General properties of transition metals

Transition Metals 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.

Why are Zn and Sc not transition metals?
Zn can only form a +2 ion. In this ion the Zn2+ has a complete d orbital and so does not meet the criteria of having an incomplete d orbital in one of its compounds.
Sc only forms a +3 ion with the electronic structure
The Sc3+ 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. H2O, NH3 and Cl-) which can form one coordinate bond per ligand or bidentate (e.g. NH2CH2CH2NH2 and ethanedioate ion C2O4 2-) which have two atoms with lone pairs and can form two coordinate bonds per ligand or multidentate (e.g. EDTA4- which can form six coordinate bonds per ligand).

The EDTA4- anion has the formula

\[
\text{O} \quad \text{C} \quad \text{H}_2 \quad \text{N} \quad \text{CH}_2 \quad \text{H}_2 \quad \text{O} \\
\text{O} \quad \text{C} \quad \text{H}_2 \quad \text{N} \quad \text{CH}_2 \quad \text{H}_2 \quad \text{O} \\
\]

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

Equations to show formation of bidentate and multidentate complexes

\[
\begin{align*}
\text{Cu(H}_2\text{O)}_6^{2+} + 3\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2 & \rightarrow \text{[Cu(NH}_2\text{CH}_2\text{CH}_2\text{NH}_2)}_3^{2+} + 6\text{H}_2\text{O} \\
\text{Cu(H}_2\text{O)}_6^{2+} + 3\text{C}_2\text{O}_4^{2-} & \rightarrow \text{[Cu(C}_2\text{O}_4)}_3^{4+} + 6\text{H}_2\text{O} \\
\text{Cu(H}_2\text{O)}_6^{2+} + \text{EDTA}^{4-} & \rightarrow \text{[Cu(EDTA)}_2^{2-} + 6\text{H}_2\text{O}
\end{align*}
\]

When forming ions lose 4s before 3d

Sc 1s2s2p63s23p6 4s23d1
Ti 1s2s2p63s23p6 4s23d2
V 1s2s2p63s23p6 4s23d3
Cr 1s2s2p63s23p6 4s23d6
Mn 1s2s2p63s23p6 4s23d5
Fe 1s2s2p63s23p6 4s23d6
Co 1s2s2p63s23p6 4s23d7
Ni 1s2s2p63s23p6 4s23d6
Cu 1s2s2p63s23p6 4s23d10
Zn 1s2s2p63s23p6 4s23d10

When forming ions lose 3s before 3d

Mn 3s23p63d5
Fe 3s23p63d6
Co 3s23p63d7
Ni 3s23p63d8
Cu 3s23p63d9
Zn 3s23p63d10

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Formation of coloured ions

**Colour** changes arise from changes in
1. oxidation state,
2. co-ordination number
3. ligand.

\[
\text{Co(H}_2\text{O)}_6^{2+} + 4\text{Cl}^- \rightarrow [\text{CoCl}_4]^{2-} + 6\text{H}_2\text{O}
\]

pink → blue

In this equation both ligand and co-ordination number are changing

**How colour arises**

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.

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

**Compounds without colour**

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.

**Spectrophotometry**

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 spectrophotometry 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.
Variable oxidation states

Transition elements show variable oxidation states When transition metals form ions they lose the 4s electrons before the 3d

General trends

• 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., MnO₄⁻
• Compounds with low oxidation states are often reducing agents, e.g., V²⁺ & Fe²⁺

Reducing Chromium

Cr³⁺ (green) and Cr²⁺ (blue) are formed by reduction of Cr₂O₇²⁻ (orange) by zinc in acid solution, but Fe²⁺ will only reduce it to Cr³⁺. The Fe²⁺ and Cr₂O₇²⁻ in acid solution reaction can be used as a quantitative redox titration. This does not need an indicator.

Manganese redox titration

The redox titration between Fe²⁺ with MnO₄⁻ (purple) is a very common exercise. This titration is self-indicating because of the significant color 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 Cl₂ by MnO₄⁻. It cannot be nitric acid as this is an oxidising agent. It cannot be conc H₂SO₄ 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⁺).

