Science Centre

Curriculum Management
and eLearning Department

CHEMISTRY FORM IV SYLLABUS

For State Schools

Commencing Scholastic Year 2011 - 2012

Introduction

In view of the changes to the Chemistry SEC 06 syllabus for 2013, there has been the need to update the State School Chemistry syllabi. This syllabus for Form IV students reflects the changes in the SEC syllabus. This new syllabus will affect Form 4 students as from scholastic year 2011 – 2012.
The complete version of the SEC 06 Chemistry syllabus can be accessed directly from the URL address http://www.um.edu.mt/matsec

General Aims

This syllabus aims to:

- stimulate students and sustain their interest in, and enjoyment of, the learning of chemistry and its role in our everyday lives
- provide a relevant chemical background for those students who intend to terminate their study of chemistry at secondary level and also lay a sound foundation for those who intend to pursue their studies in chemistry or related subjects further
- enable students to acquire a knowledge and understanding of basic chemical principles and patterns
- encourage students to apply their chemical knowledge and understanding to familiar and unfamiliar situations
- improve students' abilities to perform experiments through a guided development of relevant practical skills whilst having due regard to correct and safe laboratory practice
- develop students' investigative competence in relation to problem solving situations
- develop students' ability to communicate their chemical knowledge and findings in appropriate ways
- develop students' appreciation of the environmental and technological contributions and applications of chemistry

Scheme of assessment

The annual examination paper consists of a 1 hour 30 minutes written paper. The paper has two sections: Section A and Section B.
Section A comprises about eight questions with a total of 60 marks. All the questions in this section are compulsory. Section B includes three questions of the free response type of which the students are asked to select two. Each question in Section B carries 20 marks. Thus section B includes 40 marks.
The questions in both sections shall test both the recall of chemical concepts as well as the application of knowledge. The questions in the annual examination paper will comprise all the topics covered in form IV.
The final mark of the annual examination is worked out by calculating the total theory/exam mark out of 85% and then adding to it the mark attained by the student in the practical/laboratory work (out of 15%).

Practical work

The mark attained for the practical work is based on an average of all the practical reports presented by the student during the scholastic year. In Form 4 students must carry out and present a total of at least **five** practicals, one of which **must** involve a problem solving investigation (which carries 30 marks). The average mark is calculated by dividing the total marks for say five practicals by six. It is vital that practical work is ongoing and laboratory reports are regularly marked throughout the scholastic year, such that the average mark for the coursework is finalised prior to the Annual Examination. The students' laboratory report files must be available for the possibility of moderation by the Education Officer or College Head of Department for Chemistry at least a week before the annual examination.

Chemical Laboratory Experience

A requisite of the SEC 06 Chemistry syllabus (2013) is that the laboratory reports of the **thirteen** experiments presented by candidates must include **one experiment from each of ten specified sections**, [Refer to Section 5.4 (A) to (J) of the SEC 06 2013 Chemistry syllabus], and three other experiments.

It is important to note that, **not more than three** reports of experiments can be presented **from the same section**. Furthermore, **two** of the thirteen experiments must be of the **investigative (problem solving)** type of practical (which carry 30 marks in order to reflect the greater amount of work required). **Although candidates can present up to three experiments from the same sections, (A) to (J), candidates should not present more than one experiment (for example one experiment and one investigation) for the same sub-topic. For example an experiment and a problem-solving investigation both based on electrolysis of aqueous solutions will not be accepted.**

The experiments which are listed in the SEC syllabus and that can be carried out in Form 4 are as follows:

Unit / Topic	**Nature of experiment**	**SEC syllabus Section 5.4**
Unit 7 Topic 7.2	Electrolysis of three aqueous solutions such as dilute sulfuric acid, potassium iodide solution and copper (II) sufate solution.	(B) (ii)
Topic 7.3	Investigating the reaction of metals (e.g. magnesium, zinc, iron, copper) with dilute hydrochloric acid and dilute sulfuric acid.	(E) (i)
Topic 7.3	Investigating the reaction of metals (e.g. magnesium, zinc, iron, copper) with solutions of their salts in order to determine a displacement series.	(E) (ii)
Unit 8 Topic 8.3	Reaction of Group 2 metals (e.g. Mg, Ca) with water and dilute hydrochloric acid	(F) (ii)
Topic 8.4	Reaction of halide ions (Cl^-, Br^-, I^-) with (i) chlorine water (acidified bleaching solution) and (ii) with lead nitrate solution	(F) (i)
Unit 9 Topic 9.1	Effect of heat on materials such as: magnesium; zinc oxide; sodium nitrate; lead (II) nitrate; sodium hydrogencarbonate; sodium carbonate; copper (II) carbonate; hydrated copper (II) sulfate.	(B) (i)
Topic 9.2	An investigation involving the analysis of both the cation and anion in three 'unknown substances'. The unknowns may be either supplied as solids or in solution	(J) (i)
Topic 9.3	Experiments in acid-base titrimetry using volumetric glassware (pipettes, burettes, volumetric flasks).	(I) (i)
Unit 10 Topic 10.1	Preparation of ammonia from an ammonium salt. Investigations should include the reactions of aqueous ammonia.	(C) (iv)
	Preparation of sulfur dioxide; simple investigations can include the reactions of the gas as an acid and as a reducing agent [e.g. with potassium dichromate (VI)].	(C) (v)

Note: The preparations of gases are intended to be carried out on a test-tube scale and accompanied by simple investigations of the properties of the gases.

Practical work should not be limited to carrying out only those experiments specified in the requirements of the SEC 06 Syllabus Section 5.4.
It is expected that students will be given the opportunity to carry out more than the minimum number of **five** experiments.

CHEMISTRY FORM 4 SYLLABUS

UNIT 7		*ELECTROCHEMISTRY*
Topic	7.1	Ionic theory. Ionic equations. Redox reactions.
	7.2	Action of electricity on materials. Electrolysis.
	7.3	The Reactivity series.
UNIT 8		*THE PERIODIC TABLE PERIODICITY*
Topic	8.1	General structure of the Periodic Table.
	8.2	Trends in properties across a typical period.
	8.3	Trends down typical metallic groups. Groups 1 and 2.
	8.4	Trends down a typical non-metal group. Group 7.
	8.5	The noble gases.
UNIT 9		*CHEMICAL ANALYSIS*
Topic	9.1	Action of heat on materials.
	9.2	Qualitative analysis.
	9.3	Volumetric analysis.
UNIT 10		*THE CHEMISTRY OF CHARACTERISTIC NON-METALS AND THEIR COMPOUNDS*
Topic	10.1	Nitrogen. Ammonia. Oxides of nitrogen. Nitric acid. Nitrates.
	10.2	Sulfur. Oxides of sulfur. Sulfuric acid. Sulfates and hydrogensulfates. Sulfites.
	10.3	Halogens - Chlorine. Hydrogen chloride. Halides.
UNIT 11		*THE GAS LAWS*
Topic	11.1	Volume changes due to physical factors.
	11.2	Molar volume. Volume changes due to chemical factors.

