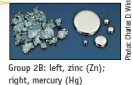
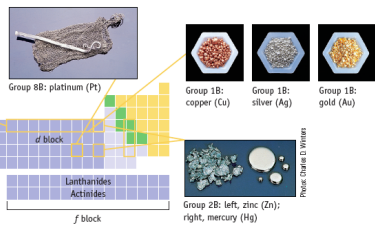


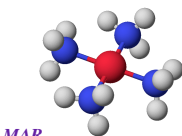
Chemistry of Coordination Compounds - Chapter 19



Fourth-period transition metals: left to right, Ti, V, Cr, Mn, Fe, Co, Ni, Cu



Group 2B: left, zinc (Zn); right, mercury (Hg)



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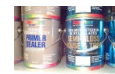
Last update: 4/29/24

Chemistry 223
Professor Michael Russell

Color Theory

Where does the color of objects come from?

From the paint covering the object!



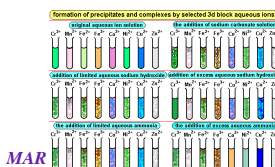
Where does the paint gets its color?

From the paint pigments!



What are the pigments?

Colorful **transition metal** compounds!



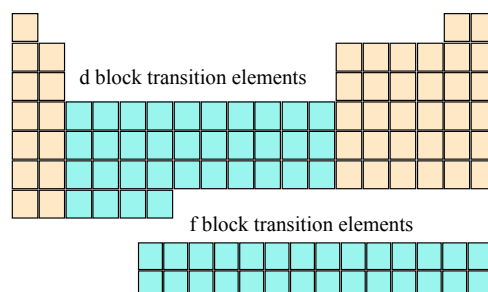
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Why Study Transition Metals

- ◆ Transition metals found in nature
 - ▲ Rocks and minerals contain transition metals
 - * Red rubies (Cr), blue sapphires (Fe and Ti)
 - ▲ Many biomolecules contain transition metals
 - * Vitamin B12 (Co), Hemoglobin, myoglobin, and cytochrome C (all Fe)
- ◆ Transition metals used in industry
 - ▲ Material science (steel, alloys)
 - ▲ Transition metal compounds are used as pigments
 - * TiO₂ (white), PbCrO₄ (yellow), Fe₄[Fe(CN)₆]₃ (Prussian blue)

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Periodic Table



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d-Block Transition Elements

IIIB	IVB	VB	VIB	VIIIB	VIII B			IB	IIB
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd
La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg

Most have partially occupied d subshells in common oxidation states

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Transition Metals

- ◆ General Properties
 - ▲ Have typical metallic properties (malleable, etc.)
 - ▲ Not as reactive as alkali and alkaline earth metals
 - ▲ Have high melting points, high boiling points, high density
 - ▲ Have 1 or 2 s electrons in valence shell
 - ▲ Differ in # d electrons in n-1 energy level
 - ▲ Exhibit multiple oxidation states
 - ▲ Both **paramagnetic** and **diamagnetic** ions exist
 - ▲ Most ions deeply colored (**crystal field theory**)

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Electronic Configurations

Element	Configuration
Sc	[Ar]3d ¹ 4s ²
Ti	[Ar]3d ² 4s ²
V	[Ar]3d ³ 4s ²
Cr	[Ar]3d ⁵ 4s ¹
Mn	[Ar]3d ⁵ 4s ²
Fe	[Ar]3d ⁶ 4s ²
Ni	[Ar]3d ⁸ 4s ²
Cu	[Ar]3d ¹⁰ 4s ¹
Zn	[Ar]3d ¹⁰ 4s ²

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Oxidation States of Transition Elements

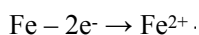
Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
							+1	+1	
	+2	+2	+2	+2	+2	+2	+2	+2	+2
+3	+3	+3	+3	+3	+3	+3	+3	+3	
	+4	+4	+4	+4	+4		+4		
		+5	+5	+5	+5				
			+6	+6	+6				
				+7					

loss of ns e^s loss of ns and (n-1)d e^s

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Electronic Configurations of Transition Metal Ions

Electronic configuration of Fe ions:



valence *ns* electrons
removed first,
then *n-1 d* electrons

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Paramagnetism and Diamagnetism

Paramagnetic species have unpaired electrons while **diamagnetic** species have all paired electrons

Paramagnetic species are attracted or repulsed by magnetic fields; the magnitude of the effect depends on the number of unpaired electrons

Examples:

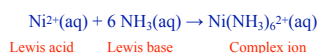
Paramagnetic Cr [Ar] $\uparrow\uparrow\uparrow\uparrow\uparrow\uparrow$ \uparrow
 3 d 4 s

Diamagnetic Pd [Kr] $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$
 4 d 5 s

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Coordination Chemistry

Transition metals act as **Lewis acids** and form **complexes** or **complex ions** with **Lewis bases** or **ligands**



Complex contains central metal ion bonded to one or more molecules or anions

Lewis acid = **metal** = center of coordination

Lewis base = **ligand** = molecules/ions covalently bonded to metal in complex

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Coordination Chemistry

A **coordination compound** has one or more complexes.

Examples: [Co(NH₃)₆]Cl₃, [Cu(NH₃)₄][PtCl₄],
[Pt(NH₃)₂Cl₂]

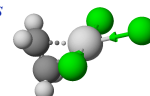
The **coordination number** is the number of donor atoms bonded to the central metal atom or ion in the complex.

