

E403 CHEMISTRY (YEAR 12) – 2008-2009

Rationale

Chemistry is concerned with the physical and chemical properties of substances, the nature of matter, and the study of chemical reactions and processes. It is an investigative subject, and through observation and experiment, a body of factual information has accumulated from which principles, laws and theories have been developed.

This subject is intended to provide students with a balanced perspective of chemical science which will enable them to understand and interpret the chemistry of their surroundings and appreciate the impact of chemical knowledge and technology on society. It contains material which is relevant to terminal chemistry students and students who intend to study chemistry at the tertiary level.

The subject attempts to provide a balance between useful facts, important concepts and theories. Concepts and theories are introduced to facilitate an understanding of the properties of substances, reactions and processes.

Where appropriate, the physical and chemical properties of substances are related to important uses. As well, some emphasis is placed on industrial processes of major local or world-wide significance and their impact on our society.

Chemistry is an experimental science, and laboratory work is an essential part of the syllabus. This should be used to enable students to investigate the properties and reactions of substances and to provide opportunities for the learning and testing of concepts and principles.

It is considered that a thorough knowledge of the Year 11 Chemistry syllabus or its equivalent is an essential requirement for the successful study of this subject.

Educational Objectives

The learning outcomes set out below represent the objectives of the subject.

Knowledge, understandings and intellectual skills

Students should be able to:

- state and apply fundamental facts and principles of chemistry dealing with
 - the nature and properties of matter
 - the physical and chemical properties and related uses of chemical substances

- changes which occur in chemical reactions
- important industrial chemical processes
- use appropriate conceptual and theoretical frameworks to explain relationships between the structure of matter and its properties and changes which occur in chemical reactions
- perform chemical calculations using the mole as a unifying concept
- demonstrate competence in the process skills of science i.e. observing, inferring, classifying, communicating, predicting, measuring, formulating and testing hypotheses, controlling variables, interpreting data, designing and performing experiments.

Manipulative skills

Students should be able to:

- safely and confidently handle chemicals and recognise and respect those which are dangerous
- use a range of manipulative skills associated with laboratory work.

Affective outcomes

Students should be able to:

- demonstrate an interest in chemistry and an appreciation of the impact of chemistry on the society in which they live
- recognise the investigative nature of chemistry and identify virtues of the scientific method
- describe chemistry as a developing science which is the product of the work and thought of generations of scientists
- explain the need to possess evidence before making decisions and exhibit a willingness to accept evidence contrary to established theories.

Teaching – Learning Program

The topics, or objectives within topics, can be taught in any order in keeping with the needs of teachers and students.

In the syllabus statement the content has been divided into seven main areas. This is not in any way intended to represent a teaching sequence but is an attempt to organise the content in a logical way. Teachers should adopt whatever sequence they feel is most appropriate for their students.

Subject objectives

1. Macroscopic properties of matter

Solutions and solubility

Cognitive objectives:

- 1.1 Define 'concentration' as a quantity of solute dissolved in a fixed quantity of solution.
- 1.2 Identify the following as units for expressing the concentration of a solution:
 - moles per litre (mol L^{-1})
 - grams per litre (g L^{-1})
 - parts per million (p.p.m.)
 - percentage composition.
- 1.3 Perform calculations relating the mass of solute and the mass of solution to the concentration of a solution in
 - a) p.p.m. (mg of solute per kg of solution)
 - b) percentage composition (g of solute per 100 g of solution).
- 1.4 Perform calculations relating the mass or number of moles of solute and volume of solution to the concentration of a solution in
 - a) g L^{-1}
 - b) mol L^{-1}
- 1.5 Solve problems involving mass, molar mass, number of moles of solute, and concentration and volume of solution.
- 1.6 Perform calculations of concentration and volume involved in dilution of solutions.
- 1.7 Calculate the concentrations of ions in a solution of a strong electrolyte of known concentration, or in a solution resulting from the mixture of two solutions of strong electrolytes.
- 1.8 Apply the solubility rules (Table 5) for the nitrates, bromides, chlorides, iodides, sulfates, sulfides, hydroxides, phosphates and carbonates of the following positive ions (cations): aluminium, ammonium, barium, cadmium, calcium, chromium(III), cobalt, copper(II), iron(II), iron(III), lead(II), magnesium, manganese, mercury(I), mercury(II), nickel, potassium, silver, sodium, strontium, tin(II), zinc.
- 1.9 Predict if precipitation will occur when two dilute (approximately 0.2 mol L^{-1}) solutions of different ionic compounds are mixed, and write equations.

2. Atomic structure and bonding

Atomic structure

Cognitive objectives:

- 2.1 Outline the structure of the atom in terms of its constituent protons, neutrons and electrons.

- 2.2 Describe the electron configuration of atoms in terms of shells and subshells (principal energy levels and energy sublevels).
- 2.3 Use the s, p, d, f notation to identify subshells.
- 2.4 Describe orbitals as regions of space within subshells.
- 2.5 Apply that the order of subshell filling in any shell is s before p before d before f.
- 2.6 Write the ground state electron configuration (shell and subshell notation) of atoms up to $Z = 20$.
- 2.7 Identify the outer shell electrons in any atom as the valence electrons.
- 2.8 Relate the electron configuration of any element to its position in the periodic table and determine, using a periodic table, the number of valence electrons in the main group elements.
- 2.9 Apply that the number of valence electrons in the atoms of an element is the most important factor in determining the chemical properties of the element.
- 2.10 Identify the s, p, d, and f blocks in the periodic table.
- 2.11 State that noble (inert) gases have the outer shell configuration s^2p^6 (except helium, s^2).
- 2.12 Infer the stability of the noble gas electron configuration from the lack of reactivity of the noble gases, and apply this concept to explain the tendency of atoms to attain the electron configuration of the nearest noble gas in the periodic table.
- 2.13 Explain the formation of ions and molecules in terms of the tendency of atoms to attain the electron configuration of the nearest noble gas in the periodic table.
- 2.14 State and apply that valency is a measure of the bonding capacity of an atom.
- 2.15 Relate the valency of a main group element to its position in the periodic table.

Bonding

Cognitive objectives:

- 2.16 Define 'ionisation energy' as the energy required to remove a mole of electrons from a mole of atoms or ions in the gas phase.
- 2.17 State and explain the variation in the first ionisation energy
 - a) down a group
 - b) across a period in the periodic table.
- 2.18 Describe and explain the variation in successive ionisation energies of an element in terms of
 - a) increasing positive charge on the ion
 - b) the number of valence electrons and the presence of shells in the atom.