Unit 7 ELECTROCHEMISTRY

Objectives

(i) To extend the knowledge of some aspects of chemistry covered in Form 3 by giving an ionic explanation for certain concepts.

(ii) To demonstrate the migration of ions and stress the need for ‘free ions’ for conduction by ionic compounds.

(iii) To show that in many chemical changes, the reaction follows a general pattern.

(iv) To show that many chemical reactions involve a transfer of electrons between ions and other particles and, that such reactions can be represented by means of ionic equations.

(v) To define oxidation, reduction, oxidising agent and reducing agent, and apply these terms to redox reactions.
To establish that oxidation and reduction occur together.

(vi) To demonstrate the general chemical principles involved in electrolytic processes with reference to specific substances.

(vii) To stress the applications of electrolysis with reference to important industrial electrolytic processes.

(viii) To extend the qualitative investigation of electrolysis to include quantitative aspects of electrolytic cells.

(ix) To deduce a reactivity series for the common metals based on experimental evidence of their reactions.

(x) To apply knowledge of the Activity Series in predicting the reactivity of a metal and its compounds .

(xi) To present the simple cell as a means of producing electrical energy from a chemical reaction.

Unit 7 Ionic Theory. Action of electricity on materials - Electrolysis. The Activity Series.

Topic	Item	Description	Additional notes
7.1	Ionic theory.	- Writing formulae for cations and anions; writing ionic formulae for ionic substances. Representing the change from atoms/molecules to ions, and vice/versa by means of ionic half equations.	
		- Writing balanced ionic half equations for synthesis reactions between a metal and a non-metal.	
		- Writing ionic equations, omitting spectator ions, for specific types of reaction: neutralisation, acid on a carbonate, acid on a sulfite, an alkali on an ammonium salt, precipitation.	~ Students should be able to give the ionic equation for these types of reaction, or to identify the type of reaction represented by a general ionic equation.
		- Writing ionic equations (omitting spectator ions) and the ionic half equations for displacement reactions.	~ To include: a metal displacing hydrogen from an acid; a metal displacing a less reactive metal from a salt solution; a halogen displacing a less reactive halogen from a halide.
		- Thermal decomposition and combustion as two other types of reaction.	~ Represented only by stoichiometric equations.
		- Describing the following terms and processes in terms of ions: acids / alkalis and their strength according to the degree of ionisation; hydrogen chloride dissolved in water as opposed to hydrogen chloride dissolved in methylbenzene; solution; neutralisation; precipitation.	e.g. H_2SO_4 / HNO_3 and CH_3COOH; NaOH / KOH and NH_4OH ~ simple treatment in terms of ionisation / non-ionisation of HCl
		- Defining oxidation and reduction in terms of oxygen and hydrogen gain / loss respectively, and in terms of electron transfer.	~ Identifying the substance oxidised / reduced and the oxidising / reducing agent in a given equation; supporting the choice with a reason.
		- Rules for assigning oxidation numbers and using them to identify whether a substance is oxidized / reduced by changes in oxidation number.	~ Limited to simple binary compounds between a metal and non-metal. Finding the oxidation number of a metal/non-metal in a complex ion or radical is not required.

Topic	Item	Description	Additional notes
7.2	Electrolysis	- An investigation of the action of electricity on solid materials leading to the classification of conductors and non-conductors.	The ability of metals and graphite to conduct electricity due to free electrons in their structure.
		- An investigation of the conduction by molten substances (e.g. lead, sulfur, wax and lead iodide) to establish the need of ‘free ions’ for electrolysis. Describing the electrolysis of a molten ionic compound. Redox reactions at the electrodes.	Can include a demonstration of the migration of ions. N.B. Electrolysis of lead iodide is not recommended unless the laboratory is equipped with adequate safety facilities. In terms of the ions present and the electrode reactions.
		- Electrolytes and non-electrolytes - Strong/weak electrolytes e.g. H_2SO_4 / CH_3COOH - An investigation of the conduction of electricity by some aqueous solutions, using inert electrodes, leading to the factors that influence the preferential discharge of ions in aqueous electrolytes.	Related to type of particles in solution Explained in terms of complete/partial ionization. Using the results to emphasise the role of water in electrolysis. namely: the position of the ion in the electrochemical series, concentration of the ions, and the nature of the electrodes.
		- A description of specific examples of electrolysis: dilute sulfuric acid (acidified water), dilute hydrochloric acid, concentrated aqueous sodium chloride (all using inert electrodes); aqueous copper (II) sulfate using carbon (inert) electrodes and using copper electrodes.	Descriptions to include: a diagram of a suitable apparatus, the ions present in solution, observations at the electrodes, a description of the preferential discharge of ions, the ionic half equations for the electrode reactions, names of products, and changes in the electrolyte.
		- Industrial applications of electrolysis: extraction of aluminium from purified bauxite; the electrolytic purification of copper; electroplating (basic principle only).	To include: starting materials and essential conditions, identity of electrodes, equations for the electrode reactions, uses of the products and a simplified outline diagram. Technical details of the industrial plants are not required.
		- The Faraday as the charge carried by one mole of electrons. Calculating the quantity of charge used during electrolysis. Calculating the mass of a substance / or volume of gas liberated during electrolysis.	Formal definitions of Faraday’s Laws, and experiments on the determination of the laws, are not required. The quantity of electricity required to liberate one mole of an element as a multiple of 96500 coulombs.

Topic	Item	Description	Additional notes
7.3	The Activity Series	The reactivity series related to the tendency of metals to form positive ions. Deriving a reactivity series by investigating the reaction of metals with: - air (oxygen) - water or steam - dilute acids (hydrochloric and sulfuric) - displacement reactions a metal displacing a less reactive metal from an aqueous metallic salt; a metal reducing an oxide of a less reactive metal The stability of metallic compounds related to the position of the metals in the reactivity series, illustrated by: - the ease of obtaining metals from their ores and the method of extraction used - the reduction, if any, of metallic oxides by carbon and by hydrogen, - the thermal decomposition of the hydroxides, carbonates and nitrates of the listed metals. The simple cell as a means of transforming chemical energy into electrical energy	To include the metals: K, Na, Ca, Mg, Al, Zn, Fe, Pb, Cu and Ag. Account for the apparent lack of reactivity of aluminium in terms of the presence of an oxide layer which adheres to the metal. (Thermit type of reaction) ~ Details of the extraction of each metal is not required. (the extraction of iron and aluminium are included in Topics 7.2 and 8.4 respectively). ~ To emphasise the thermal stability of the compounds of the very reactive metals. e.g. Zn / Cu electrodes in dilute sulfuric acid. Explanation should be limited to the half equation showing the reaction of the more reactive metal (e.g. of zinc) resulting in electrons passing through the external circuit to the less reactive metal (e.g. to copper).

