Same metal, two colours

both coloured by Cr^{3+}



Ruby (red) and emerald (green) owe their colour to the **same ion, Cr^{3+}**.

Transition-metal ions are coloured because their **d electrons absorb visible light**.

White light = all colours



Remove **one band** from white light...

...and the remaining colours mix into a **new colour**.

Ligands split the d orbitals



In a free ion the five d orbitals have **equal energy**.

Ligands (water in solution) split them into a lower and an upper level with an **energy gap**.

The d–d transition



Ti^{3+} ($3d^1$): electron now in the upper level

A d electron absorbs light whose **energy = the gap** and jumps from the lower to the upper level.

For most ions this light is in the **visible** region.

We see the complementary colour



absorbed: green-yellow

seen: purple (Ti^{3+})

absorbed: red-orange

seen: blue (Cu^{2+})

The colour we see is **what is left**, not what is absorbed.

$Ti^{3+}(aq)$ absorbs ≈ 500 nm (green-yellow) → **purple**.

d^0 and d^{10}: colourless



$3d^0$: nobody to jump

$3d^{10}$: no room above

No electron to jump (d^0: Sc^{3+}, Ti^{4+}) or no vacancy above (d^{10}: Zn^{2+}): no d–d transition.

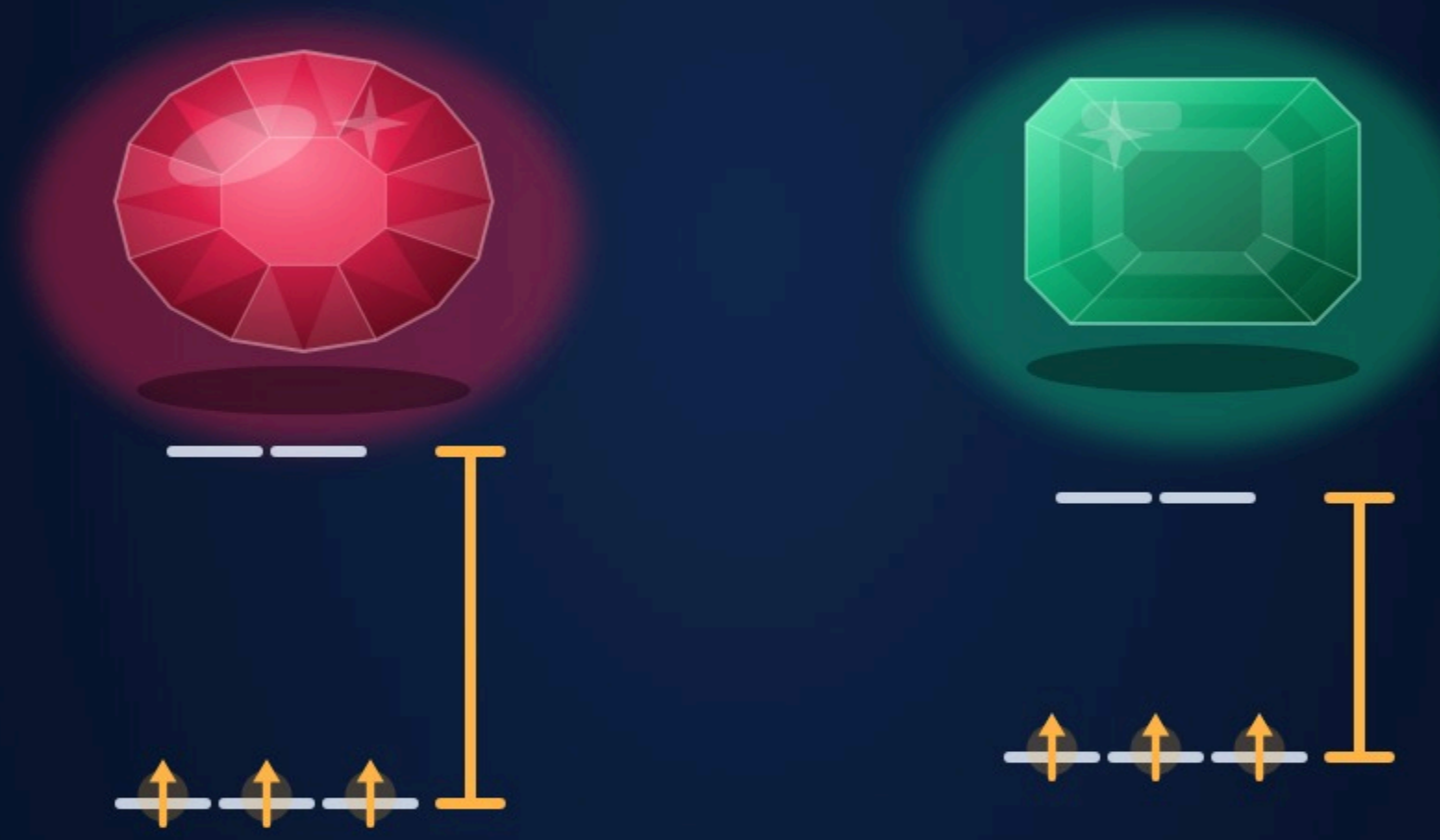
All light passes: the solution is **colourless**.

Colours of aquated ions (NCERT Table 4.8)

- Sc^{3+} — $3d^0$ · colourless
- Ti^{3+} — $3d^1$ · purple
- V^{3+} — $3d^2$ · green
- Cr^{3+} — $3d^3$ · violet
- Mn^{2+} — $3d^5$ · pink
- Fe^{3+} — $3d^5$ · yellow
- Fe^{2+} — $3d^6$ · green
- Co^{2+} — $3d^7$ · pink
- Ni^{2+} — $3d^8$ · green
- Cu^{2+} — $3d^9$ · blue
- Zn^{2+} — $3d^{10}$ · colourless

Only $3d^0$ and $3d^{10}$ ions are colourless.

The ligand sets the gap



eats yellow-green

eats red-orange

Ruby: bigger gap → absorbs yellow-green → red. Emerald: smaller gap → absorbs red-orange → green.

Your turn: Ti^{3+} or Ti^{4+}, which is colourless?