Part I

If the following electrolytes were electrolyzed, predict what half reactions would occur at the anode and cathode. (Remember, 1.0 M NaCl implies an aqueous solution of NaCl.)

1. Molten ZnCl₂; inert electrodes.

\[
\text{cathode: } \text{Zn}^{2+} + 2e^- \rightarrow \text{Zn} \\
\text{anode: } 2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^- 
\]

2. 1.0 M Na₂SO₄; inert electrodes.

\[
\text{anode: } \text{H}_2\text{O} \rightarrow \frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2e^- \\
\text{cathode: } 2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2\text{OH}^- 
\]

3. Molten CuCl₂; inert electrodes.

\[
\text{cathode: } \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \\
\text{anode: } 2\text{I}^- \rightarrow \text{I}_2 + 2e^- 
\]

4. 1.0 M KI; inert electrodes.

\[
\text{anode: } 2\text{I}^- \rightarrow \text{I}_2 + 2e^- \\
\text{cathode: } 2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2\text{OH}^- 
\]

5. 1.0 M NiSO₄; nickel electrodes.

\[
\text{anode: } \text{Ni} \rightarrow \text{Ni}^{2+} + 2e^- \\
\text{cathode: } \text{Ni}^{2+} + 2e^- \rightarrow \text{Ni} 
\]

6. 1.0 M HI; lead electrodes.

\[
\text{anode: } \text{Pb}(s) \rightarrow \text{Pb}^{2+} + 2e^- \\
\text{cathode: } 2\text{H}_2 + 2e^- \rightarrow \text{H}_2(g) 
\]
7. 1.0M ZnSO₄; inert electrodes.

**Anode:**  \[ H_2O \rightarrow \frac{1}{2}O_2 (g) + 2H^+ + 2e^- \]

**Cathode:**  \[ Zn^{2+} + 2e^- \rightarrow Zn(s) \quad \text{overpotential effect} \]

8. 1.0M NaCl; inert electrodes.

**Anode:**  \[ 2Cl^- \rightarrow Cl_2 (g) + 2e^- \quad \text{overpotential effect} \]

**Cathode:**  \[ 2H_2O + 2e^- \rightarrow H_2 (g) + 2OH^- \]

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**Part II**

9. When a 1.0M KCl solution is electrolyzed using silver electrodes, a precipitate forms at the anode. Explain this result.

**Anode:**  \[ Ag(s) \rightarrow Ag^+ + e^- \]

*The Ag⁺ ions react with the Cl⁻ ions to produce AgCl(s).*

10. If a 1.0M HCl solution is electrolyzed using platinum electrodes (inert), what minimum voltage must be applied? Predict the anode and cathode half-reactions, as well as the overall cell reaction.

**Anode:**  \[ 2Cl^- \rightarrow Cl_2 (g) + 2e^- \quad E^0 = -1.36 \text{V} \]

**Cathode:**  \[ 2H^+ + 2e^- \rightarrow H_2 (g) \quad E^0 = 0.00 \text{V} \]

**Overall:**  \[ 2Cl^- + 2H^+ \rightarrow Cl_2 (g) + H_2 (g) \quad E^{\text{cell}} = -1.36 \text{V} \]

*Therefore, 1.36 V must be applied.*

11. a. In the electrorefining of lead, lead bullion is used as the anode and pure lead is used as the cathode in an electrolytic solution containing Pb²⁺ ions. Lead bullion is primarily lead, but it does contain impurities such as silver and gold. What happens to these three metals at the anode during electrolysis?

*The Pb is oxidized to Pb²⁺. The Ag + Au are left behind. These metals can be recovered in subsequent electrolytic processes.*
11. b. Lead bullion may also contain trace amounts of impurities such as zinc metal. Describe what happens to this zinc during electrolysis, and explain why the pure lead cathode does not become contaminated with zinc.

Although both the Zn(s) and the Pb(s) are oxidized at the anode, the Zn\textsuperscript{2+} ions are "trapped" in solution. Since, at the cathode, the Pb\textsuperscript{2+} ions have a higher reduction potential than the Zn\textsuperscript{2+}, only the Pb\textsuperscript{2+} ions are reduced to Pb(s).

12. Why can aluminum metal not be produced by electrowinning Al(s) from an aqueous solution containing Al\textsuperscript{3+} ions? Write the cathode half-reaction that would occur.

Even with hydrogen's high overpotential, H\textsubscript{2}O is reduced before Al\textsuperscript{3+}.

\text{cathode:} \quad 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-