Unit 7. **Ionic Theory. Action of electricity on materials - Electrolysis. The Reactivity Series. Learning Outcomes**

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
7.1	Ionic theory.	understand the correlation of charges on metal/non-metal ions to the number of electrons lost or gained respectively				
		recall the charges on common ions met in the syllabus and apply them to write ionic formulae for appropriate substances				
		represent the formation of ions from corresponding atoms or molecules, and vice-versa, by means of ionic half equations; deduce the ionic half equations for a given synthesis reaction				
		recognise types of reaction represented by given chemical equations				
		appreciate that many chemical reactions involve a transfer of electrons between ions and other particles				
		write ionic equations, omitting spectator ions, for reactions that follow a general pattern				
		derive and write ionic equations, omitting spectator ions, for precipitation and displacement reactions; write ionic half equations for displacement reactions				
		identify spectator ions in given equations and rewrite the equations omitting spectator ions				
		use the ionic theory and ionic equations to describe and explain processes involving ionic interactions				
		define oxidation and reduction, as the loss or gain of hydrogen/electrons; gain or loss of oxygen; increase or decrease in oxidation number respectively				
		Work out the oxidation number of elements in simple binary compounds				
		apply the concepts of oxidation and reduction to identify the substance being oxidised/reduced and the oxidising/reducing agent in given reactions				
		work out the oxidation numbers of elements in simple binary compounds; describe and explain oxidation/reduction in terms of changes in oxidation number.				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
7.2	Electrolysis	define anode, cathode, electrolyte, electrolysis				
		understand an electric current as a flow of electrons; distinguish the different ways in which a current passes through a solid conductor and through an electrolyte				
		distinguish between strong/weak electrolytes and non-electrolytes in terms of the degree of ionisation or absence of ions respectively				
		recall a simple experiment to distinguish between strong/weak electrolytes and non-electrolytes				
		show familiarity with the apparatus and apply related terms used for the electrolysis of molten salts or aqueous solutions				
		describe the electrolysis of a specified binary compound in the molten state				
		recall the factors that influence preferential discharge of ions and hence the products of electrolysis				
		recall the general principle that cations migrate to and are discharged at the negative electrode with the formation of metals and hydrogen, while anions migrate to and are discharged at the positive electrode with formation of non-metals (other than hydrogen)				
		apply rules for discharge of ions to deduce/predict the substances liberated at anode and cathode for a given solution				
		use balanced ionic half equations to represent and explain electrode reactions				
		describe in detail the laboratory electrolysis of the following solutions: dilute sulfuric acid, concentrated aqueous sodium chloride, aqueous copper (II) sulfate – all using inert electrodes; aqueous copper (II) sulfate – using copper electrodes				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
7.2 contd	Applications of electrolysis	state that ores ore naturally-occurring compounds of metals				
		describe the extraction of aluminium from purified bauxite, and the purification of copper, including simple cell diagram, nature of electrolyte and electrodes, and reactions				
		relate the properties of aluminium and copper to their uses in industry and in everyday life				
		explain the need for recycling metals in terms of the finite nature of metal resources				
		appreciate that large scale electrolytic processes are expensive				
		understand and explain the chemical principles involved in electroplating and appreciate the need for this process				
	Calculations	recall the factors that affect the quantity of the products obtained during electrolysis				
		utilise Q =It to calculate the quantity of charge used in electrolysis from given data				
		calculate the mass or volume of an element liberated during electrolysis, given appropriate information				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
7.3	Activity Series	recall the activity series as specified in the syllabus				
		understand the activity series as related to the tendency of a metal to form its positive ion				
		justify the position of metals in the activity series by reference to (i) their reactions with water, steam, acids and aqueous solutions of metal ions (ii) the reduction, if any, of their oxides with carbon and with hydrogen				
		understand that reactions of metals with acids, and aqueous solutions of metal ions, are particular examples of displacement reactions; recall that changes in temperature often accompany displacement reactions				
		deduce an order of reactivity from a given set of experimental results				
		predict the position in the activity series of unfamiliar metals, given a description of their reactions; or predict the reactions of an unfamiliar metal given its position in the activity series				
		relate the order of discovery, and the method or ease of extracting a metal from its ores, to its position in the activity series				
		recall that electrolysis is used to extract more reactive metals; while carbon and carbon monoxide can reduce the oxides of less reactive metals				
		relate and explain the ‘Thermit reaction’ in terms of the relative reactivity of metals				
		predict the ease of decomposition of common metallic compounds in relation to the position of the metal in the activity series				
		relate the pattern in the reactions of the metals and their compounds, covered elsewhere in the syllabus, to the activity series				
		use balanced stoichiometric and ionic equations to represent the reactions of metals in this topic				

Unit 8 The Periodic Table. Periodicity.

There has been a drastic change, both to the teaching approach and in the content of the subject matter of the Periodic Table.
A detailed knowledge of the preparation and properties of the individual compounds of each element in a group is not required.
This reduces unnecessary detail and recall, whilst placing emphasis on the similarity of elements and trends going down Groups 1, 2 and 7.
However, some important compounds and related properties/uses are specifically mentioned in the syllabus and are expected to be known.

Objectives

(i) To present the Periodic Table as a classification of the elements in order of their atomic numbers, resulting in areas of the table containing elements with similar properties.

(ii) To show that there is a general trend in properties of the elements, and their compounds, across a typical period.

(iii) To investigate the similarity in properties and the reactivity trend of the elements in typical metallic groups, (Group 1 and Group 2).

(iv) A knowledge of the typical properties of transition metals and their compounds ~ exemplified by iron and copper.

(v) To show the similarity in properties and the reactivity trend of elements in a typical non-metallic group, (Group 7).

Unit 8 The Periodic Table. Periodicity.

Topic	Item	Description	Additional notes
8.1	General structure of the Periodic Table.	-The table as a series of vertical Groups and horizontal periods.	The relationship between the number of electrons in outer shell and Group number; the number of electron shells and period number.
		- Areas of the table which contain reactive metals, transition metals, characteristic non-metals, and the noble gases.	Students should be able to identify these areas as shown on an outline diagram of the Periodic Table.
		- Relating the Group number / electron configuration to metallic / non-metallic character.	Using the Group number of an element to predict its valency, bonding with other elements, properties, etc.
		- Awareness of metalloids (e.g. silicon and germanium)	Simply as borderline elements which show both metallic and non-metallic characteristics. Details of their reactions are not required.
8.2	Trends in properties across a typical period (Na to Ar).	Change from metallic to non-metallic character in elements across a period.	
		- Trend in valency related to the Group number and to the electronic configuration of the element.	To include the charge on ions; the possibility of more than one valency in covalent bonding.
		- Trend in oxides (simple treatment).	Namely the change from basic, through amphoteric to acidic nature; type of bonding and related properties.
		- Electrovalency and covalency in compounds in relation to the position of the element in the table	

