GCE Chemistry – 3.2.5 Transition Metals

Learning about this topic is important because: The 3d block contains 10 elements, all of which are metals. Unlike the metals in Groups 1 and 2, the transition metals Ti to Cu form coloured compounds and compounds where the transition metal exists in different oxidation states. Some of these metals are familiar as catalysts. The properties of these elements are studied in this section with opportunities for a wide range of practical investigations.	
This builds on: 3.1.1 atomic structure, 3.1.3 Bonding, 3.2.3 Periodicity	
This leads to: 3.1.8 Thermodyamics	
We will learn: 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. Transition metals can bond with ligands to form complexes. What are monodentate, bidentate and multidentate ligands? (Using examples outlined in the specification. Ligands can be exchanged in substitution reactions and this may affect the shape and charge of the complex. How haemoglobin is a complex with Fe(II) and how this can be affected by carbon monoxide. Transition metal complexes can exhibit both optical and cis-trans isomerism. Colour in transition metal complexes arises through the absorption of wavelengths of light by the movement of electrons from a ground state to an excited state. Changes in oxidation state, co-ordination number and	We will develop/practise skills including: Identify whether an element is a transition metal and be able to describe it's properties. Be able to draw a metal complex, including shape, bond angles and charge, with given information and for the ligands outlined in the specification Write equations for ligand substitution reactions. Explain the chelate effect in terms of entropy and enthalpy change Be able to identify and draw isomers of transitional metal complexes The energy difference between the ground state and the excited state of the d electrons is given by: $\Delta E = hv = hc/\lambda$ Determine the concentration of a coloured complex ion by colorimetry. How to carry out redox titrations.

GCE Chemistry – 3.2.5 Transition Metals

ligand can lead to a change in colour. A simple colorimeter can be used to determine the concentration of coloured ions in solution. 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 $[Ag(NH3)2]^{+}$ (Tollens' reagent) to metallic silver is used to distinguish between aldehydes and ketones. The redox titrations of Fe^{2+} and $C2O4^{2-}$ with $MnO4^{-}$	
Some of the vocabulary that we will use includes: Ligand Complex Sub-level Co-ordination number monodentate Bidentate Multidentate Colorimetry Redox Oxidation state Cis-trans	You could learn more about this topic by: AQA Chemistry A-level Inorganic II Revision - PMT (physicsandmathstutor.com)
Your teacher will assess your knowledge & understanding throughout the topic by looking at your work, questioning, discussion and giving you feedback in lots of different ways. The key pieces of work in this topic are: End of topic test Whiteboard quizzes Assessed homework	

GCE Chemistry – 3.2.6 Reactions of Ions in Aqueous Solution

GCE Chemistry – 3.2.6 Reactions of Ions in Aqueous Solution

Learning about this topic is important because:
The reactions of transition metal ions in aqueous solution provide a practical opportunity to understand how metal ions can be identified using simple test-tube reactions in the laboratory.

This builds on:
3.1.3 Bonding, 3.1.7 Oxidation, reduction and redox equations, 3.2.5 Transition metals

This leads to:
Required practical skills and the identification of ions in practical and analytical chemistry.

We will learn:

In aqueous solution, metal ions exist as metal-aqua ions

The following metal-aqua ions are required:

$[Fe(H_2O)_6]^{2+}$ and $[Cu(H_2O)_6]^{2+}$

$[Al(H_2O)_6]^{3+}$ and $[Fe(H_2O)_6]^{3+}$

The acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$.

How the charge/size ratio of the metal ion affects the acidity of the metal-aqua ion.

Some metal hydroxides show amphoteric character and can dissolve in both acids and bases, including aluminium hydroxide.

The simple test-tube reactions of Fe^{2+}, Cu^{2+}, Al^{3+} and Fe^{3+} ions with:

OH^-, NH_3 and CO_3^{2-}.

We will develop/practise skills including:

Describe and explain the test-tube reactions of metal-aqua ions with NaOH, NH_3 and Na_2CO_3.

Identify Fe^{2+}, Cu^{2+}, Al^{3+} and Fe^{3+} ions from observations such as solution colour, precipitate colour and solubility in excess reagent.

Write equations for the reactions of metal-aqua ions.

Explain why 3+ metal-aqua ions are more acidic than 2+ metal-aqua ions in terms of charge/size ratio and polarisation of O–H bonds.

Explain the amphoteric behaviour of aluminium hydroxide.

Carry out simple test-tube reactions safely and accurately to identify transition metal ions in aqueous solution.

