CHEM 212
Coordination Chemistry

Chapter 24

Why Study Descriptive Chemistry of Transition Metals

- Transition metals are found in nature
  - Rocks and minerals contain transition metals
  - The color of many gemstones is due to the presence of transition metal ions
    - Rubies are red due to Cr
    - Sapphires are blue due to presence of Fe and Ti
  - Many biomolecules contain transition metals that are involved in the functions of these biomolecules
    - Vitamin B12 contains Co
    - Hemoglobin, myoglobin, and cytochrome C contain Fe

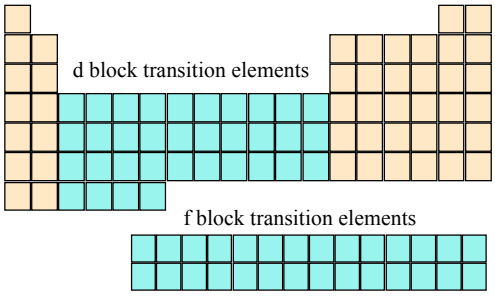
Why Study Descriptive Chemistry of Transition Metals

- Transition metals and their compounds have many useful applications
  - Fe is used to make steel and stainless steel
  - Ti is used to make lightweight alloys
  - Transition metal compounds are used as pigments
    - TiO_2 = white
    - $PbCrO_4$ = yellow
    - $Fe_4[Fe(CN)_6]_3$ (prussian blue)= blue
  - Transition metal compounds are used in many industrial processes

Why Study Descriptive Chemistry of Transition Metals

- To understand the uses and applications of transition metals and their compounds, we need to understand their chemistry.
- Our focus will be on the 4th period transition elements.

Periodic Table



Transition Metals

- General Properties
  - Have typical metallic properties
  - Not as reactive as Grp. IA, IIA metals
  - Have high MP's, high BP's, high density, and are hard and strong
  - Have 1 or 2 s electrons in valence shell
  - Differ in # d electrons in n-1 energy level
  - Exhibit multiple oxidation states

d-Block Transition Elements

IIIB	IVB	VB	VIB	VIIB	VIIIB			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

Electronic Configurations

Element	Configuration
Sc	$[Ar]3d^14s^2$
Ti	$[Ar]3d^24s^2$
V	$[Ar]3d^34s^2$
Cr	$[Ar]3d^54s^1$
Mn	$[Ar]3d^54s^2$

$[Ar] = 1s^22s^22p^63s^23p^6$

Electronic Configurations

Element	Configuration
Fe	$[Ar]\ 3d^64s^2$
Co	$[Ar]\ 3d^74s^2$
Ni	$[Ar]\ 3d^84s^2$
Cu	$[Ar]3d^{10}4s^1$
Zn	$[Ar]3d^{10}4s^2$

$[Ar] = 1s^22s^22p^63s^23p^6$

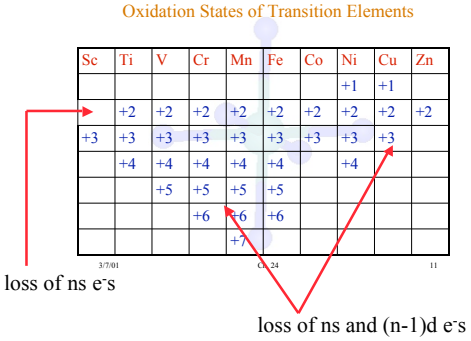
Transition Metals

- Characteristics due to d electrons:
  - Exhibit multiple oxidation states
  - Compounds typically have color
  - Exhibit interesting magnetic properties
    - paramagnetism
    - ferromagnetism

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					

Oxidation States of Transition Elements



Electronic Configurations of Transition Metal Ions

- Electronic configuration of Fe^{2+}

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$$Fe - 2e^- \rightarrow Fe^{2+}$$

Electronic Configurations of Transition Metal Ions

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$$Fe - 2e^- \rightarrow Fe^{2+}$$
$$[Ar]3d^64s^2$$

valence ns e⁻'s removed first

Electronic Configurations of Transition Metal Ions

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Electronic Configurations of Transition Metal Ions

- Electronic configuration of Fe^{3+}

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Electronic Configurations of Transition Metal Ions

- Electronic configuration of Fe^{3+}

$$Fe - 3e^- \rightarrow Fe^{3+}$$
$[Ar]3d^64s^2$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Fe^{3+}

$$Fe - 3e^- \rightarrow Fe^{3+}$$
$[Ar]3d^64s^2$ $[Ar]3d^5$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Co^{3+}

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Co^{3+}

$$Co - 3e^- \rightarrow Co^{3+}$$

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Co^{3+}

$$Co - 3e^- \rightarrow Co^{3+}$$
$[Ar]3d^74s^2$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Co^{3+}

$$Co - 3e^- \rightarrow Co^{3+}$$
$[Ar]3d^74s^2$ $[Ar]3d^6$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Mn^{4+}

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Mn^{4+}

$$Mn - 4e^- \rightarrow Mn^{4+}$$

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Mn^{4+}

$$Mn - 4e^- \rightarrow Mn^{4+}$$
$[Ar]3d^5 4s^2$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Electronic Configurations of Transition Metal Ions

- Electronic configuration of Mn^{4+}

$$Mn - 4e^- \rightarrow Mn^{4+}$$
$[Ar]3d^5 4s^2$ $[Ar]3d^3$

valence ns e⁻'s removed first, then n-1 d e⁻'s

Coordination Chemistry

- Transition metals act as Lewis acids
  - Form complexes/complex ions

$Fe^{3+}(aq) + 6CN^-(aq) \rightarrow Fe(CN)_6^{3-}(aq)$
Lewis acid Lewis base Complex ion

$Ni^{2+}(aq) + 6NH_3(aq) \rightarrow Ni(NH_3)_6^{2+}(aq)$
Lewis acid Lewis base Complex ion

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

Coordination Chemistry

- Transition metals act as Lewis acids
  - Form complexes/complex ions

$Fe^{3+}(aq) + 6CN^-(aq) \rightarrow [Fe(CN)_6]^{3-}(aq)$
Lewis acid Lewis base Complex ion