Topic	Item	**Description**	Additional notes
8.3	Trends down typical metallic groups.	**Group 1 ~ The Alkali metals Li, Na and K** • Variations down the group: - increase in size and mass of atom; trend in reactivity. • Similarities in the group: - physical properties typical to Group 1 - <u>principle</u> of extraction by electrolysis of fused salts - reaction with oxygen - reaction with water - reaction with non-metals - prediction of the properties of other Group 1 elements from the Group trend.	Exemplified by sodium and potassium To include detailed knowledge of an experiment to show the trend in reactivity, e.g. reacting the metals with water. soft metals, relatively low melting point, low density, etc. N.B. technical details of extraction are not required. exposure to air; burning in air (to form simple oxide M_2O only). e.g. with chlorine to form typical ionic salts. e.g. to predict the properties and reactivity of rubidium.
		Group 2 ~ The Alkaline earth metals Mg and Ca • Variations down the group: - increase in size and mass of atom; trend in reactivity. • Similarities in the group: - <u>principle</u> of extraction by electrolysis of fused salts - physical properties typical to Group 2 Chemical properties ~ reaction with - oxygen - water or steam - dilute hydrochloric acid - prediction of the properties of other Group 2 elements from the Group trend • Limestone - limestone as a raw material; uses of limestone conversion of limestone to quicklime and subsequently to slaked lime	To include detailed knowledge of an experiment to show the trend in reactivity, e.g. reacting the metals with water or with dilute hydrochloric acid. technical details and diagram of method of extraction are not required. higher M. pt. and density, and harder, relative to Group 1 metals e.g. to predict the properties and reactivity of barium the reactions of calcium and its compounds can be presented in the form of a reaction scheme

Topic	Item	Description	Additional notes
8.4	The Transition metals.	The transition elements as a collection of metals which have properties which are distinctly different from those of Group 1 and Group 2. • **Iron** - principle of the extraction from haematite in the Blast furnace, including function of limestone in removing main impurity. - uses of pig iron and steel related to the difference in their physical properties. - chemical properties of iron: reaction with steam, hydrogen chloride and chlorine. • Compounds of iron: Iron (II) and Iron (III) hydroxide - colour; formation by precipitation; reaction with dilute acids; oxidation of iron (II) hydroxide on exposure to air	Can include - hardness and strength, high melting points and densities, variable valency (different charge on ion), coloured compounds, catalytic activity. Technical details are not required however students should be capable of giving a simple, labeled outline diagram of the Blast furnace. Details of methods for converting pig iron to steel are not required. To emphasise that Iron (III) chloride is obtained with chlorine Revise rusting and its prevention. N.B. Students should be capable of following a reaction scheme, where the metal or its compounds are converted from one to the other.
		• **Copper** - revision of electrolytic purification - chemical properties: reaction with oxygen; oxidation by concentrated sulfuric acid; oxidation by nitric acid. - uses of copper and associated properties • Simple compounds of copper: - copper (II) oxide as a typical basic oxide and its use in preparing copper (II) salts by reaction with dilute acids - reduction of copper (II) oxide by hydrogen - thermal decomposition of copper (II) carbonate and copper (II) nitrate to give copper (II) oxide	covered in Unit 7, Topic 7.2 lack of reactivity with water; with dilute hydrochloric and dilute sulfuric acids. Detailed formal equations for the reactions with nitric acid are not required. N.B. Students should be capable of following a reaction scheme, where the metal or its compounds is/are converted from one to the other.

Topic	Item	Description	Additional notes
8.5	Trends down a typical non-metal group.	**Group 7 - The Halogens Cl, Br and I** Exemplified by trends in properties of chlorine, bromine and iodine. • Variations down the group: - size of atom; physical state (increase in melting point and boiling point); relative reactivities. • Similarities in the Group: - coloured, molecular, diatomic - reactions with water Group 1 and Group 2 metals Iron	N.B. A detailed study of the chemistry of Chlorine and its compounds will be covered in Unit 10, Topic 10.3. In this section the emphasis should be on the halogens as a family of elements showing patterns in the group. Theoretical discussion of an experiment to show the relative reactivities, e.g. reaction with iron wool; bleaching effect on water containing an indicator; halogen / halide displacement reactions.
8.6	The noble gases	A group of special non-metals which are unreactive. • General properties • Uses	Relation between lack of reactivity and full outer shell of electrons.

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
8.1 and 8.2	General structure of the Periodic Table. Trends in properties across a typical period (Na to Ar).	appreciate that the Periodic Table is a classification of elements in order of increasing atomic number and its use to predict properties of elements				
		understand the divisions of the Periodic Table; distinguish between a Group (as a vertical column of elements that show similar properties) and a Period (as a row of elements that show a gradual change from metallic to non-metallic character across the period)				
		describe characteristic properties of metals and non-metals and be aware that some elements exhibit a mixture of properties of metals and non-metals; classify an element as ‘metal’ or ‘non-metal’ on the basis of its properties				
		identify the following families of elements and recall their positions in the Periodic Table: the alkali metals, alkaline earth metals, transition metals, halogens, noble gases				
		make the relationship between group number, number of valency electrons and metallic/non-metallic character; understand the correlation between the electron configuration of elements, and charges on ions, to the position of an element in the Periodic Table				
		utilise the sequence in the valencies of the elements across a period to predict formulae of common compounds of the elements				
		recall that the oxides of the elements in Period 3 change from basic, through amphoteric to acidic				
		relate the similarity in chemical properties of elements to the number of electrons in the outer shell				
		state that the reactivity of the elements in group 1 and 2 increases going down the group; describe reactions or experiments to support this statement				
		state that the reactivity of the elements in group 7 decreases going down the group; describe reactions or experiments to support this statement				
		utilise the trends in the Periodic Table to predict properties of other elements, or their compounds, found in the same group or area of the Periodic Table				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
8.3	Trends down typical metallic groups.	**Group 1 ~ The Alkali metals Li, Na and K**				
		recall the typical 'non-metallic' physical properties (soft, low density, etc.) of these metals				
		utilise the results of reacting lithium, potassium and sodium with water to show their similarity as a 'family' of elements and confirm their order of reactivity				
		recall that the oxides of the alkali metals are basic, and the aqueous hydroxides are alkaline; that these are neutralised by acids to form salts				
		predict the properties and reactivity of other metals in this group				
		predict the *general patterns of behaviour* of the common compounds of these elements by making links to other related topics in the syllabus				
		Group 2 ~ The alkaline earth metals Mg and Ca				
		describe and understand the reactions of magnesium and calcium with oxygen, with water and dilute hydrochloric acid; magnesium with steam; utilise the results of these reactions to show the similarity of the elements and confirm their order of reactivity				
		recall that limestone is a naturally-occurring form of calcium carbonate and explain the link between limestone districts and hardness of water				
		describe the thermal decomposition of calcium carbonate to produce calcium oxide				
		describe the effect of water on calcium oxide and recall that the solution produced is alkaline				
		understand that oxides and hydroxides of calcium, that appear to be insoluble, are in fact 'sparingly soluble'				
		explain why calcium oxide and calcium hydroxide are used to neutralise soil acidity				
		describe the reactions by which the common salts of magnesium and calcium can be prepared by making links to the general methods of preparing salts covered in Unit 5, Topic 5.4				
		make predictions about the properties and reactivity of other metals in this group				
		predict the *general patterns of behaviour* of other common compounds of these elements by making links to related topics in the syllabus				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
8.4	The Transition metals. **Fe and Cu**	appreciate that the study of these two metals and their compounds illustrates typical transition metal properties				
		recall that transition metals have high melting points, high densities, a variable valency and their compounds are generally coloured as exemplified by iron and copper				
		recall specific examples of the use of transition metals and their compounds as catalysts				
		state that iron is obtained from iron ore and describe the extraction of iron by the process of smelting in the blast furnace; including an outline diagram, and a description of the key reactions occurring in different parts of the blast furnace				
		recall and explain why iron from the blast furnace has limited uses; relate the uses of steel to its properties				
		apply the term rusting to the corrosion of iron; explain that rusting is an example of oxidation				
		describe the action of steam, hydrogen chloride and chlorine on iron				
		describe the precipitation of iron (II) hydroxide and iron (III) hydroxide				
		explain, (or identify descriptions of), the conversion of iron (II) to iron (III) compounds and vice-versa; identify and explain the oxidation/reduction reactions occurring				
		describe the oxidation of copper by using concentrated nitric acid				
		discuss the preparation, and simple chemical properties, of the common compounds of copper, limited to items specified in this section and related topics in the syllabus				
		state that an alloy is a mixture of metals, or of metals with non-metals, giving examples				
		interpret and explain a given reaction scheme, where the metal or its compounds is/are converted from one to the other.				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
8.5	Trends down a typical non-metal group.	**Group 7 ~ The Halogens Cl, Br and I**				
		recall the physical characteristics such as state at room temperature and variation in colour				
		show that they are a 'family' of elements by discussing similarities in their chemical properties				
		describe and explain the relative reactivity of the halogens as illustrated by halogen-halide displacement reactions; explain this relative reactivity of the elements as oxidising agents				
		predict the properties and relative reactivity of other halogens				
		state or predict typical properties of halides of common metals				
8.6	The noble gases	**The Noble Gases**				
		state that the noble gases are a family of very unreactive elements				
		relate this lack of reactivity to their having full outer shells of electrons				
		recall the monoatomic nature of noble gases				
		relate some uses of noble gases to their physical properties and lack of chemical reactivity				