- 2.19 Determine the number of valence electrons of an element from the magnitudes of successive ionisation energies.
- 2.20 State that ions can result from the transfer of electrons from atoms of low ionisation energy to atoms of high ionisation energy.
- 2.21 Describe an ionic solid as a lattice of positive and negative ions and explain the existence of such an ionic lattice in terms of the electrostatic attraction between oppositely charged ions (ionic bonding).
- 2.22 Identify that, and explain why, metals from groups I and II are most likely to lose electrons and become positive ions (cations), while non-metals from groups VI and VII are most likely to gain electrons and become negative ions (anions) .
- 2.23 Write electron configurations for monatomic ions (up to $Z = 20$).
- 2.24 Write the symbols and names of the monatomic ions as shown in Table 1.
- 2.25 Write the formulae of the polyatomic ions as shown in Table 2, and the names of the polyatomic ions not asterisked.
- 2.26 State that groups I and II monatomic positive ions and groups VI and VII monatomic negative ions are colourless in aqueous solution.
- 2.27 Use the colours of the monatomic positive ions.
 $\text{Cr}^{3+}(\text{aq})$, $\text{Mn}^{2+}(\text{aq})$, $\text{Fe}^{2+}(\text{aq})$, $\text{Fe}^{3+}(\text{aq})$,
 $\text{Co}^{2+}(\text{aq})$, $\text{Ni}^{2+}(\text{aq})$, $\text{Cu}^{2+}(\text{aq})$, $\text{Cd}^{2+}(\text{aq})$,
 $\text{Zn}^{2+}(\text{aq})$, $\text{Ag}^{+}(\text{aq})$, $\text{Hg}_2^{2+}(\text{aq})$, $\text{Hg}^{2+}(\text{aq})$.
 of the transition metals in Table 4 to infer and describe the products of reactions.
- 2.28 Use the colours of the polyatomic positive and negative ions
 $\text{CrO}_4^{2-}(\text{aq})$, $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$, $\text{MnO}_4^{-}(\text{aq})$,
 $*[\text{Cu}(\text{NH}_3)_4]^{2+}(\text{aq})$,
 $*[\text{Zn}(\text{OH})_4]^{2-}(\text{aq})$, $*[\text{Zn}(\text{NH}_3)_4]^{2+}(\text{aq})$,
 $*[\text{Ag}(\text{NH}_3)_2]^{+}(\text{aq})$.
 of the transition metals to infer and describe the products of reactions.
- 2.29 Write the formulae of compounds formed from the ions listed above in 2.24 to 2.28, given the names of the compounds, and name the compounds given the formulae (except where asterisked).
- 2.30 State that polyatomic ions exhibit covalent bonding between the atoms within the ion.
- 2.31 Draw electron dot diagrams of simple ionic compounds (e.g. NaCl and CaBr_2) and of common polyatomic ions.
- 2.32 Use the nature of the ionic lattice structure and the attraction between the oppositely charged ions to explain physical properties of ionic lattices such as:
- high melting and boiling temperatures
 - electrical conductivity of the molten state and aqueous solutions
 - brittleness.
- 2.33 Explain the transfer of electrons in the formation of ionic substances and the sharing of electrons in the formation of covalent substances in terms of atoms attaining noble gas electron configurations.
- 2.34 Describe the electrostatic nature of the covalent bond in terms of attraction of both nuclei for the shared electron pair(s).
- 2.35 State and apply that covalent bonding is restricted to atoms having similar (high) electronegativities and that such elements are generally non-metals.
- 2.36 Explain the formation of single and multiple covalent bonds in simple molecules of elements (e.g. F_2 , O_2 , N_2), relating this to the electron configurations of the atoms involved.
- 2.37 Explain that single and multiple covalent bonds may also exist in compounds, and cite examples of such compounds.
- 2.38 State and apply that the number of covalent bonds formed by a particular atom can be related to its position in the periodic table.
- 2.39 Write the formulae and names of the molecular substances as shown in Table 3.
- 2.40 State the colours of the substances listed in 2.39, and the colours of chlorine, bromine and iodine in aqueous solution and in chlorinated hydrocarbons.
- 2.41 Draw appropriate electron dot diagrams for simple molecular substances.
- 2.42 Define the electronegativity of an atom in terms of its ability to attract electrons within a covalent bond.
- 2.43 Outline the trends in electronegativity across a period and down a group in the periodic table.
- 2.44 Relate the polarity of a covalent bond to the difference in electronegativity of the atoms involved in bond formation, leading to the unequal sharing of electrons in the bond.
- 2.45 Explain that there is a continuum from pure covalent bonding, through polar covalent bonding, to ionic bonding.
- 2.46 Describe the valence shell electron pair repulsion theory and use this theory to predict the shapes of simple molecules (limited to a maximum of four regions of charge around each atom e.g. NH_3 , H_2O , CO_2 , CH_4 , HCl , C_2H_4 , C_2H_2 , etc.).
- 2.47 State and apply that the conditions necessary for molecular polarity are
- bond polarity
 - molecular asymmetry.
- 2.48 Predict the molecular polarity of simple covalent molecular substances.
- 2.49 Distinguish between covalent molecular substances and covalent network substances.