$Ni^{2+}(aq) + 6NH_3(aq) \rightarrow [Ni(NH_3)_6]^{2+}(aq)$
Lewis acid Lewis base Complex ion

Complex with a net charge = complex ion

Complexes have distinct properties

Coordination Chemistry

- Coordination compound
  - Compound that contains 1 or more complexes
  - Example
    - $[Co(NH_3)_6]Cl_3$
    - $[Cu(NH_3)_4][PtCl_4]$
    - $[Pt(NH_3)_2Cl_2]$

Coordination Chemistry

- Coordination sphere
  - Metal and ligands bound to it
- Coordination number
  - number of donor atoms bonded to the central metal atom or ion in the complex
    - Most common = 4, 6
    - Determined by ligands
      - Larger ligands and those that transfer substantial negative charge to metal favor lower coordination numbers

Coordination Chemistry

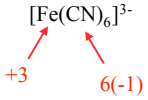
Complex charge = sum of charges on the metal and the ligands

$[Fe(CN)_6]^{3-}$

Coordination Chemistry

Complex charge = sum of charges on the metal and the ligands

$[Fe(CN)_6]^{3-}$



Coordination Chemistry

Neutral charge of coordination compound = sum of charges on metal, ligands, and counterbalancing ions

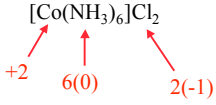
$[Co(NH_3)_6]Cl_2$

neutral compound

Coordination Chemistry

Neutral charge of coordination compound = sum of charges on metal, ligands, and counterbalancing ions

$[Co(NH_3)_6]Cl_2$



Coordination Chemistry

- Ligands
  - classified according to the number of donor atoms
  - Examples
    - monodentate = 1
    - bidentate = 2
    - tetradentate = 4
    - hexadentate = 6
    - polydentate = 2 or more donor atoms

Coordination Chemistry

- Ligands
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    - polydentate = 2 or more donor atoms

chelating agents

Ligands

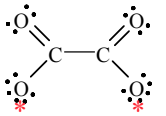
- Monodentate
  - Examples:
    - H_2O, CN^-, NH_3, NO_2^-, SCN^-, OH^-, X^- (halides), CO, O^{2-}
  - Example Complexes
    - $[Co(NH_3)_6]^{3+}$
    - $[Fe(SCN)_6]^{3-}$

Ligands

- Bidentate
  - Examples
    - oxalate ion = $C_2O_4^{2-}$
    - ethylenediamine (en) = $NH_2CH_2CH_2NH_2$
    - ortho-phenanthroline (o-phen)
  - Example Complexes
    - $[Co(en)_3]^{3+}$
    - $[Cr(C_2O_4)_3]^{3-}$
    - $[Fe(NH_3)_4(o\text{-}phen)]^{3+}$

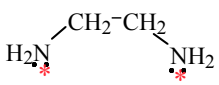
Ligands

oxalate ion

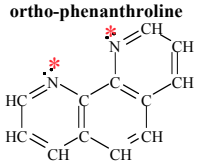


2-

ethylenediamine



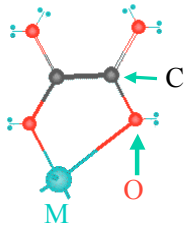
ortho-phenanthroline



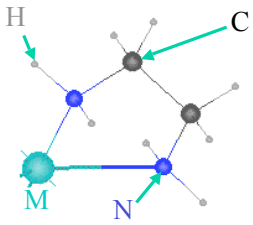
Donor Atoms

Ligands

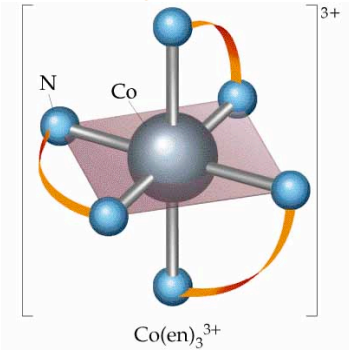
oxalate ion



ethylenediamine



Ligands



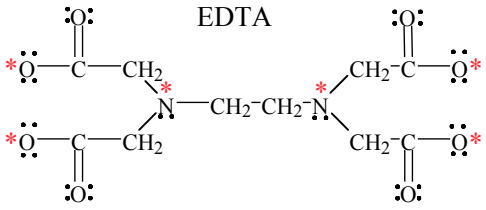
$Co(en)_3^{3+}$

Ligands

- Hexadentate
  - ethylenediaminetetraacetate (EDTA) = $(O_2CCH_2)_2N(CH_2)_2N(CH_2CO_2)_2^{4-}$
  - Example Complexes
    - $[Fe(EDTA)]^{-1}$
    - $[Co(EDTA)]^{-1}$

Ligands

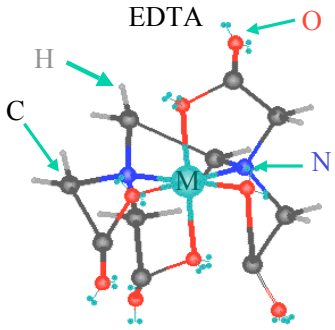
EDTA



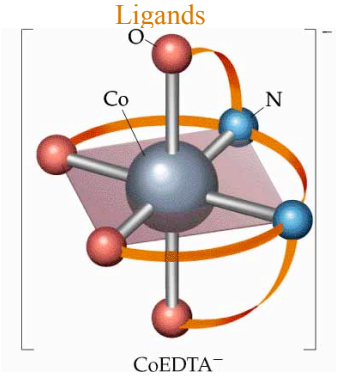
Donor Atoms

Ligands

EDTA

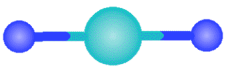


Ligands



$CoEDTA^-$

Common Geometries of Complexes

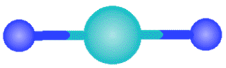
Coordination Number	Geometry
2	

Linear

Common Geometries of Complexes

Coordination Number: 2

Geometry: Linear

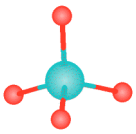


Example: $[Ag(NH_3)_2]^+$

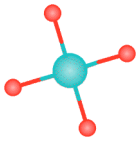
Common Geometries of Complexes

Coordination Number: 4

Geometry: tetrahedral (most common)



