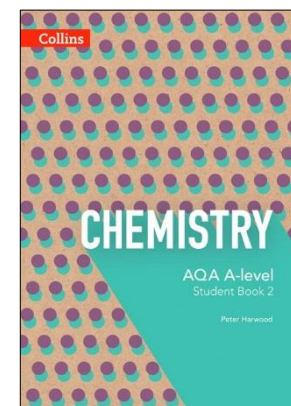


Collins

AQA A-level Chemistry Year 2

Scheme of Work



Scheme of Work

AQA A-level Chemistry Year 2 of A-level

This course covers the requirements of the second year of the AQA AS and A-level Chemistry specification. These schemes of work are designed to accompany the use of Collins' AQA A-level Chemistry Year 2 Student Book.

Each chapter is matched to the Specification Content and we have shown in which chapters the six Required Practicals may be carried out, to help you plan for these and the sourcing of necessary equipment. We have assumed that 120 one-hour lessons will be taught during the year to cover the specification.

The schemes suggested are of course flexible, and editable, to correspond with your timetabling and to enable you to plan your own route through the course.

AQA A-level Chemistry Year 2 of A-level: 120 hours

| Chapters | Specification Content | Required Practicals |
|--|---|---------------------|
| <p>CHAPTER 1 Thermodynamics</p> | <p>3.1.8.1 Born–Haber cycles</p> <p>Lattice enthalpy can be defined as either enthalpy of latticedissociation or enthalpy of lattice formation.</p> <p>Born–Haber cycles are used to calculate lattice enthalpies using the following data:</p> <ul style="list-style-type: none"> • enthalpy of formation • ionisation energy • enthalpy of atomisation • bond enthalpy • electron affinity. <p>Students should be able to:</p> <ul style="list-style-type: none"> • define each of the above terms and 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. <p>Cycles are used to calculate enthalpies of solution for ionic compounds from lattice enthalpies and enthalpies of hydration.</p> | |

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| | <p>Students should be able to:</p> <ul style="list-style-type: none"> • define the term enthalpy of hydration • perform calculations of an enthalpy change using these cycles. <p>3.1.8.2 Gibbs free-energy change, ΔG, and entropy change, ΔS</p> <p>ΔH, whilst important, is not sufficient to explain feasible change.</p> <p>The concept of increasing disorder (entropy change, ΔS).</p> <p>ΔS accounts for the above deficiency, illustrated by physical changes and chemical changes.</p> <p>The balance between entropy and enthalpy determines the feasibility of a reaction given by the relationship:</p> <p>$\Delta G = \Delta H - T\Delta S$ (derivation not required).</p> <p>For a reaction to be feasible, the value of ΔG must be zero or negative.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • 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. | |
| CHAPTER 2 Rate equations | 3.1.9.1 Rate equations | Required practical 7: Measuring the rate of reaction: |

The rate of a chemical reaction is related to the concentration of reactants by a rate equation of the form:

$$\text{Rate} = k[\text{A}]^m [\text{B}]^n$$

where m and n are the orders of reaction with respect to reactants A and B and k is the rate constant.

The orders m and n are restricted to the values 0, 1, and 2.

The rate constant k varies with temperature as shown by the equation:

$$k = Ae^{-E_a/RT}$$

where A is a constant, known as the Arrhenius constant, E_a is the activation energy and T is the temperature in K.

Students should be able to:

- define the terms order of reaction and rate constant
- perform calculations using the rate equation
- explain the qualitative effect of changes in temperature on the rate constant k
- perform calculations using the equation $k = Ae^{-E_a/RT}$
- understand that the equation $k = Ae^{-E_a/RT}$ can be rearranged into the form $\ln k = -E_a/RT + \ln A$ and know how to use this rearranged equation with experimental data to plot a straight line graph with slope $-E_a/R$

These equations and the gas constant, R , will be given when required.

3.1.9.2 Determination of rate equation

- by an initial rate method
- by a continuous monitoring method.

