

AM SYLLABUS (2021)

CHEMISTRY

AM 06

SYLLABUS

AM06 Chemistry
Syllabus Addendum

Mitigating factors for 2021 MATSEC Examinations Session

Changes in Subject Content	<p>The following content will not be assessed:</p> <ul style="list-style-type: none"> ● Section 6.2: Deviations from Raoult's Law. Azeotropic mixtures. Fractional distillation of non-ideal mixtures. Fractional distillation of ideal mixtures may still be assessed. ● Section 6.3 ● Section 12.6 except the following which may still be assessed: <ul style="list-style-type: none"> ✓ describe the chromate(VI)-dichromate(VI) equilibrium and the effect of pH change on this equilibrium; ✓ know about the toxicity of Cr(VI) compounds; ✓ explain the use of KMnO_4 as a strong oxidising agent in titrimetry and in organic chemistry; ✓ give reasons for the non-existence of iron(III) iodide and iron(III) carbonate; ✓ explain the acidity of hexaaquairon(III) ion and compare with that of the corresponding iron(II) species; ✓ state how $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$ is involved in qualitative analysis for Fe(III); ✓ describe how Cu_2O is formed in Fehling's test for aldehydes and how this oxide can be converted into CuCl and CuI; ✓ describe the instability of $\text{Cu}^+(\text{aq})$ with respect to disproportionation and to explain how ligands other than water stabilize this ion. ● Section 13.2: steam distillation; column chromatography; mixed melting point technique ● Section 14.9 ● Section 14.11 (d) and (e) ● Section 15.2 complexometric titrations involving EDTA; titrations with silver nitrate will not be set ● Section 15.3
Changes in Coursework	N/A
Changes in Exam Paper(s)	<p>Paper 1: No changes Paper 2: No changes Paper 3: No practical examination. This practical paper will be replaced by a written paper as exemplified in the sample paper below. It will carry 20% of the total score and will be of 1 ½ hour duration. A copy of the Periodic Table will be provided. Paper 3 will consist of three compulsory questions of 15, 15 and 20 marks respectively. Questions will be based on practical work related to section 15.2 excluding complexometric titrations involving EDTA; titrations with silver nitrate.</p>

MATSEC BOARD
October 2020



SUBJECT:	Chemistry
PAPER NUMBER:	III
DATE:	
TIME:	1.5 hours

Directions to Candidates

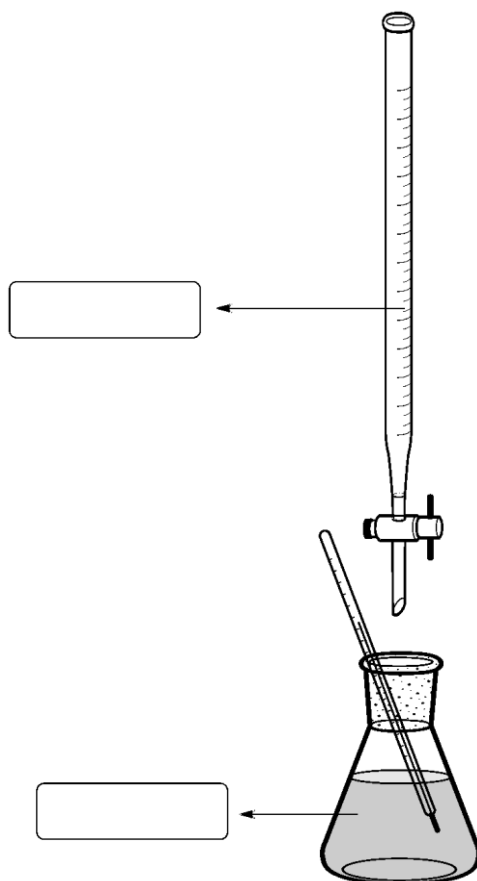
- Write your index number in the space at the top left-hand corner of this page.
 - Answer all questions. Write all your answers in this booklet. Drawings and graphical representations of data are to be made on the appropriate pages within this booklet.
 - The marks allotted to parts of question are indicated.
 - You are reminded of the necessity for good English and orderly presentation in your answers.
 - In calculations you are advised to show all the steps in your working, giving your answer at each stage.
 - The use of electronic calculators is permitted.
 - The use of a copy of a periodic table is permitted.
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For examiners' use only:

Question				Total
Score				
Maximum				50

b) Draw and label the experimental set-up for the filtration of CaC_2O_4 as described in step (II).

- d) (i) Label the solutions that would be present in the glassware, as set-up below, before the titration is carried in step (III). (2)



- (ii) Explain why a thermometer is present in the conical flask before the start of the titration, as shown in the set-up in part (d)(i)?

(3)

- (iii) Describe how the end-point is determined.

(1)

Total: 20 marks

2. An experiment was carried out to determine the percentage purity of calcium carbonate in an impure sample by a back titration. An impure sample of calcium carbonate of mass 1.356 g was transferred quantitatively in a 250.0 cm³ volumetric flask. To this sample was added 50.0 cm³ of a hydrochloric acid solution of concentration 1.20 mol dm⁻³. Shaking was performed till the reaction was complete, after which the solution was made up to 250.0 cm³.

This solution containing unreacted acid was placed in a burette and was titrated against 25.0 cm³ of aqueous sodium hydroxide of concentration 0.100 mol dm⁻³. The titration was repeated three more times. The results obtained are shown in the table below.

	Approximate	2	3	4
Final burette reading / cm ³	18.10	18.05	18.65	18.70
Initial burette reading / cm ³	0.00	0.50	0.75	1.25
Titre / cm ³				

- a) Complete the table above by writing down the titre values. (1)

- b) Select the appropriate titre values from the table and calculate their mean.

_____ (1)

- c) Calculate the number of moles of HCl initially added to the impure CaCO₃ in the volumetric flask.

_____ (2)

- d) Calculate the number of moles of HCl that reacted in the titration with aqueous NaOH.

_____ (3)

e) Calculate the number of moles of HCl remaining in 250.0 cm³ after reaction with CaCO₃.

(2)

f) Calculate the number of moles of HCl that reacted with the CaCO₃.

(1)

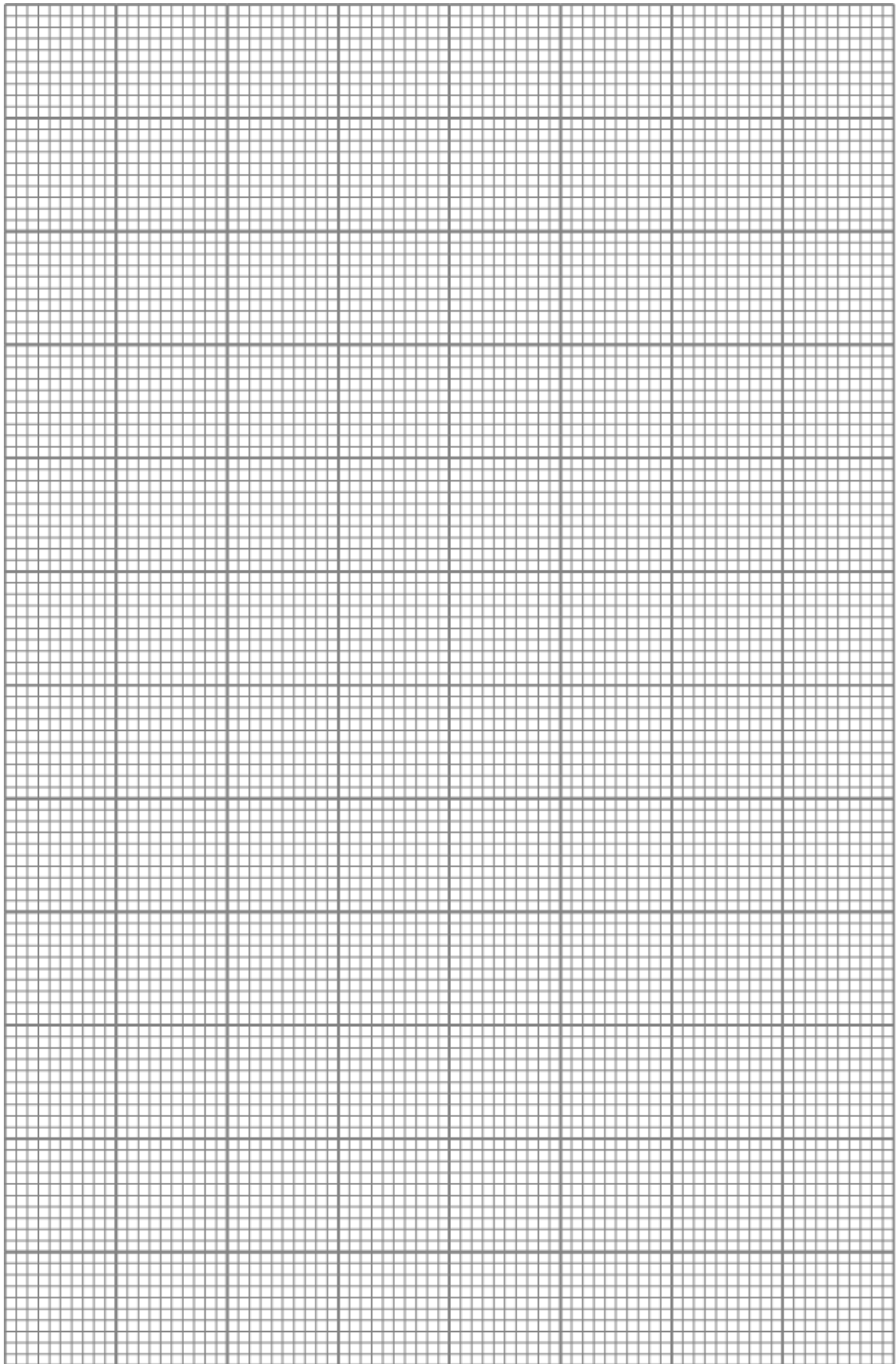
g) Calculate the number of moles of CaCO₃ in the impure sample used.

(2)

h) Calculate the percentage purity of the CaCO₃ sample, given that the molar mass of CaCO₃ is 100 g mol⁻¹. Assume there are no impurities that react with HCl.

