The Little Red Book 10th Edition

GLOSSARY OF TERMS USED IN CHEMISTRY PAPERS

It is hoped that the glossary (which is relevant only to science papers) will prove helpful to candidates as a guide, i.e. it is neither exhaustive nor definitive. The glossary has been deliberately kept brief not only with respect to the number of terms included but also to the descriptions of their meanings. Candidates should appreciate that the meaning of a term must depend in part on its context.

1. *Calculate* is used when a numerical answer is required. In general, working should be shown, especially where two or more steps are involved.

2. *Classify* requires candidates to group things based on common characteristics.

3. *Comment* is intended as an open-ended instruction, inviting candidates to recall or infer points of interest relevant to the context of the question, taking account of the number of marks available.

4. *Compare* requires candidates to provide both similarities and differences between things or concepts.

5. *Construct* is often used in relation to chemical equations where a candidate is expected to write a balanced equation, not by factual recall but by analogy or by using information in the question.

6. *Define (the term(s)...)* is intended literally. Only a formal statement or equivalent paraphrase being required.

7. *Describe* requires candidates to state in words (using diagrams where appropriate) the main points of the topic. It is often used with reference either to particular phenomena or to particular experiments. In the former instance, the term usually implies that the answer should include reference to (visual) observations associated with the phenomena. In the latter instance the answer may often follow a standard pattern, e.g. Apparatus, Method, Measurement, Results and Precautions.

   In other contexts, *describe and give an account of* should be interpreted more generally, i.e. the candidate has greater discretion about the nature and the organisation of the material to be included in the answer. *Describe and explain* may be coupled in a similar way to *state and explain.*

8. *Determine* often implies that the quantity concerned cannot be measured directly but is obtained by calculation, substituting measured or known values of other quantities into a standard formula.

9. *Discuss* requires candidates to give a critical account of the points involved in the topic.

10. *Estimate* implies a reasoned order of magnitude statement or calculation of the quantity concerned, making such simplifying assumptions as may be necessary about the points of principle and about values of quantities not otherwise included in the question.

11. *Explain* may imply reasoning or some reference to theory, depending on the context.

12. *Find* is a general term that may be variously interpreted as calculate, measure, determine etc.

13. *List* requires a number of points, generally each of one word, with no elaboration. Where a given number of points is specified, this should not be exceeded.

14. *Measure* implies that the quantity concerned can be directly obtained from a suitable measuring instrument, e.g. length, using a rule, or angle, using a protractor.

15. *Outline* implies brevity, i.e. restricting the answer to giving essentials.

16. *Predict* or *deduce* implies that the candidate is not expected to produce the required answer by recall but by making a logical connection between other pieces of information. Such information may be wholly given in the question or may depend on answers extracted from an earlier part of the question. *Predict* also implies a concise answer with no supporting statement required.

17. *Sketch*, when applied to graph work, implies that the shape and/or position of the curve need only be qualitatively correct, but candidates should be aware that, depending on the context, some quantitative aspects may be looked for, e.g. passing through the origin, having the intercept, asymptote or discontinuity at a particular value.

    In diagrams, *sketch* implies that a simple, freehand drawing is acceptable; nevertheless, care should be taken over proportions and the clear exposition of important details.

18. *State* implies a concise answer with little or no supporting argument, e.g. a numerical answer that can be obtained 'by inspection'.

19. *Suggest* is used in two main contexts, i.e. either to imply that there is no unique answer, or to imply that candidates are expected to apply their general knowledge to a 'novel' situation, one that may be formally 'not in the syllabus'.

20. *What do you understand by/What is meant by (the term(s)...)* normally implies that a definition should be given, together with some relevant comment on the significance or context of the term(s) concerned, especially where two or more terms are included in the question. The amount of supplementary comment intended should be interpreted in light of the indicated mark value.

1 Experimental Chemistry

1.1 Experimental design

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(a) name appropriate apparatus for the measurement of time, temperature, mass and volume, including burettes, pipettes, measuring cylinders and gas syringes	See table below	

Type of quantity to measure	Apparatus	Accuracy
Time	Electronic Stopwatch (0 d.p.)	Nearest s (0 d.p.)
Temperature	Thermometer x	Nearest 0.5 °C (1 d.p.)
Mass	Electronic Balance	Nearest 0.01 g (2 d.p.)
Volume of solution	Burette	Nearest 0.05 cm^3 (2 d.p.)
	Pipette	Fixed volumes of 10.0, 20.0, 25.0 or 50.0 cm^3 (1 d.p.)
	Measuring cylinder	Nearest 0.5 cm^3 (1 d.p.)
Volume of gas	Gas syringe	-

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(b) suggest suitable apparatus, given relevant information, for a variety of simple experiments, including collection of gases and measurement of rates of reaction (Refer to 6.1 Speed of Reaction)	See tables below	**Tip:** The density of a gas is roughly proportional to its relative molecular mass. Air has an average M_r of about 29, so any gas with M_r heavier than that will be denser than air.

Method of collection of gases	Suitable gases to collect
Downward delivery of gas	Only for gas that are **denser than air**. The gases are CO_2, SO_2, Cl_2, O_2 and HCl.
Upward delivery of gas	Only for gas that are **less dense than air**. The gases are H_2 and NH_3
Downward displacement of water	Only for gas that **hardly dissolves in water**. The gases are CO_2, O_2 and H_2. *Cannot collect Cl_2, SO_2, HCl and NH_3 using this method as these gases are soluble in water.
Collection using gas syringe	Suitable for all gases

Drying agent	Gases
Fused/Anhydrous calcium chloride	Can be used for most gases EXCEPT for NH_3
Calcium oxide (Quicklime)	Can be used for NH_3 and all neutral gases (e.g. O_2 and H_2)
Concentrated sulfuric acid	Can be used for most gases EXCEPT alkaline gases such as NH_3 acid-base

alkali
ammonia - base
acid reaction

1.2 Methods of purification and analysis

(a) describe methods of separations and purification for the components of the following types of mixtures:
(i) solid-solid
(ii) solid-liquid
(iii) liquid-liquid (miscible and immiscible)
Techniques to be covered for separations and purification include:
(i) use of a suitable solvent, filtration and crystallisation or evaporation
(ii) sublimation
(iii) distillation and fractional distillation
(iv) use of a separating funnel
(v) paper chromatography

Type of Mixtures	**Methods of separation**	**Example**
Two solids, one of which is soluble in a suitable solvent	Use suitable solvent to dissolve, then filter	Sand and salt, by using water
Insoluble solid + liquid/solution	Filtration	Sand and water
Solution (solid dissolved in liquid (solvent))	Crystallisation (for solid which will decompose upon heating)	Obtain copper (II) sulfate crystals from saturated copper(II) sulfate solution
	Evaporation to dryness (for solid that **will not** be decomposed upon heating)	Obtain NaCl from seawater
	Simple Distillation (to obtain liquid (solvent) from the solution)	Obtain methanol from mixture of methanol and water
Two solids, one of which sublimes	Sublimation	Iodine and sand Substances that sublime: Iodine (I_2), ammonium chloride, dry ice (solid CO_2)
Two immiscible liquids	Use of separating funnel	Oil and water
Two miscible liquids (liquid dissolved in liquid)	Fractional distillation	Air (to obtain nitrogen for Haber Process), fractional distillation of crude oil, ethanol from water (fermentation)
Mixture of organic compounds such as coloured dyes and amino acids present in <u>small amounts</u>	Paper chromatography	Coloured dyes in inks and food

Note:

Miscible liquids are liquids that dissolve in each other completely to form a solution.

Immiscible liquids are liquids that do not mix/ cannot dissolve in each other.

Students' commonly mistook the word 'Filtrate' for the process of filtration, when in fact 'Filter' should be the correct term to be used. Note that 'Filtrate' is the liquid that is obtained after filtration.

	Eg. Separating a mixture of sand and salt. Add water to mixture to dissolve salt. Filter the mixture. Sand is obtained as residue. Evaporate the filtrate to dryness to obtain salt. ***Eg. Separating a mixture of sand and water*** Filter mixture to remove sand as residue. Water will be obtained as filtrate. ***Eg. Separating a mixture of calcium carbonate (insoluble in water) and copper (II) sulfate solution.*** Filter mixture to remove calcium carbonate as residue. Evaporate **most** water from filtrate to obtain a saturated copper(II) sulfate solution. Cool the hot saturated solution to obtain pure copper(II) sulfate crystals. **Filter the solution to obtain the crystals, wash the crystals with a little cold distilled water and dry crystals between pieces of filter paper.** ***Eg. Separating a mixture of methanol and water (Simple distillation)*** Heat mixture to vaporise the two liquids. The hot vapour rise and enter the condenser. The liquid with lower boiling point (methanol) will be distilled first while the one with the higher boiling point (water) will distil later. ***Eg. Separating a mixture of iodine and sand.*** Place mixture in an evaporating dish and place an inverted funnel over the mixture. Upon gentle heating, all iodine sublimes and deposits on the cold filter funnel. Sand is left in the evaporating dish after the sublimation. ***Eg. Separating a mixture of oil and water.*** Place the mixture in the separating funnel. The less dense liquid (oil) will be on top of the denser liquid (water). The 2 liquids can be separated by manipulating the tap. ***Eg. Separating petrol from crude oil. (Fractional distillation)*** Refer to 11.1(b) Fuels and Crude Oil	**Note:** The solid that remains on the filter paper is called **residue**. The liquid or solution that passes through the filter paper is called the **filtrate**.
(b) describe paper chromatography and interpret chromatograms including comparison with 'known' samples and the use of Rf values	A spot of the food colouring is applied to the chromatography paper. The chromatography paper is dipped in the solvent with the pencil line above the solvent level. The solvent travels up the paper, carrying the dyes along. A dye that is more soluble in the solvent will travel further. $R_f = \frac{\text{distance travelled by the substance}}{\text{distance travelled by the solvent}}$ The R_f value of a substance does not change as long as chromatography is carried out under the same conditions (i.e. same solvent and same temperature).	R_f value has no units. It must be expressed in decimal places and not as fractions.

(c) explain the need to use locating agents in the chromatography of colourless compounds	Locating agents are used to analyse the components in the sample which are colourless **by making the colourless components visible.**	
(d) deduce from the given melting point and boiling point the identities of substances and their purity	Pure substances have fixed melting and boiling points. Impure substance melts/boils over a range of temperature. mixture The melting point is lowered and boiling point is increased when impurities are present.	

1.3 Identification of ions and gases

(a) describe the use of aqueous sodium hydroxide and aqueous ammonia to identify the following aqueous cations: aluminium, ammonium, calcium, copper(II), iron(II), iron(III), lead(II) and zinc (formulae of complex ions are not required)

Cation present	**Addition of a few drops of NaOH (aq)**	**Addition of excess NaOH (aq)**	**Ionic equation for precipitate formed**
Calcium, Ca^{2+}	White precipitate of $Ca(OH)_2$ is formed.	White precipitate formed is insoluble in excess NaOH (aq).	$Ca^{2+}(aq) + 2OH^-(aq) \rightarrow Ca(OH)_2(s)$
Zinc, Zn^{2+}	White precipitate of $Zn(OH)_2$ is formed.	White precipitate formed is soluble in excess NaOH (aq) to give a colourless solution.	$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$
Aluminium, Al^{3+}	White precipitate of $Al(OH)_3$ is formed.	White precipitate formed is soluble in excess NaOH (aq) to give a colourless solution.	$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$
Lead(II), Pb^{2+}	White precipitate of $Pb(OH)_2$ is formed.	White precipitate formed is soluble in excess NaOH (aq) to give a colourless solution.	$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$

Tip:

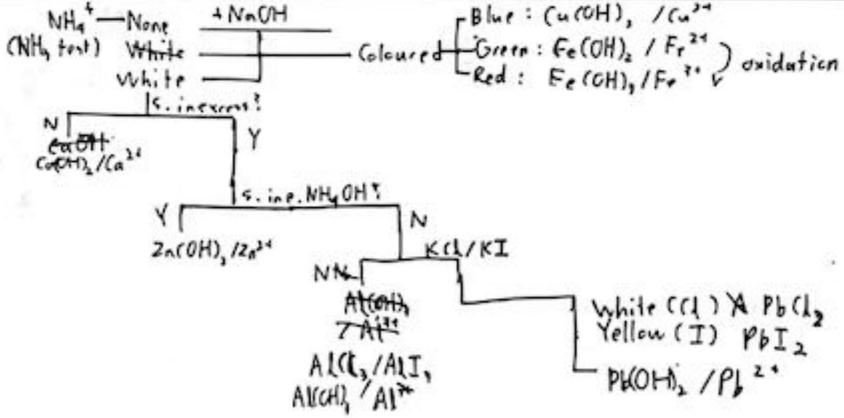
Na^+, K^+ and NH_4^+ cations will not form any precipitates when NaOH(aq) or NH_3(aq) is added.

How do we distinguish Al^{3+} from Pb^{2+}?

KCl(aq) / KI(aq) can be used to distinguish Al^{3+} from Pb^{2+}.

Pb^{2+} will react with KCl(aq) to form white precipitate ($PbCl_2$) or KI(aq) to form yellow precipitate (PbI_2) while Al^{3+} will not form a precipitate when KCl(aq) / KI(aq) is added.

Copper(II), Cu^{2+}	Light blue precipitate of $Cu(OH)_2$ is formed.	Light blue precipitate formed is insoluble in excess NaOH (aq).	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
Iron(II), Fe^{2+}	Dirty green precipitate of $Fe(OH)_2$ is formed.	Dirty green precipitate formed is insoluble in excess NaOH (aq).	$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
Iron(III), Fe^{3+}	Reddish brown precipitate of $Fe(OH)_3$ is formed.	Reddish brown precipitate formed is insoluble in excess NaOH (aq).	$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Ammonium, NH_4^+	No precipitate is formed. Upon heating, colourless and pungent gas is evolved. Gas turns damp red litmus paper blue. Gas is NH_3.	No visible reaction	No precipitate is formed.



Cation present	**Addition of a few drops of $NH_3(aq)$**	**Addition of excess $NH_3(aq)$**	**Ionic equation for precipitate formed**
Calcium, Ca^{2+}	No visible reaction		
Zinc, Zn^{2+}	White precipitate of $Zn(OH)_2$ is formed.	White precipitate formed is soluble in excess NH_3 (aq) to give a colourless solution.	$Zn^{2+}(aq) + 2OH^-(aq) \rightarrow Zn(OH)_2(s)$
Aluminium, Al^{3+}	White precipitate of $Al(OH)_3$ is formed.	White precipitate formed is insoluble in excess NH_3 (aq).	$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$
Lead(II), Pb^{2+}	White precipitate of $Pb(OH)_2$ is formed.	White precipitate formed is insoluble in excess NH_3 (aq).	$Pb^{2+}(aq) + 2OH^-(aq) \rightarrow Pb(OH)_2(s)$
Copper(II), Cu^{2+}	Light blue precipitate of $Cu(OH)_2$ is formed.	Light blue precipitate formed is soluble in excess NH_3 (aq) to give a dark blue solution.	$Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)$
Iron(II), Fe^{2+}	Dirty green precipitate of $Fe(OH)_2$ is formed.	Dirty green precipitate formed is insoluble in excess NH_3 (aq).	$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$
Iron(III), Fe^{3+}	Reddish-brown precipitate of $Fe(OH)_3$ is formed.	Reddish-brown precipitate formed is insoluble in excess NH_3 (aq).	$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$
Ammonium , NH_4^+	No reaction.		

(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 lead(II) nitrate); nitrate (by reduction with aluminium and aqueous sodium hydroxide to ammonia and subsequent use of litmus paper) and sulfate (by reaction of an aqueous solution with nitric acid and aqueous barium nitrate)

Anion present	**Test**	**Observations**	**Ionic equation**
Carbonate, CO_3^{2-}	Add dilute any /acid hydrochloric acid. Pass the gas evolved into limewater.	Effervescence is seen. Colourless, odourless gas evolved. Gas forms a white precipitate when passed through limewater. Gas is CO_2.	$2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$
Chloride, Cl^-	Add dilute nitric acid. Then add silver nitrate solution.	A white precipitate of silver chloride is formed.	$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
Iodide, I^-	Add dilute nitric acid. Then add silver nitrate solution.	A yellow precipitate of silver iodide is formed.	$Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$
Nitrate, NO_3^-	Add aqueous sodium hydroxide. Then add a piece of aluminium foil and warm the mixture. Test the gas evolved with damp red litmus paper.	A colourless, pungent gas is evolved. Gas turns damp red litmus paper blue. Gas is NH_3.	$3NO_3^-(aq) + 8Al(s) + 5OH^-(aq) + 18H_2O(l) \rightarrow 3NH_3(g) + 8[Al(OH)_4]^-(aq)$
Sulfate, SO_4^{2-}	Add dilute nitric acid. Then add barium nitrate solution.	A white precipitate of barium sulfate is formed.	$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

Note:

Addition of nitric acid is for the removal of soluble carbonate ions present in the solution before the test of anions.

* Reason For $AgNO_3$ / $BaNO_3$: NO_3^- salts always soluble

(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))

Gas	Test	Observations
Hydrogen, H_2	Place a lighted splint at the mouth of the test tube.	Colourless and odourless gas evolved. Gas extinguished **lighted** splint with a 'pop' sound.
Oxygen, O_2	Insert a glowing splint into the test tube.	Colourless and odourless gas evolved. The **glowing** splint is rekindled.
Carbon dioxide, CO_2	Bubble gas through limewater.	Colourless and odourless gas evolved. Gas forms white precipitate when passed through limewater.
Chlorine, Cl_2	Place a piece of damp blue litmus paper at the mouth of the test tube.	Greenish-yellow and pungent gas evolved. Gas turns damp blue litmus paper red and then bleaches it.
Sulfur dioxide, SO_2	Place a piece of filter paper soaked with acidified potassium manganate (VII) at the mouth of the test tube.	Colourless and pungent gas evolved. The gas turns aqueous acidified potassium manganate (VII) from purple to colourless.
Ammonia, NH_3	Place a piece of damp red litmus paper at the mouth of the test tube.	Colourless and pungent gas evolved. The damp red litmus paper turns blue.

* Can use damp red litmus paper for Cl_2 test,
- Gas bleaches litmus paper

* Cl_2 bleaches all liquid pH indicators

<table>
<tr><td colspan="3">**2 The Particulate Nature of Matter**</td></tr>
<tr><td colspan="3">**2.1 Kinetic particle theory**</td></tr>
<tr><td>**Learning Outcomes**</td><td>**Key Concepts**</td><td>**TIPS/ Common mistakes**</td></tr>
<tr><td>(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</td><td>
<table>
<tr><td></td><td>Solid</td><td>Liquid</td><td>Gas</td></tr>
<tr><td>Arrangement</td><td>**Particles are closely packed** in a **regular/orderly** arrangement</td><td>**Particles are less closely packed than in solid** in a **disorderly** arrangement</td><td>**Particles are disorderly arranged and very far apart**</td></tr>
<tr><td>Movement</td><td>**vibrate and rotate about their fixed positions**</td><td>**slide over neighbouring particles** and move freely throughout the liquid</td><td>**move freely in all directions at high speed**</td></tr>
</table>

Melting and boiling:
During melting and boiling, particles absorb heat energy. Hence, particles **gain kinetic energy** and begin to **move faster** and **further apart**.
Temperature remains constant at melting/boiling point as **heat energy absorbed/taken in** by the particles is **equal** to the energy used to **overcome the forces of attraction between the particles**.

