3.1.1 Atomic Structure

I should be able to...	Revised	Revised	Revised
determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge.			
explain the existence of isotopes.			
describe the principles of a simple time of flight (TOF) mass spectrometer, limited to ionisation, acceleration to give all ions constant kinetic energy, ion drift, ion detection and data analysis.			
interpret simple mass spectra of elements.			
calculate relative atomic mass from isotopic abundance, limited to mononuclear ions.			
deduce electron configurations of atoms and ions up to Z = 36 in terms of shells and sub–shells (orbitals) s, p and d.			
define first ionisation energy.			
write equations for first and successive ionisation energies.			
explain how first and successive ionisation energies in Period 3 (Na–Ar) and in Group 2 (Be–Ba) give evidence for electron configuration in sub–shells and in shells.			

Notes

3.1.2 Amount of Substance

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
define relative atomic mass (A_r).			
define relative molecular mass (M_r).			
carry out calculations using the Avogadro constant.			
carry out calculations using mass of substance, M_r, and amount in moles.			
carry out calculations using concentration, volume and amount of substance in a solution.			
use the ideal gas equation in calculations.			
calculate empirical formula from data giving composition by mass or percentage by mass.			
calculate molecular formula from the empirical formula and relative molecular mass.			
write balanced equations for the reactions studied.			
balance equations for unfamiliar reactions when reactants and products are specified.			
use balanced equations to calculate; masses; volumes of gases; percentage yields; percentage atom economies; concentrations and volumes for reactions in solutions.			
(Required Practical 1) describe how to make up a volumetric solution and carry out a simple acid–base titration.			

Notes

3.1.3 Bonding

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
predict the charge on a simple ion using the position of the element in the Periodic Table.			
recall formulas of compound ions, e.g. sulfate, hydroxide, nitrate, carbonate and ammonium.			
construct formulas for ionic compounds.			
represent a covalent bond using a line and a coordinate bond using an arrow.			
relate the melting point and conductivity of materials to the type of structure and the bonding present.			
explain the energy changes associated with changes of state.			
draw diagrams to represent the four crystal structures involving specified numbers of particles.			
explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs of electrons) surrounding the central atom.			
use partial charges to show that a bond is polar.			
explain why some molecules with polar bonds do not have a permanent dipole.			
explain the existence of permanent dipole–dipole forces, induced dipole–dipole (van der Waals, dispersion, London) forces and hydrogen bonding between familiar and unfamiliar molecules.			
explain how melting and boiling points are influenced by intermolecular forces.			

Notes

3.1.4 Energetics

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
define standard enthalpy of combustion ($\Delta_c H^\theta$).			
define standard enthalpy of formation ($\Delta_f H^\theta$).			
use the equation $q = mc\Delta T$ to calculate the molar enthalpy change for a reaction.			
use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.			
define the term mean bond enthalpy.			
use mean bond enthalpies to calculate an approximate value of ΔH for reactions in the gaseous phase.			
explain why values from mean bond enthalpy calculations differ from those determined using Hess's law.			
(Required Practical 2) describe a method to measure an enthalpy change.			

Notes

3.1.6 Chemical Equilibria, Le Chatelier's Principle and K_c

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium.			
explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used.			
construct an expression for K_c for a homogeneous system in equilibrium.			
calculate a value for K_c from the equilibrium concentrations for a homogeneous system at constant temperature.			
perform calculations involving K_c.			
predict the qualitative effects of changes of temperature on the value of K_c.			

Notes

3.1.7 Oxidation, Reduction and Redox Equations

I should be able to...	Revised	Revised	Revised
work out the oxidation state of an element in a compound or ion from the formula.			
write half-equations identifying the oxidation and reduction processes in redox reactions.			
combine half-equations to give an overall redox equation.			