Manganese titration example

A 2.41g nail made from an alloy containing iron is dissolved in 100 cm³ acid. The solution formed contains Fe(II) ions. 10 cm³ portions of this solution are titrated with potassium manganate (VII) solution of 0.02 M. 9.80 cm³ of KMnO₄ were needed to react with the solution containing the iron. What is the percentage of iron by mass in the nail?

MnO₄⁻(aq) + 8H⁺(aq) + 5Fe²⁺ → Mn²⁺(aq) + 4H₂O + 5Fe³⁺ Purple colourless

Step 1: find moles of KMnO₄

moles = conc x vol

0.02 x 9.8/1000
= 1.96 x 10⁻⁴ mol

Step 2: find moles Fe²⁺ in 10 cm³

= moles of KMnO₄ x 5
= 9.8 x 10⁻⁴ mol

Step 3: find moles Fe²⁺ in 100 cm³

= 9.8 x 10⁻⁴ mol x 10
= 9.8 x 10⁻³ mol

Step 4: find mass of Fe in 9.8 x 10⁻³ mol

mass = moles x RAM = 9.8 x 10⁻³ x 55.8 = 0.547 g

Step 5: find % mass

%mass = 0.547/2.41 x 100
= 22.6%

Other useful manganate titrations

With hydrogen peroxide

Ox H₂O₂ → O₂ + 2H⁺ + 2e⁻
Red MnO₄⁻(aq) + 8H⁺(aq) + 5e⁻ → Mn²⁺(aq) + 4H₂O
Overall 2MnO₄⁻(aq) + 6H⁺(aq) + 5H₂O₂ → 5O₂ + 2Mn²⁺(aq) + 8H₂O

With ethanedioate

Ox C₂O₄²⁻ → 2CO₂ + 2e⁻
Red MnO₄⁻(aq) + 8H⁺(aq) + 5e⁻ → Mn²⁺(aq) + 4H₂O
Overall 2MnO₄⁻(aq) + 16H⁺(aq) + 5C₂O₄²⁻ → 10CO₂ + 2Mn²⁺(aq) + 8H₂O
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.

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 Cr(OH)₆³⁻ → CrO₄²⁻:
- Add H₂O to balance O: Cr(OH)₆³⁻ → CrO₄²⁻ + 2H₂O + 3e⁻
- Add H⁺ to balance H: Cr(OH)₆³⁻ → CrO₄²⁻ + 2H₂O + 2H⁺ + 3e⁻
- Add OH⁻ to both sides to cancel out H⁺:
  \[ \text{Cr(OH)}_6^{3-} + 3\text{OH}^- \rightarrow \text{CrO}_4^{2-} + 2\text{H}_2\text{O} + 3\text{e}^- \]

Ammonical oxidation of Cobalt

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

H₂O₂ could also bring about the oxidation.

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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: \(2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3\)

Step 1 \(\text{SO}_2 + \text{V}_2\text{O}_5 \rightarrow \text{SO}_3 + \text{V}_2\text{O}_4\)

Step 2 \(2\text{V}_2\text{O}_4 + \text{O}_2 \rightarrow 2\text{V}_2\text{O}_5\)

Cr_2O_3 catalyst is used in the manufacture of methanol from carbon monoxide and hydrogen.

\(\text{CO} + 2\text{H}_2 \rightarrow \text{CH}_3\text{OH}\)

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.

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 \( S_2O_8^{2-} \) catalysed by Fe²⁺

**overall** \( S_2O_8^{2-} + 2I⁻ \rightarrow 2SO_4^{2-} + I₂ \)

Catalysed alternative route

**stage 1** \( S_2O_8^{2-} + 2Fe^{2+} \rightarrow 2SO_4^{2-} + 2Fe^{3+} \)

**stage 2** \( 2I⁻ + 2Fe^{3+} \rightarrow 2Fe^{2+} + I₂ \)

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 \( C_2O_4^{2-} \) with \( MnO_4^- \)

**overall** \( 2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O \)

Catalysed alternative route

**Step 1** \( 4Mn^{2+} + MnO_4^- + 8H^+ \rightarrow 5Mn^{3+} + 4H_2O \)

**Step 2** \( 2Mn^{3+} + C_2O_4^{2-} \rightarrow 2Mn^{2+} + 2CO_2 \)

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.

