1. Experimental Chemistry

1.1 Experimental Design

(a) name appropriate apparatus for the measurement of time, temperature, mass and volume; including burettes, pipettes, measuring cylinders and gas syringes

Time: stopwatch (s), Mass: Electronic balance (kg, g), Temperature: Alcohol thermometer (k), Volume: Pipette (accurate volume 10.0cm3), Volumetric flask (large accurate volume 100cm3), Measuring cylinder (nearest 0.5cm3), Burette (nearest 0.05cm3)

(b) suggest suitable apparatus, given relevant information, for a variety of simple experiments, including drying and collection of gases and measurement of rates of reaction
Water displacement: Insoluble/slightly soluble (h2, o2, co2), Downward delivery: Denser than air (cl2, hcl, so2). Upward delivery: Less dense than air (nh3)

Concentrated sulfuric acid: Remove water from acidic gas, Quicklime [freshly heated before use]: Remove water from alkaline gas, Calcium chloride [freshly heated before use]: Neutral gas (neither acidic nor alkaline)

1.2 Methods of Purification and Analysis

(a) describe methods of separation and purification for the components of mixtures, to include:

(i) use of a suitable solvent, filtration and crystallization or evaporation

Use of suitable solvent: Separate solids of different solubility (s-s mixture, only one solid is soluble in solvent used)

Filtration: Separate insoluble solids from liquids (Passes through filter paper is filtrate, remains on filter paper is residue)

Crystallization: If solute is easily decomposed, this method is used. Obtaining pure solid from saturated solution. (saturated is when no more solute can be dissolved.) Cool saturated solution until crystals appear. Pour mixture to funnel to collect crystals. Crystals are washed with cold water to remove impurities and dried between 2 sheets of filter paper.

Evaporation: Heat cause water to be evaporated, leaving salt behind. Substance with lower bp will turn into gas first, leaving another behind. Heat the solute till a solid is left behind.

(ii) Sublimation

Sublimation: Separate a substance that changes from solid to gaseous state directly.

(iii) distillation and fractional distillation (see also **11.1**(b))

Distillation: Separate pure solvent from a solution. Relies on the diff in bp (l-g)

Salt water heated and vapor rises through the condenser and exit to the conical flask. Water vapor cools in the condenser and condenses back to the liquid. Pure water is collected as distillate. As more water vaporizes, the salt concentration gets more concentrated and a solid residue remains in the distillation flask .

Boiling chips: Smooth boiling

Fractional distillation: Separate miscible liquids with diff bp. When solution is heated, both ethanol and water vapor rises and the water with a higher bp will condense on the cool surface of the fractionating column and returns back to the liquid. Hot ethanol will cool and condense into liquid in the condenser and flow into conical flask.

Glass bead: Provide large surface for vapor to condense.

Thermometer: Measure temperature (temp stays constant during boiling)

Condenser: Condense vapor, water enters from the bottom (goes against gravity to make it more efficient)

(iv) use of a separating funnel

Separating funnel: separate immiscible liquids, each component can be removed by opening the tap at the bottom and collect in diff beakers.

(v) paper chromatography

Chromatography: Used to separate a mixture of substances which have different solubilities in given solvent.

As solvent is absorbed by paper, it travels towards opp edge. More soluble travel more rapidly towards solvent front and the less soluble will more at a slower rate. (when they travel at diff speed, they get separated.)

Rf is the distance travelled by substance/distance travelled by solvent (ratio of distance)

Diff solvent have diff Rf

Starting line drawn with pencil: Pencil is graphite and will not be separated.

Small sample: Prevent spots from overlapping

Solvent front as high as possible: components separated fully

Solvent front do not reach the edge of paper: Allow actual distance to be measured

(b) suggest suitable separation and purification methods, given information about the substances involved in the following types of mixtures:

(i) solid-solid

(diff in solubility, size, magnetism, state)

(ii) solid-liquid

(diff in solubility, bp)

(iii) liquid-liquid (miscible and immiscible)

(diff in bp. Immiscible, rf)

(c) interpret paper chromatograms including comparison with 'known' samples and the use of R_f values

Rf can compare chromatograms with known substance under the same temperature and solvents. If solvent is diff. the rf of the same substance will change.

(d) explain the need to use locating agents in the chromatography of colorless compounds

(knowledge of specific locating agents is **not** required)

They are chemicals which reacts with colorless substance to from colored spot.

Examine under UV light will also work.

(e) deduce from given melting point and boiling point data the identities of substances and their purity

Pure substances has a specific mp or bp under fixed boiling point.

Mixtures melt/boil over a range of temperature. The greater the impurities, the larger the change in mp/bp.

Identify purity with chromatography, melting, boiling point.

(f) explain the importance of measuring the purity in substances used in everyday life, e.g. foodstuffs and drugs.

Impurities in medicines may cause harmful side effects. Silicon in computer chips need to have a certain purity in order to function effectively.

2. The Particulate Nature of Matter

2.1 Kinetic Particle Theory

(a) describe the solid, liquid and gaseous states of matter and explain their interconversion in terms of the kinetic particle theory and of the energy changes involved

Solid state: Closely packed in an orderly arrangement. Vibrate or rotate about their fixed position and have a low KE. The particles experience strong attractive forces, so a lot of energy is needed to break the regular arrangement. Have a definite volume and shape.

Liquid state: Their arrangement is closely packed in an disorderly arrangement. The particles slide past one another freely. The particles spacing in liquids is larger than solid that the attractive forces are less strong. Liquid has no definite volume but no definite shape.

Gaseous state: The particles are found far apart in an disorderly manner. The particles can move quickly and randomly in any direction. The particles spacing are large that the attractive force between them is weak.

(b) describe and explain evidence for the movement of particles in liquids and gases

(the treatment of Brownian motion is **not** required)

Diffusion: It is the net movement of particles from a region of higher

concentration to a lower concentration.

Movement of particles in gases: As the particles of air move about, they collide

with the solid particles of dust and make the dust appear to move randomly

under light

(c) explain everyday effects of diffusion in terms of particles, e.g. the spread of perfumes and cooking aromas; tea and coffee grains in water

Tea: The particles of tea diffuses out of the bag into the water. The particles move from higher concentrations to lower concentrations.

Perfume: When the liquid is sprayed, the violatile substance vaporize immediately and the other substance falls as liquids onto the surface of skin. Perfume diffuse away from bottle.

(d) state qualitatively the effect of molecular mass on the rate of diffusion and explain the dependence of rate of diffusion on temperature.
The smaller the mr, the faster it moves as it require lesser KE to move at a given speed, therefore it diffuses faster
The higher the Mr, there is more thermal energy converted to KE, increasing diffusion

2.2 Atomic Structure

(a) state the relative charges and approximate relative masses of a proton, a neutron and an electron
Proton: Mass 1, charge +1
Neutron: Mass 1, charge 0
Electrons: 0, Charge -1

(b) describe, with the aid of diagrams, the structure of an atom as consisting of protons and neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels)
Shells: 2,8,8,8

(c) define *proton (atomic) number* and *nucleon (mass) number*
Proton: No. of proton in nucleus
Mass number: Total no. of protons and nucleus

(d) interpret and use nuclide notations such as ${}^{12}_{6}C$
Top: mass number, bottom: proton number

(e) define the term *isotopes*
Atoms of the same elements that have the same no. of proton but diff no. of nucleon no. They have diff no. of neutrons.

(f) deduce the numbers of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers.

3. Chemical Bonding and Structure

Ionic Bonding

(a) describe the formation of ions by electron loss/gain and that these ions usually have the electronic configuration of a noble gas

Atoms that do not have duplet of octet electronic configuration have their outer shells to undergo changes that allow them to become more stable. They do so by losing, gaining or sharing e-. Ions have the configuration of noble gas as they are stable. When atom loses one or more e-, an ion that is positively charged is formed as no. of protons is more than e-. When atom gains one or more e-, an ion that is negatively charged is formed as no. of e- is more than protons.

(b) describe, including the use of 'dot-and-cross' diagrams, the formation of ionic bonds between metals and non-metals, e.g. $NaCl$; MgCl2
Ions of opp charges can join by ionic bonding, there is mutual electrostatic attraction between ions of opposite charges. It makes them move towards each other and remain in position. The force is very strong at close range and holds the ions together. Ionic compounds have no net charges as the total positive charge must be equal to the total negative charge from the negative ion. Ionic bonds are mostly formed between metals and non-metals.

(c) state that ionic materials contain a giant lattice in which the ions are held by electrostatic attraction, e.g. NaC*l*\
Positive and negative ion exert electrostatic forces all around themselves. The positive ion that is attracted to a negative ion will continue to attract another negative ion. In solid state, the ionic compound exist as a 3d giant lattice structure. The structure holds an uncountably large no. of sets which are regularly spaced and held together by ionic bonds.
Giant ionic lattice structure: 3D structure of alternating positive and negative ion.