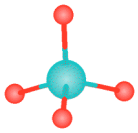
square planar (characteristic of metal ions with 8 d e⁻'s)



Common Geometries of Complexes

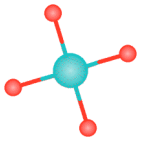
Coordination Number: 4

Geometry: tetrahedral



Examples: $[Zn(NH_3)_4]^{2+}$, $[FeCl_4]^-$

square planar

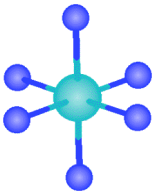


Example: $[Ni(CN)_4]^{2-}$

Common Geometries of Complexes

Coordination Number: 6

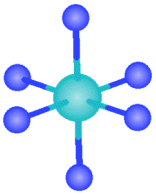
Geometry: octahedral



Common Geometries of Complexes

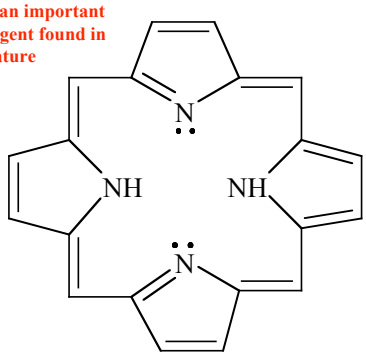
Coordination Number: 6

Geometry: octahedral

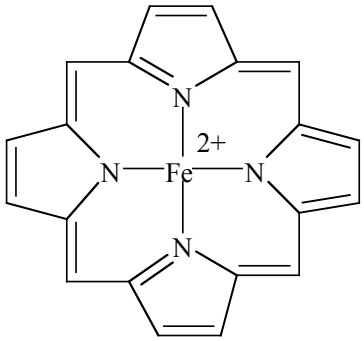


Examples: $[Co(CN)_6]^{3-}$, $[Fe(en)_3]^{3+}$

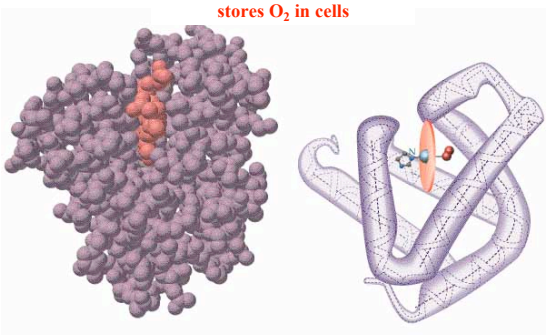
Porphine, an important chelating agent found in nature



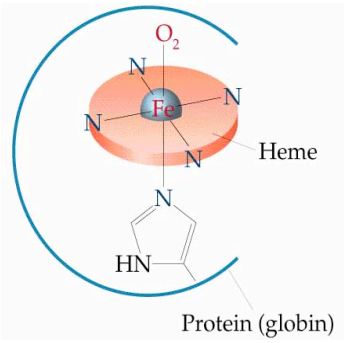
Metalloporphyrin



Myoglobin, a protein that stores O_2 in cells



Coordination Environment of Fe^{2+} in Oxymyoglobin and Oxyhemoglobin



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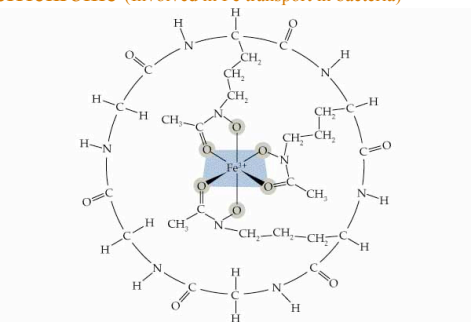
Ferrichrome (Involved in Fe transport in bacteria)



Nomenclature of Coordination Compounds: IUPAC Rules

- The cation is named before the anion
- When naming a complex:
  - Ligands are named first
    - alphabetical order
  - Metal atom/ion is named last
    - oxidation state given in Roman numerals follows in parentheses
  - Use no spaces in complex name

Nomenclature: IUPAC Rules

- The names of anionic ligands end with the suffix -o
  - -ide suffix changed to -o
  - -ite suffix changed to -ito
  - -ate suffix changed to -ato

Nomenclature: IUPAC Rules

Ligand	Name
bromide, Br^-	bromo
chloride, Cl^-	chloro
cyanide, CN^-	cyano
hydroxide, OH^-	hydroxo
oxide, O^{2-}	oxo
fluoride, F^-	fluoro

Nomenclature: IUPAC Rules

Ligand	Name
carbonate, CO_3^{2-}	carbonato
oxalate, $C_2O_4^{2-}$	oxalato
sulfate, SO_4^{2-}	sulfato
thiocyanate, SCN^-	thiocyanato
thiosulfate, $S_2O_3^{2-}$	thiosulfato
Sulfite, SO_3^{2-}	sulfito

Nomenclature: IUPAC Rules

- Neutral ligands are referred to by the usual name for the molecule
  - Example
    - ethylenediamine
  - Exceptions
    - water, H_2O = aqua
    - ammonia, NH_3 = ammine
    - carbon monoxide, CO = carbonyl

Nomenclature: IUPAC Rules

- Greek prefixes are used to indicate the number of each type of ligand when more than one is present in the complex
  - di-, 2; tri-, 3; tetra-, 4; penta-, 5; hexa-, 6
- If the ligand name already contains a Greek prefix, use alternate prefixes:
  - bis-, 2; tris-, 3; tetrakis-,4; pentakis-, 5; hexakis-, 6
  - The name of the ligand is placed in parentheses

Nomenclature: IUPAC Rules

- If a complex is an anion, its name ends with the -ate
  - appended to name of the metal

