

Transition Metals

General properties of transition metals

transition metal characteristics of elements Ti → Cu arise from an **incomplete d sub-level** in atoms or ions

these characteristics include

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

Sc $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$
 Ti $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$
 V $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$
 Cr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
 Mn $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$
 Fe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
 Co $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$
 Ni $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
 Cu $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
 Zn $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

→
When forming ions lose 4s before 3d

Sc $3+ [Ar] 4s^0 3d^0$
 Ti $3+ [Ar] 4s^0 3d^1$
 V $3+ [Ar] 4s^0 3d^2$
 Cr $3+ [Ar] 4s^0 3d^3$
 Mn $2+ [Ar] 4s^0 3d^5$
 Fe $3+ [Ar] 4s^0 3d^5$
 Co $2+ [Ar] 4s^0 3d^7$
 Ni $2+ [Ar] 4s^0 3d^8$
 Cu $2+ [Ar] 4s^0 3d^9$
 Zn $2+ [Ar] 4s^0 3d^{10}$

Why are Zn and Sc not transition metals?

Zn can only form a +2 ion. In this ion the Zn^{2+} has a **complete** d orbital and so does not meet the criteria of having a incomplete d orbital in one of its compounds.

Sc only forms a +3 ion with the electronic structure

The Sc^{3+} ion had an **empty** d orbital and so also does not meet the criteria

Complex formation

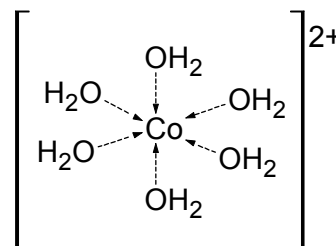
complex : is a central metal ion surrounded by ligands.

ligand.: An atom, ion or molecule which can donate a **lone electron pair**

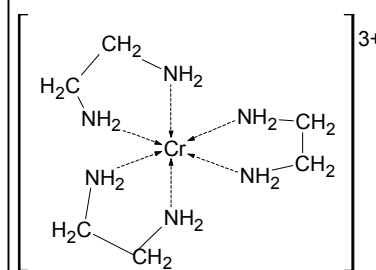
Co-ordinate bonding is involved in complex formation.

Co-ordinate bonding is when **the shared pair of electrons** in the covalent bond come from **only one of the bonding atoms**.

Co-ordination number: The number of co-ordinate bonds formed to a central metal ion



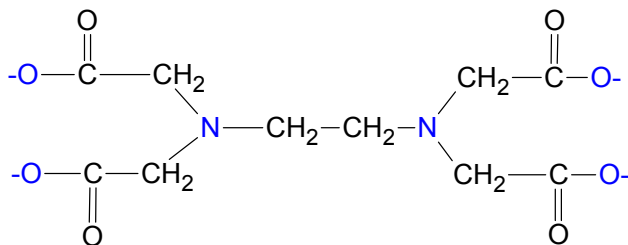
ligands can be **unidentate** (e.g. H_2O , NH_3 and Cl^-) which can form one coordinate bond per ligand or **bidentate** (e.g. $NH_2CH_2CH_2NH_2$ and ethanedioate ion $C_2O_4^{2-}$) which have two atoms with lone pairs and can form two coordinate bonds per ligand or **multidentate** (e.g. $EDTA^{4-}$ which can form six coordinate bonds per ligand).



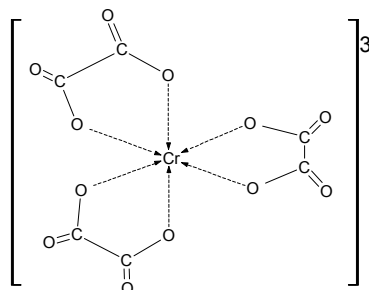
A complex with bidentate ligands e.g. $[Cr(NH_2CH_2CH_2NH_2)_3]^{3+}$

It still has a coordination number of 6

The $EDTA^{4-}$ anion has the formula

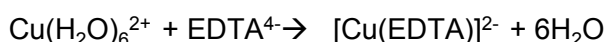
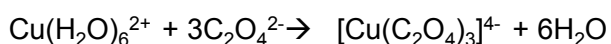
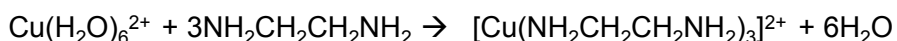


with six donor sites(4O and 2N) and forms a 1:1 complex with metal(II) ions



A complex with bidentate ethanedioate ligands e.g. $[Cr(C_2O_4)_3]^{3-}$

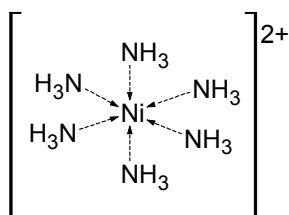
Equations to show formation of bidentate and multidentate complexes



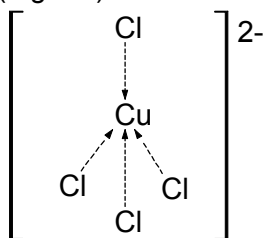
Learn the two bidentate ligands mentioned above but it is not necessary to remember the structure of EDTA

haem is an iron(II) complex with a multidentate ligand.

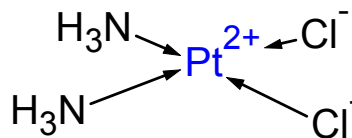
transition metal ions
commonly form **octahedral**
complexes with small
ligands (e.g. H_2O and NH_3).