Example: [Co(NH₃)₆]³⁺ coordination number = 6

Example: [PtCl₄]²⁻ coordination number = 4

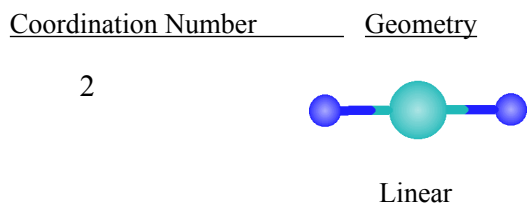
Most common = 6, 4 and 2

Coordination number determined by ligands



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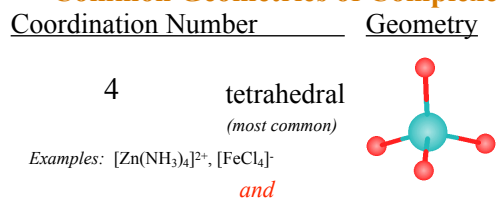
Common Geometries of Complexes



Example: $[\text{Ag}(\text{NH}_3)_2]^+$

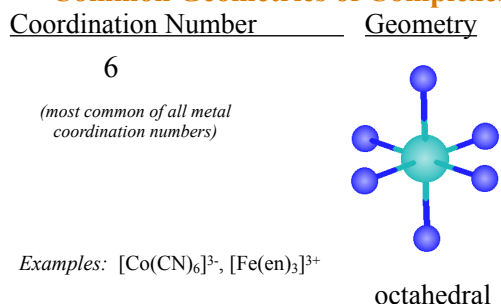
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Common Geometries of Complexes



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Common Geometries of Complexes



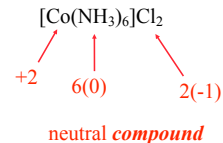
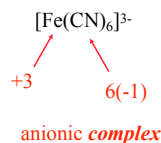
Examples: $[\text{Co}(\text{CN})_6]^{3-}$, $[\text{Fe}(\text{en})_3]^{3+}$

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Coordination Chemistry

Charge of **complex** = sum of charges on the metal and the ligands

Charge of coordination **compound** = sum of charges on metal, ligands, and counterbalancing ions



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Coordination Chemistry

Ligands - aka Lewis bases

- ▲ classified according to the number of bonds to central metal
 - ▲ "dentate" = "tooth"
 - ▲ Examples
 - * monodentate = 1
 - * bidentate = 2
 - * tetradentate = 4
 - * hexadentate = 6
 - * polydentate = 2 or more donor atoms
- ← chelating agents

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Monodentate Ligands

Monodentate ligands possess only one accessible donor group.

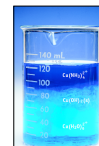
H_2O is a good example since *all* metal ions exist as *aqua* complexes in water

Monodentate Ligand Examples:

- ▲ H_2O , CN^- , NH_3 , NO_2^- , SCN^- , OH^- , X^- (halides), CO , O^{2-}

Example Complexes:

- ▲ $[\text{Co}(\text{NH}_3)_6]^{3+}$
- ▲ $[\text{Fe}(\text{SCN})_6]^{3-}$



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Bidentate Ligands

Bidentate Ligands have "two teeth", able to bond with metal at two separate places

Bidentate Examples:

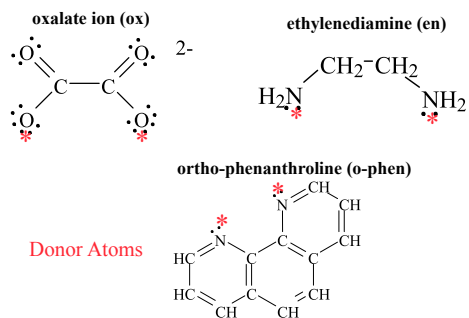
- ▲ oxalate ion = $C_2O_4^{2-}$
- ▲ ethylenediamine (en) = $NH_2CH_2CH_2NH_2$
- ▲ ortho-phenanthroline (o-phen)

Example Complexes:

- ▲ $[Co(en)_3]^{3+}$
- ▲ $[Cr(C_2O_4)_3]^{3-}$
- ▲ $[Fe(NH_3)_4(o-phen)]^{3+}$

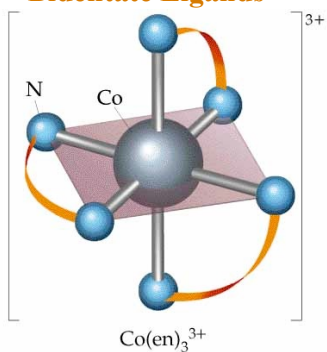
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Bidentate Ligands



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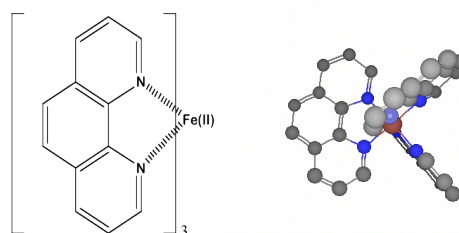
Bidentate Ligands



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Bidentate Ligands

$Fe(o-phen)_3^{2+}$:

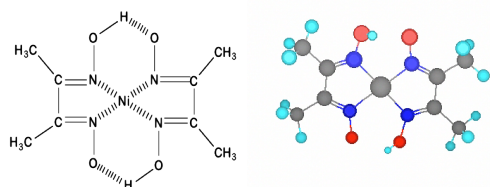


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Bidentate Ligands

Another bidentate ligand:
Dimethylglyoxime (dmg)

$Ni(dmg)_2$ (from lab):



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EDTA

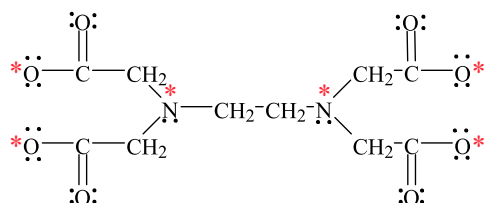
- A commonly used *polydentate* ligand
- Forms 1:1 complexes with most metals except Group IA
- Forms stable, water soluble complexes
- High formation constants
- A *primary standard* material
- See the movie "Blade" (!)

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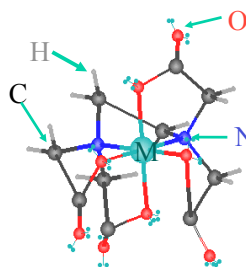
EDTA

Six Donor Regions

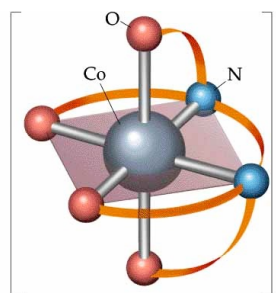


EDTA (Ethylenediamine tetraacetic acid) will adapt its coordination number to bond with the metal

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EDTA



CoEDTA⁻

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EDTA

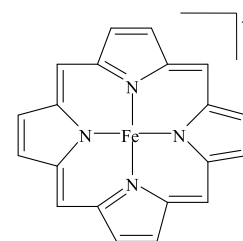
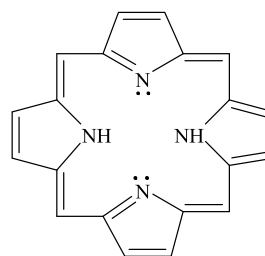
Formation constants (K_f) for some metal - EDTA complexes.

