2. Colour in Coordination Compounds

Color arises due to following reasons:-

(i) d-d transition:-
Here, d^0 & d^{10} → colourless

(ii) Charge Transfer Spectra:-

- LMCT → Ligand to Metal Charge Transfer $[MnO_4^-, Cr_2O_7^{2-}, CrO_4^{2-}]$
- MLCT → Metal to Ligand Charge Transfer
- MMCT → Metal to Metal Charge Transfer $Fe_4[Fe(CN)_6]_3$

Here,

$\underline{Fe^{3+}}$ $\underline{[Fe(CN)_6]^{4-}_3}$ (Charge transfer bet^n Fe^{3+} & Fe^{2+})

$3d^5$ $Fe^{2+}(3d^6)$

(iii) Polarization :- AgBr (pale yellow), AgI (yellow)

By Fazan's rule, Size of cation $\propto$ $\frac{1}{\text{Size of Anion}}$

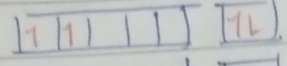
In Coordination compounds,

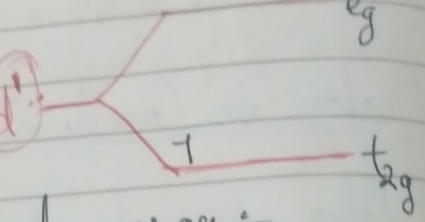
Main reason for color ⇒ d-d transition

Note :- VBT doesn't explain colour of coordination compounds
But,
CFT explains colour of such compounds

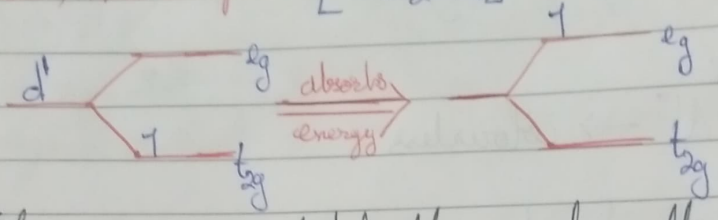
d-d transition :-

Eg $[Ti(H_2O)_6]^{3+}$: purple colour

↳ $Ti^{3+} \rightarrow$ 

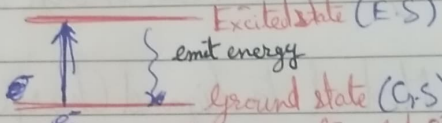
∴ 

Now, This complex $[Ti(H_2O)_6]^{3+}$ absorbs energy :-



When energy is absorbed by this complex, then :-
The e⁻ in t_{2g} gets excited & moves to e_g orbital

* Energy state :-



We have two states :- Ground State & Excited State

∴ In this example, ~~then~~

1) When e⁻ absorbs energy, this e⁻ jumps from G.S to E.S

2) The e⁻ then returns to the G.S & in the process, it emits energy → $E = h\nu$ or $E = \frac{hc}{\lambda}$

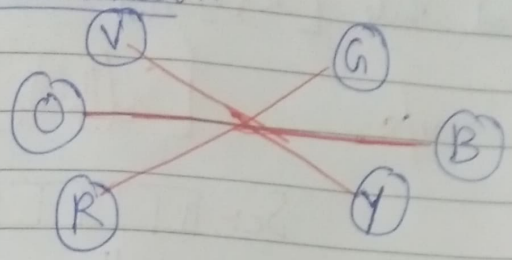
This wavelength falls under Visible spectrum (VIBGYOR)

Hence, coordination compounds show colour this way

∴ e⁻ emits energy in the form of radiation

* d^0 & d^{10} complexes are colourless because d-d transition can't occur :- d^0 → no e⁻ ∴ no d-d transition
d^{10} → fully filled e⁻ ∴ no d-d transition

Marshall Wheel :-



In this wheel, opposite colours are Complementary

This means :- If we pass Violet light to a coordination complex then :-

The complex appears to be Yellow in color

Here, V Y :- Violet ⟷ Yellow
G R :- Green ⟷ Red
O B :- Orange ⟷ Green

Wavelength of Visible Spectrum :-

~~1. Violet :-~~

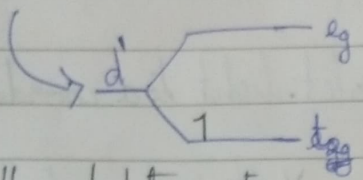
1. Violet :- 380 to 450 nm
2. Indigo :- 445 to 450 nm
3. Blue :- 450 to 495 nm
4. Green :- 495 to 570 nm
5. Yellow :- 570 to 595 nm
6. Orange :- 590 to 620 nm
7. Red :- 620 to 750 nm

Q.1. $[Ti(H_2O)_6]^{3+}$ is coloured while $[Sc(H_2O)_6]^{3+}$ is colourless. Why?

Sol:- $[Ti(H_2O)_6]^{3+}$ v/s $[Sc(H_2O)_6]^{3+}$

Ti:- ↑ | | | | | TL

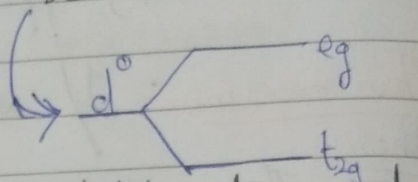
Ti^{3+}:- ↑ | | | | | □



Here, d-d transition can occur

Sc:- | | | | | | TL

Sc^{3+}:- | | | | | | □



Here, d-d transition cannot occur (∵ There is no e⁻ in d-orbital)

Hence, $[Ti(H_2O)_6]^{3+}$ → coloured & $[Sc(H_2O)_6]^{3+}$ → colourless

Q2. Arrange the following complexes in the increasing order of CFS:- $[Fe(H_2O)_6]^{2+}$, $[Fe(H_2O)_6]^{3+}$, $[FeCl_6]^{4-}$

~~Sol:- for $[Fe(H_2O)_6]^{2+}$ ~~Fe^{2+}~~ ~~| | | | |~~ ~~Charge on Fe~~~~

Sol:- Magnitude of Δ_0 depends on:- 1) Charge of metal ∝ CFS
2) Nature of ligand

(I) $[\underset{\downarrow}{Fe}(H_2O)_6]^{2+}$, (II) $[\underset{\downarrow}{Fe}(H_2O)_6]^{3+}$, (III) $[\underset{\downarrow}{Fe}Cl_6]^{4-}$

(I) $x = +2$; (II) $x = +3$; (III) $x = +4+(-1) = +2$

Here, Charge on metal is different for (II) & and (I) & (III)

(II) has more charge ∴ More splitting

② Charge on metal same for (I) & (III)
We check for nature of ligand (Spectrochemical series)

(I) has H_2O & (III) has Cl. $Cl < H_2O$ [in terms of strength] (In this series)

∴ Δ_0 arrangement:- $[FeCl_6]^{4-} < [Fe(H_2O)_6]^{2+} < [Fe(H_2O)_6]^{3+}$

[(III) < (I) < (II)]