

Transition Metal Chemistry

Transition metals as d-block elements

For the elements up to Ca the 3d orbitals are higher in energy than the 4s orbital. Therefore, after argon (element 18), the 4s orbital is filled: Ca has electron configuration [Ar] 4s².

From scandium on, the 3d orbitals are filled, until they have ten electrons at zinc.

The term “d-block elements” refers to those elements in which this d-subshell is filling (Sc–Zn), but the term “**transition elements**” is used for d-block elements which form one or more stable ions with a **partially-filled d-subshell**.

A transition element is a d-block element that forms one or more stable ions that have incomplete d-orbitals.

This excludes Sc and Zn, since their only common oxidation states are Sc³⁺ (3d⁰) and Zn²⁺ (3d¹⁰).

This distinction is made because the main features of the chemistry of the transition elements depend largely on this partially filled d-subshell.

Electron configurations of the d-block elements and their simple ions.

You will be expected to use your Periodic Table to deduce the electronic configurations of atoms and ions.

Remember that:

- (i) the stability of the half-filled sub-shell means that d⁵ and d¹⁰ configurations are particularly stable e.g. Cr is 1s² 2s² 2p⁶ 3s² 3p⁶ **4s¹ 3d⁵** (not 4s² 3d⁴);
Cu is 1s² 2s² 2p⁶ 3s² 3p⁶ **4s¹ 3d¹⁰** (not 4s² 3d⁹);
- (ii) in ions the 4s electrons are always lost before 3d electrons: Fe²⁺ is 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁶.

<i>Element</i>	<i>Symbol</i>	<i>Electronic structure of atom</i>	<i>Common ion(s)</i>	<i>Electronic structure of ion</i>
Scandium	Sc	(Ar)3d ¹ 4s ²	Sc ³⁺	(Ar)
Titanium	Ti	(Ar)3d ² 4s ²	Ti ⁴⁺ Ti ³⁺	(Ar) (Ar)3d ¹
Vanadium	V	(Ar)3d ³ 4s ²	V ³⁺	(Ar)3d ²
Chromium	Cr	(Ar)3d ⁵ 4s ¹	Cr ³⁺	(Ar)3d ³
Manganese	Mn	(Ar)3d ⁵ 4s ²	Mn ²⁺	(Ar)3d ⁵
Iron	Fe	(Ar)3d ⁶ 4s ²	Fe ²⁺ Fe ³⁺	(Ar)3d ⁶ (Ar)3d ⁵
Cobalt	Co	(Ar)3d ⁷ 4s ²	Co ²⁺	(Ar)3d ⁷
Nickel	Ni	(Ar)3d ⁸ 4s ²	Ni ²⁺	(Ar)3d ⁸
Copper	Cu	(Ar)3d ¹⁰ 4s ¹	Cu ⁺ Cu ²⁺	(Ar)3d ¹⁰ (Ar)3d ⁹
Zinc	Zn	(Ar)3d ¹⁰ 4s ²	Zn ²⁺	(Ar)3d ¹⁰

Note Fe³⁺ is more stable than Fe²⁺ as it has an electron configuration without electron repulsion in the partially filled d sub-shell.

General Properties of Transition Metals.

Transition metals have higher melting points, higher boiling points and higher densities than other metals. Transition metals also show the following characteristic properties:

1. Variable oxidation states:- Transition metals have electrons of similar energy in both the 3d and 4s levels. This means that one particular element can form ions of roughly the same stability by losing different numbers of electrons. Thus, all transition metals from titanium to copper can exhibit two or more oxidation states in their compounds.

Oxidation states of some Transition Metals:

Titanium- +2, +3, +4

Chromium- +2, +3, +6

Iron- +2, +3

Nickel- +2, +3, +4

Vanadium- +2, +3, +4, +5

Manganese- +2, +3, +4, +5, +6, +7

Cobalt- +2, +3

Copper- +1, +2

2. Formation of complex ions:- As a lot of the transition metals have some empty spaces in their 3d-orbitals, they can receive lone pairs of electrons and form dative covalent bonds thus producing complex compounds.

3. Coloured compounds:- When electrons move from a d-orbital (with lower energy) to another d-orbital (with higher energy), energy is taken in. This energy is in the form of visible light. The transition metal appears the complementary colour to the light absorbed, thus producing coloured compounds.

4. Catalytic properties:- For any element its higher oxidation states give rise to covalent compound formation. As Transition Metals have variable oxidation states, they tend to have catalytic properties.