Nomenclature: IUPAC Rules

Transition Metal	Name if in Cationic Complex	Name if in Anionic Complex
Sc	**Scandium**	**Scandate**
Ti	**titanium**	**titanate**
V	**vanadium**	**vanadate**
Cr	**chromium**	**chromate**
Mn	**manganese**	**manganate**
Fe	**iron**	**ferrate**
Co	**cobalt**	**cobaltate**
Ni	**nickel**	**nickelate**
Cu	**Copper**	**cuprate**
Zn	**Zinc**	**zincate**

Isomerism

- Isomers
  - compounds that have the same composition but a different arrangement of atoms
- Major Types
  - structural isomers
  - stereoisomers

Structural Isomers

- Structural Isomers
  - isomers that have different bonds

Structural Isomers

- Coordination-sphere isomers
  - differ in a ligand bonded to the metal in the complex, as opposed to being outside the coordination-sphere

Coordination-Sphere Isomers

- Example

  $[Co(NH_3)_5Cl]Br$ vs. $[Co(NH_3)_5Br]Cl$

Coordination-Sphere Isomers

- Example

  $[Co(NH_3)_5Cl]Br$ vs. $[Co(NH_3)_5Br]Cl$
- Consider ionization in water

$[Co(NH_3)_5Cl]Br \rightarrow [Co(NH_3)_5Cl]^+ + Br^-$

$[Co(NH_3)_5Br]Cl \rightarrow [Co(NH_3)_5Br]^+ + Cl^-$

Coordination-Sphere Isomers

- Example

  $[Co(NH_3)_5Cl]Br$ vs. $[Co(NH_3)_5Br]Cl$
- Consider precipitation

$[Co(NH_3)_5Cl]Br(aq) + AgNO_3(aq) \rightarrow [Co(NH_3)_5Cl]NO_3(aq) + AgBr(s)$

$[Co(NH_3)_5Br]Cl(aq) + AgNO_3(aq) \rightarrow [Co(NH_3)_5Br]NO_3(aq) + AgCl(aq)$

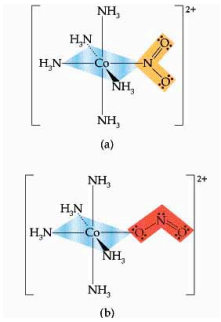
Structural Isomers

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

Linkage Isomers

- Example
  - $[Co(NH_3)_5(ONO)]^{2+}$ vs. $[Co(NH_3)_5(NO_2)]^{2+}$

Linkage Isomers



Linkage Isomers

- Example
  - $[Co(NH_3)_5(SCN)]^{2+}$ vs. $[Co(NH_3)_5(NCS)]^{2+}$
    - Co-SCN vs. Co-NCS

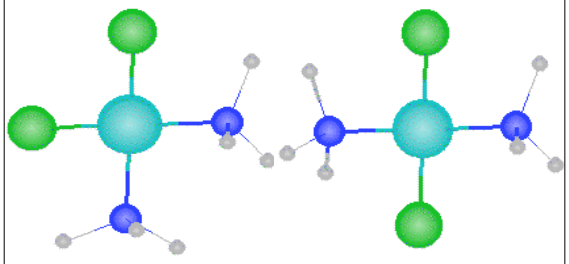
Stereoisomers

- Stereoisomers
  - Isomers that have the same bonds, but different spatial arrangements

Stereoisomers

- Geometric isomers
  - Differ in the spatial arrangements of the ligands

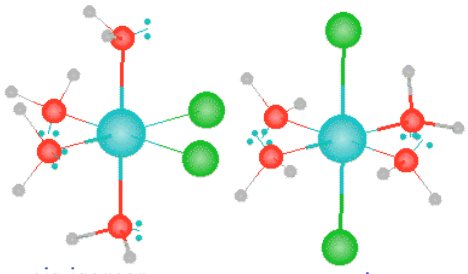
Geometric Isomers



cis isomer trans isomer

$Pt(NH_3)_2Cl_2$

Geometric Isomers



cis isomer trans isomer

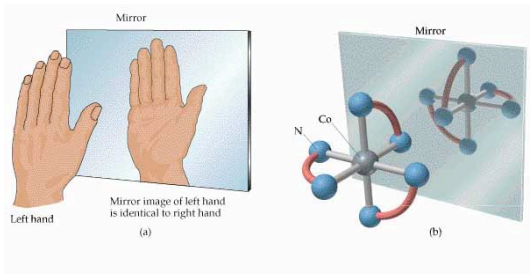
$[Co(H_2O)_4Cl_2]^+$

Stereoisomers

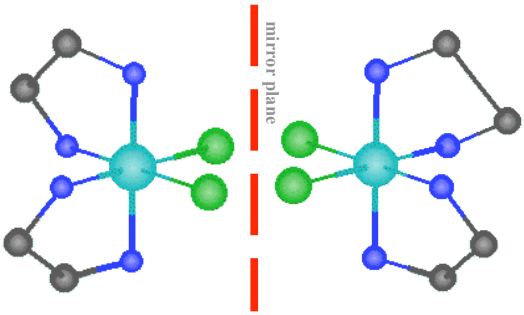
- Geometric isomers
  - Differ in the spatial arrangements of the ligands
  - Have different chemical/physical properties
    - different colors, melting points, polarities, solubilities, reactivities, etc.

Stereoisomers

- Optical isomers
  - isomers that are nonsuperimposable mirror images
    - said to be "chiral" (handed)
    - referred to as enantiomers
  - A substance is "chiral" if it does not have a "plane of symmetry"