Freezing and condensation:
During freezing and condensation, particles lose heat energy. Particles **lose kinetic energy** and move **slower** and **closer**.

Sublimation: change of state from solid to gas **without going through liquid state**.</td><td>**Tip:**
Candidate should describe the contrast in **arrangement** and **movement of particles** during the interconversion of states.</td></tr>
<tr><td>(b) describe and explain evidence for the movement of particles in liquids and gases (the treatment of Brownian motion is not required)</td><td>Diffusion is the **process** by which **particles move freely to fill up any available space.**

Diffusion **is the movement of particles** from a region of **higher concentration** to a region of **lower concentration.**</td><td></td></tr>
<tr><td colspan="2">(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</td><td></td></tr>
</table>

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<tr>
<td>(d) state qualitatively the effect of molecular mass on the rate of diffusion and explain the dependence of rate of diffusion on temperature.</td>
<td>Effect of **relative molecular mass** on rate of diffusion:
Molecules with **lower relative molecular mass diffuse faster / higher relative molecular mass diffuse slower.**
Eg. The **relative molecular mass** of ammonia gas, **17** is **lower** than that of carbon dioxide which has **relative molecular mass of 44**. Hence, ammonia will **diffuse** faster and have a **higher rate of diffusion** than carbon dioxide.

Effect of **temperature** on rate of diffusion:
When the temperature is higher, the particles **gain kinetic energy** and diffuse **faster**. Hence the **rate of diffusion increases**. The higher the temperature, the higher the rate of diffusion.</td>
<td>**Tip:**
Candidates should indicate the relative molecular masses of the substances compared.
Mr is used, **NOT** density/mass /weight.
Mr has no unit.</td>
</tr>
<tr>
<td colspan="3">**2.2 Atomic structure**</td>
</tr>
<tr>
<td>(a) state the relative charges and approximate relative masses of a proton, a neutron and an electron</td>
<td>
<table>
<tr><th></th><th>Proton</th><th>Neutron</th><th>Electron</th></tr>
<tr><td>Relative charge</td><td>+1</td><td>0</td><td>-1</td></tr>
<tr><td>Relative mass</td><td>1</td><td>1</td><td>$\frac{1}{1840}$</td></tr>
</table>
</td>
<td></td>
</tr>
<tr>
<td colspan="2">(b) describe, with the aid of diagrams, the structure of an atom as containing protons and neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels)

Note: The chemical properties (how it reacts) of an element depend on the number of **valence** electrons.</td>
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</tr>
<tr>
<td>(c) define proton (atomic) number and nucleon (mass) number</td>
<td>Proton (atomic) number is the number of **protons** in an atom.
Nucleon number is the total number of **protons** and **neutrons** in an atom.</td>
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</tr>
<tr>
<td>(d) interpret and use symbols such as $^{12}_{6}C$</td>
<td colspan="2">$$\begin{array}{r} \text{Mass number (total no. of protons and neutrons)} \longleftarrow 12 \\ \text{proton (atomic) number (no. of protons)} \longleftarrow 6 \end{array} C$$</td>
</tr>
<tr>
<td>(e) define the term isotopes</td>
<td colspan="2">Isotopes are **atoms of the same element** with the **same number of protons** but **different number of neutrons.**
E.g. ^{37}Cl and ^{35}Cl have 17 protons each but ^{37}Cl has 20 neutrons and ^{35}Cl has 18 neutrons.</td>
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<tr>
<td>(f) deduce the numbers of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers</td>
<td>No. of protons = atomic or proton number from Periodic Table
No. of electrons = no of protons (for **atoms**; not for ions)
No. of neutrons in an atom = nucleon number – proton number

Tip: An atom and ion of the same element, E.g. F & F$^-$, have the same number of protons and neutrons (so they have the same mass number). The only difference is in the number of electrons. F has 9 electrons, F$^-$ has 10 electrons.</td>
<td></td>
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</table>

2.3 Structure and properties of materials

(a) describe the differences between elements, compounds and mixtures	An element is a pure substance that **cannot be broken down into simpler substances** by chemical processes or by electricity. A compound is a pure substance made of **two or more different elements chemically combined** in a **fixed ratio.** A mixture is made up of **two or more substances** that are **not chemically combined**.	**Note:** Noble gases exist as **monatomic** elements (e.g. He). Most non-metal elements exist as **molecules** (e.g. H_2). A **molecule** consists of **two or more non-metal atoms chemically combined together.** (e.g. Cl_2, H_2O).

	Mixture	**Compound**
Separation	The components of a mixture can be separated by **physical methods**.	The elements in a compound can only be separated by **chemical reactions** or by **using electricity**.
Properties	The chemical properties of a mixture are the **same** as those of its components.	The physical and chemical properties of a compound are **different** from those of the elements in the compound.
Energy Change	No chemical reaction takes place when a mixture is formed -- usually there is little or **no energy change**.	A chemical reaction takes place when a compound is formed -- usually there is an **energy change**.
Composition	The components of a mixture can be mixed in **any** proportion.	The elements in a compound are always combined in a **fixed** proportion.

(b) compare the structure of simple molecular substances, e.g. methane; iodine, with those of giant molecular substances, e.g. poly(ethene); sand (silicon dioxide); diamond; graphite in order to deduce their properties	E.g. Comparing boiling points of methane and silicon dioxide in terms of **bonding and structure**. Methane has a **simple molecular structure** while silicon dioxide has a **giant molecular structure**. **Less energy** is required to **overcome the weak intermolecular forces of attraction** in methane than to **break the strong covalent bonds** between the silicon and oxygen **atoms**. Methane and silicon dioxide exist **as molecules** and **do not conduct electricity in any states** due to the **absence of mobile electrons** and **ions.**

<table>
<tr>
<td>(c) compare the bonding and structures of diamond and graphite in order to deduce their properties such as electrical conductivity, lubricating or cutting action (candidates will not be required to draw the structures)</td>
<td>**Electrical conductivity:**
Each carbon atom in **graphite** is covalently bonded to 3 other carbon atoms. 1 **electron** per C atom is delocalised. The presence of **delocalised mobile electrons** allow graphite to conduct electricity.

Each carbon atom in **diamond** is covalently bonded to 4 other carbon atoms. All valence electrons are involved in bonding and hence **no mobile electrons** to conduct electricity.

Use as lubricant:
Little amount of energy is required to **overcome the weak van der Waals' forces of attraction** between the **layers of carbon atoms** in graphite. Hence the **layers** of atoms can **slide over each other easily**.

Use as cutting tool:
Diamond is hard due to the **strong covalent bonds between the carbon atoms** in the giant molecular structure.</td>
<td>**Note:**

Mobile electrons OR mobile ions are to be present to conduct electricity.

Note:

Each carbon atom forms **4** single covalent bonds.</td>
</tr>
<tr>
<td>(d) deduce the physical and chemical properties of substances from their structures and bonding and vice versa</td>
<td>
<table>
<tr><th>Physical properties</th><th>Metal</th><th>Ionic</th><th>Simple molecular</th><th>Giant molecular</th></tr>
<tr><td>**Meting & boiling points**</td><td>High</td><td>High</td><td>Low</td><td>High</td></tr>
<tr><td>**Electrical conductivity**</td><td>Conducts in solid and molten states</td><td>Conducts in molten and aqueous states</td><td>Does not conduct in any states</td><td>All are non-conductors EXCEPT graphite</td></tr>
<tr><td>**Solubility in water**</td><td>Insoluble (some metals can react with water)</td><td>Depends (see solubility table)</td><td>Refer to solubility of gases</td><td>Insoluble</td></tr>
</table>

Refer to 2.3 (b), (c), 2.4 (c), (e) and 2.6 (a), (b) for explanations in terms of bonding and structures.</td>
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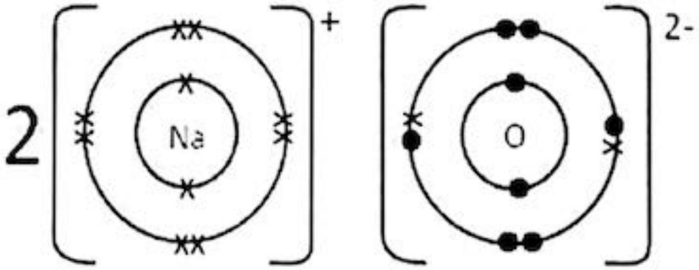
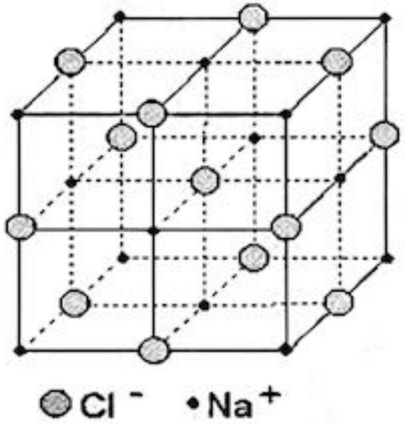
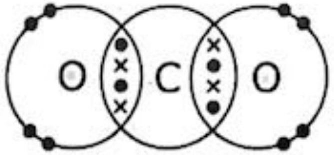
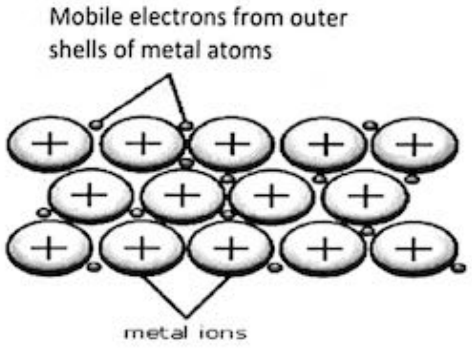
<tr>
<td colspan="3">**2.4 Ionic bonding**</td>
</tr>
<tr>
<td>(a) describe the formation of ions by electron loss/gain in order to obtain the electronic configuration of a noble gas</td>
<td>**Ionic structures:**
Correct number of valence electrons; diagram should show the next completely filled valence shell.
Correct charge on respective ions;
Correct number of electrons gained or lost.
 </td>
<td></td>
</tr>
</table>

(b) describe the formation of ionic bonds between metals and non-metals, e.g. NaCl; MgCl$_2$	Metals lose electrons to form cations (positively-charged ions). Non-metals gain electrons to form anions (negatively charged ions).	
(c) state that ionic materials contain a giant lattice in which the ions are held by electrostatic attraction, e.g. NaCl (candidates will not be required to draw diagrams of ionic lattices)	**Ionic compounds** have **giant ionic structures** where the **oppositely charged ions** are held by **strong electrostatic forces of attraction**.	
(d) deduce the formulae of other ionic compounds from diagrams of their lattice structures, limited to binary compounds	Eg.  Each Na$^+$ ion is surrounded by six Cl$^-$ ions. Each Cl$^-$ ion is surrounded by six Na$^+$ ions. Therefore, Ratio of Na$^+$: Cl$^-$ = 1:1, therefore the formula of sodium chloride is NaCl.	
(e) relate the physical properties (including electrical property) of ionic compounds to their lattice structure	High melting point: Ionic compounds have **giant ionic** lattice **structure. Large amount of energy** is required to **overcome the strong electrostatic forces of attraction** between the **oppositely charged ions**. Electrical conductivity: Ionic compounds **can conduct electricity in the molten and aqueous states** due to the presence of **mobile ions**. Ionic compounds **cannot conduct electricity in solid state** as **ions are at their fixed positions** in solid state, hence there is **no mobile ions** to conduct electricity.	

2.5 Covalent bonding		
(a) describe the formation of a covalent bond by the sharing of a pair of electrons in order to gain the electronic configuration of a noble gas	Note: Every pair of electrons shared constitutes a single covalent bond.	
(b) describe, using 'dot-and-cross' diagrams, the formation of covalent bonds between nonmetallic elements, e.g. H_2; O_2; H_2O; CH_4; CO_2	Correct number of valence electrons; diagram should show the next completely filled valence shell. Correct number of shared electrons. 1m - bonding e^- 1m - total e^- 	
(c) deduce the arrangement of electrons in other covalent molecules		
(d) relate the physical properties (including electrical property) of covalent substances to their structure and bonding	refer to 2.3(b)	
2.6 Metallic Bonding		
(a) describe metals as a lattice of positive ions in a 'sea of electrons'	Structure of metals: a lattice of regularly arranged **positive ions surrounded by a sea of mobile electrons.** 	**Note:** Candidate to draw the same number of positive ions and electrons (1:1 ratio). The positive ions are arranged in regular pattern and the sea of electrons is in random arrangement to symbolise the mobility of electrons.
(b) relate the electrical conductivity of metals to the mobility of the electrons in the structure	Metals conduct electricity in solid and molten states due to the presence of **mobile electrons** (NOT mobile ions!).	

3 Formulae, Stoichiometry and the Mole Concept

Learning Outcomes

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

Key Concepts

Ionic compounds – formed when bonding occurs between metals and non-metals
In general, metals lose electrons to form cations and non-metals gain electrons to form anions.
Hydrogen can lose/gain electrons to form H^+ or H^- ions.
Group I, II and III elements form X^+, X^{2+} and X^{3+} ions respectively.
Group V, VI and VII elements form Y^{3-}, Y^{2-} and Y^- ions respectively.

Name of ion	*Formula of ion*	*Name of ion*	*Formula of ion*
potassium	K^+	chloride	Cl^-
sodium	Na^+	bromide	Br^-
calcium	Ca^{2+}	iodide	I^-
magnesium	Mg^{2+}	oxide	O^{2-}
aluminium	Al^{3+}	hydroxide	OH^-
copper (II)	Cu^{2+}	carbonate	CO_3^{2-}
iron (II)	Fe^{2+}	sulfate	SO_4^{2-}
iron (III)	Fe^{3+}	hydrogen carbonate	HCO_3^-
zinc	Zn^{2+}	sulfite	SO_3^{2-}
silver	Ag^+	sulfide	S^{2-}
ammonium	NH_4^+	nitrate	NO_3^-
hydrogen	H^+	nitride	N^{3-}
		hydride	H^-

Covalent compounds – formed when bonding occurs between non-metals

Element	*Chemical formula*
fluorine	F_2
chlorine	Cl_2
bromine	Br_2
iodine	I_2
hydrogen	H_2
oxygen	O_2
nitrogen	N_2

Compound	*Chemical formula*
carbon dioxide	CO_2
carbon monoxide	CO
sulfur dioxide	SO_2
sulfur trioxide	SO_3
hydrogen chloride	HCl (g)
ammonia	NH_3
methane	CH_4

TIPS/ Common mistakes

In general, non-metals share electrons to attain stable noble gas electronic configuration.

E.g. 1 hydrogen atom shares 1 electron.
1 carbon atom (Group IV) shares 4 electrons.
1 nitrogen atom (Group V) shares 3 electrons.
1 oxygen atom (Group VI) shares 2 electrons.
1 chlorine atom (Group VII) shares 1 electron.

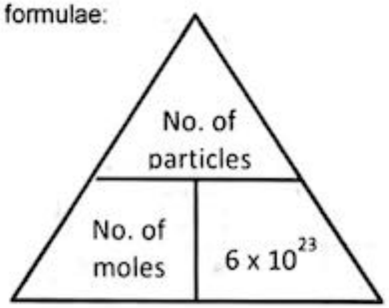
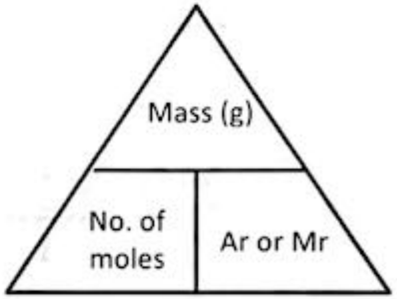
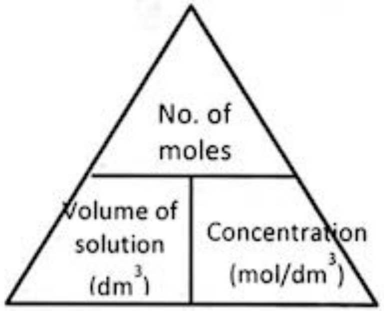
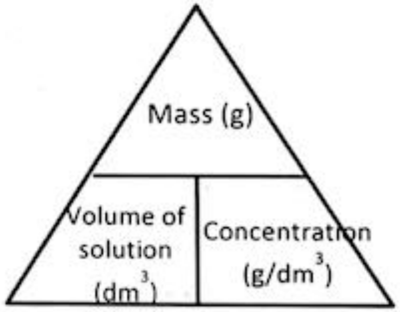
(b) deduce the formulae of simple compounds from the relative numbers of atoms present and vice versa	Eg. Sulfuric acid (formula: H_2SO_4) is made up of 2 atoms of H, 1 atom of S and 4 atoms of O. Magnesium hydroxide (formula: $Mg(OH)_2$) is made up of 1 atom of Mg, 2 atoms of O and 2 atoms of H.	
(c) deduce the formulae of ionic compounds from the charges on the ions present and vice versa	Eg. Formula of aluminium sulfate from Al^{3+} and SO_4^{2-} ions is $Al_2(SO_4)_3$.	
(d) interpret chemical equations with state symbols	E.g. $NaOH(aq) + NH_4Cl(aq) \rightarrow NaCl(aq) + NH_3(g) + H_2O(l)$ (s) - solid, (l) - liquid, (g) -gas and (aq) – aqueous The state symbols must be written on the same line as the chemical formulae.	
(e) construct chemical equations, with state symbols, including ionic equations	How to write ionic equations: Eg: hydrochloric acid reacts with sodium hydroxide to produce sodium chloride and water. 1) Write out the balanced chemical equation with state symbols. (Know your solubility table well!) 2) For the reactants and products in aqueous states, break them down into their respective ions. 3) Cancel out the spectator ions (ions that remain the same before and after the reaction). 4) Rewrite the equation. (note: **all ionic equations must have state symbols**). 1) Chemical Equation with state symbols: HCl (aq) + $NaOH$ (aq) → $NaCl$ (aq) + H_2O (l) 2) H^+ (aq) + Cl^- (aq) + Na^+ (aq) + OH^-(aq) → Na^+ (aq) + Cl^- (aq) +H_2O (l) 3) H^+ (aq) + ~~Cl^- (aq) + Na^+ (aq)~~ + OH^-(aq) → ~~Na^+ (aq) + Cl^- (aq)~~ +H_2O (l) 4) Ionic Equation (MUST have state symbols): H^+ (aq) + OH^-(aq) → H_2O (l) **Practice:** 1. aqueous silver nitrate reacts with aqueous sodium chloride Ans: Ag^+ (aq) + Cl^- (aq) → $AgCl$ (s) 2. aqueous lead (II) nitrate reacts with potassium chloride Ans: Pb^{2+} (aq) + $2Cl^-$ (aq) → $PbCl_2$ (s)	Tip: For neutralisation reactions between acids and bases, the ionic equation is usually H^+ (aq) + OH^-(aq) → H_2O (l)
(f) define relative atomic mass, Ar	The relative atomic mass (A_r) of an atom is the average mass of one atom of that element compared to $\frac{1}{12}$ of the mass of one carbon-12 atom. E.g. A_r of F = 19	A_r has no unit as it is a relative value. Tip: It's the nucleon number from the Periodic Table.