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| | <p>The rate equation is an experimentally determined relationship.</p> <p>The orders with respect to reactants can provide information about the mechanism of a reaction.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • use concentration–time graphs to deduce the rate of a reaction • use initial concentration–time data to deduce the initial rate of a reaction • use rate–concentration data or graphs to deduce the order (0, 1 or 2) with respect to a reactant • derive the rate equation for a reaction from the orders with respect to each of the reactants • use the orders with respect to reactants to provide information about the rate determining/limiting step of a reaction. | |
| <p>CHAPTER 3 Equilibrium constant</p> | <p>3.1.10 Equilibrium constant K_p for homogeneous systems</p> <p>The equilibrium constant K_p is deduced from the equation for a reversible reaction occurring in the gas phase.</p> <p>K_p is the equilibrium constant calculated from partial pressures for a system at constant temperature.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • derive partial pressure from mole fraction and total pressure | |

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| | <ul style="list-style-type: none"> • 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. | |
| <p>CHAPTER 4 Acids and bases</p> | <p>3.1.12.1 Brønsted–Lowry acid–base equilibria in aqueous solution</p> <p>An acid is a proton donor.</p> <p>A base is a proton acceptor.</p> <p>Acid–base equilibria involve the transfer of protons.</p> <p>3.1.12.2 Definition and determination of pH</p> <p>The concentration of hydrogen ions in aqueous solution covers a very wide range. Therefore, a logarithmic scale, the pH scale, is used as a measure of hydrogen ion concentration.</p> <p>$\text{pH} = -\log_{10}[\text{H}^+]$</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • convert concentration of hydrogen ions into pH and vice versa | <p>Required practical 9:</p> <p>Investigate how pH changes when a weak acid reacts with a strong base and when a strong acid reacts with a weak base.</p> |

- calculate the pH of a solution of a strong acid from its concentration.

3.1.12.3 The ionic product of water, K_w

Water is slightly dissociated.

K_w is derived from the equilibrium constant for this dissociation.

$$K_w = [\text{H}^+][\text{OH}^-]$$

The value of K_w varies with temperature.

Students should be able to use K_w to calculate the pH of a strong base from its concentration.

3.1.12.4 Weak acids and bases K_a for weak acids

K_a is the dissociation constant for a weak acid.

$$\text{p}K_a = -\log_{10} K_a$$

Students should be able to:

- 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 $\text{p}K_a$ and vice versa.

3.1.12.5 pH curves, titrations and indicators

Titrations of acids with bases.

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| | <p>Students should be able to perform calculations for these titrations based on experimental results.</p> <p>Typical pH curves for acid–base titrations in all combinations of weak and strong monoprotic acids and bases.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • sketch and explain the shapes of typical pH curves • use pH curves to select an appropriate indicator. <p>3.1.12.6 Buffer action</p> <p>A buffer solution maintains an approximately constant pH, despite dilution or addition of small amounts of acid or base.</p> <p>Acidic buffer solutions contain a weak acid and the salt of that weak acid.</p> <p>Basic buffer solutions contain a weak base and the salt of that weak base.</p> <p>Applications of buffer solutions.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • explain qualitatively the action of acidic and basic buffers • calculate the pH of acidic buffer solutions. | |
| <p>CHAPTER 5 Optical isomerism</p> | <p>3.3.7 Optical isomerism</p> <p>Optical isomerism is a form of stereoisomerism and occurs as a result of chirality in molecules, limited to molecules with a single chiral centre.</p> | |

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| | <p>An asymmetric carbon atom is chiral and gives rise to optical isomers (enantiomers), which exist as non superimposable mirror images and differ in their effect on plane polarised light.</p> <p>A mixture of equal amounts of enantiomers is called a racemic mixture (racemate).</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • draw the structural formulas and displayed formulas of enantiomers • understand how racemic mixtures (racemates) are formed and why they are optically inactive. | |
| <p>CHAPTER 6 Aldehydes and ketones</p> | <p>3.3.8 Aldehydes and ketones</p> <p>Aldehydes are readily oxidised to carboxylic acids.</p> <p>Chemical tests to distinguish between aldehydes and ketones including Fehling's solution and Tollens' reagent.</p> <p>Aldehydes can be reduced to primary alcohols, and ketones to secondary alcohols, using NaBH_4 in aqueous solution.</p> <p>These reduction reactions are examples of nucleophilic addition.</p> <p>The nucleophilic addition reactions of carbonyl compounds with KCN, followed by dilute acid, to produce hydroxynitriles.</p> <p>Aldehydes and unsymmetrical ketones form mixtures of enantiomers when they react with KCN followed by dilute acid.</p> | |

