1. Using the spectrochemical series, fill the electrons in the following crystal field diagram for the following tetrahedral complex. Is this the ligand a strong field ligand or a weak field ligand? Also predict if the compound is paramagnetic or diamagnetic. (15 pts)

![Spectrochemical Series]

\[
\begin{align*}
\text{Spectrochemical Series} \\
\Gamma < \text{Br}^- < \text{Cl}^- < \text{F}^- < \text{OH}^- < \text{H}_2\text{O} < \text{NH}_3 < \text{en}^- < \text{NO}_2^- < \text{CN}^- < \text{CO} \\
\text{Tetrahedral} \\
\text{Zn}^{2+} \\
\text{d}^{10} \text{ion} \\
\text{NH}_3 \text{ is a strong field ligand}
\end{align*}
\]

2. The complex ion \([\text{Cr(NH}_3)_6]^{3+}\) is greenish-yellow in solution. Calculate the crystal field splitting energy in kJ/mol for this ion. (10 pts)

\[
\Delta E = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ Js})(3.00 \times 10^{8} \text{ m/s})}{(400 \text{ nm} \times \frac{10^{-9} \text{ m}}{\text{nm}})} \times \frac{6.02 \times 10^{23} \text{ ion/mol}}{1000 \text{ J/kJ}} = 299 \text{ kJ/mol}
\]
3. Balance the following reactions using the half-reaction method: (10 pts)

a) \( \text{Fe}^{2+} + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} \) [acidic solution]

\[
5 \times (\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-) \\
5\text{e}^- + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}
\]

\[
5\text{Fe}^{2+} + \text{MnO}_4^- + 8\text{H}^+ \rightarrow 5\text{Fe}^{3+} + \text{Mn}^{2+} + 4\text{H}_2\text{O}
\]

b) \( \text{Mn}^{2+} + \text{H}_2\text{O}_2 \rightarrow \text{MnO}_2 + \text{H}_2\text{O} \) [basic solution]

\[
2\text{H}_2\text{O} + \text{Mn}^{2+} \rightarrow \text{MnO}_2 + 4\text{H}^+ + 2\text{e}^- \\
2\text{e}^- + 2\text{H}^+ + \text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O}
\]

\[
2\text{OH}^- + \text{Mn}^{2+} + \text{H}_2\text{O}_2 \rightarrow \text{MnO}_2 + 2\text{H}^+ + 2\text{OH}^- \\
2\text{OH}^- + \text{Mn}^{2+} + \text{H}_2\text{O}_2 \rightarrow \text{MnO}_2 + 2\text{H}_2\text{O}
\]

4. State the oxidation number and the coordination number of the transition metals for each of the following compounds. (10 pts)

a) K[CoEDTA]  
C.N. = 6  
O.N. = +3

b) Ag(NH$_3$)$_2$\(^+\)  
C.N. = 2  
O.N. = +1

c) [Co(en)$_3$]PO$_4$  
C.N. = 6  
O.N. = +3

d) MnF$_6$\(^-\)  
C.N. = 6  
O.N. = +5
5. Draw the following transition metal complexes in their correct molecular geometry and state the hybrid orbital involved in the bonding. Name the molecular geometry shape and label the bond angles to receive full credit. (15 pts)

*aSome of the charges and ligands may not be correct due to the drawing program used.*

a) PdCl$_4^{2-}$

\[ \text{d}^8 \text{ion} \rightarrow \text{square planar} \]

\[
\begin{array}{c}
\text{Cl} \\
\text{Pd}^{2-} \\
\text{Cl} \\
\text{Cl} \\
\end{array}
\]

\[ \text{bond angles} = 90^\circ \]

b) $[\text{Co(NH}_3)_5\text{Br}_3]^{3+}$

\[ \text{sp}^3 \text{d} \]

\[
\begin{array}{c}
\text{H}_2\text{N} \\
\text{Co}^{3+} \\
\text{NH}_2 \\
\text{NH}_2 \\
\end{array}
\]

\[ \text{bond angles} = \text{axial} 90^\circ, \text{equatorial} 120^\circ \]

c) $[\text{Cr(H}_2\text{O})_6]\text{Cl}_3$

\[ \text{Octahedral} \]

\[
\begin{array}{c}
\text{H}_2\text{O} \\
\text{Cr}^{3+} \\
\text{OH} \\
\text{OH} \\
\end{array}
\]

\[ \text{bond angles} = 90^\circ \]

d) FeCl$_3$

\[ \text{trigonal planar} \]

\[
\begin{array}{c}
\text{Cl} \\
\text{Fe}^{-} \\
\text{Cl} \\
\text{Cl} \\
\end{array}
\]

\[ \text{bond angles} = 120^\circ \]