<table>
<tr>
<td>(g) define relative molecular mass, Mr, and calculate relative molecular mass (and relative formula mass) as the sum of relative atomic masses</td>
<td>Relative molecular mass, M_r, of a molecular substance is the average mass of one molecule of that element OR compound compared to $\frac{1}{12}$ of the mass of one carbon-12 atom.
E.g. M_r of fluorine gas, $F_2 = 19 \times 2 = 38$
Eg. M_r of calcium carbonate, $CaCO_3 = A_r \text{ of Ca} + A_r \text{ of C} + (3 \times A_r \text{ of O})$
$= 40 + 12 + (3 \times 16)$
$=100$
E.g. M_r of Aluminium sulfate, $Al_2(SO_4)_3 = (2 \times A_r \text{ of Al}) + (3 \times A_r \text{ of S}) + (12 \times A_r \text{ of O})$
$= (2 \times 27) + (3 \times 32) + (12 \times 16)$
$= 342$</td>
<td>M_r has no unit as it is a relative value.</td>
</tr>
<tr>
<td>(h) calculate the percentage mass of an element in a compound when given appropriate information</td>
<td>The mass of an element in a compound
$= \frac{\underline{\text{No. of atoms of the element x Ar}}}{\text{Mr of the compound}}$ x mass of the compound
Eg. % of O in 5g $CaCO_3 = \frac{3 \times A_r \text{ of O}}{M_r \text{ of } CaCO_3} \times 5\text{g} = \frac{16 \times 3}{100} \times 5\text{ g}$
$= \underline{2.40\text{ g}}$
The % by mass of an element in a compound
$= \frac{\underline{\text{No. of atoms of the element x Ar}}}{\text{Mr of the compound}}$ x 100 %

Eg. % of O in $CaCO_3 = \frac{3 \times A_r \text{ of O}}{M_r \text{ of } CaCO_3} \times 100\% = \frac{16 \times 3}{100} \times 100\%$
$= \underline{48.0\%}$</td>
<td>All final answers to 3 significant figures</td>
</tr>
<tr>
<td>(i) calculate empirical and molecular formulae from relevant data

Multiply by lowest possible number throughout to convert the simplest mole ratio to whole number</td>
<td><u>Empirical formula</u>
Vitamin C (ascorbic acid) contains 40.92 % C, 4.58 % H, and 54.50 % O, by mass. The experimentally determined relative molecular mass is 176. What is the empirical and molecular formula for ascorbic acid?
<table>
<tr><th></th><th>C</th><th>H</th><th>O</th></tr>
<tr><td>Mass in 100g (%)</td><td>40.92</td><td>4.58</td><td>54.50</td></tr>
<tr><td>A_r</td><td>12</td><td>1</td><td>16</td></tr>
<tr><td>No. of moles</td><td>40.92 / 12 = 3.41</td><td>4.58 / 1 = 4.58</td><td>54.50 / 16 = 3.406</td></tr>
<tr><td>Simplest mole ratio</td><td>3.41 / 3.406 = 1</td><td>4.58 / 3.406 = 1.344</td><td>3.406/3.406 = 1</td></tr>
<tr><td></td><td>1 x 3 = 3</td><td>1.344 x 3 = 4.02 ≈ 4</td><td>1 x 3 = 3</td></tr>
</table>
Hence, the empirical formula is $C_3H_4O_3$. [Be sure to write out the empirical formula]

Let the molecular formula be $(C_3H_4O_3)_n$.
$n = 176 / (3 \times 12 + 4 \times 1 + 3 \times 16) = 2$
Hence the molecular formula is $(C_3H_4O_3)_2 = C_6H_8O_6$.</td>
<td><u>Common mistakes:</u>
- Do not round off the simplest ratio unnecessarily, e.g., 4.12 cannot be rounded off to 4 but 4.02 can be rounded off to 4.
Hence, it should be
1.5 : 1 : 2 ⇒ 3 : 2 : 4
[multiply by 2]
1.33 : 1: 2 ⇒ 4 : 3 : 6
[multiply by 3]
1.25 : 1 : 2 ⇒ 5 : 4 : 8
[multiply by 4]
1.2 : 1 : 2 ⇒ 6 : 5 : 10
[multiply by 5]</td>
</tr>
</table>

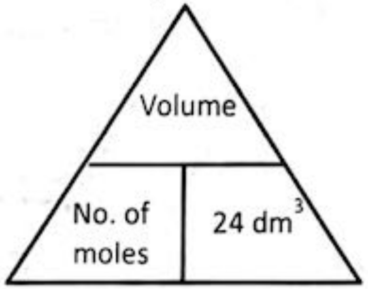
(j) calculate stoichiometric reacting masses and volumes of gases (one mole of gas occupies 24 dm³ at room temperature and pressure); calculations involving the idea of limiting reactants may be set	When given a problem to solve: 1) Write out the balanced chemical equation (you need this to get the mole ratio). 2) If data for both reactants are given, find the limiting reactant first. Useful formulae:   Note: Particles refer to atoms / molecules / ions / electrons etc. **To find concentration of solutions:**   **Concentration (g/dm^3) = Concentration (mol/dm^3) x M_r**	

For gases only (at r.t.p.):

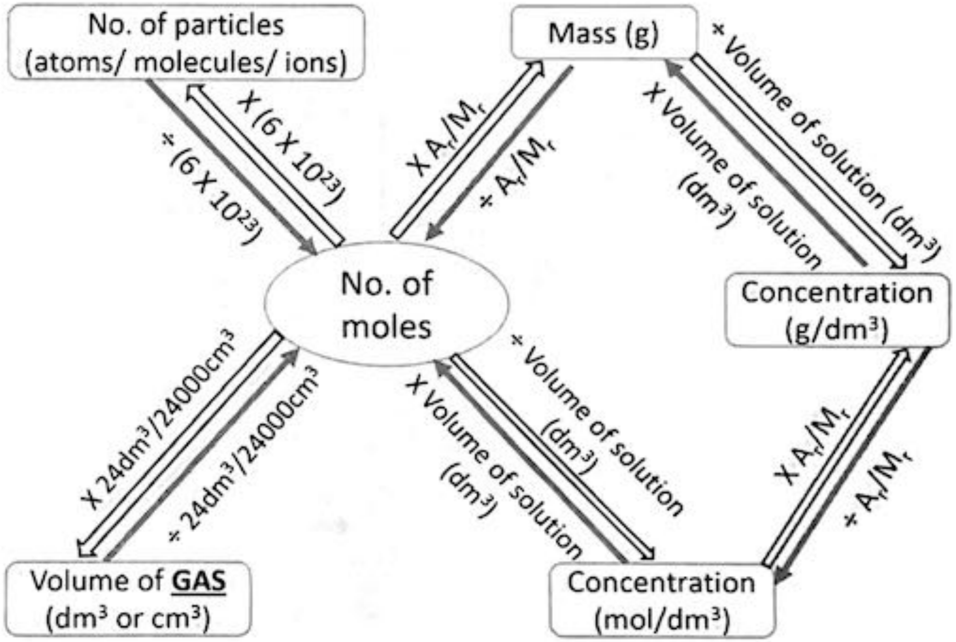
One mole of any gas occupies **24 dm³** at r.t.p. (1 dm³ =1000 cm³)
E.g. 24 cm³ of hydrogen reacts with 12 cm³ of chlorine gas. What volume of gases was left after the reaction?

Ans: $H_2 (g) + Cl_2 (g) \rightarrow 2HCl (g)$
Mole 1 : 1 : 2
Volume/cm³ 12 : 12 : 24

H_2 was present in excess and Cl_2 was the limiting reactant. Hence 12 cm³ of H_2 and 24 cm³ of HCl was left after the reaction. Total: 36 cm³



Summary



(k) apply the concept of solution concentration (in mol/dm³ or g/dm³) to process the results of volumetric experiments and to solve simple problems(Appropriate guidance will be provided where unfamiliar reactions are involved.)	$\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$ where M = concentration in mol/dm³ ; V = volume ; n = mole ratio in balanced equation **Eg:** Sodium ethanedioate, $Na_2C_2O_4$, can be made into a solution of an exact and reliable concentration and so can be used as a standard solution for checking the concentration of a solution containing sulfuric acid. The reaction is summarized by the equation: $Na_2C_2O_4 + H_2SO_4 \rightarrow Na_2SO_4 + (CO_2H)_2$ In a particular experiment 10.0 cm³ of a 0.50 mol/dm³ solution of sodium ethanedioate was found to react with 5.0 cm³ of sulfuric acid. Find the concentration in mol/dm³ of the sulfuric acid. **Ans**: Method 1: $\frac{M_AV_A}{n_A} = \frac{M_BV_B}{n_B}$ $n_B = 1$; $M_B = 0.50$ mol/dm³ ; $V_B = 0.010$ dm³ $n_A = 1$; $M_A = ?$; $V_A = 0.0050$ dm³ Hence $M_A = 1.00$ mol/dm³ Method 2: No. of moles of $Na_2C_2O_4$ = 0.50 mol/dm³ x 0.010 dm³ = 0.005 Since 1 mol of $Na_2C_2O_4$ reacts with 1 mol of H_2SO_4, no. of moles of H_2SO_4 = 0.005 Concentration of H_2SO_4 = 0.005 / 0.005 dm³ = 1.00 mol/dm³	Final answer to 3 significant figures.
(l) calculate % yield and % purity	**Percentage yield** **Eg:** When 3.7 g of ethanol was reacted with 6.0 g of ethanoic acid and the resultant ethyl ethanoate purified by distillation, the mass of ethyl ethanoate obtained was 5.8 g. What is the percentage yield?	DO NOT round off to 3 significant figures for intermediate steps. Only round off for the FINAL answer.

Ans:
Step 1: Write a balanced equation.

$CH_3CH_2OH + CH_3COOH \rightarrow CH_3COOCH_2CH_3 + H_2O$

Step 2: Find limiting reagent.

No. of moles of ethanol: 3.7 / (2 x 12 + 6 + 16) = 0.08043 mol
No. of moles of ethanoic acid = 6.0 / (2 x 12 + 2 x 16 + 4) = 0.1 mol
Since 1 mol of ethanol reacts with 1 mol of ethanoic acid, the limiting reactant is ethanol. Hence, 0.08043 mol of ethanoic acid reacts with 0.08043 mol of ethanol to produce 0.08043 mol of ethyl ethanoate.

Theoretical yield of ethyl ethanoate = 0.08043 x (4 x 12 + 2 x 16 + 8) = 7.0778g

% yield = experimental yield / theoretical yield x 100 %
= 5.8 / 7.0778 x 100%
= 81.9% (3 sf)

Usually, the experimental yield is given in the question. You have to calculate theoretical yield. Use the mass of the reactant(s) given and *work forwards*, applying mole ratio. (If you get a percentage yield of > 100%, it is wrong.)

Percentage purity

Usually, you are required to find the % purity of an impure reactant. The mass of the impure reactant is usually given. The impure substance will undergo a reaction to produce a pure product. Typically, the mass of the pure product is also given. ***Work backwards.***

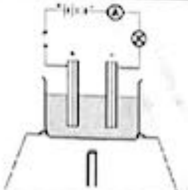
Solution: Use the mass of the pure product to calculate the no. of moles of the pure product. Apply mole ratio to find the no. of moles of the reactant that is pure. Convert it to mass and apply

% purity = mass of pure substance / mass of impure substance x 100%

An impure sample of NaCl of mass 0.50g, gave on treatment with an excess of $AgNO_3$ solution, 0.90 g of AgCl as precipitate. Calculate the percentage purity of the sample.

$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
No of mol of pure NaCl = No of mol of AgCl produced = 0.90 / (108 + 35.5) = 0.006272 mol
Mass of pure NaCl = 0.006272 x (23 + 35.5) = 0.3669 g
Percentage purity = 0.3669 / 0.5 x 100% = 73.4 % (3 sf)

4 Electrolysis

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(a) describe electrolysis as the conduction of electricity by an ionic compound (an electrolyte), when molten or dissolved in water, leading to the decomposition of the electrolyte	Electrolysis is the process of using **electricity** to **decompose** a **compound**. Electrolyte is a substance that conducts electricity in the **molten** or **aqueous** state and **is not limited to ionic compounds**; eg. **HCl (aq). Mobile ions** present in the electrolyte allow it to conduct electricity.	Common mistake: To mistake electrons as the particles responsible for electrical conductivity in electrolyte.
(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	 If substance in the beaker is **solid sodium chloride**, lamp **does not light up** because the **ions are held in fixed positions in the lattice structure**. Lamp **lights up** when the sodium chloride is **molten or in aqueous solution** because **the ions are now mobile and are able to move around.**	
(c) describe, in terms of the mobility of ions present and the electrode products, the electrolysis of molten sodium chloride, using inert electrodes	**Anode** is **positive electrode** connected to **positive terminal of the battery.** Chloride ions (anions) move towards/are attracted to the anode and are **discharged/oxidised/lose electrons** to form chlorine gas. $2Cl^-(l) \rightarrow Cl_2(g) + 2e^-$ **Cathode** is **negative electrode** connected to **negative terminal of the battery.** Sodium ions (cations) move towards/are attracted to the cathode and are **discharged/reduced/gain electrons** to form **molten** sodium. $Na^+(l) + e^- \rightarrow Na(l)$ Overall reaction is the decomposition of sodium chloride by electricity to form molten sodium metal & chlorine gas. $\mathbf{2NaCl(l) \rightarrow 2Na(l) + Cl_2(g)}$	
(d) predict the likely products of the electrolysis of a molten binary compound	Electrolysis of molten ionic compound produces **metals at the cathode** & **non-metals at the anode**. Possible observations: reddish-brown gas (Br_2), violet vapour (I_2), **silvery** globules (Na, K, Mg), **grey** globules(Pb), **pink** globules(Cu). Use globules for molten metal produced.	**Common mistakes:** For the electrolysis of **molten** ionic compounds, 1. Writing (aq) instead of **(l)** for state symbol for the ions. 2. Writing (s) instead of **(l) for molten metal** formed at cathode. $Cu^{2+}(l) + 2e^- \rightarrow Cu(l)$ [correct] $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ [wrong] 3. Writing name of product when question asks for **observations**; eg. Oxygen gas instead of **effervescence**.

<table>
<tr>
<td>(e) apply the idea of selective discharge based on
(i) cations: linked to the reactivity series (see 9.2)</td>
<td>**H^+ ions will be preferentially discharged/reduced producing hydrogen gas** because **hydrogen** is **less reactive than the other metal in the Reactivity Series**.

Cu^{2+}/Ag^+ ions are preferentially discharged/reduced to form solid copper /silver metal as silver and copper are **less reactive than hydrogen in the Reactivity Series.**

Electrolysis of **aqueous** solution using **inert** electrode will produce **hydrogen gas or metal** at the **cathode**.</td>
<td>Reactive metals such as Na & K are never produced during electrolysis of the **aqueous solution**</td>
</tr>
<tr>
<td>(ii) anions: halides, hydroxides and sulfates (e.g. aqueous copper(II) sulfate and dilute sodium chloride solution (as essentially the electrolysis of water))</td>
<td>Ease of discharge: **OH^-** > I^- > Br^- > Cl^- > SO_4^{2-} > NO_3^-

SO_4^{2-} & NO_3^- are NOT discharged.</td>
<td>**Common mistake:**
$4OH^-(l) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$
The state symbol for hydroxide ions should be **(aq)** instead of (l).</td>
</tr>
<tr>
<td>(iii) concentration effects (as in the electrolysis of concentrated and dilute aqueous sodium chloride)</td>
<td colspan="2">At the anode, **OH^- ions will be preferentially discharged/oxidised to form oxygen gas and water** except for **concentrated halide solutions where the halide ions will be preferentially discharged/oxidised to form halogen gas/liquid.**

<table>
<tr><td>Dilute NaCl – oxygen gas evolved at the anode</td><td>Concentrated NaCl – chlorine gas evolved at the anode</td></tr>
<tr><td>Anode: $4OH^-(aq) \rightarrow 2H_2O(l) + O_2(g) + 4e^-$</td><td>Anode: $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$</td></tr>
</table>
</td>
</tr>
<tr>
<td>(f) predict the likely products of the electrolysis of an aqueous electrolyte, given relevant Information</td>
<td colspan="2">When given **reactive electrode like copper anode**, copper oxidises, $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ instead of OH^- ions being discharged.</td>
</tr>
<tr>
<td>(g) construct ionic equations for the reactions occurring at the electrodes during the electrolysis, given relevant information</td>
<td colspan="2">**Common mistakes:**

1. Writing the electrons on the wrong side of the equation.
2. Discharge of product instead of ions.
3. Equation not balanced.
4. Wrong state symbols.</td>
</tr>
</table>

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

Electrolysis of aqueous copper(II) sulfate

Type of electrodes	**Component**	**Observation**	**Explanation**
Carbon or platinum	Anode	**Effervescence** produced which rekindles a glowing splint.	**Oxygen** is produced.
	Cathode	**Pink/ reddish-brown solid** deposited at the cathode.	**Copper metal** is deposited.
	Electrolyte	**Blue colour of electrolyte becomes lighter / turns colourless**.	Cu^{2+} ions are discharged from the solution, **concentration of Cu^{2+} ions decreases.**
Copper	Anode	The anode **dissolves / becomes smaller**.	**Copper anode is oxidised** to form Cu^{2+}(aq).
	Cathode	**Pink/ reddish-brown solid** deposited at the cathode.	**Copper metal** is deposited.
	Electrolyte	**Electrolyte remains blue in colour.** As **copper oxidises at the anode, and Cu^{2+} ions are discharged at the cathode**, the **concentration of Cu^{2+} ions remains constant.**	For every 1 mol of Cu anode that is oxidised to form 1 mol of Cu^{2+}, 1 mol of Cu^{2+} is reduced to form 1 mol of Cu. Hence, **Concentration of Cu^{2+} ions in the solution remains unchanged.**

To purify impure copper, **anode** is made of **impure copper**, **cathode** is **pure copper** and **copper (II) sulfate solution** is used as **electrolyte**.

(i) describe the electroplating of metals, e.g. copper plating, and state one use of electroplating

The **object** to be plated must be placed at the **cathode**, the **anode** must be the **pure plating metal**, with the electrolyte containing **ions of the plating metal**. eg. Plating a spoon with silver

Spoon is made the cathode by connecting to negative terminal of battery.
Anode is made of silver metal.

Electrolyte is **aqueous silver nitrate**. (solution used must contain the ion of the coating metal)
Objects are electroplated to **protect them from corrosion** (tin electroplated onto iron are used in food cans to prevent rusting) and to give them an **attractive appearance** (chromium electroplated onto steel used to make water taps /bicycle parts).