Record observations clearly and use them to identify unknown ions.

Key reactions and observations:

Fe^{2+}: pale green solution; green precipitate with OH^- or NH_3; precipitate turns brown on standing; green precipitate with CO_3^{2-}.

Cu^{2+}: blue solution; blue precipitate with OH^-; blue precipitate with NH_3 which dissolves in excess NH_3 to form a deep blue solution; blue-green precipitate with CO_3^{2-}.

Fe^{3+}: purple solution (often yellow-brown); brown precipitate with OH^- or NH_3; brown precipitate and CO_2 with CO_3^{2-}.

Some of the vocabulary that we will use includes:

Aqueous solution	Metal-aqua ion
Complex	Ligand
Co-ordinate bond	Hydrolysis
Charge/size ratio	Acidity
Precipitate	Amphoteric
Alkali	Excess
Solubility	Carbonate

GCE Chemistry – 3.2.6 Reactions of Ions in Aqueous Solution

Al^{3+}: colourless solution; white precipitate with OH^- or NH_3; precipitate dissolves in excess OH^-; white precipitate and CO_2 with CO_3^{2-}. Required practical 11: Carry out simple test-tube reactions to identify transition metal ions in aqueous solution.	You could learn more about this topic by: AQA A-level Chemistry – Inorganic chemistry specification AQA teaching guide: Reactions of metal ions in aqueous solution
Your teacher will assess your knowledge & understanding throughout the topic by looking at your work, questioning, discussion and giving you feedback in lots of different ways. The key pieces of work in this topic are: Whiteboard quizzes Exam Questions Homework End of topic Test	

KS5 Chemistry – 3.1.8 Thermodynamics

Learning about this topic is important because:
The further study of thermodynamics builds on the Energetics section and is important in understanding the stability of compounds and why chemical reactions occur. Enthalpy change is linked with entropy change enabling the free-energy change to be calculated.

This builds on:
3.1.4 Energetics- Hess Cycle, Bond energy Calculations and Calorimetry.

This leads to:
3.1.9 Rate Equations

We will learn:

- What are the definitions for each of the enthalpy changes and lattice enthalpy?
- Construct a 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
- Perform calculations of an enthalpy change using these cycles.
- 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.

We will develop/practise skills including:

- Constructing Born-Haber cycles to calculate missing values from known enthalpy changes.
- Placing known enthalpy values with the correct enthalpy change on a Born-Haber cycle.
- Calculating entropy changes from absolute entropy values.
- Explaining the difference between theoretical and experimental values for lattice formation and what this means in terms of covalent character.
- Rearranging the equation $\Delta G = \Delta H - T\Delta S$ to find unknown values.
- Determining ΔS and ΔH from a graph of ΔG versus T.

Some of the vocabulary that we will use includes:

Enthalpy, Entropy, Born-Haber, Gibbs Free Energy

You could learn more about this topic by:
Videos - Topic 1.08 Thermodynamics - AQA Chemistry A-level - PMT
1.8. Thermodynamics.pdf (physicsandmathstutor.com)
Flashcards - AQA Chemistry A-level Topic 1.8 Thermodynamics - PMT

Your teacher will assess your knowledge & understanding throughout the topic by looking at your work, questioning, discussion and giving you feedback in lots of different ways. The key pieces of work in this topic are:
End of topic assessed questions, retrieval questions, mini whiteboards, exercise questions.

KS5 Chemistry – 3.1.9 Rate Equations

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Learning about this topic is important because:
In rate equations, the mathematical relationship between rate of reaction and concentration gives information about the mechanism of a reaction that may occur in several steps.

This builds on:
3.1.5 Kinetics- Collision Theory and Factors which affect rate.

This leads to:
3.1.10 Equilibrium constant Kp for homogeneous systems

We will learn:

- 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 $k = Ae^{-Ea/RT}$
- Understand that the equation $k = Ae^{-Ea/RT}$ can be rearranged into the form $\ln k = -Ea/RT + \ln A$ and know how to use this rearranged equation with experimental data to plot a straight-line graph with slope $-Ea/R$
- 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.