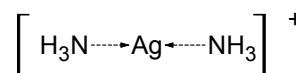
transition metal ions commonly form **tetrahedral** complexes with larger ligands (e.g. Cl^-).



square planar
complexes are also
formed, e.g.
cisplatin

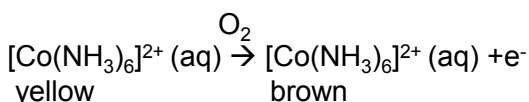


Ag⁺ commonly forms **linear** complexes
e.g. [Ag(NH₃)₂]⁺,
[Ag(S₂O₃)₂]³⁻ and
[Ag(CN)₂]⁻ (all
colourless).

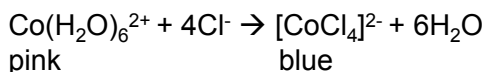


Colour changes arise from changes in

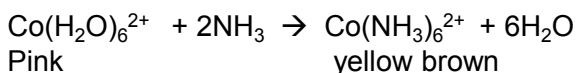
1. oxidation state,
2. co-ordination number
3. ligand.



In this equation only oxidation state is changing



In this equation both ligand and coordination number are changing

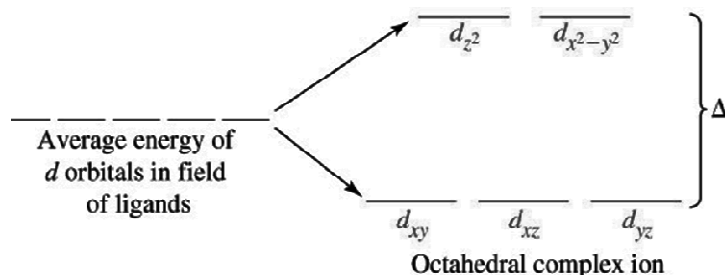


In this equation only the ligand is changing

Colour arises from **electronic transitions** from the **ground state to excited states**: between different d orbitals

$$\Delta E = h\nu. \quad (\nu = \text{frequency of light emitted})$$

A **portion of visible light is absorbed** to promote d electrons to higher energy levels. The light that is not absorbed is transmitted to give the substance colour



Ligands cause 5 d orbitals to split into two energy levels

The energy needed to excite electrons to a higher level depends on the oxidation state of the metal and the type of ligand.

Scandium is a member of the d block, its ion (Sc^{3+}) hasn't got any d electrons left to move around. So there is not an energy transfer equal to that of visible light

In the case of Zn^{2+} ions and Cu^+ ions the d shell is full e.g. $3d^{10}$ so here there is no space for electrons to transfer. So there is not an energy transfer equal to that of visible light

If visible light of increasing frequency is passed through a sample of a coloured complex ion, some of the light is absorbed.

The amount of light absorbed is proportional to the concentration of the absorbing species
Some complexes have only pale colours and do not absorb light strongly. In the cases a suitable ligand is added to intensify the colour.

absorption of visible light is used in spectrometry to determine the concentration of coloured ions

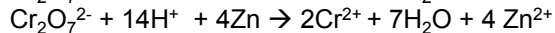
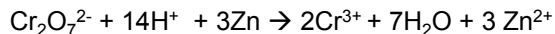
method

- Add an appropriate ligand to intensify colour
- Make up solutions of known concentration
- Measure absorption or transmission
- Plot graph of results or calibration curve
- Measure absorption of unknown and compare

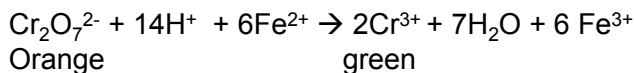
Transition elements show variable oxidation states

- Relative stability of +2 state with respect to +3 state increases across the series
- Compounds with high oxidation states tend to be oxidising agents e.g. MnO_4^-
- Compounds with low oxidation states are often reducing agents e.g. V^{2+} & Fe^{2+}

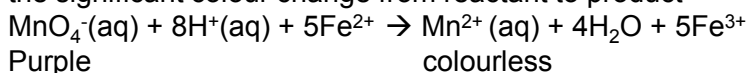
Cr^{3+} (green) and Cr^{2+} (blue) are formed by reduction of $\text{Cr}_2\text{O}_7^{2-}$ (orange) by **zinc** in acid solution, but **Fe^{2+}** will **only** reduce it to



The Fe^{2+} and $\text{Cr}_2\text{O}_7^{2-}$ in acid solution reaction can be used as a quantitative redox titration. This does not need an indicator



The redox titration between Fe^{2+} with MnO_4^- (purple) is a very common exercise. This titration is self indicating because of the significant colour change from reactant to product



Only use **dilute sulphuric acid** for manganate titration

It cannot be conc HCl as the Cl^- ions would be oxidised to Cl_2 by MnO_4^-

It cannot be nitric acid as this is an oxidising agent. It oxidises Fe^{2+} to Fe^{3+}

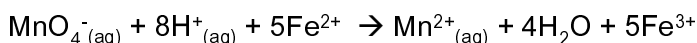
It cannot be conc H_2SO_4 as this is an oxidising agent.
It cannot be ethanoic acid as this is a weak acid and
not supply the large amount of acid needed (8H^+)

be able to perform calculations for these titrations and for others when the reductant and its oxidation product are given.

A 2.41g nail made from an alloy containing iron is dissolved in 100cm³ acid. The solution formed contains Fe(II) ions.

10cm³ portions of this solution are titrated with potassium manganate (VII) solution of 0.02M. 9.80cm³ of KMnO₄ were needed to react with the solution containing the iron.

What is the percentage of Iron by mass in the nail?