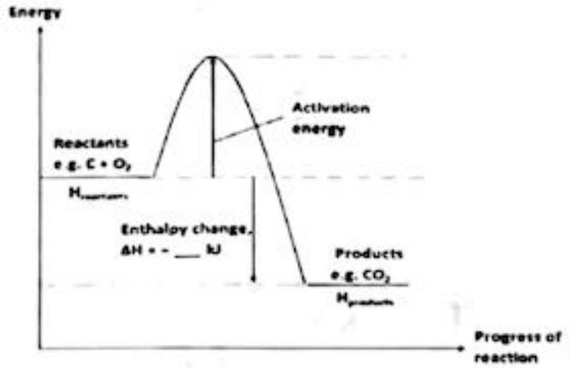
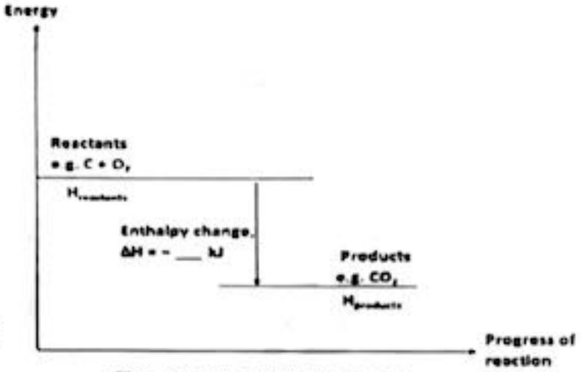
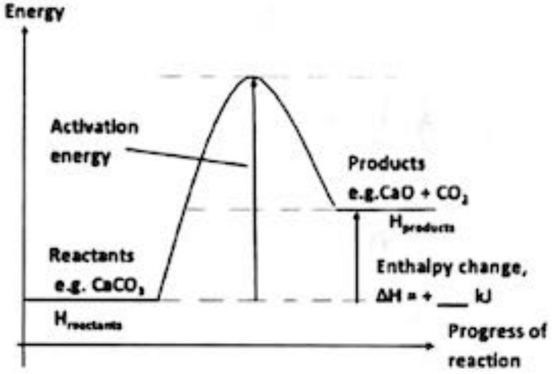
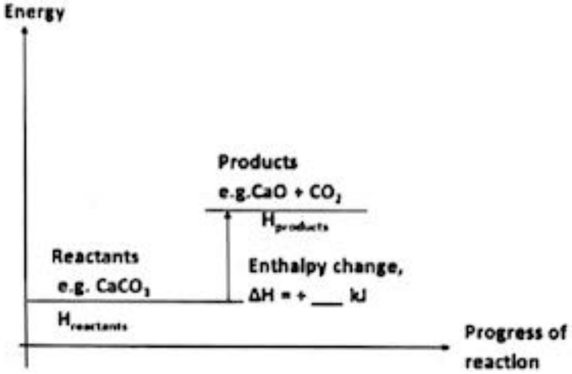
<table>
<tr>
<td>(j) describe the production of electrical energy from simple cells (i.e. two electrodes in an electrolyte) linked to the reactivity series (see 9.2) and redox reactions (in terms of electron transfer)</td>
<td>The more reactive metal is the negative electrode where oxidation occurs.

E.g., Zn and Fe chemical cell,
Zn is the negative electrode ($Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$)
The electrons then flow to the positive electrode where reduction occurs.

The further apart the two metals are in the reactivity series, the greater the voltage produced.
<table>
<tr><th>Property</th><th>Electrolytic cell (with battery)</th><th>Simple cell (It acts as THE BATTERY!!!)</th></tr>
<tr><td>Energy change</td><td>Electrical energy → chemical energy</td><td>Chemical energy → electrical energy</td></tr>
<tr><td>Negative electrode</td><td>Cathode where reduction occurs.
Cations are discharged by gaining electrons.</td><td>Anode where oxidation occurs.
Metal electrode dissolves & forms metal ions by losing electrons.</td></tr>
<tr><td>Positive electrode</td><td>Anode where oxidation occurs.
Anions are discharged by losing electrons to form molecules.</td><td>Cathode where reduction occurs. Positive ions are discharged by gaining electrons to form neutral atoms.</td></tr>
<tr><td>Flow of electrons in external circuit</td><td>Electrons flow from positive electrode (anode) to negative electrode (cathode) via wire.</td><td>Electrons flow from negative electrode (more reactive metal) to positive electrode (less reactive metal) via wire.</td></tr>
</table>
</td>
</tr>
</table>

5 Energy from Chemicals

Learning Outcomes	**Key Concepts**	**TIPS/ Common mistakes**
(a) describe the meaning of enthalpy change in terms of exothermic (ΔH negative) and endothermic (ΔH positive) reactions	An **exothermic** reaction is where the **enthalpy change is negative** and the **products have a lower energy level than the reactants**. An **endothermic** reaction is where the **enthalpy change is positive** and the **products have a higher energy level than the reactants**.	

(b) represent energy changes by energy profile diagrams, including reaction enthalpy changes and activation energies (see 6.1(c),6.1(d)) 1m - Shape + Labelling of products 1m - Labelling of A.E. & ΔH	**Exothermic**  Energy profile diagram  Energy Level diagram **Endothermic**  Energy profile diagram  Energy Level diagram	Things to note: 1) Label **axes** y-axis as Energy x-axis as Progress of reaction 2) **Shape** of graph 3) Label reactants and products as per the equation given 4) Label **enthalpy change**, ΔH (arrow always from reactants to products), include value if given DOWNWARD arrow if exothermic, UPWARD arrow if endothermic. 5) Label **activation energy**, (arrow always UPWARD from reactants to peak of curve) 6) **Enthalpy change** (ΔH) must always include the **sign** +/-

(c) describe bond breaking as an endothermic process and bond making as an exothermic process	• **Bond breaking** in reactants takes in energy: **endothermic** • **Bond forming** in products gives out energy: **exothermic**	
(d) explain overall enthalpy changes in terms of the energy changes associated with the breaking and making of covalent bonds	During reaction, **$\Delta H = \Delta H_{products}$ (always positive) + $\Delta H_{reactants}$ (always negative)** When **net heat energy is taken in** by the reaction, it is termed as an **endothermic reaction**. When **net heat energy is released** by a chemical reaction, the reaction is termed as an **exothermic reaction**. **Endothermic reaction** • The **energy taken in** for **breaking of bonds in reactants (NAME) is greater than energy given out for making of bonds in products (NAME).** • **Temperature** of the surrounding **decreases.** **Exothermic reaction** • E.g. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$, $\Delta H = -484kJ$ • The **energy taken in/absorbed** for **breaking of bonds in H_2 and O_2 (write what reactants) is less than energy given out/released** for **forming of bonds in H_2O (write what products).** • **Temperature** of the surrounding **increases.** **More examples of exothermic and endothermic reactions** **Exothermic:** • Combustion of fuel e.g. wood, coal, H_2 (in rockets), methane • Respiration • Neutralisation • Corrosion of metals (e.g. rusting of iron) • Condensation/Freezing • Dissolving of acids in water (e.g. concentrated sulfuric acid) • Dissolving of anhydrous salts (e.g. anhydrous sodium carbonate) **Endothermic :** • Thermal decomposition $CaCO_3 \xrightarrow{heat} CaO + CO_2$ • Photochemical reaction $2AgBr \xrightarrow{light} 2Ag + Br_2$ • Photosynthesis • Evaporation/ boiling • Melting • Dissolving of some ionic compounds in water (e.g. NH_4Cl, KCl, $CuSO_4 \bullet 5H_2O$)	Note: **ALL chemical reactions involve BOND BREAKING AND BOND FORMING – processes which take in heat energy and release heat energy**. Endothermic and exothermic is decided upon the ***net* release or taking in of heat energy**. **Common mistake:** Use of 'used/required' instead of 'absorbed/released' to explain the energy change in reactions. **TIP:** Don't be too bothered by temperature "of surroundings". In general, temp increases for exo and decreases for endo.

(e) describe hydrogen, derived from water or hydrocarbons, as a potential fuel, reacting with oxygen to generate electricity directly in a fuel cell (details of the construction and operation of a fuel cell are not required)	Fuels are substances that burn easily in air to give out energy. Fossil fuels are commonly used fuels. However, they are **non-renewable resources**. There are limited amounts of fossil fuels and they will eventually run out. Hydrogen as a potential fuel: • Hydrogen is used as an alternative fuel. When hydrogen burns, the product is **steam only**. • Hydrogen can be produced from electrolysis of water & cracking of crude oil **Fuel cell:** • A chemical cell in which reactants (usually a fuel which is hydrogen and oxygen) are continuously supplied **to produce electricity directly** is called a fuel cell. • The best known example is the **hydrogen-oxygen fuel cell** which is used as a source of **electrical energy** for the spacecraft and water for the astranauts. • In the hydrogen-oxygen fuel cell, **hydrogen and oxygen is converted to water through redox reactions.** At the **anode**, each **H_2 molecule is oxidised** to form $2H^+$ ions and 2 electrons. The electrons flow through the wire to generate electricity. At the cathode, the **H^+ ions are reduced** by reacting with the oxide ions to form water molecules. **$2H_2 (g) + O_2 (g) \rightarrow 2H_2O (l)$ (overall equation)**	**Common mistake:** It is wrong to state that hydrogen is a component of air and can be obtained from fractional distillation of liquid air.
6 Chemical Reactions		
6.1 Speed of reaction		
Learning Outcomes	**Key Concepts**	
(a) describe the effect of concentration, pressure, particle size and temperature on the speeds of reactions and explain these effects in terms of collisions between reacting particles	Reactions occur based on the **Collision Theory.** Reactant particles will successfully react to form products when they collide with sufficient energy (more than activation energy). This is called an ***effective collision***. Otherwise, they will just bounce off one another without any reaction, which are called *ineffective collisions.*	**TIPS/ Common mistakes** In coal mines, the coal dusts are very fine (large surface area) and very susceptible to spark explosive reactions.

FACTORS AFFECTING RATE OF A CHEMICAL REACTION

Factor	**Concentration (solutions)**	**Pressure (gaseous)**	**Surface Area/ Particle Size (Solid)**	**Temperature (all)**	**Catalyst (all)**
Effect	↑ concentration, ↑ rate	↑ pressure, ↑ rate	↓size , ↑ S.A, ↑ rate	↑ temperature, ↑ rate	Catalyst present, ↑ rate
Reason (Standard answers)	- concentration of reactant increases, - the **number of reactant particles per unit volume** increases, - increasing frequency of collisions, - increasing frequency of effective collisions.	- pressure increases / volume decreases. - The **number of reactant particles per unit volume** increases, - increasing frequency of collisions, - increasing frequency of effective collisions.	- smaller particles - **surface area for reaction** increases, -increasing frequency of collisions, -increasing frequency of effective collisions.	- At higher temperatures, the particles gain **kinetic energy** and move faster - **More particles now have energy equal to or greater than activation energy.** - increasing frequency of collisions, - increasing frequency of effective collisions.	A catalyst **provides an alternative reaction pathway** of **lowered activation energy** of the reaction. - **More particles now have energy equal to or greater than activation energy.** - increasing frequency of collisions, - increasing frequency of effective collisions.

(b) define the term catalyst and describe the effect of catalysts (including	**Catalyst** is a substance which increases the **speed of a chemical reaction** but remains ***chemically unchanged*** at the end of the reaction. They **increase the speed of reaction** by providing an **alternative pathway of lowered activation energy**.	**Note:** A catalyst increases the speed of a chemical reaction **BUT does not**

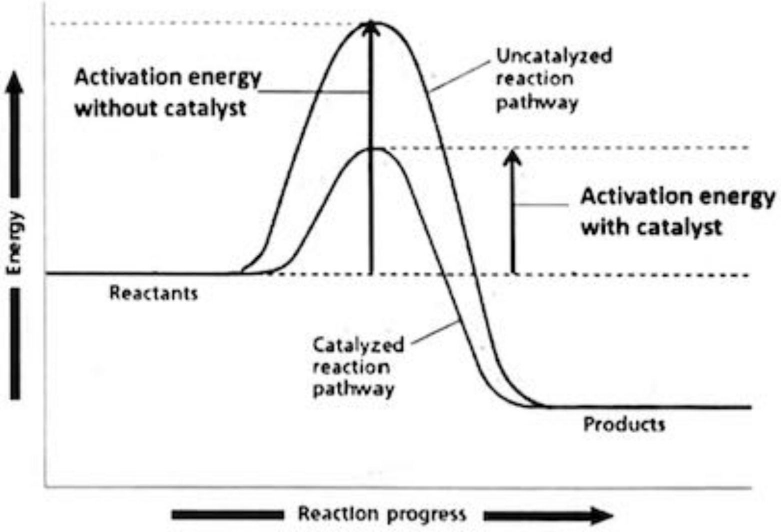
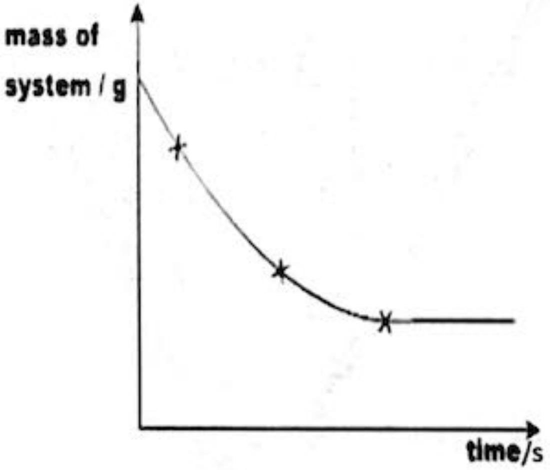
<table>
<tr>
<td><i>enzymes) on the speeds of reactions &
(c) explain how pathways with lower activation energies account for the increase in speeds of reactions</i></td>
<td>


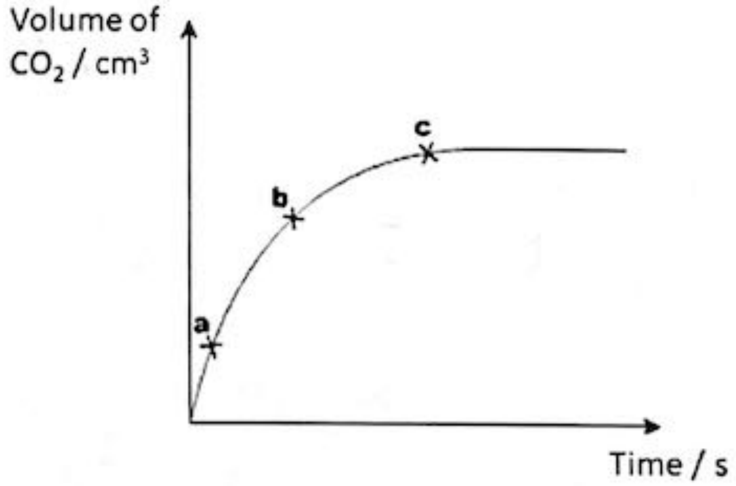
Figure 1: A catalysed reaction has a <u>lower activation energy</u> than an uncatalysed reaction. Thus the former <u>proceeds faster</u>. (Refer to table above for explanation)
</td>
<td>increase the yield of the product. The same quantity of product is obtained with or without the catalyst.</td>
</tr>
<tr>
<td><i>(d) state that some compounds act as catalysts in a range of industrial processes and that enzymes are biological catalysts (see 5(b), 6.1(c) and 10(d))</i></td>
<td colspan="2">Enzymes are <u>biological catalysts</u>. They have the following properties:

specific in their actions (like lock and key)
are usually proteins
sensitive to temperature changes (can be denatured at high temperatures, but inactive at low temperatures)
sensitive to pH changes

<u>Examples of catalysts in industrial process:</u>
<table>
<tr><th>Industrial Process</th><th>Catalyst</th></tr>
<tr><td>Haber Process</td><td>Iron (finely divided)</td></tr>
<tr><td>Addition of hydrogen to alkenes (making of margarine)</td><td>Nickel</td></tr>
<tr><td>Addition of steam to alkenes</td><td>Phosphoric(V) acid (H_3PO_4)</td></tr>
<tr><td>Fermentation of glucose to manufacture ethanol</td><td>Yeast (enzyme)</td></tr>
</table>
</td>
</tr>
<tr>
<td><i>(e) suggest a suitable method for investigating the effect of a given</i></td>
<td>
Reactions occur when reactant particles combine and react to give products.
Rate of reaction is actually the speed of the reaction. It is measured as the average <u>amount of change in a reaction per unit time.</u>
</td>
<td>These changes are measurable or observable (i.e. can easily be used to</td>
</tr>
</table>

variable on the speed of a reaction & *(f) interpret data obtained from experiments concerned with speed of reaction*	• Examples of the 'change' : ▪ Disappearance of reactants/ appearance of products (e.g. precipitate) ▪ **Mass (heavy gas given off, reaction mixture gets lighter)** ▪ **Volume of gas produced** ▪ Temperature (temperature rises or falls) ▪ Colour (reaction is accompanied by change in colour) ▪ pH Graphs Consider the reaction: $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$ There are two changes (mass of system and volume of gas produced), which can be tracked to find the rate of reaction. They are plotted as Graphs A and B below. • Change in mass:  **Graph A:** Plot of mass of system against time for reaction of $CaCO_3$ with HCl.	calculate the rate of reaction). **Recall:** In studying the shape of a graph, you need to note the **gradient** and the **height**. The **gradient** refers to the **speed of reaction** and is **affected by the factors** such as temperature, concentration, surface area and pressure while the **height** refers to the **yield of products** and is **dependent on the mole of the limiting reactant.**

- Volume of CO_2 formed:

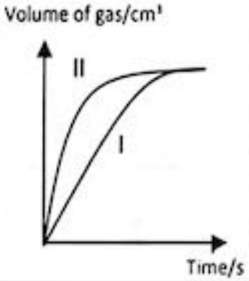
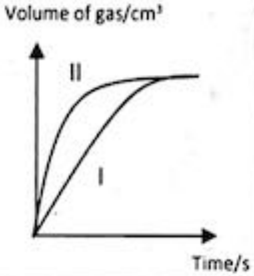
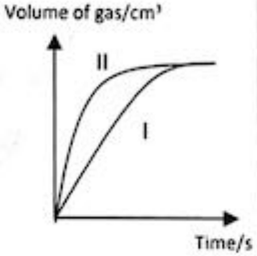


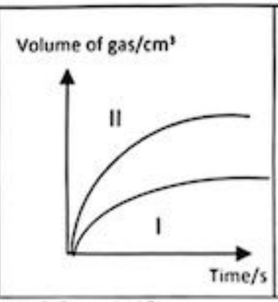
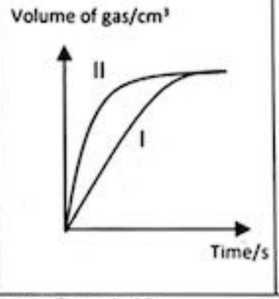
Graph B: Plot of volume of gas produced against time for reaction of $CaCO_3$ with HCl.