(3)

(Total: 15 marks)

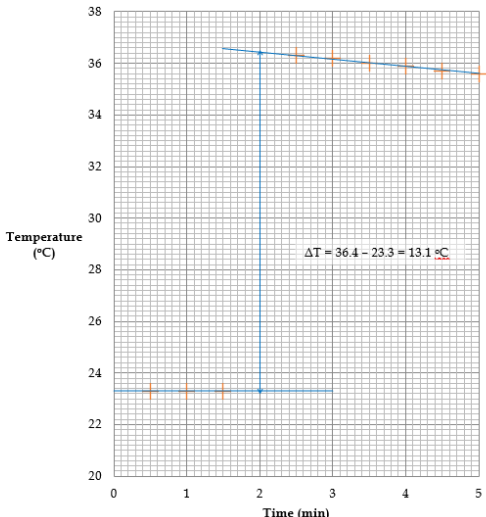


(Total: 15 marks)

Sample Paper 3 – Suggested Marking Scheme

Question No.	Suggested Answer	Mark	Comments					
1	a Record the mass of the weighing boat plus sample Transfer the solid sample to a conical flask. Re-weigh the empty weighing boat. The difference in mass is equal to the mass of the added sample. A glass funnel is inserted inside the conical flask and the HCl added slowly whilst stirring Stirring is continued until all the effervescence stops. The mixture is filtered, and the residue rinsed twice with distilled water. The pH of the filtrate can be checked either via a pH meter or else by taking a small drop on a glass rod and using pH paper. Dilute NaOH is added dropwise until the pH is close to 7.	1½	Award 1 mark only for "quantitative transfer"					
		½						
		1						
		1						
		1						
		1						
		1						
b	Drawing of conical flask/beaker with funnel Fluted / filter paper. Residue showing CaC ₂ O ₄ (s) Filtrate showing aqueous solution (mainly CaCl ₂ solution.	1						
		1						
		1						
		1						
c	Transfer the filtrate quantitatively into a 250 cm ³ volumetric flask using a clean funnel. Rinse the conical flask/beaker at least three times with distilled water and transfer each wash into the volumetric flask. Rinse the funnel with distilled water and transfer the rinse into the volumetric flask. Make up the solution with distilled water so that the bottom of the meniscus is on the mark. This must be done with the meniscus viewed at eye level. Stopper the volumetric flask and shake the flask a number of times / approximately 1 minute to homogenize the contents.	1						
		½						
		½						
		½						
		½						
di	Burette – potassium manganate(VII) solution Conical flask – acidified calcium ethanedioate solution.	1						
		1						
dii	The reaction between manganate(VII) ions and ethanedioate is autocatalytic. The reaction is initially slow and must be heated to increase the rate of reaction	1						
		1						
diii	The first permanent pink colour	1						
Total		20						
2	a	Approx.	2	3	4	1	Deduct ½ mark if value is not quoted to two decimal places.	
		Final burette reading / cm ³	18.10	18.05	18.65			18.70
		Initial burette reading / cm ³	0.00	0.50	0.75			1.25
		Titre / cm³	18.10	17.55	17.90			17.45

b	$\frac{17.55 + 17.45}{2} = 17.50 \text{ cm}^3$	1	Do not award any marks if one of the chosen values is incorrect. Deduct ½ mark if the correct values are chosen but the calculation is incorrect.
c	<p>In 1000 cm³ of solution there are 1.20 mol of HCl</p> <p>In 50.0 cm³ of solution there are $\frac{1.20 \times 50.0}{1000} = 0.06000$ mol of HCl</p>	2	Deduct ½ mark if the final value is given as 0.06 (an incorrect number of significant figures). Allow follow through.
d	<p>In 1000 cm³ of solution there are 0.100 mol of NaOH</p> <p>In 25.0 cm³ of solution there are</p> $\frac{0.100 \times 25.0}{1000} = 2.500 \times 10^{-3} \text{ mol of NaOH}$ <p>NaOH (aq) + HCl (aq) → NaCl (aq) + H₂O (l)</p> <p>Since NaOH and HCl react in 1:1 mole proportion, the number of moles of HCl that react in the titration is 2.500 × 10⁻³</p>	3	Award 2 marks for correct calculation of moles of NaOH. Award 1 mark for correct deduction of number of moles of HCl. Allow follow through.
e	<p>In 17.50 cm³ of solution there are 2.500 × 10⁻³ mol of HCl</p> <p>In 250.0 cm³ of solution there are $\frac{(2.500 \times 10^{-3}) \times 250.0}{17.50} = 0.03571$ mol of HCl</p>	2	Allow follow through.
f	Moles of acid that reacted with carbonate = 0.06000 – 0.03571 = 0.02429 mol	1	Allow follow through.
g	<p>CaCO₃ (s) + 2HCl (aq) → CaCl₂ (aq) + H₂O (l) + CO₂ (g)</p> <p>2 mol HCl react with 1 mol CaCO₃</p> <p>0.02429 mol HCl react with $\frac{0.02429}{2} = 0.01215$ mol CaCO₃</p>	2	Award ½ mark for correct deduction of mole proportion. Award the rest of the marks for the calculation. Allow follow through.

	h	<p>1 mol of CaCO_3 weighs 100 g 0.01215 mol of CaCO_3 weigh $100 \times 0.01215 = 1.215$ g</p> <p>$\% \text{ purity } (\text{CaCO}_3) = \frac{1.215}{1.356} \times 100 = 89.6\%$</p>	3	<p>Award 1 mark for correct calculation of mass of CaCO_3. Award 2 marks for correct calculation of percentage purity. Allow follow through. Deduct $\frac{1}{2}$ mark if the final value is not given to three significant figures.</p>
Total			15	
3	ai		6	<p>Award 1 mark for each axis Award 3 marks for the correct plotting Award 1 mark for fitting the best line</p>
	aii	<p>extrapolation reading and calculating the correct temperature change</p>	1 2	
	b	<p>$\Delta H_{\text{rxn}} = -mc\Delta\theta$ $m = 0.10$ kg</p> <p>$\Delta H_{\text{rxn}} = -0.1 \times 4.18 \times 13.1$ $\Delta H_{\text{rxn}} = -5.48$ kJ</p> <p>$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$ 0.1 mol 0.1 mol 0.1 mol</p> <p>0.1 mol = -5.48 kJ 1 mol = -54.8 kJ</p> <p>$\Delta H^{\circ}_{\text{rxn}} = -54.8 \text{ kJ mol}^{-1}$</p>	1 1 1/2 1/2 1 1 1/2 1/2	
Total			15	

**Chemistry AM06
Syllabus**

Paper I (3 hours) + Paper II (3 hours) + Paper III (3 hours)

Chemistry is not only a subject of academic study which has engaged scholars for centuries worldwide but its products and processes have indeed enabled the very existence of societies and allowed their development. It is a central science which informs several other disciplines including biology, medicine, materials science, agricultural science and others. Chemical warfare and pollution of the environment as well as drug abuse, adulteration of food and similar examples of misuse or improper use or disposal of chemicals often cast shadows on this scientific enterprise and generate mistrust of chemicals by the general public.

The main purpose of this syllabus is to impart a proper understanding of the core principles and unifying ideas of chemical science, supported and illustrated by a significant body of essential factual knowledge that should help students acquire a meaningful experience of chemistry to serve them adequately in their future studies. The syllabus also provides an opportunity for learners to become properly acquainted with this important science in order for them to be able to better appraise both the opportunities and the challenges which chemistry and its products present to society at large.

The syllabus for Chemistry at this level assumes and builds on knowledge of the subject at a level that is equivalent to that covered by the Secondary Education Certificate syllabus.

The syllabus is presented in two columns, describing Subject Content and Learning Outcomes. This design is intended to assist teachers and students determine the depth of treatment for each subject matter mentioned.

Teachers may of course go beyond the learning outcomes in class but the examination will test the outcomes as described in this syllabus.

The course is intended to build both a theoretical knowledge base in chemistry that prepares students for further higher studies in the subject as well as to develop basic practical skills. With respect to practical chemistry, students should learn how to perform laboratory work in a manner which respects both their health and the environment; teachers should evaluate the risks associated with practical work with a view to safeguard student safety and environmental integrity.

The Examination

The examination consists of two written papers and a practical paper. Each of the papers is of three hours duration.

Paper I consists of six to ten compulsory structured questions and carries 40% of the total score. Students will write their answers on the examination paper in the spaces provided.

Paper II consists of two sections each containing four extended response questions. Five questions are to be answered from this paper, two questions from each section and any one other. Each question carries equal marks. This paper carries 40% of the total score.

Paper III is a practical examination. This is an open book examination and candidates may use any printed material which assists them in their work. This paper carries 20% of the total score.

Candidates may use an electronic calculator in all parts of the examination.

Familiarity with Periodic Table

Candidates are expected to be familiar with the structure of the Periodic Table and the group affiliation of each element with atomic number from 1 to 30. However, a copy of the Periodic Table will be provided for Paper II of the examination although not for Paper I. Relative atomic masses will always be given where necessary.

Mathematical skills

In order to understand certain concepts required by the syllabus, candidates will need to be able to:

- (i) Recognise and use expressions in decimal and standard form; use ratios, fractions and percentages and find arithmetic means; make estimates of results; use an appropriate number of significant figures; use calculators to find and use x^y , \sqrt{x} , $1/x$, $\log_{10} x$.

- (ii) Change the subject of an equation; substitute numerical values into algebraic equations using units for physical quantities; use logarithms in relation to quantities which range over several orders of magnitude.
- (iii) Appreciate angles and shapes in regular two and three dimensional structures and to represent three dimensional forms in two dimensions; understand the symmetry of two dimensional and three dimensional shapes.
- (iv) Plot two variables from given data; understand that $y = mx + c$ represents a linear relationship and be able to determine the slope and intercept of a line; draw and use the slope of a tangent to a curve as a measure of rate of change.

Language skills

Candidates are reminded of the importance of the use of good English in answering questions in the examination. In essay-type answers and also where calculations are involved, orderly and detailed presentation will be rewarded.

Syllabus

Content	Learning outcomes and comments
1 <i>Atomic structure</i>	
1.1 The fundamental particles: protons, electrons and neutrons, their charges and relative masses.	Students should be able to: describe the structure of an atom in terms of fundamental particles using the Bohr model;
1.2 The nuclear structure of the atom. Proton (or atomic) number and nucleon (or mass) number. Isotopes and relative atomic masses. The carbon-12 scale.	Students should be able to: deduce the relative atomic mass of an element from the isotopic masses and relative abundances; describe the use of isotopes as tracers in mechanistic studies using the example of ^{18}O labelling in esterification;
1.3 Radioactivity: alpha and beta particles and gamma rays. Half-life. Nuclear equations.	Students should be able to: apply the concept of half life in determining the age of a carbon-containing artefact; determine the half-life of a radioactive isotope using an exponential decay curve; describe C-14 dating. Positron emission will not be tested. Calculations requiring use of the equation for radioactive decay will not be set.
1.4 Introductory treatment of quantised energy levels in atoms: evidence from atomic spectra; s, p and d orbitals.	Students should be able to: draw the electronic configuration of isolated atoms of elements H to Kr using 1s, 2s, 2p etc notation and electrons- in-boxes notation and applying the 'building-up' (aufbau) principle. Questions on interpretation of atomic spectra will not be set although students are expected to explain how transitions between energy levels can give rise to flame colours.
1.5 Classification of elements into periods, groups and 'blocks', including the first row d-block elements.	
1.6 Shapes of s and p orbitals. Hybridised orbitals.	Students should know the shape of hybrid orbitals sp, sp ² and sp ³ and relate it to molecular geometry.
1.7 Ionisation energies and electron affinity and relation to electronic configuration.	Students should be able to: explain how ionisation energies vary across the second and third period, both the general trends and irregularities and down a group in the Periodic Table; be able to explain why second and subsequent electron affinities are endothermic.