Step1 : find moles of KMnO_4

$$\text{moles} = \text{conc} \times \text{vol}$$

$$0.02 \times 9.8/1000$$

$$= 1.96 \times 10^{-4} \text{ mol}$$

Step 2 : find moles Fe^{2+} in 10cm^3

$$= \text{moles of KMnO}_4 \times 5$$

$$= 9.8 \times 10^{-4} \text{ mol}$$

Step 3 : find moles Fe^{2+} in 100cm^3

$$= 9.8 \times 10^{-4} \text{ mol} \times 10$$

$$= 9.8 \times 10^{-3} \text{ mol}$$

Step 4 : find mass of Fe in 9.8×10^{-3} mol

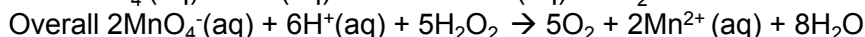
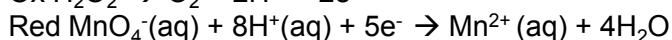
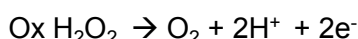
$$\text{mass} = \text{moles} \times \text{RAM} = 9.8 \times 10^{-3} \times 55.8 = 0.547\text{g}$$

Step 5 ; find % mass

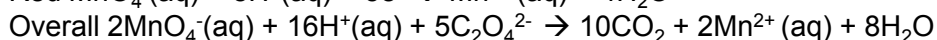
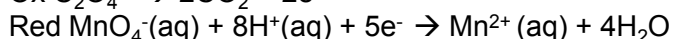
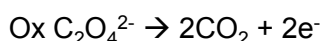
$$\% \text{mass} = 0.547 / 2.41 \times 100$$

$$= 22.6\%$$

With hydrogen peroxide



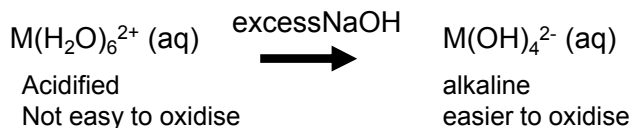
With ethanedioate



Oxidation in alkaline solution

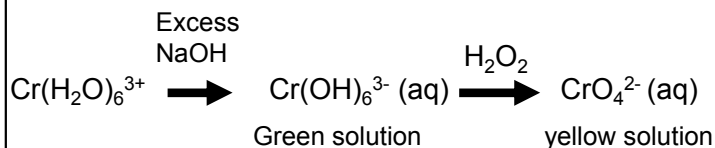
When transition metals in low oxidation states are in alkaline solution they are more easily oxidised than when in acidic solution

It is easier to remove an electron from a negatively charged ion



The metal ions can be oxidised by using oxidising agents such as hydrogen peroxide and sometimes by standing in air

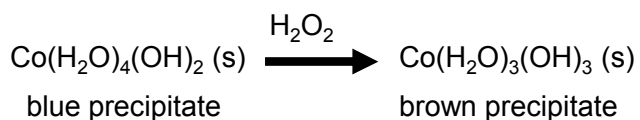
Chromium and cobalt compounds can be oxidised by using the **oxidising agent** hydrogen peroxide



Reduction: $\text{H}_2\text{O}_2 + 2\text{e}^- \rightarrow 2\text{OH}^-$

Oxidation: $\text{Cr(OH)}_6^{3-} + 2\text{OH}^- \rightarrow \text{CrO}_4^{2-} + 3\text{e}^- + 4\text{H}_2\text{O}$

$2\text{Cr(OH)}_6^{3-} + 3\text{H}_2\text{O}_2 \rightarrow 2\text{CrO}_4^{2-} + 2\text{OH}^- + 8\text{H}_2\text{O}$



Re: $\text{H}_2\text{O}_2 + 2\text{e}^- \rightarrow 2\text{OH}^-$

Ox: $\text{Co(H}_2\text{O)}_4(\text{OH})_2 + \text{OH}^- \rightarrow \text{Co(H}_2\text{O)}_3(\text{OH})_3 + \text{e}^- + \text{H}_2\text{O}$

$2\text{Co(H}_2\text{O)}_4(\text{OH})_2 + \text{H}_2\text{O}_2 \rightarrow 2\text{Co(H}_2\text{O)}_3(\text{OH})_3 + 2\text{H}_2\text{O}$

Half equations in alkaline conditions:

These are more difficult to do than half equations under acidic conditions. The easiest way of doing it is to balance as if under acidic conditions then add OH^- ions to both sides to convert to alkaline

For change $\text{Cr(OH)}_6^{3-} \rightarrow \text{CrO}_4^{2-}$

Add H_2O to balance O: $\text{Cr(OH)}_6^{3-} \rightarrow \text{CrO}_4^{2-} + 2\text{H}_2\text{O} + 3\text{e}^-$

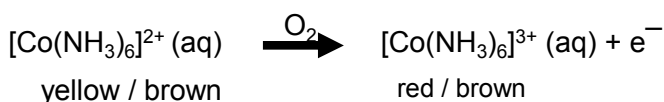
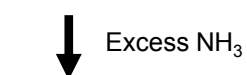
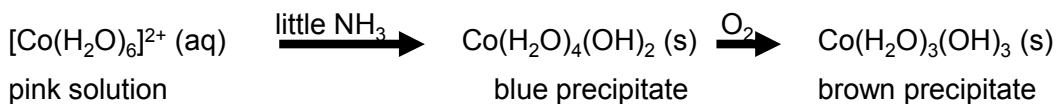
Add H^+ to balance H: $\text{Cr(OH)}_6^{3-} \rightarrow \text{CrO}_4^{2-} + 2\text{H}_2\text{O} + 2\text{H}^+ + 3\text{e}^-$

Add OH^- to both sides $\text{Cr(OH)}_6^{3-} + 2\text{OH}^- \rightarrow \text{CrO}_4^{2-} + 2\text{H}_2\text{O} + 2\text{H}^+ + 2\text{OH}^- + 3\text{e}^-$
to cancel out H^+ :

$\text{Cr(OH)}_6^{3-} + 2\text{OH}^- \rightarrow \text{CrO}_4^{2-} + 3\text{e}^- + 4\text{H}_2\text{O}$

Ammonical oxidation of Cobalt

Ammonia ligands make the Co(II) state unstable. Air oxidises Co(II) to Co(III) .



