



Science  
Centre

Curriculum Management  
and eLearning Department

# **CHEMISTRY FORM III SYLLABUS**

## **For State Schools**

**Commencing Scholastic Year 2010 - 2011**

## 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 III students reflects the changes in the SEC syllabus. This new syllabus will affect Form 3 students as from scholastic year 2010 – 2011.

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 III.

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 the first year of chemistry studies, students must carry out and present at least a total of **five** practicals, one of which **must** involve a problem solving investigation.

**N. B. Since the investigation carries 30 marks, the average mark should be calculated by dividing the marks for, say five experiments, 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 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 chromatography will not be accepted.**

**The average mark is to be calculated from the total sum of the marks awarded for the 11 experiments and the two investigations by dividing this total sum by 15.**

The experiments which are listed in the SEC syllabus and that can be carried out in Form 3 are as follows:

Unit / Topic	Nature of experiment	SEC syllabus Section 5.4
Unit 1 Topic 1.7	Extraction of a soluble salt from a mixture by solution, filtration and crystallisation	(A) (i)
	Separation of coloured compounds by paper chromatography	(ii)
	Simple distillation	(iii)
Unit 3 Topic 3.4	Preparation of oxygen using the catalytic decomposition of hydrogen peroxide	(C) (i)
Unit 5 Topic 5.4	Preparing any three insoluble salts by precipitation	(D) (i)
	Preparing a soluble salt by one of the following methods – metal + acid; insoluble base + acid; insoluble carbonate + acid	(ii)
Unit 5 Topic 5.5	Preparation of hydrogen	(C) (ii)

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 in the initial stages of Form 3, students will be given the opportunity to handle equipment and chemicals in order to acquire basic laboratory skills and become familiar with routine experimental methods. It is also important that students are made aware of the hazards associated with common laboratory procedures and the use of chemicals and hence, the precautions or safety measures that need to be taken.

It is assumed that other experiments will be demonstrated (possibly with student involvement), to support the teaching and learning of other topics.

## CHEMISTRY   FORM 3   SYLLABUS

### UNIT 1      BASIC CHEMISTRY

- Topic 1.1    States of matter. Change of state.  
1.2    Physical and chemical change.  
1.3    Elements. Metals and non-metals.  
1.4    Elements and compounds.  
1.5    Compounds and mixtures.  
1.6    Mixed states.  
1.7    Separation techniques.

### UNIT 2      NATURE OF MATTER   ATOMIC STRUCTURE   CHEMICAL BONDING

- Topic 2.1    Nature of matter. The kinetic particle theory.  
2.2    Structure of the atom. Isotopy.  
2.3    Electrovalent bonding.  
2.4    Covalent bonding.  
2.5    Chemical nomenclature. Chemical formulae.

### UNIT 3      THE ATMOSPHERE .   COMBUSTION AND RUSTING OXYGEN AND ITS COMPOUNDS.   OZONE

- Topic 3.1    The composition of air. Air pollution.  
3.2    Combustion.  
3.3    Rusting.  
3.4    Oxygen and its compounds. Ozone.

UNIT 4      WATER

- Topic 4.1 Occurrence. Physical properties. Impurities in water and their removal.  
Water pollution.
- 4.2 Solvent action of water. Solution and solubility.  
Water of crystallisation; Efflorescence, deliquescence and hygroscopy.
- 4.3 Tests for water.
- 4.4 Chemical properties of water.
- 4.5 Hardness of water.

UNIT 5      ACIDS, BASES AND SALTS.

- Topic 5.1 Acids. Investigation of the properties of acids. Chemical equations.
- 5.2 Bases and alkalis. Investigation of the properties of alkalis.
- 5.3 Neutralisation. Salts and acid salts.
- 5.4 Preparation of salts.
- 5.5 Hydrogen and its compounds.

UNIT 6      THE MOLE CONCEPT.

- Topic 6.1 Relative atomic mass. The mole of atoms.  
Molar mass and Avogadro's constant.
- 6.2 Relative molecular mass and relative formula mass.  
Mole of molecules. Mole of ions.
- 6.3 Practical applications of the mole concept. Gravimetric Analysis.
- 6.4 Percentage composition by mass.
- 6.5 Mass relationships in chemical reactions.

## Unit 1      Basic Chemistry

### Objectives

- (i) To provide students with the basic chemical terms and tools which are essential to follow a chemistry course.
- (ii) To create an awareness of the variety of materials and establish the need for scientific classification of substances based on their characteristic properties.
- (iii) To help students appreciate that the study of chemistry is not confined to the substances that they handle in the laboratory, but must extend to everyday materials used at home and in industry.
- (iv) To establish simple ways of determining the purity of a substance.
- (v) Introducing the idea of temporary and permanent changes, leading to an understanding of chemical reactions.
- (vi) To introduce the criteria by which a chemical reaction may be recognised.
- (vii) To familiarise students with the apparatus required for basic chemical investigations, and provide all students with the opportunity of handling equipment and chemicals whilst having due regard for safe laboratory practice.
- (viii) To instil a scientific method of investigation and develop students' skills of observation, and their ability to record accurately and interpret those observations.
- (ix) To provide students with the opportunity to carry out an investigative (problem solving) type of practical .
- (x) As a result of practical work in Unit 1, students should be able to name and identify appropriate apparatus for routine experimental work and for accurate measurements.

**Unit 1      Basic Chemistry                      Content**

Topic	Item	Description	Additional notes
1.1	The states of matter. Change of state.	Physical characteristics of each state. Solid, liquid and gas interconversions.  The state in which a substance will exist at room temperature in terms of melting and boiling point. Melting point and boiling point as criteria of purity.  Interpreting a heating/cooling curve.	An explanation of the physical characteristics of each state and, their interconversion, in terms of the kinetic particle theory is to be dealt with in Unit 2.  Knowledge of experimental details to determine melting point and boiling point are not required.
1.2	Physical and chemical change.	The differences between these types of change to be deduced through suitable experiments.	Identifying physical / chemical changes. Word equations can be introduced at this stage to describe a chemical reaction.
1.3	Elements. Metals and non-metals.	The idea of an element. Use of symbols to represent atoms. Practical work resulting in classifying elements by means of their main <i>physical</i> differences.	An awareness of exceptions in physical properties, e.g. potassium and carbon, is expected. Knowledge of differences in chemical properties is not required at this stage.
1.4	Elements and compounds.	Formation of compounds from elements. Differences between compounds and elements.	The idea of atoms forming molecules may be introduced at this stage.
1.5	Compounds and mixtures.	Experimental work leading to the main differences between them.	