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 from 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

\[ Pt(NH_3)_2Cl_2 + 2H_2O \rightarrow [Pt(NH_3)_2Cl(H_2O)]^+ + Cl^- \]

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

\[ [Ag(NH_3)_2]^+ \] 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 donor

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

**Metals and aqua ions**

Metal aqua ions are formed in aqueous solution

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

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

\[
[M(H_2O)_6]^{2+} + H_2O \rightleftharpoons [M(H_2O)_5(OH)]^{2+} + H_3O^+
\]

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

\[
\text{Cu(H}_2\text{O)}_6^{2+}(aq) + 2\text{OH}^- (aq) \rightarrow \text{Cu(H}_2\text{O)}_4\text{(OH)}_2(s) + 2\text{H}_2\text{O(l)}
\]

\[
\text{Co(H}_2\text{O)}_6^{2+}(aq) + 2\text{NH}_3(aq) \rightarrow \text{Co(H}_2\text{O)}_4(\text{OH})_2(s) + 2\text{NH}_4^+(aq)
\]

\[
\text{Fe(H}_2\text{O)}_6^{3+}(aq) + 3\text{OH}^- (aq) \rightarrow \text{Fe(H}_2\text{O)}_3(\text{OH})_3(s) + 3\text{H}_2\text{O(l)}
\]

\[
\text{Cr(H}_2\text{O)}_6^{3+}(aq) + 3\text{NH}_3(aq) \rightarrow \text{Cr(H}_2\text{O)}_3(\text{NH})_3(s) + 3\text{NH}_4^+(aq)
\]

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

\[
\text{e.g. } \text{Cr(H}_2\text{O)}_6^{3+}(aq) + \text{OH}^- (aq) \rightarrow [\text{Cr(H}_2\text{O)}_6(\text{OH})]^{2+}(aq) + \text{H}_2\text{O(l)}
\]

**Reaction with excess OH$^-$**

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

Al becomes \([\text{Al(OH)}_4]^-(aq)\) colourless solution

\[
\text{Cr(H}_2\text{O)}_6(\text{OH})_3(s) + 3\text{OH}^- (aq) \rightarrow [\text{Cr(OH)}_6]^{3-}(aq) + 3\text{H}_2\text{O(l)}
\]

\[
\text{Al(H}_2\text{O)}_6(\text{OH})_3(s) + \text{OH}^- (aq) \rightarrow [\text{Al(OH)}_4]^-(aq) + 3\text{H}_2\text{O(l)}
\]

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$_2$O are similar in size and are uncharged. Ligand exchange occurs without change of co-ordination number for Co and Cr

Cr becomes \([\text{Cr(NH)}_3]^{3+}\) purple solution

Co becomes \([\text{Co(NH)}_3]^{2+}\) pale yellow solution

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

Cu becomes \([\text{Cu(NH)}_3]^{2+}\) deep blue solution

\[
\text{Cr(OH)}_3(\text{H}_2\text{O})_3(s) + 6\text{NH}_3(aq) \rightarrow [\text{Cr(NH)}_3]^{3+}(aq) + 3\text{H}_2\text{O(l)} + 3\text{OH}^- (aq)
\]

\[
[\text{Cu(OH)}_3(\text{H}_2\text{O})_3](s) + 4\text{NH}_3(aq) \rightarrow [\text{Cu(NH)}_3]^{2+}(\text{H}_2\text{O})_2(aq) + 2\text{H}_2\text{O(l)} + 2\text{OH}^- (aq)
\]

\[
[\text{Co(OH)}_2(\text{H}_2\text{O})_4](s) + 6\text{NH}_3(aq) \rightarrow [\text{Cr(NH)}_3]^{2+}(aq) + 4\text{H}_2\text{O(l)} + 2\text{OH}^- (aq)
\]
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 \(\text{H}_2\text{O}\) and \(\text{NH}_3\) ligands and that ligand exchange can involve a change of coordination number.

