MARINA INTERNATIONAL SCHOOL

CHEMISTRY SCHEME OF WORK

FORM 6 - TERM 1

WEEK	TOPIC	TOPIC DETAILS
1.1	Relative masses of atoms and molecules	define and use the terms relative atomic, isotopic, molecular and formula masses, based on the 12C scale
1.2	The mole and the Avogadro constant	define and use the term mole in terms of the Avogadro constant
1.3	Periodicity of physical properties of the elements in Period 3	describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements (see the Data Booklet) b) explain qualitatively the variation in atomic radius and ionic radius c) interpret the variation in melting point and electrical conductivity in terms of the presence of simple molecular, giant molecular or metallic bonding in the elements d) explain the variation in first ionisation energy (see the Data Booklet) e) explain the strength, high melting point and electrical insulating properties of ceramics in terms of their giant structure; to include magnesium oxide, aluminium oxide and silicon dioxide
2.1	The determination of relative atomic masses, Ar	a) analyse mass spectra in terms of isotopic abundances (knowledge of the working of the mass spectrometer is not required) b) calculate the relative atomic mass of an element given the relative abundances of its isotopes, or its mass spectrum
2.2	The calculation of empirical and molecular formulae	define and use the terms empirical and molecular formula b) calculate empirical and molecular formulae, using combustion data or composition by mass

WEEK	TOPIC	TOPIC DETAILS
2.3	Periodicity of chemical properties of the elements in Period 3	a) describe the reactions, if any, of the elements with oxygen (to give Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2, SO_3), chlorine (to give $NaCl$, $MgCl_2$, Al_2Cl_6, $SiCl_4$, PCl_5) and water (Na and Mg only) b) state and explain the variation in oxidation number of the oxides (sodium to sulfur only) and chlorides (sodium to phosphorus only) in terms of their outer shell (valence shell) electrons c) describe the reactions of the oxides with water (treatment of peroxides and superoxides is not required) d) describe and explain the acid/base behaviour of oxides and hydroxides including, where relevant, amphoteric behaviour in reactions with acids and bases (sodium hydroxide only) e) describe and explain the reactions of the chlorides with water f) interpret the variations and trends in terms of bonding and electronegativity g) suggest the types of chemical bonding present in chlorides and oxides from observations of their chemical and physical properties
3.1	Reacting masses and volumes (of solutions and gases)	a) write and construct balanced equations b) perform calculations, including use of the mole concept, involving: (i) reacting masses (from formulae and equations) (ii) volumes of gases (e.g. in the burning of hydrocarbons) (iii) volumes and concentrations of solutions When performing calculations, candidates' answers should reflect the number of significant figures given or asked for in the question. When rounding up or down, candidates should ensure that significant figures are neither lost unnecessarily nor used beyond what is justified c) deduce stoichiometric relationships from calculations
3.2	Chemical periodicity of other elements	predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity b) deduce the nature, possible position in the Periodic Table and identity of unknown elements from given information about physical and chemical properties
4.1	Particles in the atom	a) identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses b) deduce the behaviour of beams of protons, neutrons and electrons in electric fields c) describe the distribution of mass and charge within an atom d) deduce the numbers of protons, neutrons and electrons present in both atoms and ions given proton and nucleon numbers (atomic and mass numbers) and charge

WEEK	TOPIC	TOPIC DETAILS
4.2	. The nucleus of the atom	a) describe the contribution of protons and neutrons to atomic nuclei in terms of proton (atomic) number and nucleon (mass) number b) distinguish between isotopes on the basis of different numbers of neutrons present c) recognise and use the symbolism x yA for isotopes, where x is the nucleon (mass) number and y is the proton (atomic) number
4.3	Similarities and trends in the properties of the Group 2 metals, magnesium to barium, and their compounds	a) describe the reactions of the elements with oxygen, water and dilute acids b) describe the behaviour of the oxides, hydroxides and carbonates with water and dilute acids c) describe the thermal decomposition of the nitrates and carbonates d) interpret, and make predictions from, the trends in physical and chemical properties of the elements and their compounds e) state the variation in the solubilities of the hydroxides and sulfates
5.1	Electrons: energy levels, atomic orbitals, ionisation energy, electron affinity	a) describe the number and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2 and 3 and also the 4s and 4p orbitals b) describe and sketch the shapes of s and p orbitals c) state the electronic configuration of atoms and ions given the proton (atomic) number and charge, using the convention 1s2 2s2 2p6 , etc. d) (i) explain and use the term ionisation energy (ii) explain the factors influencing the ionisation energies of elements (iii) explain the trends in ionisation energies across a period and down a group of the Periodic Table e) deduce the electronic configurations of elements from successive ionisation energy data f) interpret successive ionisation energy data of an element in terms of the position of that element within the Periodic Table
5.2	Some uses of Group 2 compounds	a) describe and explain the use of calcium hydroxide and calcium carbonate (powdered limestone) in agriculture
6.1	Ionic bonding	a) describe ionic bonding, using the examples of sodium chloride, magnesium oxide and calcium fluoride, including the use of ‘dot-and cross’ diagrams