Ion	log K	Ion	log K	Ion	log K
Fe ³⁺	25.1	Pb ²⁺	18.0	La ³⁺	15.4
Th ⁴⁺	23.2	Cd ²⁺	16.5	Mn ²⁺	14.0
Cr ³⁺	23.0	Zn ²⁺	16.5	Ca ²⁺	10.7
Bi ³⁺	22.8	Co ²⁺	16.3	Mg ²⁺	8.7
Cu ²⁺	18.8	Al ³⁺	16.1	Sr ²⁺	8.6
Ni ²⁺	18.6	Ce ³⁺	16.0	Ba ²⁺	7.8

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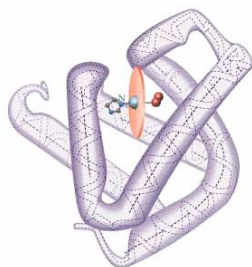
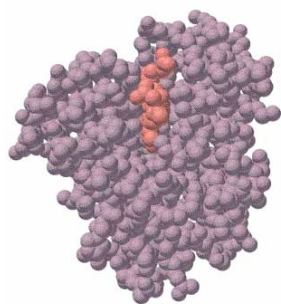
Porphine, an important chelating agent found in nature

Other Chelating Agents



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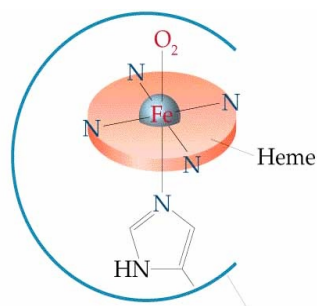
Myoglobin



Myoglobin, an Fe-containing protein that stores O₂ in cells

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Coordination Environment of Fe²⁺ in Oxymyoglobin and Oxymyoglobin



Oxymyoglobin

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Nomenclature of Coordination

The cation is named before the anion. When naming a complex:

- ◆ Cations named first, then anions. Anion metal gets *-ate* ending
- ◆ Ligands are named *first* in alphabetical order. Most ligands have *-o* ending; multiple ligands use Greek prefixes
- ◆ Metal atom/ion is named last with the oxidation state in Roman numerals.
- ◆ Use no spaces in complex name
- ◆ See *Coordination Compounds Handout*

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Nomenclature: IUPAC Rules

Neutral ligands are referred to by their usual name with these exceptions:

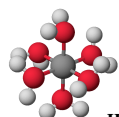
- ▲ water, H₂O = **aqua**
- ▲ ammonia, NH₃ = **ammine**
- ▲ carbon monoxide, CO = **carbonyl**
- ▲ hydroxide, OH⁻ = **hydroxo**

If the ligand name already contains a Greek prefix, use alternate prefixes for multiple occurrences:

- ▲ bis-, 2; tris-, 3; tetrakis-, 4; pentakis-, 5; hexakis-, 6
- ▲ The name of the ligand is placed in parentheses; i.e. *bis(ethylenediamine)*

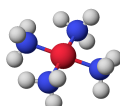
MAR

Nomenclature

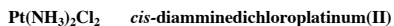
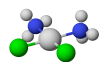


Hexaaquacobalt(II) ion

H₂O as a ligand is **aqua**



Tetraamminecopper(II) ion

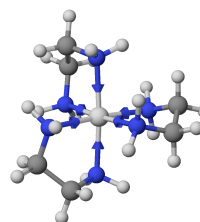


NH₃ as a ligand is **ammine**

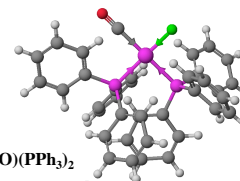
Most metals in this section octahedral, some tetrahedral/square planar (Cu, Pt)

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Nomenclature



Tris(ethylenediamine)nickel(II) ion



IrCl(CO)(PPh₃)₂
Vaska's compound

Carbonylchlorobis(triphenylphosphine)iridium(I)

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Nomenclature: IUPAC Rules



pentaamminechlorocobalt(III) chloride



Potassium tetracyanocadmiate(II)



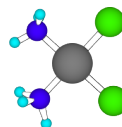
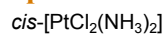
Dichlorobis(ethylenediamine)nickel(II)



Trichlorotris(tetrahydrofuran)chromium(III)

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Cisplatin



square planar Pt(II)
coordination number 4
cis-isomer



the first of a series of platinum coordination complex-based anti-cancer drugs (Platinol-AQ)

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Isomerism

Isomers are compounds that have the same composition but a different arrangement of atoms in space

Important in inorganic *and* organic chemistry

Major Types include:

- ▲ coordination sphere isomers
- ▲ linkage isomers
- ▲ geometric isomers (*cis*, *trans*, *mer*, *fac*)
- ▲ stereoisomers



BRACE YOURSELVES



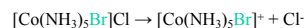
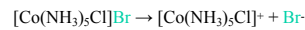
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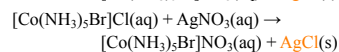
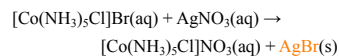
Coordination Sphere Isomers

Coordination sphere isomers differ if a ligand is bonded to the metal in the *complex*, as opposed to being outside the coordination sphere (*ionic bonding*). Different reactivities!

Consider $[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{Br}$ vs. $[\text{Co}(\text{NH}_3)_5\text{Br}]\text{Cl}$



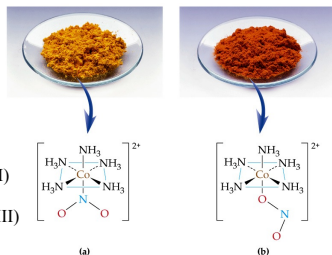
With AgNO_3 :



Linkage isomers differ in which atom of a ligand is bonded to the metal in the complex

Consider $[\text{Co}(\text{NH}_3)_5(\text{ONO})]^{2+}$

and $[\text{Co}(\text{NH}_3)_5(\text{NO}_2)]^{2+}$



a: pentaammine**nitro**cobalt(III)
(yellow color)

b: pentaammine**nitrito**cobalt(III)
(red color)

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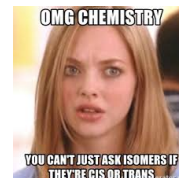
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Geometric Isomers