(candidates will **not** be required to draw diagrams of ionic lattices)

(d) relate the physical properties (including electrical property) of ionic compounds to their lattice structure (see also **3.4**(g)).

MP and BP: As the mutual electrostatic attraction between oppositely charged ions are strong, a large amount of energy is needed to overcome these forces. Hence it have a very high MP and BP, solid at rtp.

Hardness: When enough force is applied the ions move away from the lattice position and ions of the same charge approach each other, The repulsive forces between ions of the same charge become larger than the attractive forces, and the lattice will shatter, hence it is brittle.

Solubility: Most ionic compounds are soluble in water and insoluble in organic solvents. When ionic compounds are dissolved in water, they are in aq state.

Conductivity: When ionic compound are in solid state, the ions in lattice are only able to vibrate about their fixed position. Hence the ions are not mobile. When the compound is in molten, the ions are mobile and can conduct electricity.

3.2 Covalent Bonding

(a) describe the formation of a covalent bond by the sharing of a pair of electrons and that the atoms in the molecules usually have the electronic configuration of a noble gas
In covalent bonding, the atoms share e- and join. The overlap of the pair of e- is the bonding pair. It is a single covalent bond. Covalent bonds can occur as the e- can be shared between atoms. It can share e- with more than one atom.

(b) describe, using 'dot-and-cross' diagrams, the formation of covalent bonds between non-metallic elements, e.g. H_2; O_2; H_2O; CH_4; CO2

(c) deduce the arrangement of electrons in other covalent molecules
Covalent molecules are mostly simple molecules which have a fixed ratio. Some covalent molecules exist as a giant molecule which has uncountable number of atoms

(d) relate the physical properties (including electrical property) of covalent substances to their structure and bonding (see also **3.4**(g)).
Simple covalent molecules:
The atoms are held together by strong covalent bonds but the molecules are held together by weak intermolecular forces of attraction.

MP and BP the weak intermolecular forces of attraction can be easily overcome with small amounts of energy. As the molecule gets larger, the intermolecular forces of attraction are stronger and the MP and BP gets higher. Electrical conductivity: The molecules do not have mobile charge carriers to conduct electricity, however some simple covalent substance dissociate in water resulting it to conduct electricity in water.

Giant covalent structure consists of many atoms that are all connected to one another by strong covalent bonds.

Diamond: Each carbon atoms forms covalent bonds with four other carbon atoms. Thiis gives a 3D network structure.

Graphite: Each carbon atoms forms covalent bonds with three other carbon atoms. This gives a layered structure consisting of hexagonal rings of carbon atoms.
Giant covalent structures have numerous bonds holding the atoms together making it hard and rigid. Graphite has weak intermolecular forces between the layers, causing layers to be easily separated but each layer is hard to break apart.

Graphite is soft and slippery with a small amount of energy needed to overcome the weak intermolecular forces of attraction between the layers of carbon atoms.
Giant covalent bonds with no mobile e- do not conduct electricity. In graphite each carbon is bonded to three other carbon atoms, it leaves one unbonded e- per carbon which is free and mobile to conduct electricity.

MP and BP: Very resistant to heat due to the large no. of bonds in the network. To melt this giant covalent substance, a large amount of energy is required to break the strong covalent bonds. They are solid at rtp.

3.3 Metallic Bonding

(a) describe metals as a lattice of positive ions in a 'sea of electrons'

Metal atoms are held strongly together by metallic bonding, in solid state these atoms form a giant metallic structure. In this structure, metal atoms lose their outermost e- and become positively charged ions. These e- no longer belong to any metal atoms and become delocalized. They can move freely between the metal ions like a cloud of negative charges. Metallic bond: Mutal electrostatic attraction between positively charged ions in metal and "sea of mobile e-"

(b) describe the general physical properties of metals as solids having high melting and boiling points, malleable, good conductors of heat and electricity in terms of their structure (see also **3.4**(g)).

Pure metal have a regular structure. The layers of atoms can slide over one another easily. When hammered pure metals can be bent or flattened into thin sheet. It can be pulled into a wire and the sea of delocalized e- will continue holding the metal ions together unless much greater force is applied.
MP and BP: Metals have high MP and BP as the atoms are held together in a lattice by strong metallic bonds.

3.4 Structure and Properties of Materials

(a) describe the differences between elements, compounds and mixtures

Element: Pure substance that cannot be broken into 2 or more simpler substance by chemical methods.

Compound: A pure substance containing 2 or more elements that are chemically combined in a fixed ratio, it is chemically bonded together in covalent or ionic compound.

Mixture: Do not have substance in fixed ratio, it can consists of elements and compounds.

(b) describe an alloy as a mixture of a metal with another element, e.g. brass; stainless steel
Alloy is a mixture of metal with one or more other elements. The irregular structure arrangement in alloys kead to the difference in properties between alloys and pure metals.

(c) identify representations of metals and alloys from diagrams of structures

(d) explain why alloys have different physical properties to their constituent elements

In alloy, the atoms are of different sizes, The regular lattice arrangement in pure metal is disrupted, so a larger force is needed to make the layers slide over each other. Alloys tend to be less malleable and less ductile than the pure metals. Alloys are mixture that melt over a range of temperature unlike metals which have fixed MP.

(e) compare the structures of the following substances in order to deduce their properties:
  (i) simple molecular substances, e.g. methane, iodine
  (ii) macromolecules, e.g. poly(ethene)
  (iii) giant covalent substances, e.g. sand (silicon dioxide), diamond, graphite

(see also **3.4**(g))

(f) compare the bonding and structures of diamond and graphite in order to deduce their properties such as electrical conductivity, lubricating or cutting action (see 3.2d)

(candidates will **not** be required to draw the structures)

(g) deduce the physical and chemical properties of substances from their structures and bonding and vice versa (see also **3.1**(d), **3.2**(d), **3.3**(b) and **3.4**(e)).

4. Chemical Calculations

4.1 Formulae and Equation Writing

(a) state the symbols of the elements and formulae of the compounds mentioned in the syllabus

(b) deduce the formulae of simple compounds from the relative numbers of atoms present and vice versa

(c) deduce the formulae of ionic compounds from the charges on the ions present and vice versa

(d) interpret chemical equations with state symbols

s- solid, l- liquid, aq- aqueous, g- gas

(e) construct chemical equations, with state symbols, including ionic equations.

4.2 The Mole Concept and Stoichiometry

(a) define relative atomic mass, A_r
It is the average mass of one atom of that element relative to 1/12 the mass of carbon-12.

(b) define relative molecular mass, M_r, and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses

Mr of ionic compound is the average mass of one unit of that compound relative to 1/12 the mass of an atom of carbon-12.

(c) define the term *mole* in terms of the Avogadro constant
Mole counts the no. or particles. One mole is 6.02 x10(23), The number is called Avogadro constant.

(d) calculate the percentage mass of an element in a compound when given appropriate information
(No. of atom x ar of element) / (mr of compound) x100%

(e) calculate empirical and molecular formulae from relevant data

1. Fill in the percentage. 2. Fill in the ar and calculate the mole present. 3. Divide the mole value by the smallest mole value which will give the ratio of the elements.

(f) calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24 dm^3 at room temperature and pressure); calculations involving the idea of limiting reactants may be set

(knowledge of the gas laws and the calculations of gaseous volumes at different temperatures and pressures are **not** required)

(g) apply the concept of solution concentration (in mol/dm^3 or g/dm^3) to process the results of volumetric experiments (e.g. titration) and to solve simple problems

m1v1/m2v2 = mol1/mol2 (shortcut)

(appropriate guidance will be provided where unfamiliar reactions are involved)

(h) calculate % yield and % purity.
Actual/theory, mass of pure substance/ total mass of impure substance.

