

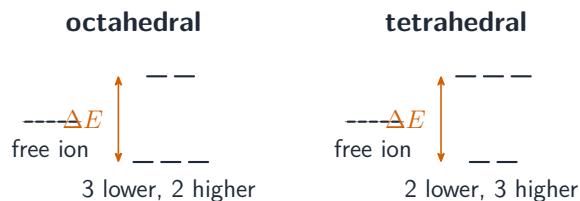
Exercise walk-through

28.3

Colour of complexes

A-Level Chemistry

Why complexes are coloured



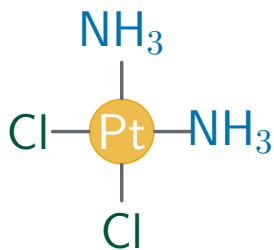
the complex absorbs light of energy ΔE , so we see the **complementary colour**

Ligands split the five **degenerate** d orbitals into two sets, a gap ΔE apart.

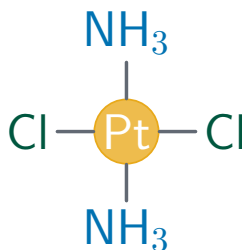
The complex absorbs light of energy ΔE ; we see the **complementary colour**.
Different ligands give a different ΔE – so the colour changes.

Stereoisomerism in complexes

square planar $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$



cis (same side)



trans (opposite)

- **cis/trans** in square planar (e.g. $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$) & octahedral complexes
- **optical** in octahedral complexes with bidentate ligands

Isomerism in complexes, and predicting polarity

geometrical isomerism

cis and *trans* in square planar and octahedral complexes

optical isomerism

non-superimposable mirror images, in octahedral complexes with bidentate ligands

polarity

a *cis* form **may be polar**; the matching *trans* form is often **non-polar**, because its bond dipoles cancel

ligand exchange colours

$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ is blue; with concentrated HCl it becomes **yellow** $[\text{CuCl}_4]^{2-}$, and cobalt(II) behaves similarly

redox feasibility

use $E^\ominus$ values to predict whether a transition-metal redox reaction will go

The standard titrations to be able to calculate are MnO_4^- against Fe^{2+} , and thiosulfate against iodine. Both are asked most years.

Stability constants

K_{stab} : equilibrium for forming a complex

A **large stability constant** means a very stable complex – ligand exchange moves towards the larger K_{stab} .

Use $E^\ominus$ values to predict redox feasibility (e.g. $\text{MnO}_4^-/\text{Fe}^{2+}$ titrations, purple to colourless).

Four key properties (and why)

variable oxidation states

act as catalysts

form complex ions

coloured compounds

The four key properties of the transition elements

Exercises 2.1-2.3

2.1 Define **degenerate** d orbitals.

[1]

2.2 State how the five d orbitals split in an **octahedral** complex.

[1]

2.3 Explain why Zn^{2+} complexes are colourless.

[2]

Exercises 2.1-2.3

2.1 Define **degenerate** d orbitals.

[1]

2.1

- d orbitals that have the same energy

2.2 State how the five d orbitals split in an **octahedral** complex.

[1]

2.2

- into a lower set of three and a higher set of two orbitals

2.3 Explain why Zn^{2+} complexes are colourless.

[2]

2.3

- Zn^{2+} has a full $3d^{10}$ sub-shell; there is no vacancy to promote an electron into, so no d-d transition occurs and no visible light is absorbed

Exercises 2.4-2.5

2.4 State the relationship between the energy gap ΔE and the frequency of light absorbed. [1]

2.5 Name three factors that change the size of ΔE . [3]

Exercises 2.4·2.5

2.4 State the relationship between the energy gap ΔE and the frequency of light absorbed. [1]

2.4

- $\Delta E = h\nu$, so a larger gap corresponds to a higher frequency absorbed

2.5 Name three factors that change the size of ΔE . [3]

2.5

- The identity of the ligand; the oxidation state of the metal ion; the coordination number or the shape of the complex

Exercises 3.1·3.2

3.1 A transition-metal complex is coloured because an electron is promoted:

[1]