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| | <p>The hazards of using KCN.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • write overall equations for reduction reactions using [H] as the reductant • outline the nucleophilic addition mechanism for reduction reactions with NaBH₄ (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. | |
| <p>CHAPTER 7 Carboxylic acids and derivatives</p> | <p>3.3.9.1 Carboxylic acids and esters</p> <p>The structures of:</p> <ul style="list-style-type: none"> • carboxylic acids • esters. <p>Carboxylic acids are weak acids but will liberate CO₂ from carbonates.</p> <p>Carboxylic acids and alcohols react, in the presence of an acid catalyst, to give esters.</p> <p>Common uses of esters (eg in solvents, plasticisers, perfumes and food flavourings).</p> <p>Vegetable oils and animal fats are esters of propane-1,2,3-triol (glycerol).</p> | <p>Required practical 10:</p> <p>Preparation of:</p> <ul style="list-style-type: none"> • a pure organic solid and test of its purity • a pure organic liquid. |

Esters can be hydrolysed in acid or alkaline conditions to form alcohols and carboxylic acids or salts of carboxylic acids.

Vegetable oils and animal fats can be hydrolysed in alkaline conditions to give soap (salts of long-chain carboxylic acids) and glycerol.

Biodiesel is a mixture of methyl esters of long-chain carboxylic acids.

Biodiesel is produced by reacting vegetable oils with methanol in the presence of a catalyst.

3.3.9.2 Acylation

The structures of:

- acid anhydrides
- acyl chlorides
- amides.

The nucleophilic addition–elimination reactions of water, alcohols, ammonia and primary amines with acyl chlorides and acid anhydrides.

The industrial advantages of ethanoic anhydride over ethanoyl chloride in the manufacture of the drug aspirin.

Students should be able to outline the mechanism of nucleophilic addition–elimination reactions of acyl chlorides with water, alcohols, ammonia and primary amines.

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| | <p>The nature of the bonding in a benzene ring, limited to planar structure and bond length intermediate between single and double.</p> <p>Delocalisation of p electrons makes benzene more stable than the theoretical molecule cyclohexa-1,3,5-triene.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • use thermochemical evidence from enthalpies of hydrogenation to account for this extra stability • explain why substitution reactions occur in preference to addition reactions. <p>3.3.10.2 Electrophilic substitution</p> <p>Electrophilic attack on benzene rings results in substitution, limited to monosubstitutions.</p> <p>Nitration is an important step in synthesis, including the manufacture of explosives and formation of amines.</p> <p>Friedel–Crafts acylation reactions are also important steps in synthesis.</p> <p>Students should be able to outline the electrophilic substitution mechanisms of:</p> <ul style="list-style-type: none"> • nitration, including the generation of the nitronium ion • acylation using AlCl_3 as a catalyst. | |
| CHAPTER 9 Amines | 3.3.11.1 Preparation | |

Primary aliphatic amines can be prepared by the reaction of ammonia with halogenoalkanes and by the reduction of nitriles.

Aromatic amines, prepared by the reduction of nitro compounds, are used in the manufacture of dyes.

3.3.11.2 Base properties

Amines are weak bases.

The difference in base strength between ammonia, primary aliphatic and primary aromatic amines.

Students should be able to explain the difference in base strength in terms of the availability of the lone pair of electrons on the N atom.

3.3.11.3 Nucleophilic properties

Amines are nucleophiles.

The nucleophilic substitution reactions of ammonia and amines with halogenoalkanes to form primary, secondary, tertiary amines and quaternary ammonium salts.

The use of quaternary ammonium salts as cationic surfactants.

The nucleophilic addition–elimination reactions of ammonia and primary amines with acyl chlorides and acid anhydrides.