Unit 9 Chemical Analysis

Objectives

(i) To develop an analytical approach to the investigation of heat on substances.

(ii) To encourage students to make their own observations and identify patterns in the effect of heat on substances.

(iii) To show that the results of the action of heat on an unknown solid may lead to its identification.

(iv) To carry out specified tests for cations and anions.

(v) To show that different ions may give the same result with a particular reagent, hence the need for further distinguishing tests to confirm the identity of such ions.

(vi) To draw inferences from information derived from qualitative tests.

(vii) To apply knowledge of qualitative tests to identify the cation and anion in ‘unknown substances’.

(viii) To use information derived from qualitative analysis to draw inferences.

(ix) To encourage group discussion and develop the skill of communicating the results of practical work effectively.

(x) To practice the correct technique of handling and reading volumetric apparatus, with due regard for safety.

(xi) To emphasise all the necessary steps required to ensure precision and accuracy during a titration experiment.

(xii) To be able to perform numerical calculations involving solutions of known volume and concentration.

Unit 9 Chemical Analysis.

Topic	Item	Description	Additional notes
9.1	Action of heat on materials.	To investigate the effect of heating: • elements - sublimation of iodine - metals ~ magnesium, copper - non-metals ~ carbon, sulfur	This is mainly a revision of the reaction of oxygen with specified elements covered in Unit 3, Topic 3.5. The types of oxide formed should also be revised.
		• compounds that - are thermally stable: e.g. NaCl, anhydrous Na_2CO_3 and oxides	Can include substances that only undergo a physical change e.g. ZnO
		- decompose (into two or more simpler substances) e,g, $CuCO_3$, $Ca(OH)_2$, $NaHCO_3$, $NaNO_3$, $Pb(NO_3)_2$.	It is important to stress that most of these substances decompose according to a general pattern, which enables us to predict the decomposition of similar compounds
		- undergo a reversible change e.g. hydration/dehydration of silica gel and hydrated copper (II) sulfate.	or hydrated cobalt (II) chloride
		• Revision of the combustion of a hydrocarbon e.g. butane (covered in Unit 3, Topic 3.2).	
		A knowledge of the changes in mass (increase, decrease, unchanged), accompanying the action of heat on substances.	

N.B. It is preferable if this topic is covered through an experimental approach whereby the students make deductions from their own observations.
Questions may be set in which students will be required to identify a substance, and the products formed, from a description of the action of heat on it.

Topic	Item	Description	Additional notes
9.2	Qualitative analysis.	• **Tests for cations** ~ - flame test for K^+, Na^+ (considered as a confirmatory test) and for Ca^{2+}, Cu^{2+} (considered as indicative only)	
		- Identification of $Cu^{2+}_{(aq)}$, $Fe^{2+}_{(aq)}$, $Fe^{3+}_{(aq)}$ using $NaOH_{(aq)}$	To include ionic equations for the formation of the insoluble hydroxide.
		- Determining the presence of $Mg^{2+}_{(aq)}$, or $Ca^{2+}_{(aq)}$, using $NaOH_{(aq)}$	Distinguishing which of these two ions is present using a flame test.
		- Determining the presence of $Pb^{2+}_{(aq)}$, or $Al^{3+}_{(aq)}$, using $NaOH_{(aq)}$. (the amphoteric character shown by the hydroxides should be mentioned.)	(Formulae and equations for formation of hydroxometallates are not required.) Distinguishing which of these two ions is present by using potassium iodide solution.
		- Identification of the ammonium ion using $NaOH_{(aq)}$ and testing for ammonia evolved on warming.	
		• **Tests for anions** ~ to identify - CO_3^{2-} by reaction with dilute acid and then limewater - SO_3^{2-} by reaction with dilute acid and then acidified potassium dichromate (VI) - SO_4^{2-} by reaction with acidified barium chloride	To include ionic equations for the reactions involved. Students should also be familiar with the results of testing for a soluble CO_3^{2-}, SO_3^{2-}, or SO_4^{2-}, by first adding barium chloride solution followed by dilute hydrochloric acid. An explanation of the chemical reactions involved, including appropriate equations, should also be given.
		- Cl^-, Br^- and $I^-_{(aq)}$, by addition of acidified silver nitrate,	The relative solubility of the precipitates in ammonia solution is not required.
		- NO_3^- by reduction with aluminium and sodium hydroxide solution to liberate ammonia.	The formal redox equation for this test is not required. The brown ring test is not required.
		• **Tests for gases** ~ oxygen, hydrogen, carbon dioxide, chlorine, ammonia, nitrogen dioxide, sulfur dioxide and hydrogen chloride.	Students should be able to describe and perform these tests, provided all safety measures are taken. Formal equations for the reducing action of sulfur dioxide are not required.