Topic	Item	Description	Additional notes
1.6	Mixed states.	<p>Solid/liquid mixtures:</p> <ul style="list-style-type: none"> <li>- soluble and insoluble substances;</li> <li>- solution - (solvent, solute); suspension.</li> </ul> <p>Liquid/liquid mixtures: miscible/immiscible</p> <p>Solid/solid mixtures:</p> <ul style="list-style-type: none"> <li>- one of the solids is soluble in a specific solvent;</li> <li>- one solid stable to heat while the other sublimes;</li> <li>- a mixture of soluble dyestuffs.</li> <li>- alloys as mixtures of metals with other elements.</li> </ul>	<p>The solubility rule can be included as a tool to predict if a substance is soluble, or not, in water.</p> <p>Experimental determination of solubility is not required.</p>
1.7	Separation techniques.	<p>Separation of:</p> <ul style="list-style-type: none"> <li>• solid/liquid mixtures <ul style="list-style-type: none"> <li>- solute from solution ~ evaporation to dryness; partial evaporation to the point of crystallisation;</li> <li>- solvent from solution ~ simple distillation;</li> <li>- suspension ~ filtration.</li> </ul> </li> <li>• liquid/liquid mixtures <ul style="list-style-type: none"> <li>- immiscible ~ separating funnel;</li> <li>- miscible ~ fractional distillation.</li> </ul> </li> <li>• solid/solid mixtures <ul style="list-style-type: none"> <li>- soluble/insoluble ~ combination of solution, filtration and evaporation;</li> <li>- volatile/ non-volatile ~ sublimation;</li> <li>- soluble dyestuffs ~ paper chromatography.</li> </ul> </li> <li>• gaseous mixtures <ul style="list-style-type: none"> <li>- one is easily liquified ~ liquefaction;</li> <li>- gas mixed with water vapour ~ drying agent</li> </ul> </li> </ul>	<p>Students will be expected to apply a combination of techniques, e.g. solvent extraction, followed by filtration and crystallisation.</p> <p>Students will also be expected to suggest suitable purification techniques, given information about the components of a mixture.</p> <p>Simple treatment only, exemplified by alcohol and water.</p> <p>Fractional crystallisation is not to be included.</p> <p>To include interpretation of simple chromatograms.</p> <p>e.g. NO<sub>2</sub> / O<sub>2</sub> passed through a freezing mixture.</p> <p>N.B. Form 3 students will not be required to give chemical methods of separation, e.g. CO / CO<sub>2</sub> mixture.</p>

**Unit 1. Basic Chemistry Learning Outcomes**

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
1.1	States of matter.	Recall that there are three states of matter – solid, liquid and gas; identify and apply state symbols to substances				
		Describe or list the general physical properties of solids, liquids and gases				
		Discuss the interconversion of solid liquid and gas; recall and apply correctly the terms that describe these interconversions				
		Understand that melting and boiling absorb energy while condensing and freezing liberate energy				
		Interpret a simple heating / cooling curve				
		Show an awareness that we can identify substances or assess their purity from melting point and boiling point information				
1.2	Physical and chemical change.	Classify changes as physical or chemical, giving a convincing reason ; discuss the differences between physical and chemical changes				
		Show an awareness of chemical changes that occur in our day-to-day lives				
		Identify a chemical reaction by a change in appearance or by a detectable energy change; state that a chemical change involves the conversion of one or more substances into new product(s)				
1.3	Elements. Metals and non-metals.	Define an element; classify elements as metals or non-metals on the basis of their physical properties				
		Recall the general physical differences between metals and non-metals; show an awareness of typical exceptions				
1.4	Elements and compounds.	Distinguish between an element and a compound; show an awareness of the use of symbols to represent elements				
1.5	Compounds and mixtures.	Recall the general differences between mixtures and compounds with specific reference to an iron / sulfur mixture and the compound iron (II) sulfide				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
1.6	Mixed states.	Show familiarity with different types of solid/liquid, liquid/liquid and solid/solid mixtures				
		State that a solution is formed when a solute dissolves in a solvent; explain the difference between a solution and a suspension				
		Distinguish between homogeneous and heterogeneous mixtures with reference to simple examples				
		Describe an alloy as a mixture of a metal with other elements				
1.7	Separation techniques.	Recall and describe techniques for separation of solid/liquid mixtures - filtration, crystallisation or evaporation to dryness, distillation				
		Describe the separation of a solid/solid mixture, including the use of a suitable solvent, and by sublimation				
		Describe paper chromatography and interpret simple chromatograms				
		Describe methods of purifying liquid/liquid mixtures, namely fractional distillation and the use of the separating funnel				
		Describe simple laboratory methods of separating or purifying a gas/gas mixture				
		Suggest suitable separation/purification techniques, or a combination of techniques, given information about the substances in a mixture				
		Plan a method of isolating a pure substance and test the method experimentally (problem solving situation)				
Basic skills.		Recall and correctly use basic chemical terms				
		Identify and name appropriate apparatus for measurement or for performing basic routine chemical techniques				
		Draw clear diagrams of single items of apparatus or experimental set-ups				
		Handle equipment and chemicals correctly, whilst showing due regard to safety				

**Unit 2      Nature of matter. Atomic structure. Chemical bonding.**

**Objectives**

- (i) To establish the particulate nature of matter through experimental evidence.
- (ii) To deduce the particle arrangement in the states of matter and explain their interconversion in terms of the kinetic particle theory.
- (iii) Using the kinetic theory to explain diffusion and Brownian motion.
- (iv) To give facts about the relative charges and relative masses of the sub-atomic particles and their arrangement in the atom.  
To explore models of atoms leading to the diagrammatic representations of atoms.
- (v) To use the atomic number and the simple structure of atoms to explain the basis of arranging elements in eight groups of the Periodic table.
- (vi) To define isotopy. To distinguish between mass number and relative atomic mass with reference to isotopes.
- (vii) To emphasise the stability and lack of reactivity of the noble gases in relation to full outer shells of electrons.
- (viii) To relate the formation of ions to the attainment of a stable electronic (noble gas) configuration.  
To describe the formation of ionic bonds between metallic and non-metallic elements as a result of transfer of electrons.
- (ix) To describe the formation of single covalent bonds as the sharing of pairs of electrons leading to the noble gas configuration.
- (x) To establish the relationship between electron configuration and valency.  
Valency in terms of the loss, gain or sharing of electrons by a particular element leading to chemical formulae.