Students should be able to outline the mechanisms of:

- these nucleophilic substitution reactions

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| | <ul style="list-style-type: none"> • the nucleophilic addition–elimination reactions of ammonia and primary amines with acyl chlorides. | |
| <p>CHAPTER 10 Polymers</p> | <p>3.3.12.1 Condensation polymers</p> <p>Condensation polymers are formed by reactions between:</p> <ul style="list-style-type: none"> • dicarboxylic acids and diols • dicarboxylic acids and diamines • amino acids. <p>The repeating units in polyesters (eg Terylene) and polyamides (eg nylon 6,6 and Kevlar) and the linkages between these repeating units.</p> <p>Typical uses of these polymers.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • 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. <p>3.3.12.2 Biodegradability and disposal of polymers</p> <p>Polyalkenes are chemically inert and non-biodegradable.</p> <p>Polyesters and polyamides can be broken down by hydrolysis and are biodegradable.</p> | |

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| | <p>The advantages and disadvantages of different methods of disposal of polymers, including recycling.</p> <p>Students should be able to explain why polyesters and polyamides can be hydrolysed but polyalkenes cannot.</p> | |
| <p>CHAPTER 11 Amino acids, proteins and DNA</p> | <p>3.3.13.1 Amino acids</p> <p>Amino acids have both acidic and basic properties, including the formation of zwitterions.</p> <p>Students should be able to draw the structures of amino acids as zwitterions and the ions formed from amino acids:</p> <ul style="list-style-type: none"> • in acid solution • in alkaline solution. <p>3.3.13.2 Proteins</p> <p>Proteins are sequences of amino acids joined by peptide links.</p> <p>The importance of hydrogen bonding and sulfur–sulphur bonds in proteins.</p> <p>The primary, secondary (α-helix and β-pleated sheets) and tertiary structure of proteins.</p> <p>Hydrolysis of the peptide link produces the constituent amino acids.</p> <p>Amino acids can be separated and identified by thin-layer chromatography.</p> <p>Amino acids can be located on a chromatogram using developing agents such as ninhydrin or ultraviolet light and identified by their R_f values.</p> | |

Students should be able to:

- 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 in diagrams
- explain how these structures are maintained by hydrogen bonding and S–S bonds
- calculate R_f values from a chromatogram.

3.3.13.3 Enzymes

Enzymes are proteins.

The action of enzymes as catalysts, including the concept of a stereospecific active site that binds to a substrate molecule.

The principle of a drug acting as an enzyme inhibitor by blocking the active site.

Computers can be used to help design such drugs.

Students should be able to explain why a stereospecific active site can only bond to one enantiomeric form of a substrate or drug.

3.3.13.4 DNA

The structures of the phosphate ion, 2-deoxyribose (a pentose sugar) and the four bases adenine, cytosine, guanine and thymine are given in the Chemistry Data Booklet.

A nucleotide is made up from a phosphate ion bonded to 2-deoxyribose which is in turn bonded to one of the four bases adenine, cytosine, guanine and thymine.

A single strand of DNA (deoxyribonucleic acid) is a polymer of nucleotides linked by covalent bonds between the phosphate group of one nucleotide and the 2-deoxyribose of another nucleotide. This results in a sugar-phosphatesugar-phosphate polymer chain with bases attached to the sugars in the chain.

DNA exists as two complementary strands arranged in the form of a double helix.

Students should be able to explain how hydrogen bonding between base pairs leads to the two complementary strands of DNA.

3.3.13.5 Action of anticancer drugs

The Pt(II) complex cisplatin is used as an anticancer drug.

Cisplatin prevents DNA replication in cancer cells by a ligand replacement reaction with DNA in which a bond is formed between platinum and a nitrogen atom on guanine.

Appreciate that society needs to assess the balance between the benefits and the adverse effects of drugs, such as the anticancer drug cisplatin.

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| | <p>Students should be able to:</p> <ul style="list-style-type: none"> • explain why cisplatin prevents DNA replication • explain why such drugs can have adverse effects. | |
| CHAPTER 12 Organic synthesis | <p>3.3.14 Organic synthesis</p> <p>The synthesis of an organic compound can involve several steps.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • 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. | |
| CHAPTER 13 NMR and chromatography | <p>3.3.15 Nuclear magnetic resonance spectroscopy</p> <p>Appreciation that scientists have developed a range of analytical techniques which together enable the structures of new compounds to be confirmed.</p> <p>Nuclear magnetic resonance (NMR) gives information about the position of ^{13}C or ^1H atoms in a molecule.</p> <p>^{13}C NMR gives simpler spectra than ^1H NMR.</p> <p>The use of the δ scale for recording chemical shift.</p> <p>Chemical shift depends on the molecular environment.</p> | <p>Required practical 12:</p> <p>Separation of species by thin-layer chromatography.</p> |

Integrated spectra indicate the relative numbers of ^1H atoms in different environments. ^1H NMR spectra are obtained using samples dissolved in deuterated solvents or CCl_4

The use of tetramethylsilane (TMS) as a standard.