H_2O_2 could also bring about the oxidation

Catalysis

Catalysts increase reaction rates without getting used up. They do this by **providing an alternative route** with a **lower activation energy**

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

A **heterogeneous catalyst** is in a different phase from the reactants

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

Heterogeneous catalysis

Heterogeneous catalysts are usually solids whereas the reactants are gaseous or in solution. The reaction occurs at the surface of the catalyst.

Strength of adsorption

The strength of adsorption helps to determine the effectiveness of the catalytic activity
Some metals e.g. **W** have **too strong** adsorption and so the products cannot be released
Some metals e.g. **Ag** have **too weak** adsorption, and the reactants do not adsorb in high enough concentration
Ni and Pt have about the right strength and are most useful as catalysts

Surface area:

Increasing the surface area of a solid catalyst will improve its effectiveness. A support medium is often used to maximise the surface area and minimise the cost (e.g. Rh on a ceramic support in catalytic converters).

Examples of heterogeneous catalysts

V_2O_5 is used as a catalyst in the Contact Process.

Overall equation : $2SO_2 + O_2 \rightarrow 2SO_3$

step 1 $SO_2 + V_2O_5 \rightarrow SO_3 + V_2O_4$

step 2 $2V_2O_4 + O_2 \rightarrow 2V_2O_5$

Cr_2O_3 catalyst is used in the manufacture of methanol from carbon monoxide and hydrogen
 $CO + 2H_2 \rightarrow CH_3OH$

Poisoning Catalysts

catalysts can become poisoned by impurities and consequently have reduced efficiency

It is important to ensure the purity of the reactants if poisoning can occur

Homogeneous catalysis

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

Adsorption of reactants at active sites on the surface may lead to catalytic action. In this case the reactants tend to adsorb (chemically bond) on to the surface of the catalyst. This can result in the bonds within the reactant molecules becoming weaker, or the molecules being held in a more reactive configuration. There will also be a higher concentration of reactants at the solid surface so leading to a higher collision frequency

Transition Metals can use the 3d and 4s e- of atoms on the metal surface to form weak bonds to the reactants.

Steps in Heterogeneous Catalysis

1. Reactants form bonds with atoms at **active sites** on the surface of the catalyst (adsorbed onto the surface)
2. As a result bonds in the reactants are weakened and break
3. New bonds form between the reactants held close together on catalyst surface
4. This in turn weakens bonds between product and catalyst and product leaves (desorbs)

Learn the equations for this mechanism. Note the oxidation number of the vanadium changes and then changes back. It is still classed as a catalyst as it returns to its original form

Fe is used as a catalyst in the Haber Process
 $N_2 + 3H_2 \rightarrow 2NH_3$

Poisoning has a cost implication e.g. poisoning by sulphur in the Haber Process and by lead in catalytic converters in cars means that catalysts lose their efficiency and may need to be replaced

Leaded petrol cannot be used in cars fitted with a catalytic converter since lead strongly absorbs onto the surface of the catalyst

The intermediate will often have a different oxidation state to the original transition metal. At the end of the reaction the original oxidation state will reoccur. This illustrates importance of variable oxidation states of transition metals in catalysis

Examples of homogenous catalysts

Learn these 2 examples and equations carefully

Reaction between iodide and persulphate ions

The reaction between I^- and $\text{S}_2\text{O}_8^{2-}$ catalysed by Fe^{2+}
overall $\text{S}_2\text{O}_8^{2-} + 2\text{I}^- \rightarrow 2\text{SO}_4^{2-} + \text{I}_2$

Catalysed alternative route

stage 1 $\text{S}_2\text{O}_8^{2-} + 2\text{Fe}^{2+} \rightarrow 2\text{SO}_4^{2-} + 2\text{Fe}^{3+}$

stage 2 $2\text{I}^- + 2\text{Fe}^{3+} \rightarrow 2\text{Fe}^{2+} + \text{I}_2$

The uncatalysed reaction is very slow because the reaction needs a collision between **two negative ions**. **Repulsion** between the ions is going to hinder this – meaning **high activation energy**

Both of the individual stages in the catalysed mechanism involve collision between positive and negative ions and will have lower activation energies.

Autocatalytic reaction between Ethanedioate and Manganate ions

The autocatalysis by Mn^{2+} in titrations of $\text{C}_2\text{O}_4^{2-}$ with MnO_4^-

overall $2\text{MnO}_4^- + 5\text{C}_2\text{O}_4^{2-} + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 10\text{CO}_2 + 8\text{H}_2\text{O}$

Catalysed alternative route

Step 1 $4\text{Mn}^{2+} + \text{MnO}_4^- + 8\text{H}^+ \rightarrow 5\text{Mn}^{3+} + 4\text{H}_2\text{O}$