Complex ions

Water molecules, hydroxide ions, ammonia molecules and cyanide ions can all link on to transition metal ions to form **complex ions**. They do so by donating a lone pair to form a bond – this is a dative covalent bond. The ions or molecules that form these bonds are called ligands.

A **complex ion** is one in which a central positive ion is surrounded by ligands, which are co-ordinately (datively) bonded to it; e.g. $\text{Cr}(\text{H}_2\text{O})_6^{3+}$, $\text{Fe}(\text{CN})_6^{4-}$.

A **ligand** is a molecule or negative ion which has a lone pair of electrons, and can use its lone pair to form co-ordinate bonds to a metal ion; e.g. H_2O , CN^- .

The transition metals are not unique in forming complexes (there are small numbers formed by metals in groups 2, 3 and 4), but they form a much wider range than other elements.

This is because the transition metal ions are small and polarising, since their nuclei are poorly shielded, and so they attract ligands strongly.

Naming Complex Ions

The names of complex ions contain four main components.

First part	Second part	Third part	Fourth part
Number of each type ligand	Name of ligand	Name of transition metal (ending in -ate if it is a negative ion)	Charge on transition metal

Ion formula	Name of ion
$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$	hexaaquacopper(II)
$[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$	hexaaquachromium(III)
$[\text{CuCl}_4]^{2-}$	Tetrachlorocuprate(II)
$[\text{Cu}(\text{NH}_3)_4]^{2+}$	tetraamminecopper(II)
$[\text{Fe}(\text{CN})_6]^{4-}$	hexacyanoferrate(II)

When writing formulae, the central atom is put first, then the negative ions and then follow any neutral molecules. Everything is then put in square brackets and the charge added. For example, tetraaquachloro copper (II) would be written as $[\text{CuCl}(\text{H}_2\text{O})_4]^+$

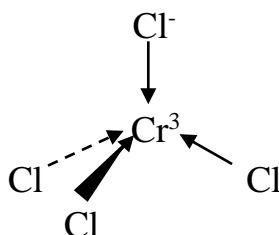
Shapes of Complex Ions

Complex ions can be tetrahedral in shape, but the majority have an octahedral shape.

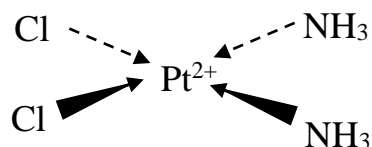
Linear Dichlorocuprate(I) $[\text{CuCl}_2]^-$



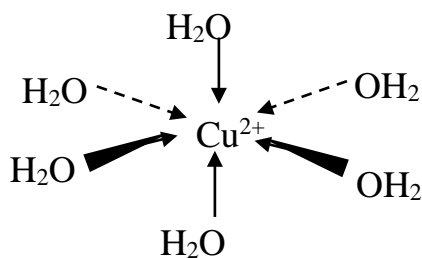
Tetrahedral Tetrachlorochromate(III), $[\text{CrCl}_4]^-$



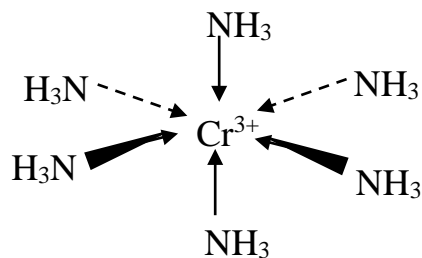
Planar Dichlorodiamminoplatinum(II), $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$



Octahedral Hexaaquacopper(II), $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$



Octahedral Hexaamminechromium(III), $[\text{Cr}(\text{NH}_3)_6]^{3+}$

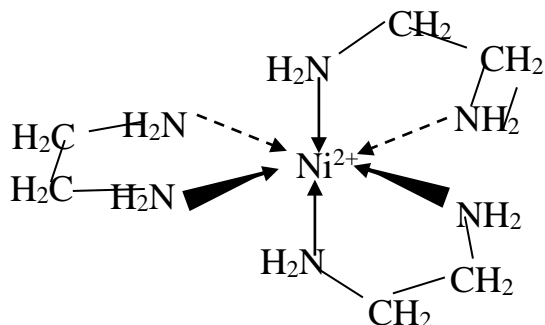


Bidentate and Polydentate ligands

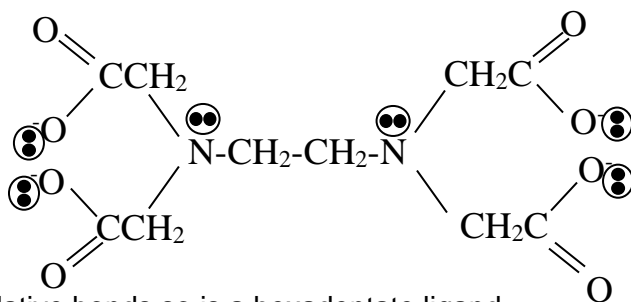
All the above complexes contain monodentate ligands (ligands that have one set of teeth with which they bite onto the transition metal ion) as they form one dative covalent bond.