5. Acid-Base Chemistry

5.1 Acids and Bases

(a) describe the meanings of the terms acid and alkali in terms of the ions they produce in aqueous solution and their effects on Universal Indicator
Acid is a substance which produce hydrogen ions in aqueous solution
Alkali is a substance which produce hydroxide ion in aqueous solution
In universal indicator, it will turn from green to orange for acid

In universal indicator, it will turn from green to purple for alkali

(b) describe neutrality and relative acidity and alkalinity, in terms of

(i) relative H^+ and OH^- ion concentrations,

Acid-> high H+ low OH- Alkali -> low H+ high OH- Neutral-> H+ = OH-

(ii) colour in Universal Indicator, and

Acid: Red to light green

Alkali: Dark green to purple

(iii) the pH scale

Acid: 0-6.9

Alkali: 7.1-14

(calculation of pH from hydrogen ion concentration is **not** required)

(c) describe qualitatively the difference between strong and weak acids in terms of the extent of ionization

Strong acid: Acid that is completely ionized in aqueous solution

Weak acid: Acid that is only partially ionized in aqueous solution

(d) describe the characteristic properties of acids as in reactions with metals, bases and carbonates to form salts
Metal + Acid-> Salt + Hydrogen
Base + Acid -> Salt + Water
Carbonate + Acid -> Salt + Water + Carbon Dioxide
Acid + Base -> Salt + Water

(e) describe the reaction between hydrogen ions and hydroxide ions to produce water,
$H^+ + OH^- \rightarrow H_2O$, as neutralisation

(f) describe the importance of controlling the pH in soils and how excess acidity can be treated using calcium hydroxide

Controlling pH is important for optimal growth of plant

Acidity can be controlled by CaOH as the base reacts with the acids in the soil and raise the pH so the plants can grow healthily. Adding too much CaOH will make it too alkaline and unsuitable for growth.

(g) describe the characteristic properties of bases in reactions with acids and with ammonium salts
Base + Ammonium salts -> Salt + Water + Ammonia

(h) classify oxides as acidic, basic, amphoteric or neutral based on metallic/non-metallic character.

Metal oxide can be classified as basic or amphoteric. Mose metal oxides are basic oxides which are insoluble in water while some dissolve in water to form alkali . Metal oxides reacts with acid to form salt and water. (All metal oxides except ZAP)

Amphoteric oxides are metallic oxides which reacts with both acids and bases. (ZAP: Zn Al Pb)

Non-metallic oxides can be classified as acidic or neutral oxides. Most non-metallic oxides are acidic oxides which will dissolve in water to form acids. (All except H2O, CO, NO)

Neutral oxides show neither basic or acidic properties and are insoluble in water. (H2O, CO, NO)

5.2 Salts

(a) describe the techniques used in the preparation, separation and purification of salts as examples of some of the techniques specified in Section **1.2**(a) (methods for preparation should include precipitation and titration together with reactions of acids with metals, insoluble bases and insoluble carbonates)

Method: **Reaction with acid** Salts: Soluble, Starting materials: 1 insoluble 1 soluble.
Acid + Excess metal, Acid + Excess insoluble base, Acid + Excess insoluble carbonate
Add acid and excess solid (must be in excess so that the acid is used up otherwise the salt produced will be contaminated with the acid) Filter the mixture and collect the filtrate and crystalize the filtrate.

Method: **Titration** Salts: Soluble, Starting materials: Both soluble.
Acid + Alkali, Acid + Soluble carbonate
The volume of starting materials must be in exact as it cannot be removed by any method. Titrate and collect the solution and pour into evaporating dish and hit till saturated. Cool and crystalize and collect crystals

Method: **Precipitation** Salts: Insoluble, Starting materials: Both soluble.
Mixing 2 solutions to form insoluble solid that separates out from reaction mixture. When solution is mixed, the soluble salt are formed and can be separated from the starting mixture through filtration. Wash the ppt with distilled water and allow the ppt to dry between 2 pieces of filter paper.

(b) describe the general rules of solubility for common salts to include nitrates, chlorides (including those of silver and lead), sulfates (including those of barium, calcium and lead), carbonates, hydroxides, salts of Group 1 cations and ammonium salts
All nitrates, ammonium, potassium, sodium (G1 ammonium) are soluble
All chlorides are soluble except silver and lead (AP)
All sulfates are soluble except Barium, Lead, Calcium (PBC)
All carbonates are insoluble except Sodium, Potassium, Ammonium (G1, NH_4)

(c) suggest a method of preparing a given salt from suitable starting materials, given appropriate information.

5.3 Ammonia

(a) describe the use of nitrogen, from air, and hydrogen, from the cracking of crude oil, in the manufacture of ammonia
$N_2 + 3H_2 \leftrightarrow 2NH_3$
A reversible reaction that can go forward and backward at the same time

(b) state that some chemical reactions are reversible, e.g. manufacture of ammonia

(c) interpret data relating to the conditions used in industry for processes involving reversible reactions, e.g. manufacture of ammonia by the Haber Process,

Hydrated copper 2 sulfate

(knowledge of Le Chatelier's Principle is **not** required).

Optimal conditions: 250atm, 450 degrees Celsius,

presence of finely divided catalyst

6. Qualitative Analysis

(a) describe the use of aqueous sodium hydroxide and/or aqueous ammonia to identify the following aqueous cations through the formation of precipitates (if any) and their subsequent solubility: aluminium, ammonium (together with evolution of ammonia gas upon warming), calcium, copper(II), iron(II), iron(III) and zinc

(formulae of complex ions are **not** required)

Cation, Excess NaOH, Excess NH_3

Aluminum, White ppt soluble in excess, White ppt insoluble in excess

Zinc, White ppt soluble in excess, White ppt soluble in excess

Calcium, White ppt insoluble in excess, No change observed

Ammonium, No ppt (NH_3 given off upon heating), no change observed

Copper (II), Light blue ppt insoluble in excess, Light blue ppt soluble in excess to form dark blue solution

Iron (II), Green ppt insoluble in excess, Green ppt insoluble in excess (turn brown upon standing due to oxidation)

Iron (III), Brown ppt insoluble in excess, Brown ppt insoluble in excess

(b) describe tests to identify the following anions: carbonate (by the addition of dilute acid and subsequent use of limewater); chloride (by reaction of an aqueous solution with nitric acid and aqueous silver nitrate);

iodide (by reaction of an aqueous solution with nitric acid and aqueous silver nitrate); nitrate (by reduction with aluminium in aqueous sodium hydroxide to ammonia and subsequent use of damp red litmus paper) and sulfate (by reaction of an aqueous solution with nitric acid and aqueous barium nitrate)

Carbonate, Acidify, Bubble in limewater, White ppt formed (CO2 produced)

Chloride, Acidify, add aq silver nitrate, White ppt formed (AgCl formed)

Iodide, Acidify, add aq silver nitrate, yellow ppt formed (AgI formed)

Sulfate, Acidify, add aq barium nitrate, white ppt formed (BaSO4 formed)

Nitrate, Add NaOH, aluminum and warm mixture, turn damp red litmus paper blue (NH3 produced)

(c) describe tests to identify the following gases: ammonia (using damp red litmus paper); carbon dioxide (using limewater); chlorine (using damp litmus paper); hydrogen (using a burning splint); oxygen (using a glowing splint) and sulfur dioxide (using acidified potassium manganate(VII)).

Hydrogen, Place burning splint, Extinguish with a pop sound
Oxygen, Place glowing splint, Rekindled (catches fire again)
Carbon dioxide, Bubble gas into limewater, White ppt formed (ppt will dissolve upon further bubbling)

Pungent gases
Chlorine, Place damp blue litmus paper, Will turn red them bleached
Sulfur dioxide, Place soaked acidified KMnO4 at the mouth, It will turn colorless
Ammonia, Place damp red litmus paper, Damp red litmus paper turn blue

7. Redox Chemistry

7.1 Oxidation and

Reduction

(a) define *oxidation* and *reduction* (redox) in terms of oxygen/hydrogen gain/loss
Oxidation is the loss of hydrogen and the gain of oxygen
Reduction is the gain of hydrogen and the loss of oxygen
Redox reaction refers to reaction that involve oxygen being directly gained or loss where reduction and oxidation take place at the same time.

(b) define *redox* in terms of electron transfer and changes in oxidation state

Remember OiL RiG where Oxidation is the loss of e- and reduction is the gain of e-

For reaction that do no involve the gain/lose of hydrogen/oxygen and gain/loss of e- use oxidation state (e.g. covalent bonding)

Oxidation state is the charge of an atom/element would have if it existed as an ion in compound. It can be +ve, -ve, or even 0. Increase in oxidation state refers to the loss of e- therefore oxidation. Decrease in oxidation state refers to the gain of e- therefore reduction.

Rules: Atoms of same element will have 0 oxidation state.

G1 and G2 will be +1 and +2 respectively

Fluorine in compound will always be -1 as it is the most reactive non-metal.

Hydrogen in compound will usually be +1. When it forms compound with non-metal, it will be -1

Oxygen in compound will usually be -2. In peroxides, it will be -1. (second most reactive after fluorine)

Element from G3 to G7 is often based on their valency, Metallic elements generally positive, non-metallic elements generally positive or negative. (compound has no net charge so oxidation state will be 0 and the ion will either have a positive or negative charge that will add up to the charge number.)

Elements that are transition metals can have different oxidation state. The roman numerals will indicate their oxidation state.

(c) identify redox reactions in terms of oxygen/hydrogen gain/loss, electron gain/loss and changes in oxidation state
Redox is when both oxidation and reduction occur at the same time. E.g. Respiration and photosynthesis.