**Unit 2      Nature of matter. Atomic structure. Chemical bonding.**

Topic	Item	Description	Additional notes
2.1	Nature of matter. The kinetic particle theory.	<ul style="list-style-type: none"> <li>- Evidence for the existence of particles. Explaining simple observations using the particle theory.</li> <li>- Models of particles in the states of matter. Explaining the characteristic properties of solids, liquids and gases in terms of the particle arrangement in each state. Interconversion between the three states of matter in terms of the kinetic theory.</li> <li>- Diffusion and Brownian movement explained in terms of simple kinetic theory.</li> <li>- Idea of size of particles.</li> </ul>	<p>e.g. dilution of coloured solution; reduction in volume on mixing equal volumes of ethanol and water.</p> <p>Detailed kinetic theory of gases and Gas laws not to be covered. [ Form 4 Unit 11 ]</p> <p>Energy requirements for change of state in terms of the structures not to be included for Form 3 students. [ Form 5 Unit 15 ]</p> <p>To include a comparison of the rate of diffusion of two different gases, e.g. ammonia and hydrogen chloride.</p> <p>Qualitative treatment only; recall of size of numerical values of particles is not required. The 'oil drop experiment' is not required.</p>

Topic	Item	Description	Additional notes
2.2	Structure of the atom.  Isotopy.	<ul style="list-style-type: none"> <li>- Nuclear model of the atom. Sub-atomic particles; proton, electron and neutron; relative charges and relative masses.</li> <li>- Mass (nucleon) number and Atomic (proton) number.</li> <li>- Electron configuration. Diagrams representing the structure of atoms (restricted to elements of atomic number 1 to 18).</li> <li>- Stability of the noble gas electronic structures.</li> <li>- Isotopes and relation of isotopy to relative atomic mass. Effect of difference in neutrons on mass of atoms exemplified by <math>^{35}\text{Cl}</math> and <math>^{37}\text{Cl}</math>.</li> </ul>	<p>Do not discuss the mass spectrometer.</p> <p>To include predicting if the atom represents a metal, a non-metal or a noble gas, from the number of electrons in the outer shell of the atom.</p> <p>To include determining the relative atomic mass of an element, given the mass numbers and relative abundance of its isotopes (or given the RAM to indicate which isotope is more abundant) – limited to <u>two</u> isotopes.</p>
2.3	Electrovalent bonding.	<ul style="list-style-type: none"> <li>- Relation of electron configuration to electrovalency. Ionic bonding - formation of ions by loss or gain of electrons. Relationship between number of electrons lost or gained, and the charge on the ion .</li> <li>- Diagrammatic representation of ionic bonding showing (i) all electron shells (ii) only the outer shell electrons</li> </ul>	<p>As governed by the octet rule.</p> <p>Writing of ionic formulae not to be included. [ Form 4 Unit 7 ]</p>
		<ul style="list-style-type: none"> <li>- To deduce the number of protons, neutrons, electrons (and electron configuration) present in ions given the atomic number and mass number, and charge on the ion.</li> </ul>	<p>Also to deduce if a particle is an atom or ion (and its charge), given relevant information about the number of protons and electrons.</p>
		<ul style="list-style-type: none"> <li>- Physical properties of ionic compounds – high melting points, solubility in water, conductivity.</li> </ul>	<p>Structure of an ionic crystal lattice e.g. NaCl not to be covered</p> <p>- Form 3 students will not be expected to give reasons for the properties by referring to the structure. [ Form 5 Unit 16 ]</p>

Topic	Item	Description	Additional notes
2.4	Covalent bonding.	<p>Covalent bond formation in simple molecules by sharing of electron pairs leading to the noble gas configuration – as governed by the octet rule.</p> <p>(a) in elements – H<sub>2</sub>, Cl<sub>2</sub>, O<sub>2</sub>, and N<sub>2</sub>.</p> <p>(b) in compounds – e.g. HCl, H<sub>2</sub>O, NH<sub>3</sub>, CH<sub>4</sub>, CO<sub>2</sub></p> <p>‘Dot-and-cross diagrams’ to show the sharing of valency electrons, as well as non-bonding electrons, in covalent compounds.</p> <p>- Physical properties of covalent substances.</p> <p>Reasons for these properties to be explained in relation to weak forces between molecules and the fact that these particles are neutral.</p>	<p>Dative covalency not to be covered.</p> <p>Single, double and triple bonds.</p> <p>A discussion of bond polarity and hydrogen bonding is not to be included.</p> <p>All intermolecular forces to be simply referred to as ‘weak forces’.</p> <p>Shapes of molecules are not required.</p>
2.5	<p>Chemical nomenclature.</p> <p>Chemical formulae.</p>	<p>Naming of substances, given their formulae.</p> <p>Recall of names and formulae of common substances, including gases and acids.</p> <p>Use of prefixes mono, di, etc. and Roman numerals (I), (II), etc. in the naming of substances.</p> <p>Use of the symbols and valency to write formulae.</p> <p>Relationship between the valency and the electron configuration. Valency as the number of electrons lost, gained or shared by an atom.</p>	<p>Identifying whether symbols/formulae represent atoms/ molecules of an element; molecules of a compound; or the formula of an ionic compound.</p> <p>N.B. Form 3 students are not required to write ionic formulae. [ Form 4 Unit 7 ]</p> <p>Reference to the Periodic table is suggested (e.g. the relationship between electron configuration/valency and the Group number in which the elements are found). However, direct questions on Periodic table will not be set in the Annual Exam paper for Form 3.</p>