Students should be able to:

- explain why TMS is a suitable substance to use as a standard
- use ^1H 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 ^1H 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.

3.3.16 Chromatography

Chromatography can be used to separate and identify the components in a mixture.

Types of chromatography include:

- thin-layer chromatography (TLC) – a plate is coated with a solid and a solvent moves up the plate
- column chromatography (CC) – a column is packed with a solid and a solvent moves down the column

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| | <ul style="list-style-type: none"> • gas chromatography (GC) – a column is packed with a solid or with a solid coated by a liquid, and a gas is passed through the column under pressure at high temperature. <p>Separation depends on the balance between solubility in the moving phase and retention by the stationary phase. Retention times and R_f values are used to identify different substances.</p> <p>The use of mass spectrometry to analyse the components separated by GC.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • calculate R_f values from a chromatogram • compare retention times and R_f values with standards to identify different substances. | |
| <p>CHAPTER 14 Electrochemical cells</p> | <p>3.1.11 Electrode potentials and cells</p> <p>IUPAC convention for writing half-equations for electrode reactions.</p> <p>The conventional representation of cells.</p> <p>Cells are used to measure electrode potentials by reference to the standard hydrogen electrode.</p> <p>The importance of the conditions when measuring the electrode potential, E (Nernst equation not required).</p> <p>Standard electrode potential, E^\ominus, refers to conditions of 298 K, 100 kPa and 1.00 mol dm^{-3} solution of ions.</p> <p>Standard electrode potentials can be listed as an electrochemical series.</p> | <p>Required practical 8:</p> <p>Measuring the EMF of an electrochemical cell.</p> |

Students should be able to:

- use E^\ominus values to predict the direction of simple redox reactions
- calculate the EMF of a cell
- write and apply the conventional representation of a cell.

3.1.11.2 Commercial applications of electrochemical cells

Electrochemical cells can be used as a commercial source of electrical energy.

The simplified electrode reactions in a lithium cell:

Positive electrode: $\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightarrow \text{Li}^+[\text{CoO}_2]^-$

Negative electrode: $\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$

Cells can be non-rechargeable (irreversible), rechargeable or fuel cells.

Fuel cells are used to generate an electric current and do not need to be electrically recharged.

The electrode reactions in an alkaline hydrogen–oxygen fuel cell.

The benefits and risks to society associated with using these cells.

Students should be able to:

- use given electrode data to deduce the reactions occurring in non-rechargeable and rechargeable cells
- deduce the EMF of a cell

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| | <ul style="list-style-type: none"> • explain how the electrode reactions can be used to generate an electric current. | |
| CHAPTER 15 Aqueous chemistry of inorganic compounds | <p>3.2.4 Properties of Period 3 elements and their oxides</p> <p>The reactions of Na and Mg with water.</p> <p>The trends in the reactions of the elements Na, Mg, Al, Si, P and S with oxygen, limited to the formation of Na₂O, MgO, Al₂O₃, SiO₂, P₄O₁₀, SO₂ and SO₃</p> <p>The trend in the melting point of the highest oxides of the elements Na–S</p> <p>The reactions of the oxides of the elements Na–S with water, limited to Na₂O, MgO, Al₂O₃, SiO₂, P₄O₁₀, SO₂ and SO₃, and the pH of the solutions formed.</p> <p>The structures of the acids and the anions formed when P₄O₁₀, SO₂ and SO₃ react with water.</p> <p>Students should be able to:</p> <ul style="list-style-type: none"> • 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. <p>3.2.5.1 General properties of transition metals</p> | <p>Required practical 11:</p> <p>Carry out simple test-tube reactions to identify transition metal ions in aqueous solution.</p> |

Transition metal characteristics of elements Ti–Cu arise from an incomplete d sub-level in atoms or ions.

The characteristic properties include:

- complex formation
- formation of coloured ions
- variable oxidation state
- catalytic activity.

A ligand is a molecule or ion that forms a co-ordinate bond with a transition metal by donating a pair of electrons.

A complex is a central metal atom or ion surrounded by ligands.

