1. (1 pt) Which of the following substances is the best reducing agent? Briefly explain your answer.
   a) Na⁺  b) Zn  c) Li⁺  d) Ag  e) Al
   It is IMPOSSIBLE for Na⁺ and Li⁺ to act as reducing agents – that would produce Na²⁺ and Li²⁺. So you can limit your consideration to Zn, Ag, and Al. Of these three, Al has the most positive oxidation potential (reverse of reduction potential) = +1.66 V

2. (1 pt) Which of the following substances is the best oxidizing agent? Briefly explain your answer.
   a) O₂  b) Li⁺  c) Cl₂  d) Ag⁺  e) F⁻
   It is IMPOSSIBLE for F⁻ to act as an oxidizing agent – that would produce F²⁻. Li⁺ is very happy where it is, and has a very negative (non-spontaneous) reduction potential, so that can be ruled out as well. That leaves O₂, Cl₂, and Ag⁺. Cl₂ has the most positive reduction potential (+1.36 V), so that is your best answer.

3. (1 pt) Which of the following substances is the best reducing agent? Briefly explain your answer.
   a) F₂  b) Mg  c) Li⁺  d) Na  e) Zn²⁺
   It is IMPOSSIBLE for F₂, Li⁺, and Zn²⁺ to act as reducing agents – that would produce F⁺, Li²⁺, and Zn³⁺. So you can limit your consideration to Mg and Na. Of these three, Na has the most positive oxidation potential (+2.71 V), so that is your best answer.

4. (1 pt) Which of the following substances is the best oxidizing agent? Briefly explain your answer.
   a) H⁺  b) Al³⁺  c) Ag⁺  d) Li  e) O₃
   It is IMPOSSIBLE for Li to act as an oxidizing agent – that would produce Li⁻. Of the others, O₃ has the most positive reduction potential (+2.07 V), so that is your best answer.

5. (3 pts) Balance the following rxn in acidic solution (add water or H⁺ as needed). Clearly show your work.
   Oxidation states:  -1  +7  -2
   I⁻ (aq) + MnO₄⁻(aq) ⇄ I₂(aq) + MnO₂(s)

   The two half cell rxns are:
   Oxidation:  2I⁻ → I₂ + 2e⁻
   Reduction:  MnO₄⁻ + 3e⁻ → MnO₂
   Note that you have to have at least 2I⁻ in order to make one I₂. It is important to perform this internal balancing within the half cell first.
   The common factor is 6e⁻, so we need to multiply the first half reaction by 3 and the second by 2, then add the reactions together and cancel out the 6e⁻ on each side to give the overall core redox balanced reaction:
   6I⁻ + 2MnO₄⁻ → 3I₂ + 2MnO₂

   Now, check that the # of oxygen atoms on each side matches!

   There are 8 oxygens on the reactant side, and 4 oxygens on the product side. Add 4 water molecules to the product side that is missing oxygens, then add 8H⁺ to balance the hydrogens on the reactant side to produce our final balanced equation. Note that the spectator cations are not included.
   8H⁺ + 6I⁻ + 2MnO₄⁻ → 3I₂ + 2MnO₂ + 4H₂O
6. (3 pts) Balance the following rxn in basic solution (add water or OH\(^{-}\) as needed). Clearly show your work.

\[
\text{Oxidation states:} \quad +6 \quad -2 \quad +1 \quad -2
\]

\[
\text{Cr}_2\text{O}_7^{2-}(aq) \quad + \quad \text{ClO}^{-}(aq) \quad \underset{\text{Red half rxn:}}{\longrightarrow} \quad \underset{\text{Ox half rxn:}}{\longrightarrow} \quad \text{Cr}^{3+}(aq) \quad + \quad \text{ClO}_4^{-}(aq)
\]

Write out the REDOX half cells and balance:

**Ox half rxn:** \( \text{ClO}^{-} \rightarrow \text{ClO}_4^{-} + 6e^{-} \)

**Red half rxn:** \( \text{Cr}_2\text{O}_7^{2-} + 6e^{-} \rightarrow 2\text{Cr}^{3+} \)

Note that we need to add \( 2\text{Cr}^{3+} \) because the \( \text{Cr}_2\text{O}_7^{2-} \) has two redox active \( \text{Cr}^{6+} \) atoms in it. Each \( \text{Cr}^{6+} \) reacts with \( 3e^{-} \) for a total of \( 6e^{-} \). Thus, each half cell uses \( 6e^{-} \) and we are balanced:

\[
\text{Cr}_2\text{O}_7^{2-} + \text{ClO}^{-} \rightarrow 2\text{Cr}^{3+} + \text{ClO}_4^{-}
\]

Now, check that the # of oxygen atoms on each side matches!

There are 8 oxygens on the reactant side, and 4 oxygens on the product side. Add 4 water molecules to the product side that missing 4 oxygens, then add \( 8\text{H}^{+} \) to the reactant side to balance the hydrogens.

7. (5 pts) Write the oxidation state for the underlined element in the box following each compound.

- a) NaH \( \underline{-1} \)
- b) KNO\(_3\) \( +5 \)
- c) Na\(_2\)PtCl\(_6\) \( +4 \)
- d) Ca\(_3\)(PO\(_3\))\(_2\) \( +3 \)
- e) Na(NCS) \( -2 \)

8. (5 pts) Calculate the redox potentials for the following reactions. Show the two half cell reactions used to calculate the overall potential.

