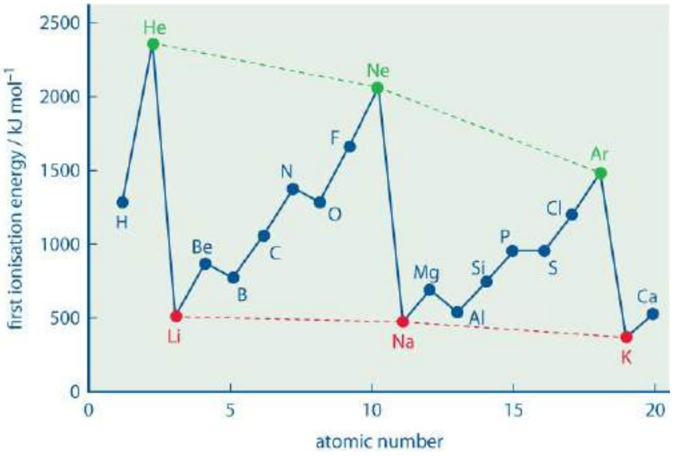
Unit 1: Periodicity

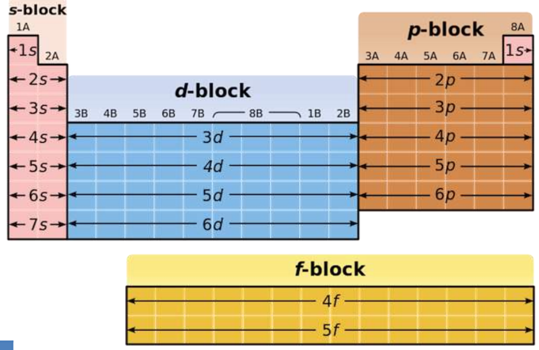
1. Keywords

Periodicity	The regular occurrence of elements with similar properties giving repeated patterns when compared to atomic number
First ionisation energy	The energy needed to remove 1 electron from each atom in 1 mole of gaseous atoms. $M(g) \rightarrow M^+(g) + e^-$
Second ionisation energy:	The energy needed to remove 1 electron from each atom in 1 mole of gaseous +1 ions. $M^+(g) \rightarrow M^{2+}(g) + e^-$
Successive ionisation energies:	Removing each electron in turn from a mole of gaseous atoms. Provides evidence of energy levels and orbitals
Atomic radius	a measure of the size of its atoms, usually the mean distance from the centre of the nucleus to the outermost electron shell



2. The periodic table.

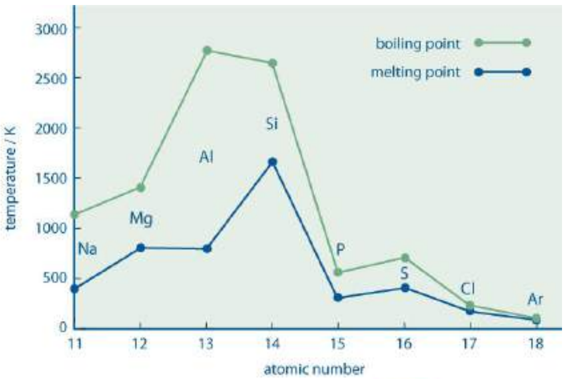
An element is classified as s, p, d or f block according to its position in the Periodic Table, which is determined by its proton number so on the nature of the orbital into which the last electron of the atom enters.



3. Melting and boiling point across period 3

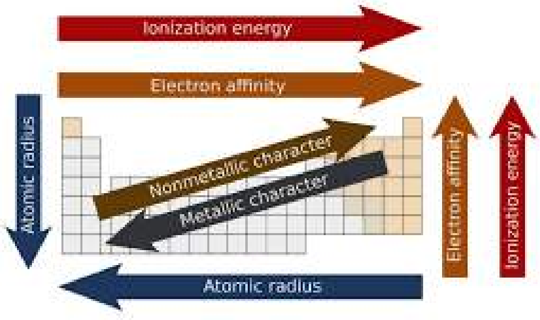
Trends in melting and boiling point are explained based on the intermolecular forces between the molecules.

	Structure	MP/BP
Metal (Na,Mg,Al)	Electrostatic	Low
Semi Metal (Si)	Macromolecular, strong covalent bonds	High
Non Metal (P_4,S_8,Cl_2)	Small molecular substances, weak van der Waals, depend on size of molecule	Low
Noble gas (Ar)	Monoatomic, weak van der Waals	Low



4. Atomic radius

Decreases across period due to increased nuclear charge attracting shells.
Increases down group due to more shells



5. Ionisation energy

Ionisation energy increases across a period as similar distance from nucleus, same shielding but increased nuclear charge so stronger attraction.
Group 3 drop due to p orbital higher energy than s.
Group 6 drop due to repulsion of p orbital pair
Ionisation energy decreases going down a group as more shielding and greater atomic radius.

