution and environmental conservation*

Pollution - The act of making the air, soil and water unfit for use

Types of pollution - Water pollution, air pollution, land (terrestrial) pollution

Pollutants - Any substance produced and released into the environment that makes it unfit for use

Causes of water pollution - Detergents, sewage, farm runoff, industrial waste, oil

4.4.2 Terrestrial pollution

Effects of farming of on the environment - Erosion, destroys bacteria in the soil, overgrazing, deforestation, chemicals, artificial fertilizers tend to increase soil acidity, monoculture destroys fertility

Terrestrial Pollution - The process of dumping garbage and waste materials on the land making the land dirty or uninhabitable

Materials which contribute to terrestrial pollution - Plastic bags, buckets, glass, cans, industrial waste, plastic wrapping

Causes of terrestrial pollution - Littering, agrochemicals (e.g. DDT, chlorinated hydrocarbons), toxic waste from industry or hospitals

Methods of preventing terrestrial pollution - Use of biodegradable insecticide, returning old bottles and glass containers for recycling and reusing, construction of free flowing sanitation systems, control of toxic waste dumping

4.4.4 Aerial Pollution

Main sources of aerial pollution - Industry, internal combustion engines (cars/lorries)

Substances causing air pollution - Dust, smoke, sulphur dioxide, carbon monoxide, carbon dioxide, lead

Acid Rain - Sulphur dioxide and oxides of nitrogen from cars or industry form acids in clouds which rain damaging buildings and wildlife

Gases which cause acid rain - Carbon dioxide (CO_2), sulphur dioxide (SO_2), nitrogen dioxide (NO_2)

Effects of acid rain on the environment - Increases soil acidity which reduces soil fertility, kills trees and vegetables

4.4.5 Environmental Conservation

Biodegradable Substances - Substances which can be broken down in the environment by bacteria or naturally occurring chemicals

Importance of environmental conservation - Plants can be used as medicine, biological control of pests to protect food supplies, beauty of the natural environment, tourism, protection of wildlife, reduction of the rate of loss of natural resources, natural resources can be renewed

4.4.7 Ozone layer destruction

Chlorofluorocarbons (CFC) - Gases which cause the destruction of the ozone layer

Greenhouse Effect - Refers to the role that the atmosphere plays in the warming of the Earth's surface by preventing radiation from leaving Earth's surface. Gases in the atmosphere reflect energy back to the Earth instead of letting it go out into space, thereby increasing the temperature of the Earth. The main gas which causes this is carbon dioxide (CO_2)

Effect of ultraviolet radiation (UV) on living organisms - Causes mutation leading to changes in the phenotype of an organism (UV is a mutagen), it is a carcinogen (causes cancer), kills unicellular organisms like bacteria

Function of ozone (O_3, ozone layer) in the atmosphere - Prevents the penetration of harmful UV radiation from the sun onto the Earth's surface

Ways to prevent ozone layer destruction - Proper handling of waste gases, planting of trees which use CO_2 gas produced by industries or automobiles, reducing number of harmful practices used by industries, building of more efficient cars, use of renewable energy (e.g. wind power)

Effects of destroying the ozone layer - Allows harmful UV radiation to reach the earth which can cause skin burns or cancer

Examples of nonmetallic oxides which cause pollution - Carbon dioxide (CO_2), sulphur dioxide (SO_2)

4.5.0 Qualitative Analysis

4.5.1 Qualitative analysis

What will happen if -

Yellow flowers are placed in a container with chlorine gas - The yellow colour will disappear due to the bleaching effects of chlorine

A glass rod which was dipped in concentrated HCl is introduced into a jar with ammonia gas - Dense white fumes will be seen due to the reaction of $NH_3 + HCl \rightarrow NH_4Cl$

Sulphur dioxide gas is bubbled through a yellow acidified potassium dichromate solution - The yellow colour of dichromate will change green due to the reduction of Cr^{6+} to Cr^{3+}

A dilute solution of calcium chloride and sodium carbonate are mixed - A white precipitate of calcium carbonate is formed

An iron earring dropped into a container of copper sulphate solution - Iron will displace copper from its sulphate producing $FeSO_4$ which will change the solution from blue to green

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_4 + Cu$$

$$\text{Blue} \qquad\qquad \text{Green}$$

A copper knife is dipped into a zinc nitrate solution - Nothing will happen because copper cannot reduce zinc since it is at a higher position in the electrochemical series

Copper turnings are dropped into a container of dilute hydrochloric acid - There will be no reaction because copper cannot replace hydrogen in acids since it is below hydrogen in the reactivity series