☐

A out of the atom

B between two non-degenerate d orbitals

C into the nucleus

D between s orbitals

3.2 The colour we see is:

[1]

☐

A the colour of light absorbed

B the complementary colour of the light absorbed

C always green

D white

Exercises 3.1-3.2

3.1 A transition-metal complex is coloured because an electron is promoted:

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A out of the atom

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3.1 B

- An electron is promoted between two non-degenerate d orbitals

3.2 The colour we see is:

[1]

☐

A the colour of light absorbed

B the complementary colour of the light absorbed

C always green

D white

3.2 B

- We see the complementary colour of the light absorbed

Exercises 3.3-3.4

3.3 Explain, in terms of d orbitals, why $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ is coloured. [3]

3.4 Explain why changing the ligand around a transition-metal ion changes the colour of the complex. [2]

Exercises 3.3-3.4

3.3 Explain, in terms of d orbitals, why $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ is coloured. [3]

3.3

- the ligands split the d orbitals into two sets separated by ΔE ; the complex absorbs light of energy equal to ΔE , promoting a d electron to the higher set; the colour seen is the complementary colour of the light absorbed

3.4 Explain why changing the ligand around a transition-metal ion changes the colour of the complex. [2]

3.4

- a different ligand gives a different splitting energy ΔE ; so a different frequency of light is absorbed, and a different complementary colour is seen

Exercise 3.5 (a)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(a) Explain, in terms of d orbitals, why this complex is coloured.

[4]

Exercise 3.5 (a)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(a) Explain, in terms of d orbitals, why this complex is coloured. [4]

(a)

- The ligands split the five degenerate 3d orbitals into two levels separated by ΔE ; Co^{2+} has a partially filled d sub-shell, so an electron can be promoted; it absorbs a photon whose energy equals ΔE , a d–d transition; the light not absorbed is transmitted, and the complementary colour is what is seen

Exercise 3.5 (b)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(b) The pink solution absorbs green light. State the relationship between the colour absorbed and the colour seen. [1]

Exercise 3.5 (b)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(b) The pink solution absorbs green light. State the relationship between the colour absorbed and the colour seen. [1]

(b)

- The colour seen is the complement of the colour absorbed

Exercise 3.5 (c)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(c) Concentrated hydrochloric acid is added and the solution turns blue, forming $[\text{CoCl}_4]^{2-}$.
Explain the colour change, referring to both the ligand and the shape. [4]

Exercise 3.5 (c)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(c) Concentrated hydrochloric acid is added and the solution turns blue, forming $[\text{CoCl}_4]^{2-}$. Explain the colour change, referring to both the ligand and the shape. [4]

(c)

- Chloride ligands replace the water by ligand exchange; chloride is lower in the spectrochemical series than water, so ΔE decreases; the coordination number also falls from 6 to 4 and the shape changes from octahedral to tetrahedral, which splits the orbitals differently; a different frequency is therefore absorbed and the complementary colour seen changes from pink to blue

Exercise 3.5 (d)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(d) A solution of $[\text{Sc}(\text{H}_2\text{O})_6]^{3+}$ is colourless. Explain why.

[2]

Exercise 3.5 (d)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(d) A solution of $[\text{Sc}(\text{H}_2\text{O})_6]^{3+}$ is colourless. Explain why. [2]

(d)

- Sc^{3+} has an empty $3d^0$ sub-shell; with no d electrons there can be no d–d transition, so no visible light is absorbed

Exercise 3.5 (e)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(e) Predict, with a reason, whether $[\text{Co}(\text{CN})_6]^{4-}$ would absorb light of a higher or lower frequency than $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$. [2]

Exercise 3.5 (e)

3.5 Aqueous cobalt(II) chloride is pink because of the complex $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$.

(e) Predict, with a reason, whether $[\text{Co}(\text{CN})_6]^{4-}$ would absorb light of a higher or lower frequency than $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$. [2]

(e)

- A higher frequency; cyanide is higher in the spectrochemical series than water, so it produces a larger ΔE , and $\Delta E = h\nu$