Unit 2: Group II

1. Physical properties

Trends in group II, Beryllium is not typical of the group and it is not considered here.

Atomic radius	Increases	More shells
Melting point	Decreases	Weaker metallic bonds due to larger atoms
First ionisation energy	Decreases	Outer electron in higher energy level, larger atom, more shells more shielding

2. Reactivity with water

Reactivity with water INCREASES down the group.

Magnesium reacts slowly with liquid water, but rapidly when heated in the presence of steam.

$\overset{0}{M}(s) + 2\overset{+1}{H}_2O(l) \rightarrow \overset{+2}{M}(OH)_2(aq) + \overset{0}{H}_2(g)$

All equations similar down group: beware state symbol of hydroxide

3. Hydroxides and sulphates - solubility

Hydroxides – $M(OH)_2$

- Varying solubility in water.
- Solubility INCREASES as you descend the group.
- pH of the hydroxide in water varies.
- pH increases as you descend the group.

Solubility of $Mg(OH)_2$ described as sparingly soluble

Sulphates – MSO_4

- Colourless solids
- Solubility DECREASES as you descend the group.
- Carbonates thermally decompose to form $MO(s)$ and $CO_2(g)$.
- Thermal stability increases as you descend the group.

$BaSO_4$ is described as insoluble

4. Application of group II compounds

Chemical	Common name	Applications
$Mg(OH)_2$	Milk of magnesia	Treat indigestion, heartburns and wind.
$Ca(OH)_2$	Slaked lime	Neutralise fields and polluted lakes.
CaO and $CaCO_3$		Remove SO_2 from flue gases
$BaSO_4$	Barium meal	Contrast medium for gut X-ray.
$BaCl_2$		Test for sulfate ions: acidified with hydrochloric acid to remove carbonate ions

Unit 3a: Group VII

1. Keywords

Mean bond Enthalpy:	The average enthalpy change when one mole of a specific bond is broken in a range of different gaseous compounds.
Displacement:	A displacement reaction is a type of reaction in which part of one reactant is replaced by another reactant.
Electronegativity:	The power of an atom to attract the electrons in a covalent bond.

2. Physical properties

Trends in group VII, Fluorine often does strange things due to its size and electron density

Atomic radius	Increases	More shells
Boiling point	Increases	Larger molecules have stronger van der Waals forces which need more energy to break
Electronegativity	Decreases	Larger atoms so bonding pair further from nucleus, less strongly attracted
Bond enthalpy of X_2 and HX	Decreases	Atoms larger, bond longer, less energy to break

3. Physical states

The physical state of the halogens are summarised below.

Symbol	In pure form	In non-polar solvents	In water
F	Pale yellow gas	(Reacts with solvents)	(Reacts with water)
Cl	Pale green gas	Pale green solution	Pale green solution
Br	Dark red liquid	Orange solution	Orange solution
I	Grey solid	Purple solution	Insoluble

4. Oxidising abilities – Displacement reactions

The oxidising ability of the halogens decreases down the group. You cannot investigate the oxidising ability of Fluorine in aqueous solution because it reacts with water.

	F^-	Cl^-	Br^-	I^-
F_2	-	yes	yes	yes
Cl_2	no	-	yes	yes
Br_2	no	no	-	yes
I_2	no	no	no	-

Unit 3b: Reactivity of halide ions and chlorine.

1. Keywords

Disproportionation:	a reaction in which a substance is simultaneously oxidised and reduced, giving two different products.
Precipitate (ppt):	deposited solid formed in a solution.

2. Reducing strength

The reducing ability of the halide ions increases down the group.

Symbol	Atomic radius (nm)
F^-	0.133
Cl^-	0.180
Br^-	0.195
I^-	0.215

4. Reaction with silver ions

All metal halides (except fluoride) react with silver ions to form an insoluble precipitate. Dilute nitric acid is added before the reaction to get rid of any carbonate or hydroxide impurities.

$Ag^+(aq) + X^-(aq) \rightarrow AgX(s)$

Symbol	Observation	Halide salt solubility
Cl^-	White ppt	Dilute NH_3
Br^-	Cream ppt	Concentrated NH_3
I^-	Pale yellow ppt	Insoluble in NH_3

3. Reaction with concentrated sulphuric acid

Solid halides react with concentrated sulphuric acid giving the hydrogen halide. This then reacts with more acid to give different products based on their reducing powers. Iodide is more reducing so gives more products

Halide ion	Reaction A to form $MeHSO_4$	Reaction B to form SO_2	Reaction C to form S	Reaction D to form H_2S	Observations
Cl^-	✓	×	×	×	- Steamy fumes (HCl) - ppt ($MeHSO_4$)
Br^-	✓	✓	×	×	- Steamy fumes (HBr) - Brown fumes (Br_2) - Pungent gas (SO_2)
I^-	✓	✓	✓	✓	- Steamy fumes (HI) - Black ppt (I_2) - Rotten egg smell (H_2S) - Yellow ppt (S)