- 2.50 State that covalent network solids consist of arrays of atoms covalently bonded to one another, citing examples of such solids – elements (e.g. silicon and carbon as in graphite and diamond) and compounds (e.g. silicon carbide and silica), and recognise structural diagrams of these solids.
- 2.51 Use the structure of covalent network solids and the existence of strong covalent bonds between atoms to explain properties such as:
- high melting and boiling points
 - hardness
 - lack of electrical conductivity
 - the unusual electrical conductivity and lubricating properties of graphite.
- 2.52 Describe the electrostatic nature of metallic bonding in terms of the attraction of metal ions for delocalised valence electrons.
- 2.53 Describe a metal lattice as an array of positive metal ions interspersed with delocalised valence electrons.
- 2.54 Use the nature of the metallic lattice and particularly the presence of delocalised electrons in the metallic lattice to explain properties such as:
- wide range of melting and boiling points
 - high thermal and electrical conductivity
 - malleability and ductility.
- 2.55 Distinguish between intramolecular covalent bonding and the weaker intermolecular van der Waals' forces.
- 2.56 State that van der Waals' forces may include dispersion forces (induced dipole interactions), dipole-dipole interactions and hydrogen bonding.
- 2.57 State and apply that van der Waals' forces are significant attractive forces between molecules in the solid and liquid states.
- 2.58 Use a simple model to explain the existence of dispersion forces between atoms and molecules.
- 2.59 State and apply that dispersion forces will be present in all substances, but will not be significant in ionic, metallic or covalent network solids because of the relative weakness of dispersion forces compared with the strength of ionic, metallic and covalent bonding.
- 2.60 Describe and apply the relationship between the strength of dispersion forces and molecular shape and molecular size.
- 2.61 State that dipole-dipole forces result from the electrostatic attraction between polar molecules.
- 2.62 Identify nitrogen, oxygen and fluorine as highly electronegative elements and state that polar molecules involving these elements are likely to exhibit strong dipole-dipole attractions.
- 2.63 Identify hydrogen bonding as an extreme form of dipole-dipole interaction, and state that hydrogen bonding involves the attraction between the hydrogen atom attached to a nitrogen, oxygen or fluorine atom in one molecule and a nitrogen, oxygen or fluorine atom in another molecule.
- 2.64 State and apply that, for molecules of similar size, hydrogen bonding is generally regarded as the strongest of the van der Waals' forces, but that it is much weaker than ionic, covalent or metallic bonding.
- 2.65 Describe the trend in melting and boiling points of the hydrides of groups V, VI, and VII, indicating the reasons for these trends, and explaining the anomalous behaviour of NH_3 , H_2O and HF in terms of hydrogen bonding.
- 2.66 Compare the melting and boiling points of CH_4 with those of NH_3 , H_2O and HF and explain the differences.
- 2.67 State that for molecules having comparable shape and molar mass, generally hydrogen bonding is stronger than dipole-dipole attraction, which in turn is stronger than dispersion forces.
- 2.68 State that covalent molecular solids consist of arrays of molecules (containing covalent intramolecular bonds) with van der Waals' intermolecular forces.
- 2.69 Use the structure of covalent molecular solids and the existence of van der Waals' forces between molecules to explain properties such as:
- low melting and boiling points
 - lack of electrical conductivity in the solid or molten states.
- 2.70 Distinguish between polar and non-polar solvents.
- 2.71 Classify solutes as ionic, polar or non-polar.
- 2.72 Propose simple explanations and predict (in terms of interactions between solute particles, between solvent molecules, and between solute particles and solvent molecules):
- the general solubility of ionic and polar solutes in polar solvents
 - the general insolubility of non-polar solutes in polar solvents, and of polar solutes in non-polar solvents
 - the general solubility of non-polar solutes in non-polar solvents.
- 2.73 Explain (in terms of the interactions included in Cognitive objective 2.72) the behaviour of water as a solvent and its ability to dissolve ionic, polar and non-polar solutes.
- Laboratory work objectives:**
- 2.L.1 Demonstrate and relate the electrical conductivity of some liquids and solutions to the types of chemical substances present.

Table 1**Cognitive objective 2.24: Symbols and names of monatomic ions.**

1+		2+		3+		4+	
hydrogen	H ⁺	magnesium	Mg ²⁺	aluminium	Al ³⁺	tin(IV)	Sn ⁴⁺
lithium	Li ⁺	calcium	Ca ²⁺	iron(III)	Fe ³⁺	lead(IV)	Pb ⁴⁺
sodium	Na ⁺	barium	Ba ²⁺	chromium(III)	Cr ³⁺		
potassium	K ⁺	manganese(II)	Mn ²⁺	gold(III)	Au ³⁺		
silver	Ag ⁺	iron(II)	Fe ²⁺				
copper(I)	Cu ⁺	copper(II)	Cu ²⁺				
gold(I)	Au ⁺	zinc	Zn ²⁺				
		mercury(II)	Hg ²⁺				
		tin(II)	Sn ²⁺				
		lead(II)	Pb ²⁺				
		strontium	Sr ²⁺				
		nickel(II)	Ni ²⁺				
		cobalt(II)	Co ²⁺				
		cadmium(II)	Cd ²⁺				
1-		2-		3-			
hydride	H ⁻	oxide	O ²⁻	nitride	N ³⁻		
fluoride	F ⁻	sulfide	S ²⁻				
chloride	Cl ⁻						
bromide	Br ⁻						
iodide	I ⁻						

Table 2**Cognitive objective 2.25: Formulae and names of polyatomic ions.**

1-		2-		3-	
hydroxide	OH ⁻	Carbonate	CO ₃ ²⁻	phosphate	PO ₄ ³⁻
nitrate	NO ₃ ⁻	Sulfate	SO ₄ ²⁻		
nitrite	NO ₂ ⁻	Sulfite	SO ₃ ²⁻		
hydrogencarbonate	HCO ₃ ⁻	Dichromate	Cr ₂ O ₇ ²⁻		
hydrogensulfate	HSO ₄ ⁻	Chromate	CrO ₄ ²⁻		
ethanoate (acetate)	CH ₃ COO ⁻	Hydrogenphosphate	HPO ₄ ²⁻		
permanganate	MnO ₄ ⁻	Oxalate	C ₂ O ₄ ²⁻		
cyanide	CN ⁻	peroxide	O ₂ ²⁻		
dihydrogenphosphate	H ₂ PO ₄ ⁻	*tetrahydroxozincate	[Zn(OH) ₄] ²⁻		
*perchlorate	ClO ₄ ⁻				
*chlorate	ClO ₃ ⁻				
*chlorite	ClO ₂ ⁻				
*hypochlorite	ClO ⁻				
*tetrahydroaluminate	[Al(OH) ₄] ⁻				
*tetrahydrochromate(III)	[Cr(OH) ₄] ⁻				
)					
*dicyanoaurate(I)	[Au(CN) ₂] ⁻				
1+		2+			
Ammonium	NH ₄ ⁺	mercury(I)	Hg ₂ ²⁺		
*diamminesilver(I)	[Ag(NH ₃) ₂] ⁺	*tetraamminecopper(II)	[Cu(NH ₃) ₄] ²⁺		
		*tetraamminezinc	[Zn(NH ₃) ₄] ²⁺		

* These names do not need to be learned.

Table 3**Cognitive objective 2.39: Formulae and names of molecular substances**

Elements		Compounds	
hydrogen	H ₂	carbon monoxide	CO
nitrogen	N ₂	carbon dioxide	CO ₂
oxygen	O ₂	sulfur dioxide	SO ₂
fluorine	F ₂	sulfur trioxide	SO ₃
chlorine	Cl ₂	water	H ₂ O
bromine	Br ₂	ammonia	NH ₃
iodine	I ₂	hydrogen sulfide	H ₂ S
phosphorus	P ₄	hydrogen peroxide	H ₂ O ₂
sulfur	S ₈	hydrogen fluoride	HF
		hydrogen chloride	HCl
		hydrogen bromide	HBr
		hydrogen iodide	HI
		nitrogen monoxide	NO (nitric oxide)
		nitrogen dioxide	NO ₂
		dinitrogen monoxide	N ₂ O (nitrous oxide)
		dinitrogen tetroxide	N ₂ O ₄
		nitric acid	HNO ₃
		phosphoric acid	H ₃ PO ₄
		sulfurous acid	H ₂ SO ₃
		sulfuric acid	H ₂ SO ₄
		*hypochlorous acid	HClO
		organic compounds as in 6.4	

*This does not need to be learned.