(d) describe the use of aqueous potassium iodide and acidified potassium manganate(VII) in testing for oxidising and reducing agents from the resulting colour changes.

Oxidizing agent will oxidize other substance and reduce itself as it takes away e- from the other substance for itself.
e.g. $KMnO_4$, O_2, Cl_2
Reducing agent will reduce other substance and oxidize itself as it loses e- to the other substance.
e.g. H_2, reactive metals, C, KI

$KMnO_4$ will turn from purple to colourless if reducing agent is present. The ions are reduced from -4(MnO_4^-_ to 2+(Mn^{2+})

KI will turn from colorless to yellow brown if oxidizing agent is present. The ions are oxidized from –1 ($2I^-$) to 0(I_2).
If unsure, add few drops of starch and it will turn dark blue if iodine is present.

7.2 Electrochemistry

(a) describe electrolysis as the conduction of electricity through an ionic compound (an electrolyte), when molten or dissolved in water, leading to chemical changes (including decomposition) at the electrodes

Electrolysis is the process of passing an electric current through a compound to chemically separate its compounds.

Electrolytic cell is a device that converts electrical energy into chemical energy.

Power supply: Acts as electron pump as it causes e- to move from anode to cathode. E- enters the battery from the anode and pumped out to the cathode. Creates positive charge on the anode amd negative charge at the cathode.

Electrolyte: Electrically conductive substance in the molten or aq state. Contains mobile charge carriers to conduct electricity.

Electrodes: Contains delocalized e- to conduct electricity. Anode is at the positive end while the cathode is at the negative end.

In electrolysis, charges through the external circuit by e- and through the electrolute by the ions. In the external circuit, e- move form the negative terminal to the positive terminal.

At the anode: Anions are attracted to the positively charged anode. The anions can lose e- to the anode and become oxidized.

At the cathode: The cations are attracted to the negatively charged cathode. The cations can gain e- to the cathode and become reduced.

(b) describe electrolysis as evidence for the existence of ions which are held in a lattice when solid but which are free to move when molten or in solution

A circuit is only complete if mobile charges are able to pass through all parts of the circuit. Using electrolysis to prove that ionic compounds can conduct electricity only in molten and aqueous state.

(c) describe, in terms of the mobility of ions present and the electrode products, the electrolysis of molten sodium chloride, using inert electrodes

In solid state, NaCl do not conduct electricity. In molten and aqueous state, it is able to conduct electricity. Therefore, the sodium and chloride ions are mobile and conduct electricity. It contain positive and negative ions that move to conduct charge. Electrolytes undergo redox reaction ate the electrodes to form new substance.

(d) predict the likely products of the electrolysis of a molten binary ionic compound using inert electrodes
Inert electrodes are electrodes that do not take part in the electrolysis reaction. They are mostly graphite and

platinum. They do not undergo chemical changes. However other metal anodes may will generally undergo oxidation. Molten binary ionic compound us typically a salt containing only one cation and anion in the liquid state.
When the anions and cations are oxidized and reduced respectively, they from atoms or molecules and they are discharged.
Electrolysis of molten sodium chloride. Cl^- are attracted to the anode as each cl^- lose one e^- to the anode and becomes oxidized. The na^+ are attracted to the cathode. Each Na^+ gain one e^- from the cathode and gets reduced.
$2NaCl \rightarrow 2Na + Cl_2$

(e) apply the idea of selective discharge based on

(i) cations: linked to the reactivity series (see also **8.4**)

The more reactive a metal, the more stable its ion, which is harder to convert the ion back into the metal.

The less reactive the metal, the less stable its ion which is more easily converted back into metal.

Ag^+ will be selectively discharged at the cathode over Cu^{2+}. Cu^{2+} will only begin to discharge only after all the Ag^+ ions have been discharged from the solutions. It is almost impossible to discharge metals that are above hydrogen in the series from the aq as almost all aq solutions contains H^+ ions.

(ii) anions: halides, hydroxides and sulfates (e.g. aqueous copper(II) sulfate and dilute sodium chloride solution (as essentially the electrolysis of water))

In dilute solution, the lower the solution of the anion in the electrochemical series the more likely the anion will be selectively discharged. From most reactive to least reactive: Sulfate, Nitrate, Chloride, Bromide, Iodide, Hydroxide.

Chloride: $2Cl^- \rightarrow Cl_2 + 2e^-$

Bromide: $2Br^- \rightarrow Br_2 + 2e^-$

Iodide: $2I^- \rightarrow I_2 + 2e^-$

Hydroxide: $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$

(iii) concentration effects (as in the electrolysis of concentrated and dilute aqueous sodium chloride)

(in all cases above, **inert** electrodes are used)

A high concentration increase the ease of discharge of cl^-, br^-, i^-. This can sometimes override the electrochemical series. Cl^- ions can be selectively discharged over OH^- if concentrated and chloride gas will be evolved instead of oxygen gas.

(f) predict the likely products of the electrolysis of an aqueous electrolyte, given relevant information

Based on selectively discharged anion and cation.

(g) construct ionic equations for the reactions occurring at the electrodes during the electrolysis, given relevant information

Construct the half ionic equations and make sure it is in a ratio of 1:1

(h) describe the electrolysis of aqueous copper(II) sulfate with copper electrodes as a means of purifying copper

Reactive electrodes can take part in the electrolysis process. The reactive electrodes is often used to introduce a metal cation into the electrolyte. When reactive metal used as anode, it is oxidized to form its cation, the cation is then reduced back into the original metal on the cathode.

Electrolysis using copper electrode and copper(II) sulfate solutions.

At the anode: SO_4^{2-} and OH^- ions are attracted to the anode. However copper is not an inert electrode so it is oxidized instead. $Cu \rightarrow Cu^{2+}$ $2e^-$ the copper anode dissolve to form cu^{2+} ions and decrease in mass.

At the cathode: cu^{2+} and H^+ ions are attracted to the cathode. Since copper is below hydrogen in the reactivity series, cu^{2+} ions are selectively discharged over h^+ ions. $Cu^{2+} + 2e^- \rightarrow Cu$ Red brown copper is deposited into the copper cathode and increase in mass.

The copper gets deposited onto the cathode which mainly comes from the electrolyte. The copper anode is constantly oxidizing and dissolving into the electrolyte replenishing cu^{2+} ions which are reduced at the cathode.

As the concentration of cu^{2+} ions remains unchanged, the color intensity of the blue electrolyte remains the same.

Purification of impure copper: The impure copper anode dissolve and decrease in mass. Pure copper is deposited into the cathode. Impurities present in the impure copper fall to the bottom as slime. If the anode is replaced with inert electrode, some pure copper can still be deposited at the cathode as the copper is produced from the cu^{2+} from the electrolyte. The color intensity changed from blue to colorless as the cu^{2+} reservoir has been deposited.

(i) describe the electroplating of metals, e.g. copper plating, and state one use of electroplating
Electroplating allows to coat a thin layer of metal onto object. Electroplating is where the plating metal is made the anode and the object to be coated with the metal is made the cathode. If it is non- conductive object is to be coated, it need to be immersed in graphite to make it conductive.
Use of electroplating: To improve the attractiveness, protect them from corroding

(j) describe the production of electrical energy from simple cells (i.e. two electrodes in an electrolyte) linked to the reactivity series (see also **8.4**) and redox reactions (in terms of electron transfer)

The more reactive metal acts as the anode which is preferentially oxidized forming cations that enter the electrolyte and releases the e- that flow through the external circuit.

The less reactive metal acts as the cathode and causes cation from the electrolyte to gain e- and gets reduced.

This result in the movement of e-. The e- move spontaneously from where they are produced to where they are consumed. This movement of e- produces electricity.

The further the two metals apart, the greater the voltage produced.

(k) describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generate electricity directly in a hydrogen fuel cell

at the cathode : o2 +2h2o +4e- -> 4oh-

At the anode: h2 +2oh- -> 2h2o + 2e-

8. Patterns in the Periodic Table

8.1 Periodic Trends

(a) describe the Periodic Table as an arrangement of the elements in the order of increasing proton (atomic) number

(b) describe how the position of an element in the Periodic Table is related to proton number and electronic configuration

Electronic configuration: No. of shells = period
G1 -> 1 valence e-

(c) describe the relationship between number of outer (valence) electrons and the ionic charge of an ion for the first twenty elements

(d) explain the similarities between the elements in the same group of the Periodic Table in terms of their electronic configuration

elements in the same group will have the same amt of valence e- and the elements with the same number of valence e- will have similar chemical properties.