1.8 The mass spectrometer. An elementary treatment of mass spectrometry including the use of the mass spectrometer to determine the relative atomic mass and as a tool for determination of molecular structure (see also section 2.10).	Students should be able to: draw and label a diagram showing the principle of operation of a magnetic sector mass spectrometer; recognize and interpret a simple mass spectrum of an atomic and a molecular sample, explaining the peak height ratios for the mass spectra of Cl ₂ and Br ₂ , and restricting discussion to singly-charged peaks; use the mass spectrum of an element to obtain the relative atomic mass.
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2	<i>Electronic theory and chemical bonding</i>	
2.1	The ionic (electrovalent), covalent and coordinate (or dative) bond. Electronegativity. Intermediate bonding: ion polarisation and bond polarisation.	The concept of electron sharing and electron transfer between atoms as complete or partial processes depending on the electronegativity of the atomic species concerned with reference to Pauling's electronegativity scale. Dative bonding exemplified by molecular compounds such as CO, HNO ₃ , Al ₂ Cl ₆ , NH ₃ BF ₃ and ions H ₃ O ⁺ and NH ₄ ⁺ .
2.2	Polar covalent bonds and electrical dipoles in molecules (qualitative treatment only). Polar covalent bonds which may, or may not, give rise to molecules with a permanent dipole.	Vector representation of dipole moments to explain why polar covalent bonds in molecules can lead to molecules having a permanent dipole or a zero dipole: qualitative treatment only. Use of symbols δ ⁺ and δ ⁻ to represent partial charges in bonds.
2.3	Nature of forces in bonding. Comparison between ionic and covalent bonding. Multiple bonding.	Students should be able to describe: Lewis structures drawn showing electrons as dots and crosses; covalent bonding in terms of overlap of orbitals, including hybridised orbitals to produce σ and π bonds.
2.4	Delocalisation of electrons.	Evidence for delocalisation from bond lengths and thermochemical data. Delocalised π bonds in, for example, benzene, buta-1,3-diene, ozone and carbonate, nitrate and alkanoate ions. Students should be familiar with resonance structures (canonical forms) and resonance hybrids (delocalised structures).
2.5	Metallic bonding	Electron sea model; use of model to account for malleability, ductility and thermal and electrical conductivity of metals. Students should be able to relate the strength of metallic bonding to attraction between positive ions and delocalised valence electrons and to explain the variation of metallic bonding down a group.
2.6	Lattice structures of NaCl and CsCl as typical ionic solids. Coordination numbers.	Students should be able to relate the ratio of cationic radius to anionic radius to type of crystal lattice and to understand the concept of the unit cell. Students should be able to state the approximate (not exact) bond angles in molecules and ions having both lone and bonding pairs surrounding the central atom with up to six electron pairs
2.7	Shapes of molecules and ions and the valence shell electron pair repulsion (VSEPR) theory.	
2.8	Molecular crystals and macromolecular structures.	Students should be able to : account for physical properties (melting point, boiling point, electrical conductivity) in terms of the structure for such substances as iodine, carbon dioxide, carbon (diamond), carbon (graphite) and silicon(IV) oxide.
2.9	Intermolecular forces: hydrogen bonding, and van der Waals forces (permanent dipole-permanent dipole and induced dipole-induced dipole forces).	Students should be able to: explain the effects of these forces on specific properties of molecular compounds: e.g. boiling points of simple hydrides in Groups 4, 5, 6 and 7; variation of the boiling points along a homologous series of organic compounds; formation of carboxylic acid dimers; the open structure of ice leading to its density being lower than that of liquid

	water; distinguish between intermolecular and intramolecular hydrogen bonding as in 4-nitrophenol and 2-nitrophenol isomers.
2.10 Mass spectrometer as a method of determining molecular structure: interpretation of simple mass spectra involving only singly charged ions; the molecular (parent) ion and fragment ions.	Use of the mass spectrum of a simple molecular substance to obtain information about the molecular structure of the substance. Knowledge of simple fragmentation only (no knowledge of peak formation via rearrangements is expected).

3	<i>States of Matter</i>	
3.1	The ideal gas law.	Use of ideal gas law in the determination of relative molecular mass for gases and volatile liquids: use of gas syringe.
3.2	Dalton's Law of Partial Pressures; diffusion and effusion.	Graham's Law of diffusion will not be examined.
3.3	Kinetic theory of gases; distribution of molecular speeds.	Qualitative treatment only. Students should be able to draw distribution diagrams showing speed of gas molecules for different temperature and to relate this to the concept of activation energy of chemical reactions. See Section 10. The Zartmann experiment will not be tested.
3.4	Real gases and deviation from ideal behaviour. Van der Waals equation.	Basic assumptions of the ideal gas model. No numerical questions will be set to test knowledge of van der Waals equation.
3.5	Vapour pressure and saturation vapour pressure; vaporisation and fusion in terms of the kinetic molecular model.	Students should know about the relationship between vapour pressure and temperature.

4	<i>Quantity of matter</i>	
4.1	Moles of substance. Molar concentration in terms of moles dm ⁻³ or mol L ⁻¹ ; mass concentration in terms of g dm ⁻³ or g L ⁻¹ . Volume measured in dm ³ or L and cm ³ or mL.	Students should be able to define the mole and to perform chemical calculations in terms of molar quantities. Matter described in terms of atoms, molecules, ions.
4.2	Avogadro constant. Empirical and molecular formulae.	
4.3	Chemical equations, both full and ionic and the use of these equations in calculations on reacting substances in terms of amounts and concentrations measured in moles and mol dm ⁻³ respectively.	Students should be able to: use in calculations the concept of limiting reagent; calculate percentage yield and state why reactions do not always produce 100% yield; calculate the percentage purity of a substance from given data.
4.4	Titrimetric analysis: acid-base, precipitation (silver nitrate – chloride) redox and complexometric (EDTA) reactions.	Students should be able to explain the experimental details of titrimetry and be able to carry out a titrimetric analysis (for more details see Section 14).
4.5	Back titrations	Students should be able to explain how to carry out back titration involving acid/base and redox systems.

5	<i>Energetics</i>	
5.1	Changes of energy accompanying phase changes and chemical changes. Energy level (enthalpy) diagrams. Joule (J) as the unit of energy.	
5.2	Standard enthalpy change of reaction, formation, atomisation, ionization, combustion, solution (or dissolution), neutralisation and solvation including hydration. Electron affinity. Bond enthalpy terms and bond dissociation enthalpies; lattice enthalpies; enthalpies of solution and relation to lattice enthalpy and enthalpies of hydration.	<p>Students should be able to:</p> <p>define the enthalpy changes listed in this section;</p> <p>distinguish between first and second electron affinity for dinegative ions;</p> <p>explain the difference between bond enthalpy terms and bond dissociation enthalpies;</p> <p>use bond enthalpies to estimate the enthalpy change of a reaction and explain why values obtained in this manner are approximate.</p>
5.3	Hess's law and its use in simple calculations. Born-Haber cycle.	<p>Students should be able to:</p> <p>state Hess's Law and use it to calculate enthalpy changes indirectly by construction of simple enthalpy cycles or energy level diagrams from given thermochemical data;</p> <p>construct Born-Haber cycles and carry out associated calculations.</p>
5.4	Comparison between theoretical and experimental values of enthalpy change and interpretation in terms of structure and bonding.	<p>Silver halides provide good examples of deviation between theoretical and experimental lattice enthalpies and interpretation of difference in terms of a degree of covalency in this formally ionic structure.</p> <p>The difference between expected and experimental values of enthalpies of formation or combustion of benzene and similar simple hydrocarbons provide evidence of delocalised structures.</p>
5.5	Calorimetry: determination of enthalpy change of neutralization, solution and reaction. Calorific value of fuels and food. Simple thermometric acid-base titrations.	<p>Students should be able to:</p> <p>interpret temperature-time curves to determine temperature change and to use such data in the relationship $\Delta H = - mc\Delta T$;</p> <p>interconvert Joule and calorie as units of enthalpy change.</p>
5.6	Concept of system and its surroundings. Entropy, Gibbs free energy and spontaneity of chemical change. Familiarity and use of the relationship between free energy, enthalpy and entropy, namely, $\Delta G = \Delta H - T\Delta S$. Kinetic versus thermodynamic stability.	<p>Students should be able to:</p> <p>state that ΔH alone is not able to fully predict spontaneous change;</p> <p>consider entropy as a measure of disorder of a system and predict whether ΔS is positive, negative or nearly zero in simple changes such as state change, combustion, dissolution, dimerization;</p> <p>calculate ΔS from absolute entropy values;</p> <p>explain that processes are spontaneous when the free energy change is negative and to calculate the temperature at which processes start or cease to be spontaneous by use of the equation $\Delta G = \Delta H - T\Delta S$.</p>

6 Phase Equilibria	
6.1	Phase changes for one-component systems: H ₂ O, CO ₂ . Pressure-temperature phase diagrams and boiling, melting and triple points. Critical temperature and supercritical fluids.
6.2	Two-component systems: mixtures of two miscible liquids and Raoult's Law. Pressure-composition and temperature-composition diagrams. Deviations from Raoult's Law. Azeotropic mixtures. Fractional distillation of ideal and non-ideal mixtures.
6.3	Immiscible liquids and steam distillation.
6.4	Osmosis and osmotic pressure: use of osmotic pressure, Π , for the determination of relative molecular mass from relation $\Pi V = nRT$ where n is the amount, in moles, of particles of solute. Reverse osmosis.

Students should be able to use phase diagrams to explain sublimation and know about the use of supercritical fluids as 'green' solvents which can replace toxic organic solvents, e.g. in the food industry.

Students should be able to:

give examples of binary mixtures which are approximately ideal and which show both positive and negative deviations from Raoult's Law;

use Raoult's law to calculate the vapour pressure of a component above an ideal binary mixture;

employ Dalton's law to determine the composition of the vapour above an ideal binary mixture.

explain why a mixture of two immiscible liquids boil at a temperature below the boiling point of both liquids;

describe the advantage of steam distillation over simple distillation.

deduce an equation which relates the mass ratio of the substances in the vapour phase to the vapour pressures of the liquids and their relative molecular masses.

Osmosis as an example of a colligative property of solutions. Questions on the determination of partial degree of dissociation/association will not be set.

Students should be able to relate reverse osmosis to local desalination of sea water and compare it to other methods of purification of water in terms of energy considerations.