The experiment is repeated to show the effects of surface area, temperature, concentration and catalyst on the speed of the reaction.

In comparing the speeds of two reactions, only one factor is varied and the other factors are kept constant.

Examples:

Set	**Condition that is varied**	**Conditions that are fixed**	**Shape of graph**	**Observation**
A	**<u>Size of zinc particles</u>** *Experiment I:* Granulated zinc *Experiment II:* Powdered zinc	• Mass of zinc • Volume of dilute HCl • Concentration of dilute HCl • Temperature • No catalyst		- Graph II has a steeper gradient => Reaction II is a faster reaction. - The smaller the particle size, the faster the reaction.
B	**<u>Temperature of aqueous HCl</u>** *Experiment I:* HCl (aq) at 25°C *Experiment II:* HCl (aq) at 35°C	• Mass of zinc powder • Volume of dilute HCl • Concentration of dilute HCl No catalyst		- Graph II has a steeper gradient => Reaction II is a faster reaction. - The higher the temperature, the faster the reaction.
C	Presence of catalyst, aqueous copper (II) sulfate *Experiment I:* Without catalyst *Experiment II:* With catalyst	• Mass of zinc powder • Volume of dilute HCl • Concentration of dilute HCl • Temperature		- Graph II has a steeper gradient => Reaction II is a faster reaction. - In the presence of a catalyst, the reaction is faster. - **<u>A catalyst has no effect on the yield of the product.</u>**

		D	<u>**Concentration of dilute HC*l***</u> *Experiment I:* 1 mol/dm^3 HC*l* *Experiment II:* 2 mol/dm^3 HC*l*	• Mass of zinc powder (excess) • Volume of dilute hydrochloric acid • Temperature • No catalyst		- Graph II has a steeper gradient => Reaction II is a faster reaction. - The higher the concentration, the faster the reaction.		
			Why is the height of Graph II twice that of Graph I? **Ans:** Moles of HC*l* in experiment II is doubled. Since HC*l* is the limiting reagent, volume of H_2 produced will be doubled.					
			<u>**Concentration of dilute HC*l***</u> *Experiment I:* 1 mol/dm^3 HC*l* *Experiment II:* 2 mol/dm^3 HC*l*	• Mass of zinc powder (limiting) • Volume of dilute hydrochloric acid • Temperature • No catalyst		- Graph II has a steeper gradient => Reaction II is a faster reaction. - The higher the concentration, the faster the reaction.		
			Why is the height of Graph II the same as Graph I? **Ans:** Zinc powder is the limiting reagent. Since the number of moles of zinc for both reactions are the same, the number of moles of gas produced will be the same. Therefore, volume of gas produced will be the same.					

6.2 Redox

Learning Outcomes	**Key Concepts**	**TIPS/ Common mistakes**
(a) define oxidation and reduction (redox) in terms of oxygen/hydrogen gain/loss	**Oxidation and Reduction as loss or gain of Hydrogen and Oxygen.** • **Oxidation is the:** (a) Gain of Oxygen e.g. $C + O_2 \rightarrow CO_2$ C is oxidized. (b) Loss of Hydrogen e.g. $H_2S + Cl_2 \rightarrow S + 2HCl$ H_2S is oxidized. • **Reduction is the:** (a) Loss of Oxygen e.g. $CuO + H_2 \rightarrow Cu + H_2O$ CuO is reduced. (b) Gain of Hydrogen e.g. $H_2 + Cl_2 \rightarrow 2HCl$ Cl_2 is reduced.	
(b) define redox in terms of electron transfer and changes in oxidation state	**Oxidation and Reduction as change in Oxidation States and Electron Transfer** • **Oxidation States (O.S.) / Oxidation Number (O.N.)** (a) O.S. of simple ion = its charge e.g. ZnS : Zn (+2) and S (-2) NaCl : Na (+1) and Cl (-1) (b) O.S. of elements in uncombined state = 0 e.g. Cl_2 : Cl (0) K : K (0) • **Oxidation and Reduction as change in O.S. and electron transfer** Fill in the oxidation states of the above oxidation and reduction reactions and notice that: - Oxidation is accompanied by an **increase in O.S. (or the loss of electrons)**. E.g. $C + O_2 \rightarrow CO_2$ C is oxidised. (O.S. of C increases from 0 to +4.)	**Note:** The **oxidation state** of Zn in ZnS is **+2** while the **charge** of Zn in ZnS is **2+**.

E.g. $Fe^{2+} \rightarrow Fe^{3+} + e^-$

Fe^{2+} is oxidised. (O.S. of Fe increases from +2 to +3.)

- Reduction is accompanied by a **decrease in O.S. (or the gain of electrons).**

E.g. $CuO + H_2 \rightarrow Cu + H_2O$

CuO is reduced.(O.S. of Cu decreases from +2 to 0.)

E.g. $Cl_2 + 2e^- \rightarrow 2Cl^-$

Cl_2 is reduced to Cl^-. (O.S. of Cl decreases from 0 to -1.)

What is oxidation number/oxidation state?

The oxidation number or state is the *charge an atom of an element would have if it existed as an ion in a compound* (even if it is actually covalently bonded)

Rules to determine oxidation number: * Elements, not ions

1. The oxidation state of free element is zero.
   E.g. Cu = 0; C = 0; Cl_2 = 0
2. The oxidation state of oxygen is always -2, unless in its peroxides (where it is -1).
   E.g. In CuO, oxidation state of O = -2
   but in H_2O_2 (hydrogen peroxide), oxidation state of O = -1
3. The oxidation state of hydrogen is always +1, unless in its hydrides.
   E.g. In HCl, oxidation state of H = +1
   but in NaH (sodium hydride), oxidation state of H = -1
4. The oxidation state of a simple ion is the same as the charge on the ion.
   E.g. In K^+, oxidation state of K = +1 ; In S^{2-}, oxidation state of S = -2
5. The oxidation state of all the atoms present in a formula of a compound adds up to zero.
   e.g. HCl , oxidation state of H = +1 and Cl = -1 → (+1) + (-1) =0
   $CaCO_3$, oxidation state of Ca = +2, C = +4 , O = -2
   ⇒ (+2) + (+4) + 3(-2) = 0
6. The total of the oxidation state of the atoms in a polyatomic ion is equal to the charge on the ion. e.g. CO_3^{2-} ; oxidation state of C = +4, O = -2
   ⇒ (+4) + 3(-2) = -2

<table>
<tr>
<td>(c) identify redox reactions in terms of oxygen/hydrogen gain/loss, electron gain/loss and changes in oxidation state</td>
<td>
<table>
<tr><th></th><th>Oxidation</th><th>Reduction</th></tr>
<tr><td>Oxygen</td><td>Gain of Oxygen</td><td>Loss of Oxygen</td></tr>
<tr><td>Hydrogen</td><td>Loss of Hydrogen</td><td>Gain of Hydrogen</td></tr>
<tr><td>Electrons</td><td>Loss of electrons</td><td>Gain of electrons</td></tr>
<tr><td>Oxidation State</td><td>Increase</td><td>Decrease</td></tr>
</table>

A reaction where **both oxidation and reduction occur simultaneously** is a **redox** reaction. *

E.g., **iron in Fe_2O_3** is **reduced** as the **oxidation state of iron decreases** from **+3 in Fe_2O_3** to 0 in **Fe**. **Carbon in CO_2** is **oxidised** as **oxidation state of carbon increases** from **+2 in CO** to **+4 in CO_2**.
</td>
<td>
TIP:

Common Mnemonics :

- OIL RIG (Oxidation is loss of electrons, Reduction is gain of electrons)
- LEO GER (Lose electrons Oxidation, Gain electrons Reduction)
</td>
</tr>
<tr>
<td>(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</td>
<td>
Oxidising agents (O.A.) and Reducing agents (R.A.)

Oxidising agents oxidise others while reducing agents reduce others.

- **Oxidising agent (O.A.)** oxidises others and itself gets reduced (**gain electrons**) in the process. (electron acceptors)
  They are also known as ***oxidants***.

  E.g. $2Ca + O_2 \rightarrow 2CaO$
  $4HCl + MnO_2 \rightarrow MnCl_2 + 2H_2O + Cl_2$

  O_2 is the oxidising agent. It oxidises Ca to CaO (0 →+2) while itself is being reduced to CaO (0 →-2).
  Other common O.A.: Cl_2, $KMnO_4$, acidified $K_2Cr_2O_7$

- **Reducing agent (R.A.)** reduces others and itself gets oxidised (**lose electrons**) in the process. (Electron donors)
  They are also known as ***reductants***.

  E.g. $Zn\ (s) + Cu^{2+}\ (aq) \rightarrow Zn^{2+}\ (aq) + Cu\ (s)$
  $C_2H_4 + H_2 \rightarrow C_2H_6$

  Zn is the reducing agent. It reduces Cu^{2+} to Cu, while it itself is being oxidised to Zn^{2+}.
  Other common R.A.: H_2, C, CO, H_2S, SO_2, KI, $Na_2S_2O_3$.
</td>
<td>
Note:

* Elements in the highest O.S. are usually O.A. e.g. MnO_4^-, Fe^{3+}.
* Elements in the lowest O.S. are usually R.A. e.g. KI, metals.
* Elements in the intermediate O.S. can act as both O.A. or R.A. e.g. MnO_2.
* Ionic equations must be accompanied by state symbols.
* transition elements can exhibit multiple oxidation states.
</td>
</tr>
</table>

- **Test for O.A.**

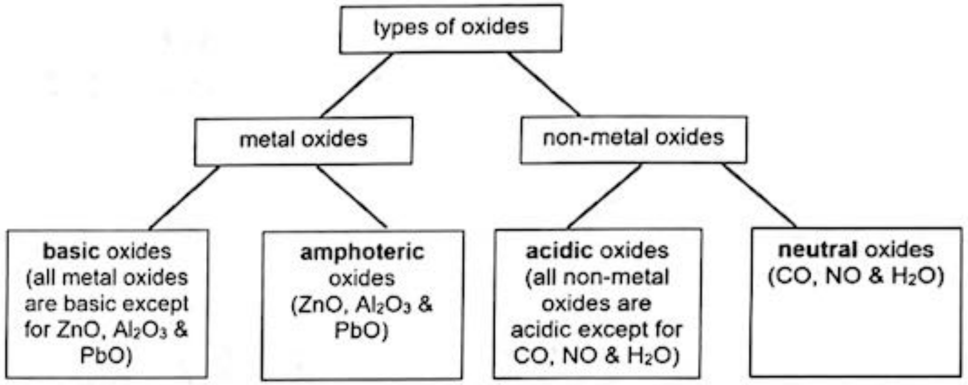
Test	Observation
Add KI	$2I^-$ (aq) → I_2 (aq) Colourless Brown

- **Tests for R.A.**

Test	Observation
a. Add acidified $KMnO_4$	MnO_4^- (aq) → Mn^{2+} (aq) Purple Colourless
b. Add acidified $K_2Cr_2O_7$	$Cr_2O_7^{2-}$ (aq) → Cr^{3+} (aq) Orange Green

7 Acids, Bases and Salts		
7.1 Acids and bases		
Learning Outcomes	**Key Concepts**	**TIPS/ Common mistakes**
(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	<u>**Acids dissociate in water**</u> to <u>**produce H^+ ions**</u>; <u>**Alkalis dissociate in water**</u> to produce <u>**OH^- ions**</u>; Colours of Universal Indicator in: -strong acids: red -weak acids: orange/yellow - neutral: green -weak alkalis: blue -strong alkalis: purple/violet	
(b) describe how to test hydrogen ion concentration and hence relative acidity using Universal Indicator and the pH scale	The <u>**pH is calculated based on number of hydrogen ions or hydroxide ions present per unit volume in a solution.**</u>	
(c) describe qualitatively the difference between strong and weak acids in terms of the extent of ionisation	<u>**Strong acids dissociate completely in water**</u> to give <u>**high concentration of H^+ ions**</u>. Common examples of strong acids: sulfuric acid, nitric acid, hydrochloric acid. E.g. HCl (aq) → H^+ (aq) + Cl^- (aq) H_2SO_4 (aq) → $2H^+$ (aq) + SO_4^{2-} (aq) It is important to note that H_2SO_4 is <u>**NOT**</u> a stronger acid than HCl (they are equal in strength). <u>**Weak acids dissociate partially in water**</u> to give <u>**low concentration of H^+ ions.**</u> Common example of weak acid: ethanoic acid.	

	E.g. $CH_3COOH\ (aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$ A strong acid can be dilute; a weak acid can be concentrated. Strength (degree of dissociation – cannot be changed) is not related to concentration of the acid (which can be changed)!	
(d) describe the characteristic properties of acids as in reactions with metals, bases and carbonates	1. **Acids + reactive metals (above hydrogen in metal reactivity series) → salt + hydrogen gas** **Observation:** Metal dissolves. Effervescence observed. Gas evolved extinguishes lighted splint with a 'pop' sound. 2. **Acids + bases (defined as any metal oxides or hydroxides) → salt + water (neutralisation reaction, exothermic in nature)** **Observation:** Beaker feels warm. Solid base dissolves. 3. **Acids + carbonates → salt + water + carbon dioxide gas** **Observation:** Metal carbonate dissolves. Effervescence observed. Gas evolved forms white precipitate in limewater.	NOTE: Candidates have to select suitable reagents for the reactions, taking into account the feasibility of reaction, availability of reagents and safety issues. **Eg. DO NOT react highly reactive metals such as sodium, potassium, calcium with hydrochloric acid to form the chlorides as the reaction is too violent.**
(e) state the uses of sulfuric acid in the manufacture of detergents and fertilisers; and as a battery acid	**Sulfuric acid** is **used in making detergents, in making fertilisers** and **used as the electrolyte in car batteries.**	
(f) describe the reaction between hydrogen ions and hydroxide ions to produce water, $H^+ + OH^- \rightarrow H_2O$, as neutralisation	**Ionic equation** for any **neutralization reaction involving acid and alkali** is as follows: **$H^+\ (aq) + OH^-\ (aq) \rightarrow H_2O\ (l)$**	Note: must include state symbols for ionic equations
(g) describe the importance of controlling the pH in soils and how excess acidity can be treated using calcium hydroxide	Plants grow well in soils of certain pH. In cases whereby the soil is too acidic, **calcium hydroxide (slaked lime) or calcium oxide (quicklime) can be added to neutralise the excess acidity in soil.** Calcium hydroxide **cannot** be added to the soil after it has been treated with fertilisers containing ammonium compounds. This is because calcium hydroxide will **react** with the ammonium compound and produce **ammonia gas**, resulting in **loss of nitrogen** which is an **essential element that aids plant growth.**	

(h) describe the characteristic properties of bases in reactions with acids and with ammonium salts	1. **Acids + bases (defined as any metal oxides or hydroxides) → salt + water (neutralisation reaction, exothermic in nature)** **Observation:** Beaker feels warm. Solid base dissolves. 2. **Alkalis + any ammonium salts (e.g NH_4Cl, NH_4Br) → Salt + water + ammonia gas** **Observation:** Effervescence ~~was~~ observed. Pungent gas evolved turns moist red litmus paper blue. E.g. $NaOH + NH_4Cl \rightarrow NaCl + NH_3 + H_2O$ 3. **Alkalis + metal salt → metal hydroxide + salt** (basis of test for cations) E.g. $2NaOH + CuCl_2 \rightarrow Cu(OH)_2 + 2NaCl$
(i) classify oxides as acidic, basic, amphoteric or neutral based on metallic/non-metallic character	 1. **Acidic oxides react with alkalis to form a salt and water** E.g. $SO_2 + 2\ NaOH \rightarrow Na_2SO_3 + H_2O$ Most acidic oxides dissolve in water to form an acid. Examples include **carbon dioxide, sulfur trioxide and phosphorus (V) oxide**. 2. **Basic oxides react with acids to form a salt and water** E.g. $CaO + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O$ Most basic oxides are insoluble in water. Few basic oxides such as **sodium oxide and potassium oxide** dissolve in water to form alkalis. E.g. $Na_2O + H_2O \rightarrow 2NaOH$ 3. **Amphoteric oxides are metallic oxides that react with both acids and bases to form salt and water.** Examples include **ZnO, Al_2O_3, PbO.** 4. **Neutral oxides (some non-metal oxides) show neither basic nor acidic properties. They are insoluble in water.** Examples are **water, carbon monoxide, CO and nitrogen monoxide, NO.**

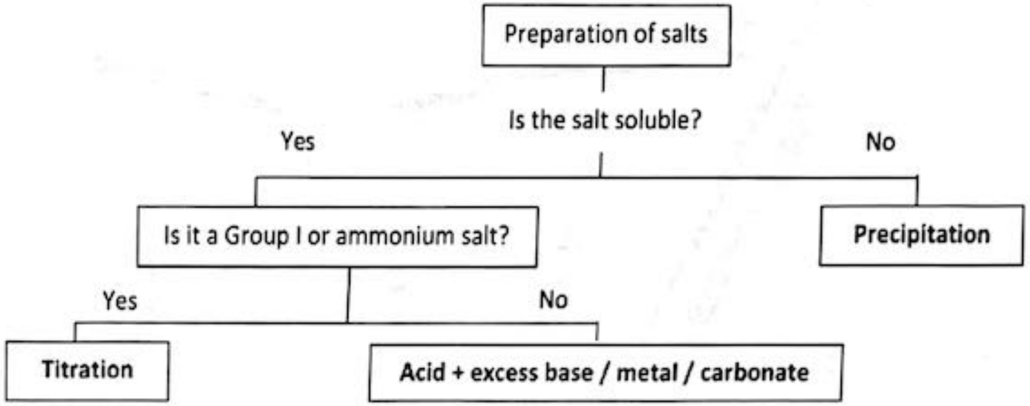
7.2 Salts

Learning Outcomes

(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)

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

Key Concepts



Acid + excess base / metal / carbonate (Note: To prepare all other soluble salts EXCEPT SPA and Group I salts)
First, decide whether a salt is soluble or not. If the salt is soluble in water, we react an acid with a suitable metal base or carbonate to get the salt.
Note: When using metals, this method cannot be used for reactive metals such as potassium, sodium and calcium as they react violently with acids. Furthermore, this method cannot be used for unreactive metals such as lead, copper and silver as they do not react with dilute acids. Hence, only moderately reactive metals such as zinc, magnesium and iron can be used.

E.g. Add **excess** zinc to dilute sulfuric acid and stir. **Filter the mixture to remove excess unreacted zinc.** **Collect the filtrate** and **heat it to evaporate** most of the water to obtain a **saturated solution**. Allow the hot solution to **cool** and **filter to collect the crystals**. **Wash the crystals with cold distilled water**. **Dry the crystals with filter paper.**

Titration method (Note: To prepare SPA and Group I salts)
E.g. **Fill up the burette** with aqueous hydrochloric acid. **Pipette 25.0cm^3** of aqueous sodium hydroxide into the conical flask. Add a **few drops of methyl orange**. Titrate the solution until the indicator turns from **yellow to orange permanently**. Record the volume of hydrochloric acid used and **repeat the titration using the same volume without adding the indicator.** **Heat** the salt solution **to evaporate** most of the water to obtain **a saturated solution**. Allow the hot solution to **cool** and **filter to collect the crystals**. **Wash the crystals with cold distilled water**. **Dry the crystals with filter paper.**

Precipitation (Note: To prepare insoluble salts using two aqueous solutions as reagents)
E.g. **Mix aqueous** sodium chloride with **aqeuous** lead (II) nitrate. **Filter** the mixture to obtain lead (II) chloride as the **residue**. **Wash the residue with distilled water** and **dry** with filter paper.