Other ligands can form more than one dative bond and are called polydentate.

An example of a polydentate ligand is 1,2-diaminoethane; this can form two dative bonds so can also be referred to as bidentate.



Another polydentate ligand is called by letters which come from the old name of the ethylenediaminetetraacetate, or EDTA.



This can form six dative bonds so is a hexadentate ligand.

Colour in Complex Ions

When ligands pack around a metal ion, the d-orbitals no longer have exactly the same energies. If they are partially filled, it is possible for an electron to jump from a lower-energy d-orbital to an unoccupied higher-energy d-orbital.

These “**d-d transitions**” are of an energy corresponding to absorption in the visible region, and so the compound appears to be coloured. The colour is that of the light which is *not* absorbed:

e.g. copper(II) ions look blue because they absorb red light.

The energies of d-orbitals, and so the colour of the complex, are very sensitive to the ligand present.

Visible light stretches from purple/blue at 400–500nm, via yellow at around 600nm to red at 650nm. A complex absorbing from 400–550nm will look red (like $\text{Fe}(\text{H}_2\text{O})_5\text{SCN}^{2+}$ below) while one absorbing from 600-650nm will look blue (e.g. $\text{Cu}(\text{H}_2\text{O})_6^{2+}$). One absorbing in the middle, from 500-600nm will be a blue/red mix i.e. purple.

Compounds of d-block elements which are not transition elements cannot undergo these electron transitions, so do not have coloured compounds.

The compounds of zinc and scandium are therefore white and colourless in solution.

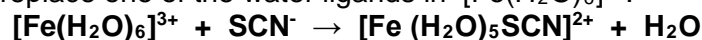
Copper(I) which has a full d-shell also has white compounds.

Ligand Exchange Processes

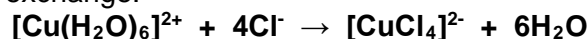
It is possible for one type of ligand to be replaced by another type.

Ligand exchange reactions often have a colour change associated with them.

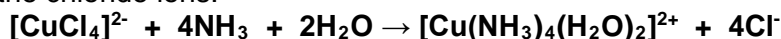
Thiocyanate ions can replace one of the water ligands in $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$.



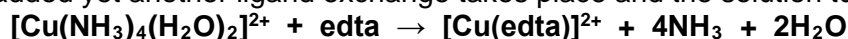
An aqueous Copper(II) sulphate solution is blue in colour because of the presence of $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ ions. When concentrated hydrochloric acid is added to this solution the colour changes from blue to green. This happens because the $[\text{CuCl}_4]^{2-}$ ion is produced. The Cl^- ions have replaced the H_2O molecules in a ligand exchange.