Topic	Item	Description	Additional notes
9.3	Volumetric analysis.	- Definition of standard solution and concentration. Expressing concentration in g dm^{-3} or mol dm^{-3}. The practical steps in preparing a standard solution, including all precautions to ensure accuracy. Numerical examples: Calculating (a) the mass of solute required to prepare a solution of known molarity, (b) the molarity of a solution, given the mass of solute (or the number of moles of solute) dissolved in solution, (c) the number of moles of solute in a solution of known volume and concentration. - The correct technique of handling and reading volumetric apparatus, including safety precautions. - The correct technique of carrying out a titration, emphasising the solutions/liquid to be used for washing specific items of apparatus, selecting an appropriate indicator, recording burette readings and carrying out sufficient number of titres to ensure accuracy of results. Calculations: - to find the unknown concentration of a solution from the results of volumetric analysis. - Given the volume and concentration of a solution (reactant), to find (i) the mass of a solid formed or (ii) the volume of a gas produced.	Numerical examples: converting concentration expressed as a molarity to g dm^{-3} or vice-versa. This can be in $1dm^3$ solution or less. In (b) and (c), the volume of solution will usually be less than 1 dm^3. It is assumed that students themselves will carry out a number of acid/alkali titrations where, in each experiment with a good end-point, two titres should agree to within 0.1 cm^3. Students should be able to perform all calculations from basic principles. Calculation to determine concentration should not be carried out using the formula $\frac{MaVa}{\text{mole ratio (a)}} = \frac{MbVb}{\text{mole ratio (b)}}$ The solution need not be an acid or alkali.

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
9.1	Action of heat on materials.	broadly classify the action of heat on solids as – thermally stable, forming a more complex substance, thermal decomposition, undergoing a stage of a reversible chemical change				
		describe and explain the action of heat on solids specified in the syllabus				
		identify a substance from a description of the result of heating it				
		predict the result/observation of heating compounds similar to those investigated in this topic				
		be aware of potential hazards when heating substances and the related safety measures				
9.2	Qualitative analysis.	understand that many precipitation reactions form the basis of simple tests for ions in solution				
		recall the results of tests for Na^+ and K^+ in solids or solution using a flame test				
		recall tests for the following cations in solution: Ca^{2+}, Mg^{2+}, Pb^{2+}, Al^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, NH_4^+				
		describe the test/result to distinguish between the pairs of ions Ca^{2+} and Mg^{2+}; Pb^{2+} and Al^{3+}				
		describe tests and results to identify CO_3^{2-} and SO_3^{2-} ions in a solid by the action of dilute hydrochloric acid and testing for the gas liberated				
		describe tests/results to identify CO_3^{2-}, SO_3^{2-} and SO_4^{2-} ions in solution using barium chloride solution and dilute hydrochloric acid, or using acidified barium chloride solution				
		describe tests/results to identify/distinguish Cl^-, Br^- and I^- ions in solution using silver nitrate solution and dilute nitric acid, or using acidified silver nitrate solution				
		describe the test/result to identify the NO_3^- ion in solution using sodium hydroxide solution and aluminium turnings and identifying the gas liberated				
		explain and write ionic equations for the precipitation reactions involved in the tests for Ca^{2+}, Mg^{2+}, Pb^{2+}, Al^{3+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, using $OH^-_{(aq)}$ ions				
		write ionic equations for the liberation of gases in tests for CO_3^{2-} and SO_3^{2-} using $H^+_{(aq)}$ ions; and the test for NH_4^+ using $OH^-_{(aq)}$ ions				
		write ionic equations for the precipitation reactions of CO_3^{2-}, SO_3^{2-} and SO_4^{2-} ions with Ba^{2+} ions; and the reaction of halide ions with $Ag^+_{(aq)}$ ions				
		identify unknown simple salts containing two ions, or acids and alkalis, from the experimental results of qualitative analysis				

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
9.3	Volumetric analysis	define concentration and standard solution				
		calculate concentrations in g dm^{-3} and mol dm^{-3} from given quantities of solute and solvent				
		convert concentration expressed in g dm^{-3} to mol dm^{-3} and vice-versa				
		recall the procedure for carrying out acid-base titrations, emphasising the steps to ensure accuracy of results				
		perform calculations from the results of a titration in which guidance on the method may not be given				
		use mole ratios in equations to calculate the mass of a solid formed, or the volume of a gas produced, from the given volume and concentration of a reactant				

Unit 10 The Chemistry of Characteristic non-metals and their compounds.

Objectives

(i) To revise the principle of isolating nitrogen in industry and its uses.

(ii) To prepare dry ammonia, and ammonia solution, in the laboratory by the displacement of ammonia from its salts.
To investigate the properties of ammonia gas and ammonia solution.

(iii) To discuss the principle for the manufacture of ammonia by the Haber process. To discuss the important uses of ammonia.

(iv) To demonstrate the laboratory preparation of nitrogen dioxide and its properties.

(v) An awareness that nitric acid shows some differences, to other dilute acids, in its reactions as a result of its strong oxidising action.

(vi) A knowledge of the general methods of preparation and general properties of nitrates; their use in nitrogenous fertilisers.

(vii) Sulfur as an element that exhibits allotropy. Properties of sulfur and its importance in the chemical industry.

(viii) To prepare sulfur dioxide, and its solution, in the laboratory.
To investigate the reactions of sulfur dioxide which exemplify its acidic nature and its reducing action.

(ix) To outline the essential stages in the manufacture of sulfuric acid by the Contact process; the uses of sulfuric acid.

(x) To describe the properties of dilute sulfuric acid as a typical acid; and concentrated sulfuric acid as having distinct properties.

(xi) A knowledge of the general methods of preparing sulfates and their general properties.

(xii) To describe a method of preparing chlorine in the laboratory and a knowledge of its properties.

(xiii) To discuss the industrial importance of chlorine.

(xiv) Dilute hydrochloric acid as a typical acid.

(xv) A knowledge of the general methods of preparation and the general properties of common metallic chlorides.

Unit 10 The Chemistry of Characteristic non-metals and their compounds.

Topic	Item	Description	Additional notes
10.1	Nitrogen and its compounds.	• **Nitrogen** - Occurrence. Revision of the principle of the industrial isolation of nitrogen from the atmosphere, by fractional distillation of liquid air. - Nitrogen as an unreactive gas and its uses.	Covered in Unit 3, Topic 3.1 Technical details are not required.
		• **Ammonia** Revision of the general techniques/ apparatus required for the generation, drying and collection of gases.	
		- Laboratory preparation of the dry gas; test for the gas.	Drawing and labeling of the apparatus required
		- Manufacture of ammonia by the Haber process, emphasising: raw materials, starting materials, essential conditions, isolation of the product and its uses.	Ideas of dynamic equilibrium, and choice of conditions in terms of Le Chatelier's Principle, are not to be included at this stage. (To be covered in Form 5 – Unit 13.)
		Chemical properties: - catalytic oxidation (to give NO and H_2O); reaction with hydrogen chloride; reducing action on copper (II) oxide.	
		- Laboratory preparation of ammonia solution and an investigation of its properties.	
		- Chemical properties of aqueous ammonia: alkalinity, neutralisation of acids, precipitation of insoluble metallic hydroxides.	Linked to identification of cations (Unit 9, Topic 9.2)
		- Ammonium salts: preparation by neutralisation, reaction with alkalis, sublimation/thermal dissociation of ammonium chloride, use of ammonium salts as fertilisers.	