Reaction A	$MeX(s) + H_2SO_4(l) \rightarrow MeHSO_4(s) + HX(g)$	
Reaction B	$2H^+(aq) + 2X^-(aq) + H_2SO_4(l) \rightarrow SO_2(g) + 2H_2O(l) + X_2(l)$	Acid-base
Reaction C	$6H^+(aq) + 6X^-(aq) + H_2SO_4(l) \rightarrow S(s) + 4H_2O(l) + 3X_2(s)$	Acid-base and redox
Reaction D	$8H^+(aq) + 8X^-(aq) + H_2SO_4(l) \rightarrow H_2S(g) + 4H_2O(l) + 4X_2(s)$	Acid-base and redox

5. Reactivity of chlorine

Reactivity with water:	
$Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$ "chlorine water“	Disproportionation reaction
Reactivity with water in sunlight:	
$2Cl_2(g) + 2H_2O(l) \rightarrow 4HCl(g) + O_2(g)$	Goes from pale green to colourless
Alternative chlorination of swimming pools:	
$NaClO(s) + H_2O(l) \rightleftharpoons Na^+(aq) + OH^-(aq) + HClO(aq)$	Water is kept slightly acidic
Reactivity with alkali:	
$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$	Disproportionation reaction

Unit 4: Period 3

1. Keywords

Amphoteric:	able to react both as a base and as an acid.

2. Reactivity

Reactivity with water (only Na, Mg, Cl_2):

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$ Faster with steam

$Cl_2(g) + H_2O(l) \rightleftharpoons HCl(aq) + HClO(aq)$

Reactivity with oxygen:

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$	Bright yellow flame, white solid sodium oxide
$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$	Bright white flame, white solid magnesium oxide.
$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$	Bright white flame in a gas jar, white solid Al_2O_3.
$Si(s) + O_2(g) \rightarrow SiO_2(s)$	Reacts only when heated.
$4P(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$	White flame and white fumes
$S(s) + O_2(g) \rightarrow SO_2(g)$	Blue flame, smelly gas

3. Period 3 oxides melting points

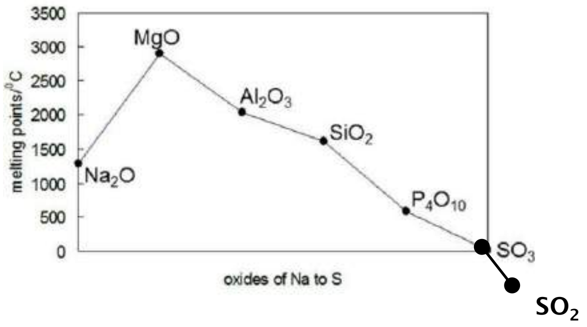
Na_2O/1548 K	Strong attraction between oppositely charged ions in ionic lattice = high melting point
MgO/ 3125K	Smaller and more highly charged than Na so stronger attraction to the oxide ions
Al_2O_3 /2345K	Ionic structure with covalent character. Very high charge density on the small Al
SiO_2/1883K	Giant covalent structure. Strong covalent bonds Si-O.
P_4O_{10} /573K	Simple molecular compounds. Weak van der Waals = low melting point
SO_3/290K	Simple molecule Weak van der Waals = low melting point
SO_2/200K	Simple molecule. Weak van der Waals = low melting point

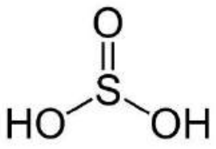
4. Reactivity of Period 3 oxides with acids/bases

Na_2O	$Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$	Alkali
MgO	$MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$	Alkali
Al_2O_3	$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$ $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$	Amphoteric
SiO_2	$SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(l)$	Weak acid
P_4O_{10}	$P_4O_{10}(s) + 12NaOH(aq) \rightarrow 4Na_3PO_4(aq) + 6H_2O(l)$	Acid
SO_2	$SO_2(g) + 2NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$	Acid
SO_3	$SO_3(g) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$	Acid

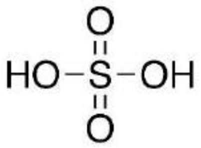
5. Reactivity of Period 3 oxides with water

Na_2O	$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$	pH = 14
MgO	$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$	pH = 9
Al_2O_3	Insoluble no reaction	pH = 7
SiO_2	Insoluble	pH = 7
P_4O_{10}	$P_4O_{10}(s) + 6H_2O\ (l) \rightarrow 4H_3PO_4(aq)$	pH = 1-2
SO_2	$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$	pH = 2-3
SO_3	$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$	pH = 0-1

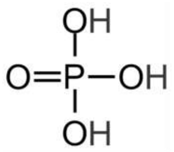




Sulfuric(IV) acid
H_2SO_3



Sulfuric(VI) acid
H_2SO_4



Phosphoric(V) acid
H_3PO_4

Unit 5a: Transition metals

1. Keywords

Complex ions:	central metal atom or ion surrounded by ligands, which are held by dative covalent bonds.
Coordination number:	is the number of co-ordinate bonds to the central metal atom or ion.
Ligand:	is an ion or molecule that forms a co-ordinate bond with a transition metal by donating a pair of electrons
Chelate:	compound containing a ligand bonded to a central metal atom at two or more points.
Coordinate (dative) bond	A type of covalent bond where both electrons are donated by one atom

2. Classification

A transition metal is an element that forms at least one stable ion with a partially full d-shell.