**Unit 2**    **Nature of matter. Atomic structure. Chemical bonding.**                      **Learning Outcomes**

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
2.1	Nature of matter. The kinetic particle theory.	Recall simple experiments leading to the smallness of particles and to their motion				
		Describe and draw the arrangement of particles in solids, liquids and gases				
		Explain the properties of solids, liquids and gases in terms of the kinetic particle theory				
		Describe changes in state in terms of the change in movement and distance between particles				
		Describe and explain diffusion and Brownian motion in terms of the movement of particles				
		Recall that small, light particles diffuse faster than large, heavy ones				
2.2	Structure of the atom. Isotopy.	State that every element is made up of very small particles called atoms				
		Appreciate that atoms are the building blocks which make up all substances				
		Recall that atoms are made from protons, electrons and neutrons				
		Recall the relative mass and the relative charge of the sub-atomic particles and give their location in an atom				
		Understand and explain the terms atomic number and mass number in relation to the respective sub-atomic particles				
		Find the number of protons, electrons and neutrons in an atom from the mass number and atomic number; use and interpret symbols such as ${}_{11}^{23}\text{Na}$				
		Deduce and state the electronic configuration of the first eighteen elements (given their atomic number)				
		Use the number of electrons in the outer shell to predict if an element is a metal, non-metal or noble gas				
		Use the mass number and atomic number, together with the build up of electrons in shells to draw a model of an atom; describe the structure of an atom				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
2.2 contd.	Structure of the atom. Isotopy.	Define and explain the existence of isotopes				
		Understand that the relative atomic mass of an element takes into account the different proportions of each naturally occurring isotope				
		State which of two isotopes is more abundant, given their mass numbers and the relative atomic mass of the element				
		Calculate the relative atomic mass of an element, given the mass numbers and the respective percentage abundance of two isotopes				
2.3 and 2.4	Electrovalent bonding and Covalent bonding.	Appreciate the importance of the noble gas electron configuration (stability of complete outer octet) and hence their inertness				
		Understand that atoms of different elements can combine to form compounds by the formation of new chemical bonds				
		Recall that compounds contain two or more types of atom chemically bonded together				
		Understand that chemical bonds can be made by the transfer or sharing of electrons leading to the noble gas configuration				
		Recall that that electron transfer occurs between metals and non-metals and results in the ionic bond				
		Recall that ions are charged particles; recall that metals form cations by electron loss, that non-metals form anions by electron gain; link electron configuration with ionic charge				
		Explain, using dot/cross diagrams the formation of $\text{Na}^+$ and $\text{Cl}^-$ ions				
		Show the formation of ions in other ionic compounds				
		Know the characteristic properties of ionic compounds				
		State that (usually) only atoms of non-metal elements bond to form molecules				
		Recall that electron sharing between atoms of non-metals results in the formation of covalent bonds				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
2.3 and 2.4 contd.	Electrovalent bonding and Covalent bonding contd.	Understand that a single covalent bond is a pair of shared electrons while a double bond is two pairs of shared electrons				
		Recall that sharing of electrons between atoms of the same non-metal results in simple molecules of the element; define atomicity using examples of elements that exist as molecules				
		Explain, using dot / cross diagrams, the formation of single covalent bonds in compounds, including HCl, H <sub>2</sub> O, NH <sub>3</sub> and CH <sub>4</sub> .				
		Know the characteristic properties of covalent substances				
		Compare the differences in volatility, solubility and electrical conductivity between ionic and covalent compounds				
		Predict the type of bonding in a substance from its physical properties				
		Be aware that chemical reactions can be shown by word equations				
		Become familiar with writing simple symbol/formula equations and understand the need for balancing equations				
2.5	Chemical nomenclature. Chemical formulae.	Recall that symbols represent atoms of elements				
		Recall and write correct formulae for simple molecular elements and simple molecular compounds relevant to the syllabus				
		Know the relationship between valency and the electron configuration of elements				
		Use the symbols and valency of elements to write formulae of compounds				
		Name and give the formulae of the most common radicals encountered in the syllabus; write formulae for compounds containing radicals				
		Deduce the number of atoms of each element present in the formula of a compound				
		Deduce the valencies of elements and radicals in a formula				
		Understand a system of naming of substances; use correct chemical nomenclature				

**Unit 3**      **The atmosphere. Combustion and rusting.**  
**Oxygen and oxides.**

**Objectives**

- (i) To encourage experimental investigations in order to establish the role of air in chemical processes.
- (ii) To show that air is a mixture of gases and demonstrate the presence of these gases in air.  
To discuss how some of these gases can be isolated in the laboratory and in industry.
- (iii) To create an awareness about the main pollutants in air - their source, effect on the environment and possible remedies.
- (iv) To plan a series of experiments, using metals, in order to confirm that air is essential for combustion; to establish that only part of the air is used up during combustion and that this active part is oxygen.
- (v) Using suitable experiments, to show that the products of complete combustion of a hydrocarbon fuel are carbon dioxide and water.
- (vi) To investigate the conditions necessary for rusting and to show that this process has some aspects in common with combustion.
- (vii) To familiarise students with handling techniques for the preparation, drying and collection of gases.
- (viii) To guide students in carrying out the preparation of oxygen and investigating its properties.
- (ix) To investigate the properties of oxides.
- (x) To introduce the existence of allotropes. Ozone as an allotrope of oxygen.





**Unit 3 The atmosphere. Combustion and rusting. Oxygen and Oxides. Learning Outcomes**

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
3.1	The composition of air. Air pollution.	Evaluate evidence to show that air is a mixture of gases				
		Recall the gases present in clean air and their approximate percentage by volume				
		Describe an experiment to determine the approximate percentage of oxygen and nitrogen in the atmosphere				
		State that the main gases in air are nitrogen (78%) and oxygen (21%)				
		Describe experiments to show the presence of other gases in air (e.g. carbon dioxide, water vapour)				
		Recall and describe the principles involved in the industrial isolation of nitrogen, oxygen and noble gases from liquid air by fractional distillation; recall some uses of nitrogen, oxygen and noble gases				
		Name the common pollutants in the air; state their source and their adverse effect on structures, organisms or health				
		Describe the conditions leading to ‘acid rain’ and ‘global warming’				
		Suggest possible remedies for the effects of pollutant gases in air				
3.2	Combustion.	Describe, in simple terms, ideas related to combustion; state that combustion is a reaction of a substance with oxygen, giving out energy				
		Describe a simple experiment that shows that (a) oxygen is required (b) oxygen is used up during combustion				
		Describe an experiment to show that the products of complete combustion of a hydrocarbon are carbon dioxide and water				
		Understand that the production of carbon dioxide and water vapour, on burning, indicates the presence of carbon and hydrogen in the original fuel				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
3.3	Rusting.	Recall and describe an experiment to investigate the conditions under which iron rusts				
		Show an understanding of and describe simple experiments to show the chemical similarities between combustion and rusting				
		Recall methods by which rusting of iron may be prevented				
3.4	Oxygen. Oxides.	Describe the laboratory preparation of oxygen from hydrogen peroxide				
		State a simple test for oxygen				
		Recall the physical characteristics of oxygen				
		Discuss the reactions with oxygen of magnesium, copper, carbon and sulfur				
		Discuss the type of oxides formed in the above reactions				
		Classify oxides as either basic, acidic, amphoteric or neutral related to metallic / non-metallic character				
		Discuss the properties of the different types of oxide				
		Define allotropes; recall that ozone is an unstable allotrope of oxygen and compare these allotropes				
		Show an awareness of the importance of ozone in the higher regions of the atmosphere in contrast to its toxicity if present close to Earth				
		Write balanced chemical equations for all reactions covered in this Unit				