- a) \( 3\text{H}_2(g) + 2\text{Al}^{3+}(aq) \underset{\text{Non-spontaneous}}{\longrightarrow} 6\text{H}^{+}(aq) + 2\text{Al}(s) \)

\[
\begin{align*}
\text{H}_2 & \rightarrow 2\text{H}^{+} + 2e^{-} \\
\text{Al}^{3+} + 3e^{-} & \rightarrow \text{Al}
\end{align*}
\]

\( E^\circ = 0.00 \text{ V} \)

\( E^\circ = -1.66 \text{ V} \)

- b) \( 2\text{AgCl}(s) + \text{Mg}(s) \underset{\text{Non-spontaneous}}{\longrightarrow} 2\text{Ag}(s) + \text{Mg}^{2+}(aq) + 2\text{Cl}^{-}(aq) \)

\[
\begin{align*}
\text{AgCl} + 1e^{-} & \rightarrow \text{Ag} + \text{Cl}^{-} \\
\text{Mg} & \rightarrow \text{Mg}^{2+} + 2e^{-}
\end{align*}
\]

\( E^\circ = 0.22 \text{ V} \)

\( E^\circ = 2.36 \text{ V} \)

- c) \( 2\text{F}_2(g) + 2\text{H}_2\text{O}(l) \underset{\text{Non-spontaneous}}{\longrightarrow} 4\text{F}^{-}(aq) + 4\text{H}^{+}(aq) + \text{O}_2(g) \)

\[
\begin{align*}
\text{F}_2 + 2e^{-} & \rightarrow 2\text{F}^{-} \\
2\text{H}_2\text{O} & \rightarrow \text{O}_2 + 4\text{H}^{+} + 4e^{-}
\end{align*}
\]

\( E^\circ = -2.87 \text{ V} \)

\( E^\circ = -1.23 \text{ V} \)

- d) \( \text{Mg}^{2+}(aq) + \text{Cu}(s) \underset{\text{Non-spontaneous}}{\longrightarrow} \text{Mg}(s) + \text{Cu}^{2+}(s) \)

\[
\begin{align*}
\text{Mg}^{2+} + 2e^{-} & \rightarrow \text{Mg} \\
\text{Cu} & \rightarrow \text{Cu}^{2+} + 2e^{-}
\end{align*}
\]

\( E^\circ = -2.36 \text{ V} \)

\( E^\circ = -0.34 \text{ V} \)

- e) \( \text{Li}(s) + \text{Ag}^{+}(aq) \underset{\text{Non-spontaneous}}{\longrightarrow} \text{Li}^{+}(aq) + \text{Ag}(s) \)

\[
\begin{align*}
\text{Ag}^{+} + 1e^{-} & \rightarrow \text{Ag} \\
\text{Li} & \rightarrow \text{Li}^{+} + 1e^{-}
\end{align*}
\]

\( E^\circ = 0.80 \text{ V} \)

\( E^\circ = 3.05 \text{ V} \)

\( 3.85 \text{ V} \)
9. (4 pts) A MgCl$_2$ solution containing a Mg electrode is connected by means of a salt bridge to a CuCl$_2$ solution containing a copper electrode. Sketch out a Galvanic Cell showing this and clearly indicate the movement of anions, cations, electrons, which electrode is dissolving, and which is forming a metallic deposit. Label the anode and cathode and show the cell potential.

The spontaneous rxn is:  
\[ \text{Mg(s)} + \text{Cu}^{2+} \longrightarrow \text{Mg}^{2+} + \text{Cu(s)} \]

<table>
<thead>
<tr>
<th>Anode Reaction</th>
<th>Cathode Reaction</th>
<th>( E^° )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg(s) ( \rightarrow ) Mg$^{2+}$(aq) + 2e$^-$</td>
<td>Cu$^{2+}$(aq) + 2e$^-$ ( \rightarrow ) Cu(s)</td>
<td>2.36 V</td>
</tr>
</tbody>
</table>

Cell potential = 2.70 V

10. (3 pts) How long (in hours) will it take to electrodeposit 1 mole of Al metal by passing a current of 9.65 amps through a solution of Al$^{3+}$? Please clearly show all your work and put a box around your final answer.

\[ \text{#sec} = \frac{\text{(#moles) (96,485 C/mole e$^-$)(3e$^-$)}}{\text{(9.65 C/sec)}} \]
\[ \text{#sec} = 29,995 \approx 30,000 \]
\[ \text{#hrs} = \frac{30,000 \text{ sec}}{3600 \text{ sec/hr}} = 8.33 \text{ hr} \]

11. (3 pts) What is the concentration of [Ag$^+$] in a half-cell if the reduction potential of the Ag$^+/Ag$ couple is observed to be 0.40 V? Clearly show all your work and put a box around your final answer.

\[ \text{Ag}^+(aq) + e^- \longrightarrow \text{Ag(s)} \hspace{1cm} +0.80 \hspace{1cm} Q = \frac{1}{[\text{Ag}^+]} \]

Rearrange the Nernst equation and solve for \( Q \), from which we can calculate [Ag$^+$]:
\[ E = E^° - \frac{0.052}{n} \log(Q) \]
\[ \log(Q) = \frac{n(E° - E)}{0.0592} = \frac{(1)(0.4)}{0.0592} = 6.757 \]

Take the antilog (10$^x$) of each side to get the value of Q:  
\[ Q = 10^{6.767} = 5.71 \times 10^6 \]

\[ [\text{Ag}^+] = \frac{1}{Q} = \frac{1}{5.71 \times 10^6} = 1.75 \times 10^{-7} \text{ M} \]