7 <i>Chemical Equilibrium</i>		
7.1	Concept of dynamic equilibrium. Characteristics of an equilibrium mixture. Equilibrium constant in terms of concentrations and of pressures: K_C and K_P including units as appropriate. Homogeneous and heterogeneous equilibria and their associated equilibrium constants.	<p>Students should be able to:</p> <p>identify the main characteristic of the equilibrium state as constancy of concentrations/pressures resulting from the equality of the rates of the forward and reverse reactions;</p> <p>deduce the extent of reaction from the magnitude of the equilibrium constant;</p> <p>write equilibrium constant expressions involving condensed (liquid or solid) phases in heterogeneous equilibria;</p>
7.2	Degree of dissociation	<p>write an expression for the degree of dissociation of molecular covalent substances converting into smaller molecules;</p> <p>relate the degree of dissociation to K_C and K_P. Problems requiring solving quadratic equations using the formula will not be set.</p>
7.3	Experimental methods of investigating chemical equilibrium.	<p>Students should be able to describe how the following equilibrium reactions can be investigated practically: (a) esterification reaction of a carboxylic acid and an alcohol; (b) dissociation of hydrogen iodide into the elements</p>
7.4	Le Chatelier's Principle. The effect of concentration, temperature, and (where relevant) pressure on: (a) wholly gaseous equilibria, (b) solid-gas equilibria, (c) equilibria in the liquid phase (e.g. esterification) or in solution. Effects of presence of catalysts on equilibrium.	<p>Students should be able to:</p> <p>use the Haber Process for the manufacture of ammonia to illustrate this principle and to explain how and why compromise conditions are required. (Details of manufacturing plant will not be tested);</p> <p>describe and predict the qualitative effect of changes of temperature or concentration or pressure on the equilibrium position and on the value of the equilibrium constant;</p> <p>describe the effect of the introduction of an inert gas on an equilibrium mixture at constant volume.</p>
7.5	Distribution of a non-volatile solute between two immiscible solvents: partition constant; solvent extraction.	<p>Students should be able to solve numerical questions on partition constant K_d related to solvent extraction.</p> <p>Discussion of association and dissociation effects on partition equilibrium are not required.</p>

8 <i>Ionic Equilibria</i>		
8.1	<p>Acids and bases: Arrhenius, Bronsted-Lowry and Lewis definitions; conjugate pairs; proticity or 'basicity' of acids and 'acidity' of bases (e.g. sulfuric acid as a 'diprotic' or 'dibasic' acid and carbonate as a 'diacid' base); strong and weak acids and bases; amphoteric compounds; dissociation constants K_a and K_b and their units. Ionic product of water, K_w; the pK convention; pH of aqueous solutions; hydrolysis of salts.</p> <p>Indicators, typified by phenolphthalein and methyl orange.</p> <p>pH curves</p> <p>Buffer solutions</p>	<p>Students should be able to:</p> <p>use pH to determine the $[H_3O^+]$ and $[OH^-]$ of solutions and <i>vice versa</i>;</p> <p>distinguish between strong and weak acids and bases in terms of the values of the constants concerned;</p> <p>relate K_w to K_a and K_b;</p> <p>calculate concentrations of species using dissociation constants for weak acids and bases;</p> <p>recognize that dibasic acids dissociate in a stepwise manner with an associated K_a value for each step;</p> <p>describe use of pH meter and indicators to measure pH; decide on use of appropriate indicators for titrations and recall the pH range of phenolphthalein and methyl orange;</p> <p>use K_a values of indicators to justify their use in acid-base titrations.</p> <p>Note: pH curves are restricted to monobasic acid – monoacid base reactions.</p> <p>Students are expected to be able to:</p> <p>explain a buffer system and how it works to keep the pH fairly stable;</p> <p>explain applications of buffers (e.g. in blood, in seawater; in industry);</p> <p>perform calculations involving pH of buffer solutions.</p>
8.2	<p>Role of solvent in equilibria involving ionisation of molecular solutes. Degree of ionisation.</p> <p>Qualitative explanation of conductivity of solutions of strong and weak electrolytes.</p>	<p>Students should be able to:</p> <p>calculate the degree of ionization of weak electrolytes which are acidic or basic;</p> <p>recognize the connection between conductivity of a solution and the presence of dissolved ions;</p> <p>explain the special case represented by the conductivity of the proton and the hydroxide ion in aqueous systems;</p> <p>explain how to carry out and interpret the results of a conductimetric titration (not involving weak acids with weak bases).</p>
8.3	<p>Heterogeneous ionic equilibria: solubility product K_{sp}, its units and relation to molar solubility.</p>	<p>Students should be able to:</p> <p>write solubility product expressions for sparingly soluble salts;</p> <p>solve mathematical equations involving K_{sp} and molar solubility of ionic sparingly soluble solutes of the type 1:1 (e.g. AgCl or BaSO₄);</p>

Factors influencing the solubility product equilibrium.	describe the effect on the solubility product equilibrium by addition (common ion effect) or removal of an ion involved in the equilibrium by acid-base or complex ion formation.
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9 Redox Equilibria	
9.1	<p>Redox reactions: balancing of equations using half equations. Disproportionation reaction. Oxidation number (state).</p>
9.2	<p>Electrodes, galvanic cells and standard electrode potentials. Standard hydrogen electrode as a reference electrode.</p>
9.3	<p>Cell diagrams</p> <p>The equation $\Delta G = -z E^{\circ}F$</p>
9.4	<p>Application of electrode potentials to the predict redox change. Electrochemical series.</p> <p>Fuel cells</p> <p>Corrosion as an electrochemical process: sacrificial protection.</p>

Students should be able to:

distinguish between oxidation number (state) and the actual charges on atoms;

explain the connection between oxidation number and bonding, e.g. for carbon in CO, CO₂ and hydrocarbons; oxygen in H₂O and H₂O₂ sulfur in S₂O₃²⁻ and S₄O₆²⁻ ions;

state common oxidising and reducing agents as are used in the laboratory and in the chemical industry.

Students are expected to be able to describe the hydrogen electrode as a typical standard electrode.

Students should be able to:

draw a cell diagram for a galvanic couple, for example for the copper-zinc cell, as follows:
 $\text{Cu(s)} \mid \text{Cu}^{2+}(\text{aq}) \parallel \text{Zn}^{2+}(\text{aq}) \mid \text{Zn(s)}$ and determine for such a cell the polarity and the emf;

describe the connection between the free energy change of a reaction its E° value and also employ electrode potentials to determine whether redox reactions are energetically possible.

Students should be able to:

use the Nernst Equation to predict redox change under non-standard conditions (although the equation will be provided where its use is required);

recognize that a pH meter makes use of the Nernst equation to measure the hydrogen ion concentration in a solution (the theory behind the glass electrode will not be tested);

describe theory of fuel cells using the hydrogen – air and methanol – air cells as examples (operational details and construction of fuel cells will not be tested).

10 Reaction Kinetics	
10.1 The experimental investigation of reaction rate for simple reactions.	<p>Students should be able to describe suitable methods (including their relative merits and demerits) for the determination of reaction rates and be able to interpret the data obtained.</p> <p>The methods that shall be tested are:</p> <p>gas volume changes by gas syringe; colorimetry; conductance measurements; changes of pH for (slow) reactions involving consumption or formation of hydroxonium ions; titrimetry.</p>
10.2 Order and rate coefficients and their measurements, including the initial rate method; rate equation.	<p>Students should be able to:</p> <p>define order of reaction and rate constant;</p> <p>state the units of rate constants;</p> <p>derive the rate equation for a reaction from data provided;</p> <p>present and interpret kinetic data from graphical representations or extract information from graphical representations of kinetic data.</p> <p>Integrated rate equations and fractional and pseudo orders will not be tested.</p> <p>Half-life for first order reactions.</p> <p>Students should be able to:</p> <p>state that the half life of a first order reaction is constant;</p> <p>determine half life of a first order reaction from given data and draw the decay curve from the half life.</p>
10.3 Effects of pressure, concentration, surface area, temperature and catalysts on reaction rate. Activation energy of reactions. Maxwell-Boltzmann distribution of energies and the collision theory of reactions.	<p>Students should recognize the connection between the reaction rate, temperature and/or activation energy as represented by the Arrhenius equation.</p> <p>The equation will be provided when required and numerical calculations will not be set.</p>
10.4 Concept of rate-determining step in a multistep reaction and its relevance to the mechanism of the reaction. The molecularity of a reaction.	<p>Students should be able to:</p> <p>distinguish between order and molecularity of a reaction;</p> <p>describe examples of mechanisms with a slow step followed by a number of fast steps.</p> <p>Third order reactions and notion of the improbability of three-particle collisions.</p> <p>Mechanisms with a fast reversible first step followed by a slow step will not be examined.</p>
10.5 Photochemical reactions and free radical mechanisms.	<p>Students should be able to describe initiation, propagation and termination steps of free radical mechanisms typified by the chlorination of methane, the polymerisation of ethene and the formation of ozone in the stratosphere (an important reaction in view of the role of ozone in filtering ultraviolet radiation).</p>

- 10.6 Homogeneous and heterogeneous catalysis; examples of the use of catalysts in industrial processes; catalytic converters; autocatalysis. Students should be able to:
- distinguish between homogeneous and heterogeneous catalysts;
 - describe the role of catalysts in speeding up a reaction by providing an alternative pathway with a lower activation energy;
 - describe and explain their lack of influence on the equilibrium composition of a reversible reaction;
 - draw enthalpy level diagrams (energy profiles) for reactions with and without catalysts;
 - describe the role of the catalytic converter in changing unburnt hydrocarbons, CO and nitrogen oxides in car exhaust into less harmful products (details of different type of converters will not be tested);
 - describe the phenomenon of catalyst poisoning.

11 Chemical Periodicity	
<p>Periodic classification in terms of electronic structure. Periodic relationships amongst the elements Li to Ar.</p>	<p>Students should be able to:</p> <p>describe how properties of elements exhibit periodic behaviour as illustrated by their trends in melting and boiling points, electrical conductivity and ionisation energy; and by their reaction with oxygen, chlorine and water;</p> <p>write formulae of the oxides and describe their physical state and relate it to their structure and bonding and state the acid-base character of oxides/hydroxides of metals and oxides of non-metals for the following Na_2O, MgO, Al_2O_3, NaOH, Mg(OH)_2, Al(OH)_3, SiO_2, P_4O_{10}, SO_2, SO_3, Cl_2O;</p> <p>write formulae of chlorides (Li to N and Na to P), describe their physical state and relate it to their structure and bonding and state the reactions of chlorides with water;</p> <p>write formulae of the simple hydrides (Li to F and Na to Cl but excluding boron and aluminium hydrides), describe their physical state and relate it to their structure and bonding.</p>

12 Inorganic chemistry	
<p>Comparative study of the s-block elements (i) lithium, sodium and potassium; (ii) beryllium, magnesium, calcium, strontium and barium.</p> <p>12.1 Fixed oxidation states and electronic configuration in terms of orbitals.</p> <p>Physical properties of elements; flame tests; reaction with water and with oxygen.</p> <p>Diagonal relationships: Li and Mg; Be and Al (see also section 12.2 on Al).</p> <p>Oxides, peroxides, superoxides, hydroxides, nitrides, carbonates, hydrogencarbonates, chlorides, nitrates(V), nitrates(III), sulfates(VI) and sulfates(IV).</p> <p>Trends in thermal stability of nitrates(V) and carbonates and solubility of hydroxides and sulfates(VI).</p>	<p>Students should be able to explain the reasons for diagonal relations and give examples to illustrate the relation.</p> <p>Students should be able to:</p> <p>draw structures of these ions, including any covalent bonds and any delocalization features;</p> <p>draw the crystal lattices of NaCl and CsCl and to state the coordination number of the metal ion in these chlorides;</p> <p>describe a method of preparation of the following compounds: oxides of Li and group II metals; peroxides of Na and Ba; KO_2; $\text{Mg}(\text{OH})_2$ by precipitation reactions;</p> <p>describe the effect of heat on nitrates(V), hydrogencarbonates and carbonates;</p> <p>explain trends in solubility on the basis of hydration and lattice enthalpies.</p>
<p>12.2 Aspects of the chemistry of aluminium: hydrated aluminium ion, oxide and hydroxide; chlorides, alums, nitrates.</p>	<p>Students should be able to explain:</p> <p>the amphoteric nature of the metal and its oxide and hydroxide;</p> <p>passivation of the metal;</p> <p>structure and acidity of the hydrated ion;</p> <p>a method of preparation of anhydrous AlCl_3 (by direct union of the elements) and hydrated aluminium chloride;</p> <p>the structure of the dimeric form of gaseous Al_2Cl_6 and know that it exists in equilibrium with the monomeric form;</p> <p>alums as double salts of general formula $\text{M}^{\text{I}}\text{M}^{\text{III}}(\text{SO}_4)_2 \cdot 12\text{H}_2\text{O}$ where M^{III} could be an ion other than aluminium (preparation of alums is not required).</p>
<p>12.3 Chemistry of carbon, silicon, tin and lead.</p> <p>Allotropy of the elements</p> <p>Main oxidation states and inert pair effect</p>	<p>Students should be able to:</p> <p>describe the structure of the allotropes of carbon (graphite, diamond and fullerenes typified by C_{60}) and the properties of diamond and graphite and relation to structure and bonding;</p> <p>explain the variation in ionization energies;</p>