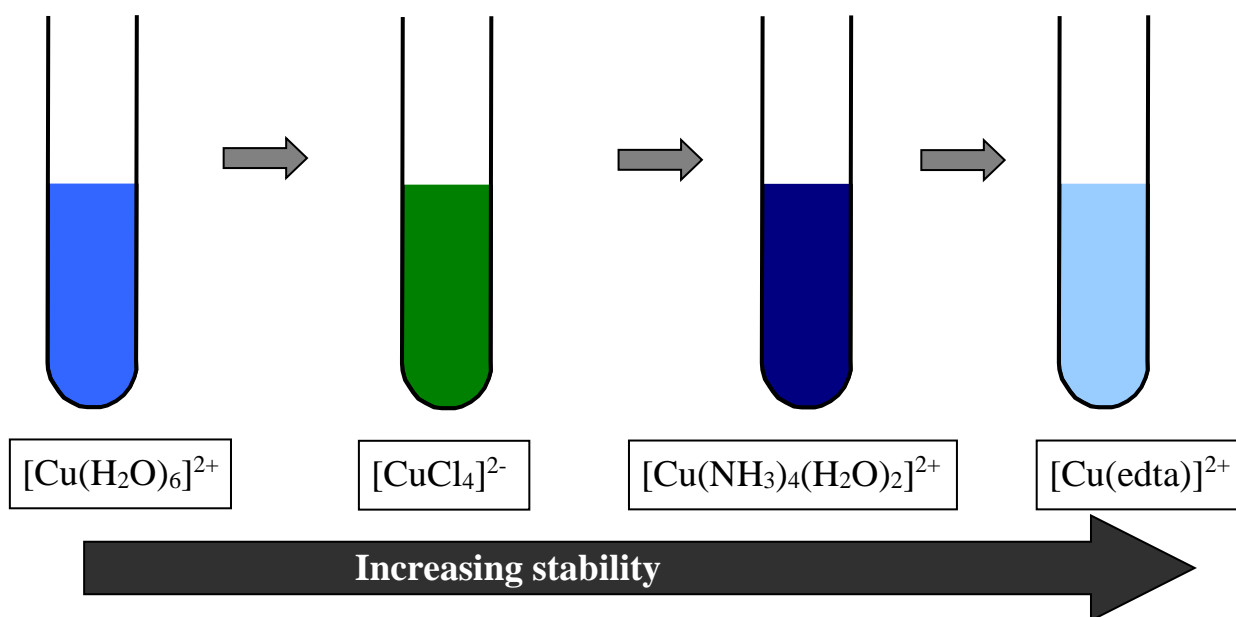
When ammonia is added a further change from green to deep blue takes place as ammonia molecules replace the chloride ions.



If EDTA is then added yet another ligand exchange takes place and the solution turns pale blue.



These changes take place because the complexes become more stable.



The stability of the $[\text{Cu}(\text{edta})]^{2+}$ complex can be understood in terms of entropy.

As the one edta molecule replaces the six smaller ligands, the small ligands are released into solution and therefore have greater freedom of movement, so greater disorder and consequently the entropy increases.

Oxidation states of transition elements – Cu and Cr

Copper [Ar]3d¹⁰4s¹

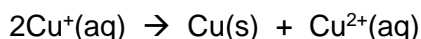
Copper, 3d¹⁰, is the only member of the transition series to have a significant +1 oxidation state, and even here the +1 state is only stable if in a complex ion, or in an insoluble compound – in solution, it disproportionates.

The +1 state, with a full d sub-shell, is not coloured (apart from Cu₂O).

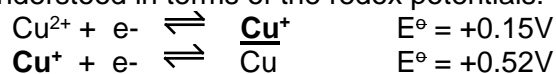
The +2 state, with its familiar blue and green complexes, is the normal stable state.

Cu(I)

- Cu₂O, as made by reduction of Fehling's or Benedict's solution with a reducing sugar, is a red insoluble solid.
- CuCl and Cu₂SO₄ are white solids. Both of these, when dissolved in water disproportionate:



- This can be understood in terms of the redox potentials:



There is a reaction between the two underlined species, i.e. the disproportionation.

So when a copper(I) compound is dissolved in water a blue solution (Cu²⁺_(aq)) and a red-brown solid (Cu_(s)) are formed.

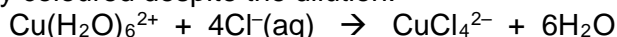
Cu(II)

- Most copper(II) compounds are blue, and in solution they give blue Cu(H₂O)₆²⁺ ions.
- When **copper(II) sulphate** solution is treated with dilute aqueous **ammonia**, the solution starts blue because of the Cu(H₂O)₆²⁺ ion. It first forms a pale blue precipitate of Cu(OH)₂, and then this dissolves to give a deep blue coloured solution, containing the Cu(NH₃)₄(H₂O)₂²⁺ ion.

[N.B. the hydroxide is formed first because ammonia solution is alkaline, due to the reaction NH₃ + H₂O ⇌ NH₄⁺ + OH⁻. Then the high concentration of NH₃ molecules displaces the equilibrium in favour of forming the ammonia complex.]