Notes

3.1.8 Thermodynamics

I should be able to...	Revised	Revised	Revised
define each of the following terms: • enthalpy of formation • ionisation energy • enthalpy of atomisation • bond enthalpy • electron affinity • lattice enthalpy			
construct Born–Haber cycles to calculate lattice enthalpies using these enthalpy changes.			
construct Born–Haber cycles to calculate one of the other enthalpy changes.			
compare lattice enthalpies from Born–Haber cycles with those from calculations based on a perfect ionic model to provide evidence for covalent character in ionic compounds.			
define the term enthalpy of hydration.			
calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration.			
calculate entropy changes from absolute entropy values.			
use the relationship $\Delta G = \Delta H - T\Delta S$ to determine how ΔG varies with temperature.			
use the relationship $\Delta G = \Delta H - T\Delta S$ to determine the temperature at which a reaction becomes feasible.			

Notes

3.1.10 Equilibrium Constant K_p for Homogeneous Systems

I should be able to...	Revised	Revised	Revised
derive partial pressure from mole fraction and total pressure.			
construct an expression for K_p for a homogeneous system in equilibrium.			
perform calculations involving K_p.			
predict the qualitative effects of changes in temperature and pressure on the position of equilibrium.			
predict the qualitative effects of changes in temperature on the value of K_p.			
understand that, whilst a catalyst can affect the rate of attainment of an equilibrium, it does not affect the value of the equilibrium constant.			

Notes

3.1.11 Electrode Potentials and Electrochemical Cells

I should be able to...	Revised	Revised	Revised
use E^{θ} values to predict the direction of simple redox reactions.			
calculate the EMF of a cell.			
write and apply the conventional representation of a cell.			
use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells.			
deduce the EMF of a cell.			
explain how the electrode reactions of a lithium–ion cell can be used to generate an electric current.			
explain how the electrode reactions of an alkaline hydrogen–oxygen fuel cell can be used to generate an electric current.			
(Require Practical 8) measure the EMF of an electrochemical cell.			

Notes

3.1.12 Acids and Bases

I should be able to...	Revised	Revised	Revised
define the term 'Bronstead-Lowery acid'.			
define the term 'Bronstead-Lowery base'.			
convert concentration of hydrogen ions into pH and vice versa.			
calculate the pH of a solution of a strong acid from its concentration.			
use K_w to calculate the pH of a strong base from its concentration.			
construct an expression for K_a.			
perform calculations relating the pH of a weak acid to the concentration of the acid and the dissociation constant, K_a.			
convert K_a into pK_a and vice versa.			
perform calculations for these titrations based on experimental results.			
sketch and explain the shapes of typical pH curves.			
use pH curves to select an appropriate indicator.			
explain qualitatively the action of acidic and basic buffers.			
calculate the pH of acidic buffer solutions.			
(Required Practical 9) investigate how pH changes when a weak acid reacts with a strong base and when a strong acid reacts with a weak base.			

Notes

3.2.1 Periodicity

I should be able to...	Revised	Revised	Revised
classify elements as s, p, d or f block according to their position in the Periodic Table or by their electron configurations.			
explain the trends in atomic radius and first ionisation energy across a Period.			
explain the melting point of the elements in terms of their structure and bonding.			

Notes

3.2.2 Group 2 – The Alkaline Earth Metals

I should be able to...	Revised	Revised	Revised
explain the trends in atomic radius as you go down Group 2.			
explain the trend in first ionisation energy as you go down Group 2.			
explain the melting point of the elements in Group 2 in terms of their structure and bonding.			
write equations for the reactions of Group 2 elements with water.			
describe the trend in solubilities of the hydroxides of Group 2 elements in water.			
describe the trend in solubilities of the sulfates of Group 2 elements in water.			
describe the use of magnesium in the extraction of titanium.			
describe the use of $Mg(OH)_2$ in medicine.			
describe the use of $Ca(OH)_2$ in agriculture.			
describe the use of CaO and $CaCO_3$ in the removal of SO_2 from flue gases.			
describe the use of $BaSO_4$ in medicine.			
explain why $BaCl_2$ solution is used to test for sulfate ions and why it is acidified.			