<p>Principal oxides</p> <p>Hydrides</p> <p>Halides: preparation and hydrolytic behaviour; formation of complex ions.</p> <p>Toxicity of lead</p>	<p>predict the relative stability of the two main oxidation states on descending the group;</p> <p>explain the redox character of simple compounds of Sn(II) and Pb(IV);</p> <p>predict the nature of the bonding in a given oxide and describe the acid/base character of oxides including the amphotericity of tin(II) and lead(II);</p> <p>state how CH_4 can be produced in the lab (by decarboxylation reaction or by hydrolysis of Al_4C_3) and how SiH_4 can be made from Mg_2Si and acid. Predict relative stability towards heat;</p> <p>describe a method of preparation of the halides (excluding fluorides) of the elements (both in the IV and II states);</p> <p>know how tetra- and dichlorides (restricted to Sn and Pb) react with water and recognize why CCl_4 does not hydrolyse;</p> <p>describe the formation of PbCl_4^{2-}, SnCl_6^{2-};</p> <p>explain the use of lead tetraethyl in petrol engines and why it has been phased out.</p>
<p>12.4 Chemistry of fluorine, chlorine, bromine and iodine.</p> <p>Laboratory preparation of chlorine, bromine and iodine and simple tests for these elements.</p> <p>Trends in physical and chemical properties for the series X_2 and HX ($\text{X} = \text{F}, \text{Cl}, \text{Br}$ and I).</p> <p>Trends in the ionisation energies and electron affinities of the halogens.</p> <p>Reactions between ionic halides X^- ($\text{X} = \text{Cl}, \text{Br}, \text{I}$) and silver ions, phosphoric(V) and sulfuric(VI) acids.</p> <p>Acidity of HX ($\text{X} = \text{F}, \text{Cl}, \text{Br}, \text{I}$) and anomalous behaviour of HF.</p> <p>Oxidation states of the halogens Cl, Br and I in their compounds and the relative oxidising strength of halogens X_2 ($\text{X} = \text{Cl}, \text{Br}, \text{I}$). Oxo anions as oxidising agents.</p>	<p>Students should be able to:</p> <p>describe one method of preparation including the practical details; recall the bleaching property of chlorine and bromine, the colour of solutions of bromine and iodine in organic solvents; the starch test for iodine;</p> <p>describe and explain the trends in boiling points and bond dissociation enthalpies;</p> <p>explain why the solubility of AgX ($\text{X} = \text{Cl}, \text{Br}, \text{I}$) is influenced in the presence of ammonia and thiosulfate and recognize the formation of the corresponding silver(I) complexes;</p> <p>explain the reaction of the halide ions with concentrated sulfuric(VI) acid in relation to the reducing properties of X^- and HX;</p> <p>explain the trend in acid strength for HX and describe and explain the hydrogen bonding in HF and in the ion HF_2^-;</p> <p>write ionic equations describing the disproportionation reactions of chlorine and chlorate(I) (OCl);</p> <p>describe the structures of chlorate(V) and chlorate(VII) ions;</p> <p>describe the effect of heat on KClO_3;</p>

	<p>describe the oxidation of iodide by chlorate(V);</p> <p>describe the reaction of iodine with thiosulfate ion and the use of this reaction in redox titrimetry.</p>
<p>Iodine in titrimetry; triiodide ion.</p>	
<p>12.5 The chemistry of the transition elements.</p> <p>Difference between transition and d-block elements.</p> <p>A general overview of the transition metals emphasising features they have in common.</p> <p>Complex ions</p> <p>Shapes of di-, tetra-, and hexa-coordinated systems; isomerism in complex ions.</p> <p>Stability of complex ions and ligand exchange.</p> <p>Stability of chelating ligands.</p> <p>Nomenclature of complex ions.</p>	<p>Students should be able to:</p> <p>identify as a 'transition element' that which forms at least one species in which the d-orbital is incompletely filled;</p> <p>give the electronic configurations of the metals Ti through to Cu (given their atomic number);</p> <p>account for their metallic character, variable oxidation states explained on the basis of the electron configuration of these elements, catalytic roles, formation of coloured compounds and formation of complex ions - explanation of colour in terms of electronic transitions between d-orbitals of different energy will not be examined;</p> <p>describe bonding in complex ions in terms of the electrostatic model or else in terms of dative covalency between ligand and central metal ion;</p> <p>describe and illustrate structural and stereoisomerism in octahedral complexes;</p> <p>write equilibrium reactions related to stepwise formation constants and predict ligand exchange from values of formation constants;</p> <p>know the bidentate ligands ethylenediamine, ethanedioate and glycinate ion and the hexadentate ligand EDTA⁴⁻ and describe how these ligands easily displace monodentate ligands to form more stable chelate complexes;</p> <p>name complex ions (using as a guide the publication <i>Chemical Nomenclature for Use in Matriculation Examinations</i>).</p>
<p>12.6 More detailed chemistry of the following d-block elements: chromium, manganese, iron and copper.</p> <p><i>Chromium:</i></p> <p>Reaction of metal with acids, including passivation with HNO₃, with alkalis to form chromate(III).</p> <p>Chromium(II) state and its strong reducing character.</p> <p>Chromium(III) compounds</p>	<p>Students should be able to :</p> <p>write relevant chemical equations to describe these changes and explain passivation;</p> <p>describe, using a balanced chemical equation, the formation of Cr(II) from Cr(VI) by reduction with Zn/H⁺;</p> <p>describe how to prepare from readily available precursors (e.g. simple salts) the oxide and hydroxide of Cr(III) and chromates(III);</p> <p>describe the conversion of Cr(III) to chromate(VI);</p> <p>write the possible isomers from the hydrates of CrCl₃ (method of preparation of isomers will not be tested);</p> <p>recognize that Cr(III) can form alums;</p>

Chromium(VI) compounds	<p>describe and illustrate the structures of the chromate(VI) and dichromate(VI) ions;</p> <p>describe the chromate(VI)-dichromate(VI) equilibrium and the effect of pH change on this equilibrium;</p> <p>describe the use of chromate(VI) in qualitative analysis as a precipitant and dichromate(VI) in titrimetry and as an oxidant in organic chemistry;</p> <p>know how ammonium dichromate(VI) decomposes on heating;</p> <p>know how CrO_3 can be made and how it can be converted into chromic(VI) acid;</p> <p>know about the toxicity of Cr(VI) compounds.</p>
<p><i>Manganese:</i></p> <p>Reactivity of metal exhibited by its reaction with dilute acids to form Mn^{2+} and hydrogen.</p> <p>Manganese(II) compounds</p> <p>Manganese(IV) and manganese(VI) compounds</p> <p>Manganese(VII) compounds</p>	<p>Students should be able to:</p> <p>describe the oxide and hydroxide of Mn(II) and explain how $\text{Mn}(\text{OH})_2$ oxidises easily to Mn(III) and Mn(IV);</p> <p>describe how manganese(II) can be oxidised to manganese(VII) using bismuthate and persulfate ion;</p> <p>recall that MnO_2 is a dark brown solid that can be prepared from Mn(II) in the lab and that it is a strong oxidising agent capable of oxidising HCl to chlorine, ethanedioate to carbon dioxide and alkylbenzenes to the corresponding aldehyde;</p> <p>describe how MnO_2 can be converted into manganate(VI) and how this species disproportionates;</p> <p>recall the tetrahedral structure of MnO_4^-;</p> <p>explain the use of KMnO_4 as a strong oxidising agent in titrimetry and in organic chemistry.</p>
<p><i>Iron:</i></p> <p>Reaction of metal with oxidising and non-oxidising acids.</p> <p>Reaction of iron with steam to produce hydrogen and Fe_3O_4 (iron(II) diiron(III) oxide)</p> <p>Reaction of iron with non-metals including C to form steels; rusting of iron</p> <p>Iron(II) compounds; oxidation to iron(III) compounds.</p> <p>Iron(III) compounds; reduction to iron(II) compounds.</p>	<p>Students should be able to:</p> <p>describe Fe_3O_4, (which occurs naturally as magnetic iron ore) as an example of a mixed oxide containing iron(II) and iron(III) ions;</p> <p>describe the rusting of iron as an electrochemical process; explain how rusting can be prevented including the use of sacrificial anodizing;</p> <p>describe methods of formation of the oxide (a very unstable, easily oxidisable substance), hydroxide, iron(II) sulfate-7-water and the double salt ammonium iron(II) sulfate-6-water;</p> <p>describe methods of oxidizing iron(II) compounds; describe methods of formation of the oxide and of anhydrous FeCl_3;</p> <p>describe the reaction of iron(III) with alkali, including aqueous ammonia;</p>

<p>Complex ions of iron</p>	<p>describe methods of reducing iron(III) compounds;</p> <p>give reasons for the non-existence of iron(III) iodide and iron(III) carbonate;</p> <p>explain the acidity of hexaaquairon(III) ion and compare with that of the corresponding iron(II) species;</p> <p>describe the formation of $[\text{Fe}(\text{CN})_6]^{4-}$ by reaction of iron(II) and cyanide and explain how it can be converted into $[\text{Fe}(\text{CN})_6]^{3-}$ by oxidation with chlorine;</p> <p>explain the use of hexacyanoferrates in qualitative inorganic analysis;</p> <p>state how $[\text{Fe}(\text{SCN})(\text{H}_2\text{O})_5]^{2+}$ is involved in qualitative analysis for Fe(III);</p> <p>describe the ion $[\text{Fe}(\text{NO})(\text{H}_2\text{O})_5]^{2+}$ as an Fe(II)-NO complex and explain how it forms during the brown ring test for nitrate(V).</p>
<p><i>Copper:</i></p> <p>Reaction with oxidising acids</p> <p>Copper(I) compounds</p> <p>Copper(II) compounds</p> <p>Complex ions of copper(I) and copper(II)</p>	<p>Students should be able to:</p> <p>write chemical equations for reactions of Cu with nitric(V) acid of various concentrations and with concentrated sulfuric(VI) acid;</p> <p>describe how Cu_2O is formed in Fehling's test for aldehydes and how this oxide can be converted into CuCl and CuI;</p> <p>describe the instability of $\text{Cu}^+(\text{aq})$ with respect to disproportionation and to explain how ligands other than water stabilize this ion;</p> <p>describe methods of preparation of the oxide, hydroxide, sulfate(VI), nitrate(V) and basic carbonates of Cu(II);</p> <p>explain the formation of complex ions of Cu(I) and Cu(II) with ammonia and chloride ions, the formation of the glycinate complex of Cu(II);</p> <p>describe the shape of Cu(II) complexes as either square planar or distorted octahedral species; tetrahedral CuCl_4^{2-};</p> <p>explain the interconversion of blue aqua copper(II) species into yellow chlorocuprate(II).</p>
<p>12.7 A more detailed study of the inorganic chemistry of the following non-metals.</p> <p><i>Hydrogen</i></p> <p>Manufacture of hydrogen</p> <p>Laboratory preparation</p> <p>Properties as reducing agent</p>	<p>Students should be able to:</p> <p>describe the manufacture of the element from natural gas by steam reformation and by electrolysis of water;</p> <p>explain methods of preparation of hydrogen on a small scale using reactions involving metal + acid or Al + alkali;</p> <p>describe reactions of hydrogen acting as a reducing agent with non-metals (oxygen, nitrogen and sulphur) and metal</p>