## Unit 4    Water.

### Objectives

- (i) To show that naturally occurring water is not pure since it contains both solid and gaseous impurities which accumulate over time, hence the need for methods of purifying water depending on its intended use.
- (ii) To emphasize that, although we tend to take water for granted, the availability of natural water in Malta is limited and has to be supplemented by other sources such as reverse osmosis.
- (iii) To emphasise the solvent action of water and related properties.
- (iv) To create an awareness of the pollutants in water and their effect on the environment.
- (v) To show that crystals of various salts contain water of crystallisation and that the shape of the crystals is dependant on this water. To distinguish between the phenomenon of efflorescence, deliquescence and hygroscopy.
- (vi) To show that the presence of water can be detected by means of chemical tests but that definite physical tests are required to confirm that a liquid is pure water.
- (vii) To emphasize that water is a chemical and takes part in a variety of chemical reactions.
- (viii) To show that metals with a high affinity for oxygen will remove it from water or steam, forming the oxide or hydroxide respectively, and liberating hydrogen in the process.
- (ix) To establish that hardness of water is caused by soluble salts of calcium or magnesium, and to appreciate the natural processes which result in these specific substances accumulating in water. To clarify the terms temporary and permanent hardness.
- (x) To relate hardness of water to some natural processes and also its effect on our everyday lives.

**Unit 4    Water.**

Topic	Item	Description	Additional notes
4.1	<p>Occurrence, physical properties.</p> <p>Naturally occurring impurities.</p> <p>Pollution of water.</p>	<p>- The occurrence of water; physical properties of pure water.</p> <p>- Solid and gaseous impurities due to natural processes; methods of purifying water including a simple, qualitative treatment of the technique of reverse osmosis for purifying brackish water and sea water.</p> <p>- An awareness of the common water pollutants and their effects.</p>	<p>Students should be aware that air dissolved in water has a different composition from ordinary air due to the difference in solubility of nitrogen and oxygen.</p> <p>An experimental determination of the proportion of air in water is not required.</p> <p>This can include fertilisers, pesticides, detergents, sewage, heavy metals, crude oil and thermal pollution.</p> <p>Note that 'Pollution of water' is not included in the SEC 06 syllabus.</p>
4.2	<p>Solvent action of water. Solution and solubility.</p> <p>Water of crystallisation.</p> <p>Efflorescence, deliquescence and hygroscopy.</p>	<p>- Revision of terms: solvent, solute and solution.</p> <p>- The solubility rule.</p> <p>- Saturated solutions and interpretation of solubility curves.</p> <p>- Hydrated salts; to show the presence of water of crystallisation in a hydrated salt e.g. <math>\text{CuSO}_4 \cdot 5\text{H}_2\text{O}</math>.</p> <p>- Specific examples can include: washing soda, anhydrous calcium chloride, sodium hydroxide, concentrated sulfuric acid and silica gel.</p>	<p>Quantitative work (both experiments and calculations) to determine the solubility of a salt in water, to be left out.</p> <p>The experimental determination of a solubility curve is not required. The polar nature of water is not required.</p> <p>To include the fact that the action of heat on hydrated copper (II) sulfate is a reversible reaction.</p> <p>To include reversibility of dehydration/hydration of silica gel.</p>

Topic	Item	Description	Additional notes
4.3	Tests for water.	Chemical tests for the <i>presence</i> of water and physical tests for <i>pure</i> water.	
4.4	Chemical properties of water.	<ul style="list-style-type: none"> <li>- Reaction of metals with               <ul style="list-style-type: none"> <li>(i) water e.g. K, Na, Ca and Mg</li> <li>(ii) steam e.g. Mg, Zn and Fe.</li> </ul> </li> <li>- Reaction with non-metals               <ul style="list-style-type: none"> <li>(i) chlorine with water</li> <li>(ii) carbon with steam.</li> </ul> </li> <li>- Reaction with (soluble) basic and acidic oxides to form alkaline and acidic solutions respectively.</li> </ul>	<p>Including observations and equations.</p> <p>This was also included in Topic 3.5</p>
4.5	<p>Hardness of water.</p> <p>Effects of hardness.</p>	<ul style="list-style-type: none"> <li>- The cause of temporary and permanent hardness of water; methods of removing both types of hardness.</li> <li>- The formation of scum with soap.</li> <li>- Advantages of synthetic detergents over soap when used with hard water.</li> <li>- The formation of stalactites. Furring of kettles and scaling in water pipes.</li> <li>- Advantages and disadvantages of hard water.</li> </ul>	<p>Simple treatment - not to include structure of soap or synthetic detergent molecules.</p>

## Unit 4 Water

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
4.1	Occurrence. Physical properties. Impurities and their removal.	Describe, in outline, methods of purifying water supplies from naturally occurring impurities in terms of filtration, distillation, reverse osmosis and chlorination				
		Describe a chemical test for the presence of water				
		Recall that boiling point can be used to test for purity of water				
4.2	Solution and solubility. Water of crystallisation.	Recall and use the terms: solute, solvent, solution and saturated solution				
		Recall the general rules which describe the solubility of common types of salt in water				
		Appreciate that different substances have a different solubility in water and that solubility may vary with temperature				
		Recall that the solubility of gases decreases with increase of temperature				
		Use solubility rules to predict whether a specified salt is soluble or not				
		Understand the significance of and interpret simple, continuous solubility curves				
		Understand that, for some salts, the crystalline shape (and in some cases even the colour), is dependant on water combined in the crystals				
		Distinguish between the terms deliquescence, hygroscopy and efflorescence, using suitable examples				
4.3	Tests for water.	State a chemical test for the presence of water and a physical test for pure water				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
4.4	Chemical properties of water.	Describe the reactions of specific metals and non-metals with water or steam				
		State that metal oxides which dissolve in water produce alkaline solutions				
		State that non-metal oxides which dissolve in water produce acid solutions				
		Describe the reactions of common basic oxides and acidic oxides with water				
4.5	Hardness of water.	Define hard water as water that does not readily lather with soap but gives a scum				
		Recall that dissolved calcium and magnesium salts cause hardness				
		Understand that temporary hardness results from contact of rain water with limestone				
		Explain the chemical reaction involved in the formation of temporary hardness				
		Recall that boiling removes temporary but not permanent hardness				
		Explain the chemical reaction underlying the removal of temporary hardness by boiling				
		Relate the formation of stalactites / stalagmites to the presence of temporary hardness				
		Understand that the presence of temporary hardness leads to industrial and domestic problems caused by the deposition of scale				
		Recall that both types of hardness can be removed by ion exchange, by distillation and by the addition of sodium carbonate				
		Appreciate that hard water wastes soap but does not affect soapless detergents				
		Compare/contrast the advantages/disadvantages of soaps and synthetic detergents				
		Write balanced chemical equations for all reactions covered in this Unit				
		Name some of the uses of water in the home and in industry				