Topic	Item	Description	Additional notes
10.1 contd.	Nitrogen and its compounds (contd.)	• **Oxides of nitrogen, NO and NO_2** - Nitrogen monoxide: conversion of nitrogen monoxide to nitrogen dioxide (test for the gas). - Revision of the formation of oxides of nitrogen in the internal combustion engine and their removal by a catalytic converter. - Laboratory preparation of nitrogen dioxide by the thermal decomposition of lead (II) nitrate; physical properties and test for the gas. - Chemical properties of nitrogen dioxide: acidity and action with water (linked to acid rain). • **Nitric Acid** - uses. - Nitric acid as a dilute acid, exemplified by its reaction with bases and carbonates emphasising the difference in its reaction with metals. - Oxidising action of the concentrated acid a) with metals (exemplified by its reaction with copper) b) with an iron (II) salt. • **Metallic Nitrates** Revision of: - general methods of preparation - solubility - action of heat on nitrates - test for NO_3^-	Experimental details of the laboratory preparation of the gas is not required. Covered in Unit 3, Topic 3.1 i.e. identification due to its colour The laboratory preparation of nitric acid using a retort is not required. Recall of equations for the reactions with metals is not required; simply emphasise its oxidising action. The formal equations for the redox reactions are not required; simply give ionic half equations for oxidation of copper atoms to copper (II) ions, and iron (II) to iron (III) ions. These items will have been covered in previous Topics.

Topic	Item	Description	Additional notes
10.2	Sulfur and its compounds.	• **Sulfur** - Allotropy of sulfur. Sources and uses of the element. - Chemical properties: burning of sulfur; reaction with metals, e.g. iron	The extraction of sulfur (Frasch process) and details of the laboratory preparation of the allotropes are not required.
		• **Hydrogen sulfide** - its formation by the action of dilute acids on sulfides; toxic nature of the gas	A description of the laboratory preparation of hydrogen sulfide is not required. No experiments involving hydrogen sulfide are to be carried out.
		• **Oxides of Sulfur** - The laboratory preparation of dry sulfur dioxide by the reaction of a dilute acid on a sulfite. Test for the gas (linked to reducing action of the gas.) - Preparation of aqueous sulfur dioxide; acidic nature of the solution (linked to acid rain). - Reducing action of sulfur dioxide, exemplified by its reactions with acidified potasium dichromate (VI) and iron (lll) sulfate. - Uses of sulfur dioxide.	Using acidified potasium dichromate (VI) only. By the inverted filter funnel method. Students will not be expected to recall the formal stoichiometric equation for the reactions with this reagent.
		- **Sulfur trioxide –** acidic nature of the gas	The laboratory preparation of SO_3 is not required.

Topic	Item	Description	Additional notes
10.2	Sulfur and its compounds. (contd.)	• **Sulfuric Acid** - The manufacture of sulfuric acid by the Contact process, to include: raw materials, starting materials, chemical reactions, essential conditions, isolation of the product and its uses. - Properties a) as a typical dilute acid b) when concentrated: hygroscopic nature; dehydrating action (exemplified by action on sugar, and on hydrated copper (II) sulfate); oxidising action (exemplified by the reaction with copper). Action on metallic chlorides to liberate hydrogen chloride. • **Sulfates and Hydrogensulfates** -General methods of preparing salts as applied to sulfates. Preparation of sodium sulfate by titration method; modifying the results of the titration to prepare sodium hydrogensulfate. - Solubility - Revise of action of heat on hydrated copper (II) sulfate. - Some important sulfates: $CaSO_4.2H_2O$ Gypsum and Plaster of Paris, $(NH_4)_2SO_4$ in fertilisers. - Test for SO_3^{2-} and SO_4^{2-}	Ideas of dynamic equilibrium are not required at this stage. The conversion of SO_2 to SO_3 is not to be discussed in terms of Le Chatelier's Principle; simply give the actual conditions used. Dehydration of ethanol not to be covered at this stage (to be covered with Form 5 syllabus – Unit 14). N.B. Not to include details of the laboratory preparation of hydrogen chloride. covered in Unit 5, Topic 5.4 covered in Unit9, Topic 9,1 also mentioned in Unit 5, Topic 5.4. covered in Unit 9, Topic 9.2

Topic	Item	Description	Additional notes
10.3	The Halogens - exemplified by chlorine and its compounds	Revision of: - similarities and trends in properties of the elements in Group 7 ~ The Periodic Table. - displacement of one halogen by another • **Chlorine** - Laboratory preparation of pure, dry chlorine by the oxidation of hydrochloric acid using managanese (IV) oxide. Test for chlorine. Physical properties. Uses of chlorine - Chemical properties, namely the reaction of chlorine: with water and the bleaching action of the solution; with metals e.g. iron; with hydrogen. Oxidising action of chlorine, e.g. on metals and on iron (II) chloride. - Domestic bleaches: their alkalinity and oxidizing action • **Hydrogen chloride and Hydrochloric acid** - test for the gas. - Reaction of hydrogen chloride gas with ammonia gas. - Reactions of dilute hydrochloric acid as a typical acid • **Chlorides of common metals** - General methods of preparing metallic chlorides from dilute hydrochloric acid. - formation of iron (II) chloride and iron (III) chloride. - deliquescence, e.g. of anhydrous calcium chloride;. - the reaction of chlorides with conc. sulfuric acid - the of test for halide ions in solution.	N.B. Limited to chlorine, bromine and iodine. (A discussion of fluorine and its compounds is not required). Described in terms of the relative reactivity of the elements and their oxidising action. (method using potassium manganate (VII) is not required) Exemplified by the liberation of iodine from potassium iodide solution Linked to general methods covered in Unit 5 – Topic 5.4 Covered in Unit 8, Topic 8.4. related to its use as a drying agent Covered in Unit 9, Topic 9.2

Nitrogen and its compounds

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
10.1	Nitrogen	describe the principle of isolation of nitrogen from the atmosphere				
		recall the reaction of nitrogen with oxygen to form nitrogen monoxide and the link to motor car engines				
		state the uses of nitrogen and relate them to its relative inertness				
		make links between this topic and the material covered in Unit 3, Topic 3.1				
	Ammonia	outline the industrial synthesis of ammonia by the Haber process; recall the operating conditions used in the process				
		describe in detail the laboratory preparation of pure, dry ammonia				
		state a simple chemical test for ammonia				
		recall that ammonia is very soluble and show how it can be dissolved safely; recall other physical properties of ammonia				
		recall that ammonia dissolves in water to form a weakly alkaline solution				
		recall and describe the chemical properties of ammonia gas				
		recall and describe chemical reactions involving ammonium hydroxide solution				
		recall that nitrogenous fertilisers can be manufactured by neutralising ammonia with nitric or sulfuric acid				
		recall other uses of ammonia to illustrate its importance in industry and in everyday life				