<p>Ionic, covalent and interstitial hydrides</p> <p>Complex hydrides: LiAlH_4 and NaBH_4</p> <p>Water and its physical properties; solvent action</p> <p>Deuterium oxide and deuteroderivatives (eg: DCl, ND_3, C_2D_2).</p>	<p>oxides as well as with organic molecules, e.g. unsaturated hydrocarbons;</p> <p>know about the old concept of 'nascent hydrogen' and its modern explanation in terms of reduction by electrons released in the metal-acid reaction;</p> <p>describe the bonding in salt-like hydrides (containing H^- ion) and covalent hydrides and explain the connection between interstitial hydrides and catalysis in hydrogenation reactions;</p> <p>explain the structure and bonding in LiAlH_4 (lithium tetrahydridoaluminate) and how it can be made in the laboratory;</p> <p>use of LiAlH_4 as a reducing agent in organic chemistry for the $>\text{C}=\text{O}$ and the $-\text{COOH}$ groups but which does not reduce the $\text{C}=\text{C}$ double bond;</p> <p>account for the physical properties of water in terms of hydrogen bonding and for its solvent properties in terms of its polar structure</p> <p>describe how simple deuteroderivatives can be synthesised using D_2O as a source of deuterium (the method of formation of deuterium oxide will not be tested).</p>
<p><i>Nitrogen</i></p> <p>Structure and bonding of N_2</p> <p>Ammonia and its reactions as base, ligand and reducing agent; ammonium salts</p> <p>Oxides of nitrogen N_2O, NO, NO_2, N_2O_4, N_2O_5</p> <p>Nitric(III) (nitrous) acid and salts</p> <p>Nitric(V) (nitric) acid and salts</p>	<p>Students should be able to:</p> <p>explain the lack of reactivity of nitrogen as a consequence of the strength of the triple bond;</p> <p>describe the properties of ammonia as a weak Lowry-Brønsted base, a good monodentate ligand (with Cu^{2+}, Zn^{2+}, Ni^{2+} in qualitative analysis) and a reducing agent;</p> <p>explain the thermal instability of ammonium nitrate(III) and ammonium nitrate(V) as resulting from the redox reaction between the reducing cation and the oxidising anion;</p> <p>describe a laboratory preparation of these oxides of nitrogen;</p> <p>draw electronic structures to show the bonding in these oxides and to recognise the presence of odd electron systems;</p> <p>draw molecular shapes of these oxides;</p> <p>describe the disproportionation of NO_2 when dissolved in water or alkali;</p> <p>describe the laboratory preparation of nitric(III) acid and nitrates(III) of s-block elements;</p> <p>describe the laboratory preparation of nitric(V) acid by reaction of an involatile acid on nitrates(V);</p> <p>describe reactions of nitric(V) acid acting as an acid, an oxidising agent and a nitrating agent in organic chemistry;</p>

	describe the reduction of nitrate(III) and nitrate(V) to ammonia using 'nascent hydrogen' type systems and the brown ring test for nitrate(V).
<p><i>Oxygen</i></p> <p>Allotropes of oxygen: O₂ and O₃ (trioxygen or ozone)</p> <p>Laboratory preparation of oxygen</p> <p>Reaction of oxygen with metals and non-metals</p> <p>Hydrogen peroxide and peroxides</p> <p>Superoxide ion</p>	<p>Students should be able to:</p> <p>state that oxygen is a main component of the atmosphere;</p> <p>test for oxygen in the lab;</p> <p>explain how trioxygen forms in the stratosphere from photochemical reactions and its action as a screen of ultraviolet radiation;</p> <p>state that trioxygen forms in the lower atmosphere as an air pollutant (chemistry of formation will not be tested);</p> <p>describe the preparation of oxygen gas from oxides and peroxides;</p> <p>describe the formation of compounds of oxygen by direct union between oxygen and various elements and to classify oxides as acidic, basic, amphoteric or neutral;</p> <p>distinguish between normal and mixed oxides (typified by Fe₃O₄ and Pb₃O₄) in terms of the species present in these compounds;</p> <p>describe a laboratory preparation of hydrogen peroxide from barium peroxide;</p> <p>describe the main properties of hydrogen peroxide namely, as oxidising agent, as a reducing agent with oxidants stronger than itself (MnO₄⁻ and Cr₂O₇²⁻), and as a thermally unstable liquid;</p> <p>recognize the relatively weak nature of the peroxide bond and the ease of its homolysis to form free radicals;</p> <p>describe the structure and bonding of the peroxide ion in sodium and barium peroxide (as typical compounds) and its reaction with water, acid and with carbon dioxide;</p> <p>describe the structure and bonding of the superoxide ion in KO₂ and its reaction with water.</p>
<p><i>Sulfur</i></p> <p>Allotropes of sulfur (S₈): rhombic and monoclinic sulfur</p> <p>Hydrogen sulfide</p> <p>Sulfur dioxide and sulfuric(IV) (sulfurous) acid, sulfates(IV) (sulfites)</p>	<p>Students should be able to:</p> <p>recognize the effect of temperature on the stability of the two allotropes;</p> <p>recognize the odour and toxicity of hydrogen sulfide;</p> <p>describe how H₂S can be prepared from sulfides and explain its properties as a weak dibasic acid, a reducing agent and as a precipitant for insoluble sulphides;</p> <p>describe how sulfur dioxide, sulfuric(IV) and sulfates(IV) can be made and interconverted in the laboratory and explain their reducing character;</p>

Sulfur trioxide and sulfuric(VI) (sulfuric) acid	describe how sulfur trioxide, sulfuric(VI) and sulfates(VI) can be made and interconverted in the laboratory;
Thiosulfate	describe reactions of concentrated sulfuric(VI) acid as an involatile proton donor capable of displacing more volatile acids from their salts, oxidising agent, strong dehydrating agent and sulfonating agent in organic chemistry;
Structure of oxoanions of sulfur	describe the formation of thiosulfate ion from sulfate(IV) and its reaction with acids, iodine and chlorine;
Sulfur oxides as environmental pollutants	draw the electronic structures and shapes of the ions: sulfite, sulfate and thiosulfate and recognize the presence of electronic delocalization in these structures; recognize the fact that the two sulfur atoms in thiosulfate are present in different structural environments and the implication of this on the oxidation state; explain why sulfur oxides are often present in the atmosphere as contaminants, their origin in fossil fuels and their role in the formation of acid rain and its environmental impact.

<p>13.1 Organic chemistry: general principles</p>	
<p>13.1 Functional groups and homologous series. molecular, empirical and structural formulae including graphical representations</p> <p>Nomenclature of organic compounds</p>	<p>Students should be able to:</p> <p>distinguish between molecular and empirical formulae and write possible structural formulae from molecular formulae;</p> <p>identify functional groups in the following classes: alkenes, alkynes, arenes, , alcohols, ethers, phenols, halogenoalkanes and halogenoarenes, aldehydes, ketones, carboxylic acids and derivatives, amines and diazonium compounds;</p> <p>recognize as structural formulae systems such as $\text{CH}_3\text{CH}_2\text{CH}_3$ or CH_3CHO as well as graphical (displayed) formulae where individual bonds are shown explicitly;</p> <p>use wedge and dashed/hatched line convention to represent three dimensional structure;</p> <p>give systematic names of organic compounds from their structure and vice versa (refer to publication <i>Chemical Nomenclature for Use in Matriculation Examinations</i>).</p>
<p>13.2 Purification of compounds</p> <p>Preparation of derivatives and their use for characterisation and purification</p> <p>Determination of melting points as a test for purity</p> <p>Mixed melting point technique</p>	<p>Students should be able to:</p> <p>describe the following techniques used in the purification of organic compounds: solvent extraction, recrystallisation, drying (for gases, liquids and solids), simple, fractional and steam distillation, sublimation; column, and thin layer chromatography (one dimensional);</p> <p>describe the process of derivatisation for characterization purposes (in aldehydes, ketones and alcohols) and for purification purposes (hydrogensulfite addition compounds for aldehydes);</p> <p>recognize that presence of impurities cause a melting point to deviate from the standard value;</p> <p>describe the mixed melting point technique as a method of identification of compounds.</p>
<p>13.3 The determination of empirical, molecular and structural formulae from analytical information.</p>	<p>Experimental techniques are not required but students should be able to use percentage composition data to deduce the empirical formula and, through relative molecular mass, obtain the molecular formula.</p>
<p>13.4 Structure deduced from chemical and physical properties and from information derived from instrumental techniques, namely, mass spectrometry and infra red spectroscopy.</p>	<p>Students should be able to:</p> <p>deduce the presence of specific functional groups from results of reactions with reagents;</p> <p>interpret infra red information (including spectra) for the identification of simple functional groups using given tables of frequencies/wave numbers (details of instrumentation are not required).</p>

13.5 Catenation in carbon compounds and spatial distribution of the bonds in simple carbon compounds.	<p>Students should be able to:</p> <p>explain why the element carbon is unique in exhibiting catenation;</p> <p>explain the shape of the bonds and bond angles in alkanes (tetrahedral), alkenes and carbonyl groups (planar) and alkynes (linear) in terms of hybridized sp^3, sp^2 and sp orbitals.</p>
13.6 Isomerism in organic compounds	<p>Students should be able to:</p> <p>describe structural and stereoisomerism namely: cis-trans (geometrical) restricted to C=C and C=N systems; optical isomerism (restricted to enantiomerism) and conformational isomerism (restricted to cyclohexane);</p> <p>describe the property of optical activity in terms of dextrorotatory (+) and laevorotatory (-) compounds and connect it to chiral molecules being those which do not possess a plane of symmetry (including systems which possess two chiral carbons but still having a plane of symmetry), use of the descriptors D and L and R and S are not required;</p> <p>explain the term racemate/racemic mixture of optical isomers and distinguish between internal and external compensation (resolution of racemic mixture will not be tested);</p> <p>describe the phenomenon of keto-enol tautomerism and be able to predict its occurrence in structures that may be novel to them;</p> <p>explain how size, shape and polarity of molecular isomers can influence physical properties including comparison of volatility of linear and branched isomers;</p> <p>describe the effect of intramolecular versus intermolecular hydrogen bonding on boiling point as typified by the nitrophenols;</p>
13.7 Delocalisation of electrons in organic molecules	<p>Students should be able to:</p> <p>describe the benzene ring in terms of sigma (σ) and pi (π) bonds and to recognize the significance of delocalization in this and similar molecules;</p> <p>predict the presence of delocalisation in conjugated systems and know that delocalization lowers the energy of the system and is thus a favourable structural feature;</p> <p>use canonical formulae, including Kekule structures for aromatic molecules, for the description of delocalization in organic molecules.</p>