TIPS/ Common mistakes

<table>
<tr>
<td>(b) describe the general rules of solubility for common salts to include nitrates, chlorides (including silver and lead), sulfates (including barium, calcium and lead), carbonates, hydroxides, Group I cations and ammonium salts</td>
<td>
Solubility Table
<table>
<tr><th>Soluble Salts</th><th>Insoluble Salts</th></tr>
<tr><td>All sodium, potassium, ammonium, Group I and nitrate salts</td><td>---</td></tr>
<tr><td>All chloride salts</td><td>Except lead (II) chloride, silver chloride</td></tr>
<tr><td>All sulfate salts</td><td>Except calcium sulfate, barium sulfate and lead (II) sulfate</td></tr>
<tr><td>SPA carbonates</td><td>All other carbonate salts</td></tr>
<tr><th>Soluble Bases (Alkalis)</th><th>Insoluble Bases</th></tr>
<tr><td>Group I hydroxides
Group II hydroxides (sparingly)
Note: metal hydroxides are not salts (they are bases)</td><td>All oxides (but Group I oxides react with water to form alkalis)
Aluminium hydroxide and Group III hydroxides</td></tr>
</table>
</td>
</tr>
</table>

7.3 Ammonia

Learning Outcomes	**Key Concepts**	**TIPS/ Common mistakes**
(a) describe the use of nitrogen, from air, and hydrogen, from the cracking of crude oil, in the manufacture of ammonia.	**Nitrogen is extracted from air by fractional distillation of liquid air. Hydrogen is produced from the cracking of petroleum.** $N_2 + 3H_2 \rightleftharpoons 2NH_3$	
(b) state that some chemical reactions are reversible, e.g. manufacture of ammonia	Some chemical reactions are reversible. Note the use of the reversible arrow (⇌). $\text{nitrogen} + \text{hydrogen} \xrightleftharpoons{\text{finely divided iron}} \text{ammonia}$	
(c) describe the essential conditions for the manufacture of ammonia by the Haber process	**temperature of 450 °C ; pressure of 250 atm ; finely divided iron catalyst.**	
(d) describe the displacement of ammonia from its salts	**Alkalis + any ammonium salts (e.g NH_4Cl, NH_4Br) → Salt + water + ammonia gas** E.g. $NH_4Cl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) + NH_3(g)$	

8 The Periodic Table

8.1 Periodic trends

8.2 Group properties

<table>
<tr><th>Learning Outcomes</th><th>Key Concepts</th><th>TIPS/ Common mistakes</th></tr>
<tr><td>(a) describe the Periodic Table as an arrangement of the elements in the order of increasing proton (atomic) number</td><td>Across the period, the proton number of the element increases. From left to right across a period, there is a decrease in metallic properties and an increase in non-metallic properties.</td><td>NOTE: Candidates to describe general trends and highlight exceptions.</td></tr>
<tr><td>(b) describe how the position of an element in the Periodic Table is related to proton number and electronic structure</td><td colspan="2">Elements in the Periodic Table are arranged in order of increasing proton number.

Periods:
There are 7 periods (horizontal row of elements) in the Periodic Table. Elements in the same period has the same number of electron shells.

Groups:
There are 8 groups (vertical column of elements) in the Periodic Table. The group numbers are written as Roman numerals I, II, III etc. Elements in the same group will have the same number of valence electrons.</td></tr>
<tr><td>(c) describe the relationship between group number and the ionic charge of an element</td><td colspan="2"><table>
<tr><td>Gp I</td><td>Gp II</td><td>Gp III</td><td>Gp IV</td><td>Gp V</td><td>Gp VI</td><td>Gp VII</td><td>Gp 0</td></tr>
<tr><td>1+</td><td>2+</td><td>3+</td><td>4+/4-</td><td>3-</td><td>2-</td><td>1-</td><td>0</td></tr>
</table></td></tr>
<tr><td>(d) explain the similarities between the elements in the same group of the Periodic Table in terms of their electronic structure</td><td colspan="2">Elements in the same group have the same number of valence electrons and hence have similar chemical properties.
<table>
<tr><th>Element</th><th>Magnesium</th><th>Calcium</th><th>Carbon</th><th>Silicon</th><th>Oxygen</th><th>Sulfur</th></tr>
<tr><td>Electronic configuration</td><td>2,8,2</td><td>2,8,8,2</td><td>2,4</td><td>2,8,4</td><td>2,6</td><td>2,8,6</td></tr>
<tr><td>Group</td><td>II</td><td>II</td><td>IV</td><td>IV</td><td>VI</td><td>VI</td></tr>
</table>
E.g.: Magnesium and calcium have similar chemical properties as they both lose two valence electrons to achieve stable electronic configuration
Exception: In Group 0, He has 2 valence electrons while the rest of the elements have 8 valence electrons</td></tr>
</table>

<table>
<tr>
<td>(e) describe the change from metallic to non-metallic character from left to right across a period of the Periodic Table</td>
<td>From left to right, there is a decrease in metallic properties and an increase in non-metallic properties.
<table>
<tr><td>Group</td><td>I</td><td>II</td><td>III</td><td>IV</td><td>V</td><td>VI</td><td>VII</td><td>0</td></tr>
<tr><td>Symbol</td><td>Na</td><td>Mg</td><td>Al</td><td>Si</td><td>P</td><td>S</td><td>Cl</td><td>Ar</td></tr>
<tr><td>Properties</td><td colspan="3">Metallic</td><td>Metalloid</td><td colspan="4">Non-metallic</td></tr>
<tr><td>Nature of oxide</td><td colspan="2">Basic</td><td>amphoteric</td><td colspan="4">Acidic</td><td></td></tr>
</table>
</td>
</tr>
<tr>
<td>(f) describe the relationship between group no., no. of valence electrons & metallic/non-metallic character</td>
<td>Elements that are metallic (Group I, II, III) lose electrons to form positive ions
Elements that are non-metallic (Group V, VI, VII) gain electrons to form negative ions.
The group number is the same as the number of valence electrons for each element.</td>
</tr>
<tr>
<td>(g) predict the properties of elements in Group I and VII using the Periodic Table</td>
<td>
<table>
<tr><th>Group Number</th><th>I</th><th>VII</th></tr>
<tr><td>Properties</td><td>- Low melting and boiling points
- Low densities
- Soft
- Compounds are ionic, soluble in water and have similar chemical formulas</td><td>- Diatomic simple covalent molecules (elements)
- Low melting and boiling points
- Coloured</td></tr>
</table>
</td>
</tr>
<tr>
<td>(h) describe lithium, sodium and potassium in Group I (the alkali metals) as a collection of relatively soft, low density metals showing a trend in melting point and in their reaction with water</td>
<td>The melting point and boiling point of alkali metals decreases down the group while the density generally increases down the group.

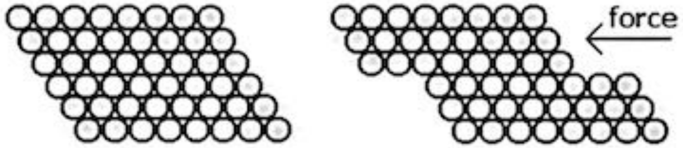
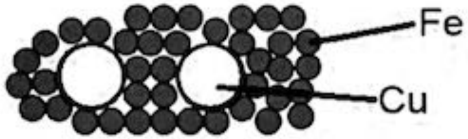
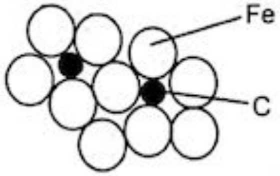
All alkali metals react with cold water to form hydrogen and an alkali.
<table>
<tr><th>Alkali metal</th><th>Observations and equations for reactions with water</th></tr>
<tr><td>Lithium</td><td>Reacts quickly with water, Lithium floats on water and no flame is seen.
$2Li\ (s) + 2H_2O\ (l) \rightarrow 2LiOH\ (aq) + H_2\ (g)$</td></tr>
<tr><td>Sodium</td><td>Reacts very quickly. The molten sodium darts around the water surface. A yellow flame is seen.
$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2\ (g)$</td></tr>
<tr><td>Potassium</td><td>Reacts violently. A lilac flame is seen.
$2K\ (s) + 2H_2O\ (l) \rightarrow 2KOH\ (aq) + H_2\ (g)$</td></tr>
</table>
Sodium atom is bigger than Lithium atom and Potassium atom is bigger than Sodium atom. Reactivity increases down Group I as the tendency to lose the valence electrons increases down the group.</td>
</tr>
</table>

(i) describe chlorine, bromine and iodine in Group VII (the halogens) as a collection of diatomic non-metals showing a trend in colour, state and their displacement reactions with solutions of other halide ions	**Melting points and boiling points increases** down the group. **Colour** of halogens become **more intense/darker** down the group. E.g. F_2 : pale yellow gas Cl_2 : greenish yellow gas Br_2 : reddish brown liquid I_2 : black solid (sublimes to form violet vapour). A displacement reaction is a reaction in which one element takes the place of another element in a compound. **A more reactive halogen will displace a less reactive halogen from its halide solution**. Since size of halogen atom increases down the group, the **reactivity of the halogens decreases down the group** as the **tendency to gain electrons decreases** down the group. Fluorine is the most reactive halogen and Astatine is the least reactive. Examples of displacement reactions: • $Cl_2(g) + 2NaBr(aq) \rightarrow 2NaCl(aq) + Br_2(aq)$ • $Br_2(g) + 2KI(aq) \rightarrow 2KBr(aq) + I_2(aq)$
(j) describe the elements in Group 0 (the noble gases) as a collection of monatomic elements that are chemically unreactive and hence important in providing an inert atmosphere, e.g. argon and neon in light bulbs; helium in balloons; argon in the manufacture of steel	Group 0 elements are referred to as inert gases / noble gases. Noble gases have **stable duplet** (for helium) **or octet configurations** (for other group 0 elements). They **do not lose, gain or share electrons**. Hence, they **do not react to form compounds**, are **monatomic** and are unreactive. Noble gases are mostly used to provide an **inert** atmosphere. Uses of noble gases: • argon or neon is used to fill light bulbs to prevent oxidation of filament • argon is used in the manufacture of steel to prevent oxidation of iron • neon is used in advertising signs • helium is used in balloons
(k) describe the lack of reactivity of the noble gases in terms of their electronic structures	Noble gases have stable duplet or octet electronic configuration with fully filled electron shells, hence they do not need to lose, gain or share electrons with other elements.

<table>
<tr><th colspan="3">8.3 Transition Elements</th></tr>
<tr><th>Learning Outcomes</th><th>Key Concepts</th><th>TIPS/ Common mistakes</th></tr>
<tr>
<td>(a) describe the transition elements as metals having high melting points, high density, variable oxidation state and forming coloured compounds</td>
<td>The transition elements (or metals) are located between Group II and III of the Periodic Table.
<table>
<tr><th colspan="2">Transition elements</th></tr>
<tr><td>Melting point</td><td>High</td></tr>
<tr><td>Density</td><td>High</td></tr>
<tr><td>Oxidation state</td><td>Variable oxidation states (i.e. more than one oxidation states) in its compounds => transition elements can form more than one compound with another element
<table>
<tr><th>Compounds of T.M.</th><th>Chemical formla</th><th>O.S. of the T.M.</th></tr>
<tr><td>Copper(I) oxide</td><td>Cu_2O</td><td>+1</td></tr>
<tr><td>Copper(II) oxide</td><td>CuO</td><td>+2</td></tr>
<tr><td>Iron(II) oxide</td><td>FeO</td><td>+2</td></tr>
<tr><td>Iron(III) oxide</td><td>Fe_2O_3</td><td>+3</td></tr>
<tr><td>Chromium(III) oxide</td><td>Cr_2O_3</td><td>+3</td></tr>
<tr><td>Potassium dichromate(VI)</td><td>$K_2Cr_2O_7$</td><td>+6</td></tr>
<tr><td>Manganese(II) chloride</td><td>$MnCl_2$</td><td>+2</td></tr>
<tr><td>Manganese(IV) oxide</td><td>MnO_2</td><td>+4</td></tr>
<tr><td>Potassium manganate(VII)</td><td>$KMnO_4$</td><td>+7</td></tr>
</table></td></tr>
<tr><td>Coloured Compounds</td><td>Transition elements form coloured compounds.
E.g.
⇨ anhydrous copper(II) sulfate is white while hydrated copper(II) sulfate is blue
⇨ Anhydrous cobalt(II) chloride is blue while hydrated cobalt(II) chloride is pink.
⇨ Used as a test for the presence of water</td></tr>
</table></td>
<td>Transition elements are NOT coloured. It is their compounds that are coloured.</td>
</tr>
<tr>
<td>(b) state that the elements and/or their compounds are often able to act as catalysts (see also 6.1(d))</td>
<td>A catalyst is a substance that increases the speed of reaction.
Transition elements and their compounds are often used as catalysts in industrial processes.
<table>
<tr><th>Industrial Process</th><th>Catalyst</th></tr>
<tr><td>Haber process to manufacture ammonia</td><td>Finely divided iron, Fe</td></tr>
<tr><td>Contact process to produce sulfur trioxide for the manufacture of sulfuric acid</td><td>Vanadium(V) oxide, V_2O_5</td></tr>
<tr><td>Conversion of harmful pollutants to harmless gases in catalytic converters</td><td>Platinum, Pt or Rhodium, Rh</td></tr>
<tr><td>Hydrogenation of vegetable oil to produce margarine</td><td>Nickel</td></tr>
</table></td>
<td></td>
</tr>
</table>

9 Metals

9.1 Properties of metals

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(a) 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	**High melting and boiling point:** Giant metallic lattice structure with **strong electrostatic forces of attraction** between regular lattice of **positive metal ions** and **sea of delocalised mobile electrons**, **high amount of energy** is needed to **overcome** the strong forces of attraction. **Metals are good conductors of heat & electricity** due to the presence of **mobile electrons.** **Malleable**: ability to be bent into different shapes Metal atoms in metal are **closely packed** in **regular** layers. When a force is applied, the **layers** of atoms will **slide** over each other **without disrupting the metallic bonding**. Hence, metal is malleable.  metal atoms when force is applied on layer of metal atoms	Common Mistakes: Negative ions instead of electrons; cannot use break instead of overcome (will be penalised)
(b) describe alloys as a mixture of a metal with another element, e.g. brass; stainless steel	An alloy is a **mixture** of a metal with one or a few other elements e.g. steel is an alloy of iron and carbon.	Common mistake: Alloy is mixture of metals only
(c) identify representations of metals and alloys from diagrams of structures	**Drawing of alloys (see examples below)**: label atoms; **atoms must still be closely packed; disrupted layers; atoms of different sizes for different elements and correct ratio of elements**   	Tip: - Draw the atom with bigger proton number as bigger sized atom
(d) explain why alloys have different physical properties to their constituent elements	Atoms of the added element have a **different size** from those of the pure iron. This **disrupts the orderly/regular arrangement** of the iron atoms, hence the **layers of** iron **atoms cannot slide over each other easily**. This makes the iron alloy harder and less malleable.	Common mistakes: Molecules or ions instead of atoms

9.2 Reactivity series

(a) place in order of reactivity calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to
(i) the reactions, if any, of the metals with water, steam and dilute hydrochloric acid,
(ii)the reduction, if any, of their oxides by carbon and/or by hydrogen

Metal	Reaction with cold water	Reaction with steam
Potassium	Reacts violently. Enough heat is produced to cause hydrogen gas to catch fire. 61	Reacts explosively.(Dangerous!) 61 C
Sodium		
Calcium	Reacts readily.	
Magnesium	Reacts very slowly. C2)	Reacts violently. A bright white glow is produced.
Aluminium	No reaction.	No reaction (due to layer of oxide that coats the metal)
Zinc		Reacts readily. (Note: Zinc oxide is yellow when hot and white when cold)
Iron		Reacts slowly. Iron must be heated constantly in order for reaction to proceed. $3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
Tin Lead Copper Silver Gold Platinum		No reaction.

Reactivity of metals with water/steam

Some metals can react with cold water to form **metal hydroxide** and **hydrogen gas**.

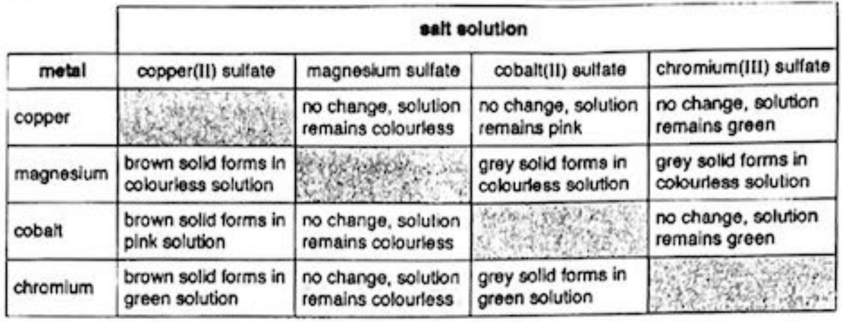
Observations: Effervescence observed and metal dissolves

Test for hydrogen gas: Extinguishes lighted splint with a 'pop' sound

Apart from reacting with water, some metals can react with steam to form **metal oxide** and **hydrogen gas**.

<table>
<tr>
<td></td>
<td>
Reactive metals will react with dilute acid. The higher the metal in the reactivity series, the more vigorous the reaction with acid.
<table>
<tr><th>Too reactive</th><th>Reactive</th><th>Unreactive</th></tr>
<tr><td>Potassium
Sodium
Calcium</td><td>Magnesium
Aluminium
Zinc
Iron
Tin</td><td>Lead*
Copper
Silver
Gold
Platinum</td></tr>
</table>
*Lead should react with dilute hydrochloric acid/dilute sulfuric acid since it is more reactive than hydrogen. However, due to the formation of an insoluble layer of lead(II) chloride/ lead(II) sulfate around lead, lead does not appear to react with the two acids.
<table>
<tr><th></th><th colspan="2">Reduction of metal oxide with</th></tr>
<tr><td>Equation</td><td><u>Carbon</u>
Metal oxide + carbon → metal + carbon dioxide/carbon monoxide</td><td><u>Hydrogen</u>
Metal oxide + hydrogen → metal + steam</td></tr>
<tr><td>Metal oxide</td><td>ZnO to CuO Cold H_2O –Al</td><td>FeO to CuO Steam NA –Al + Fe</td></tr>
</table>
</td>
</tr>
<tr>
<td>(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
(ii)the oxides of the other listed metals</td>
<td>A more reactive metal has a higher tendency to form its positive ions (higher tendency to lose e^-) compared to a less reactive metal.
A more reactive metal can displace a less reactive metal from its salt solution/its oxide.

e.g. Zn(s) + $CuSO_4$(aq) → $ZnSO_4$(aq) + Cu(s)
Zn(s) + CuO(s) → ZnO(s) + Cu(s) [requires heating]</td>
</tr>
<tr>
<td>(c) deduce the order of reactivity from a given set of experimental results</td>
<td>Experimental results (like example given below from O Level 2017 P2) may be given and candidates are expected to deduce the reactivity of the metals based on the results.</td>
</tr>
</table>



	salt solution			
metal	copper(II) sulfate	magnesium sulfate	cobalt(II) sulfate	chromium(III) sulfate
copper	▨	no change, solution remains colourless	no change, solution remains pink	no change, solution remains green
magnesium	brown solid forms in colourless solution	▨	grey solid forms in colourless solution	grey solid forms in colourless solution
cobalt	brown solid forms in pink solution	no change, solution remains colourless	▨	no change, solution remains green
chromium	brown solid forms in green solution	no change, solution remains colourless	grey solid forms in green solution	▨

Based on this set of experimental results, the most reactive metal is magnesium, followed by chromium, cobalt and the least reactive metal is copper.