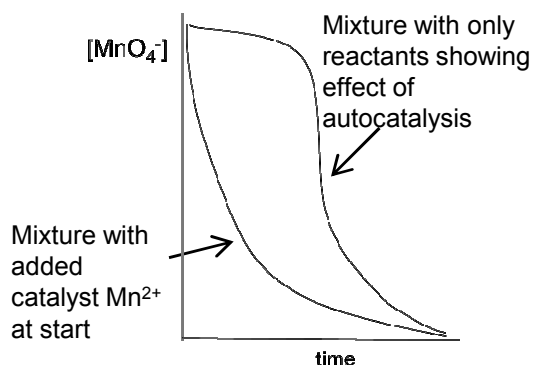
Step 2 $2\text{Mn}^{3+} + \text{C}_2\text{O}_4^{2-} \rightarrow 2\text{Mn}^{2+} + 2\text{CO}_2$

The initial uncatalysed reaction is **slow** because the reaction needs a collision between **two negative ions** which **repel each other** leading to a **high activation energy**

As the Mn^{2+} ions are produced act as an **autocatalyst** and the reaction starts to speed up as they bring about the alternative reaction route with lower activation energy.

The reaction eventually slows as the MnO_4^- concentration drops

This is an example of **autocatalysis** where one of the products of the reaction can catalyse the reaction



Following the reaction rate

This can be done by removing samples at set times and titrating to work out the concentration of MnO_4^- .

It could also be done by use of a spectrometer measuring the intensity of the purple colour. This method has the advantage that it **does not disrupt the reaction mixture**, using up the reactants and it leads to a much **quicker determination of concentration**

Other applications of transition metal complexes

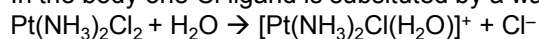
Fe(II) in haemoglobin enables oxygen to be transported in the blood,

CO is toxic to humans as CO can form a strong coordinate bond with haemoglobin. This is a stronger bond than that made with oxygen and so it prevents the oxygen attaching to the haemoglobin..

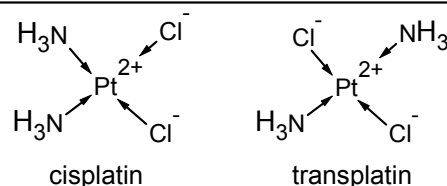
The Pt(II) complex cisplatin is used as an anticancer drug.

The cisplatin version only works as two chloride ions are displaced and the molecule joins on to the DNA. In doing this it stops the replication of cancerous cells.

In the body one Cl ligand is substituted by a water molecule



Be able to apply your knowledge of bonding to given information in the question to explain how it bonds to DNA molecule- generally a combination of dative covalent bonding and hydrogen bonding



It can also prevent the replication of healthy cells by bonding on to healthy DNA which may lead to unwanted side effects like hair loss

$[\text{Ag}(\text{NH}_3)_2]^+$ is used in Tollen's reagent to distinguish between aldehydes and ketones

Reactions of Inorganic Compounds in Aqueous Solution

Lewis acids and bases

definitions: **Lewis acid**: electron pair acceptor
Lewis base: electron pair donator

In the formation of complex ions the ligand is the Lewis base as it donating a pair of electrons in the dative covalent bond and the metal ion is the Lewis acid

Metal-aqua ions

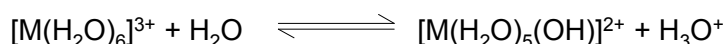
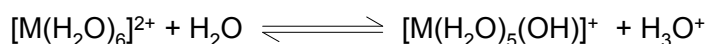
Metal aqua ions are formed in aqueous solution

$[M(H_2O)_6]^{2+}$, limited to M = Fe (green) , Co(pink) and Cu (blue);
 $[M(H_2O)_6]^{3+}$, limited to M = Al (colourless), Cr (ruby) and Fe (violet)

In solution Cr(III) often appears green and Fe(III) appears yellow/brown due to hydrolysis reactions. The ruby and violet colour is only really seen in solid hydrated salts that contain these complexes

Acidity or hydrolysis reactions

The following equilibria happen in aqueous solutions of metal ions



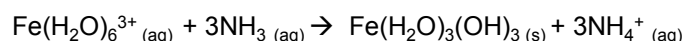
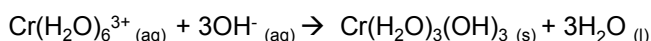
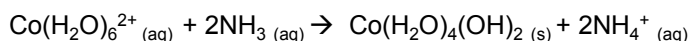
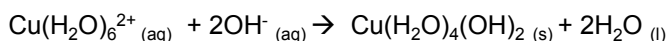
The equilibria lead to generation of acidic solutions with M^{3+} ions, and very weakly acidic solutions with M^{2+} ions.

The acidity of $[M(H_2O)_6]^{3+}$ is greater than that of $[M(H_2O)_6]^{2+}$ in terms of the greater polarising power (charge/size ratio) of the $3+$ metal ion. The greater the polarising power, the more strongly it attracts the water molecule. This weakens the O-H bond so it breaks more easily

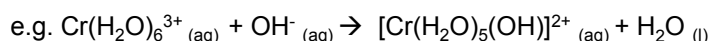
Reaction with limited OH^- and limited NH_3

The bases OH^- and ammonia when in limited amounts form the same hydroxide precipitates. They form in deprotonation acid base reactions

$M(OH)_2(H_2O)_4(s)$: Cu blue ppt, Co blue ppt, Fe (II) green ppt
 $M(OH)_3(H_2O)_3(s)$: Cr (III) green ppt, Fe(III) brown ppt, Al white ppt