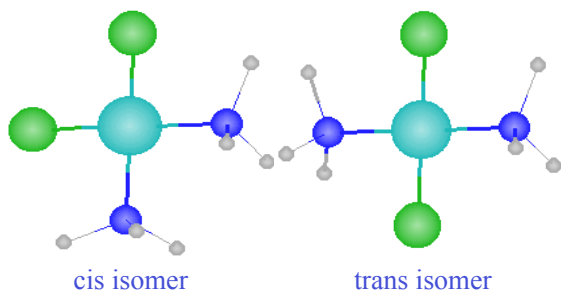
Geometric isomers have same types and number of atoms, but they differ in the spatial arrangements of the ligands

Inorganic complexes exhibit *cis* and *trans* isomerism in square planar and octahedral complexes, as well as *mer* and *fac* isomers in octahedral complexes

Geometric isomers lead to different colors, melting points, boiling points, reactivities, solubilities, etc.



Geometric Isomers

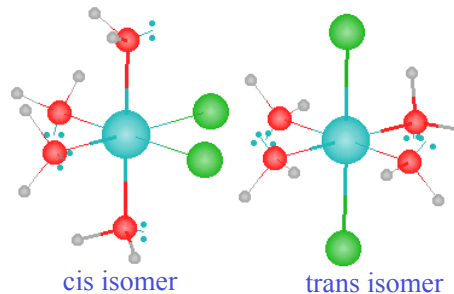


$\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ - square planar

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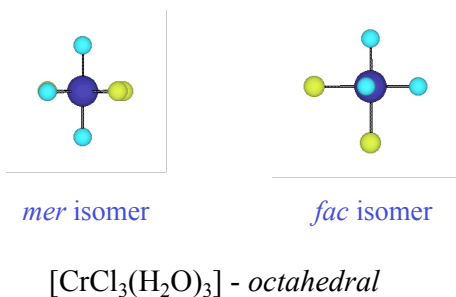
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Geometric Isomers



$[\text{Co}(\text{H}_2\text{O})_4\text{Cl}_2]^+$ - octahedral

Geometric Isomers



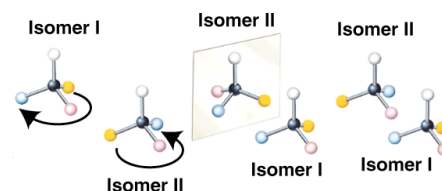
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Optical Isomers

Optical isomers are isomers that are nonsuperimposable mirror images

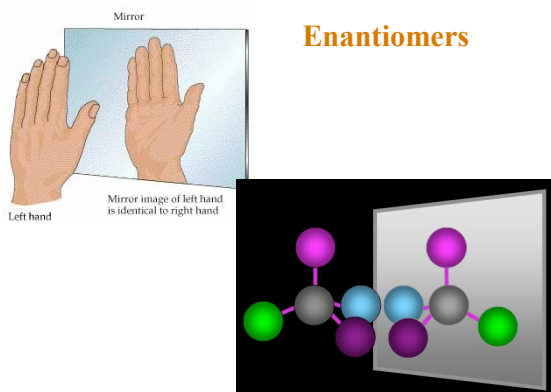
Optical isomers are called **chiral** (*handed*) and referred to as **enantiomers**

A substance is chiral if it does not have a *plane of symmetry*



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Enantiomers



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Properties of Optical Isomers

Optical isomers or enantiomers possess many identical properties

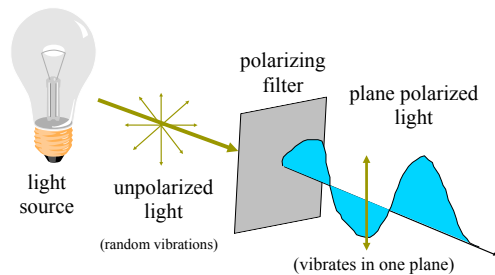
Enantiomers have similar solubilities, melting points, boiling points, colors, chemical reactivity (with nonchiral reagents)

Enantiomers differ in their interactions with plane polarized light and other chiral reagents



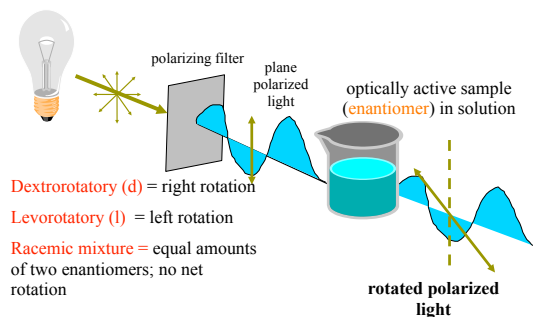
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Optical Isomers



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Optical Isomers

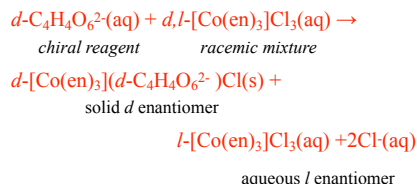


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Properties of Optical Isomers

Differences in enantiomer reactivity with other chiral reagents can lead to separation methods

Example:



Very complex procedures, but important (Ritalin, Thalidomide, etc.)

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Crystal Field Theory

Crystal Field Theory is a model for bonding in transition metal complexes

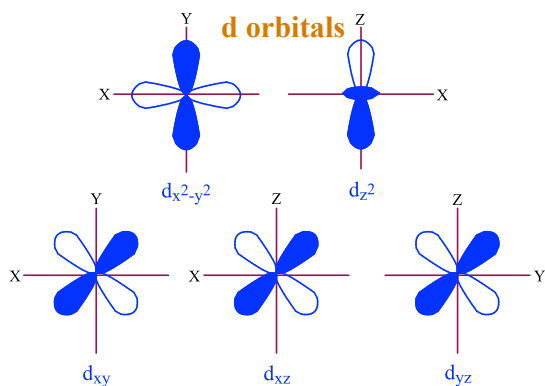
Accounts for color and magnetism of complexes quite well

Focuses on d-orbitals of metals. Ligands are considered to be point negative charges

Assumes ionic bonding (no covalent bonding)

As in organic chemistry, molecular orbital theory is better but much more complex

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Crystal Field Theory

Electrostatic attractions between the positive metal ion and the negative ligands bond them together, lowering the energy of the whole system