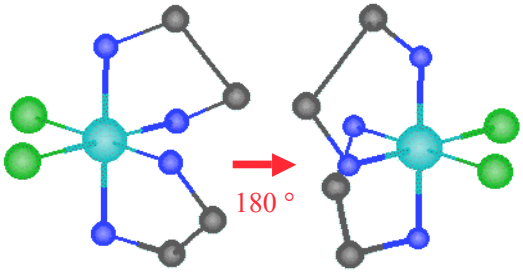
Example 1



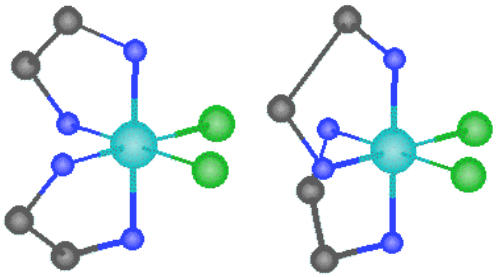
cis-$[Co(en)_2Cl_2]^+$

Example 1

rotate mirror image 180°

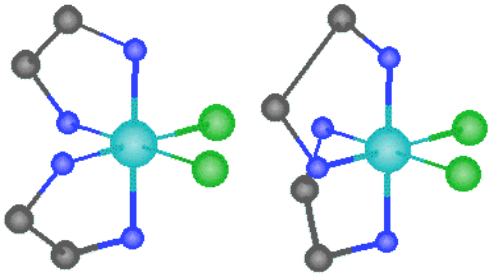


Example 1 nonsuperimposable



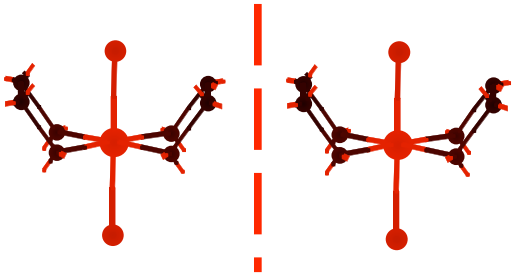
cis-$[Co(en)_2Cl_2]^+$

Example 1 enantiomers



cis-$[Co(en)_2Cl_2]^+$

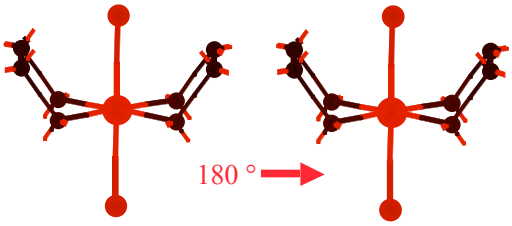
Example 2



trans-$[Co(en)_2Cl_2]^+$

Example 2

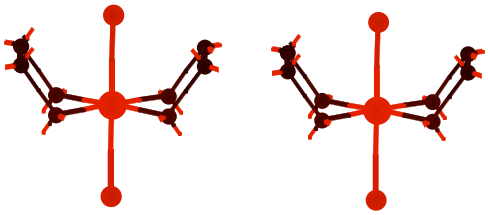
rotate mirror image 180°



trans-$[Co(en)_2Cl_2]^+$

Example 2

Superimposable-not enantiomers

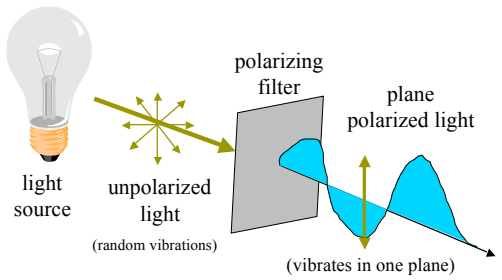


trans-$[Co(en)_2Cl_2]^+$

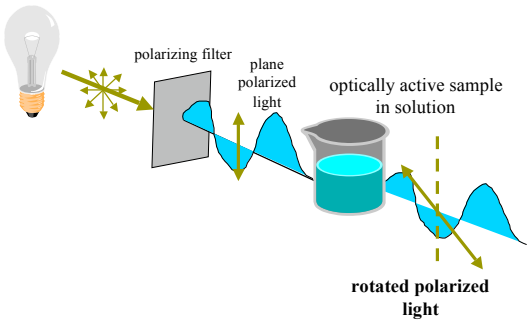
Properties of Optical Isomers

- Enantiomers
  - possess many identical properties
    - solubility, melting point, boiling point, color, chemical reactivity (with nonchiral reagents)
  - different in:
    - interactions with plane polarized light

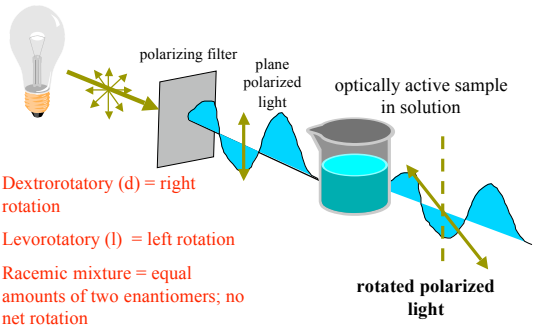
Optical Isomers



Optical Isomers



Optical Isomers



Dextrorotatory (d) = right rotation

Levorotatory (l) = left rotation

Racemic mixture = equal amounts of two enantiomers; no net rotation

Properties of Optical Isomers

- Enantiomers
  - possess many identical properties
    - solubility, melting point, boiling point, color, chemical reactivity (with nonchiral reagents)
  - different in:
    - interactions with plane polarized light
    - reactivity with "chiral" reagents

Example

$d\text{-}C_4H_4O_6^{2-}(aq) + d,l\text{-}[Co(en)_3]Cl_3(aq) \rightarrow$

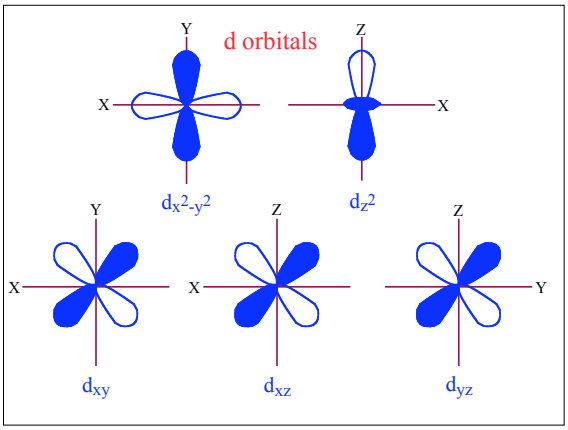
$d\text{-}[Co(en)_3](d\text{-}C_4H_4O_6^{2-})Cl(s) + l\text{-}[Co(en)_3]Cl_3(aq) + 2Cl^-(aq)$