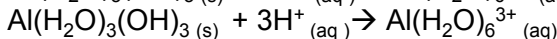
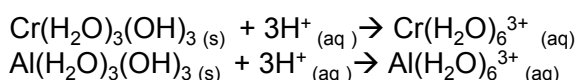
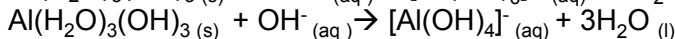
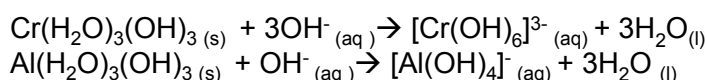
This process can happen step wise removing one proton at a time. Be able to write equations for this too



Reaction with excess OH^-

With excess NaOH, the Cr and Al hydroxides dissolve.
Cr becomes $[Cr(OH)_6]^{3-}(aq)$ green solution
Al becomes $[Al(OH)_4]^-(aq)$ colourless solution

These hydroxides are classed as **amphoteric** as they dissolve in both acids and bases



Reaction with excess NH_3

With **excess NH_3** substitution reactions occur with Cu, Co and Cr and their precipitates dissolve

The ligands NH_3 and H_2O are similar in size and are uncharged. Ligand exchange occurs without change of co-ordination number for Co and Cr

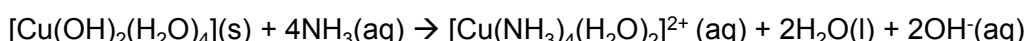
Cr becomes $[Cr(NH_3)_6]^{3+}$ purple solution

Co becomes $[Co(NH_3)_6]^{2+}$ pale yellow solution

Remember this Co complex is oxidised to $+3$ on standing in air

Cu becomes $[Cu(NH_3)_4(H_2O)_2]^{2+}$ deep blue solution

This substitution may, however, be incomplete as in the case with Cu



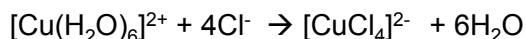
Reactions with chloride ions

Addition of a high concentration of chloride ions (from conc HCl) to an aqueous ion lead to a substitution reaction.

The Cl^- ligand is larger than the uncharged H_2O and NH_3 ligands and that ligand exchange can involve a change of co-ordination number

Addition of HCl to aqueous ions of Cu and Co lead to a change in coordination

number from 6 to 4	These are
$[\text{CuCl}_4]^{2-}$ yellow solution	tetrahedral in
$[\text{CoCl}_4]^{2-}$ blue solution	shape



Reactions with carbonate solution

The 2+ ions react differently to the 3+ ions with carbonate solutions

The 2+ ions with **carbonate** solution results in MCO_3 ppt being formed (Cu blue, Co pink, Fe(II) green)

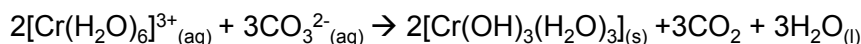


These are precipitation reactions

The 3+ ions with **carbonate** solution form a $M(OH)_3$ ppt and CO_2 gas is evolved

Al forms white ppt of $\text{Al}(\text{OH})_3 \cdot (\text{H}_2\text{O})_3 + \text{CO}_2$
 Cr (III) forms green ppt of $\text{Cr}(\text{OH})_3 \cdot (\text{H}_2\text{O})_3 + \text{CO}_2$
 Fe(III) forms brown ppt of $\text{Fe}(\text{OH})_3 \cdot (\text{H}_2\text{O})_3 + \text{CO}_2$

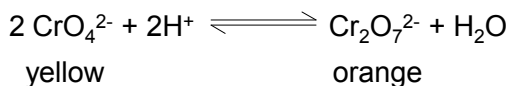
MCO_3 is formed with 2+ ions but $\text{M}_2(\text{CO}_3)_3$ is not formed with 3+ ions. The difference is explained by the greater polarising power of the 3+ ion due to its higher charge density.



These are classed as acidity reactions

Chromate/ dichromate equilibrium

The chromate and dichromate ions can be converted from one to the other by the following equilibrium reaction

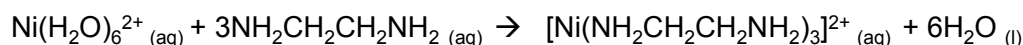


This is not a redox reaction as both the chromate and dichromate ions have an oxidation number of +6. This is an acid base reaction

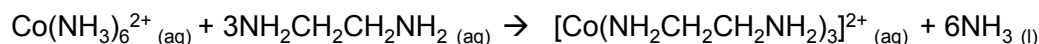
Stability of complexes

The substitution of unidentate ligand with a bidentate or a multidentate ligand leads to a more stable complex.

This chelate effect can be explained in terms of a positive entropy change in these reactions as more molecules of products than reactants