The lone pair e⁻'s on ligands repulsed by e⁻'s in metal d orbitals; this interaction is called the crystal field

The repulsion influences d orbital energies, but not all d orbitals influenced the same way

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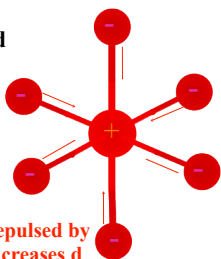
Crystal Field Theory

Octahedral Crystal Field

(-) Ligands attracted to (+) metal ion; provides stability

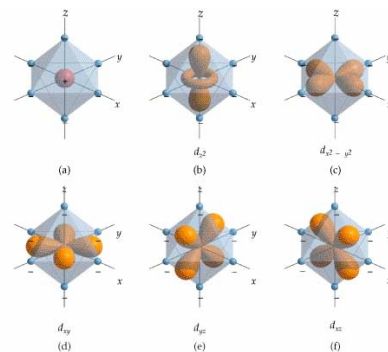
d orbital e⁻'s repulsed by (-) ligands; increases d orbital potential energy

ligands approach along x, y, z axes



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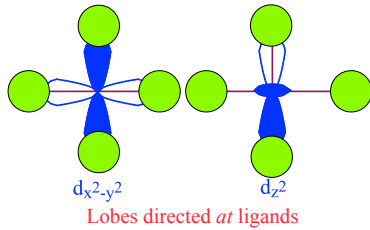
Notice how (b) and (c) exhibit direct overlap between d orbitals and ligands



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d Orbitals and the Ligand Field

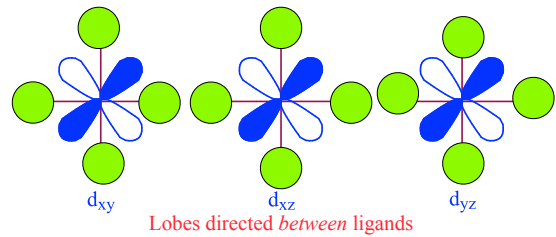
Crystal Field Theory



greater electrostatic repulsion = higher potential energy

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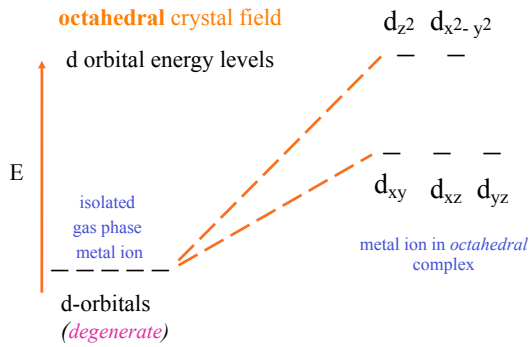
Crystal Field Theory



less electrostatic repulsion = lower potential energy

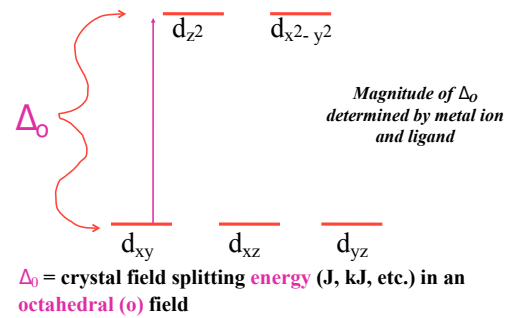
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Crystal Field Theory

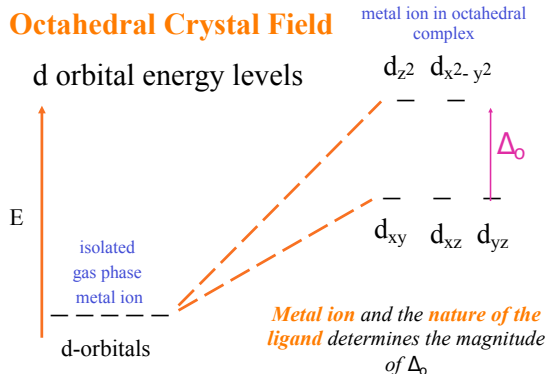


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Crystal Field Splitting Energy



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Crystal Field Theory

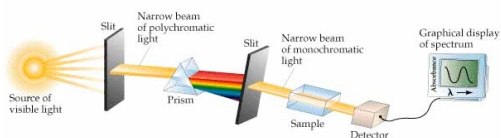
Crystal Field Theory can be used to account for the **colors** of complexes as well as **magnetic properties**

A complex *must* have a partially filled metallic d subshell to exhibit color; i.e. a complex with d^0 or d^{10} is colorless

Amount of paramagnetism depends on ligands; i.e. $[\text{FeF}_6]^{3-}$ has five unpaired electrons while $[\text{Fe}(\text{CN})_6]^{3-}$ has only one

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Visible Spectrum



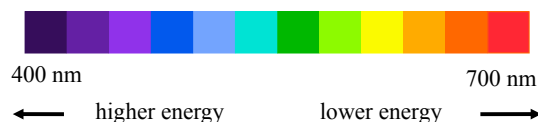
Compounds/complexes with color *absorb* specific wavelengths of visible light (400 –700 nm); wavelengths not absorbed are *transmitted*

MAR

Visible Spectrum

wavelength, nm

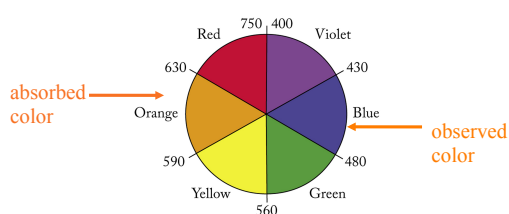
(Each wavelength corresponds to a different color)



White = all the colors (wavelengths)

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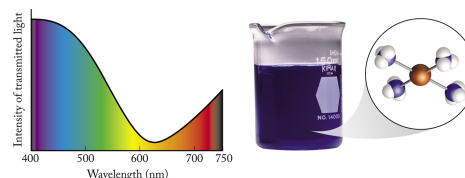
More on Colors



The color observed in a compound/complex is actually the *complementary* color of the color absorbed

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More on Colors



absorbs yellow, orange, and red...

...so observed color of complex is navy blue.