- When the fourth period metals form ions the 4s electrons are the first to be lost then 3d electrons.
- Scandium and Zinc are not considered to be transition metals. Scandium forms only a 3+ ion [Ar] $4s^0$ $3d^0$
- Zinc forms only a 2+ ion [Ar] $4s^0$ $3d^{10}$
- Copper is a transition metal because its +2 ion has an incomplete d orbital. [Ar] $4s^0$ $3d^9$

Chemical properties
- Form coloured ions
- Form complexes
- Catalysts
- Variable oxidation state

Physical properties
- Shiny
- Strong
- Hard
- Good conductors

3. Complex ion formation

All transition metals for co-ordinate bonds by accepting electron pairs from other ions or molecules generally called ligands.
Different ligands form different strength bonds.

Monodentate ligands (single co-ordinate bond):
- H_2O (Aqua) } small, similar size
- NH_3 (Ammino) } and uncharged
- OH^- (Hydroxy)
- CN^-
- Cl^- large

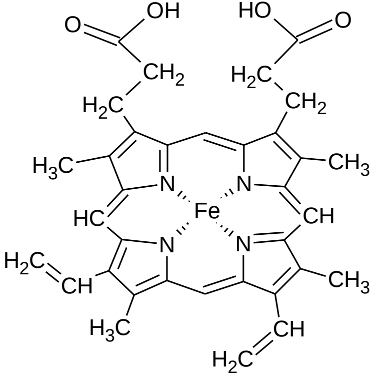
Bidentate (two atoms that each donate a lone pair of electrons to form co-ordinate bonds):
- $H_2NCH_2CH_2NH_2$ ethane-1,2-diamine
- $C_2O_4^{2-}$ ethanedioate

Multidentate (many co-ordinate bonds):
- $EDTA^{4-}$ which can form 6 co-ordinate bonds

The charge of the complex ion depends on the charge of the transition metal and on the charge and number of the ligand.

4. Haem

Haem is an iron(II) complex with a multidentate ligand. Oxygen forms a co-ordinate bond to Fe(II) in haemoglobin, enabling oxygen to be transported in the blood. Carbon monoxide is toxic because it replaces oxygen co-ordinately bonded to Fe(II) in haemoglobin.

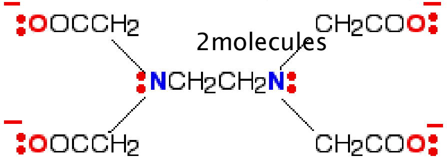


5. Chelate effect

When bidentate or multidentate ligands take the place of monodentate, there are more products than reactants, the number of particles increases, hence the entropy of system increases. The products are thermodynamically more stable.
The enthalpy change of these reaction is very small but the increase in entropy makes the reaction possible.

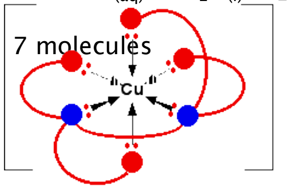
$[Cu(H_2O)_6]^{2+}_{(aq)} + EDTA^{4-}_{(aq)} \leftrightharpoons [CuEDTA]^{2-}_{(aq)} + 6H_2O_{(l)}$

2molecules



the EDTA^{4-} ion

7 molecules



the [Cu(EDTA)]$^{2-}$ ion

Gibbs free-energy equation

$\Delta G = \Delta H - T\Delta S$

Entropy increases = favourable reaction

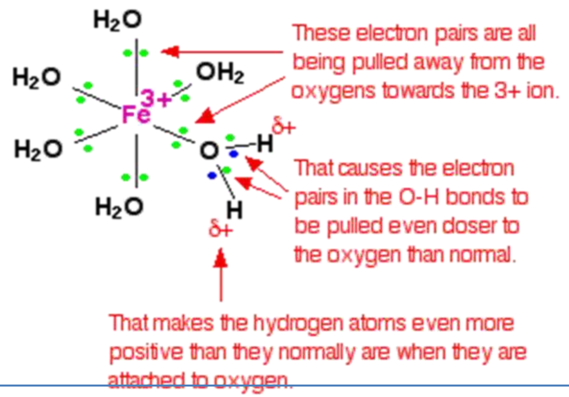
Unit 5b: Reactivity of complex ions

1. Keywords

Amphoteric	dissolves in both acids and bases so therefore has both acidic and basic character
Hydrolysis	the chemical breakdown of a compound due to reaction with water.
Metal-aqua ion	The complex ion formed when water molecules form coordinate bonds with a metal ion

2. Acidity

The acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$
M^{3+} smaller and more positive than M^{2+}
M^{3+} has greater charge density so more strongly attracts electrons on metal-oxygen bond
More polarisation of water molecules.
H^+ more easily lost as OH bond weaker



These electron pairs are all being pulled away from the oxygens towards the 3+ ion.