WEEK	TOPIC	TOPIC DETAILS
6.2	Covalent bonding and co-ordinate (dative covalent) bonding including shapes of simple molecules	a) describe, including the use of 'dot-and-cross' diagrams: (i) covalent bonding, in molecules such as hydrogen, oxygen, chlorine, hydrogen chloride, carbon dioxide, methane, ethene (ii) co-ordinate (dative covalent) bonding, such as in the formation of the ammonium ion and in the Al_2Cl_6 molecule b) describe covalent bonding in terms of orbital overlap, giving σ and π bonds, including the concept of hybridisation to form sp, sp^2 and sp^3 orbitals (see also Section 14.3) c) explain the shapes of, and bond angles in, molecules by using the qualitative model of electron-pair repulsion (including lone pairs), using as simple examples BF_3 (trigonal planar), CO_2 (linear), CH_4 (tetrahedral), NH_3 (pyramidal), H_2O (non-linear), SF_6 (octahedral), PF_5 (trigonal bipyramidal) d) predict the shapes of, and bond angles in, molecules and ions
6.3	Physical properties of the Group 17 elements	a) describe the colours and the trend in volatility of chlorine, bromine and iodine b) interpret the volatility of the elements in terms of van der Waals' forces
6.4	The chemical properties of the elements and their hydride	a) describe the relative reactivity of the elements as oxidising agents b) describe and explain the reactions of the elements with hydrogen c) (i) describe and explain the relative thermal stabilities of the hydrides (ii) interpret these relative stabilities in terms of bond energies
7.1	Intermolecular forces, electronegativity and bond properties	a) describe hydrogen bonding, using ammonia and water as simple examples of molecules containing N–H and O–H groups b) understand, in simple terms, the concept of electronegativity and apply it to explain the properties of molecules such as bond polarity , the dipole moments of molecules and the behaviour of oxides with water c) explain the terms bond energy, bond length and bond polarity and use them to compare the reactivities of covalent bonds d) describe intermolecular forces (van der Waals' forces), based on permanent and induced dipoles, as in, for example, $CHCl_3(l)$; $Br_2(l)$ and the liquid Group 18 elements
7.2	Some reactions of the halide ions	a) describe and explain the reactions of halide ions with: (i) aqueous silver ions followed by aqueous ammonia (ii) concentrated sulfuric acid
7.3	The reactions of chlorine with aqueous sodium hydroxide	a) describe and interpret, in terms of changes of oxidation number, the reaction of chlorine with cold and with hot aqueous sodium hydroxide and recognise this as a disproportionation reaction

WEEK	TOPIC	TOPIC DETAILS
7.4	Some important uses of the halogens and of halogen compounds	a) explain the use of chlorine in water purification b) state the industrial importance and environmental significance of the halogens and their compounds (e.g. for bleaches, PVC, halogenated hydrocarbons as solvents, refrigerants and in aerosols
8.1	Metallic bonding	describe metallic bonding in terms of a lattice of positive ions surrounded by delocalised electrons
8.2	Bonding and physical properties	a) describe, interpret and predict the effect of different types of bonding (ionic bonding, covalent bonding, hydrogen bonding, other intermolecular interactions, metallic bonding) on the physical properties of substances b) deduce the type of bonding present from given information c) show understanding of chemical reactions in terms of energy transfers associated with the breaking and making of chemical bonds
8.3	Nitrogen	a) explain the lack of reactivity of nitrogen b) describe and explain: (i) the basicity of ammonia (ii) the structure of the ammonium ion and its formation by an acid-base reaction (iii) the displacement of ammonia from its salts c) state the industrial importance of ammonia and nitrogen compounds derived from ammonia d) state and explain the environmental consequences of the uncontrolled use of nitrate fertilisers e) state and explain the natural and man-made occurrences of oxides of nitrogen and their catalytic removal from the exhaust gases of internal combustion engines f) explain why atmospheric oxides of nitrogen are pollutants, including their catalytic role in the oxidation of atmospheric sulfur dioxide
9.1	Sulfur: the formation of atmospheric sulfur dioxide, its role in acid rain	a) describe the formation of atmospheric sulfur dioxide from the combustion of sulfur-contaminated fossil fuels b) state the role of sulfur dioxide in the formation of acid rain and describe the main environmental consequences of acid rain