Nitrogen and its compounds

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
10.1 contd	Oxides of Nitrogen	recall the conversion of nitrogen monoxide to nitrogen dioxide and its use as a test for nitrogen monoxide				
		describe in detail the laboratory preparation of nitrogen dioxide and identification of the gas				
		describe the reaction of nitrogen dioxide with water and its contribution to acid rain				
	Nitric acid	recall the observations of the reaction of concentrated nitric acid with copper and state that this is an oxidising action of nitric acid				
		recall and describe the reactions of dilute nitric acid with insoluble bases, alkalis and carbonates				
	Nitrates	make links between preparing nitrates and the general methods of preparing salts in Topic 5.4				
		make links between this topic and the action of heat on nitrates covered in Topic 9.1				

Sulfur and its compounds

Topic	**Item**	**Learning Outcomes – at the end of this topic students should be able to:**	**Time**	**Difficulty level**		
				A	**B**	**C**
10.2	Sulfur	define allotropy with reference to sulfur; recall the names of the two main allotropes				
	Sulfur dioxide	recall and describe the laboratory preparation of sulfur dioxide				
		describe a simple chemical test for sulfur dioxide				
		recall the physical properties of sulfur dioxide				
		describe the acidic nature of sulfur dioxide and link it to the cause of acid rain				
		describe and explain the reducing action of sulfur dioxide with reference to a specific example				
	Sulfuric acid	recall and describe the stages in the manufacture of sulfuric acid; recall the actual operating conditions used in the Contact process				
		distinguish between the reactions of sulfuric acid as a dilute acid and as a concentrated acid				
		describe and explain specific examples to show concentrated sulfuric acid to be hygroscopic, an oxidising agent and a dehydrating agent				
		recall some uses of sulfuric acid to illustrate its importance in industry and in everyday life				
	Sulfites, sulfates and hydrogensulfates	describe the reaction of sulfites with dilute acids				
		make links between preparing sulfates and the general methods of preparing salts covered in Unit 5, Topic 5.4				
		state the modifications required to produce sodium hydrogensulfate after titrating sulfuric acid with sodium hydroxide solution				

Chlorine and its compounds

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
10.3	Chlorine	describe the laboratory preparation of pure, dry chlorine by oxidation of concentrated hydrochloric acid				
		state a simple test for chlorine				
		describe and explain the oxidising action of chlorine with iron and iron (II) chloride				
		describe and explain the displacing/oxidising action of chlorine on bromides and iodides				
		describe the alkalinity and oxidising action of domestic bleaches				
	Hydrogen chloride and hydrochloric acid	describe a test for hydrogen chloride				
		recall that hydrogen chloride is very soluble in water to form an acidic solution				
		appreciate the difference between hydrogen chloride gas and hydrochloric acid				
		contrast and explain the difference in properties of hydrogen chloride in water and in methylbenzene				
		describe the properties of aqueous hydrogen chloride as a typical acid				
		make links between preparing chlorides and the general methods of preparing salts covered in Unit 5, Topic 5.4				
		Make appropriate links between this topic and the similarities and trends of the halogens covered in Unit 8, Topic 8.5				

Unit 11 The Gas Laws

Objectives

(i) To describe qualitatively the behaviour of particles in the gas state.

(ii) To discuss the basic assumptions of the kinetic theory as applied to an ‘ideal’ gas.

(iii) To describe, using a simple kinetic molecular model, the effect of temperature and pressure on gas volumes.

(iv) To perform calculations on the conversion of gaseous volumes to different temperatures and pressures.

(v) To use the molar gas volume.

(vi) To perform calculations involving stoichiometric reacting masses (or solutions) and volumes of gases.

(vii) The application of Gay-Lussac’s law and Avogadro’s law to gas-to-gas calculations.

Unit 11 The Gas Laws.

Topic	Item	Description	Additional notes
11.1	Volume changes due to physical factors.	- A revision of the behaviour of particles in the gas state. The characteristic properties of gases explained in terms of the simple kinetic theory. - A simple qualitative discussion of the assumptions that the kinetic theory makes about an 'ideal gas'. - The effect of change of temperature, and change of pressure, on gas volumes explained in terms of the kinetic particle theory. Boyle's law and Charles' law; related calculations. -Calculations on the conversion of gaseous volumes to different temperatures and pressures, using the general gas equation, PV/T = constant.	Covered in Form 3, Unit 2, Topic 2.1 e.g. ability to flow easily, gas pressure, diffusion. Effect of increasing the temperature (while keeping the volume constant); and increasing the temperature (while keeping pressure constant). The following will be provided in the 'Useful Data' section on the first page of the Annual Examination paper: stp condtions are to be taken as $0^{o}C$ and 1atm (760 mm Hg); the molar volume for gases at stp is 22.4 dm^3 .

Topic	Item	Description	Additional notes
11.2	Volume changes due to chemical factors.	- The molar volume as the volume occupied by one mole of any gas at stp. (taken to contain 6.0×10^{23} molecules.) The mass of a mole of molecules of a gas. - Calculations to find the volume of a gas that reacts, or is produced, at s.t.p. from the mole ratios of balanced stoichiometric reactions. - Calculating the volume of a gaseous product under conditions other than stp. - Gay-Lussac's law for combining volumes of gases. Avogadro's law. - Calculations involving Gay-Lussac's law: to calculate the volume of a gas that reacts, or is produced, from the relative volume ratios in balanced gaseous reactions.	The molar volume will be given as $22.4dm^3$ at s.t.p. Practice on mole/mass/volume interconversions for gases. Given known amounts of one of the starting materials or products; (a) a given mass of a substance (b) a solution of known volume and concentration. Produced from a reacting mass of a substance, or a reacting solution. Volume-volume relationships in chemical reactions. To include simple problems where the volumes of both reactants are given, but one of them being in excess. To find the total volume of the product and gases remaining.

Topic	Item	Learning Outcomes – at the end of this topic students should be able to:	Time	Difficulty level		
				A	B	C
11.1	Volume changes due to physical factors.	use the kinetic theory to explain the physical properties of gases				
		understand the quantitative relationship between the volume of a gas and temperature and pressure				
		perform simple calculations to convert the volume of a gas from one set of conditions of pressure and/or temperature, to its new volume under different conditions				
11.2	Volume changes due to chemical factors.	convert moles of gases into masses and vice-versa				
		calculate the volume of a given mass of gas (given the molar volume at stp) and vice-versa				
		calculate the mass or volume of a gas reacting or produced (at stp) from mole ratios in equations				
		recall Avogadro's law and apply it to calculate volumes of gaseous reactants and/or products in gaseous chemical reactions				