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

Metal carbonate	**Observation**
Potassium carbonate Sodium carbonate	Unaffected by heat (Except through electrolysis of molten substance)
Calcium carbonate to Copper(II) carbonate	Decompose into metal oxide and carbon dioxide on heating Eg $CaCO_3 \rightarrow CaO + CO_2$
Silver carbonate	Decomposes into silver and carbon dioxide on heating

9.3 Extraction of metals

(a) describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series

Metal	**Oxides**
K **to Al**	Electrolytic Reduction (Very reactive metals, hence electrolysis of molten metal compounds is required to obtain the pure metals)
Zn **to** **Cu**	Reduced by heating with Carbon (For Less reactive metals - Carbon acts as a reducing agent to reduce the metal compounds to form the metal, sometimes Carbon is substituted by Hydrogen in certain questions)
Ag	Reduced by heating alone (Least reactive metal)
Au and Pt	Found chemically uncombined (unreactive metal)

Note: All metals can be extracted from its ores via electrolysis but industries may not do so due to cost issues.

9.4 Recycling of metals

(a) describe metal ores as a finite resource and hence the need to recycle metals, e.g. recycling of iron

Metal ores are a finite resource. With the **increasing demand** for metals, our natural resources will **not last much longer**, hence it is important to **recycle metals**.

(b) discuss the social, economic and environmental issues of recycling metals

Social	• Building new mines reduces the land available for other important uses eg build homes, factories, offices, grow food, rear livestock.
Economic	• May save land needed to build landfill sites, land can be available for other uses. • Usually saves energy extracting new metal from ores, hence lower energy cost. • Recycling metals can be costly – cost incurred to transport scrap metal to processing plants, separate the different types of metals and sort and clean the scrap metal
Environmental	• Helps conserve the limited amount of metal ores in the Earth. • Cut down on pollution such as dumping on wastelands and leaching into soil and river water

Question 1:
Scrap iron can be recycled by adding it to the molten iron, after it leaves the blast furnace. Give one reason, other than cost, why is scrap iron recycled?

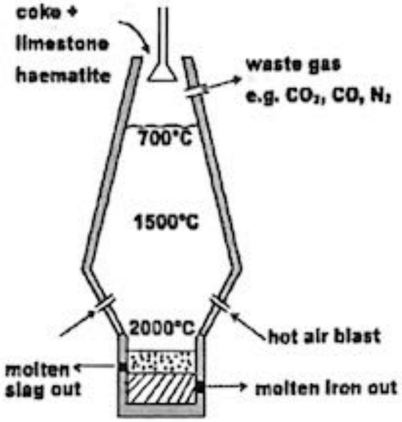
Answer 1:
Recycling helps to save the earth's finite resources for iron / Energy needed for recycling metals is less than that of extraction.

Question 2:
Give a reason why it is important to recycle steel.

Answer 2:
To conserve the finite resource of iron/ disposal of steel requires more landfills.

9.5 Iron

(a) describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace



Molten slag being **less dense** than molten iron forms a layer above the molten iron while molten iron runs to the **bottom** of the furnace. Molten

1. **$C + O_2 \rightarrow CO_2$ + Heat**
Coke (contains carbon) that **reacts with oxygen in hot air to produce carbon dioxide**.

2. **$C + CO_2 \rightarrow 2CO$**
The carbon dioxide produced will then **react with more coke to produce carbon monoxide**.

3. **$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$**
The carbon monoxide is the **reducing agent** in the blast furnace that is needed to **reduce iron(III) oxide in haematite to form iron**.

4. **$CaCO_3 \rightarrow CaO + CO_2$**
Decomposition of limestone (calcium carbonate) by heat in the blast furnace forms calcium oxide

5. **$SiO_2 + CaO \rightarrow CaSiO_3$ (Molten Slag)**
Calcium oxide (basic) reacts with acidic impurities like silicon dioxide through neutralization to form slag

(b) describe steels as alloys which are a mixture of iron with carbon or other metals and how controlled use of these additives changes the properties of the iron, e.g. high carbon steels are strong but brittle whereas low carbon steels are softer and more easily shaped

- The properties and uses of the steel vary with
  - ⇨ The metal(s) added;
  - ⇨ The amounts of carbon and metal(s) added.

Type of steel	Uses	Special properties
Mild steel (low carbon steel) (consist of up to 0.25% of carbon)	Car bodies and machinery	Both mild and high carbon steel are harder than pure iron. Comparing mild and high carbon steel, mild steel is softer and more easily shaped while high carbon steel is stronger and more brittle.
High-carbon steel (consist of 0.45% to 1.5% of carbon)	Knives, hammers, chisels, saws and other cutting tools	
Stainless steel	Cutlery, surgical instruments and equipment in chemical plants	Durable and highly resistant to corrosion

precision/no rust

<table>
<tr>
<td>(c) state the uses of mild steel, e.g. car bodies; machinery, and stainless steel, e.g. chemical plants; cutlery; surgical instruments</td>
<td colspan="2">Higher % of C means that alloy is
• Less malleable
• Less ductile
• Harder
• More brittle</td>
</tr>
<tr>
<td>(d) 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; galvanising</td>
<td colspan="2">■ Rust is a brown solid, mainly consisting of hydrated Iron(III) Oxide, $Fe_2O_3.xH_2O$
■ Essential conditions for rusting are iron in the presence of oxygen and water.
■ Rusting can be prevented by the following methods:
<table>
<tr><th>Method</th><th>Comment</th></tr>
<tr><td>Painting</td><td>If the paint on the metal surface is scratched, rusting will take place under the painted surface.</td></tr>
<tr><td>Oiling/greasing</td><td>The protective film of oil or grease gathers dust and must be renewed.</td></tr>
<tr><td>Plastic coating</td><td>If the plastic layer is torn, the iron starts to rust.</td></tr>
<tr><td>Galvanising (for zinc only)</td><td>Iron does not rust even if the zinc layer is damaged. (This is because zinc is more reactive than iron, hence zinc corrodes in place of iron.)</td></tr>
<tr><td>Tin-plating</td><td>If the tin layer is scratched, the iron beneath it rusts. (This is because tin is less reactive than iron)</td></tr>
</table>
By placing barrier around the metal (iron), it prevents iron from reacting with oxygen in air and water, hence preventing rusting of iron.</td>
</tr>
<tr>
<td>(e) 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</td>
<td>The magnesium blocks in underwater pipes offer **sacrificial protection** to prevent iron from rusting. Magnesium, being a **more reactive** metal than iron, has a higher tendency to **undergo oxidation** and **hence magnesium corrodes in place of iron**. (Note: **Corrosion** is the process whereby a metal loses electrons (is oxidised) to produce a compound, such as an oxide.)

OTHER RELATED ANSWERS TO METALS:
Aluminium is very reactive and it reacts with the oxygen in the air to form a **protective layer** of aluminium oxide, which is coated around the aluminium surface. This layer prevents the aluminium from further reaction with oxygen.

When iron is coated with tin, and the layer of tin is scratched, the iron underneath **rusts**. Iron is **more reactive** than tin, so it has a **higher tendency** than tin **to lose electrons**.</td>
<td>**ONLY iron rusts other metals corrode.**

The **corrosion of iron and iron in steel** is specifically called **rusting** due to the red-brown substance called rust that forms in the presence of water and oxygen.</td>
</tr>
</table>

ATMOSPHERE

10 Air

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(a) describe the volume composition of gases present in dry air	Air is a mixture of several gases.	

↓
- can be separated into into its constituents by fractional distillation of liquid air. ~~Nit~~
- Nitrogen is distilled over first since it has the lowest boiling point

Gas	Composition by volume
Nitrogen	About 78%
Oxygen	About 21%
Carbon dioxide Noble gases • Mostly Argon	Remaining 1%
Carbon dioxide	About 0.03%

- Clean dry air

*Water vapour is present in air even though it's not reflected in the table (Vary from 0-5%) desert / forest)

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

(c) state the sources of these pollutants

$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$

$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$

(colourless) (brown

~~carboxyhaemoglobin~~ carboxyhaemoglobin ~~carboxyhaemoglobin~~

Air pollutants	Source	Effects
Carbon monoxide	Incomplete combustion of carbon-containing substances such as petrol in car engines hydrocarbons	It combines with haemoglobin in blood to prevent haemoglobin from transporting oxygen to the rest of the body, may result in death
Nitrogen oxides (NO, NO_2)	Lightning activity and internal combustion engines (due to high temperatures – oxygen and nitrogen in air react)	❖ Irritates eyes and lungs and cause respiratory difficulties ❖ High levels lead to inflammation of the lungs (bronchitis) ❖ Formation of acid rain (which results in corrosion of buildings, reduced crop yield, harms aquatic life and plants)
Sulfur dioxide	Volcanic eruptions and combustion of fossil fuels ~~S_2~~ $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$	

Other pollutants: Unburnt hydrocarbons (released in car exhaust fumes & chemical plants), effects: ~~carcinogenic~~ carcinogenic, methane (produced when plants & animal matter decay w/ decay of rubbish in landfills), Ozone (formed when NO_2 in air reacts with unburnt hydrocarbons in presence of sunlight)

react with nitrogen oxides to form ozone
↓
effects:
- greenhouse gas which cause global warming

↓
effects:
- forms photochemical smog which irritates eyes & lungs & causes respiratory ~~difficul~~ difficulties
- damage crops

(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 6.1(d))

(ii) the use of calcium carbonate to reduce the effect of 'acid rain' and

Harmful pollutants are converted to ~~harmless~~ harmless / less harmful substances in a catalytic converter.
Carbon monoxide is **oxidised** to carbon dioxide.
Oxides of nitrogen are **reduced** to nitrogen.
Unburnt hydrocarbons are **oxidised** to carbon dioxide and water.

$2CO + O_2 \rightarrow 2CO_2$

$2NO + 2CO \rightarrow N_2 + 2CO_2$

$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

To ↓ SO_2 pollution, can remove sulfur from fossil fuels before burning, but expensive & difficult. Instead, we do flue gas ~~desulfu~~ desulfurisation

Flue gases are waste gases (sulfur dioxide) formed when fossil fuels e.g. coal (which contains sulfur as impurity) undergo combustion.

in flue gas desulfurisation	Calcium carbonate is **basic** and it reacts with **acidic gases** such as sulfur dioxide. calcium carbonate + sulfur dioxide → calcium sulfite + carbon dioxide $CaCO_3(s) + SO_2(g) \rightarrow CaSO_3(s) + CO_2(g)$ ⟷ Calcium sulfate is further oxidised to calcium sulfate by atmospheric oxygen. Alternatively, calcium oxide can also be used to remove the acidic sulfur dioxide. $2CaSO_3(s) + O_2(g) \rightarrow 2CaSO_4(s)$ calcium oxide + sulfur dioxide → calcium sulfite $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$
(e) discuss some of the effects of these pollutants on health and on the environment (i) the poisonous nature of carbon monoxide (ii) the role of nitrogen dioxide and sulfur dioxide in the formation of 'acid rain' and its effects on respiration and buildings pH of unpolluted rainwater (5.5-6.5); slightly below 7 due to $CO_2 + H_2O \rightarrow H_2CO_3(aq)$ (acid rain pH = ~~4≲~~ ≤4)	Carbon monoxide reacts with haemoglobin in blood to form carboxyhaemoglobin, which reduces the ability of haemoglobin to transport oxygen to the rest of the body. It causes headaches, fatigue, respiratory difficulties and even death. Acid rain is formed when acidic pollutants such as sulfur dioxide and nitrogen dioxide dissolve in rainwater. 1. Sulfur dioxide dissolves in water to form **sulfurous acid** (H_2SO_3). sulfur dioxide + water → sulfurous acid $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$ In the presence of oxygen in the air, sulfurous acid is slowly oxidised to **sulfuric acid** (H_2SO_4). 2. In the presence of oxygen and water, nitrogen dioxide is converted to **nitric acid**. nitrogen dioxide + water + oxygen → nitric acid $4NO_2(g) + 2H_2O(l) + O_2(g) \rightarrow 4HNO_3(aq)$ **Effects** Acid rain leaches nutrients from soil and causes plants to wither and die. Acid rain lowers pH of water bodies (e.g. lakes), which may be too acidic for aquatic life to survive. Acid rain reacts with metals and with carbonates in marble and limestone ($CaCO_3$) damaging metal bridges and limestone structures. These gases interact in the atmosphere to form fine sulfate and nitrate particles that can be transported long distances by winds and inhaled deep into people's lungs which can give rise to respiratory problems such as asthma and inflammation of lungs (bronchitis).
(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 acts as a shield or a giant sunscreen to filter out harmful ultraviolet (UV) radiation from the sun. → Compounds containing Cl, F & C Chlorofluorocarbons (CFCs) are very stable and they remain in the atmosphere for a long time. CFCs slowly diffuse through the air. In the presence of UV light, CFCs break down to form chlorine atoms that are very reactive. They react with ozone, destroying the ozone layer. If ozone depletion continues, **more** harmful UV radiation will reach the earth causing skin cancer, genetic mutations and eye damage (e.g. cataracts). Sources of CFCs: Propellants in aerosols, ~~[illegible]~~ coolants in refrigerants & air-conditioners * Ozone is a pollutant at ground level, but important to us at stratosphere (layer of atmosphere at around ≈ 20-30 km above Earth)

*O₃ on ground vs at stratosphere

(g) describe the carbon cycle in simple terms, to include (i) the processes of combustion, respiration and photosynthesis (ii) how the carbon cycle 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 discuss the possible consequences of an increase in global warming	There are two main processes which produce carbon dioxide: **1. Respiration** (aerobic respiration) Energy is released during respiration (**exothermic**). Respiration converts glucose in food we eat into carbon dioxide and water. glucose + oxygen → carbon dioxide + water + energy $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ **2. Combustion** Coal and petroleum and natural gas are examples of fuels. Most fuels contain both carbon and hydrogen. methane + oxygen → carbon dioxide + water vapour + heat energy (**complete combustion** reaction - **exothermic**) However, if limited supply of air is used, methane + oxygen → carbon monoxide + water vapour + carbon + heat energy (**incomplete combustion**) **Photosynthesis** (reverse of aerobic respiration) Plants are essential because they help to remove carbon dioxide from the atmosphere via photosynthesis. Energy is absorbed during photosynthesis (**endothermic**). carbon dioxide + water → glucose + oxygen $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ Carbon cycle is the mechanism that maintains the level of carbon dioxide in the atmosphere, meaning the rate of carbon dioxide that is removed from the atmosphere (e.g. through photosynthesis) must be balanced by the rate of carbon dioxide released to the atmosphere (e.g. through respiration and combustion of fuels). Carbon dioxide and methane are greenhouse gases. The burning fossil fuels and large scale cutting down of forests are causing an increase in amount of carbon dioxide in the atmosphere. Methane is derived from burning of domestic waste and bacterial decay of vegetation. **Effects of Excessive Greenhouse Gas** Increase in Earth's temperature ➔ leads to global warming - More areas become deserts ➔ decrease in crop yields - Melting of ice caps ➔ cause flooding of low-lying lands - More occurrences of unusual weather conditions such as warm spells, droughts, unexpected storms and hurricanes

11 Organic Chemistry

11.1 Fuels and crude oil

Learning Outcomes	Key Concepts	TIPS/ Common mistakes
(a) name natural gas as mainly methane		
(b) describe petroleum as a **mixture** of hydrocarbons and its separation into useful fractions by fractional distillation	**Heat to vapourise the crude oil.** Hot vapour rises up the fractionating column. **The crude oil is heated into vapour and separated into different fractions due to their different boiling points. Each fraction condenses at a different temperature.** **Hydrocarbons with *short* carbon chains (smaller M_r) and lower boiling points condense and escape from the top of the fractionating column.** **Hydrocarbons with *long* carbon chains (larger Mr) and higher boiling points condense and are collected at the *bottom* of the fractionating column.** **Note:** **Common misconception:** The various fractions are collected one by one, based on the lowest boiling point first. **Truth:** Fractional distillation of crude oil in the industrial process actually takes place at the same time for all fractions. Do not confuse with the fractional distillation process of ethanol and water done using common laboratory apparatus.	**Each fraction contains a MIXTURE of hydrocarbons which has a range of boiling points.** Hydrocarbons are organic compounds that contain only carbon and hydrogen.
(c) name the following fractions and state their uses	(i) **petrol** (gasoline) as a **fuel** for cars < petroleum gas (ii) **naphtha** as **feedstock (raw material)** for the chemical industry (iii) **paraffin (kerosene)** as a **fuel for heating and cooking and for aircraft engines** (iv) **diesel** as a **fuel for diesel engines** (v) **lubricating oils** as **lubricants** and for making polishes and waxes (vi) bitumen for making road surfaces	**Each fraction is STILL a MIXTURE. Eg petrol is not pure, it is a MIXTURE.**
(d) describe the issues relating to the competing uses of oil as an energy source and as a chemical feedstock	About 90% of all the petroleum produced is used as fuel to generate heat and electricity. Another 10% is used as chemical feedstock for the manufacture of petrochemicals and medicines. Petroleum is a non-renewable resource and the Earth's petroleum reserves are finite. With the supply of petroleum decreasing rapidly, there is a growing need for its conservation.	

11.2 Alkanes

(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; flammability

Naming of compounds

Prefix

Number of Carbon Atoms	**Name will start with**
1	Meth
2	Eth
3	Prop
4	But
5	Pent
6	Hex

Suffix

Name of Homologous Series	**Functional Group**	**Name of functional group**	**Name will end with**
Alkanes	Nil	Nil	-ane
Alkenes	C = C	Carbon-carbon Double bond	-ene
Alcohols	-OH	Hydroxyl	-anol
Carboxylic Acids	-COOH	Carboxyl	-anoic acid

Characteristics **of Homologous series:**	Organic compounds in the same homologous series have the following characteristics: (a) Have the **same general formula (specify the formula)** (b) **Successive members differ by a –CH_2 – unit** (c) Have the **same functional group** (state functional group) (d) Have similar chemical properties. (e) A gradual change in their physical properties as we go down the series from one member to the next.