## Unit 5      Acids, bases and salts.

### Objectives

- (i) To define acidity and alkalinity in terms of  $\text{H}^+(\text{aq})$  and  $\text{OH}^-(\text{aq})$  respectively. Introducing pH numbers as a scale to denote a degree of acidity or alkalinity as shown by an indicator.
- (ii) To emphasise the need of following instructions carefully. To instil due regard for correct practical techniques, as well as the need for safety measures, when handling acidic/alkaline solutions, both in the laboratory and in the home environment.
- (iii) To help students familiarise themselves with the general reactions of acids through practical investigations.
- (iv) To distinguish between insoluble bases and alkalis, and hence their use in practical situations.
- (v) To show the properties of alkalis through practical work.
- (vi) To investigate the process of neutralisation; to introduce titration as a method of determining the point of neutralisation.  
To distinguish between normal salts and acid salts.
- (vii) To show how the properties of acids can be utilised in different practical methods for the preparation of salts.  
To help students plan their practical work by considering the various practical difficulties and/or chemical exceptions that could influence the choice of the best method suited to prepare a particular salt.
- (viii) To make students aware that many salts are of practical use in our everyday lives.
- (ix) To take advantage of the variety of chemical reactions in this Unit, in order to give students ample practice in expressing the changes by word equations, and also in writing and balancing stoichiometric equations to represent the chemical processes.
- (x) To prepare hydrogen gas and investigate its properties.  
To discuss the use of hydrogen as an alternative to fossil fuels.



Topic	Item	Description	Additional notes
5.2	<p>Bases and alkalis.</p> <p>Investigating the properties of alkalis.</p> <p>Some useful bases.</p>	<p>- A base as a metallic oxide or hydroxide that reacts with an acid to form a salt and water only.</p> <p>- An alkali as a soluble base giving <math>\text{OH}^-</math>(aq). Strong and weak alkalis in terms of the number of <math>\text{OH}^-</math> ions formed and the pH value.</p> <p>- Effect on indicators; reaction with acids; reaction with ammonium salts to liberate ammonia; knowledge that NaOH and KOH react with 'amphoteric' metals Al and Zn to liberate hydrogen.</p> <p>Magnesium hydroxide in Milk of magnesia; bleaches and alkaline cleaners; quicklime and slaked lime to counter acidity in soil.</p>	<p>Students should be able to recall the names and formulae of specific bases and alkalis.</p> <p>Students will be expected to give equations for reactions of alkalis with ammonium salts; but equations for reactions of caustic alkalis with metals and formulae of hydroxymetallates are not required.</p> <p>Both the common name and chemical name (or formula) will be provided in the examination paper.</p>
5.3	<p>Neutralisation.</p> <p>Normal salts and acid salts.</p>	<p>- Neutralisation as the formation of neutral molecules of water from <math>\text{H}^+</math>(aq) and <math>\text{OH}^-</math>(aq), with a salt being the other product of the reaction.</p> <p>The difference between normal salts and acid salts; equations for their formation.</p>	<p>The ionic equation for neutralisation should be given.</p> <p>To include names and formulae of specific examples.</p>



Topic	Item	Description	Additional notes
5.5	Hydrogen.	<ul style="list-style-type: none"> <li>- Laboratory preparation of hydrogen from the action of a dilute non-oxidising acid on a metal. [Method of collection over water and collection of dry hydrogen should both be known.]</li> <li>- Test for hydrogen.</li> <li>- Physical properties of hydrogen.</li> <li>- Chemical properties: combustion; reaction with a metal, e.g. Na; reaction with non-metals, e.g. Cl<sub>2</sub> and N<sub>2</sub>; reduction of metallic oxides e.g. CuO.</li> <li>- Uses of hydrogen. Hydrogen as a fuel: advantages and disadvantages.</li> </ul>	<p>Exemplified by the reaction of dilute hydrochloric acid or dilute sulfuric acid on magnesium, zinc or iron.</p> <p>The industrial preparation of hydrogen is not required.</p>
	Hydrides.	Simply the knowledge that hydrogen forms ionic hydrides with metals, e.g. NaH; covalent hydrides with non-metals, e.g. HCl, NH <sub>3</sub> .	Details of properties of ionic and covalent hydrides are not required.

**Unit 5 Acids, Bases and salts. Hydrogen.**

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
5.1	Acids.  Investigation of the properties of acids.	Define the terms acids in terms of the ions they contain or produce in aqueous solution				
		Realise that substances such as hydrogen chloride (gas), sulfuric acid (liquid) and citric acid (solid), only show acidic character when encountered in aqueous solution				
		Explain the difference between strong / weak acids in terms of the degree of ionisation using hydrochloric acid and ethanoic acid as typical examples				
		Define the basicity of an acid and know typical examples of such acids : monobasic, dibasic, etc.				
		Recall the names and formulae of common acids				
		Describe the characteristic properties of acids as reactions with metals, bases or alkalis, carbonates and sulfites				
		Recall the general patterns of the products formed when acids react with metals, bases/alkalis, carbonates and sulfites e.g. carbonate + acid → salt + water + carbon dioxide				
		Appreciate that there are limitations to these general patterns of reactions (e.g. some metals do not react, some acids never liberate hydrogen, some salts are insoluble and therefore will stop the reaction between an acid and a solid reagent)				
		Know a large scale use of common mineral acids (HCl, H <sub>2</sub> SO <sub>4</sub> and HNO <sub>3</sub> ) and that hydrochloric acid is present in digestive fluid				
		Define the terms alkalis in terms of the ions they contain in aqueous solution				
Know an example of the use of carbonic, citric and ethanoic acids in our everyday life						