The color observed in a compound/complex is actually the *complementary* color of the color absorbed

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Colors of Transition Metal Complexes

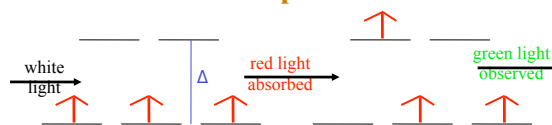
Absorption of UV-visible radiation by compound or complex occurs *only* if the radiation has the energy needed to raise an e^- from its ground state (highest occupied orbital or HOMO) to an excited state (usually the lowest unoccupied orbital or LUMO)

The light energy *absorbed* = energy difference between the ground state and excited state, or

$$\Delta E = E_{\text{LUMO}} - E_{\text{HOMO}} = h\nu = hc/\lambda_{\text{max}}$$

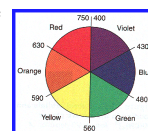
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Colors of Transition Metal Complexes



For transition metal complexes, Δ corresponds to energies of visible light.

Absorption raises an electron from the lower d subshell to the higher d subshell.



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Colors of Transition Metal Complexes

The **magnitude** of Δ depends on the **ligand field** (i.e. types of ligands) and the **metal**, and Δ determines the **color** of the complex

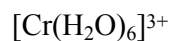
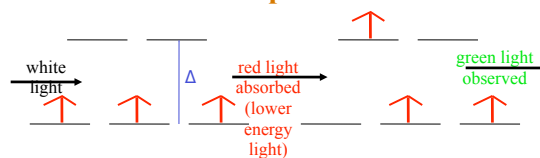
Different complexes exhibit different colors due to the magnitude of Δ

larger Δ = higher energy light absorbed, shorter wavelengths

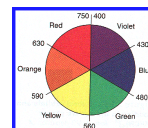
smaller Δ = lower energy light absorbed, longer wavelengths

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Colors of Transition Metal Complexes

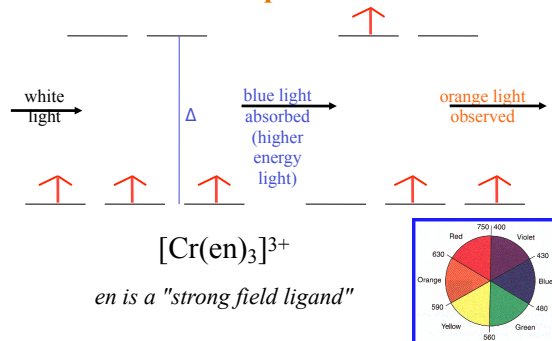


H_2O is a "weak field ligand"



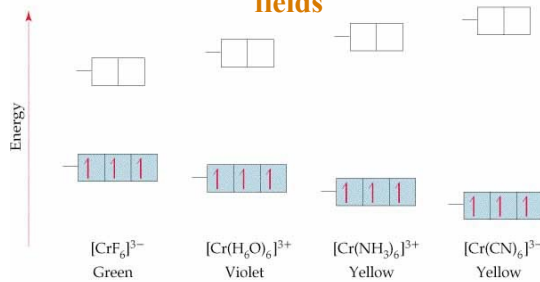
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Colors of Transition Metal Complexes



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Chromium(III) with different ligand fields

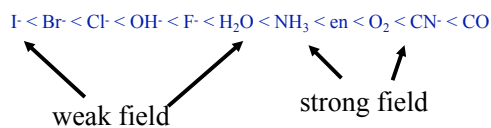


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Colors of Transition Metal Complexes

The Spectrochemical Series:

Smallest Δ $\xrightarrow{\Delta \text{ increases}}$ Largest Δ

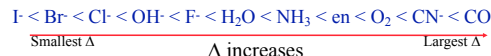
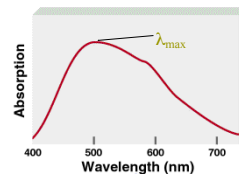


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Spectrochemical Series

Examples with chromium(III):

Complex	λ_{max} (nm)	Notes
CrCl_6^{3-}	736	largest λ_{max} , smallest Δ
$\text{Cr}(\text{H}_2\text{O})_6^{3+}$	573	
$\text{Cr}(\text{NH}_3)_6^{3+}$	462	
$\text{Cr}(\text{CN})_6^{3-}$	380	smallest λ_{max} , largest Δ



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Paramagnetism of Transition Metal Complexes

Many complexes / compounds are **paramagnetic** due to unpaired d electrons

The **degree** of paramagnetism dependent on ligands ("ligand field")

Example with Fe^{3+} :

$[Fe(CN)_6]^{3-}$ has **1** unpaired d electron

$[FeF_6]^{3-}$ has **5** unpaired d electrons

Crystal Field Theory answers this discrepancy

MAR

Electronic Configurations of Transition Metal Complexes

Filling electron shells via CH 221:

- ◆ lowest energy vacant orbitals are occupied first
- ◆ electrons fill degenerate orbitals singly until no longer possible (**Hund's rule**), then pair (**Pauli Exclusion**)

These rules help minimize repulsions between electrons.

These rules work well for *gas-phase transition metal ions*, but they are not always followed by *transition metal complexes in a ligand field*

MAR

Electronic Configurations of Transition Metal Complexes

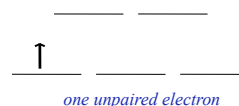
For complexes in a ligand field, d orbital occupancy depends on Δ and **pairing energy, P**

- ◆ Electrons assume the configuration with the lowest possible energy "cost"
- ◆ If $\Delta > P$ (Δ large; **strong field** ligand), e-'s pair up in **lower energy d subshell first**, referred to as a **low spin** complex
- ◆ If $\Delta < P$ (Δ small; **weak field** ligand), e-'s **spread out among all d orbitals** before any pair up, referred to as a **high spin** complex

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d-orbital energy level diagrams octahedral complex

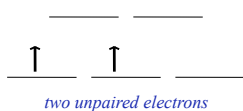
d^1



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d-orbital energy level diagrams octahedral complex

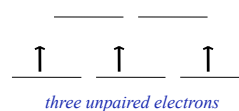
d^2



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d-orbital energy level diagrams octahedral complex