That causes the electron pairs in the O-H bonds to be pulled even closer to the oxygen than normal.

That makes the hydrogen atoms even more positive than they normally are when they are attached to oxygen.

3. Reactions with OH^-

Reaction with OH- to produce precipitate is called hydrolysis- often indicated by use of EXCESS NaOH
Reactions are equilibria so can get single and multiple substitutions
If precipitate not produced, product is aqueous and charged
Only aluminium precipitate dissolves- amphoteric: acting as an acid

M^{3+} aqua ion	$[M(H_2O)_6]^{3+}{}_{(aq)} + 3OH^-{}_{(aq)} \rightarrow M(H_2O)_3(OH)_3{}_{(s)} + 3H_2O_{(l)}$
M^{2+} aqua ion	$[M(H_2O)_6]^{2+}{}_{(aq)} + OH^-{}_{(aq)} \rightarrow M(H_2O)_4(OH)_2{}_{(s)} + H_2O_{(l)}$
Ammonia ppt dissolving	$Al(H_2O)_3(OH)_3{}_{(s)} + OH^-{}_{(aq)} \rightarrow [Al(H_2O)_2(OH)_4]^-{}_{(aq)}$

4. Reactions with NH_3

Ammonia in water releases OH^-
$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+ + OH^-{}_{(aq)}$
Reactions the same as with OH^-
Aluminium precipitate does not dissolve
Copper precipitate dissolves- ligand exchange

M^{3+} aqua ion	$[M(H_2O)_6]^{3+}{}_{(aq)} + 3OH^-{}_{(aq)} \rightarrow M(H_2O)_3(OH)_3{}_{(s)} + 3H_2O_{(l)}$
M^{2+} aqua ion	$[M(H_2O)_6]^{2+}{}_{(aq)} + OH^-{}_{(aq)} \rightarrow M(H_2O)_4(OH)_2{}_{(s)} + H_2O_{(l)}$
Copper ppt dissolving	$Cu(H_2O)_4(OH)_2{}_{(s)} + 4NH_3{}_{(aq)} \rightarrow [Al(H_2O)_2(NH_3)_4]^-{}_{(aq)} + 2OH^-{}_{(aq)} + 2H_2O_{(l)}$

5. Reactions with Na_2CO_3

M^{3+} ions fizz due to CO_2 released- they are acting as an acid
M^{2+} ions do not fizz

M^{3+} aqua ion	$2[M(H_2O)_6]^{3+}{}_{(aq)} + 3CO_3^{2-}{}_{(aq)} \rightarrow 2M(H_2O)_3(OH)_3{}_{(s)} + 3H_2O_{(l)} + 3CO_2{}_{(g)}$
M^{2+} aqua ion	$[M(H_2O)_6]^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow MCO_3{}_{(s)} + 6H_2O_{(l)}$

6. Amphoteric aluminium

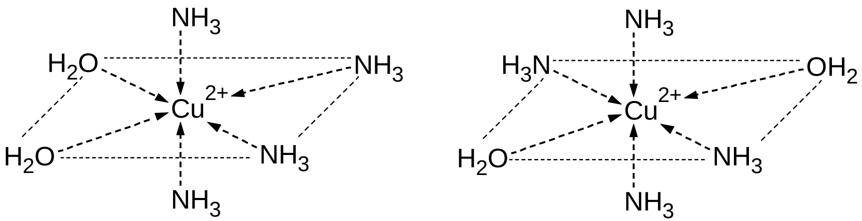
Acting as an acid	$Al(H_2O)_3(OH)_3{}_{(s)} + OH^-{}_{(aq)} \rightarrow [Al(H_2O)_2(OH)_4]^-{}_{(aq)}$
Acting as a base.	$Al(H_2O)_3(OH)_3{}_{(s)} + 3HCl_{(aq)} \rightarrow [Al(H_2O)_6]^{3+}{}_{(aq)} + 3Cl^-{}_{(aq)}$ $Al(H_2O)_3(OH)_3{}_{(s)} + H^+{}_{(aq)} \rightarrow [Al(H_2O)_4(OH)_2]^+{}_{(aq)} + H_2O_{(l)}$