Notes

3.2.3 Group 7 – The Halogens

I should be able to...	Revised	Revised	Revised
explain the trend in electronegativity as you go down Group 7.			
explain the trend in the boiling point of the elements in Group 7 in terms of their structure and bonding.			
explain, in the identification of halide ions, why silver nitrate solution is used, why the silver nitrate solution is acidified and why ammonia solution is added.			
describe the trend in the oxidising ability of halogens as you go down Group 7.			
describe the trend in reducing ability of halide ions as you go down Group 7, including the reactions of solid sodium halides with concentrated sulfuric acid.			
write equations for the reactions of chlorine with water.			
assess the role of chlorine in water treatment.			
write an equation for the reaction of chlorine with cold, dilute, aqueous NaOH and describe the uses of the solution formed.			
(Required Practical 4) describe simple test–tube reactions to identify: • cations – Group 2, NH_4^+ • anions – Group 7 (halide ions), OH^-, CO_3^{2-}, SO_4^{2-}			

Notes

3.2.4 Properties of Period 3 Elements and Their Oxides

I should be able to...	Revised	Revised	Revised
explain the trend in the melting point of the oxides of the elements Na–S in terms of their structure and bonding.			
explain the trends in the reactions of the oxides with water in terms of the type of bonding present in each oxide.			
write equations for the reactions that occur between the oxides of the elements Na–S and given acids and bases.			

Notes

3.2.5 Transition Metals

I should be able to...	Revised	Revised	Revised
define the terms transition element and ligand.			
recall the characteristic properties of transition elements.			
represent the 3D structure of transition metal complex ions.			
explain the chelate effect, in terms of the balance between the entropy and enthalpy change in these reactions.			
explain how the colours of transition metal ions arise.			
describe how a simple colorimeter can be used to determine the concentration of coloured ions in solution.			
perform calculations for titrations of Fe^{2+} and $C_2O_4^{2-}$ with MnO_4^- and similar redox reactions.			
explain the importance of variable oxidation states in catalysis.			
explain, with the aid of equations, how V_2O_5 acts as a catalyst in the Contact process.			
explain, with the aid of equations, how Fe^{2+} ions catalyse the reaction between I^- and $S_2O_8^{2-}$.			
explain, with the aid of equations, how Mn^{2+} ions autocatalyse the reaction between $C_2O_4^{2-}$ and MnO_4^-.			

Notes

3.2.6 Reactions of Ions in Aqueous Solution

I should be able to...	Revised	Revised	Revised
explain, in terms of the charge/size ratio of the metal ion, why the acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$.			
describe and explain the simple test–tube reactions of: $M^{2+}(aq)$ ions, limited to M = Fe and Cu, and of $M^{3+}(aq)$ ions, limited to M = Al and Fe, with the bases OH^-, NH_3 and CO_3^{2-}.			
(Required Practical 11) describe simple test–tube reactions to identify transition metal ions in aqueous solution.			

Notes

3.1.1 Atomic Structure

I should be able to...	Revised	Revised	Revised
determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge.			
explain the existence of isotopes.			
describe the principles of a simple time of flight (TOF) mass spectrometer, limited to ionisation, acceleration to give all ions constant kinetic energy, ion drift, ion detection and data analysis.			
interpret simple mass spectra of elements.			
calculate relative atomic mass from isotopic abundance, limited to mononuclear ions.			
deduce electron configurations of atoms and ions up to Z = 36 in terms of shells and sub–shells (orbitals) s, p and d.			
define first ionisation energy.			
write equations for first and successive ionisation energies.			
explain how first and successive ionisation energies in Period 3 (Na–Ar) and in Group 2 (Be–Ba) give evidence for electron configuration in sub–shells and in shells.			