Table 4**Cognitive objective 2.27: Colours of aqueous ions of selected transition elements**

Cr ³⁺ _(aq)	deep green	Cd ²⁺ _(aq)	colourless
Mn ²⁺ _(aq)	very pale pink	Zn ²⁺ _(aq)	colourless
Fe ²⁺ _(aq)	pale green	Ag ⁺ _(aq)	colourless
Fe ³⁺ _(aq)	brown	Hg ₂ ²⁺ _(aq)	colourless
Ni ²⁺ _(aq)	green	Hg ²⁺ _(aq)	colourless
Co ²⁺ _(aq)	pink	[Zn(OH) ₄] ²⁻	colourless
Cu ²⁺ _(aq)	blue	[Zn(NH ₃) ₄] ²⁺	colourless
[Cu(NH ₃) ₄] ²⁺ _(aq)	deep blue	[Ag(NH ₃) ₂]	colourless
CrO ₄ ²⁻ _(aq)	yellow		
Cr ₂ O ₇ ²⁻ _(aq)	orange		
MnO ₄ ⁻ _(aq)	purple		

Note:

- It is acceptable if a student infers the colour of the solid from the colour of the constituent ions.
- It is acceptable if a student assumes that a white solid forms from colourless ions.

Table 5

Cognitive objective 1.8: Solubility rules for ionic solids in water**Soluble in water**

Soluble	Exceptions	
	Insoluble	Slightly soluble
Most chlorides	AgCl, Hg ₂ Cl ₂	PbCl ₂
Most bromides	AgBr, Hg ₂ Br ₂ , HgBr ₂	PbBr ₂
Most iodides	AgI, Hg ₂ I ₂ , HgI ₂ , PbI ₂	
All nitrates	Nil	
Most sulfates	SrSO ₄ , BaSO ₄ , HgSO ₄ , PbSO ₄	CaSO ₄ , Ag ₂ SO ₄

Insoluble in water

Insoluble	Exceptions	
	Soluble	Slightly soluble
Most hydroxides	NaOH, KOH, Ba(OH) ₂ (NH ₄ OH does not exist)	Ca(OH) ₂ , Sr(OH) ₂
Most carbonates	Na ₂ CO ₃ , K ₂ CO ₃ , (NH ₄) ₂ CO ₃	
Most phosphates	Na ₃ PO ₄ , K ₃ PO ₄ , (NH ₄) ₃ PO ₄	
Most sulfides	Na ₂ S, K ₂ S, (NH ₄) ₂ S	

Soluble	=	more than 0.1 mole dissolves per litre
Slightly soluble	=	between 0.01 and 0.1 mole dissolves per litre
Insoluble	=	less than 0.01 mole dissolves per litre

3. Chemical reactions**Reactions, equations and stoichiometry****Cognitive objectives:**

- 3.1 Describe chemical reactions as processes in which substances called reactants are changed to different substances called products.
- 3.2 Represent chemical reactions by equations which feature only those species consumed in the reaction and the new species produced, whether they be **ions** [for example Ag⁺(aq)], **molecules** [for example NH₃(g), NH₃(aq), CH₃COOH(l), CH₃COOH(aq)], or **solids** [for example BaSO₄(s), Cu(s), Na₂CO₃(s)].
- 3.3 Identify and use (s), (l), (g), and (aq), representing solid, liquid, gas and aqueous solution respectively, in chemical equations.
- 3.4 State and apply that mass, atoms and charge are conserved during a chemical reaction.
- 3.5 Balance chemical equations for numbers of atoms and charge by using coefficients in front of the formulae in the equation.
- 3.6 State and apply that the relationship between the numbers of individual particles and also between the numbers of moles of particles in a reaction is shown by a balanced chemical equation.
- 3.7 Derive chemical equations from information about the molar ratios of reacting substances, or from information that would allow these molar ratios to be determined.
- 3.8 Perform the following stoichiometric calculations using the mole as a unifying concept:
- mole/mole calculations
 - mass/mass calculations (e.g. gravimetric analysis, industrial processes)
 - mass/gas volume* calculations
 - gas volume*/gas volume* calculations
 - mass/solution volume and concentration calculations (e.g. precipitation reactions)
 - gas volume*/solution volume and concentration calculations
 - solution volume and concentration/solution volume and concentration calculations (e.g. acid-base and redox titrations).
- ***Note:** Gas volume calculations may involve the conversion of volumes of gases from S.T.P. to non-S.T.P. conditions and vice versa *as part of a longer calculation.*
- 3.9 Apply the concept of limiting reagent in calculations in which non-stoichiometric quantities of reactants are present.

Reaction rates

Cognitive objectives:

- 3.10 Describe the effect of each of the following factors on the rates of chemical reactions:
- nature of reactants
 - concentration of reactants
 - state of subdivision of reactants
 - temperature
 - catalyst.
- 3.11 Describe the transition state (activated complex) in a reaction as the highest energy state for the reacting system which:
- corresponds to some stage in the reaction at which bond-breaking and bond-formation is taking place
 - is unstable, having no more than a temporary existence, and explain its significance.
- 3.12 Define the activation energy of a reaction and explain its significance to the reaction rate.
- 3.13 Draw and interpret potential energy diagrams for exothermic and endothermic processes showing the transition state, activation energy and heat of reaction.
- 3.14 Explain the Collision Theory of reaction rates identifying that as prerequisites for any successful reaction:
- the reacting particles must have an appropriate collision orientation, and
 - the reacting particles must collide with sufficient energy.
- 3.15 Use the Collision Theory and a potential energy diagram or kinetic energy distribution curve, where appropriate, to explain the effect of the following factors on rates of chemical reactions:
- nature of the reactants
 - concentration of reactants
 - state of subdivision of reactants
 - temperature
 - catalyst.
- 3.16 Describe catalysts as agents which increase the rate of chemical reactions by providing an alternative reaction pathway with a lower activation energy and are not permanently consumed in the reaction.

Laboratory work objectives:

- 3.L.1 Demonstrate the effects of concentration, temperature, and catalysts on the rate of a chemical reaction.

Chemical equilibrium

Cognitive objectives:

- 3.17 State that the characteristics of chemical equilibrium in a closed system at constant temperature are:
- a constancy of macroscopic properties, in particular that the concentrations of all species are constant
 - a dynamic condition in which the rates of the forward and reverse reactions are equal.
- 3.18 Identify from a qualitative description whether a particular chemical system is at equilibrium.

- 3.19 Write the equilibrium law expression for a reaction at equilibrium. No calculations are required.
- 3.20 Explain that the equilibrium constant (K) for a chemical reaction
- indicates the relative proportions of products to reactants at equilibrium,
 - is a constant for that reaction at a given temperature
 - provides no information about the rate of a particular reaction.
- 3.21 Predict the effect on the relative proportions of products to reactants at equilibrium of the following changes:
- temperature
 - concentration (or partial pressure of a gas in the mixture)
 - volume (or partial pressure of all gases in the mixture).
- 3.22 State and apply that changes in the mass of a solid, the volume of a liquid and the presence of a catalyst have no effect on the relative proportions of products to reactants at equilibrium.
- 3.23 Explain how industrial processes such as the Haber process and Contact process often involve a compromise of rate, equilibrium yield and economic considerations.