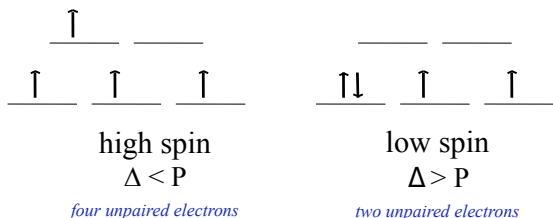
d^3



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**d-orbital energy level diagrams
octahedral complex**

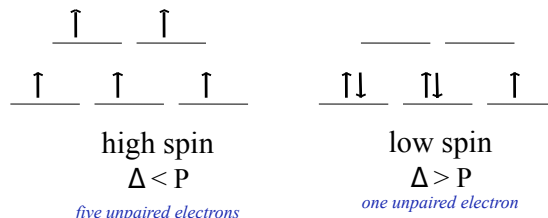
d^4



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**d-orbital energy level diagrams
octahedral complex**

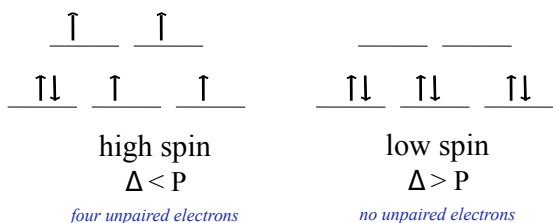
d^5



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**d-orbital energy level diagrams
octahedral complex**

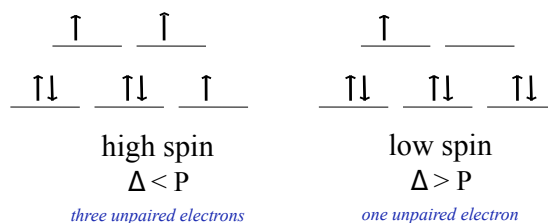
d^6



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**d-orbital energy level diagrams
octahedral complex**

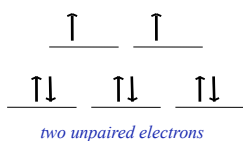
d^7



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**d-orbital energy level diagrams
octahedral complex**

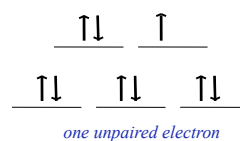
d^8



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**d-orbital energy level diagrams
octahedral complex**

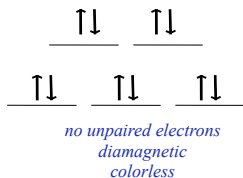
d^9



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d-orbital energy level diagrams octahedral complex

d^{10}



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Electronic Configurations of Transition Metal Complexes

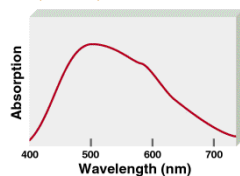
To determine which d-orbital energy level diagram to use on a complex or compound:

- ▲ determine the oxidation # of the metal
- ▲ determine the # of d e⁻'s
- ▲ determine if ligand is weak field or strong field
- ▲ draw energy level diagram

MAR

Example with $\text{Co}(\text{NH}_3)_6^{3+}$

The absorbance spectrum of $\text{Co}(\text{NH}_3)_6^{3+}$ is shown to the right.



1. What is the **name** of the complex?
2. What is the **coordination number**?
3. How many **d electrons** does the complex possess?
4. Will this be a **high spin** or **low spin** complex?
5. Is the compound **paramagnetic**?
6. What is the **value of Δ** for this complex?
7. What will be the **observed color** of the complex? What **color is absorbed**?

MAR

Example with $\text{Co}(\text{NH}_3)_6^{3+}$

1. What is the **name** of the complex?
Each ammonia is neutral; hence, this is a **cobalt(III)** complex.
Ammonia is named "ammine"; hence,
Name = **hexaamminecobalt(III) ion**
2. What is the **coordination number**?

There are six monodentate ammonia ligands around the cobalt(III) ion; hence, this is a **six coordinate compound**, or the **coordination number = 6**

MAR

Example with $\text{Co}(\text{NH}_3)_6^{3+}$

3. How many **d electrons** does the complex possess?

Cobalt(III) ions have the electron configuration:

$[\text{Ar}]3d^6$ (two 4s and 1 3d electron removed)

This complex has **six d electrons (d^6)**

4. Will this be a **high spin** or **low spin** complex?

To answer this, we need to look at the **spectrochemical series**.

NH_3 is considered a **strong field ligand**, which implies that Δ , the splitting energy, will be greater than the electron pairing energy (P).

The **octahedral NH_3 ligand field** makes this a **low spin complex**

MAR

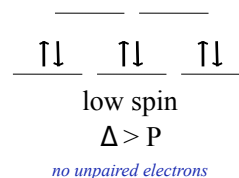
Example with $\text{Co}(\text{NH}_3)_6^{3+}$

5. Is the compound **paramagnetic**?

This is a **d^6 low spin complex** in an **octahedral field**.

Three degenerate d orbitals are filled first, followed by the remaining two orbitals.

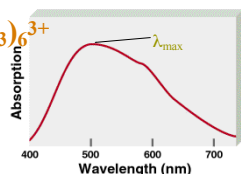
Since each orbital holds two electrons, the three lower orbitals are full and the complex is **diamagnetic**



MAR

Example with $\text{Co}(\text{NH}_3)_6^{3+}$

The absorbance spectrum of $\text{Co}(\text{NH}_3)_6^{3+}$ is shown to the right.



6. What is the value of Δ for this complex?

Notice that $\lambda_{\text{max}} = 500 \text{ nm}$ in the spectrum, and technically this is a question asking for Δ since the complex is in an octahedral field

$$\Delta_o = \Delta E = h\nu = hc/\lambda_{\text{max}}$$

$$\Delta_o = 6.626 \times 10^{-34} \text{ J s} \times 2.998 \times 10^8 \text{ m s}^{-1} / 500. \times 10^{-9} \text{ m}$$

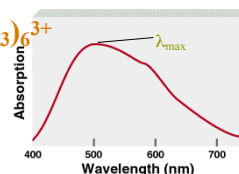
$$\Delta_o = 3.97 \times 10^{-19} \text{ J or}$$

$$\Delta_o = 239 \text{ kJ/mol}$$

MAR

Example with $\text{Co}(\text{NH}_3)_6^{3+}$

The absorbance spectrum of $\text{Co}(\text{NH}_3)_6^{3+}$ is shown to the right.