(e) describe the change from metallic to non-metallic character from left to right across a period of the Periodic Table
Metals are on the left and the non-metals are on the right. There is a decrease in metallic properties and an increase in non-metallic properties across a period. As the atom is less likely to lose e-. As it requires more energy to lose e-.

(f) describe the relationship between number of outer (valence) electrons and metallic/non-metallic character

(g) predict the properties of elements in Group 1 and Group 17 using the Periodic Table.

8.2 Group Properties

(a) describe lithium, sodium and potassium in Group 1 (the alkali metals) as a collection of relatively soft, low density metals showing a trend in melting point and in their reaction with water

Group 1 are alkali metals which have similar chemical properties. They are soft and be cut easily. They have low MP and BP, low densities where they can float on water. Going down the group, the MP and BP decreases. As it goes down the group the size of atom increases, and it is easier to lose the valence e-. The reactivity will then increase.

(b) describe chlorine, bromine and iodine in Group 17 (the halogens) as a collection of diatomic non-metals showing a trend in color, state and their displacement reactions with solutions of other halide ions
Group 17 elements are Halogens and have low MP and BP, where it is colored.
Going down the group, the MP and BP increases, and the color intensifies.
Chlorine – Yellow-green gas, Bromine – Red-brown liquid, Iodine – Purple black solid
A displacement is a reaction in which one element takes the place of another element in the compound. A more reactive halogen will displace a less reactive halogen from the halide solution. When chlorine water is added to aq sodium bromide, a red brown solution is obtained. (the red is from the bromine displaced from the bromide solution.)

(c) describe the elements in Group 18 (the noble gases) as a collection of monoatomic elements that are chemically unreactive and hence important in providing an inert environment, e.g. argon and neon in light bulbs; helium in balloons; argon in the manufacture of steel

Noble gases are monoatomic non-metals and are colorless at rtp and have low MP and BP. They are insoluble in water and unreactive. Noble gases are unreactive as they have fully filled valence shell and do not gain or share e-.

(d) describe the lack of reactivity of the noble gases in terms of their electronic configurations.

8.3 Transition Elements

(a) describe typical transition elements as metals having high melting point, high density, variable oxidation state and forming coloured compounds
Transition have high MP and BP and densities. With these properties this allows them to be used in industries. As they are lightweight and withstand high temperatures.
They have variable oxidation state.
They are colored compound which are different at different oxidation state.

(b) state that the elements and/or their compounds are often able to act as catalysts (see also **10**(d)).

They are good catalyst for many reactions. Catalyst is a substance that increase the rate of chemical reactions and remain chemically unchanged at the end of the reaction.

8.4 Reactivity Series

(a) place in order of reactivity calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to

Reactivity series: G1, G2, Zinc iron lead, copper silver.

(i) the reactions, if any, of the metals with water, steam and dilute hydrochloric acid

Reaction of metals with cold water. Metal + water -> Metal hydroxide + hydrogen

Some metals react with cold water and more reactive metals react violently with cold water. Potassium reacts **very violently**, Sodium reacts **violently**, Calcium reacts **readily**, Magnesium reacts **very slowly**.

Reaction of metals with steam. Metal + steam -> Metal oxide + Hydrogen

K, Na, Ca reacts **explosively** which should not be carried out. Hot magnesium (White glow produced) reacts **violently**, Hot zinc (Yellow when hot, white when cold) reacts **readily**, red hot iron reacts **slowly** (needs to be constantly heated as it is a reversible reaction).

Reaction of metal with dilute hydrochloric acid. Metal + HCl -> metal chloride + Hydrogen

K, Na reacts **explosively**. Calcium reacts **violently**, Magnesium reacts **rapidly**, Zinc reacts **moderately fast**, Iron **reacts slowly**, lead should react but in the initial reaction lead chloride which is insoluble is formed and will prevent further reaction from taking place.

(ii) the reduction, if any, of their oxides by carbon and/or by hydrogen

Reduction of metal oxides with carbon. Metal oxide + carbon -> Metal + carbon dioxide. The lower the metal in the reactivity series from the carbon, the more readily the reduction of the metal oxide will occur.

Reduction of metal oxides with hydrogen. Metal oxide + Hydrogen -> Metal + Steam. Hydrogen will reduce the oxides of metals from iron and below. Oxides of iron will require the highest temperature for reduction.

(b) describe the reactivity series as related to the tendency of a metal to form its positive ion, illustrated by its reaction with

(i) the aqueous ions of the other listed metals

A more reactive metal has a greater tendency to form positive ions compared to a less reactive metal.

(ii) the oxides of the other listed metals

A more reactive metal can displace a less reactive metal from its salt solution. When iron fillings are added to copper (II) sulfate where copper metal is displaced out of the solutions and red-brown solid is formed. The solution turns green. As iron has displaced copper from copper (II) sulfate solution. A more reactive metal forms positive ion more readily, as iron is more reactive than copper, iron atoms become iron(II) ions and form iron (II) sulfate. If copper is dipped into iron (II) sulfate, no reaction occurs. Copper is less reactive than iron and has a lower tendency to form ions compared to irons.

(c) deduce the order of reactivity from a given set of experimental results

(d) describe the action of heat on the carbonates of the listed metals and relate thermal stability to the reactivity series

The more reactive a metal, the more difficult it is to decompose its carbonate by heat.
Potassium and sodium carbonate is unaffected by heat.
Calcium, Magnesium, Zinc, Iron, Lead, Copper carbonate can decompose into metal oxide and carbon dioxide on heating.
Silver carbonate decompose into silver oxide and carbon dioxide upon heating. As silver oxide it thermally unstable, it is further decompose into silver.

(e) describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series
Metals can be extracted by electrolysis and extracted by reduction with carbon. Metals can only be extracted by electrolysis as the compounds are very difficult to break down. Metals which are less reactive are extracted by carbon.

(f) describe the essential conditions for the corrosion (rusting) of iron as the presence of oxygen and water; prevention of rusting can be achieved by placing a barrier around the metal, e.g. painting; greasing; plastic coating; galvanizing
Rusting is the process that produced rust. It needs oxygen and water to happen.

(g) describe the sacrificial protection of iron by a more reactive metal in terms of the reactivity series where the more reactive metal corrodes preferentially, e.g. underwater pipes have a piece of magnesium attached to them.
Sacrificial protection is the protection of iron and steel against rusting using a more reactive metal such as zinc. As zinc is more reactive than iron and can react in place of iron. It is galvanizing and attaching a more reactive metal. Galvanizing is dipping the iron object into molten zinc and protects the object by corroding in place of the iron and steel. When attaching a more reactive metal the more reactive metal corrodes in place of the iron and steel.
As zinc has a higher tendency to form positive ions. Zinc atoms lose e- in preference to iron atoms and prevents iron from forming iron (III) oxide. More reactive metal will corrode in place or iron. As long zinc is present, iron will not rust

9. Chemical Energetics

(a) describe the meaning of enthalpy change in terms of exothermic (ΔH negative) and endothermic (ΔH positive) reactions
Endothermic changes (entering): absorb thermal energy from the surrounding which cause a decrease in temperature of the surroundings.
Exothermic changers (exiting): releases thermal energy to the surroundings, This causes a increase in temperature of the surroundings.

(b) represent energy changes by energy profile diagrams, including reaction enthalpy changes and activation energies (see also **10**(c), **10**(d))
Activation energy is the minimum amount of energy that colliding reactant particles must posses to react with each other. When reactant particles with sufficient energy collide, they are able to break their bonds to form new bonds.
Enthalpy changes measures the difference in energy content of the reactants and products.

Endothermic changes is the energy absorbed in bond breaking is greater than the energy released during the bond making phase.
Exothermic changes is the energy released during the bond making phase greater than the energy absorbed during the bond making phase.

(c) describe bond breaking as an endothermic process and bond making as an exothermic process

Bond breaking requires energy. This energy is absorbed from the surroundings thus it is endothermic.
Bond making releases energy to the surroundings, thus making it exothermic.

(d) explain qualitatively overall enthalpy changes in terms of the energy changes associated with the breaking and making of covalent bonds.

(e) Overall enthalpy change is bond breaking – bond making

10. Rate of Reactions

(a) describe the effect of concentration, pressure, particle size and temperature on the rates of reactions and explain these effects in terms of collisions between reacting particles

When chemical reaction occurs between particles, the collision is said to be an effective collision.
Effect of concentration: as the concentration of reactant increase, the number of reacting particles per unit volume increases and collide more frequently thus increasing the rate of effective collision.
Effect of pressure: When gases are compressed to a smaller volume, the spacing between particles in the gas decreases. As the particles are closer together, they collide more frequently, increasing the rate of effective collision.
Effect of particle size: By increasing the surface area, it allows for more effective collision to occur.
Effect of temperature: at higher temperatures, the reactant particles will have more kinetic energy and they will move and vibrate faster, increasing the rate of effective collision.
At higher temperatures, more reactant particles posses energy that is greater than or equal to the activation energy, which will increase the frequency of effective collision.