WEEK	TOPIC	TOPIC DETAILS
10.1	Formulae, functional groups and the naming of organic compounds	a) interpret and use the general, structural, displayed and skeletal formulae of the following classes of compound: (i) alkanes, alkenes and arenes (ii) halogenoalkanes and halogenoarenes (iii) alcohols (including primary, secondary and tertiary) and phenols (iv) aldehydes and ketones (v) carboxylic acids, esters and acyl chlorides (vi) amines (primary only), nitriles, amides and amino acids (Candidates are expected to recognise the shape of the benzene ring when it is present in organic compounds. Knowledge of benzene or its compounds is not required for AS Level.) b) understand and use systematic nomenclature of simple aliphatic organic molecules with functional groups detailed up to six carbon atoms (six plus six for esters and amides, straight chains only) c) deduce the possible isomers for an organic molecule of known molecular formula d) deduce the molecular formula of a compound, given its structural, displayed or skeletal formula
11.1	Characteristic organic reactions	a) interpret and use the following terminology associated with types of organic reactions: (i) functional group (ii) homolytic and heterolytic fission (iii) free radical, initiation, propagation, termination (iv) nucleophile, electrophile (v) addition, substitution, elimination, hydrolysis, condensation (vi) oxidation and reduction (in equations for organic redox reactions, the symbols [O] and [H] are acceptable for oxidising and reducing agents)
12.1	Shapes of organic molecules; σ and π bonds	a) (i) describe and explain the shape of, and bond angles in, the ethane, ethene and benzene molecules in terms of σ and π bonds (ii) predict the shapes of, and bond angles in, other related molecules

WEEK	TOPIC	TOPIC DETAILS
13.1	Isomerism: structural and stereoisomerism	a) describe structural isomerism and its division into chain, positional and functional group isomerism b) describe stereoisomerism and its division into geometrical (cis-trans) and optical isomerism (use of E, Z nomenclature is acceptable but is not required) c) describe geometrical (cis-trans) isomerism in alkenes, and explain its origin in terms of restricted rotation due to the presence of π bonds d) explain what is meant by a chiral centre and that such a centre normally gives rise to optical isomerism (Candidates should appreciate that compounds can contain more than one chiral centre, but knowledge of meso compounds, or nomenclature such as diastereoisomers is not required.) e) identify chiral centres and geometrical (cis-trans) isomerism in a molecule of given structural formula

CHEMISTRY SCHEME OF WORK

FORM 6 - TERM 2

WEEK	TOPIC	TOPIC DETAILS
1.1	The gaseous state: ideal and real gases and pV = nRT	a) state the basic assumptions of the kinetic theory as applied to an ideal gas b) explain qualitatively in terms of intermolecular forces and molecular size: (i) the conditions necessary for a gas to approach ideal behaviour (ii) the limitations of ideality at very high pressures and very low temperatures c) state and use the general gas equation pV = nRT in calculations, including the determination of Mr
1.2	Alkanes	a) understand the general unreactivity of alkanes, including towards polar reagents b) describe the chemistry of alkanes as exemplified by the following reactions of ethane: (i) combustion (ii) substitution by chlorine and by bromine c) describe the mechanism of free-radical substitution at methyl groups with particular reference to the initiation, propagation and termination reactions d) explain the use of crude oil as a source of both aliphatic and aromatic hydrocarbons e) suggest how cracking can be used to obtain more useful alkanes and alkenes of lower Mr from larger hydrocarbon molecules
2.1	The liquid state	a) describe, using a kinetic-molecular model, the liquid state, melting, vaporisation and vapour pressure