We will develop/practise skills including:

- Students use given rate data and deduce a rate equation, then use some of the data to calculate the rate constant including units.
- Students use a graph of concentration–time and calculate the rate constant of a zero-order reaction by determination of the gradient.
- Students could determine the order of reaction for a reactant in the iodine clock reaction.
- Students could be given data to plot and interpret in terms of order with respect to a reactant. Alternatively, students could just be given appropriate graphs and asked to derive order(s).
- Students calculate the rate constant of a zero-order reaction by determining the gradient of a concentration–time graph.
- Students plot concentration–time graphs from collected or supplied data and draw an appropriate best-fit curve. Students draw tangents to such curves to deduce rates at different times.
- Required practical 7 Measuring the rate of reaction:
  - by an initial rate method
  - by a continuous monitoring method

Some of the vocabulary that we will use includes:
Rate Constant
Rate Equation
Order of Reaction
Rate Determining Step

You could learn more about this topic by:
1.09. Rate Equations.pdf (physicsandmathstutor.com)
Flashcards - AQA Chemistry A-level Topic 1.9 Rate Equations - PMT (physicsandmathstutor.com)
Videos - Topic 1.09 Rate Equations - AQA Chemistry A-level - PMT (physicsandmathstutor.com)

KS5 Chemistry – 3.1.9 Rate Equations

Your teacher will assess your knowledge & understanding throughout the topic by looking at your work, questioning, discussion and giving you feedback in lots of different ways. The key pieces of work in this topic are:
End of topic assessed questions, retrieval questions, mini whiteboards, exercise questions.

A2 Chemistry – 3.1.11 Electrode Potentials and Electrochemical Cells

West Kirby Grammar School

Learning about this topic is important because:
Electrochemical cells transfer electrons indirectly through an external circuit, creating a potential difference that can drive an electric current. This chemistry underpins portable batteries, rechargeable devices, electric vehicles and fuel-cell technologies.

This builds on:
GCSE Chemistry (electrolysis, simple cells, batteries and fuel cells)
3.1.2 Amount of Substance (concentration and standard conditions)
3.1.6 Chemical Equilibria (position of equilibrium and changing conditions)
3.1.7 Oxidation, Reduction and Redox Equations (oxidation states and half-equations)

This leads to:
3.2.5 Transition Metals (redox reactions and oxidation states); Required Practical 8; synoptic physical, inorganic and practical questions in Papers 1 and 3

We will learn:

- Electrode half-equations are written as reductions using the IUPAC convention.
- Cells are represented using conventional cell notation and electrode potentials are measured relative to the standard hydrogen electrode ($E^\circ = 0.00$ V).
- Standard electrode potentials use 298 K, 100 kPa and ion concentrations of 1.00 mol dm^{-3}; changing conditions changes the electrode potential.
- Standard electrode potentials form an electrochemical series and indicate the relative tendencies of species to be reduced or oxidised.
- Cells may be non-rechargeable, rechargeable or fuel cells; fuel cells generate current while fuel and oxidant are supplied and are not electrically recharged.
- The simplified lithium-cell reactions are $Li \rightarrow Li^+ + e^-$ and $Li^+ + CoO_2 + e^- \rightarrow Li^+[CoO_2]^-$.
- An alkaline hydrogen–oxygen fuel cell combines electrode reactions to give the overall reaction $2H_2 + O_2 \rightarrow 2H_2O$.

We will develop/practise skills including:

- Writing reduction half-equations and combining them to obtain balanced overall redox equations.
- Writing and interpreting conventional cell representations, including phase boundaries, the salt bridge and inert electrodes when required.
- Using E° values to identify oxidising and reducing agents and predict the direction of simple redox reactions.
- Calculating cell EMF using $E^\circ cell = E^\circ positive\ electrode - E^\circ negative\ electrode$.
- Setting up electrochemical cells, measuring EMF and evaluating the effects of concentration or temperature in Required Practical 8.
- Using electrode data to deduce reactions in non-rechargeable and rechargeable cells and explain how electron flow generates current.
- Writing alkaline hydrogen–oxygen fuel-cell electrode equations and evaluating benefits and risks of commercial cells to society.

Some of the vocabulary that we will use includes:
redox, half-cell, electrode, electrolyte, salt bridge, standard hydrogen electrode, standard electrode potential, E°, standard conditions, cell notation, EMF, electrochemical series, oxidising agent, reducing agent, feasibility, lithium cell, rechargeable, non-rechargeable, fuel cell

You could learn more about this topic by:
PMT - 1.11 Electrode Potentials and Cells Questions
Chemguide - Redox Equilibria

Your teacher will assess your knowledge and understanding throughout the topic by looking at your work, questioning, discussion and giving you feedback in lots of different ways. The key pieces of work in this topic are:
Required Practical 8, E° and EMF calculations, cell-notation and redox-prediction problems, commercial-cell questions, an end-of-topic assessed question, retrieval questions, mini whiteboards and exam questions.