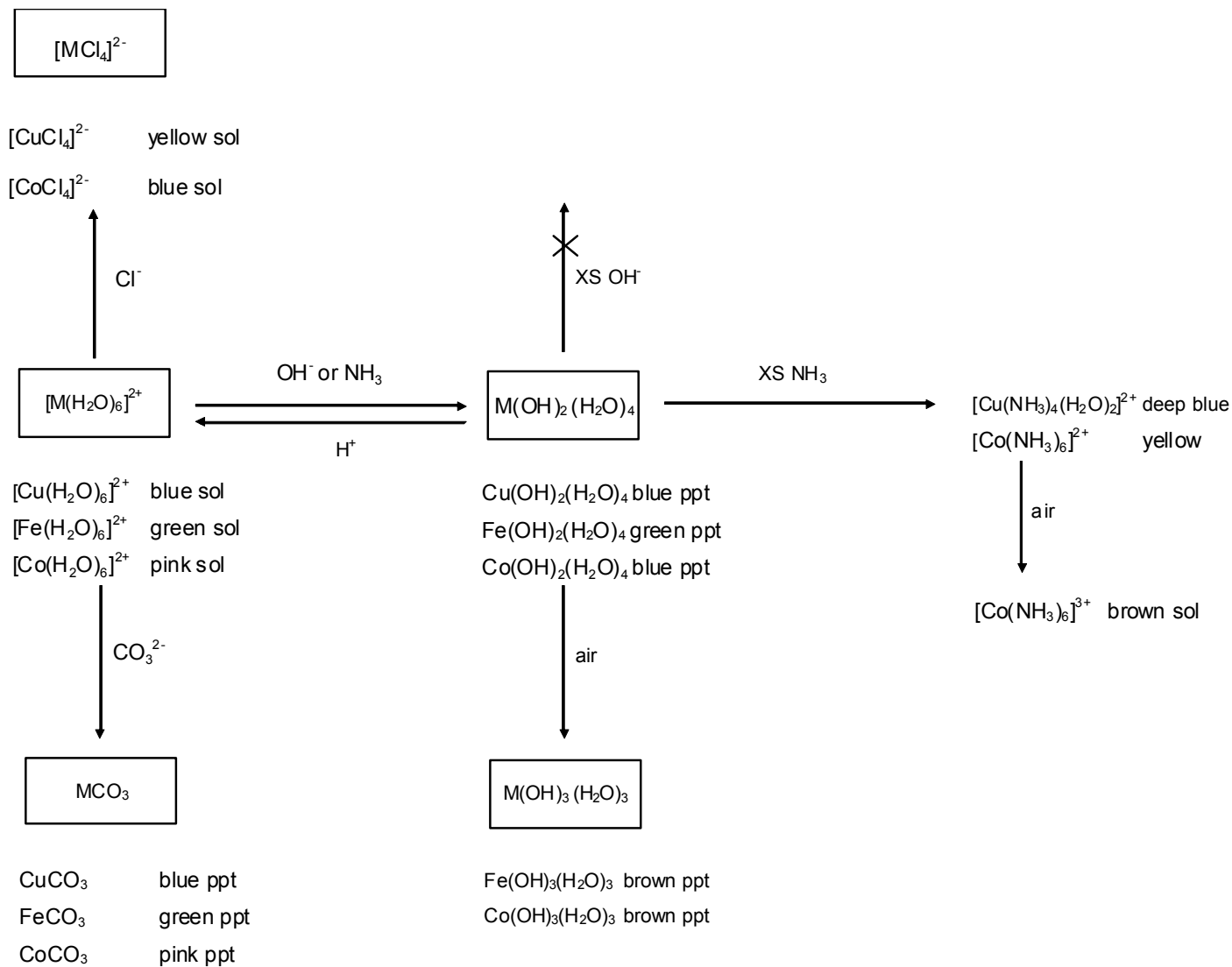


The Ni complex ion has changed from having unidentate ligands to bidentate ligands. In this reaction there is an increase in the entropy because there are more moles of products than reactants (from 4 to 7), creating more disorder.

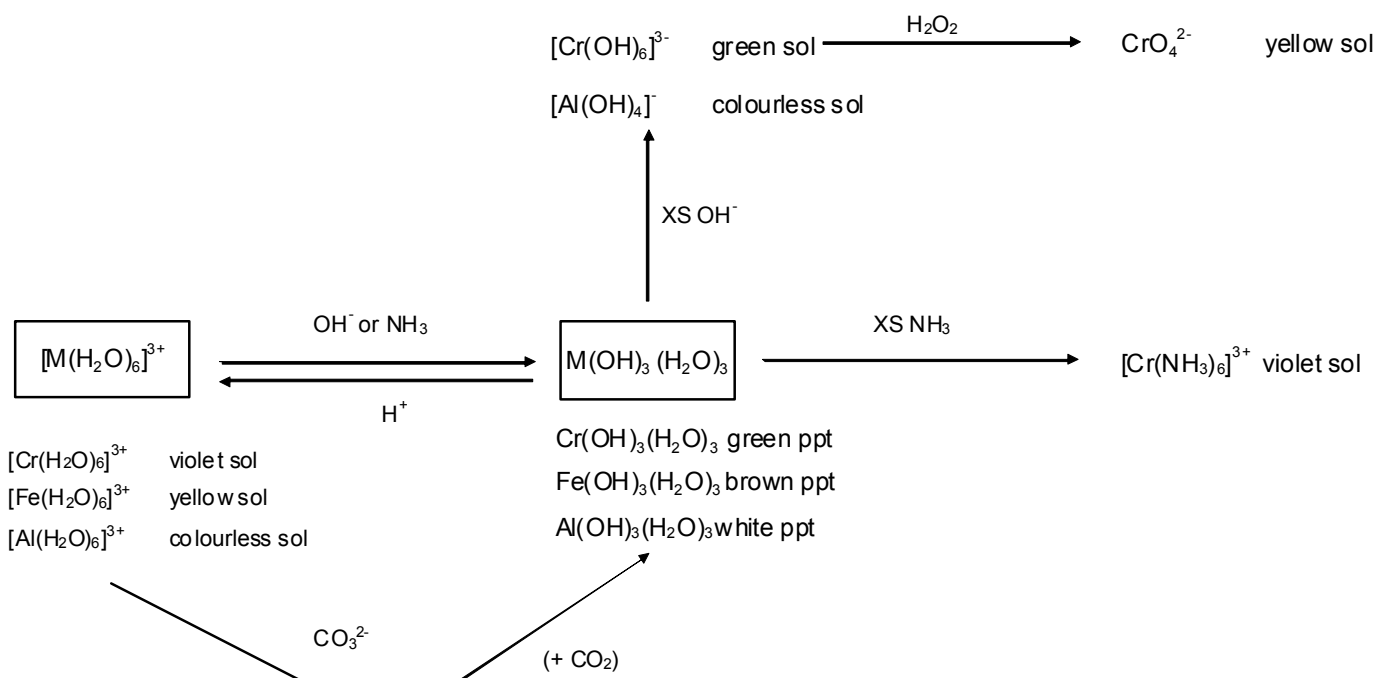


This reaction has an increase in stability and entropy as explained above. Its enthalpy change is close to zero as **the number** of dative covalent and **type** (N to metal coordinate bond) **are the same** so the energy required to break and make bonds will be the same