Notes

3.1.2 Amount of Substance

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
define relative atomic mass (A_r).			
define relative molecular mass (M_r).			
carry out calculations using the Avogadro constant.			
carry out calculations using mass of substance, M_r, and amount in moles.			
carry out calculations using concentration, volume and amount of substance in a solution.			
use the ideal gas equation in calculations.			
calculate empirical formula from data giving composition by mass or percentage by mass.			
calculate molecular formula from the empirical formula and relative molecular mass.			
write balanced equations for the reactions studied.			
balance equations for unfamiliar reactions when reactants and products are specified.			
use balanced equations to calculate; masses; volumes of gases; percentage yields; percentage atom economies; concentrations and volumes for reactions in solutions.			
(Required Practical 1) describe how to make up a volumetric solution and carry out a simple acid–base titration.			

Notes

3.1.3 Bonding

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
predict the charge on a simple ion using the position of the element in the Periodic Table.			
recall formulas of compound ions, e.g. sulfate, hydroxide, nitrate, carbonate and ammonium.			
construct formulas for ionic compounds.			
represent a covalent bond using a line and a coordinate bond using an arrow.			
relate the melting point and conductivity of materials to the type of structure and the bonding present.			
explain the energy changes associated with changes of state.			
draw diagrams to represent the four crystal structures involving specified numbers of particles.			
explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs of electrons) surrounding the central atom.			
use partial charges to show that a bond is polar.			
explain why some molecules with polar bonds do not have a permanent dipole.			
explain the existence of permanent dipole–dipole forces, induced dipole–dipole (van der Waals, dispersion, London) forces and hydrogen bonding between familiar and unfamiliar molecules.			
explain how melting and boiling points are influenced by intermolecular forces.			

Notes

3.1.4 Energetics

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
define standard enthalpy of combustion ($\Delta_c H^\theta$).			
define standard enthalpy of formation ($\Delta_f H^\theta$).			
use the equation $q = mc\Delta T$ to calculate the molar enthalpy change for a reaction.			
use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.			
define the term mean bond enthalpy.			
use mean bond enthalpies to calculate an approximate value of ΔH for reactions in the gaseous phase.			
explain why values from mean bond enthalpy calculations differ from those determined using Hess's law.			
(Required Practical 2) describe a method to measure an enthalpy change.			

Notes

3.1.6 Chemical Equilibria, Le Chatelier's Principle and K_c

This topic is also examined in Paper 2.

I should be able to...	Revised	Revised	Revised
use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium.			
explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used.			
construct an expression for K_c for a homogeneous system in equilibrium.			
calculate a value for K_c from the equilibrium concentrations for a homogeneous system at constant temperature.			
perform calculations involving K_c.			
predict the qualitative effects of changes of temperature on the value of K_c.			

Notes

3.1.7 Oxidation, Reduction and Redox Equations

I should be able to...	Revised	Revised	Revised
work out the oxidation state of an element in a compound or ion from the formula.			
write half-equations identifying the oxidation and reduction processes in redox reactions.			
combine half-equations to give an overall redox equation.			

Notes

3.1.8 Thermodynamics

I should be able to...	Revised	Revised	Revised
define each of the following terms: • enthalpy of formation • ionisation energy • enthalpy of atomisation • bond enthalpy • electron affinity • lattice enthalpy			
construct Born–Haber cycles to calculate lattice enthalpies using these enthalpy changes.			
construct Born–Haber cycles to calculate one of the other enthalpy changes.			
compare lattice enthalpies from Born–Haber cycles with those from calculations based on a perfect ionic model to provide evidence for covalent character in ionic compounds.			
define the term enthalpy of hydration.			
calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration.			
calculate entropy changes from absolute entropy values.			
use the relationship $\Delta G = \Delta H - T\Delta S$ to determine how ΔG varies with temperature.			
use the relationship $\Delta G = \Delta H - T\Delta S$ to determine the temperature at which a reaction becomes feasible.			