Properties of Transition Metal Complexes

- Properties of transition metal complexes:
  - usually have color
    - dependent upon ligand(s) and metal ion
  - many are paramagnetic
    - due to unpaired d electrons
    - degree of paramagnetism dependent on ligand(s)
      - $[Fe(CN)_6]^{3-}$ has 1 unpaired d electron
      - $[FeF_6]^{3-}$ has 5 unpaired d electrons

Crystal Field Theory

- Crystal Field Theory
  - Model for bonding in transition metal complexes
    - Accounts for observed properties of transition metal complexes
  - Focuses on d-orbitals
  - Ligands = point negative charges
  - Assumes ionic bonding
    - electrostatic interactions

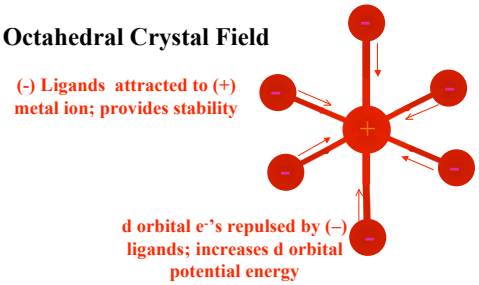


Crystal Field Theory

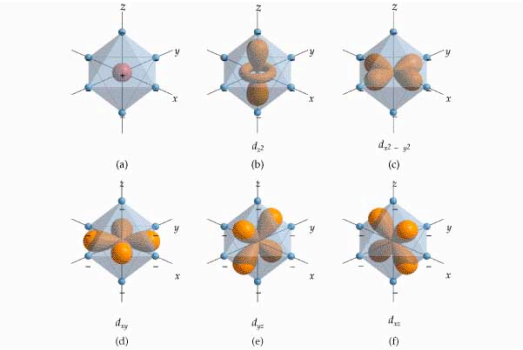
- Electrostatic Interactions
  - (+) metal ion attracted to (-) ligands (anion or dipole)
    - provides stability
  - lone pair e⁻'s on ligands repulsed by e⁻'s in metal d orbitals
    - interaction called crystal field
    - influences d orbital energies
      - not all d orbitals influenced the same way

Crystal Field Theory

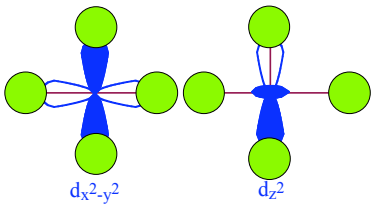
Octahedral Crystal Field



ligands approach along x, y, z axes



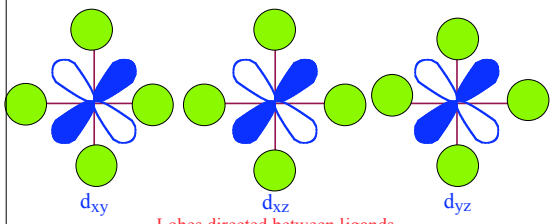
Crystal Field Theory



Lobes directed at ligands

greater electrostatic repulsion = higher potential energy

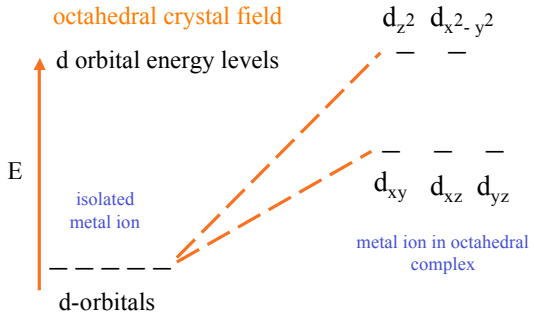
Crystal Field Theory



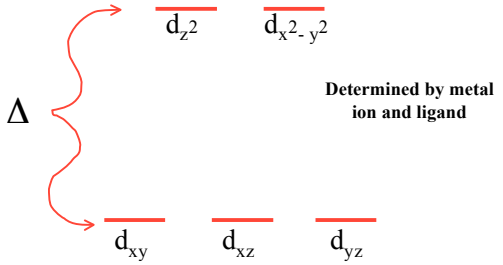
Lobes directed between ligands

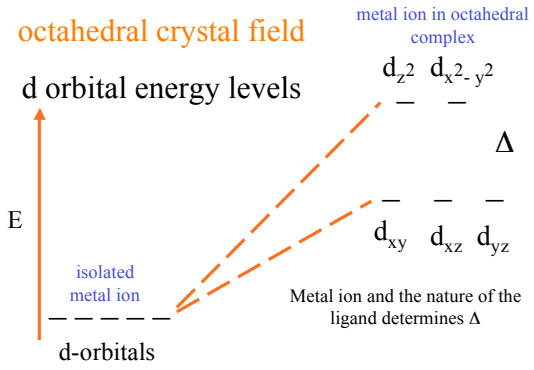
less electrostatic repulsion = lower potential energy

Crystal Field Theory



Crystal Field Splitting Energy





Properties of Transition Metal Complexes

- Properties of transition metal complexes:
  - usually have color
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  - many are paramagnetic
    - due to unpaired d electrons
    - degree of paramagnetism dependent on ligand(s)
      - $[Fe(CN)_6]^{3-}$ has 1 unpaired d electron
      - $[FeF_6]^{3-}$ has 5 unpaired d electrons

Crystal Field Theory

- Crystal Field Theory
  - Can be used to account for
    - Colors of transition metal complexes
      - A complex must have partially filled d subshell on metal to exhibit color
      - A complex with 0 or 10 d e-s is colorless
    - Magnetic properties of transition metal complexes
      - Many are paramagnetic
      - # of unpaired electrons depends on the ligand

Colors of Transition Metal Complexes

- Compounds/complexes that have color:
  - absorb specific wavelengths of visible light (400 –700 nm)
    - wavelengths not absorbed are transmitted

Visible Spectrum

wavelength, nm

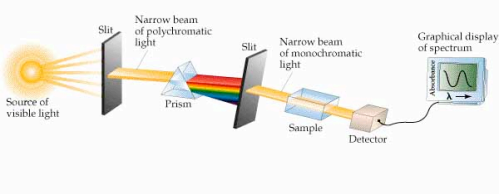
(Each wavelength corresponds to a different color)