Laboratory work objectives:

- 3.L.2 Demonstrate Le Châtelier's principle in some simple chemical systems.

4. Acids and bases in aqueous solutions

Cognitive objectives:

- 4.1 Classify water as a weak electrolyte in terms of its partial ionisation into $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$.
- 4.2 State the expression for the ionisation constant for water, $K_w = [\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]$, which equals 1.0×10^{-14} at 298 K.
- 4.3 State that the concentrations of $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ in pure water are equal to $1.0 \times 10^{-7} \text{ mol L}^{-1}$ at 298 K.
- 4.4 Use the Arrhenius and Bronsted-Lowry acid-base models to describe acid-base behaviour.
- 4.5 Identify in chemical equations the reactants which are acting as acids or bases.
- 4.6 Write equations to represent:
- acids as sources of $\text{H}^+(\text{aq})$ and bases as sources of $\text{OH}^-(\text{aq})$ in aqueous solutions
 - acids as proton donors and bases as proton acceptors.
- 4.7 Define amphoteric metals, oxides and hydroxides as those which react with both acids and bases.
- 4.8 Identify chromium, zinc and aluminium as amphoteric metals forming amphoteric oxides and hydroxides and write equations for the reactions of the oxides and hydroxides of chromium (III), zinc and aluminium with $\text{H}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$.

- 4.9 State and apply the relationship between pH and acidity and alkalinity in aqueous solutions.
- 4.10 Define 'pH' as $-\log_{10}[\text{H}^+(\text{aq})]$ and calculate the pH of strong acid solutions and strong base solutions.
- 4.11 Calculate the $[\text{H}^+(\text{aq})]$, $[\text{OH}^-(\text{aq})]$ and pH when solutions are mixed.
- 4.12 Identify strong acids including HCl, H_2SO_4 , HNO_3 and strong bases including NaOH, KOH, $\text{Ca}(\text{OH})_2$.
- 4.13 Identify weak acids including H_3PO_4 , CH_3COOH and other organic acids, and weak bases including NH_3 , Na_2CO_3 .
- 4.14 Distinguish between the following terms when referring to acids and bases:
- 'strong' and 'weak'
 - 'concentrated' and 'dilute'.
- 4.15 Define 'neutralisation' as a reaction of an acid with a base.
- 4.16 Describe the formation of salts by neutralisation reactions between acids and bases.
- 4.17 Use the concept of hydrolysis to explain and predict the acidic, basic or neutral nature of aqueous solutions of salts derived from the reaction of:
- strong acids and weak bases (e.g. ammonium chloride)
 - strong bases and weak acids (e.g.. sodium carbonate)
 - strong acids and strong bases (e.g. sodium chloride).
- 4.18 Describe the characteristics of a primary standard.
- 4.19 Define a 'standard solution' as one with an accurately known concentration.
- 4.20 Describe how to prepare a standard solution by:
- dissolving a known mass of a primary standard in a known volume of solution (e.g. anhydrous sodium carbonate)
 - standardisation of the solution (e.g. sodium hydroxide) by titration against a standard solution (e.g. hydrochloric acid).
- 4.21 Describe an indicator as a substance which exhibits two pH-dependent colour forms with particular reference to methyl orange and phenolphthalein.
- 4.22 Distinguish between the equivalence point and the end point in a titration.
- 4.23 Explain the use of phenolphthalein indicator in titrations between a strong base and a weak acid and methyl orange indicator in titrations between a strong acid and a weak base.
- 4.24 State that either phenolphthalein or methyl orange could be used as the indicator in titrations between strong acids and strong bases.
- 4.25 Describe how to use a pipette, burette and volumetric flask in volumetric analysis.
- 4.26 Describe the procedures used in volumetric analysis with particular reference to the acid-base titrations described in 4.L.5.
- Laboratory work objectives:**
- 4.L.1 Demonstrate the common properties of acids and bases.
- 4.L.2 Use indicators to determine hydrogen ion concentration and pH.
- 4.L.3 Prepare a standard solution by dissolving a known mass of a primary standard in a known volume of solution (e.g. anhydrous sodium carbonate).
- 4.L.4 Use a pipette, burette and volumetric flask in volumetric analysis.
- 4.L.5 Carry out the procedures used in volumetric analysis with particular reference to the following acid-base titrations:
- standardisation of hydrochloric acid with solid anhydrous sodium carbonate
 - standardisation of sodium hydroxide solution with standard acid solution
 - analysis of common substances such as vinegar or battery acid by titration with standard sodium hydroxide solution.

5. Oxidation and reduction

Redox reactions

Cognitive objectives:

- 5.1 Describe oxidation and reduction in terms of electron donation and acceptance using simple displacement reactions between metals and metal ions, halide ions and halogens as examples.
- 5.2 Identify the term 'redox' as one which emphasises the interdependence of oxidation and reduction processes.
- 5.3 Apply the terms 'oxidising agents (oxidants)' and 'reducing agents (reductants)' in terms of a substance's tendency to accept or donate electrons.
- 5.4 Identify common oxidising agents (oxidants) including O_2 , Cl_2 , MnO_4^- , $\text{Cr}_2\text{O}_7^{2-}$, ClO^- , H^+ , conc. sulfuric acid, conc. nitric acid and common reducing agents (reductants) including Zn, C, H_2 , Fe^{2+} , $\text{C}_2\text{O}_4^{2-}$.
- 5.5 Define and assign oxidation numbers (oxidation states).
- 5.6 Identify the species oxidised and reduced from given chemical equations.
- 5.7 Write and balance oxidation-reduction equations and half-equations using one of the following methods:
- the oxidation number method
 - the ion-electron method.
- 5.8 Use appropriate half-equations to illustrate that hydrogen peroxide can undergo both oxidation and reduction.
- 5.9 Identify disproportionation reactions such as the decomposition of hydrogen peroxide
- 5.10 Describe the use of hypochlorite ion in bleaching and in water treatment.