(b) describe the alkanes as an homologous series of saturated hydrocarbons with the general formula C_nH_{2n+2}

They are **saturated hydrocarbons** with **general formula C_nH_{2n+2}**, where n = number of carbon atoms in one molecule. Saturated as the carbon atoms in the molecule are joined to each other by **carbon-carbon single** covalent bonds only.

(c) draw the structures of branched and unbranched alkanes, C1 to C4, and name the unbranched alkanes, methane to butane

TIP: 1 Carbon can only form **4 covalent** bonds and hydrogen forms **1 covalent** bond

(d) define isomerism and identify isomers	(a) Isomers are molecules with the **same molecular formula** but **different structural formula**. (b) Isomers of the same compound have **similar chemical properties (if they are from the same homologous series/have the same functional group)** but **different physical properties** e.g. boiling points. Isomers can be due to 1. different ways C atoms are joined (branched vs unbranched), 2. different positions of the functional group, 3. different functional groups.	**TIP:** Isomers need **NOT** be from the same homologous series.
(e) describe the properties of alkanes (exemplified by methane) as being generally unreactive except in terms of burning and substitution by chlorine	**Physical Properties of alkanes**	• Melting point and boiling point increases with **increasing number of carbon atoms/ increasing molecular size** due to **stronger intermolecular forces of attraction, more energy needed to overcome attraction.** • **Viscosity (resistance to flow) increases** as the **molecular size increases**. • **Flammability decreases with increasing molecular size.** • **Insoluble** in water, **soluble** in organic solvent
	Chemical Properties of alkanes	
	1. Combustion of alkanes	**Complete Combustion** when **sufficient oxygen** is supplied, produces only **carbon dioxide and water (vapour).** Reaction is very exothermic thus giving alkanes important role as fuel. E.g. $CH_4 (g) + 2 O_2(g) \rightarrow CO_2(g) + 2 H_2O(l)$ **Incomplete Combustion** when there **is insufficient oxygen**, produces **carbon monoxide and water** / **soot (unburnt carbon) and water**. Eg. $2 CH_4(g) + 3 O_2 (g) \rightarrow 2 CO (g) + 4 H_2O (l)$ **TIP: To balance organic equations, balance the elements in the order of C, H, O.**
	2. Substitution reaction with halogens	Condition: **presence of UV (ultraviolet) light**. The **hydrogen atoms** in the alkane are **replaced** by the halogen atoms **ONE AT A TIME**, forming a **by-product.** E.g. substitution of methane in **EXCESS chlorine** gives rise to **monochloromethane(CH_3Cl), dichloromethane(CH_2Cl_2), trichloromethane ($CHCl_3$) and tetrachloromethane (CCl_4)** and the by-product, **hydrogen chloride**. Eg. substitution of methane in **LIMITED chlorine** gives rise to **monochloromethane(CH_3Cl) and hydrogen chloride only**

11.3 Alkenes

(a) describe the alkenes as an homologous series of unsaturated hydrocarbons with the general formula, C_nH_{2n}	**(Only for alkenes with 1 C=C double bond)**	
(c) 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	**Making of alkenes and hydrogen through Catalytic Cracking of crude oil/ hydrocarbons**	- a process where large long-chain alkane molecule is broken up into smaller molecules which are **more useful** and **higher in demand** - Conditions: at high temperature (**600°C**), pressure and finely divided catalyst, **Al_2O_3 or SiO_2** (NOTE: sometimes **porous pot** is used as it contains both **Al_2O_3 and SiO_2**) e.g. $C_{10}H_{22} \rightarrow C_8H_{18} + C_2H_4$ (NOTE: the **TOTAL number** of carbon and hydrogen atoms **DO NOT CHANGE after cracking**) Eg long-chain alkane →short-chain alkane + short-chain alkene / long-chain alkane → short-chain alkene + short-chain alkene + hydrogen gas
(d) describe the difference between saturated and unsaturated hydrocarbons from their molecular structures and by using aqueous bromine	**Test for saturation**	**Saturated** hydrocarbons: hydrocarbons with **carbon-carbon single** bonds **Unsaturated** hydrocarbons: hydrocarbons with **carbon-carbon double bond/ carbon-carbon triple bond** **Hydrocarbons: compounds containing the elements hydrogen and carbon only** **Test for saturation: Reddish brown aqueous bromine** turns **colourless/ is decolourised** when added to **unsaturated hydrocarbons**. **Aqueous bromine remains reddish-brown** when added to **saturated hydrocarbons**.
(e) describe the properties of alkenes (exemplified by ethene) in terms of combustion, polymerisation and the addition reactions with bromine, steam and hydrogen	**Chemical Properties of alkenes**	**1. Combustion of alkenes** **Complete combustion (sufficient oxygen) produces carbon dioxide and water.** **Incomplete combustion** (insufficient oxygen) produces **carbon monoxide and water / soot (unburnt carbon) and water**. Alkenes burns less readily due to having higher percentage of carbon. **TIP: To balance organic equations, balance the elements in the order of C, H, O.**
		2. Addition reactions of alkenes with hydrogen/ halogen/ steam/ alkene molecules (Polymerisation)

Type of reaction	**Addition of**			
Name of reaction in bracket	**Hydrogen, H_2 (Hydrogenation)**	**Steam (Hydration)**	**Aqueous Br_2**, or any halogens (bromination)	**(Addition Polymerisation)**
Reaction	$C_nH_{2n} + H_2 \rightarrow C_nH_{2n+2}$	$C_nH_{2n} + H_2O \rightarrow C_nH_{2n+1}OH$	$C_nH_{2n} + Br_2 \rightarrow C_nH_{2n}Br_2$	large numbers of identical or similar ***monomers*** (with unsaturated double (or triple) bonds) join together to form a ***long chain molecule*** without losing any atoms or molecules
Catalyst	**Ni** or Pt	**Phosphoric (V) acid** **H_3PO_4**	Nil	Catalyst required
Condition	**200°C**	**300 °C, 60 atm**	Room temperature	1000atm, 200°C
Product	Alkane	Alcohol	Alkyl halide (alkane with halogen atoms)	Poly(name of alkene)
Uses	**Margarine** is made by reacting vegetable oils (unsaturated) with hydrogen.	To **manufacture ethanol** from ethene	**Test for** the presence of **unsaturated hydrocarbons**	Make plastics

Reacts **by breaking the carbon-carbon double bond** and **adding the atoms ACROSS the double bond.**

(f) state the meaning of polyunsaturated when applied to food products (g) describe the manufacture of margarine by the addition of hydrogen to unsaturated vegetable oils to form a solid product	Polyunsaturated means presence of **more than 1 carbon-carbon double bond/ carbon-carbon triple bond.** **Addition of hydrogen, H_2 to unsaturated vegetable oil** **Conditions:** 200°C, nickel as catalyst Unsaturated vegetable oil will turn to form solid margarine (saturated)	**TIP: Presence of 1 C=C double bond reduces the no of hydrogen atoms by 2.**

11.4 Alcohols

(a) describe the alcohols as an homologous series containing the -OH group	Name of **functional group**: **hydroxyl (–OH) group**	
(c) describe the properties of alcohols in terms of combustion and oxidation to carboxylic acids	**Physical Properties**	
	Boiling Points (BP)	• BP increases as molecular size/mass increases due to **more energy** needed to **overcome stronger van der Waals' forces of attraction between molecules/intermolecular forces of attraction.** • NOTE: **covalent bonds ARE NOT broken** (it is the intermolecular forces of attraction that is overcome)
	Solubility	• Soluble in both water and organic solvents. • Solubility in water decreases as molecular size increases
	Chemical Properties / Reactions	
	Combustion	Complete/ Incomplete Combustion
	Oxidation	Alcohols are **oxidised** to the corresponding carboxylic acids. (**no of carbon atoms DO NOT change**) **Water is produced too** E.g. $C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$ purple → colourless orange → green Oxidising agents: **heated/warm** with **acidified $KMnO_4$(aq)** OR **acidified $K_2Cr_2O_7$(aq)** OR **oxygen with the help of bacteria (Fermentation)** **Note: write 2[O] as O_2 in the case of fermentation** Observation: **Purple** acidified $KMnO_4$ solution turns **colourless** - Potassium maganate (VII) **Orange** acidified $K_2Cr_2O_7$ solution turns **green** - Potassium dichromate (VI)
	Dehydration	Alcohols lose water to give alkenes. (Name of reaction: dehydration). Dehydrating agent: concentrated H_2SO_4

<table>
<tr>
<td>(d) describe the formation of ethanol by the catalysed addition of steam to ethene and by fermentation of glucose</td>
<td>
<table>
<tr>
<td>Synthesis from

ethene</td>
<td>Ethene is obtained from the <u>catalytic cracking of crude oil.</u>

Ethanol is then obtained by <u>catalytic addition of steam to ethene</u>

Conditions: 300°C, 60 atm, phosphoric (V) acid as catalyst

$CH_2{=}CH_2 + H_2O \xrightarrow[300^\circ C,\ 60\ atm]{H_3PO_4(aq)} CH_3CH_2OH$

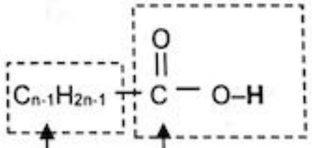
Ethene Steam Ethanol</td>
</tr>
</table>

<table>
<tr>
<td>Fermentation</td>
<td>Fermentation is the slow breakdown of starch/ glucose into ethanol.

$C_6H_{12}O_6(aq) \xrightarrow{Yeast} 2C_2H_5OH(aq) + 2CO_2(g)$

Glucose Ethanol Carbon dioxide

Conditions: <u>37°C and anaerobic/ no oxygen</u>
Source of glucose: <u>Sugar cane plants</u></td>
</tr>
</table>
Note: Fermentation can only yield a maximum of 18-20% ethanol. The yeast cells die beyond this concentration and fermentation stops. To produce ethanol of higher percentage, <u>fractional distillation</u> is used.
</td>
</tr>
<tr>
<td>(e) state some uses of ethanol</td>
<td>e.g. as a solvent for paints and perfumes; as a fuel; as a constituent of alcoholic beverages</td>
</tr>
<tr>
<td colspan="2">11.5 Carboxylic acids</td>
</tr>
<tr>
<td>(a) describe the carboxylic acids as an homologous series containing the $-CO_2H$ group</td>
<td>
<table>
<tr>
<td>Carboxylic Acids</td>
<td>For compounds, the formula is $C_{n-1}H_{2n-1}COOH$ (n=1, 2, 3, ...) <u>OR</u> $C_nH_{2n+1}COOH$ (n=0, 1, 2, 3, ...).


$C_{n-1}H_{2n-1}-\overset{\overset{\displaystyle O}{\|}}{C}-O-H$
Hydrocarbon <u>Carboxyl group</u> (functional group)</td>
</tr>
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</td>
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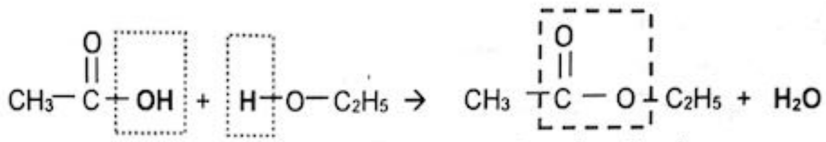
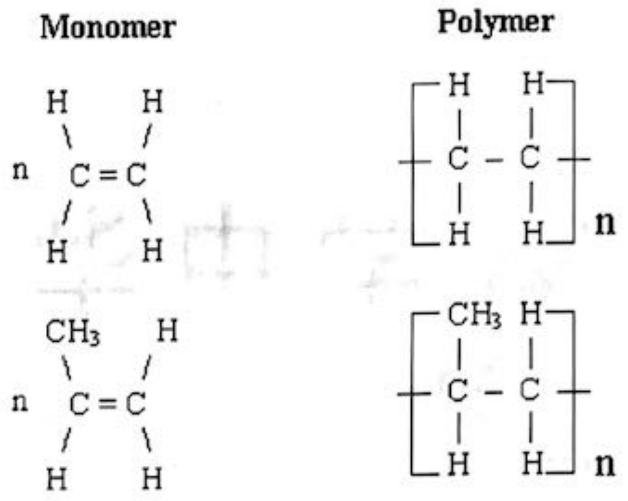
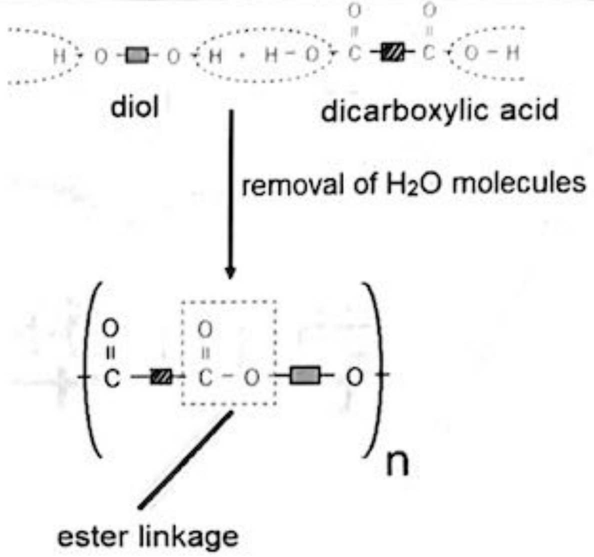
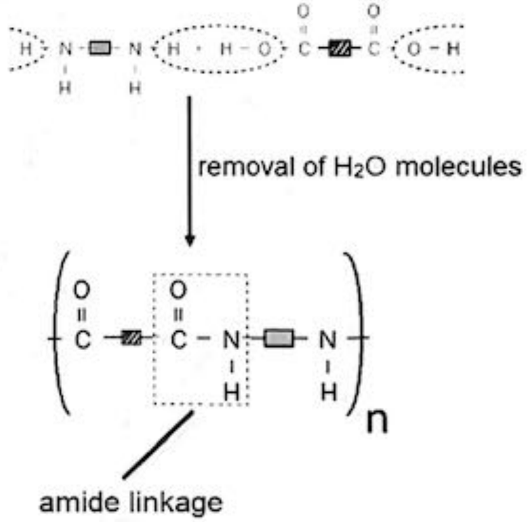
</table>

(b) draw the structures of carboxylic acids, methanoic acid to butanoic acid and name the unbranched acids, methanoic to butanoic acids (do this yourself)			
(c) describe the carboxylic acids as weak acids, reacting with carbonates, bases and some metals	**Acidic** **Properties**	Carboxylic acids, being organic acids, are weak acids which dissociates **partially in water** to give **low concentration of hydrogen Ions** E.g. $CH_3COOH (aq) \rightleftharpoons CH_3COO^- (aq) + H^+ (aq)$ Carboxylic acids behaves like acids in chemical reactions: 1. Reacts with **reactive metals (above hydrogen in metal reactivity series but excluding the very reactive metals like Na, K for safety reasons)** to give **salt and hydrogen** gas. 2. Reacts with **metal carbonates** to give **salt, carbon dioxide gas and water**. 3. Reacts with **bases** (metal oxides and metal hydroxides) to give **salt and water**. (Neutralisation)	
(d) describe the formation of ethanoic acid by the oxidation of ethanol by atmospheric oxygen or acidified potassium manganate(VII)	**Oxidation** **Use as test for identity of alcohol**	Conditions: **heated with oxidising agent, acidified $KMnO_4$ (aq) OR oxygen in air with the help of bacteria** - **write 2[O] as O_2 in this case** Alcohols are **oxidised** to the respective **carboxylic acids (e.g.ethanol is oxidised to ethanoic acid)** and **water** is produced. E.g. $C_2H_5OH + 2[O] \rightarrow CH_3COOH + H_2O$	NOTE: **SAME NUMBER of C atoms in alcohol and carboxylic acid** **TIP**: To draw structure of carboxylic acid, **remove 2 H atoms** from the **C connected to the OH group** and **replace it with a =O**
(e) describe the reaction of a carboxylic acid with an alcohol to form an ester, e.g. ethyl ethanoate	**Esterification**	Conditions: **heating with concentrated sulfuric acid** Reagents: **carboxylic acid and alcohol** Type of reaction: **condensation** reaction Products: **ester and water** **Carboxylic acid + Alcohol → Ester + water**	

	E.g. Ethanoic acid + ethanol → **ethyl ethanoate** + water  $CH_3-\overset{O}{\overset{\parallel}{C}}-OH + H-O-C_2H_5 \rightarrow CH_3-\overset{O}{\overset{\parallel}{C}}-O-C_2H_5 + H_2O$ The **OH group of the COOH in the acid** combined with the **H of the OH in the alcohol** to give water. Naming of ester: **alcohol (....yl), followed by carboxylic acid (........anoate)** **Tip**: when drawing ester, draw acid first. When naming, name alcohol first.
(f) state some commercial uses of esters	e.g. used to make perfumes; in food flavourings; in solvents for cosmetics
11.6 Macromolecules	
(a) describe macromolecules as large molecules built up from small units, different macromolecules having different units and/or different linkages	Macromolecule is a long chain molecule which contains a large number (hundreds or thousands) of **monomers** joined together by covalent bonds through a process called **polymerisation**.
(b) describe the formation of poly(ethene) as an example of addition polymerisation of ethene as the monomer	**Addition Polymerisation** • ***unsaturated monomers*** *join together to form a polymer* ***without the loss of any molecules or atoms.*** • ***Empirical formula*** of the monomer & the polymer is the **SAME**. *(Monomers need* ***NOT be the SAME*** *to be polymerized.)* ***TIP: When drawing the polymer, must change the monomer to 'H' shape with C=C double bond in the middle***
(c) state some uses of poly(ethene) as a typical plastic	e.g. to make plastic bags; clingfilm

(d) deduce the structure of the polymer product from a given monomer and vice versa	**Monomers,** **Polymer,** **Repeat Unit**	 *TIP: If question specify number of repeat units, no need to write n.*	***Note: Do NOT denote 'n' or [ ] for REPEAT UNIT.***
(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**	- *Monomers with 2 reactive functional groups (can be SAME or DIFFERENT) join together to form a polymer with the loss of small molecules such as water* *TIP: When drawing, please follow the question and* *DO NOT use* ☐	**NOTE: small molecules removed can be hydrogen chloride, etc depending on the monomers**

	Terylene			**For terylene, remove the OH from –COOH of the acid and H from the –OH of the alcohol to form water**
	Nylon			**For nylon, *Remove the OH from –COOH of the acid and H from the –NH_2 of diamine to form water***

(f) state some typical uses of man-made fibres such as nylon and Terylene	e.g. clothing; curtain materials; fishing line; parachutes; sleeping bags
(g) describe the pollution problems caused by the disposal of non-biodegradable plastics	Plastics are non-biodegradable which means they **CANNOT** be decomposed by micro-organisms, hence require landfill sites which **takes up land space** and causes **LAND POLLUTION.**

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