(b) define the term *catalyst* and describe the effect of catalysts (including enzymes) on the rates of reactions

Catalyst increases the rate of chemical reaction without being chemically changed at the end of the reaction.

Effect off catalyst: Catalyst will provide an alternative pathway of lower activation energy for the reaction to occur. With an alternative pathway, more reactant particles will have energy greater than or equal to activation energy which increases the frequency of effective collision.

(c) explain how pathways with lower activation energies account for the increase in rates of reactions (see also **9**(b))

(d) state that some compounds act as catalysts in a range of industrial processes and that enzymes are biological catalysts (see also **8.3**(b), **9**(b), **10**(c) and **12**(d)
e.g. Iron, AlO, Nickel, Platanium.

(e) suggest a suitable method for investigating the effect of a given variable on the rate of a reaction

Volume of gas produced, Change in mass.

(f) interpret data obtained from experiments concerned with rate of reaction.

Steeper gradient = faster rate of reaction

11. Organic Chemistry

Fuels and Crude Oil

a) name natural gas, mainly methane, and crude oil as non-renewable sources of energy
natural gas and crude oil are considered as non-renewable sources of energy. There are limited amounts of these sources of energy and cannot be replaced quickly enough to keep up with the rate of use.

b) describe crude oil as a mixture of hydrocarbons and its separation by fractional distillation to yield fractions which have competing uses as fuels and as a source of chemicals (see also **1.2**(a))
crude oil contains many types of hydrocarbons with different properties. They need to be separated into different group of hydrocarbons before each group can be used.
The crude oil can be separated through fractional distillation where different fractions are condensed at

different boiling points in a huge fractionating column. The component with the lowest boiling point is collected at the topmost outlet and the component with the highest boiling point is collected at the outlets down the column. The fractions with lower bp rise further up the column before they are condense and collected. Crude oil contains hydrocarbons with 1-70 carbons per molecule. The hydrocarbons that contain more carbon atoms have higher boiling points. Hydrocarbons with lower boiling points are less viscous and burn more easily.
Liquefied petroleum gas (LPG) – fuel for cooking and heating
Petrol - fuel for motorcars
Naphtha – feedstock for making petrochemicals such as plastic and detergents
Kerosene – fuel for aircraft engines, cooking using oil stoves and heating
Diesel – fuel for diesel engines in buses, lorries and trains
Paraffin/Lubricating oil – for lubricating machines/ making waxes
Bitumen – making road surface and roofing

c) describe biofuel (exemplified by bioethanol from sugarcane) as a renewable alternative to natural gas and crude oil
biofuels are alternative renewable energy sources to crude oil and natural gas. They come from plants or animals and can be replaced relatively quickly after use.

d) describe how biofuel, when compared to fossil fuels, can be more environmentally sustainable in terms of carbon dioxide emission (see also **12**(g)).
The burning of bioethanol is considered environmentally sustainable compared to fossil fuels. As the sugarcane grows, it absorbs co2 during photosynthesis which offset the co2 produced when the bioethanol is burnt.

11.1 Hydrocarbons

(a) describe a homologous series as a group of compounds with a general formula, similar chemical properties and showing a gradation in physical properties as a result of increase in the size and mass of the molecules, e.g. melting and boiling points; viscosity

homologous series consists of a family of compounds with the same general formula, and similar chemical properties because they have the same functional group. They have the same functional group, similar chemical properties, each member of series differs from the next by an additional ch2 unit. There is a gradual change in their physical properties as we go down the series from one member to the next.

(b) describe the alkanes as a homologous series of saturated hydrocarbons with the general formula C_nH_{2n+2}
alkanes are hydrocarbons that contains only carbon-carbon single bonds and carbon-hydrogen bonds. The name alkane ends with -ane. Alkane are simple molecular substances with weak forces of attraction between the molecules. The alkane have lo emp and bp. The mp and bp gradually increase down the series. The forces of attraction between alkane molecules increase as size increase. Alkane become more viscous as molecular size increase

(c) draw the structures of branched and unbranched alkanes, C_1 to C_4, and name the unbranched alkanes methane to butane

Meth(1), **E**th(2), **P**rop(3), **B**ut(4)

(d) define *isomerism* and identify isomers

compounds that have the same molecular formula but different structural formulae are called isomers.

Isomerism can occur when the length of the main carbon chain differs.

(e) describe alkanes (exemplified by methane) as being generally unreactive except in terms of combustion and substitution by chlorine
alkanes are generally unreactive because they only consist of strong carbon-carbon single bonds and strong carbon-hydrogen bonds, these bonds are hard to break. They can undergo the following reactions.

Combustion
When there is a source of ignition and can undergo **combustion**. Alkane burns excess

oxygen to produce carbon dioxide and water.
Butane + oxygen -> carbon dioxide + water
When there is insufficient oxygen, incomplete combustion takes place. Instead of carbon dioxide, carbon monoxide and soot are formed.
Butane + oxygen -> carbon monoxide + carbon + water

Substitution reactions
Alkane react with **halogen in the presence of uv light**.
Methane + chlorine -> chloromethane + hydrogen chloride
More hydrogen can be replaced with chlorine atoms form other compounds such as dichloromethane

(f) describe the alkenes as a homologous series of unsaturated hydrocarbons with the general formula C_nH_{2n}
Alkene are hydrocarbons that contains carbon-carbon double bonds. Where n>1 as the double bonds needs to take place.
Naming: ending with "ene"
MP and BP: they are simple molecular substance with weak forces of attraction between the molecules. Hence, they have low MP and BP. There is a gradual increase in MP and BP as it goes down the series. As there is a stronger forces of attraction between the molecules.
Solubility: insoluble in water but soluble in organic solvents. Alkene are not used in the solvents as they are more reactive and may interfere with the reactions.

(g) draw the structures of branched and unbranched alkenes, C_2 to C_4, and name the unbranched alkenes ethene to butene

(h) describe the manufacture of alkenes and hydrogen by cracking hydrocarbons and recognise that cracking is essential to match the demand for fractions containing smaller molecules from the refinery process

Cracking is a process in which larger hydrocarbons molecules are broken down into smaller hydrocarbons molecules.
E.g. : cracking with hexane. After cracking, hexane is separated into butane and ethane.
To speed up cracking, a catalyst is used, and which this process is known as catalytic cracking.
Conditions: Presence of Al2O3/SiO2, temperature of 600 degrees Celsius, 1atm
Importance of cracking: it converts less useful components of crude oil into shorter chain alkene(usually a starting materials) or shorter chain of alkane(usually used as fuel)
The byproducts of cracking also produces hydrogen which is used in the production of NH3

(i) describe the difference between saturated and unsaturated hydrocarbons from their molecular structures and by using aqueous bromine

Saturated hydrocarbons contain single covalent bonds between the carbon atoms. (1:4)

Unsaturated hydrocarbons contain double carbon bonds where more atoms can be added across the carbon-carbon bonds.

To test for saturation, bromination is used. Aqueous bromine/bromine water is added to the solution. Aqueous bromine will remain brown if added to saturated hydrocarbons while it will decolorize if added to unsaturated hydrocarbons.

(j) describe the reactions of alkenes (exemplified by ethene) in terms of combustion, polymerization (see also **11.4**(b)), and the addition with bromine, steam and hydrogen

Combustion
Alkene burns in **excess oxygen** to produce carbon dioxide and water.
Ethene + oxygen -> Carbon dioxide + water
As the percentage of mass oof carbon is higher, it is more likely to undergo incomplete combustion, producing carbon monoxide and carbon. It will also burn with a sootier flame.

Addition
As the carbon-carbon bonds are reactive, they will readily undergo addition reactions.
The c=c bonds will break to form new single bonds e.g. c-c.

Hydration
Alkane can react with steam to produce alcohol
Conditions: 300℃ , 60atm, Phosphoric acid

e.g. C2H4 + H2O -> C2H5OH

Hydrogenation
Alkene can react with hydrogen to form alkane
Conditions: 150℃, nickel catalyst
C2H4 + H2 -> C2H6

Bromination
Alkene can react with halogens. When ethene is bubble through liquid bromine, the brown color of bromine disappears and colorless liquid is formed.
C2H4 + Br -> CH2BrCH2Br

(k) state the meaning of *polyunsaturated* when applied to food products
Polyunsaturated fats contains hydrocarbon chain with two or more carbon-carbon double bonds in the chain.
The oil is unsaturated and margarine is saturated. Saturated fats are solid at rtp while unsaturated fats are liquid in rtp.

(l) describe the manufacture of margarine by the addition of hydrogen to unsaturated vegetable oils to form a solid product.