400 nm 700 nm

higher energy lower energy

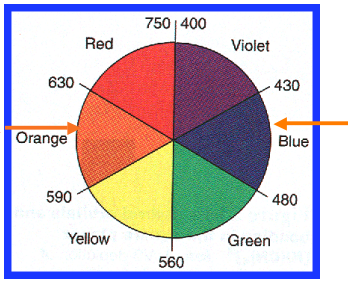
White = all the colors (wavelengths)

Visible Spectrum



Colors of Transition Metal Complexes

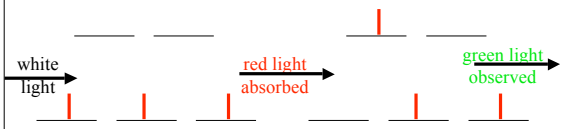
- Compounds/complexes that have color:
  - absorb specific wavelengths of visible light (400 –700 nm)
    - wavelengths not absorbed are transmitted
    - color observed = complementary color of color absorbed



Colors of Transition Metal Complexes

- Absorption of UV-visible radiation by atom, ion, or molecule:
  - Occurs only if radiation has the energy needed to raise an e^- from its ground state to an excited state
    - i.e., from lower to higher energy orbital
    - light energy absorbed = energy difference between the ground state and excited state
    - "electron jumping"

Colors of Transition Metal Complexes



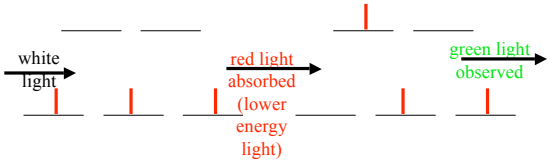
For transition metal complexes, Δ corresponds to energies of visible light.

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

Colors of Transition Metal Complexes

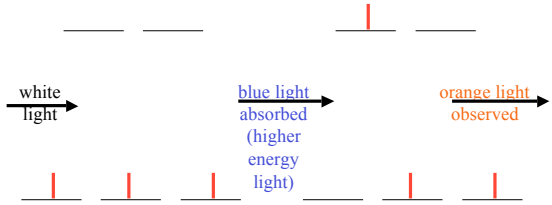
- Different complexes exhibit different colors because:
  - color of light absorbed depends on Δ
    - larger Δ = higher energy light absorbed
      - Shorter wavelengths
    - smaller Δ = lower energy light absorbed
      - Longer wavelengths
  - magnitude of Δ depends on:
    - ligand(s)
    - metal

Colors of Transition Metal Complexes

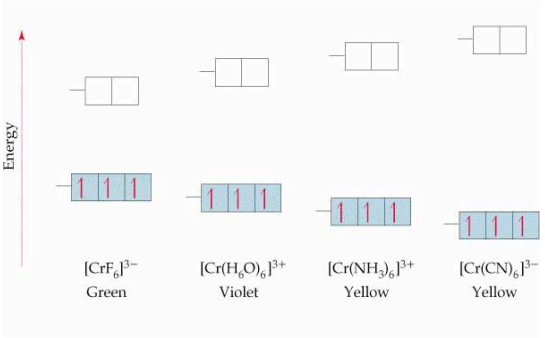


$[M(H_2O)_6]^{3+}$

Colors of Transition Metal Complexes



$[M(en)_3]^{3+}$



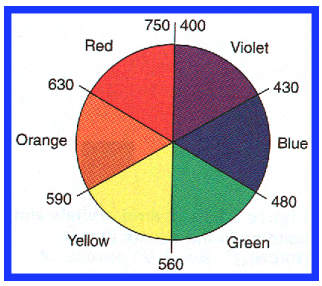
Colors of Transition Metal Complexes

Spectrochemical Series

Smallest Δ — Δ increases — Largest Δ

$I^- < Br^- < Cl^- < OH^- < F^- < H_2O < NH_3 < en < CN^-$

weak field — strong field



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Electronic Configurations of Transition Metal Complexes

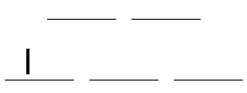
- Expected orbital filling tendencies for e⁻'s:
  - occupy a set of equal energy orbitals one at a time with spins parallel (Hund's rule)
    - minimizes repulsions
  - occupy lowest energy vacant orbitals first
- These are not always followed by transition metal complexes.

Electronic Configurations of Transition Metal Complexes

- d orbital occupancy depends on Δ and pairing energy, P
  - e⁻'s assume the electron configuration with the lowest possible energy cost
  - If Δ > P (Δ large; strong field ligand)
    - e⁻'s pair up in lower energy d subshell first
  - If Δ < P (Δ small; weak field ligand)
    - e⁻'s spread out among all d orbitals before any pair up

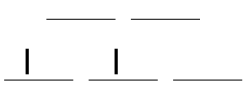
d-orbital energy level diagrams octahedral complex

d^1



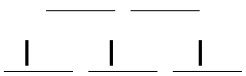
d-orbital energy level diagrams octahedral complex

d^2



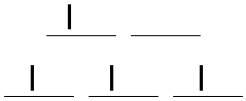
d-orbital energy level diagrams
octahedral complex

d^3

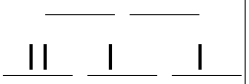


d-orbital energy level diagrams
octahedral complex

d^4



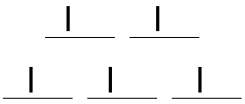
high spin
$\Delta < P$



low spin
$\Delta > P$

d-orbital energy level diagrams
octahedral complex

d^5

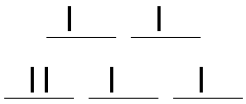


high spin
$\Delta < P$

low spin
$\Delta > P$

d-orbital energy level diagrams
octahedral complex

d^6

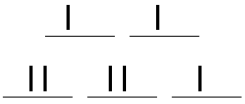


high spin
$\Delta < P$

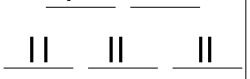
low spin
$\Delta > P$

d-orbital energy level diagrams
octahedral complex

d^7



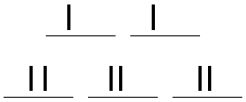
high spin
$\Delta < P$



low spin
$\Delta > P$

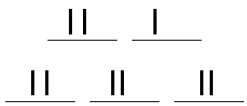
d-orbital energy level diagrams
octahedral complex