14 Chemistry of Aliphatic and Aromatic Compounds	
<p>14.1 Petroleum: hydrocarbons from petroleum</p> <p>Use of hydrocarbons as fuels</p> <p>Environmental considerations</p>	<p>Students should be able:</p> <p>to recall that alkanes, alkenes and aromatic hydrocarbons can be obtained from petroleum by fractional distillation;</p> <p>to describe the principle of the processes of cracking and reforming (no technical details of the processes are expected);</p> <p>to explain how the combustion reactions of alkanes and other hydrocarbons lead to their use as fuels in industry, in the home and in transport;</p> <p>to explain how the combustion of hydrocarbons leads to the formation of carbon dioxide, an important greenhouse gas;</p> <p>to explain the formation of carbon monoxide, oxides of nitrogen and unburnt hydrocarbons arising from the internal combustion engine and the use of catalysts to minimise their effect on pollution.</p>
<p>14.2 Alkanes, alkenes and alkynes</p> <p>Laboratory preparation of alkanes</p> <p>Alkanes and cycloalkanes: reactivity and reactions</p> <p>Concept of ring strain</p> <p>Alkenes: alkenes as examples of unsaturated hydrocarbons</p> <p>Preparation of alkenes</p> <p>Addition reactions of alkenes</p> <p>Markovnikov's rule</p>	<p>Students should:</p> <p>know that alkanes can be obtained from a number of organic compounds as indicated elsewhere in this syllabus;</p> <p>be aware of the general unreactivity of alkanes, including towards polar reagents;</p> <p>be able to describe by appropriate chemical equations free radical substitution by chlorine and bromine;</p> <p>be able to describe how ring strain in molecules, such as cyclopropane, leads to enhanced reactivity relative to corresponding non-cyclic alkanes.</p> <p>Students should be able to recognise:</p> <p>alkenes as unsaturated hydrocarbons containing a double bond made up of a σ and π bond;</p> <p>the double bond as an electron rich centre susceptible to electrophilic attack.</p> <p>Students should be able to describe the preparation of alkenes from alcohols and haloalkanes through elimination reactions.</p> <p>Students should be able to:</p> <p>describe the reaction of alkenes with halogen, hydrogen, halogen halides, and water in presence of sulfuric(VI) acid;</p> <p>identify the major product in the reaction of halogen halides with an unsymmetrical alkene and explain its formation in terms of relative carbocation stability;</p>

<p>Ozonolysis as the addition reaction of ozone to a double bond followed by hydrolysis in the presence of zinc/ethanoic acid</p> <p>Oxidation of the double bond</p> <p>Addition polymerisation</p> <p>Hydrocarbon poly(alkene)s as waste products</p> <p>Alkynes</p> <p>Preparation of alkynes</p> <p>Addition reactions of alkynes</p> <p>Combustion</p> <p>Acidity of terminal alkynes</p>	<p>draw the structure of the unstable ozonide;</p> <p>describe the reaction as producing carbonyl compounds as products of hydrolysis of the ozonide;</p> <p>describe the use of the reaction for the location of double bond;</p> <p>describe the reaction with cold, dilute, alkaline manganate(VII) ions to form the diol;</p> <p>describe the following characteristics of addition polymerisation such as that involved in formation of poly(ethene) and PVC: deduce the repeat unit of a polymer obtained from a given monomer;</p> <p>identify the monomer present in a given section of a polymer molecule;</p> <p>recognise the use of poly(alkene)s waste in recycling and as a fuel and also the nonbiodegradability of these products;</p> <p>recognise alkynes as unsaturated hydrocarbons containing a triple bond made up of a σ and two π bonds;</p> <p>describe the preparation of alkynes by dehydrohalogenation reactions;</p> <p>describe the preparation of ethyne from CaC_2;</p> <p>state the reactions with hydrogen, X_2 and HX to form partial or fully saturated products and with water to form carbonyl compounds via tautomerisation of the enol (students should recognize the importance of chloroethene as a feedstock for PVC manufacture);</p> <p>describe the exothermic reaction of ethyne with oxygen and its use in oxy-acetylene flame;</p> <p>describe the reaction of ethyne and propyne with ammoniacal silver nitrate and copper(I) chloride.</p>
<p>14.3 Arenes: benzene and alkylbenzenes</p> <p>Electrophilic substitution</p> <p>Addition reactions</p>	<p>Students should:</p> <p>be able to describe the structure and bonding in benzene and similar aromatic systems;</p> <p>be able to describe how to introduce the nitro, sulfonate, alkyl, alkanoyl and the halogen group on the aromatic ring;</p> <p>be familiar with the concept of activating and deactivating groups and their directing influence on further substitution of the aromatic ring. Such groups are to be exemplified: $-\text{CH}_3$, $-\text{OH}$, $-\text{NH}_2$, $-\text{X}$, $-\text{CHO}$, $-\text{COOH}$;</p> <p>be able to describe the addition of hydrogen and chlorine to benzene under appropriate conditions;</p>

<p>Side chain reactions of methylbenzene:</p> <p>(i) halogenation</p> <p>(ii) oxidation</p>	<p>describe the free radical substitution of the methyl side chain with chlorine or bromine;</p> <p>be able to convert the $-CH_3$ side chain to $-CHO$ and $-COOH$ using chromium(VI) dichloride dioxide and manganate(VII) respectively;</p> <p>recognise the fact that any alkyl group is oxidised to $-COOH$ upon oxidation with manganate(VII).</p>
<p>14.4 Alcohols, ethers, phenols</p> <p>Classification of alcohols</p> <p>Physical properties of alcohols</p> <p>Preparation of alcohols</p> <p>Reactions of alcohols</p> <p>Aromatic alcohols</p> <p>Polyhydric alcohols</p> <p>Industrial preparation of ethanol</p> <p>Ethers</p> <p>Phenols</p> <p>Laboratory preparation:</p> <p>Acidity of phenols:</p> <p>Reactions of phenol, including substitution reactions</p>	<p>Students should be able to:</p> <p>identify a given alcohol as primary, secondary or tertiary and distinguish between classes in terms of oxidation and triiodomethane reaction;</p> <p>know about the Lucas test and recognize that it can easily distinguish tertiary alcohols from the other two classes, (the test has dubious validity in distinguishing secondary from primary alcohols and for this reason it will not be employed in the practical examination for this purpose);</p> <p>explain the relatively high boiling point and miscibility with water of lower members in terms of hydrogen bonding in comparison to ethers;</p> <p>describe how to prepare alcohols from halogenoalkanes, alkenes, by reduction of aldehydes, ketones and carboxylic acids, from Grignard reagents;</p> <p>describe the formation of alkoxides, halogenoalkanes, esters, alkenes and ethers from alcohols;</p> <p>distinguish aromatic alcohols, typified by phenylmethanol, from phenols;</p> <p>recall ethane-1,2-diol and propane-1,2,3-triol as typical polyhydric alcohols;</p> <p>describe the formation of ethanol by the fermentation of sugars and by steam hydration of ethane;</p> <p>recall the preparation of ethers from alkoxides and halogenoalkanes;</p> <p>describe the effect of HI on ethers;</p> <p>recognise the use of ethoxyethane as a solvent and in extraction procedures despite its inflammable nature;</p> <p>describe the preparation of phenol from sulfonic acids and diazonium salts;</p> <p>recognise the weak acid nature of phenol by its failure to liberate CO_2 with sodium carbonate and sodium hydrogencarbonate (contrast with carboxylic acids);</p> <p>describe the formation of esters and ethers from phenols;</p> <p>describe substitution reactions in the ring to include tribromination, nitration;</p>

		<p>describe the reduction to benzene and the coupling reactions with diazonium ions to form azo dyes;</p> <p>recall that phenol and related compounds give a characteristic violet colouration with neutral iron(III) chloride which serves as a diagnostic test for phenols.</p>
14.5	<p>Halogenoalkanes and halogenoarenes</p> <p>Preparation</p> <p>Reactions</p> <p>Effect of structure on chemical properties</p> <p>Halogenoarenes</p> <p>Preparation</p> <p>Effect of structure on chemical properties</p> <p>(Chloromethyl)benzene</p> <p>Reaction with magnesium</p> <p>Uses of organohalogen compounds other than in synthetic organic chemistry</p>	<p>Students should be able to:</p> <p>describe the preparation of halogenoalkanes from alcohols and alkenes;</p> <p>describe the conversion of halogenoalkanes into alcohols, ethers, amines, nitriles, esters, alkanes, alkenes, alkynes, Grignard reagents and alkylarenes;</p> <p>explain the reactivity of halogenoalkanes in terms of their structure (primary, secondary, tertiary halogenoalkanes) and the halogen atom (Cl, Br, I);</p> <p>describe the preparation by direct halogenation (where appropriate) and via diazonium compounds;</p> <p>explain the unreactivity of the substituent halogen atom with respect to replacement reactions in terms of delocalization of electrons;</p> <p>explain why (chloromethyl)benzene is more susceptible to replacement reactions than nuclear halogenoarenes;</p> <p>recall the formation of Grignard reagents;</p> <p>recall the use as solvents; explain the impact on the ozone layer caused by the use of CFCs and other halogeno compounds.</p>
14.6	<p>Aldehydes and ketones</p> <p>Preparation</p> <p>Addition reactions of the carbonyl group</p> <p>Condensation reactions</p>	<p>Students should be able to:</p> <p>describe the preparation of aldehydes and ketones from corresponding alcohols;</p> <p>describe the preparation of aromatic aldehydes by oxidation of methylarenes using chromium(VI) dichloride dioxide and the preparation of aromatic ketones by Friedel-Crafts reaction;</p> <p>describe the addition reactions with H^+ (as in the hydridometallates); ROH; HCN; and $NaHSO_3$ (including the use of this reagent in the purification of aldehydes);</p> <p>describe the reactions with hydroxylamine; hydrazine, phenylhydrazine and 2,4-dinitrophenylhydrazine;</p> <p>explain the use of 2,4-dinitrophenyl-hydrazones for characterisation and identification purposes;</p> <p>describe the formation of condensation polymers from methanol;</p> <p>recall the reactions with halogen and with PCl_5; the haloform reaction; the aldol reaction; the Cannizzaro</p>

Other reactions	<p>reaction; the reduction using zinc and acid and with hydrogen;</p> <p>distinguish aldehydes and ketones on the basis of oxidation employing weak oxidant systems typified by Fehling reagent and Tollen's reagent.</p>
<p>14.7 Carboxylic acid</p> <p>Formation of carboxylic acids:</p> <p>Weak acidity</p> <p>Effect on acidity of α-substitution by halogen atoms or alkyl groups</p> <p>Salt formation and alkalinity of carboxylate anion</p> <p>Soaps</p> <p>Decarboxylation reaction</p> <p>Reduction to primary alcohols</p> <p>Exceptional behaviour of methanoic acid</p> <p>Formation of the following derivatives of carboxylic acids: esters, acid chlorides and acid anhydrides (restricted to simple anhydrides), acid amides and nitriles (but excluding isonitriles).</p>	<p>Students should be able to:</p> <p>describe the formation of carboxylic acids from alcohols and aldehydes by oxidation and from carboxylic acid derivatives by hydrolysis;</p> <p>formation from alkylbenzenes by oxidation of the side chain;</p> <p>explain the weak protic acidity of carboxylic acids in terms of their structure and relative stability of the carboxylate anion;</p> <p>account for the increase or decrease in acidity of halogen, or alkyl substituted carboxylic acids in terms of inductive effect;</p> <p>compare the acidity of ethanoic acid with that of ethanol and phenol and account for differences in terms of structure;</p> <p>describe methods of formation of carboxylate salts and explain the alkalinity through hydrolysis reactions;</p> <p>recall that sodium or potassium carboxylates having long alkyl chains can act as soaps;</p> <p>be familiar with the use of soda lime to bring about decarboxylation of carboxylic acid salts and to represent the equation involved in terms of reaction with "NaOH";</p> <p>describe the conversion of $-\text{COOH}$ to $-\text{CH}_2\text{OH}$ using LiAlH_4;</p> <p>recall the anomalous behaviour of methanoic acid and its salts as typified by reaction with soda lime, dehydration, reaction with PCl_5 and oxidising agents;</p> <p>recall that esters can be made using the reaction of acids with alcohols; acid chlorides or anhydrides with alcohols or phenols;</p> <p>recall the formation of esters from reaction of halogenoalkane with silver carboxylate;</p> <p>recall the formation of acid chlorides from reaction of carboxylic acids with phosphorus chlorides and SOCl_2 (sulfur oxide dichloride) and acid anhydrides from the acid chloride;</p> <p>describe the formation of carboxylic acid amides, $-\text{CONH}_2$, by thermal decomposition of ammonium carboxylates and reaction of esters, acid chlorides and anhydrides with ammonia;</p> <p>describe the reactions of acid chlorides and anhydrides</p>

