

Experiment / Demonstration

Title: Ligand Exchange Reactions and Colours of Transition Metal Complexes

Learning Focus

Students will:

- Observe how **ligand substitution** leads to distinct **colour changes** in transition-metal complexes
- Identify **aqua, ammine, chloro, and hydroxo complexes** of Cu^{2+} and Co^{2+}
- Relate observations to **coordination number, ligand identity, and equilibrium shifts**

Reagents (Teacher-controlled)

- Copper(II) sulfate solid, **$\text{CuSO}_4(\text{s})$**
- Cobalt(II) chloride solid, **$\text{CoCl}_2(\text{s})$**
- Distilled water
- **Aqueous ammonia, 6 mol dm^{-3}**
- **Sodium hydroxide, 2 mol dm^{-3}**
- **Concentrated hydrochloric acid, 18 mol dm^{-3}** (demonstration only)

Safety Notes (Important)

- **$18 \text{ mol dm}^{-3} \text{HCl}$** : corrosive, fumes – use in **fume hood**, teacher-handled only
- **$6 \text{ mol dm}^{-3} \text{NH}_3$** : irritant, strong smell – use dropwise
- Goggles required; students observe rather than handle concentrated reagents

Procedure

Part A: Copper(II) Complexes

1. Dissolve a small spatula of **$\text{CuSO}_4(\text{s})$** in distilled water
→ forms the **aqua complex**
2. Add **$2 \text{ mol dm}^{-3} \text{NaOH}$** dropwise
→ observe precipitate formation
3. Add a few drops of **$6 \text{ mol dm}^{-3} \text{NH}_3$** , and slowly continue adding **NH_3** until excess
→ observe dissolution and colour change
4. (Optional extension) Add **$18 \text{ mol dm}^{-3} \text{HCl}$** carefully
→ observe ligand exchange back to chloro/aqua complex

Part B: Cobalt(II) Complexes

1. Dissolve **CoCl₂(s)** in distilled water
→ observe initial colour (temperature-dependent discussion possible)
2. Add **2 mol dm⁻³ NaOH** dropwise
→ precipitate forms
3. Add a few drops of **6 mol dm⁻³ NH₃**, and slowly continue adding **NH₃** until excess
→ precipitate dissolves, new colour forms
4. Add **18 mol dm⁻³ HCl** (teacher demonstration)
→ observe colour change due to formation of chloro complex

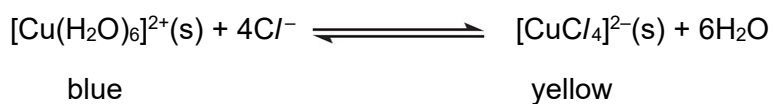
Key Teaching Points to Emphasise

- Ligand identity affects d-orbital splitting → colour
- Ligand exchange is reversible and equilibrium-based
- Precipitates vs soluble complexes
- Coordination number changes (e.g. 6 → 4)

Student Observation Table

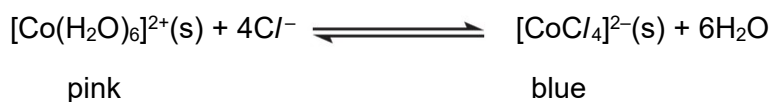
Copper(II) System

Step	Reagent added	Observation (Colour / State)	Complex / Species Formed
1	Water	Blue solution	$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ (aq)
2	NaOH	Pale blue precipitate	$\text{Cu}(\text{OH})_2$ (s)
3	A few drops NH_3	Pale blue precipitate	$\text{Cu}(\text{OH})_2$ (s)
4	Excess NH_3	Deep royal blue solution	$[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ (aq)
5	HCl	Blue fades / turns to blue-green	$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ / $[\text{CuCl}_4]^{2-}$



Cobalt(II) System

Step	Reagent added	Observation (Colour / State)	Complex / Species Formed
1	Water	Pink solution	$[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ (aq)
2	NaOH	Blue precipitate	$\text{Co}(\text{OH})_2$ (s)
3	A few drops NH_3	Blue precipitate	$\text{Co}(\text{OH})_2$ (s)
4	Excess NH_3	Straw/yellow-brown solution, turns darker brown in air	$[\text{Co}(\text{NH}_3)_6]^{2+}$ (aq)
5	HCl	Turns to violet solution	$[\text{CoCl}_4]^{2-}$ (aq)



Optional Extension Questions

- Why does **NH₃** dissolve the hydroxide precipitate?
- Which ligands cause a **tetrahedral complex**?
- How does **ligand field strength** relate to observed colour?
- Why is **[CoCl₄]²⁻** blue but **[Co(H₂O)₆]²⁺** pink?