Unit 5c: Complex ions

1. Keywords

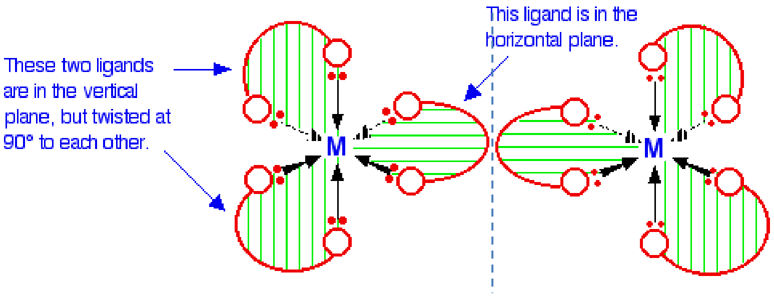
Cis trans isomerism	when an octahedral or square planar complex with monodentate ligands has two molecules with ligands in different position in space
Cis isomer	ligands next to each other
Trans isomer	ligands opposite each other
Optical isomerism:	non-superimposable mirror images with bidentate ligands, particularly $C_2O_4^{2-}$ or $NH_2CH_2CH_2NH_2$

Cis trans isomers



cis **trans**

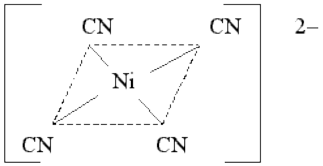
Optical isomers

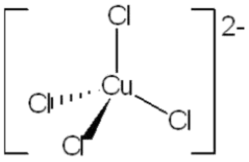


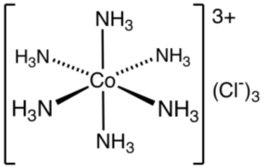
2. Complex ion shapes

The coordination number dictates the shape of the complex ions.

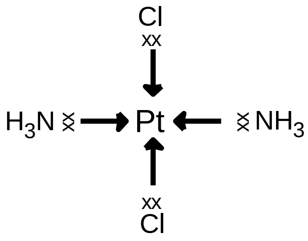
Coordination number	2	4	6
Shape	Linear	Tetrahedral or square planar	Octahedral



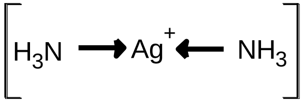




Square planar can show cis-trans isomerism if two different ligands present



Cisplatin



The silver complex in Tollens' reagent is linear

Unit 5d: Properties of complexes

1. Coloured ion formation

Ions with d^{10} or d^0 are colourless, those partly filled tend to be coloured.

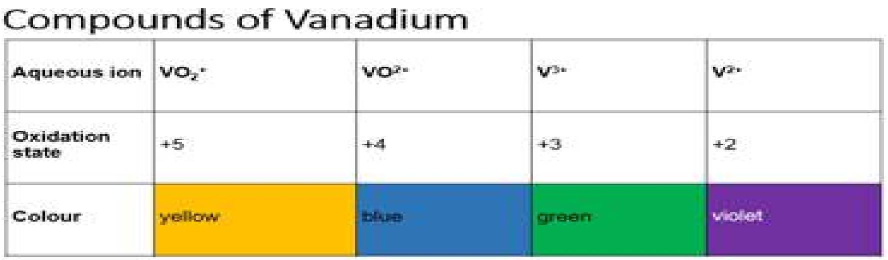
Colour arises when d electrons in the ground state absorb energy from visible light which promotes them to an excited state: a higher 3d sublevel. The rest of the frequencies are transmitted or reflected; hence the complementary colour is observed via transmission or reflection.

Factors affecting colour by altering ΔE:
- Size & type of ligand
- Oxidation state
- Complex shape
- Coordination number

The energy difference between the ground state and the excited state of the d electrons is given by

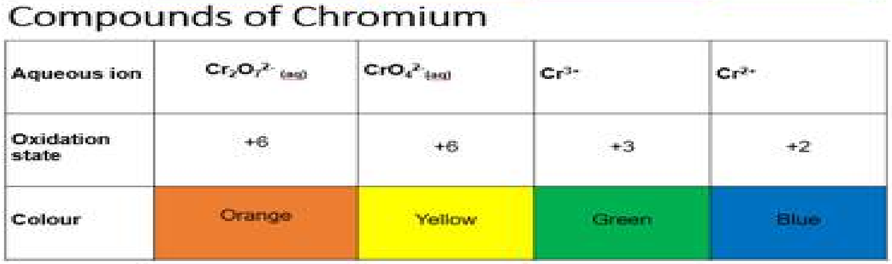
$$\Delta E = h\nu = \frac{hc}{\lambda}$$

ΔE = energy absorbed (J)
c = speed of light (3.00×10^{8} m s^{-1})
h = Planck's constant (6.63×10^{-34} J s)
λ = wavelength of light absorbed (m)
ν = frequency of light absorbed (Hz)