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level			
				A	B	C	
5.2	Bases and alkalis.	State that bases (metal hydroxides) which dissolve in water produce alkaline solutions					
		Recall that bases are the oxides or hydroxides of metals that are neutralised by acids to form salts					
	Investigating alkalis.	Understand that bases (metal oxides) which dissolve in water, also react with the water to form soluble metal hydroxides (alkalis)					
		Define ‘strong and weak’ alkalis and know typical examples e.g. sodium hydroxide and aqueous ammonia respectively					
		Recall the names and formulae of common alkalis used in the laboratory					
		Describe the characteristic properties of alkalis as reactions with acids and with ammonium salts					
		Recall the general pattern of alkalis reacting with an ammonium salt i.e. ammonium salt + alkali → salt + water + ammonia					
		Be familiar with the names of common alkalis used in the home					
		Understand why calcium oxide and calcium hydroxide are used to neutralise soil acidity and that a salt is formed in the reaction					
		Indicators.		Be aware that acid / base indicators are substances (naturally occurring or prepared) which change colour according to the pH of the solution			
				Relate pH to acidity/alkalinity and recall the use of Universal indicator to determine pH			
				Identify solutions which have a pH value of less than 7 as acidic and those with a pH value of more than 7 as alkaline; recall that solutions with a pH of 7 are neutral			
	Recall the colour changes given by other common indicators with acids / alkalis: litmus, methyl orange and phenolphthalein						

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
5.3	Neutralisation. Salts and acid salts.	State that neutralisation is the reaction of acids with bases or alkalis				
		Write the reaction of hydrogen ions and hydroxide ions to form water				
		Understand that neutralisation moves the pH of an acid towards 7				
		State that in water and in a neutral solution, the concentration of $H^+_{(aq)}$ and $OH^-_{(aq)}$ ions is the same				
		Define and distinguish between normal salts and acid salts using suitable examples				
		Give everyday examples of neutralisation and uses of salts as specified in the syllabus				
5.4	Preparation of salts.	<ul style="list-style-type: none"> <li>understand that the method selected for preparing a specific salt depends on factors such as solubility of the salt, ease of the technique and safety considerations</li> </ul>				
		<ul style="list-style-type: none"> <li>predict a suitable method for preparing a given salt</li> </ul>				
		<ul style="list-style-type: none"> <li>outline a method of preparing a salt from suitable starting materials, given appropriate information or guidance</li> </ul>				
		<ul style="list-style-type: none"> <li>describe, in detail, the usual experimental procedures to prepare pure, dry samples of salts: <ul style="list-style-type: none"> <li>soluble salts by the reaction of an excess of a suitable solid (metal, base or carbonate) with an acid</li> <li>the use of titration apparatus simply as a means of determining the point of neutralisation in preparing salts of potassium, sodium and ammonium</li> <li>preparing an insoluble salt by precipitation</li> <li>the synthesis of iron (II) sulfide and iron (III) chloride in the laboratory</li> </ul> </li> </ul>				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
5.4 contd	Preparation of salts.	Understand the importance of reacting all the acid by using excess solid reactant, when preparing many soluble salts				
		Understand that insoluble salts can be formed as precipitates by the reaction of suitable reagents in solution				
		Appreciate that synthesis reactions are limited as a means of preparing salts in the laboratory				
5.5	Handling techniques for preparing gases.  Hydrogen.	Be aware that the apparatus used for the preparation of a gas usually consists of three sections – gas generator, drying and/or purification and collection				
		Understand that the method for drying and/or collecting a gas depends on its physical properties and on whether the gas is required dry or not				
		Draw correct, clear and well labelled diagrams for the preparation of a gas				
		Be able to identify faults in the set up for the preparation of a gas				
		Perform the preparation of a gas on a test-tube scale				
		Recall that hydrogen is liberated by the reaction of dilute hydrochloric and dilute sulfuric acids on the fairly reactive metals magnesium, zinc and iron				
		Describe the generation and collection of hydrogen over water or the preparation of dry hydrogen				
		Appreciate the potential danger of explosion when hydrogen is collected or handled				
		Describe a simple test for hydrogen				
		Recall that hydrogen is liberated from the reaction between reactive metals and water; fairly reactive metals and steam; an amphoteric metal and an alkali				
		Discuss the reactions of hydrogen with oxygen, chlorine and copper (II) oxide				
State the use of hydrogen in the manufacture of ammonia, of margarine and as a fuel						

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
	Chemical equations	<ul style="list-style-type: none"> <li>be familiar with the common types of chemical reactions – synthesis, thermal decomposition, combustion, neutralisation, precipitation</li> </ul>				
		<ul style="list-style-type: none"> <li>represent chemical reactions by word equations</li> </ul>				
		<ul style="list-style-type: none"> <li>name the products formed from the reaction between a given set of reagents</li> </ul>				
		<ul style="list-style-type: none"> <li>balance a given formula equation</li> </ul>				
		<ul style="list-style-type: none"> <li>complete and balance an equation, given the formulae of the reactants</li> </ul>				
		<ul style="list-style-type: none"> <li>identify and correct mistakes in given equations</li> </ul>				
		<ul style="list-style-type: none"> <li>deduce the balanced equation for a reaction, given relevant information</li> </ul>				
		<ul style="list-style-type: none"> <li>utilise the general patterns of reactions covered in this Unit to construct balanced equations to represent the chemical reactions</li> </ul>				
		<ul style="list-style-type: none"> <li>recall and apply in equations, where appropriate, the state symbols</li> </ul>				

## Unit 6      The mole concept. Gravimetric Analysis

### Objectives

- (i) To familiarise students with the idea that masses of atoms are based on a 'relative scale', in which the standard element chosen for comparison, carbon, has an arbitrarily assigned mass. To emphasise that this relative mass is not the actual mass of a single atom.
- (ii) To introduce the concept that a 'mole' of atoms of any element contains the same number of atoms, which is given by Avogadro's constant.
- (iii) Combining the concepts of 'relative atomic mass' and 'the mole of atoms', to establish that the molar mass of an element represents the mass in grammes of Avogadro's number of atoms of that element.
- (iv) To consider the mass of different elements that would have to be taken to get the same number of atoms of each element, leading to the concept of 'counting out a number of atoms by weighing out masses of elements'.
- (v) To apply these principles in calculations on mole/mass/number of atoms interconversions.
- (vi) To extend the idea of relative atomic mass, to include the notations 'relative molecular mass' or 'relative formula mass'.  
To emphasise that, since the mole represents a fixed number of 'particles', the ideas related to moles of atoms can also be applied to molecules and ions.
- (vii) To give students practice in numerical problems involving mole/mass/particles interconversions with respect to compounds.
- (viii) To show the relevance of the mole concept to practical situations - to determine experimentally the formulae of substances by converting reacting masses of elements into the number of 'moles of atoms' of the constituent elements.
- (ix) To give students practice in problems based on percentage composition.
- (x) To introduce mole/mass relationships in chemical equations.