WEEK	TOPIC	TOPIC DETAILS
2.2	Alkenes	a) describe the chemistry of alkenes as exemplified, where relevant, by the following reactions of ethene and propene (including the Markovnikov addition of asymmetric electrophiles to alkenes using propene as an example): (i) addition of hydrogen, steam, hydrogen halides and halogens (ii) oxidation by cold, dilute, acidified manganate(VII) ions to form the diol (iii) oxidation by hot, concentrated, acidified manganate(VII) ions leading to the rupture of the carbon–carbon double bond in order to determine the position of alkene linkages in larger molecules (iv) polymerisation b) describe the mechanism of electrophilic addition in alkenes, including using bromine/ethene and hydrogen bromide/propene as examples c) describe and explain the inductive effects of alkyl groups on the stability of cations formed during electrophilic addition d) describe the characteristics of addition polymerisation as exemplified by poly(ethene) and PVC e) deduce the repeat unit of an addition polymer obtained from a given monomer f) identify the monomer(s) present in a given section of an addition polymer molecule g) recognise the difficulty of the disposal of poly(alkene)s, i.e. nonbiodegradability and harmful combustion products
3.1	The solid state: lattice structures	a) describe, in simple terms, the lattice structure of a crystalline solid which is: (i) ionic, as in sodium chloride and magnesium oxide (ii) simple molecular, as in iodine and the fullerene allotropes of carbon (C60 and nanotubes only) (iii) giant molecular, as in silicon(IV) oxide and the graphite, diamond and graphene allotropes of carbon (iv) hydrogen-bonded, as in ice (v) metallic, as in copper b) discuss the finite nature of materials as a resource and the importance of recycling processes c) outline the importance of hydrogen bonding to the physical properties of substances, including ice and water (for example, boiling and melting points, viscosity and surface tension) d) suggest from quoted physical data the type of structure and bonding present in a substance
3.2	Hydrocarbons as fuels	a) describe and explain how the combustion reactions of alkanes make them suitable to be used as fuels in industry, in the home and in transport b) recognise the environmental consequences of: (i) carbon monoxide, oxides of nitrogen and unburnt hydrocarbons arising from the internal combustion engine and of their catalytic removal (ii) gases that contribute to the enhanced greenhouse effect c) outline the use of infra-red spectroscopy in monitoring air pollution

WEEK	TOPIC	TOPIC DETAILS
4.1	Enthalpy change, ΔH	a) explain that chemical reactions are accompanied by energy changes, principally in the form of heat energy; the energy changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive) b) explain and use the terms: (i) enthalpy change of reaction and standard conditions, with particular reference to: formation, combustion, hydration, solution, neutralisation, atomization (ii) bond energy (ΔH positive, i.e. bond breaking)) calculate enthalpy changes from appropriate experimental results, including the use of the relationship $\Delta H = -mc\Delta T$
4.2	Halogenoalkanes	a) recall the chemistry of halogenoalkanes as exemplified by: (i) the following nucleophilic substitution reactions of bromoethane: hydrolysis, formation of nitriles, formation of primary amines by reaction with ammonia (ii) the elimination of hydrogen bromide from 2-bromopropane b) describe the SN1 and SN2 mechanisms of nucleophilic substitution in halogenoalkanes including the inductive effects of alkyl groups c) recall that primary halogenoalkanes tend to react via the SN2 mechanism; tertiary halogenoalkanes via the SN1 mechanism; and secondary halogenoalkanes by a mixture of the two, depending on structure
5.1	Hess' Law, including Born-Haber cycles	a) apply Hess' Law to construct simple energy cycles, and carry out calculations involving such cycles and relevant energy terms, with particular reference to: (i) determining enthalpy changes that cannot be found by direct experiment, e.g. an enthalpy change of formation from enthalpy changes of combustion (ii) average bond energies b) construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy
5.2	Relative strength of the C–Hal bond	a) interpret the different reactivities of halogenoalkanes (with particular reference to hydrolysis and to the relative strengths of the C–Hal bonds) b) explain the uses of fluoroalkanes and fluorohalogenoalkanes in terms of their relative chemical inertness c) recognise the concern about the effect of chlorofluoroalkanes on the ozone layer
6.1	Redox processes: electron transfer and changes in oxidation number (oxidation state)	a) calculate oxidation numbers of elements in compounds and ions b) describe and explain redox processes in terms of electron transfer and changes in oxidation number c) use changes in oxidation numbers to help balance chemical equation