Addition of HCl to aqueous ions of Cu and Co lead to a change in coordination number from 6 to 4. These are tetrahedral in shape.

\[
\text{[Cu(H}_2\text{O)}_3]^{2+} + 4\text{Cl}^{-} \rightarrow \text{[CuCl}_4]^{2-} + 6\text{H}_2\text{O}
\]

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).

The 3+ ions with carbonate solution form a M(OH)\(_3\) ppt and CO\(_2\) gas is evolved.

MCO\(_3\) is formed with 2+ ions but M\(_2\)(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.

Chromate/dichromate equilibrium

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

\[
2\text{CrO}_4^{2-} + 2\text{H}^+ \rightleftharpoons \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}
\]

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.

\[
\text{Ni(H}_2\text{O)}_6^{2+} (aq) + 3\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2 (aq) \rightarrow \text{[Ni(NH}_2\text{CH}_2\text{CH}_2\text{NH}_2)}_3]^{2+} (aq) + 6\text{H}_2\text{O} (l)
\]

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.

\[
\text{Co(NH}_3)_6^{2+} (aq) + 3\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2 (aq) \rightarrow \text{[Co(NH}_2\text{CH}_2\text{CH}_2\text{NH}_2)_3]^{2+} (aq) + 6\text{NH}_3 (l)
\]

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

[M(H₂O)₆]²⁺

[CuCl₄]²⁻ yellow sol
[CoCl₄]²⁻ blue sol

Cl⁻

[Cu(H₂O)₆]²⁺ blue sol
[Fe(H₂O)₆]³⁺ green sol
[Co(H₂O)₆]³⁺ pink sol

OH⁻ or NH₃

M(OH)₂(H₂O)₄

Cu(OH)₂(H₂O)₄ blue ppt
Fe(OH)₂(H₂O)₄ green ppt
Co(OH)₂(H₂O)₄ blue ppt

air

[Co(NH₃)₆]³⁺ brown sol

3+ ion summary

[Cr(OH)₆]³⁻ green sol
[Al(OH)₃]³⁻ colourless sol

OH⁻ or NH₃

M(OH)₃(H₂O)₄

Cr(OH)₃(H₂O)₃ green ppt
Fe(OH)₃(H₂O)₃ brown ppt
Co(OH)₃(H₂O)₃ brown ppt

air

[Cr(NH₃)₆]³⁺ violet sol

H^⁺
Chromium summary

$\text{Cr}_2\text{O}_7^{2-} (aq)$ Orange solution $\xrightarrow{\text{acid}} \text{CrO}_4^{2-} (aq)$ Yellow solution

$\text{Zinc and HCl}$

$\text{Cr(H}_2\text{O)}_6^{3+} (aq)$ Green solution

$\text{Limited NaOH or NH}_3$

$\text{Cr(H}_2\text{O)}_6^{3+} (aq)$ Green precipitate

$\text{Excess NaOH}$

$\text{Cr(OH)}_6^{3-} (aq)$ Green solution

Silver Chemistry

Reactions of halides with Silver nitrate
Fluorides produce no precipitate
Chlorides produce a white precipitate
$\text{Ag}^+(aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl}(s)$
Bromides produce a cream precipitate
$\text{Ag}^+(aq) + \text{Br}^- (aq) \rightarrow \text{AgBr}(s)$
Iodides produce a pale yellow precipitate
$\text{Ag}^+(aq) + \text{I}^- (aq) \rightarrow \text{AgI}(s)$

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
$\text{AgCl}(s) + 2\text{NH}_3(aq) \rightarrow [\text{Ag(NH}_3)_2]^+ (aq) + \text{Cl}^- (aq)$

Silver bromide dissolves in concentrated ammonia to form a complex ion
$\text{AgBr}(s) + 2\text{NH}_3(aq) \rightarrow [\text{Ag(NH}_3)_2]^+ (aq) + \text{Br}^- (aq)$

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

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

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

$\text{NaBr}$

Dilute NH$_3$

Dilute NH$_3$

Conc HNO$_3$

Dilute NH$_3$

Conc NH$_3$

Conc HNO$_3$