Notes

3.1.10 Equilibrium Constant K_p for Homogeneous Systems

I should be able to...	Revised	Revised	Revised
derive partial pressure from mole fraction and total pressure.			
construct an expression for K_p for a homogeneous system in equilibrium.			
perform calculations involving K_p.			
predict the qualitative effects of changes in temperature and pressure on the position of equilibrium.			
predict the qualitative effects of changes in temperature on the value of K_p.			
understand that, whilst a catalyst can affect the rate of attainment of an equilibrium, it does not affect the value of the equilibrium constant.			

Notes

3.1.11 Electrode Potentials and Electrochemical Cells

I should be able to...	Revised	Revised	Revised
use E^{θ} values to predict the direction of simple redox reactions.			
calculate the EMF of a cell.			
write and apply the conventional representation of a cell.			
use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells.			
deduce the EMF of a cell.			
explain how the electrode reactions of a lithium–ion cell can be used to generate an electric current.			
explain how the electrode reactions of an alkaline hydrogen–oxygen fuel cell can be used to generate an electric current.			
(Require Practical 8) measure the EMF of an electrochemical cell.			

Notes

3.1.12 Acids and Bases

I should be able to...	Revised	Revised	Revised
define the term 'Bronstead-Lowery acid'.			
define the term 'Bronstead-Lowery base'.			
convert concentration of hydrogen ions into pH and vice versa.			
calculate the pH of a solution of a strong acid from its concentration.			
use K_w to calculate the pH of a strong base from its concentration.			
construct an expression for K_a.			
perform calculations relating the pH of a weak acid to the concentration of the acid and the dissociation constant, K_a.			
convert K_a into pK_a and vice versa.			
perform calculations for these titrations based on experimental results.			
sketch and explain the shapes of typical pH curves.			
use pH curves to select an appropriate indicator.			
explain qualitatively the action of acidic and basic buffers.			
calculate the pH of acidic buffer solutions.			
(Required Practical 9) investigate how pH changes when a weak acid reacts with a strong base and when a strong acid reacts with a weak base.			

Notes

3.2.1 Periodicity

I should be able to...	Revised	Revised	Revised
classify elements as s, p, d or f block according to their position in the Periodic Table or by their electron configurations.			
explain the trends in atomic radius and first ionisation energy across a Period.			
explain the melting point of the elements in terms of their structure and bonding.			

Notes

3.2.2 Group 2 – The Alkaline Earth Metals

I should be able to...	Revised	Revised	Revised
explain the trends in atomic radius as you go down Group 2.			
explain the trend in first ionisation energy as you go down Group 2.			
explain the melting point of the elements in Group 2 in terms of their structure and bonding.			
write equations for the reactions of Group 2 elements with water.			
describe the trend in solubilities of the hydroxides of Group 2 elements in water.			
describe the trend in solubilities of the sulfates of Group 2 elements in water.			
describe the use of magnesium in the extraction of titanium.			
describe the use of $Mg(OH)_2$ in medicine.			
describe the use of $Ca(OH)_2$ in agriculture.			
describe the use of CaO and $CaCO_3$ in the removal of SO_2 from flue gases.			
describe the use of $BaSO_4$ in medicine.			
explain why $BaCl_2$ solution is used to test for sulfate ions and why it is acidified.			

Notes

3.2.3 Group 7 – The Halogens

I should be able to...	Revised	Revised	Revised
explain the trend in electronegativity as you go down Group 7.			
explain the trend in the boiling point of the elements in Group 7 in terms of their structure and bonding.			
explain, in the identification of halide ions, why silver nitrate solution is used, why the silver nitrate solution is acidified and why ammonia solution is added.			
describe the trend in the oxidising ability of halogens as you go down Group 7.			
describe the trend in reducing ability of halide ions as you go down Group 7, including the reactions of solid sodium halides with concentrated sulfuric acid.			
write equations for the reactions of chlorine with water.			
assess the role of chlorine in water treatment.			
write an equation for the reaction of chlorine with cold, dilute, aqueous NaOH and describe the uses of the solution formed.			
(Required Practical 4) describe simple test–tube reactions to identify: • cations – Group 2, NH_4^+ • anions – Group 7 (halide ions), OH^-, CO_3^{2-}, SO_4^{2-}			