<p>Reactions of derivatives</p> <p>Halogenoacids</p> <p>Polyesters and polyamides</p> <p>Dicarboxylic acids:</p>	<p>with primary and secondary amines to form corresponding N-substituted amides;</p> <p>explain the formation of nitriles by dehydration of amides using phosphorus(V) oxide;</p> <p>recall that nitriles can also be made from reaction of halogenoalkanes and KCN in ethanol solution and recognise this reaction as a step-up technique in organic synthesis;</p> <p>recall that aromatic nitriles can be formed from diazonium salts;</p> <p>recall the acid and alkaline hydrolysis of the various derivatives;</p> <p>explain the relative ease of hydrolysis of acyl chlorides and halogenoalkanes ;</p> <p>compare ammonolysis reactions with hydrolysis reactions;</p> <p>explain the reactions of acid chlorides and anhydrides with alcohols;</p> <p>explain the Hofmann degradation of amide and appreciate its importance in step down synthesis;</p> <p>recall the reduction of acid amides and nitriles to amines;</p> <p>recall the reduction of esters to alcohols (reduction of acid chlorides and anhydrides will not be tested);</p> <p>recall the reaction of amides with nitric(III) acid;</p> <p>explain the formation of halogenoacids by reaction of the acid and halogen;</p> <p>explain how amino acids can be prepared from halogenoacids (practical details of the isolation of the amino acid will not be tested);</p> <p>explain the formation of polyesters and polyamides and state their major commercial use;</p> <p>explain the environmental advantage of the ease of hydrolysis of both the ester and peptide link and contrast this with the properties of poly(alkene)s;</p> <p>recall the use of ethanedioic acid and its salts as standard reducing agents in titrimetry;</p> <p>explain the formation of the acid anhydride from dicarboxylic acids that allow it on the basis of stereochemistry.</p>
<p>14.8 Amines and diazonium compounds</p> <p>Structure</p>	<p>Students should be able to:</p> <p>identify a given amine as primary, secondary or tertiary;</p> <p>identify quaternary ammonium compounds;</p>

<p>Formation of amines</p> <p>Basicity of amines</p> <p>Reactions of amines</p> <p>Diazonium salts</p> <p>Formation of diazonium compounds</p> <p>Reactions of diazonium compounds: (i) substitution reactions in which the diazo group is lost (ii) reaction in which the diazo group is retained</p>	<p>prepare amines by Hofmann degradation, reduction of simple and N-substituted amides, reduction of nitriles and hydrolysis of N-substituted amides;</p> <p>prepare aromatic amines by reduction of nitroarenes by tin and concentrated hydrochloric acid;</p> <p>explain the relative basic strength of ammonia, primary and secondary aliphatic amines and phenylamine in terms of their structure;</p> <p>explain the difference in basic strength between amines and acid amides;</p> <p>give the formation of alkylammonium salts from corresponding amines;</p> <p>explain the alkylation reaction (Hofmann reaction) and acylation of amines using acid chlorides and anhydrides;</p> <p>recognise phenylamine as an activated ring and use its reaction with aqueous bromine as an example of this reactivity;</p> <p>explain the reaction of nitric(III) acid with primary aliphatic amines to produce nitrogen as a product of decomposition of unstable aliphatic diazonium compounds and recognize that this reaction is useful as a test for primary aliphatic amines.</p> <p>Students should be able to:</p> <p>draw structures of diazonium salts and recognize their ionic character;</p> <p>describe the preparation of aryl diazonium salts by reaction with dilute hydrochloric acid and sodium nitrate(III) in ice cold conditions, account for the relative stability of the the diazonium ion on the basis of delocalization of charge;</p> <p>convert the diazonium salt to phenols, halogenoarenes, arenes and nitriles;</p> <p>give coupling reactions with phenols to form azo dyes.</p>
<p>14.9 Amino acids</p> <p>Properties of amino acids</p>	<p>Students should be able to:</p> <p>recognize amino acids as typical of difunctional molecules which are important in biological systems;</p> <p>predict the reactions of amino acids as those pertaining to the carboxyl and amino groups including esterification, acylation, etc.;</p> <p>know about the formation of dipolar ions (zwitterions) in amino acids and the influence on physical properties such as solubility in water and organic solvents and melting points;</p> <p>explain the existence of optical activity of amino acids;</p>

Preparation of amino acids	<p>describe the effect of varying pH on the composition of an aqueous solution of an amino acid: isoelectric point; (technique of electrophoresis shall not be examined);</p> <p>explain the reaction of amino acid salts as bidentate ligands as typified by the copper(II) complexes;</p> <p>describe a practical method of separation of a mixture of amino acids using chromatography;</p> <p>relate amino acids to peptides and proteins;</p> <p>describe the synthesis of amino acids from carboxylic acids or aldehydes via cyanohydrins.</p>
<p>14.10 Mechanistic aspects</p> <p>Ionic mechanisms</p> <p>Markovnikov's Rule</p> <p>Homolytic mechanisms</p>	<p>Students should be able to:</p> <p>use reaction mechanisms to rationalise the facts of organic chemistry;</p> <p>recall the following mechanistic concepts: nucleophile, electrophile, free radical; inductive and mesomeric effects;</p> <p>distinguish between bimolecular and unimolecular nucleophilic substitution reactions of halogenoalkanes;</p> <p>describe electrophilic substitution of arenes (limited to nitration, alkylation, acylation and chlorination);</p> <p>describe nucleophilic addition to carbonyl group typified by the reactions with aldehydes and ketones of HCN, hydroxylamine and the hydrazines;</p> <p>describe electrophilic addition to carbon-carbon double bond as typified by reaction with alkenes of halogens and hydrogen halides;</p> <p>explain the basis of Markovnikov's Rule in terms of the stability of carbocations;</p> <p>describe the homolytic reactions involved in the halogenation of alkanes and methylbenzene and in the polymerisation of alkenes.</p>
14.11 Organic laboratory preparations	<p>Students are expected to be familiar with detailed laboratory preparation and isolation of pure samples of the following organic substances:</p> <p>(a) a bromoalkane from the corresponding alcohol;</p> <p>(b) nitrobenzene from benzene;</p> <p>(c) ethanal from ethanol;</p> <p>(d) an ester from the reaction of a carboxylic acid and an alcohol;</p> <p>(e) phenylamine from nitrobenzene.</p>

<p>15 Practical Examination</p>	
<p>15.1 The examination will primarily attempt to test practical skills.</p>	<p>The practical examination will seek to test the ability of candidates to:</p> <p>(a) manipulate chemicals and simple apparatus in quantitative and qualitative exercises;</p> <p>(b) observe and record results of experimental work;</p> <p>(c) interpret these observations and deduce correct inferences and conclusions based both on qualitative and quantitative data.</p>
<p>15.2 Quantitative exercises including the measurement of mass, volume, temperature and time may be set.</p> <p>The following may be set:</p> <p>volumetric analysis involving acids and alkalis,</p> <p>redox titrations,</p> <p>complexometric titration involving EDTA.</p>	<p>Indicator colour changes will be specified except for colour changes involving phenolphthalein and methyl orange.</p> <p>Redox titrations may involve iodimetry (experiments involving the formation of iodine from the reaction of iodide with an oxidizing agent followed by titration with thiosulfate) or titration involving permanganate and a reducing agent.</p> <p>In such titrations, the colour change of the indicator will be specified.</p> <p>Titration with silver nitrate will not be set.</p>
<p>15.3 Qualitative exercises involving observations of reactions and requiring deductions on the chemical nature of the substances will also be set. These will involve both organic and inorganic materials.</p> <p>It will be assumed that candidates are familiar with the following species from which unknown compounds will be set for qualitative analysis:</p> <p>(i) ammonium, potassium, sodium, silver, calcium, magnesium, strontium, barium, aluminium, lead(II), zinc, chromium(III), manganese(II), manganese(IV), iron(II), iron(III), cobalt(II), nickel(II), copper(I), copper(II);</p> <p>(ii) oxide, hydroxide, carbonate, hydrogencarbonate, nitrate(V), nitrate(III), carboxylate, sulfate(IV), sulfate(VI), thiosulfate, chloride, bromide, iodide, chromate(VI), dichromate(VI), phosphate, manganate(VII);</p> <p>(iii) the more common metals (e.g. Zn, Al, Cu, Fe)</p>	<p>A knowledge of classical systematic analysis will not be required and will not be tested.</p> <p>Binary mixtures involving salts, oxides and metals may be set for qualitative analysis.</p> <p>Students should be able to test for and recognize by physical or chemical means the following gases: ammonia, carbon dioxide, carbon monoxide, hydrogen, oxygen, sulfur dioxide, hydrogen halides (HX, X = Cl, Br, I), the halogens (Cl, Br, I), nitrogen dioxide, sulfur dioxide, hydrogen sulfide and nitrogen (by elimination).</p> <p>Students should be aware of the toxicity of a number of these species.</p>

<p>Organic substances containing the following functional groups may be set as unknown compounds in qualitative analysis:</p> <p>Double and triple bonds (alkenes, alkynes), hydroxyl (alcohols and phenols), carbonyl (aldehydes and ketones) carboxyl, carboxyl derivatives (carboxylate salts, esters, amides, nitriles) and amino.</p>	<p>Both aliphatic and aromatic compounds may be set.</p> <p>A combination of functional groups may be set, e.g. amino acids, hydroxyacids etc.</p>
<p>15.4 Candidates are instructed to use labcoats and to wear eye protection (safety goggles) during the practical examination and to avoid wearing clothes or articles that increase the risk of accident in the laboratory.</p>	<p>Candidates will not be allowed into the examination room unless equipped with proper safety gear.</p>

Recommended Texts

Any of the following are suitable textbooks for the syllabus; the list is not comprehensive.

1. Lister, T and Renshaw, J, Understanding chemistry for Advanced Level, 4th Edition, Nelson Thornes Ltd., 1999.
2. Ramsden, EN, A-Level Chemistry, 4th Edition, Nelson Thornes Ltd., 2000.
3. Jones, L and Atkins, P, Chemistry: molecules, matter and change, 4th Edition, W H Freeman, 1999.
4. Hill, G and Holman, J, Chemistry in context, 6th Edition, Nelson Thornes Ltd., 2011.