WEEK	TOPIC	TOPIC DETAILS
6.2	Alcohols	a) recall the chemistry of alcohols, exemplified by ethanol, in the following reactions: (i) combustion (ii) substitution to halogenoalkanes (iii) reaction with sodium (iv) oxidation to carbonyl compounds and carboxylic acids (v) dehydration to alkenes (vi) formation of esters by esterification with carboxylic acids b) (i) classify hydroxy compounds into primary, secondary and tertiary alcohols (ii) suggest characteristic distinguishing reactions, e.g. mild oxidation c) deduce the presence of a CH3CH(OH)– group in an alcohol from its reaction with alkaline aqueous iodine to form tri-iodomethane
7.1	Chemical equilibria: reversible reactions, dynamic equilibrium	a) explain, in terms of rates of the forward and reverse reactions, what is meant by a reversible reaction and dynamic equilibrium b) state Le Chatelier's principle and apply it to deduce qualitatively (from appropriate information) the effects of changes in temperature, concentration or pressure on a system at equilibrium c) state whether changes in temperature, concentration or pressure or the presence of a catalyst affect the value of the equilibrium constant for a reaction
7.2	Chemical equilibria: reversible reactions, dynamic equilibrium	d) deduce expressions for equilibrium constants in terms of concentrations, Kc, and partial pressures, Kp (treatment of the relationship between Kp and Kc is not required) e) calculate the values of equilibrium constants in terms of concentrations or partial pressures from appropriate data f) calculate the quantities present at equilibrium, given appropriate data (such calculations will not require the solving of quadratic equations) g) describe and explain the conditions used in the Haber process and the Contact process, as examples of the importance of an understanding of chemical equilibrium in the chemical industry
7.3	Aldehydes and ketones	a) describe: (i) the formation of aldehydes and ketones from primary and secondary alcohols respectively using Cr2O7 2–/H+ (ii) the reduction of aldehydes and ketones, e.g. using NaBH4 or LiAlH4 (iii) the reaction of aldehydes and ketones with HCN and NaCN or KCN b) describe the mechanism of the nucleophilic addition reactions of hydrogen cyanide with aldehydes and ketones c) describe the use of 2,4-dinitrophenylhydrazine (2,4-DNPH) to detect the presence of carbonyl compounds d) deduce the nature (aldehyde or ketone) of an unknown carbonyl compound from the results of simple tests (Fehling's and Tollens' reagents; ease of oxidation) e) describe the reaction of CH3CO– compounds with alkaline aqueous iodine to give tri-iodomethane

WEEK	TOPIC	TOPIC DETAILS
8.1	Ionic equilibria	a) show understanding of, and use, the Brønsted-Lowry theory of acids and bases, including the use of the conjugate acid, conjugate base (acid-I base-I, acid-II base-II) concept b) explain qualitatively the differences in behaviour between strong and weak acids and bases and the pH values of their aqueous solutions in terms of the extent of dissociation
8.2	Carboxylic acids	a) describe the formation of carboxylic acids from alcohols, aldehydes and nitriles b) describe the reactions of carboxylic acids in the formation of: (i) salts, by the use of reactive metals, alkalis or carbonates (ii) alkyl esters (iii) alcohols, by the use of $LiAlH_4$
9.1	Simple rate equations, orders of reaction and rate constants	a) explain and use the term rate of reaction b) explain qualitatively, in terms of collisions, the effect of concentration changes on the rate of a reaction
9.2	Esters	a) describe the acid and base hydrolysis of esters b) state the major commercial uses of esters, e.g. solvents, perfumes, flavourings
9.3	Infra-red spectroscopy	a) analyse an infra-red spectrum of a simple molecule to identify functional groups (see the Data Booklet for the functional groups required)
10.1	Effect of temperature: on reaction rates and rate constants and the concept of activation energy	a) explain and use the term activation energy, including reference to the Boltzmann distribution b) explain qualitatively, in terms both of the Boltzmann distribution and of collision frequency, the effect of temperature change on the rate of a reaction
10.2	Homogeneous and heterogeneous catalysts including enzymes	a) explain and use the term catalysis b) explain that catalysts can be homogeneous or heterogeneous c) (i) explain that, in the presence of a catalyst, a reaction has a different mechanism, i.e. one of lower activation energy (ii) interpret this catalytic effect in terms of the Boltzmann distribution d) describe enzymes as biological catalysts (proteins) which may have specificity