Compounds of Vanadium

Aqueous ion	VO_2^+	VO^{2+}	V^{3+}	V^{2+}
Oxidation state	+5	+4	+3	+2
Colour	yellow	blue	green	violet



Compounds of Chromium

Aqueous ion	$Cr_2O_7^{2-}$ (aq)	CrO_4^{2-} (aq)	Cr^{3+}	Cr^{2+}
Oxidation state	+6	+6	+3	+2
Colour	Orange	Yellow	Green	Blue

Colorimetry: we can then determine the concentration by comparing our sample against known concentrations of the same metal ion / ligand.
1. Make up solutions of known concentration.
2. Measure absorption of each known solution
3. Plot a graph of absorption vs concentration
4. Measure absorption of unknown and compare on the graph to give a concentration reading

2. Variable oxidation state

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 $[Ag(NH_3)_2]^+$ (Tollens' reagent) to metallic silver is used to distinguish between aldehydes and ketones.

Since many transition metal compounds are coloured, we can use colour changes as self indicators in titrations.
The titration using $KMnO_4$ is a typical redox titration.
The end point is the first permanent pink i.e. the first sight of the MnO_4^- ion. The pink colour is not due to phenolphthalein!
Typical scenarios which appear in exam questions are: concentration of iron solutions, % iron in iron tablets and % iron in a steel sample.

$$Mn^{2+} + 4H_2O + 5Fe^{3+} \rightarrow MnO_4^- + 8H^+ + 5Fe^{2+}$$

$$2MnO_4^- + 16H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$$

Unit 5e: Catalysis

1. Keywords

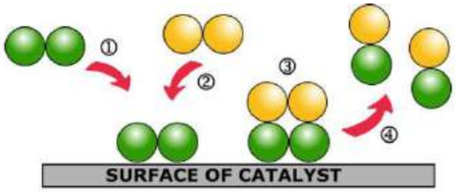
Catalyst:	is a substance that increases the rate of a chemical reaction by providing an alternative route with a lower activation energy. The catalyst is unchanged by the reaction.
Heterogeneous catalyst:	catalyst is in a different phase from the reactants and the reaction occurs at active sites on the surface
Homogeneous catalyst:	is in the same phase as the reactants. The reaction happens via an intermediate species.
Support medium	used to maximise the surface area of a catalyst and minimise cost
Catalytic poisoning	impurities blocking the active sites of a heterogeneous catalyst, reducing efficiency. Catalyst can also be lost from the support. This has a cost implication.
Autocatalysis	where the catalyst is a product of the reaction. As a result, the rate of reaction increases over time as more product, and therefore catalyst, is produced.

2. Catalytic activity

Transition metals are good as catalysts either due to a change in oxidation state or they adsorb other substances onto the surface. A partially filled d orbital can be used to form bonds with adsorbed reactants.

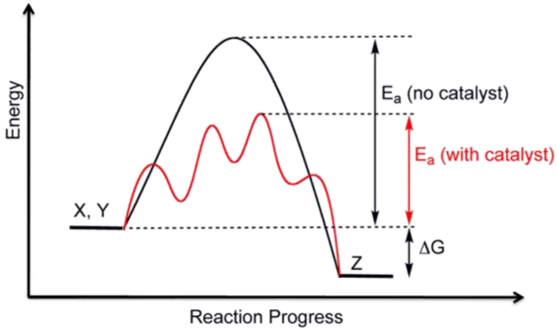
Heterogeneous examples
Haber process: $3H_2 + N_2 \rightleftharpoons 2NH_3$
Catalyst Fe

Contact process: $SO_2 + \frac{1}{2}O_2 \rightarrow SO_3$
Catalytic cycle:
$SO_2 + V_2O_5 \rightarrow SO_3 + V_2O_4$
$V_2O_4 + \frac{1}{2}O_2 \rightarrow V_2O_5$
Catalyst vanadium(V) oxide, vanadium changes oxidation state



1 Adsorption of first reactant onto catalyst active site
2 Adsorption of second reactant onto catalyst active site
3 Bonds weaken, reaction takes place
4 Products leave surface

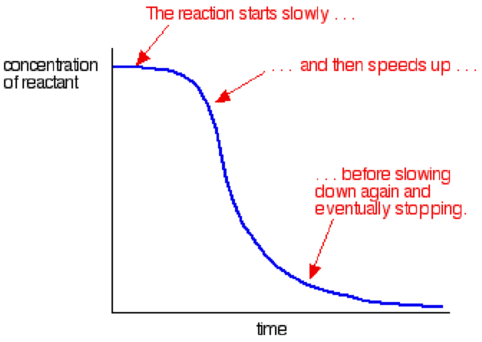
Homogeneous example
$2I^- + S_2O_8^{2-} \rightarrow 2SO_4^{2-} + I_2$
Needs catalyst due to repulsion of two negative ions, high activation energy
Catalytic cycle:
$2Fe^{2+} + S_2O_8^{2-} \rightarrow 2SO_4^{2-} + 2Fe^{3+}$
$2Fe^{3+} + 2I^- \rightarrow I_2 + 2Fe^{2+}$
Catalyst Fe^{2+} changes oxidation state