Notes

3.2.4 Properties of Period 3 Elements and Their Oxides

I should be able to...	Revised	Revised	Revised
explain the trend in the melting point of the oxides of the elements Na–S in terms of their structure and bonding.			
explain the trends in the reactions of the oxides with water in terms of the type of bonding present in each oxide.			
write equations for the reactions that occur between the oxides of the elements Na–S and given acids and bases.			

Notes

3.2.5 Transition Metals

I should be able to...	Revised	Revised	Revised
define the terms transition element and ligand.			
recall the characteristic properties of transition elements.			
represent the 3D structure of transition metal complex ions.			
explain the chelate effect, in terms of the balance between the entropy and enthalpy change in these reactions.			
explain how the colours of transition metal ions arise.			
describe how a simple colorimeter can be used to determine the concentration of coloured ions in solution.			
perform calculations for titrations of Fe^{2+} and $C_2O_4^{2-}$ with MnO_4^- and similar redox reactions.			
explain the importance of variable oxidation states in catalysis.			
explain, with the aid of equations, how V_2O_5 acts as a catalyst in the Contact process.			
explain, with the aid of equations, how Fe^{2+} ions catalyse the reaction between I^- and $S_2O_8^{2-}$.			
explain, with the aid of equations, how Mn^{2+} ions autocatalyse the reaction between $C_2O_4^{2-}$ and MnO_4^-.			

Notes

3.2.6 Reactions of Ions in Aqueous Solution

I should be able to...	Revised	Revised	Revised
explain, in terms of the charge/size ratio of the metal ion, why the acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$.			
describe and explain the simple test–tube reactions of: $M^{2+}(aq)$ ions, limited to M = Fe and Cu, and of $M^{3+}(aq)$ ions, limited to M = Al and Fe, with the bases OH^-, NH_3 and CO_3^{2-}.			
(Required Practical 11) describe simple test–tube reactions to identify transition metal ions in aqueous solution.			

Notes

3.3.1 Introduction to Organic Chemistry

I should be able to...	Revised	Revised	Revised
draw structural, displayed and skeletal formulas for given organic compounds.			
apply IUPAC rules for nomenclature to name organic compounds limited to chains and rings with up to six carbon atoms each.			
apply IUPAC rules for nomenclature to draw the structure of an organic compound from the IUPAC name limited to chains and rings with up to six carbon atoms each.			
write balanced equations for the steps in a free-radical mechanism.			
outline mechanisms by drawing the structures of the species involved and curly arrows to represent the movement of electron pairs.			
define the term structural isomer.			
draw the structures of chain, position and functional group isomers.			
define the term stereoisomer.			
draw the structural formulas of *E* and *Z* isomers.			
apply the CIP priority rules to *E* and *Z* isomers.			

Notes

3.3.2 Alkanes

I should be able to...	Revised	Revised	Revised
explain the economic reasons for cracking alkanes.			
write balanced equations for the complete and incomplete combustion of alkanes.			
explain why sulfur dioxide can be removed from flue gases using calcium oxide or calcium carbonate.			
explain the reaction of methane with chlorine as a free-radical substitution mechanism involving initiation, propagation and termination steps.			

Notes

3.3.3 Halogenoalkanes

I should be able to...	Revised	Revised	Revised
outline the nucleophilic substitution mechanisms of the reactions of halogenoalkanes with the nucleophiles OH^-, CN^- and NH_3.			
explain why the carbon–halogen bond enthalpy influences the rate of reaction.			
explain the role of potassium hydroxide as both nucleophile and base in the concurrent substitution and elimination reactions with a halogenoalkane.			
outline the mechanisms of the substitution and elimination reactions of a halogenoalkane with potassium hydroxide.			
use equations, such as the following, to explain how chlorine atoms catalyse decomposition of ozone: $Cl\bullet + O_3 \rightarrow ClO\bullet + O_2$ and $ClO\bullet + O_3 \rightarrow 2O_2 + Cl\bullet$.			