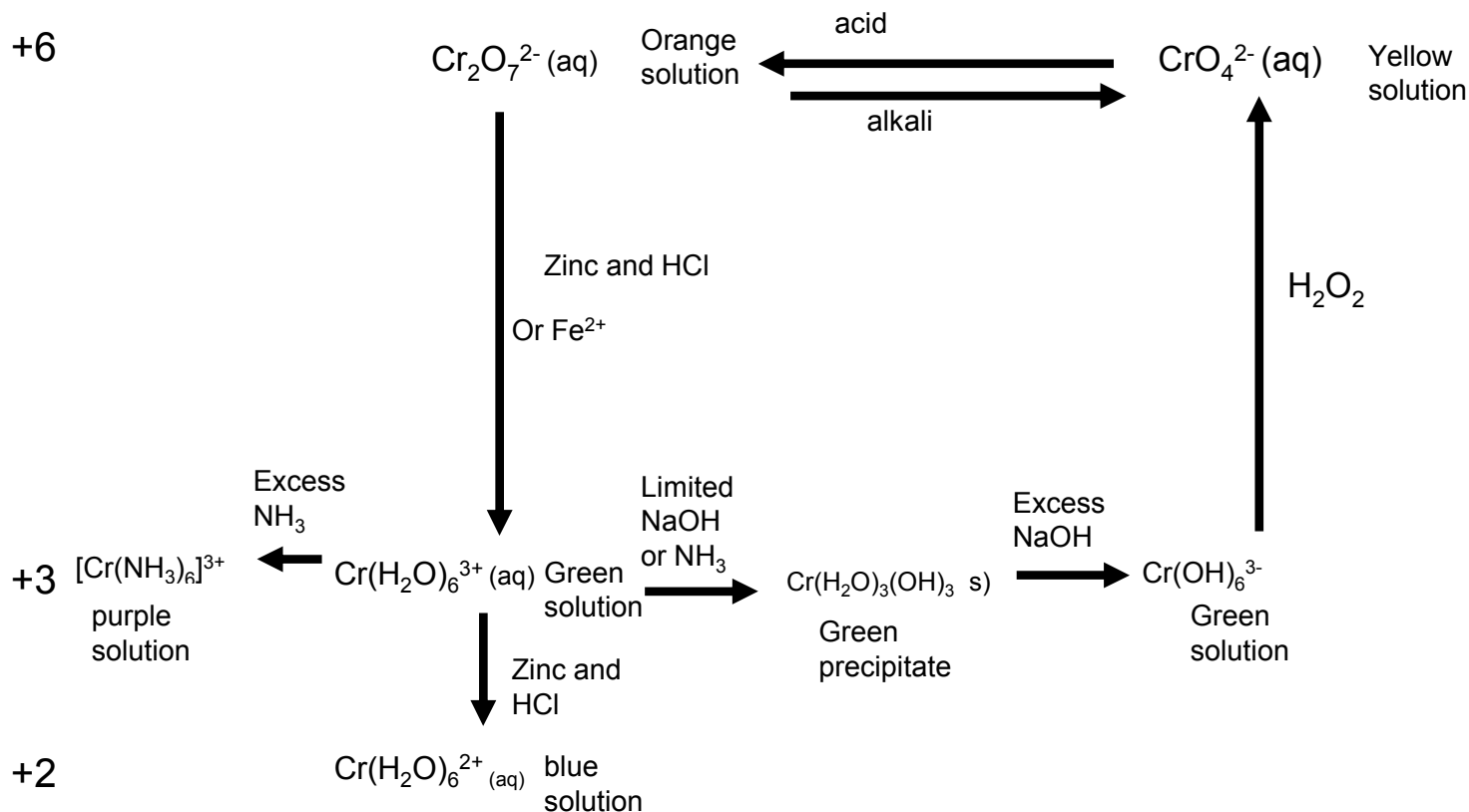
2+ ion summary



3+ ion summary



Chromium summary

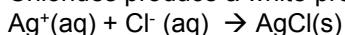


Silver Chemistry

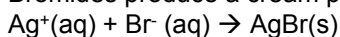
Reactions of halides with Silver nitrate

Fluorides produce no precipitate

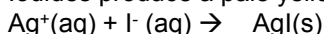
Chlorides produce a white precipitate



Bromides produce a cream precipitate



Iodides produce a pale yellow precipitate

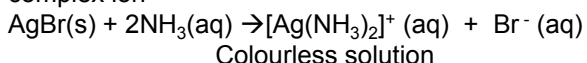


The silver halide precipitates can be treated with ammonia solution to help differentiate between them if the colours look similar:

Silver chloride dissolves in **dilute ammonia** to form a complex ion



Silver bromide dissolves in **concentrated ammonia** to form a complex ion



Silver iodide does not react with ammonia – it is too insoluble.

A2 questions often link module 2 halide silver nitrate chemistry to transition metals