Electrochemical cells

Cognitive objectives:

- 5.11 Describe an electrochemical cell as a system for transforming the chemical potential energy of an oxidation-reduction reaction to electrical energy.
- 5.12 Define the 'anode' as the electrode at which oxidation occurs and the 'cathode' as the electrode at which reduction occurs.
- 5.13 Describe the components and operation of an electrochemical cell with regard to:
- anode processes
 - cathode processes

- c) electrolyte
 - d) salt bridge and ion migration
 - e) electron flow in the external circuit.
- 5.14 Define 'electromotive force (EMF)' of a cell.
- 5.15 Describe and explain how an electrochemical cell can be considered as two half-cells.
- 5.16 Define standard reduction potentials for half-cells.
- 5.17 Explain the use of the hydrogen half-cell as the standard for determining half-cell reduction potentials.
- 5.18 State that standard reduction potential (E°) values are measured at solution concentrations of 1 mol L^{-1} at 298 K and gas pressure of 101.3 kPa.
- 5.19 Use standard reduction potential (E°) tables to:
- a) predict whether reactions will occur
 - b) calculate standard cell EMFs.
- 5.20 State the limitations of the use of standard reduction potential (E°) tables.
- 5.21 Describe the construction and operation of the following electrochemical cells:
- a) the dry cell
 - b) the lead-acid accumulator
 - c) the fuel cell e.g. hydrogen/oxygen fuel cell.
- 5.22 Describe the conditions and reactions for the corrosion of iron.
- 5.23 Write the chemical equations and half-equations for the corrosion of iron and the formation of rust.
- 5.24 Describe methods for the prevention of corrosion of iron using:
- a) cathodic protection
 - b) sacrificial anodes e.g. use of magnesium or zinc
 - c) exclusion of air e.g. painting and plating.

Laboratory work objectives:

- 5.L.1 Demonstrate and identify some typical oxidation-reduction reactions.
- 5.L.2 Construct some simple electrochemical cells and measure their cell emfs.
- 5.L.3 Identify the factors affecting the corrosion of iron.

Electrolysis**Cognitive objectives:**

- 5.25 Describe electrolysis as a process in which electrical energy is used to produce chemical change.
- 5.26 Describe an electrolytic cell as a system for transforming electrical energy to chemical potential energy.
- 5.27 Describe the components and operation of an electrolytic cell with regard to:
- a) anode processes
 - b) cathode processes
 - c) electrolyte (aqueous or molten)
 - d) ion migration
 - e) electron flow in the external circuit.
- 5.28 Explain how electrolysis reactions provide evidence for the Atomic Theory.
- 5.29 Predict the products of simple electrolysis reactions using, where appropriate, a table of

standard reduction potentials (E°) as exemplified by:

- a) 1 mol L^{-1} copper(II) sulfate solution with inert or copper electrodes
 - b) 1 mol L^{-1} sulfuric acid solution with inert electrodes
 - c) 1 mol L^{-1} sodium hydroxide solution with inert electrodes
 - d) 1 mol L^{-1} sodium chloride solution with inert electrodes
 - e) molten sodium chloride with inert electrodes.
- 5.30 Describe the use of electrolysis in electrolytic refining e.g. copper, and electroplating e.g. silver.
- 5.31 Calculate the charge carried by an electric current flowing for a given time using $Q = It$.
- 5.32 Perform calculations relating moles of electrons and charge, given that the charge carried by one mole of electrons is -9.649×10^4 coulomb(C).
- 5.33 Perform calculations involving number of moles of substances, masses of substances, volumes of gases and quantities of electric charge used in electrolysis reactions.
- Note:** Gas volume calculations may involve the conversion of volumes of gases from STP to non-STP conditions and vice versa *as part of a longer calculation.*

Laboratory work objectives:

- 5.L.4 Demonstrate and identify the products of some typical electrolysis reactions.
- 5.L.5 Demonstrate the relationship between the number of moles of electrons and the number of moles of substance produced during electrolysis.

Extraction of metals**Cognitive objectives:**

- 5.34 Describe the extraction of metals, using, as examples, the production of:
- a) iron by carbon reduction
 - b) aluminium by electrolytic reduction
 - c) gold by carbon in pulp
- Relate the method of extraction to the ease of reduction of the metal concerned.

Redox titrations**Cognitive objectives:**

- 5.35 State the characteristics of a primary standard.
- 5.36 Define a 'standard solution' as one with an accurately known concentration.
- 5.37 Describe how to prepare a standard solution by:
- a) dissolving a known mass of a primary standard in a known volume of solution (e.g. ammonium iron(II)sulfate hexahydrate, oxalic acid dihydrate or an oxalate)
 - b) standardisation of the solution by titration against a standard solution.
- 5.38 Explain why an indicator is not required in potassium permanganate titrations.
- 5.39 Describe the use of a pipette, burette and volumetric flask in volumetric analysis.

5.40 Describe the procedures used in volumetric analysis with particular reference to the following redox titrations:

- standardisation of potassium permanganate solution using ammonium iron(II) sulfate hexahydrate or oxalic acid dihydrate
- analysis of hydrogen peroxide solution using standard potassium permanganate solution.

Laboratory work objectives:

5.L.6 Carry out the procedures used in volumetric analysis with particular reference to the following redox titrations:

- standardisation of potassium permanganate solution using ammonium iron(II) sulfate solution or oxalic acid solution
- analysis of hydrogen peroxide solution using standard potassium permanganate solution.

6. Organic chemistry

Hydrocarbons

Cognitive objectives:

- Relate the large number of known carbon compounds to the unlimited number of stable chains, branched chains and rings which can be formed due to:
 - strong covalent bonding between carbon atoms and between carbon and hydrogen atoms
 - the high valency (4) of carbon, which allows chains to branch and to be bonded to a variety of substituents.
- Distinguish between alkanes, alkenes and alkynes according to the type of carbon-carbon bonding.
- Describe the structure of benzene and relate the stability of its structure to the presence of delocalised electrons.

Nomenclature and formulae

Cognitive objectives:

- Write the IUPAC names and structural formulae of:
 - straight and branched chain alkanes (C_1 to C_8)
 - simple cycloalkanes (C_3 to C_6)
 - straight and branched chain alkenes (C_2 to C_8)
 - simple cycloalkenes (C_3 to C_6)
 - straight and branched chain alkynes (C_2 to C_8)
 - straight chain alkyl groups (C_1 to C_8)
 - halogen substituted alkanes (C_1 to C_8)
 - simple examples of: alcohols, aldehydes, ketones, carboxylic acids, esters and primary amines.

***Note:** It is not expected that students will be able to name such compounds as dienes, diynes, trienes, triynes etc.
- Describe and identify simple aromatic compounds as substances containing a benzene ring.
- Identify and write structural formulae of structural isomers having:
 - different carbon configurations (branched chains and cyclic compounds)

- the same carbon chain with the same functional group in different positions (e.g. the propanols)
- the same carbon chain with different functional groups (e.g. propanal and propanone).

- Identify and write structural formulae for cis/trans geometric isomers.
- Explain and apply the terms 'empirical formula', 'molecular formula' and 'structural formula'.
- Perform calculations to determine the empirical formula of an unknown substance given one of the following:
 - its elemental composition by mass
 - its elemental percentage composition
 - information concerning the quantities of products (e.g. masses and gas volumes) obtained from its combustion or other specified reactions.
- Determine the molar mass of an unknown substance from the volume of a known mass of its vapour.
- Determine the molecular formula of a substance from its molecular weight and its empirical formula.
- Draw possible structural formulae of an unknown organic compound given its molecular formula.