To produce margarine from vegetable oils, basically hydrogenation

Conditions: 150℃, Nickel catalyst

11.2 Alcohols, Carboxylic Acids and Esters

(a) describe the alcohols as a homologous series containing the –OH group

Alcohols are organic compounds with the hydroxyl functional group (-OH)

Suffix “-ol” where the hydroxyl is connected to the last carbon

The shorter chain alcohol are volatile liquids at rtp and pressure. They have higher MP and BP compared to hydrocarbons with the same no. of carbons.

Solubility decreases while MP/BP, molecular size, forces of attraction increases as length increases

(b) draw the structures of branched and unbranched alcohols, C_1 to C_4, and name the unbranched alcohols methanol to butanol

Methanol, ethanol, propanol, butanol

(c) describe the reactions of alcohols in terms of combustion and oxidation to carboxylic acids

Combustion
Undergo combustion to produce carbon dioxide and water

Oxidation
Most alcohol can undergo oxidation when heated with oxidizing agent and a carboxylic acid will be formed upon heating. E.g. of oxidizing agent: KMnO4
Conditions: Heat
Ethanol + Oxygen from oxidizing agent -> ethanoic acid + water
CH3CH2OH + 2[O] -> CH3COOH + H2

(d) describe the formation of ethanol by the catalysed addition of steam to ethene and by fermentation of glucose
Addition of stream to ethene (Hydration)
Main method for industrial production of ethanol
Conditions: 300℃, 60atm, Phosphoric acid
Ethene + Steam -> Ethanol

Fermentation of glucose
Alcohol drinks contain ethanol. To produce ethanol, fermentation is done. Microorganism act on carbohydrates in the absence of oxygen to produce ethanol and carbon dioxide. When fermentation is complete, the yeast settles to the bottom and alcohol is pumped out of the tank
Conditions: Yeast, Absence of carbon dioxide

Glucose -Yeast-> Ethanol + carbon dioxide
This is a slower reaction compared to the catalytic addition of steam to ethene. But this method is more sustainable as the starting materials are obtained from plants, which are renewable.

(e) describe the carboxylic acids as a homologous series containing the $–CO_2H$ group
CnH(2n+1)COOH
Carboxylic acids are organic compounds with the carboxyl functional group (-COOH) where n>= 0
Carboxylic acids have a higher MP and BP compared to alcohols with same number of carbons. There is a gradual increase in MP and BP as their molecular size is increasing. They are all soluble in water.

(f) draw the structures of carboxylic acids, C_1 to C_4, and name the unbranched acids methanoic acid to butanoic acid

Methanoic acid, ethanoic acid, propanoic acid, butanoic acid.

(g) describe the carboxylic acids as weak acids, reacting with carbonates, bases and some metals
Carboxylic acids are weak acids because they partially ionize in water to form a carboxylate ion and hydrogen ion. RCOOH <-> RCOO- + H+, where R is the alkyl group. They exhibit properties of acids

Reaction with metals
Ethanoic acid + Magnesium -> Magnesium ethanoate + Hydrogen
2CH3COOH + Mg -> (CH3COOH)2Mg + H2
The metal will remain at the back of the carboxylate ion when writing out the formula
Reaction with carbonates
Methanoic acid + Sodium carbonate -> Sodium methonate + Carbon dioxide + water
2HCOOH + Na2CO3 -> 2HCOONa + CO2 + H2O
Reaction with bases
Propanoic acid + Calcium hydroxide -> Calcium propanoate + water
2CH3CH2COOH + CaOH -> (CH3CH2COO)2Ca + H2O

(h) describe the formation of ethanoic acid by the oxidation of ethanol by atmospheric oxygen or acidified potassium manganate(VII)

Oxidation of ethanol by atmospheric oxygen
Ethanol reacts with the atmospheric oxygen in the presence of one type of bacteria that is usually present in the air to become ethanoic acid
Conditions: Bacteria
Ethanol + Oxygen in the air -Bacteria-> Ethanoic acid + water

Oxidation of ethanol by acidified potassium manganate(VII)
Prepared by heating mixture of ethanol and acidified and sulfuric acid and a redox reaction takes place.
Conditions: Heat
Ethanol + Oxygen from the oxidizing agent -heat-> Ethanoic acid + water
KMnO4 acts as a oxidizing agent which itself will get reduced. It will then change from purple to colorless. Ethanol acts as a reducing agent which is then oxidized to be a ethanoic acid.

(i) describe the reaction of a carboxylic acid with an alcohol to form an ester, e.g. ethyl ethanoate

Esters are organic compounds formed by the reaction of alcohols with carboxylic acids.

When carboxylic acids is warmed with alcohol and a few drops of sulfuric acid(catalyst), an ester and water is formed.

Carboxylic acid + Ethanol <-concentrated sulfuric acid-> Ester + Water

Ethanol + Ethanoic acid <-> Ethyl ethanoate + water

During the formation the -OH group from the ethanoic acid and -H from the ethanol acid is removed to form water.

The side where the C=O link is facing is the carboxylic acid part.

As ester is being formed, it can be referred to as esterification and condensation as there is loss of small molecules.

Although there is (-OH) group, it does not dissociate to form OH- ions in water, therefore alcohol is not a base and esterification is not acid-base reaction.

(j) deduce the name and formula of an ester from the unbranched carboxylic acid, C_1 to C_4, and alcohol, C_1 to C_4, and vice versa.
General formula: CnH(2n+1)COOCmH(2m+1)
Where CnH(2n+1)CO- is from the carboxylic acid and -OCmH(2m+1) is from the alcohol.
How to name: Split the formula at the COOC and the one with the c=c is the acid and the one without the double bond is the alcohol.

11.3 Polymers

(a) describe polymers as large molecules built up from small units (monomers), different polymers having different units and/or different linkages

Polymers is a very large organic molecule built up from many small units called monomers.

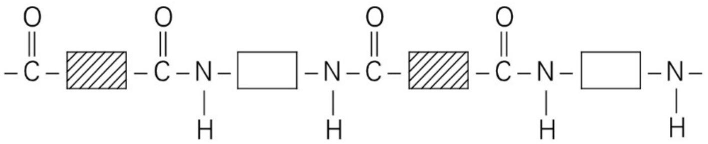
Different polymers are made up of different monomers which can also contain different linkage. The process of joining together large number of monomers to form polymer is called polymerization. The monomers in polymers are joined together by covalent bonds.

(b) describe the formation of poly(ethene) as an example of addition polymerisation of ethene as the monomer (see also **11.2**(j))
Addition polymerisation occurs when unsaturated monomers join together without losing any molecules or atoms to form addition polymers. The molecules contains carbon-carbon double/triple bonds. At high pressure and temperature, thousands of alkene join together to form polymers.
Poly(ethene) is the simplest addition. Produced from the addition of ethene molecules. At high temperatures and presence of catalyst, the C=C bonds break and each will bond with two other monomers. Poly(ethene) is then formed.

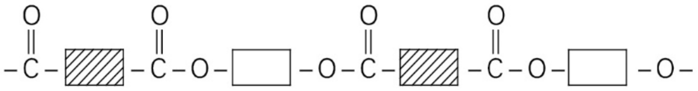
(c) state some uses of poly(ethene) as a typical plastic, e.g. plastic bags; clingfilm
Poly(ethene) can be easily molded into various shapes, thus it can be used to make cling wraps, plastic bags and toys.

(d) deduce the structure of the polymer product from a given monomer and vice versa
Monomer -> Polymer reduce the C=C bonds and ensure there is only 2 units
Polymer -> Monomer Detect the repeat units and add the C=C bonds.

(e) describe nylon, a polyamide, and *Terylene*, a polyester, as condensation polymers, the partial structure of nylon being represented as



and the partial structure of *Terylene* as



(details of manufacture and mechanisms of these polymerisations are **not** required)

Condensation polymerisation occurs when monomers combine to form condensation polymers with the removal of small molecules. There are 2 main groups of condensation polymers, Polyamides and polyesters.

Nylon-Polyamide
It can be made from a dicarboxylic acid (two -COOH) and a diamine (two -NH3)
When they react, OH is removed from the dicarboxylic acid and H is removed from the diamine where water molecules are formed for each pair of monomers.
The amide linkage is (OCNH) where is is C=O bond
Terylene – Polyester
It can be made from a dicarboxylic acid (two -COOH) and a diol (two -OH)
When they react, OH is removed from the dicarboxylic acid and H is removed from the diol and water molecule is produced from each pair of monomers.

The ester linkage is (COO) where there is one C=O bond.