Notes

3.3.4 Alkenes

I should be able to...	Revised	Revised	Revised
describe the use of bromine to test for unsaturation.			
outline the mechanisms for the reactions of alkenes with HBr, H_2SO_4 and Br_2.			
explain the formation of major and minor products by reference to the relative stabilities of primary, secondary and tertiary carbocation intermediates.			
draw the repeating unit of a polymer from a monomer structure.			
draw the repeating unit from a section of the polymer chain.			
draw the structure of the monomer from a section of the polymer.			
explain why addition polymers are unreactive.			
explain the nature of intermolecular forces between molecules of polyalkenes.			

Notes

3.3.5 Alcohols

I should be able to...	Revised	Revised	Revised
explain the meaning of the term biofuel.			
justify the conditions used in the production of ethanol by fermentation of glucose.			
write equations to support the statement that ethanol produced by fermentation is a carbon-neutral fuel and give reasons why this statement is not valid.			
outline the mechanism for the formation of an alcohol by the reaction of an alkene with steam in the presence of an acid catalyst.			
discuss the environmental (including ethical) issues linked to decision making about biofuel use.			
write equations for the oxidation reactions of primary alcohols to form aldehydes which can then be oxidised further to form carboxylic acids (equations showing [O] as oxidant are acceptable).			
write equations for the oxidation reactions of secondary alcohols to form ketones (equations showing [O] as oxidant are acceptable).			
explain how the method used to oxidise a primary alcohol determines whether an aldehyde or carboxylic acid is obtained.			
use chemical tests to distinguish between aldehydes and ketones including Fehling's solution and Tollens' reagent.			
outline the mechanism for the elimination of water from alcohols.			
(Required Practical 5) describe the distillation of a product from a reaction.			

Notes

3.3.6 Organic Analysis

I should be able to...	Revised	Revised	Revised
identify the functional groups using reactions in the specification.			
use precise atomic masses and the precise molecular mass to determine the molecular formula of a compound.			
use infrared spectra and the Chemistry Data Sheet or Booklet to identify particular bonds, and therefore functional groups, and also to identify impurities.			
(Required Practical 6) describe test–tube reactions to identify functional groups including alcohol, aldehyde, alkene and carboxylic acid.			

Notes

3.3.7 Optical Isomerism

I should be able to...	Revised	Revised	Revised
draw the structural formulas and displayed formulas of enantiomers.			
understand how racemic mixtures (racemates) are formed and why they are optically inactive.			
describe how enantiomers might be identified.			

Notes

3.3.8 Aldehydes and Ketones

I should be able to...	Revised	Revised	Revised
write overall equations for reduction reactions using [H] as the reductant.			
outline the nucleophilic addition mechanism for reduction reactions with $NaBH_4$ (the nucleophile should be shown as H^-).			
write overall equations for the formation of hydroxynitriles using HCN.			
outline the nucleophilic addition mechanism for the reaction with KCN followed by dilute acid.			
explain why nucleophilic addition reactions of KCN, followed by dilute acid, can produce a mixture of enantiomers.			

Notes

3.3.9 Carboxylic Acids and Derivatives

I should be able to...	Revised	Revised	Revised
identify and represent the structures of carboxylic acids, esters, acyl chlorides, acid anhydrides and amides.			
write equations to show that carboxylic acids will liberate CO_2 from carbonates.			
write equations for the reaction between carboxylic acids and alcohols, in the presence of a catalyst, to form esters.			
recall some common uses of esters.			
outline the mechanism of nucleophilic addition–elimination reactions of acyl chlorides with water, alcohols, ammonia and primary amines.			
assess the industrial advantages of ethanoic anhydride over ethanoyl chloride in the manufacture of the drug aspirin.			
(Required Practical 10) describe the preparation of a pure organic solid and testing its purity and the preparation of a pure organic liquid.			