Overall: Cu(H₂O)₆²⁺ + 4NH₃ ⇌ Cu(NH₃)₄(H₂O)₂²⁺ + 2H₂O

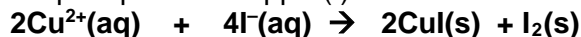
- When **copper(II) sulphate** solution is treated with **concentrated hydrochloric acid** (or sodium chloride solution), the solution starts blue because of the Cu(H₂O)₆²⁺ ion. As Cl⁻ ions are added they displace water molecules, forming the tetrahedral CuCl₄²⁻, which is yellow. The colour changes from blue through lime-green to yellow-green, and becomes more intensely coloured despite the dilution:



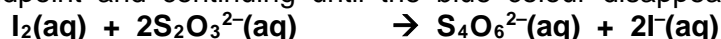
With **sodium hydroxide** Cu²⁺(aq) turns from a blue solution to give a mid/light-blue precipitate: Cu²⁺(aq) + 2OH⁻(aq) → Cu(OH)₂(s)

Estimation of copper(II)

Cu²⁺(aq) ions will react quantitatively with iodide ions, oxidising the latter to iodine and being reduced themselves to a white precipitate of copper(I) iodide:



When excess potassium iodide solution is added, the blue colour disappears and a brown solution with a white precipitate results. This can be titrated with standard sodium thiosulphate, adding starch before the endpoint and continuing until the blue colour disappears (leaving the white precipitate).



Chromium [Ar]3d⁵4s¹

Chromium has common oxidation states of 3+ and 6+, although 2+ also exists.

Chromium (II)	Chromium (III)	Chromium (VI)	
Cr ²⁺	Cr ³⁺	Chromate CrO ₄ ²⁻	Dichromate Cr ₂ O ₇ ²⁻
Blue	Green	Yellow	Orange

Cr(II)

- The Cr(H₂O)₆²⁺ ion is readily oxidised to Cr(H₂O)₆³⁺

Cr(III)

- The Cr(H₂O)₆³⁺ ion is **purple**, as are crystals of chromium(III) sulphate, Cr₂(SO₄)₃.
- Cr₂O₃ is a **green** solid, and Cr(OH)₃ is obtained as a green precipitate by adding sodium hydroxide to any solution of a chromium(III) salt. It is amphoteric and dissolves in excess sodium hydroxide to form a green solution of Cr(OH)₆³⁻.
- Hydrated chromium(III) chloride is a **green** solid, which gives a green solution with one or more Cl⁻ ions in the aqua-complex e.g. Cr(H₂O)₄Cl₂⁺.
- Cr³⁺(aq) can be oxidised to chromium(VI) by adding excess sodium hydroxide, then hydrogen peroxide, and boiling. The solution goes yellow as CrO₄²⁻ ions are formed.

Cr(VI)

- Potassium dichromate(VI), K₂Cr₂O₇, is an orange solid which dissolves in water to give an orange solution. In alkali this changes to the yellow chromate(VI) ion, and back again to orange dichromate(VI) on acidification:

$$\text{Cr}_2\text{O}_7^{2-} + 2\text{OH}^- \rightleftharpoons 2\text{CrO}_4^{2-} + \text{H}_2\text{O} \quad \text{then} \quad 2\text{CrO}_4^{2-} + 2\text{H}^+ \rightarrow \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$$

orange yellow yellow orange

 Note that this is *not* a redox reaction.
- Potassium dichromate is a good primary volumetric standard (i.e. can be obtained pure and stable, so can be weighed out to give a solution of reliably known concentration), and is often used to titrate with iron(II) ions, using a redox indicator.

$$\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{Fe}^{2+} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 6\text{Fe}^{3+}$$
- The E^o value of +1.33V for Cr₂O₇²⁻ in acid (to 2Cr³⁺) shows that it is quite a strong oxidising agent, and can be reduced by many moderate reducing agents (SO₂, Sn²⁺, ethanol on warming), when it turns from orange to green.

Uses of Cr

Chromium metal is used in making **stainless steel** is much more expensive than mild steel, resists corrosion effectively, but lacks some other useful properties (e.g strength, hardness) and so cannot always be substituted for normal steel.

Chromium is added to iron in smaller amounts to make **alloy steels** which are very hard (used for example in ball bearings).

Deprotonation reactions

Deprotonation reactions involve water ligands losing hydrogen ions (proton) to a proton acceptor such as an hydroxide ion.



Deprotonation reactions often result in the formation of a precipitate.

Reaction of complex ions with sodium hydroxide and ammonia solutions

Sodium hydroxide and ammonia solutions contain hydroxide ions. When these are added to solutions containing transition metal ions a precipitate of the metal hydroxide is formed.

