**Electrochemistry**

**ELECTROLYSIS WORKSHEET**

<table>
<thead>
<tr>
<th>Standard Reduction Potential</th>
<th>( E^0 ) (volts)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Cl}_2(g) + 2e^- \rightarrow 2\text{Cl}^-(aq) )</td>
<td>+1.36</td>
</tr>
<tr>
<td>( \text{O}_2(g) + 4\text{H}^+(aq) + 4e^- \rightarrow 2\text{H}_2\text{O}(l) )</td>
<td>+1.23</td>
</tr>
<tr>
<td>( \text{Ag}^+(aq) + e^- \rightarrow \text{Ag}(s) )</td>
<td>+0.80</td>
</tr>
<tr>
<td>( \text{I}_2(s) + 2e^- \rightarrow 2\text{I}^-(aq) )</td>
<td>+0.535</td>
</tr>
<tr>
<td>( \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s) )</td>
<td>+0.337</td>
</tr>
<tr>
<td>( \text{SO}_4^{2-}(aq) + 4\text{H}^+(aq) + 2e^- \rightarrow \text{SO}_2(g) + 2\text{H}_2\text{O} )</td>
<td>+0.20</td>
</tr>
<tr>
<td>( 2\text{H}^+(aq) + 2e^- \rightarrow \text{H}_2(g) ) (reference electrode)</td>
<td>0.00</td>
</tr>
<tr>
<td>( 2\text{H}_2\text{O}(l) + 2e^- \rightarrow \text{H}_2(g) + 2\text{OH}^- (aq) )</td>
<td>-0.828</td>
</tr>
<tr>
<td>( \text{Na}^+(aq) + e^- \rightarrow \text{Na}(s) )</td>
<td>-2.714</td>
</tr>
<tr>
<td>( \text{K}^+(aq) + e^- \rightarrow \text{K}(s) )</td>
<td>-2.93</td>
</tr>
</tbody>
</table>

1. All of the equations in the chart above are written as _reductions_ (oxidations/reductions).

2. The chemicals at the upper left (Cl\(_2\) and O\(_2\)) are the most likely to be _reduced_ (oxidized/reduced) and therefore the best _oxidizing agents_ (oxidizing agents/reducing agents).

3. The chemicals at the lower right (Na\(^+\) and K\(^+\)) are the most likely to be _oxidized_ (oxidized/reduced) and therefore the best _reducing agents_ (oxidizing agents/reducing agents).

4. In an electrolytic cell, the (-) electrode is negative because it has _too many_ (too many/too few) electrons. Chemicals that come into contact with the (-) electrode will _gain_ (gain/lose) electrons and be _reduced_ (oxidized/reduced). The (-) electrode in electrolysis is called the _cathode_ (cathode/anode).

5. Write the change that water goes through at the (-) electrode. \( 2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2\text{OH}^- \)

6. In an electrolytic cell, the (+) electrode is positive because it has _too few_ (too many/too few) electrons. Chemicals that come into contact with the (+) electrode will _lose_ (gain/lose) electrons and be _oxidized_ (oxidized/reduced). The (+) electrode in electrolysis is called the _anode_ (cathode/anode).

7. Write the change that water goes through at the (+) electrode. \( 2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4e^- \)

8. Add these two reactions together (make certain the electrons cancel) and write the overall reaction for the electrolysis of water. \( 2\text{H}_2\text{O} \rightarrow \text{O}_2 + 2\text{H}_2 \)

9. We will perform this electrolysis using an aqueous solution of sodium sulfate. Both the Na\(^+\) and H\(_2\)O will be near the (-) electrode. Which chemical is more likely to be reduced? \( \text{H}_2\text{O} \)

   \[ E^0 = -2.714 \text{ V} \]

10. Both the SO\(_4^{2-}\) and H\(_2\)O will be near the (+) electrode. Which chemical will be oxidized? \( \text{H}_2\text{O} \)

   \[ E^0 = -0.828 \text{ V} \]

\[ \text{SO}_4^{2-} \] cannot be oxidized so \( \text{H}_2\text{O} \) must be oxidized.

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

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

\[ 4\text{H}_2\text{O} + 4e^- \rightarrow \text{2H}_2 + 4\text{OH}^- \]
11. In the electrolysis of KCl(aq),
Both the K⁺ and Cl⁻ will be near the (-) electrode. Which chemical is more likely to be reduced?
Both the H₂O and H₂ will be near the (+) electrode. Which chemical is more likely to be oxidized?
Write the reactions at each electrode and the overall reaction:
Cathode:
\[ 2H_2O + 2e^- \rightarrow H_2 + 2OH^- \]
Anode:
\[ 2Cl^- \rightarrow Cl_2 + 2e^- \]
Overall:
\[ 2H_2O + 2Cl^- \rightarrow Cl_2 + H_2 + 2OH^- \]

12. In the electrolysis of CuSO₄(aq)
Both the Cu²⁺ and H₂O will be near the (-) electrode. Which chemical will be reduced?
Both the SO₄²⁻ and H₂O will be near the (+) electrode. Which chemical will be oxidized?
Write the reactions at each electrode and the overall reaction:
Cathode:
\[ Cu^{2+} + 2e^- \rightarrow Cu^{0} \]
Anode:
\[ 2H_2O \rightarrow O_2 + 4H^+ + 4e^- \]
Overall:
\[ 2Cu^{2+} + 2H_2O \rightarrow O_2 + 4H^+ + 2Cu \]

13. Silver plating occurs when electrolysis of a Ag₂SO₄ solution is used because silver metal is formed at the cathode (cathode/anode).
This is the (-) (+) electrode. The reaction at this electrode is: \[ Ag^{+} + e^- \rightarrow Ag^{0} \] (reduction)
Recall that 1 amp-sec = 1 Coulomb and 96,500 Coulombs = 1 mole e⁻ (Faraday's constant).
If a cell is run for 200. seconds with a current of 0.250 amps, how many grams of Ag⁰ will be deposited?
\[ 200 \text{ s} \times 0.250 \text{ amp} \times \frac{1 \text{ C}}{1 \text{ amp} \cdot 1 \text{ s}} \times \frac{1 \text{ mole} \text{ e}^-}{96,500 \text{ C}} \times \frac{1 \text{ mole Ag}}{1 \text{ mole e}^-} \times \frac{107.87 \text{ g}}{1 \text{ mole Ag}} = 0.056 \text{ g Ag} \]

14. A current of 10.0 amperes flows for 2.00 hours through an electrolytic cell containing a molten salt of metal X. This results in the decomposition of 0.250 mole of metal X at the cathode. The oxidation state of X in the molten salt is (X³⁺, X₂⁺, X⁺, X⁰)
\[ 10.0 \text{ amp} \times 2.00 \text{ h} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{1 \text{ C}}{1 \text{ amp} \cdot 1 \text{ s}} \times \frac{1 \text{ mole e}^-}{96,500 \text{ C}} = 0.746 \text{ mole e}^- \]
\[ \frac{0.746 \text{ mole e}^-}{0.250 \text{ mole X}^{3+}} = 3 \text{ mole e}^- \times \frac{X^{3+}}{X^{3+}} \]

15. Solutions of Ag⁺, Cu²⁺, Fe³⁺ and Ti⁴⁺ are electrolyzed with a constant current until 0.10 mol of metal is deposited. Which will require the greatest length of time? Ti⁴⁺