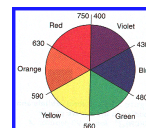


7. What will be the observed color of the complex? What color is absorbed?

Since λ_{max} is 500. nm in the spectrum, this is the wavelength that is absorbed.

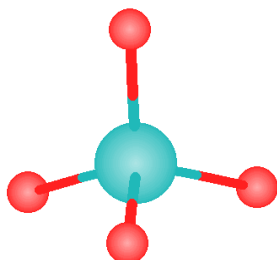
Using 500. nm, the chart to the right implies that the complex is absorbing the color green

The complementary color is the color that is actually observed by our eyes. Red is opposite green, so the observed color should be red



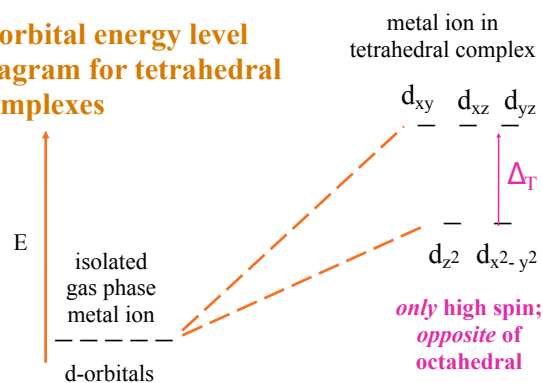
MAR

d-orbital energy level diagrams for tetrahedral complexes



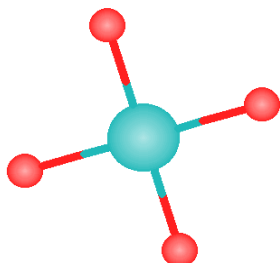
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d-orbital energy level diagram for tetrahedral complexes



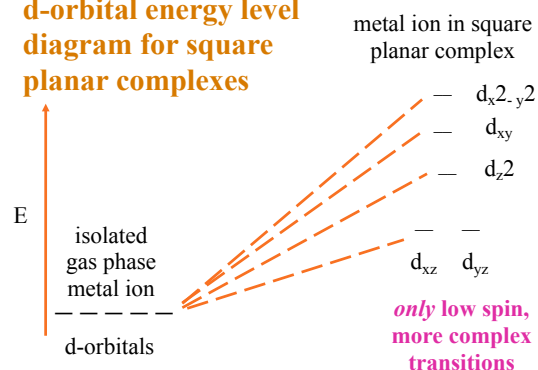
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d-orbital energy level diagrams for square planar complexes

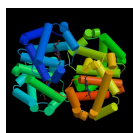


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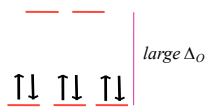
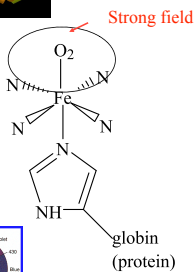
d-orbital energy level diagram for square planar complexes



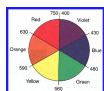
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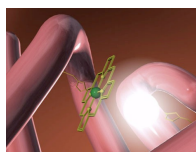
Example: Arterial Blood



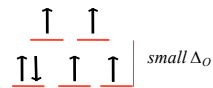
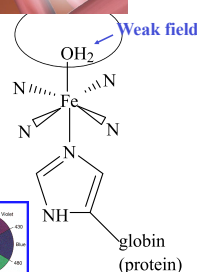
d⁶, low spin
Bright red due to absorption of greenish light



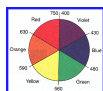
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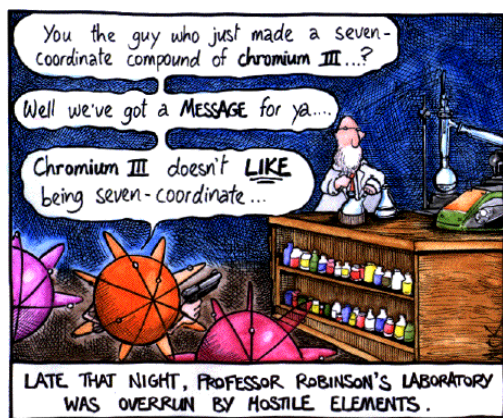
Example: Venous Blood



d⁶, high spin
Bluish color due to absorption of orange light



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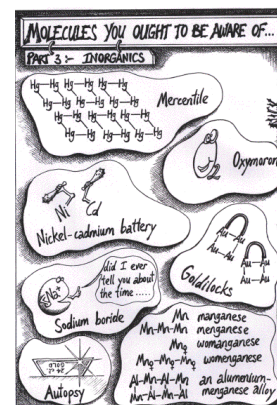


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End of Chapter 19

- See:
- * [Chapter Nineteen Study Guide](#)
 - * [Important Equations \(following this slide\)](#)
 - * [End of Chapter Problems \(following this slide\)](#)

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Important Equations, Constants, and Handouts from this Chapter:

- The primary focus of this chapter is to introduce you to coordination compound nomenclature (with a little review of metals, electron configurations, etc. as well from CH 221 and CH 222)

- Handouts:
- [Coordination Compounds Handout](#)

MAR

End of Chapter Problems: Test Yourself

1. Give the electron configuration for the Cr^{3+} ion. Is it paramagnetic or diamagnetic?
2. Which of the following ligands is expected to be monodentate and which might be polydentate? a. CH_3NH_2 b. CH_3CN c. en d. Br^- e. phen
3. Give the oxidation number of the metal ion in $[\text{Mn}(\text{NH}_3)_6]\text{SO}_4$
4. Give the oxidation number of the metal ion in $\text{Cr}(\text{en})_2\text{Cl}_2$
5. Write the formula for potassium tetrachloroplatinate(II)
6. Write the formula for tetraamminediaquairon(II)
7. Name the following: $[\text{Ni}(\text{C}_2\text{O}_4)_2(\text{H}_2\text{O})_2]^{2-}$
8. Name the following: $\text{Pt}(\text{NH}_3)_2(\text{C}_2\text{O}_4)$

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End of Chapter Problems: Answers

1. $[\text{Ar}]3d^3$, paramagnetic
2. a, b, d: monodentate c, e: polydentate (bidentate)
3. +2
4. +2
5. $\text{K}_2[\text{PtCl}_4]$
6. $[\text{Fe}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$
7. diaquabis(oxalato)nickelate(II) ion
8. diammineoxalatoplatinum(II)

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