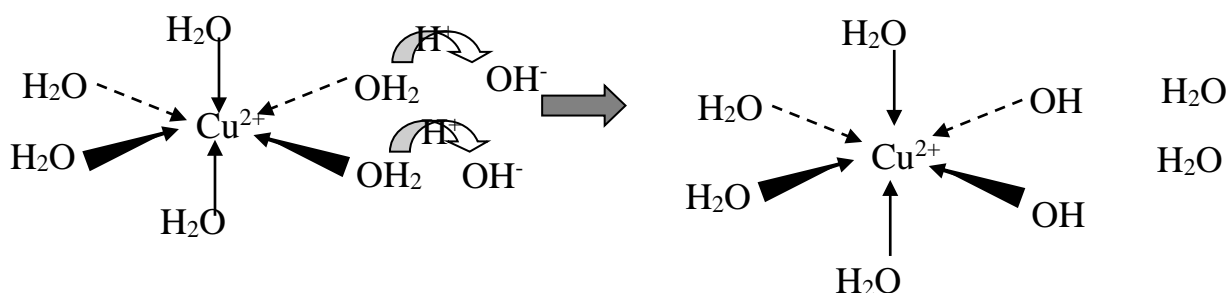
If further quantities of these reagents are added to the mixture, the precipitate, in certain cases, dissolves.

Ion in solution	Reaction with a few drops of $\text{NaOH}_{(\text{aq})}$ or $\text{NH}_3_{(\text{aq})}$	Reaction with excess $\text{NaOH}_{(\text{aq})}$	Reaction with excess $\text{NH}_3_{(\text{aq})}$
Cr^{3+}	Pale green ppt	Ppt dissolves to form a deep green solution	
Mn^{2+}	Beige ppt	No further reaction	
Fe^{2+}	Dirty green ppt	No further reaction	No further reaction
Fe^{3+}	Red-brown ppt	No further reaction	No further reaction
Ni^{2+}	Green gelatinous ppt	No further reaction	Ppt dissolves to form a blue solution
Cu^{2+}	Blue ppt	No further reaction	Ppt dissolves to form a deep blue solution
Zn^{2+}	White gelatinous ppt	Ppt dissolves to form a colourless solution	

A simple way of looking at these is as hydroxide ions adding to the transition metal ion – the number of hydroxide ions being equal to the charge on the ion.



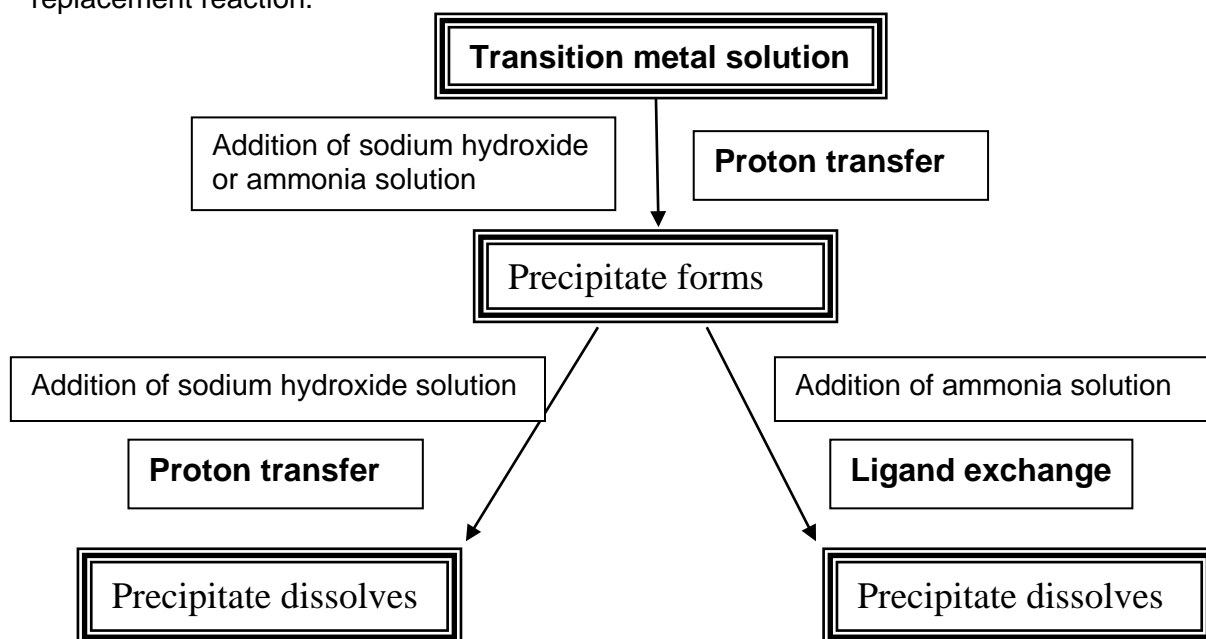
In fact rather than a water molecule leaving and a hydroxide ion joining, the process actually consists of a hydrogen ion moving.