6. Draw the isomers for the following compounds: (10 pts)

a. Optical Isomers

\[ [\text{Co(en)}_3]^{3+} \]

b. Coordination Isomers

\[ [\text{Co(NH}_3)_5\text{Cl}]\text{Br} \]

Mirror images

7. Determine the oxidation and reduction half-reactions for the following galvanic cell. Clearly label which half-reaction occurs at the anode and which occurs at the cathode. Calculate $E^\circ_{\text{cell}}$ (volts) and $\Delta G^\circ$ (kJ) for this cell. (10 pts)

<table>
<thead>
<tr>
<th>Reduction Potentials</th>
<th>$E^\circ$ (V)</th>
<th>$F = 96500 \text{ J/V mol}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Half-reactions: 2 x (Ag$^+$ + e$^-$ → Ag)</td>
<td>0.80</td>
<td>Cathode - Reduction</td>
</tr>
<tr>
<td>Cd$^{2+}$ + 2e$^-$ → Cd</td>
<td>-0.40</td>
<td>Anode - Oxidation</td>
</tr>
</tbody>
</table>

$E^\circ_{\text{cell}} = 0.80 \text{ V} + 0.40 \text{ V} = 1.20 \text{ V}$

$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -(2 \text{ mol e-})(96500 \text{ J/Vmol})(1.20 \text{ V}) = -231600 \text{ J} \text{ or } -232 \text{ kJ}$
8. Complete and name the following nuclear equations: (10 pts)

Alpha Decay

\[ ^{241}_{95}Am \rightarrow ^{237}_{93}Np + ^{4}_{2}He \]

Negatron Emission

\[ ^{233}_{91}Pa \rightarrow ^{233}_{92}U + ^{0}_{-1}e \]

Positron Emission

\[ ^{15}_{8}O \rightarrow ^{0}_{+1}e + ^{15}_{7}N \]

Electron Capture

\[ ^{51}_{24}Cr + ^{0}_{-1}e \rightarrow ^{51}_{23}V + ^{0}_{0}Y \]

9. A mammoth skeleton has a carbon-14 decay rate of 0.48 disintegrations per minute per gram (0.48 dis/min g). How long has the mammoth been dead? Living organisms have a carbon-14 decay rate of 15.3 dis/min g and carbon-14 has a half-life of 5730 years. (10 pts)

\[ k = \frac{\ln 2}{t_{1/2}} = \frac{0.693}{5730 \text{ yr}} = 1.21 \times 10^{-4} \text{ yr}^{-1} \]

\[ t = \frac{1}{k} \ln \frac{A_0}{A_t} = \frac{1}{1.21 \times 10^{-4} \text{ yr}^{-1}} \ln \frac{15.3}{0.48} = 28610 \text{ yrs} \]

10. (Extra Credit) The electron-transport chain is a series of spontaneous electrochemical reactions that are responsible for the synthesis of ATP. For the following cytochromes, write a balanced reaction for the spontaneous reaction of electron transfer. Clearly label which cytochrome is the oxidizing agent and which is the reducing agent. (5 pts)

<table>
<thead>
<tr>
<th>Reducton</th>
<th>( E^\circ ) (Volts)</th>
</tr>
</thead>
<tbody>
<tr>
<td>cytochrome ( a_3(Fe^{3+}) + e^- \rightarrow cytochrome ( a_3(Fe^{2+}) )</td>
<td>0.350</td>
</tr>
<tr>
<td>cytochrome ( c(Fe^{3+}) + e^- \rightarrow cytochrome ( c(Fe^{2+}) )</td>
<td>0.254</td>
</tr>
</tbody>
</table>

Reduction  cytochrome \( a_3(Fe^{3+}) + e^- \rightarrow cytochrome \( a_3(Fe^{2+}) \)  0.350 V  “oxidizing agent”

Oxidation  cytochrome \( c(Fe^{2+}) \rightarrow cytochrome \( c(Fe^{3+}) + e^- \)  -0.254 V  0.096 V  “reducing agent”