**Unit 6      The Mole concept. Gravimetric Analysis**

Topic	Item	Description	Additional notes
6.1	<p>Relative atomic mass.</p> <p>The mole of atoms.</p> <p>Avogadro's constant.</p>	<p>- Relative atomic mass based on <math>^{12}\text{C} = 12</math>. The difference between 'mass number' and 'relative atomic mass' as exemplified by chlorine.</p> <p>- The mole as a constant number of particles The mass of one mole of atoms.</p> <p>- Practice on mole(s) / mass and mole(s) / atoms interconversion.</p>	Or vice-versa.
6.2	<p>Relative molecular mass or relative formula mass.</p> <p>Mole of molecules.</p> <p>Mole of ions.</p>	<p>- Relative molecular mass as the sum of the relative atomic masses. The mass of a mole of molecules as the relative molecular mass expressed in grammes.</p> <p>- Calculating the mass of one mole (or a fraction or multiple of moles) of molecular elements or covalent compounds. Practice on mole(s) of molecules / mass and mole(s) / number of molecules interconversions.</p> <p>- Calculating the mass of one mole (or fraction or multiple of moles) of an ionic compound. Practice on mole(s) / mass interconversions.</p>	<p>The term 'relative formula mass' will be used for ionic compounds.</p> <p>Numerical problems to include work on a fraction or multiple of moles of molecules.</p> <p>Calculations to find the number of ions will not be set in Form 3</p>

Topic	Item	Description	Additional notes
6.3	Applications of the mole concept.  Gravimetric Analysis.	<ul style="list-style-type: none"> <li>- Experimental determination of chemical formulae of binary compounds.</li> <li>- Using moles of atoms to determine formulae from reacting masses of elements.</li> <li>- Deriving the value of x in a hydrated compound, by heating the hydrate to constant mass.</li> </ul>	<p>Including experimental details and calculations from results.</p> <p>By direct combination, e.g. MgO, or by reduction using hydrogen, e.g. CuO.</p> <p>Suitable examples could be <math>\text{MgSO}_4 \cdot x\text{H}_2\text{O}</math> or <math>\text{ZnSO}_4 \cdot x\text{H}_2\text{O}</math>.</p>
6.4	Percentage composition by mass.	<ul style="list-style-type: none"> <li>- Calculating the percentage composition by mass of an element in a compound.</li> <li>- Calculating the percentage composition by mass of water of crystallisation in a hydrated salt.</li> <li>- Determining the empirical formula of a compound from the percentage composition by mass. Given the relative molecular mass, to determine the molecular formula, from the empirical formula.</li> <li>- Deriving the formula of a hydrated compound, given its percentage composition by mass.</li> </ul>	<p>(including the percentage of water of crystallisation).</p>
6.5	Mole/mass relationships in chemical equations.  Calculations on stoichiometric reacting masses.	<ul style="list-style-type: none"> <li>- Law of conservation of mass.</li> <li>- Using the mole ratios of substances as expressed in a balanced equation, to calculate the mass of a reactant/product required/or produced from a given mass of reactant/product.</li> <li>- Deriving a balanced chemical equation by converting the masses of reactants and/or products into mole ratios.</li> </ul>	<p>Molar volume and molar solutions are not to be included. Numerical problems involving mass to volume, or volume to volume relationships not to be covered in Form 3.</p> <p>Details of experimental investigations of chemical reactions in order to establish the combining ratios of reactants and/or products, will not be required from Form 3 students.</p>

**Unit 6**      **The mole concept. Gravimetric Analysis**

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
6.1	Relative atomic mass.	Recall that atomic masses are the masses of atoms relative to $^{12}\text{C} = 12$ and are referred to as the relative atomic masses				
	The mole of atoms.	Distinguish between mass numbers and relative atomic mass				
	Molar mass and Avogadro's constant.	Define the mole				
		Understand the term 'mole of atoms' as a number of atoms equal to the Avogadro constant				
		Convert moles of atoms into masses of elements and vice-versa				
		Convert moles of atoms into number of atoms and vice-versa				
6.2	Relative molecular mass and relative formula mass.	Recall that a mole is an amount of substance which can also be expressed as the Avogadro number of particles (molecules, formulae or ions)				
	Mole of molecules.	Recall that the relative molecular / formula mass of a substance is the sum of the relative atomic masses as shown in the chemical formula				
	Mole of ions.	Recall that the mass of one mole of molecules or formula can be expressed as the relative molecular / formula mass in grams				
		Work out the relative formula mass of a substance from its formula				
		Convert moles of molecules / formula into masses of substances and vice-versa				
		Convert moles of molecules into number of molecules and vice-versa				

Topic	Item	Learning Outcomes – at the end of these topics students should be able to:	Time	Difficulty level		
				A	B	C
6.3	Practical applications of the mole concept. Gravimetric analysis.	Understand how the idea of a mole of atoms can be applied in finding chemical formulae				
		Recall experiments to find the formulae of simple compounds such as magnesium oxide and copper (II) oxide				
		Determine the formulae of simple binary compounds from information provided about the reacting masses of elements and understand that such formulae are empirical				
		Distinguish between and define the terms empirical and molecular formulae				
		Recall the practical steps required when heating a hydrated compound to constant mass				
		Determine the number of molecule of water of crystallisation in a hydrated compound from information provided about heating the compound to constant mass				
		Appreciate that the formulae of other compounds can be obtained experimentally				
6.4	Percentage composition by mass.	Calculate the percentage composition by mass of elements in compounds				
		Calculate the percentage composition by mass of water of crystallisation in hydrated compounds				
		Derive the empirical and / or molecular formula, given the percentage composition by mass and the relative formula mass				
6.5	Mass relationships in chemical reactions.	Use mole ratios from chemical equations to calculate stoichiometric masses of reactants and products				