Explain the following with equations -

White anhydrous copper (II) sulphate changes its colour to blue when water is added - This happens because hydrated copper (II) sulphate is formed which is the colour blue

$$CuSO_{4(s)} + 5H_2O_{(l)} \rightarrow CuSO_4 5H_2O$$

$$\text{White} \qquad\qquad \text{Blue}$$

Vigorous reaction takes place when a small piece of sodium metal is placed in water - Sodium metal is highly reactive with water because the sodium ionizes by electron loss and displaces the hydrogen forming NaOH and liberating hydrogen gas. Since the difference in reactivity between hydrogen and sodium metal is so high, the reaction is violent

$$2Na + 2H_2O \rightarrow 2NaOH + H_2$$

Addition of Zn metal to a solution of copper (II) sulphate results in the decolourization of the solution and deposition of a solid brown substance - Caused by the displacement of Cu by Zn since Zn is at a higher position than Cu in the electrochemical series (meaning Zn is more reactive than Cu). This causes the formation of $ZnSO_4$ which is colourless. The brown substance is the copper

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

Why freshly prepared nitric acid changes from being colourless to a yellowish brown after standing - When concentrated HNO_3 is left open it will decompose to NO_2 which is a yellowish brown colour

$$4HNO_3 \xrightarrow{air, 25C} 4NO_2 + 2H_2O + O_2$$

A mixture of ammonia chloride and sodium hydroxide solution is heated - Causes a choking smell due to the formation of ammonia

$$NH_4Cl + NaOH \xrightarrow{heat} NH_3 + NaCl + H_2O$$

Ammonium chloride crystals are heated - Causes the formation of ammonia gas

$$NH_4Cl \xrightarrow{heat} NH_3 + HCl$$

Ammonium nitrate crystals are heated - Nitrogen oxide and water forms

$$NH_4NO_3 \xrightarrow{heat} N_2O + 2H_2O$$

Ammonium sulphate crystals are heated - Ammonia sulphate crystals decompose into ammonia and ammonium hydrogen sulphate

$$(NH_4)_2SO_4 \xrightarrow{heat} NH_3 + NH_4HSO_4$$

Hydrochloric acid (HCl) is added to a test tube containing carbonate - Effervescence occurs due to the formation of carbon dioxide gas

$$Na_2CO_3 + 2HCl \longrightarrow 2NaCl + H_2O + CO_2$$

Ammonium sulphate crystals and sodium hydroxide solution are warmed together - It will smell bad due to ammonia gas

$$(NH_4)_2SO_4 + 2NaOH \xrightarrow{warm} Na_2SO_4 + 2H_2O + 2NH_3$$

Lead nitrate is heated in a test tube - A brown gas (nitrogen dioxide) and yellow residue (lead oxide) forms $2Pb(NO_3)_2 \xrightarrow{heat} 2PbO + 4NO_2 + O_2$

Identify the following substances by the given information -

A solid which is yellow when hot and white when cold - **ZnO**

When water is added to a white powder and heated, the white powder changes to blue crystals - **Anhydrous copper (II) sulphate**

An aqueous solution of greenish crystalline sulphate forms a pale green precipitate with sodium hydroxide solution and turns brown on standing when exposed to air - **Iron (II) sulphate**

A colourless gas becomes brown on exposure to the air - **Nitrogen monoxide (NO)**

A pale green solution which becomes yellowish brown on exposure to air - **Iron (II) chloride ($FeCl_2$)**

Colourless liquid, immiscible with water, dissolves sulphur - **Carbon disulphide (CS_2)**

A heavy liquid metal - **Mercury, Hg**

Explodes with a pop sound when lit by a fire - **Hydrogen**

Reignites a glowing splint - **Oxygen**

Reddish brown when hot and yellow when cold - **Lead oxide**

4.5.2 Qualitative analysis questions

1. Compound M was mixed with dilute HCL acid to form effervescence with gas Y and a solution of compound X. Normally when compound M is strongly heated forms gas Y and a solid residue W, then when residue W dissolved in warm water forms solution B. When gas Y is allowed pass through solution B for two to three minutes, solution B turned to milky colour. Identify by names compound B, M, X, W and gas Y.

Gas Y is CO_2 because the reaction between a metallic carbonate and acids results in effervescence due to the evolution of CO_2 gas

Compound B is calcium hydroxide because of the reaction between carbon dioxide and calcium hydroxide forms a solution with a milky colour

Compound M is calcium carbonate because the reaction of compound M with dilute HCL causes effervescence and because when it was heated it formed a metallic oxide Q (because it is a solid residue and gas was given out and when the residue was dissolved in water, calcium hydroxide was formed)

Compound X is calcium chloride because the reaction between calcium carbonate and HCl causes the formation of the calcium chloride, water and carbon dioxide

Compound W is calcium oxide because the decomposition of calcium carbonate forms calcium oxide and CO_2 gas