This process is called a deprotonation reaction. All the reactions in which the transition ions form **precipitates** are deprotonation reactions.

When the precipitates dissolve in excess sodium hydroxide solution, it is because a further deprotonation reaction takes place. These reactions represent a metal hydroxide reacting with an alkali, so it can be regarded as being due to the amphoteric nature of the metal hydroxide and reflects a degree of non-metal character.

When the precipitate dissolves in excess ammonia solution, it is because the ammonia molecules replace the hydroxide and water molecules around the transition metal ion, and so this is a ligand replacement reaction.



ion	aqueous sodium hydroxide	aqueous ammonia solution
Cr³⁺	$\text{Cr}^{3+}_{(aq)} + 3\text{OH}^{-}_{(aq)} \rightarrow \text{Cr(OH)}_{3(s)}$ grey green $\text{Cr(OH)}_3(\text{H}_2\text{O})_{3(s)} + 3\text{OH}^{-}_{(aq)} \rightarrow [\text{Cr(OH)}_6]^{3-}_{(aq)} + 3\text{H}_2\text{O}$ deprotonation hexahydroxochromate(III)	$\text{Cr}^{3+}_{(aq)} + 3\text{OH}^{-}_{(aq)} \rightarrow \text{Cr(OH)}_{3(s)}$ grey green
Mn²⁺	$\text{Mn}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Mn(OH)}_{2(s)}$ white/brown	$\text{Mn}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Mn(OH)}_{2(s)}$ white/brown
Fe²⁺	$\text{Fe}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Fe(OH)}_{2(s)}$ muddy/green	$\text{Fe}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Fe(OH)}_{2(s)}$ muddy/green
Fe³⁺	$\text{Fe}^{3+}_{(aq)} + 3\text{OH}^{-}_{(aq)} \rightarrow \text{Fe(OH)}_{3(s)}$ rust brown	$\text{Fe}^{3+}_{(aq)} + 3\text{OH}^{-}_{(aq)} \rightarrow \text{Fe(OH)}_{3(s)}$ rust brown
Ni²⁺	$\text{Ni}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Ni(OH)}_{2(s)}$ lime green	$\text{Ni}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Ni(OH)}_{2(s)}$ lime green ligand exchange - dissolves in excess to give blue $[\text{Ni}(\text{NH}_3)_6]^{2+}_{(aq)}$
Cu²⁺	$\text{Cu}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Cu(OH)}_{2(s)}$ pale blue	$\text{Cu}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Cu(OH)}_2(\text{H}_2\text{O})_{4(s)}$ pale blue ligand exchange - dissolves in excess to give deep blue $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}_{(aq)}$
Zn²⁺	$\text{Zn}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Zn(OH)}_{2(s)}$ white $\text{Zn(OH)}_{2(s)} + 2\text{OH}^{-}_{(aq)} \rightarrow [\text{Zn(OH)}_4]^{2-}_{(aq)}$ deprotonation	$\text{Zn}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Zn(OH)}_{2(s)}$ ligand exchange - dissolves in excess to give $[\text{Zn}(\text{NH}_3)_4]^{2+}_{(aq)}$

Catalytic Properties of Transition Elements

The ability of transition elements to change oxidation state allows them to be used as catalysts. Transition elements or their compounds are used in a number of important industrial processes.

Substance	Reaction catalysed
Iron	Haber process to convert nitrogen and hydrogen to ammonia
Nickel	Margarine production to hydrogenate unsaturated hydrocarbons
Vanadium(V) oxide	Contact process to convert oxygen and sulphur dioxide to sulphur trioxide

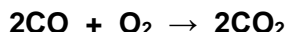
There are two main types of catalysis.

Heterogeneous catalysis is where the catalyst and the reactants are in different states.

Homogeneous catalysis is where the catalyst and the reactants are in the same state.

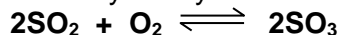
An example of heterogeneous catalysis is seen in the catalytic converters found in cars.

Platinum is used to remove harmful gases such as carbon monoxide.