d^8



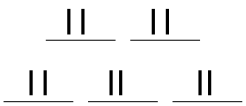
d-orbital energy level diagrams octahedral complex

d^9



d-orbital energy level diagrams octahedral complex

d^{10}



Electronic Configurations of Transition Metal Complexes

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

Colors of Transition Metal Complexes

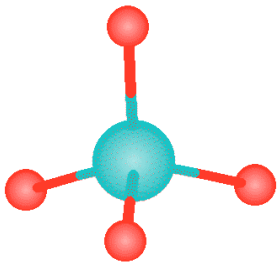
Spectrochemical Series

Smallest Δ — Δ increases — Largest Δ

$I^- < Br^- < Cl^- < OH^- < F^- < H_2O < NH_3 < en < CN^-$

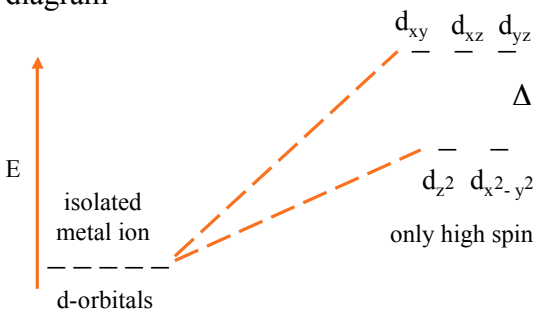
weak field — strong field

d-orbital energy level diagrams tetrahedral complex

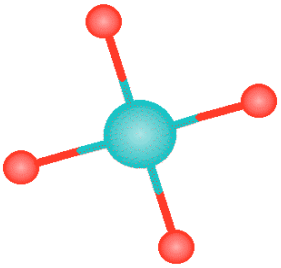


d-orbital energy level diagram

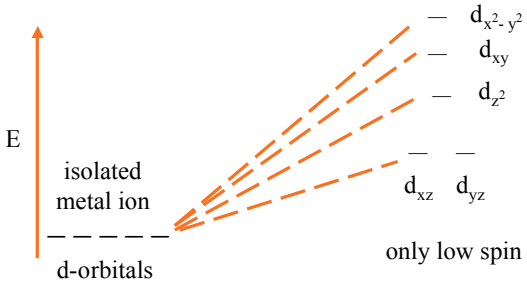
metal ion in tetrahedral complex



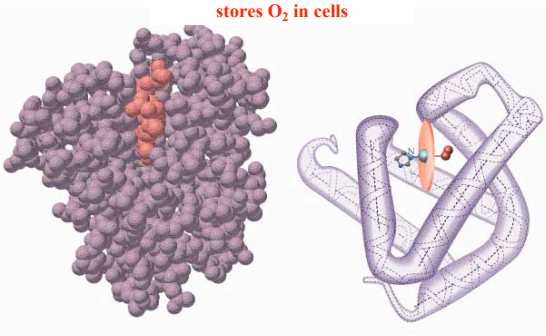
d-orbital energy level diagrams square planar complex



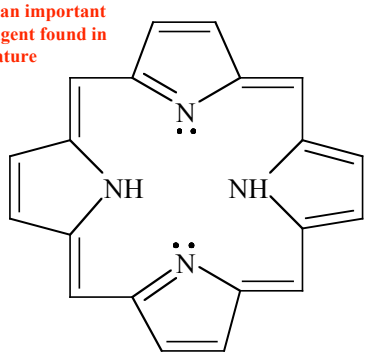
d-orbital energy level diagram



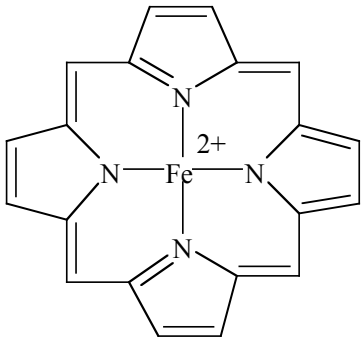
Myoglobin, a protein that stores O_2 in cells



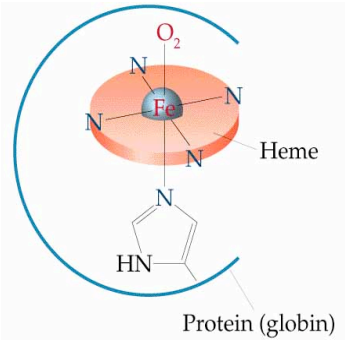
Porphine, an important chelating agent found in nature



Metalloporphyrin

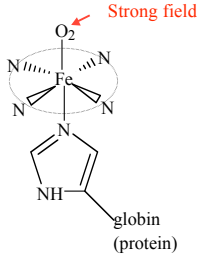


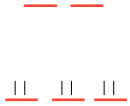
Coordination Environment of Fe^{2+} in Oxymyoglobin and Oxyhemoglobin



Arterial Blood

Strong field



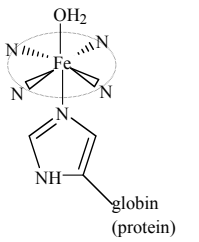


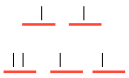
large Δ

Bright red due to absorption of greenish light

Venous Blood

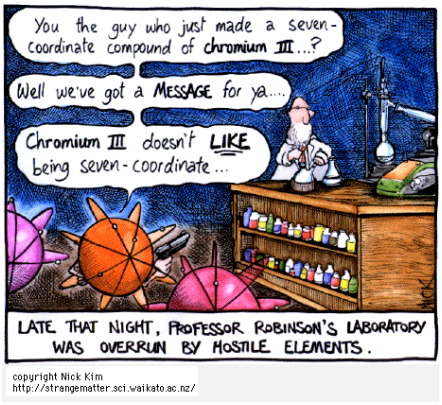
Weak field





small Δ

Bluish color due to absorption of orangish light



End of Presentation