Co-ordination number is number of co-ordinate bonds to the central metal atom or ion.

3.2.6 Reactions of ions in aqueous solution

In aqueous solution, the following metal-aqua ions are formed:

$[\text{M}(\text{H}_2\text{O})_6]^{2+}$, limited to M = Fe and Cu

$[\text{M}(\text{H}_2\text{O})_6]^{3+}$, limited to M = Al and Fe

The acidity of $[\text{M}(\text{H}_2\text{O})_6]^{3+}$ is greater than that of $[\text{M}(\text{H}_2\text{O})_6]^{2+}$

Some metal hydroxides show amphoteric character by dissolving in both acids and bases (eg hydroxides of Al^{3+}).

Students should be able to:

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| | <ul style="list-style-type: none"> • 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-} | |
| <p>CHAPTER 16 Transition metals</p> | <p>3.2.5.2 Substitution reactions</p> <p>H_2O, NH_3 and Cl^- can act as monodentate ligands.</p> <p>The ligands NH_3 and H_2O are similar in size and are uncharged.</p> <p>Exchange of the ligands NH_3 and H_2O occurs without change of co-ordination number (eg Co^{2+} and Cu^{2+}).</p> <p>Substitution may be incomplete (eg the formation of $[Cu(NH_3)_4(H_2O)_2]^{2+}$).</p> <p>The Cl^- ligand is larger than the uncharged ligands NH_3 and H_2O</p> <p>Exchange of the ligand H_2O by Cl^- can involve a change of co-ordination number (eg Co^{2+}, Cu^{2+} and Fe^{3+}).</p> <p>Ligands can be bidentate (eg $H_2NCH_2CH_2NH_2$ and $C_2O_4^{2-}$).</p> <p>Ligands can be multidentate (eg $EDTA^{4-}$).</p> <p>Haem is an iron(II) complex with a multidentate ligand.</p> <p>Oxygen forms a co-ordinate bond to $Fe(II)$ in haemoglobin, enabling oxygen to be transported in the blood.</p> | |

Carbon monoxide is toxic because it replaces oxygen co-ordinately bonded to Fe(II) in haemoglobin.

Bidentate and multidentate ligands replace monodentate ligands from complexes. This is called the chelate effect.

Students should be able to explain the chelate effect, in terms of the balance between the entropy and enthalpy change in these reactions.

3.2.5.3 Shapes of complex ions

Transition metal ions commonly form octahedral complexes with small ligands (eg H₂O and NH₃).

Octahedral complexes can display cis–trans isomerism (a special case of *E–Z* isomerism) with monodentate ligands and optical isomerism with bidentate ligands.

Transition metal ions commonly form tetrahedral complexes with larger ligands (eg Cl[–]).

Square planar complexes are also formed and can display cis–trans isomerism.

Cisplatin is the cis isomer.

Ag⁺ forms the linear complex [Ag(NH₃)₂]⁺ as used in Tollens' reagent.

3.2.5.5 Variable oxidation states

Transition elements show variable oxidation states.

Vanadium species in oxidation states IV, III and II are formed by the reduction of vanadate(V) ions by zinc in acidic solution.

The redox potential for a transition metal ion changing from a higher to a lower oxidation state is influenced by pH and by the ligand.

The reduction of $[\text{Ag}(\text{NH}_3)_2]^+$ (Tollens' reagent) to metallic silver is used to distinguish between aldehydes and ketones.

The redox titrations of Fe^{2+} and $\text{C}_2\text{O}_4^{2-}$ with MnO_4^-

Students should be able to perform calculations for these titrations and similar redox reactions.

3.2.5.6 Catalysts

Transition metals and their compounds can act as heterogeneous and homogeneous catalysts.

A heterogeneous catalyst is in a different phase from the reactants and the reaction occurs at active sites on the surface.

The use of a support medium to maximise the surface area of a heterogeneous catalyst and minimise the cost.

V_2O_5 acts as a heterogeneous catalyst in the Contact process.

Fe is used as a heterogeneous catalyst in the Haber process.

Heterogeneous catalysts can become poisoned by impurities that block the active sites and consequently have reduced efficiency; this has a cost implication.

A homogeneous catalyst is in the same phase as the reactants.

When catalysts and reactants are in the same phase, the reaction proceeds through an intermediate species.

Students should be able to:

- 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^-