Platinum provides a surface with which the carbon monoxide and oxygen can form weak bonds. The transition metal uses 3d and 4s electrons to form these bonds. The bonding of the molecules on the surface brings them close together and the formation of the weak bonds with the surface weakens the bonds within the molecules, reducing the energy required to break them.

The ability of the transition element to change its oxidation state is also important in catalysis. This is seen in the manufacture of sulphuric acid where the conversion of sulphur dioxide to sulphur trioxide is catalysed by vanadium(V) oxide.



The sulphur dioxide first reacts with the vanadium(V) oxide. $\text{SO}_{2(\text{g})} + \text{V}_2\text{O}_{5(\text{s})} \rightarrow \text{SO}_{3(\text{g})} + \text{V}_2\text{O}_{4(\text{s})}$

The vanadium(IV) formed in this reaction then reacts with the oxygen $\frac{1}{2}\text{O}_{2(\text{g})} + \text{V}_2\text{O}_{4(\text{s})} \rightarrow \text{V}_2\text{O}_{5(\text{s})}$

An example of homogeneous catalysis is seen in the catalysis by iron(II) of the reaction between the persulphate and iodide ions. $\text{S}_2\text{O}_8^{2-} + 2\text{I}^- \rightarrow 2\text{SO}_4^{2-} + \text{I}_2$

Although the persulphate is a powerful oxidizing agent, the reaction is slow because it requires negative ions to come together and this repulsion gives the reaction a high activation energy.

The iron(II) reacts first with the persulphate $2\text{Fe}^{2+} + \text{S}_2\text{O}_8^{2-} \rightarrow 2\text{SO}_4^{2-} + 2\text{Fe}^{3+}$

The iron(II) formed in this step then reacts with the iodide ions $2\text{Fe}^{3+} + 2\text{I}^- \rightarrow 2\text{Fe}^{2+} + \text{I}_2$

Development of New Catalysts

Development of new catalysts is an important area for research

Ethanoic acid is a very important industrial chemical used in the manufacture of polymers, perfumes, flavourings and pharmaceuticals. Until 1970, ethanoic acid was manufactured by oxidizing naphtha and butane at a temperature of 200°C, a pressure of 50atm and a catalyst of cobalt ethanoate. The large number of by-products gives the reaction a low atom economy.

In the 1960s a new process began to be used starting with methanol and carbon monoxide.



A catalyst of cobalt and iodine is used for this reaction. It had a theoretical atom economy of 100%, although in practice it did not reach this.

Various improvements were made to the catalyst used, using rhodium and then iridium in place of cobalt. As improvements were made the conditions required became milder, the reaction more efficient and the atom economy improved.

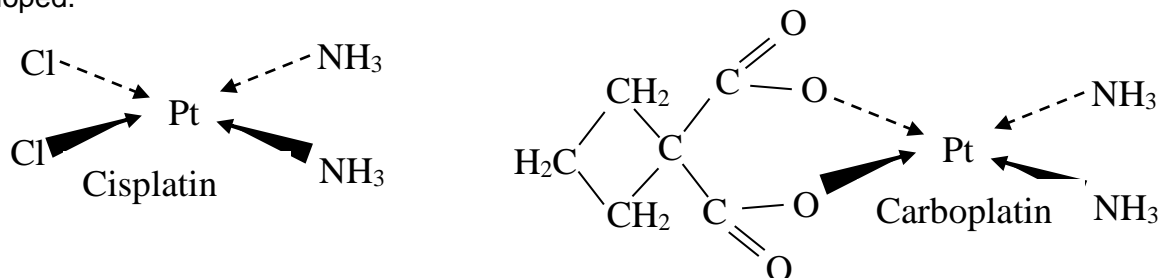
Other Uses of Transition Elements

Cancer treatment

Cancer involves cells dividing uncontrollably forming tumours.

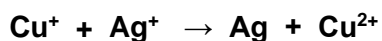
Cis $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ was found to be able to inhibit cell division and could therefore be used as a treatment for cancer.

Cisplatin as the material was called is now one of the most widely used anti-cancer drugs. One of the problems with it is its toxicity, so research continues and a new drug, carboplatin, has been developed.



Sunglasses

Photochromic sunglasses which become darker as the light intensity increases use a redox reaction. The lenses contain silver(I) chloride and copper(I) chloride. Strong light causes the following reactions to take place:



The silver produced in this reaction turns the glasses darker. When the light intensity decreases, the reverse reaction takes place and the glasses become less dark.