Functional groups

Cognitive objectives:

- Describe the importance of the functional groups present in an organic compound in determining its chemical behaviour.
- Write the structures for, and identify functional groups in the following compounds: alkenes, alkynes, alcohols, aldehydes, ketones, carboxylic acids, esters, haloalkanes and primary amines.
- Distinguish between the structure of primary, secondary and tertiary alcohols as exemplified by alcohols with the formula C_4H_9OH .
- Describe, by stating observations and writing equations, the reactions of alcohols, in particular ethanol, with the following substances:
 - sodium
 - acidified solutions of the oxidising agents $MnO_4^-(aq)$ and $Cr_2O_7^{2-}(aq)$
 - carboxylic acids.
- Differentiate between the reactions of primary, secondary and tertiary alcohols in 6.16 a) and 6.16 b).
- Identify and write equations for the substitution reactions of alkanes with halogens.
- Identify and write equations for the addition reactions of alkenes and alkynes with hydrogen and halogens.
- Write chemical equations to illustrate the following reactions:
 - the formation of aldehydes by the oxidation of primary alcohols
 - the formation of ketones by the oxidation of secondary alcohols
 - the formation of carboxylic acids by the oxidation of primary alcohols or aldehydes
 - the formation of esters by the reaction of carboxylic acids with alcohols (esterification).

Laboratory work objectives:

- 6.L.1 Demonstrate the reactions of alcohols with the following:
- sodium
 - oxidising agents such as acidified solutions of MnO_4^- and $\text{Cr}_2\text{O}_7^{2-}$.
- 6.L.2 Prepare esters by the reaction of alcohols and carboxylic acids.

Common organic substances**Cognitive objectives:**

- 6.21 Draw structural formulae to exemplify a simple soap and a sulfonate detergent.
- 6.22 Describe the action of soaps and sulfonate detergents in terms of the interaction of the polar and non-polar parts of the molecule with water and greases, etc.
- 6.23 Describe the preparation of a simple soap and write a simple equation for the reaction (ester hydrolysis).
- 6.24 Describe with examples the use of organic liquids as solvents, e.g. petrol, kerosene, chloroform, acetone, esters, mineral turpentine and methylated spirits.
- 6.25 Define the terms 'monomer' and 'polymer'.
- 6.26 Describe the formation of poly(ethene) and poly(vinyl chloride) (addition polymers) and polyesters (condensation polymers) and write equations for their formation.

Laboratory work objectives:

- 6.L.3 Prepare a simple soap and compare the properties of soaps and sulfonate detergents.
- 6.L.4 Prepare a polymer.

7. Inorganic chemistry**Periodic table****Cognitive objectives:**

- 7.1 Relate the arrangement of elements in the periodic table to their atomic number and their electron configuration.
- 7.2 Describe the following trends across rows in the periodic table as illustrated by the third (sodium-argon) row:
- ionisation energy and common oxidation states
 - physical properties in relation to the nature of the bonding present in the elements
 - oxidising and reducing properties
 - acidic and basic properties of the oxides and the hydroxides.
- 7.3 Describe the following trends down groups in the periodic table as illustrated by groups IV and VII:
- ionisation energy
 - physical properties in relation to the nature of the bonding present in the element.

Laboratory work objectives:

- 7.L.1 Demonstrate trends in the physical and chemical properties of the third row elements.

Transition metals**Cognitive objectives:**

- 7.4 Relate the existence of the transition metals (d block elements) to the filling of d orbitals.
- 7.5 State and apply that transition metals exhibit variable oxidation states.
- 7.6 State and apply that transition metals tend to form coloured compounds.
- 7.7 State and apply that transition metals tend to form complex ions.

Laboratory work objectives:

- 7.L.2 Demonstrate the following properties of some transition elements:
- variable oxidation states
 - formation of coloured compounds
 - formation of complex ions.

Industrial chemicals**Cognitive objectives:**

Describe, where required, the preparation, extraction, manufacture, properties and related uses of the following important elements and compounds. Where properties are to be described these include, where appropriate, appearance, colours of ions, reaction with oxygen, reaction with water, reaction with acids and bases, and the uses related to these properties.

<i>Element or compound</i>	<i>Knowledge required</i>
7.8 Aluminium	– extraction and purification of alumina from bauxite – extraction of aluminium from alumina by electrolysis – properties
7.9 Copper	– properties
7.10 Gold	– extraction by carbon in pulp process – properties
7.11 Iron	– extraction – properties – steels as alloys of iron that contain other elements
7.12 Zinc	– properties
7.13 Ammonia	– manufacture by the Haber process – properties – describe and write equations for the reaction of $\text{NH}_3(\text{aq})$ with $\text{Cu}^{2+}(\text{aq})$, $\text{Zn}^{2+}(\text{aq})$, $\text{Ag}^+(\text{aq})$, $\text{Cu}(\text{OH})_2(\text{s})$, $\text{Zn}(\text{OH})_2(\text{s})$ and $\text{Ag}_2\text{O}(\text{s})$
7.14 Fertilizers	– uses of ammonium sulfate, superphosphate, potassium fertilizers and trace elements
7.15 Hydrochloric acid	– manufacture – properties
7.16 Sulfuric acid	– manufacture from sulfur or sulfide ores – properties.

Laboratory work objectives:

7.L.3 Demonstrate the properties of zinc, iron and copper.

Practical work**Introduction**

In carrying out practical work, students should be encouraged to look upon their experiments as something in the nature of research or investigation, rather than to regard them as efforts to reproduce known results. Students should keep a laboratory notebook in which short descriptions of experimental conditions and details of observations should be written whilst the experiments are being carried out.

With many of the experiments, the following general concepts should be illustrated:

- the importance of measurement in scientific work
- the degree of uncertainty in measurements and the use of significant figures as an indication of uncertainty.

Students should be instructed in the safety aspects of chemicals which are scheduled for use in the laboratory work. Many of the chemicals in the following experiments are commonly encountered in everyday experience.

Experiments**Atomic structure and bonding**

Types of chemical bond related to the electrical conductivity of some liquids and solutions.

Chemical reactions

Rates of chemical reactions; the effect of temperature, concentration and catalysts.

Chemical equilibrium and the application of Le Châtelier's Principle.

Volumetric analysis – analysis by the measurement of solution volumes; characteristics and uses of primary standards, preparing standard solutions, use of pipettes and burettes, identifying the end point of titration.

Acid-base titrations:

- standardisation of hydrochloric acid with solid sodium carbonate
- standardisation of sodium hydroxide solution with standard acid solution
- analysis of common substances such as vinegar or battery acid by titration with standard sodium hydroxide solution.

Redox titrations using potassium permanganate in acid solution:

- standardisation of potassium permanganate solution using iron(II) ammonium sulfate or oxalic acid
- analysis of hydrogen peroxide solution using standard potassium permanganate solution.