(f) state some typical uses of man-made fibers such as nylon and *Terylene*, e.g. clothing; curtain materials; fishing line; parachutes; sleeping bags
Nylon: Used to make fishing lines
Terylene: Tents and sleeping bags

(g) describe the pollution problems caused by the disposal of non-biodegradable plastics
Plastics are non-biodegradable which means they cannot be broken down by bacteria and other living organisms in the soil. Where they will create land pollution when burying them in landfills creating built up waste. Water pollution: Plastic thrown into sea will endanger marine animals as they might mistaken them for food and choke on them.
Air pollution: When plastics are incinerated, it will produce poisonous gas.

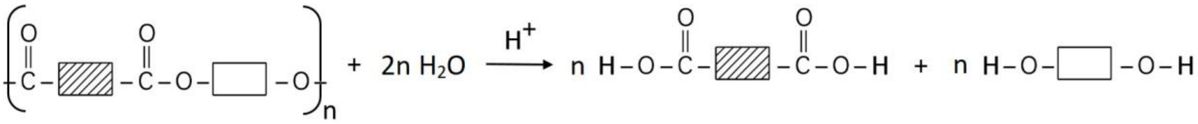
(h) describe two methods of recycling plastics as

(i) physical method (exemplified by melting small pieces of poly(ethene) waste into pellets)

Small pieces of plastic are cooled and pulled into strands and cut into pellets which can be used to make new products. The chemical composition of plastic remains unchanged.

(ii) chemical method (exemplified by depolymerization and cracking of plastic waste into chemical feedstock and fuel respectively)
Cracking: undergo cracking to form short chain of alkane and alkene where it can be used to make fuels or starting materials. This require a high temperature and the presence of catalyst.

describe depolymerisation as a process in which polymers are broken down into their monomers, exemplified by hydrolysis of polyesters using acid as a catalyst



(details of mechanisms are **not** required)

Depolymerization: process where polymers are broken into monomers. They undergo depolymerization through acid hydrolysis and down in water with the presence of acid. E.g. Polyester is broken down into dicarboxylic acid and diol.

discuss the social, economic and environmental issues of recycling plastics.

Social: More convenient to throw than recycle, many do not know the proper way to recycle plastic (contamination, increases effort too separate non recyclables), takes time for community to adopt the habit (not immediate and effective).

Economic: Expensive to do so (transporting the waste, sorting them out, require manpower, machine and energy)

Environmental: If waste water from recycling plant is not disposed of properly, it can cause water pollution.

12. Maintaining Air Quality

Learning Outcomes

Candidates should be able to:

(a) describe the volume composition of gases present in dry air as being approximately 78% nitrogen, 21% oxygen and the remainder being noble gases (with argon as the main constituent) and carbon dioxide

78% nitrogen, 21% oxygen, 0.04 Carbon dioxide), 0.96% noble gases (mainly argon), water vapor and other gases.

(b) name some common atmospheric pollutants, e.g. carbon monoxide; methane; nitrogen oxides (NO and NO_2); ozone; sulfur dioxide; unburned hydrocarbons

Common air pollutant: Sulfur dioxide, Nitrogen oxides, Carbon monoxide, Methan, Unburnt hydrocarbons, Ozone.

(c) state the sources of these pollutants as

(i) carbon monoxide from incomplete combustion of carbon-containing substances

Incomplete combustion of carbon based fuel

(ii) nitrogen oxides from lightning activity and internal combustion engines

Vehicle combustion engines and lightning

(iii) sulfur dioxide from volcanoes and combustion of fossil fuels

Combustion of fossil fuels, Volcanic eruptions

(d) describe the reactions used in possible solutions to the problems arising from some of the pollutants named in (b)

(i) the redox reactions in catalytic converters to remove combustion pollutants (see also **10**(d))

It is located midway of the exhaust system of a vehicle. It is made of platinum, palladium, rhodium that speeds up the conversion of harmful substance in the vehicle exhaust into less harmful substances. It is arranged in honeycomb structure to maximize surface area for gas.

Carbon monoxide + Nitrogen monoxide -> Carbon dioxide + Nitrogen

carbon monoxide + Oxygen -> Carbon dioxide

Nitrogen dioxide -> nitrogen + Oxygen

Octane + Oxygen -> Carbon dioxide + water vapor.

(ii) the use of calcium carbonate to reduce the effect of ‘acid rain’ and in flue gas desulfurization

Desulfurization can remove a significant proportion of sulfur dioxide in flue gas before it is released into the atmosphere. By using wet scrubbing of $CaCO3$ slurry. The slurry is formed when calcium carbonate is mixed with water. The slurry droplets sprayed through the flue gas and calcium carbonate reacts with sulfur dioxide to form calcium sulfite and carbon dioxide. It is further oxidized to form calcium sulfate by atmospheric oxygen.

It is then hydrated to form gypsum which can be used to make fertilizers.

Liming is adding $CaCO3$ to the soil to remove some of the excess acid contributed by the acid rain.

(e) discuss some of the effects of these pollutants on health and on the environment

(i) the toxic nature of carbon monoxide

It will bind irreversibly with the hemoglobin in our red blood cell which will lower the ability of hemoglobin to transport oxygen to the rest of the body. It can result in loss of consciousness and death.

(ii) the role of nitrogen dioxide and sulfur dioxide in the formation of ‘acid rain’ and its effects on respiration and buildings

Sulfur dioxide causes breathing difficulties in humans and animals.

Sulfur dioxide and nitrogen dioxide react with oxygen in the air to form acidic compounds which dissolve in rainwater, forming solutions that will kill aquatic life and corrode buildings.

$SO2$ -> $SO4$ -> $H2SO4$, NO -> $NO3$ -> $HNO3$

(f) discuss the importance of the ozone layer and the problems involved with the depletion of ozone by reaction with chlorine-containing compounds, chlorofluorocarbons (CFCs)

Ozone layer blocks out UV radiation. It can absorb UV radiation thus reducing the amount of UV radiation before reaching the earth’s surface.

Ozone is allotrope of oxygen. Ozone molecules will break to reform oxygen when they absorb UV radiation. ($3O2 \leftrightarrow 2O3$)

CFCs found in propellants and certain kinds of plastic. They contain chlorine, fluorine, carbon where they are released into the stratosphere and react with UV to produce chlorine atoms which react with the ozone.

UV breaks up CFC in the stratosphere into chlorine atoms. The chlorine atoms react with the ozone to form oxygen and chlorine oxide. The chlorine oxide continue to break up another ozone molecule. The chlorine atom is released and continues to attack another ozone molecule, forming a chain reaction. The hole in the ozone layer will start to form and will allow UV radiation to reach the surface.

(g) describe the carbon cycle in simple terms, to include

(i) the processes of combustion, respiration and photosynthesis

Combustion: Most fossil fuels contain mostly methane. They are burnt in the presence of oxygen to generate energy. Carbon dioxide is a product of combustion.

Respiration: Living organisms gain the energy for life process by breaking down glucose back into carbon dioxide and water which are released into the atmosphere.

Photosynthesis: Plants take in carbon dioxide and water to make glucose and oxygen in the presence of sunlight. Glucose is used in other process that sustain plant life.

(ii) how the carbon cycle regulates the amount of carbon dioxide in the atmosphere

Carbon cycle is a set of processes that regulates the amount of carbon dioxide in

the atmosphere.

(h) state that carbon dioxide and methane are greenhouse gases and may contribute to global warming; give the sources of these gases and describe the potential effects of increased levels of these greenhouse gases, including more extreme weather events and melting of polar ice.

Global warming: Increase in the average temperature of the earth's surface due to increasing amounts of greenhouse gases in the atmosphere.

Solar radiation reaches the earth's atmosphere and some is reflected into space and some passes through the atmosphere and becomes heat. As the earth surface is warmed, some of the heat rises and escape into space. The remaining heat is trapped by greenhouse gases, further warming the planet.

Major sources of greenhouse are carbon dioxide and methane.

Change in rainfall patterns

Shifting wind patterns can lead to changes in rainfall patterns. Where lush areas receive too little rain and turn into deserts while other areas might receive too much rain and become flooded. The fertile land would become barren and might not be able to produce food and causing food shortages.

Heat waves

It will become more common. Some parts might become too hot to live. High temperatures might make wildfires more common which can destroy the ecosystem.

Tropical storms

They will form over warm waters and move westwards, bringing strong winds, sheets of rain and high tide. It is destructive and can endanger human lives, property, farmland. The warmer the ocean, the stronger the storms.

Ocean acidification

Warmer ocean water absorb carbon dioxide more easily forming carbonic acid and decreasing the pH level of the water. The crustaceans depend on carbonate materials to form shell and the acidified water remove the carbonate material. Existing shells will also dissolve, making the organism more vulnerable to injury and predation.

Melting of polar ice caps

Increase in global temperatures will cause ice caps to melt and shrink. Increasing sea level.