1. (15 points) Fill in the blank or space with the correct response.

For a-e below, the following table is to be used.

<table>
<thead>
<tr>
<th>Half-Reaction</th>
<th>E° (V)</th>
<th>Half-Reaction</th>
<th>E° (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Co²⁺ + 2 e⁻ V Co</td>
<td>-0.277</td>
<td>Ti³⁺ + e⁻ V Ti²⁺</td>
<td>-0.369</td>
</tr>
<tr>
<td>I² + 2 e⁻ V 2 I⁻</td>
<td>+0.536</td>
<td>Sn⁴⁺ + 2 e⁻ V Sn²⁺</td>
<td>+0.154</td>
</tr>
</tbody>
</table>

(a) Which chemical species is the strongest oxidizing agent? ______________________

(b) Which chemical species is the strongest reducing agent? ______________________

(c) Which chemical species is the weakest oxidizing agent? ______________________

(d) Which chemical species is the weakest reducing agent? ______________________

(e)(2 points) Choosing two of the above eight chemical species, write the balanced redox equation in the space below that would have the largest possible equilibrium constant.

For f-g below, the following equation applies: Cu²⁺ + 2 Ag + 2 Br⁻ V Cu + 2 AgBr

(f) Which species is the oxidizing agent? ______________________

(g) Which species is the reducing agent? ______________________

(h) Give the oxidation number of sulfur in HSO₃⁻. ______________________

(i) Give the oxidation number of chromium in Cr(OH)₄⁻. ______________________

(j) Give the oxidation number of carbon in H₂C₂O₄. ______________________

For k-n below, predict the sign (positive, negative, or zero) for each.

(k) ∆S for the evaporation of water ______________________

(l) ∆H for the burning of coal ______________________

(m) ∆G for the corrosion of iron metal ______________________

(n) ∆H° for copper metal ______________________

2. (15 points) Balance the following equation in basic solution.
\[ \text{H}_2\text{O}_2 \ + \ \text{ClO}_4^- \ \rightleftharpoons \ \text{ClO}_2^- \ + \ \text{O}_2 \]

3. (20 points) Calculate the \( \Delta H^0 \) for the reaction: \( 3\text{C}(s) \ + \ 4\text{H}_2(g) \ \rightleftharpoons \ \text{C}_3\text{H}_6(g) \) using the equations

\[
\begin{align*}
\text{C}_3\text{H}_8(g) & \ + \ 5\text{O}_2(g) \ \rightleftharpoons \ 3\text{CO}_2(g) \ + \ 4\text{H}_2\text{O}(l) \quad \Delta H^0 = -2220 \text{ kJ} \\
\text{C}(s) & \ + \ \text{O}_2(g) \ \rightleftharpoons \ \text{CO}_2(g) \quad \Delta H^0 = -394 \text{ kJ} \\
\text{H}_2(g) & \ + \ \frac{1}{2}\text{O}_2(g) \ \rightleftharpoons \ \text{H}_2\text{O}(l) \quad \Delta H^0 = -286 \text{ kJ}
\end{align*}
\]
4. (25 points) Consider the following reaction and corresponding thermodynamic data at 25°C.

\[
\text{CS}_2(g) + 4\text{H}_2(g) \rightleftharpoons \text{CH}_4(g) + 2\text{H}_2\text{S}(g)
\]

<table>
<thead>
<tr>
<th></th>
<th>CS(_2)(g)</th>
<th>H(_2)(g)</th>
<th>CH(_4)(g)</th>
<th>H(_2\text{S})(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(\Delta H_f^0) (kJ/mol)</td>
<td>117</td>
<td>0</td>
<td>–74.9</td>
<td>–20.2</td>
</tr>
<tr>
<td>(S^0) (J/K·mol)</td>
<td>237.8</td>
<td>130.6</td>
<td>186.1</td>
<td>205.6</td>
</tr>
</tbody>
</table>

(a) Determine the enthalpy change for the above reaction.

(b) Determine the entropy change for the above reaction.

(c) Determine \(\Delta G^0\) for the above reaction. Is the reaction spontaneous? Why or why not?

(d) Determine the equilibrium constant for the above reaction.

\{R = 8.314 \text{ J/K·mol} \text{ and } F = 96485 \text{ C/mol}\}
5. (25 points) Consider the following cell (assume cathode is on the right and the anode on the left):

\[
\begin{align*}
\text{Ag}^+ (0.010 \text{ M}) & \rightarrow \text{Ag} \quad \text{**Cu}^{2+} (0.20 \text{ M}) & \rightarrow \text{Cu} \\
E^0 &= +0.80 \text{ V} & E^0 &= +0.34 \text{ V}
\end{align*}
\]

(a) What is the overall balanced reaction corresponding to the above cell?

(b) What is \( E^0_{\text{cell}} \)?

(c) What is the cell potential?

(d) Is the cell galvanic or electrolytic? Explain.

(e) What is the equilibrium constant corresponding to the reaction in part (a) above? \\
\{ R = 8.314 \text{ J/K·mol} \text{ and } F = 96485 \text{ C/mol}\}