Notes

3.3.10 Aromatic Chemistry

I should be able to...	Revised	Revised	Revised
use thermochemical evidence from enthalpies of hydrogenation to account for the extra stability as a result of the delocalisation of p electrons.			
explain why substitution reactions occur in preference to addition reactions.			
outline the electrophilic substitution mechanisms of nitration, including the generation of the nitronium ion.			
outline the electrophilic substitution mechanisms of acylation using $AlCl_3$ as a catalyst.			
explain how melting and boiling points are influenced by intermolecular forces.			

Notes

3.3.11 Amines

I should be able to...	Revised	Revised	Revised
explain the difference in base strength between ammonia, primary aliphatic and primary aromatic amines in terms of the availability of the lone pair of electrons on the N atom.			
outline the mechanisms of the nucleophilic substitution reactions of ammonia and amines with halogenoalkanes to form primary, secondary, tertiary amines and quaternary ammonium salts.			
outline the mechanisms of the nucleophilic addition–elimination reactions of ammonia and primary amines with acyl chlorides and acid anhydrides.			

Notes

3.3.12 Polymers

I should be able to...	Revised	Revised	Revised
draw the repeating unit from monomer structure(s).			
draw the repeating unit from a section of the polymer chain.			
draw the structure (s) of the monomer (s) from a section of the polymer.			
explain the nature of the intermolecular forces between molecules of condensation polymers.			
evaluate the advantages and disadvantages of different methods of disposal of polymers, including recycling.			
explain why polyesters and polyamides can be hydrolysed but polyalkenes cannot.			

Notes

3.3.13 Amino Acids, Proteins and DNA

I should be able to...	Revised	Revised	Revised
draw the structures of amino acids as zwitterions and the ions formed from amino acids in acid solution and in alkaline solution.			
draw the structure of a peptide formed from up to three amino acids.			
draw the structure of the amino acids formed by hydrolysis of a peptide.			
identify primary, secondary and tertiary structures of proteins in diagrams.			
explain how protein structures are maintained by hydrogen bonding and S–S bonds.			
calculate R_f values from a chromatogram.			
explain why a stereospecific active site can only bond to one enantiomeric form of a substrate or drug.			
explain how hydrogen bonding between base pairs leads to the two complementary strands of DNA.			
explain why cisplatin prevents DNA replication.			
explain why drugs such as cisplatin can have adverse effects.			

Notes

3.3.14 Organic Synthesis

I should be able to...	Revised	Revised	Revised
explain why chemists aim to design processes that do not require a solvent and that use non-hazardous starting materials.			
explain why chemists aim to design production methods with fewer steps that have a high percentage atom economy.			
use reactions in this specification to devise a synthesis, with up to four steps, for an organic compound.			

Notes

3.3.15 Nuclear Magnetic Resonance Spectroscopy

I should be able to...	Revised	Revised	Revised
explain why TMS is a suitable substance to use as a standard.			
use ^{1}H NMR and ^{13}C NMR spectra and chemical shift data from the Chemistry Data Booklet to suggest possible structures or part structures for molecules.			
use integration data from ^{1}H NMR spectra to determine the relative numbers of equivalent protons in the molecule.			
use the n+1 rule to deduce the spin–spin splitting patterns of adjacent, non-equivalent protons, limited to doublet, triplet and quartet formation in aliphatic compounds.			

Notes

3.3.16 Chromatography

I should be able to...	Revised	Revised	Revised
calculate R_f values from a chromatogram.			
compare retention times and R_f values with standards to identify different substances.			
(Required Practical 12) describe the separation of species by thin-layer chromatography.			

Notes