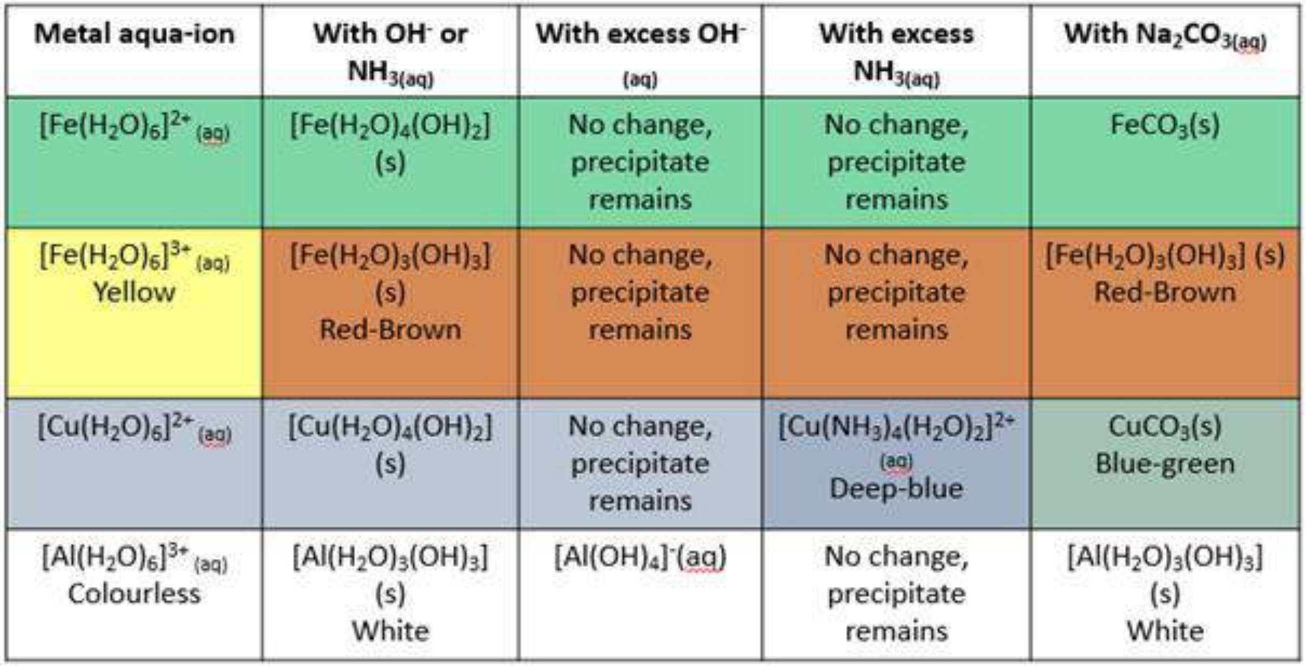
3. Autocatalysis

Autocatalysis example
$2MnO_4^- + 6H^+ + 5C_2O_4^{2-} \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

The reaction starts slowly due to no catalyst.
Catalyst Mn^{2+} speeds up reaction when produced.
Rate falls when reactants used up.



Unit 6: Reactions of ions in aqueous solution



Metal aqua-ion	With OH^- or $NH_{3(aq)}$	With excess $OH^-_{(aq)}$	With excess $NH_{3(aq)}$	With $Na_2CO_{3(aq)}$
$[Fe(H_2O)_6]^{2+}_{(aq)}$	$[Fe(H_2O)_4(OH)_2]$ (s)	No change, precipitate remains	No change, precipitate remains	$FeCO_3$(s)
$[Fe(H_2O)_6]^{3+}_{(aq)}$ Yellow	$[Fe(H_2O)_3(OH)_3]$ (s) Red-Brown	No change, precipitate remains	No change, precipitate remains	$[Fe(H_2O)_3(OH)_3]$ (s) Red-Brown
$[Cu(H_2O)_6]^{2+}_{(aq)}$	$[Cu(H_2O)_4(OH)_2]$ (s)	No change, precipitate remains	$[Cu(NH_3)_4(H_2O)_2]^{2+}_{(aq)}$ Deep-blue	$CuCO_3$(s) Blue-green
$[Al(H_2O)_6]^{3+}_{(aq)}$ Colourless	$[Al(H_2O)_3(OH)_3]$ (s) White	$[Al(OH)_4]^-$(aq)	No change, precipitate remains	$[Al(H_2O)_3(OH)_3]$ (s) White

Colours of Vanadium oxidation states