Acids and bases in aqueous solution

Use of indicators to determine hydrogen ion concentration.

Chemical properties of acids and bases.

Oxidation and reduction

Typical redox reactions.

Electrochemical cells.

The corrosion of iron.

Electrolysis – the relationship between the charge which passes through a circuit and the amount of product formed.

Organic chemistry

Some reactions of alcohols.

Preparation of an ester.

Preparation of a polymer.

Inorganic chemistry

Properties of some metals e.g. Zn, Fe, Cu.

Chemical properties of some third row elements.

Properties of some transition elements – variable oxidation states, formation of coloured compounds, formation of complex ions.

Time Allocation

The subject has been designed to be completed through a structured education program of approximately 110 hours in any suitable contexts and series of learning experiences. Typically the subject will be studied over the period of one school year. For administrative reasons schools wishing to vary this delivery pattern (e.g. over a shorter period or over a longer period up to two school years) are required to notify the Chief Executive Officer of the Curriculum Council.

Subject Completion

Students must complete the school's structured educational and assessment program for a subject in order to be eligible to receive a grade unless there are exceptional and justifiable circumstances. In situations where the school considers that insufficient information has been gathered to justify the award of a grade for the subject, a result of U (for unfinished) should be allocated. The Curriculum Council offers the flexibility for the U to be converted to a grade after the final grades have been submitted. Further details on assessment and grading are provided in Volume I of the Syllabus Manuals.

Resources

Note 1: The resources in this list were available at the time of printing, but please be aware that their subsequent availability cannot be guaranteed.

Note 2: '(OP)' identifies resources that are out of print but still valuable if teachers can locate them in their department or library.

Textbooks

One of the following is recommended as a general textbook.

Bucat, R.B., (Supervising Editor), *Elements of Chemistry*, Australian Academy of Science, Canberra, 1983.

Garnett, P.J., et al., *Foundations of Chemistry*, 2nd edn., Addison Wesley Longman, Melbourne, 1996.

Hunter, R.J., et al., *Chemical Science*, Science Press, Sydney, 1990.

Teacher References

The following books may be of use to teachers as source material, but most do not follow this syllabus closely enough to be set as student texts.

Cole, A.R.H., Watts, D.W., and Bucat, R.B., *Chemical Properties and Reactions*, Cole, Watts and Bucat, Perth, 1981.

Commons, C., et al., *Chemistry Two – Chemistry and the Marketplace, Energy and Matter*, Heinemann Education Australia, 1991.

Elvins C., et al., *Chemistry One – Materials, Chemistry in Everyday Life*, Heinemann Educational Australia, 1990.

Simpson, P.G., *A new Approach to Teaching and Assessing Practical Work in Upper School Chemistry: A Handbook for Teachers*, (rev.ed.), Science Teachers' Association of WA, Perth, 1999.

Smith, A., and Dwyer, C., *Key Chemistry Book 1: Materials and Everyday Life*, Melbourne University Press, Melbourne, 1991.

Smith, A., and Dwyer, C., *Key Chemistry Book 2: Energy, Matter and the Market Place*, Melbourne University Press, Melbourne, 1992.

Smith R., *Conquering Chemistry*, McGraw-Hill, 1987.

Teacher Guides

There is a teachers' guide for each of the student texts. They are written by the same groups of authors and published by the same publishers as the student texts.

Laboratory Manuals

Garnett, P.J., Garnett, P.J. and Anderton, J., (eds.), *Chemistry Laboratory Manual*, 2nd edn., Science Teachers' Association of WA, Perth, 1998.

The laboratory work is incorporated in the text, *Elements of Chemistry*.

Chemical Calculations

Either of the following books may serve as a useful additional source of numerical problems.

Kolomyjec, C., Garnett, P., and Anderton, J., (eds.), *Chemistry Problems*, Science Teachers' Association of WA, Perth, 1991.

Sleet, R., *Chemical Calculations: A Programmed Approach*, Science Press, NSW, 1979.(OP)

Examination Details

The examination will consist of one written paper of three hours duration.

Part One: multiple choice questions (30%).

Part Two: short answer questions (35%).

Part Three: calculations (25%).

Part Four: extended writing (10%).

Resources:

- Candidates will need a 2B, B or HB pencil for answering multiple choice questions.
- Calculators satisfying the conditions set by the Curriculum Council for this subject, which are listed on the Curriculum Council website:
<http://www.curriculum.wa.edu.au/pages/student/calculators.htm>

Note: No questions will be set that require the use of mathematical functions other than those on a scientific calculator.

Assessment Structure

Assessment structures are an integral part of all Accredited Subjects.

The structure specifies:

1. the components and learning outcomes to be included in assessment
2. weightings to be applied to these components
3. the types of assessment considered appropriate for the subject.

Table 1

Syllabus Content	Weighting percentage
Macroscopic properties of matter	60-80
Atomic structure and bonding	
Chemical reactions	
Acids and bases in aqueous solutions	
Oxidation and reduction	
Organic chemistry	10-20
Inorganic chemistry	10-20

Table 2

Learning Outcomes	Weighting percentage
Cognitive:	90-95
Knowledge, process	
Sensorimotor	5-10

Table 3

Types of Assessment	Weighting percentage
Examinations/extended tests on longer periods of work (see notes)	55-70
Topic or section tests	10-25
Laboratory performance	5-10
Laboratory reports	5-10
Assignments	0-5

The assessment program must provide students with the opportunity to demonstrate achievement of the requirements of the subject,

and

students must complete the requirements of the subject.

Notes on Table 1

The weighting reflects the time devoted to teaching the material. Practical work is applicable to all three sections of the syllabus.

Notes on Table 2

Knowledge is deemed to include the following lower-order cognitive skills: defining, recalling, relating, listing, describing, performing simple calculations.

Process is deemed to include the following higher-order cognitive skills: classifying, interpreting, applying, explaining, demonstrating, analysing, inferring, predicting, evaluating, communicating.

Sensorimotor involves laboratory performance as defined in the Notes on Table 3.

Notes on Table 3

For the purpose of this document:

Examinations/extended test on longer periods of work are formal written examinations or tests covering major portions of the syllabus (normally term or semester examinations unless only one major examination is held).

Topic or section tests are class tests covering a restricted portion of work.

Laboratory performance is to include the following:

- ability to follow directions
- use of safe procedures
- manipulative skills
- confidence and independence
- understanding the experimental procedure.

Laboratory performance may be assessed under test conditions.

Laboratory reports are the students' written records of experiments and could include such sections as:

- aim
- apparatus/equipment
- method/procedure
- results
- discussion
- conclusion
- answers to set questions.

Laboratory reports may be assessed under test conditions.

Grade-Related Descriptors

Grade-Related Descriptors describe the student performance standards that are used to award grades in this subject. Schools delivering this subject have been provided with a copy of the document. Additional copies may be purchased from the Curriculum Council.

