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, **$CuSO_4$(s)**
- Cobalt(II) chloride solid, **$CoCl_2$(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 mol dm^{-3} HCl**: corrosive, fumes – use in **fume hood**, teacher-handled only
- **6 mol dm^{-3} 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 **$CuSO_4$(s)** in distilled water
   → forms the **aqua complex**
2. Add **2 mol dm^{-3} NaOH** dropwise
   → observe precipitate formation
3. Add a few drops of **6 mol dm^{-3} NH_3**, and slowly continue adding **NH_3** until excess
   → observe dissolution and colour change
4. (Optional extension) Add **18 mol dm^{-3} HCl** carefully
   → observe ligand exchange back to chloro/aqua complex

Part B: Cobalt(II) Complexes

1. Dissolve **$CoCl_2$(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} NH_3**, and slowly continue adding **NH_3** 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	$[Cu(H_2O)_6]^{2+}$ (aq)
2	NaOH	Pale blue precipitate	$Cu(OH)_2(s)$
3	A few drops NH_3	Pale blue precipitate	$Cu(OH)_2(s)$
4	Excess NH_3	Deep royal blue solution	$[Cu(NH_3)_4(H_2O)_2]^{2+}$ (aq)
5	HCl	Blue fades / turns to blue-green	$[Cu(H_2O)_6]^{2+}$ / $[CuCl_4]^{2-}$

$[Cu(H_2O)_6]^{2+}(s) + 4Cl^- \rightleftharpoons [CuCl_4]^{2-}(s) + 6H_2O$

blue yellow

Cobalt(II) System

Step	Reagent added	Observation (Colour / State)	Complex / Species Formed
1	Water	Pink solution	$[Co(H_2O)_6]^{2+}$ (aq)
2	NaOH	Blue precipitate	$Co(OH)_2(s)$
3	A few drops NH_3	Blue precipitate	$Co(OH)_2(s)$
4	Excess NH_3	Straw/yellow-brown solution, turns darker brown in air	$[Co(NH_3)_6]^{2+}$ (aq)
5	HCl	Turns to violet solution	$[CoCl_4]^{2-}$ (aq)

$[Co(H_2O)_6]^{2+}(s) + 4Cl^- \rightleftharpoons [CoCl_4]^{2-}(s) + 6H_2O$

pink blue

Optional Extension Questions

- Why does **NH_3 dissolve the hydroxide precipitate**?
- Which ligands cause a **tetrahedral complex**?
- How does **ligand field strength** relate to observed colour?
- Why is **$[CoCl_4]^{2-}$ blue but $[Co(H_2O)_6]^{2+}$ pink**?