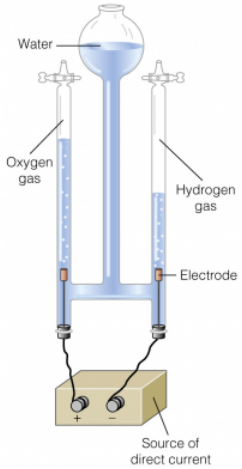
**Demonstration [5 pts]** Name_____
**Electrolytic Cells** Period _____ Date_____

**Voltaic Cells (Galvanic Cells):** Spontaneous reactions are used to produce electricity. *$E^\circ_{net}$ are positive!*
**Electrolytic Cells**: Non-spontaneous reactions are forced to occur when an input of electricity (from a battery or poser source) is used. *$E^\circ_{net}$ are negative!*




**Demonstration #1: The Electrolysis of water (Hoffman Apparatus)**

a) **Set up**: Hoffman apparatus is filled with water, $Na_2SO_4$ and the indicator, bromthymol blue. The apparatus is connected to a power supply.

b) Observe the colors in the left and right-hand tubes. **Label diagram w/ colors.**
c) Bromthymol blue indicator has the following colors:
*acidic solution: yellow* *neutral solution: green* *basic solution: blue*
Based on the colors formed**, label the diagram with $H^+$ and $OH^-$**

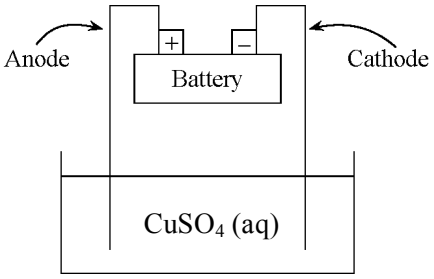
d) Fill out the oxidation and reduction half reactions that must be taking place. Balance $e^-$, write overall Rxn and fill in $E^\circ$ values and $E^\circ_{net}$.

**Ox**: $H_2O\,(l) \rightarrow$ $E^\circ_{ox} =$
**Red:** $H_2O\,(l) \rightarrow$ $E^\circ_{red} =$
**Overall Rxn:** $E^\circ_{net} =$

e) Based on where $H_2$ (g) and $O_2$ (g) are each produced, **label the anode and cathode in the diagram**.
f) How does the $E^\circ_{net}$ value prove to you that electricity was needed to do this reaction?

**Demonstration #2: Electrolysis of $CuSO_4$** (Electroplating)

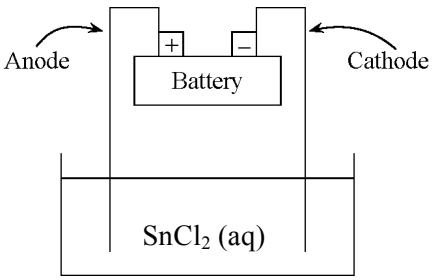
a) **Setup:** Two paper clips attached to a power source (9 V) are placed into $CuSO_4$ (aq). What ions are in the solution? _____
b) In the diagram, label what is observed at the anode and cathode.
c) After writing the half reactions below, go back and label the diagram with the substances observed at the anode and cathode.
d) Given the overall reaction below, fill in the half reactions, balance electrons, fill in $E^\circ$ values and calculate $E^\circ_{net}$.




**Ox:** $E^\circ_{ox} =$
**Red:** $E^\circ_{red} =$
**Overall:** $2\,Cu^{+2}\,(aq) + 2\,H_2O \rightarrow 2\,Cu\,(s) + O_2\,(g) + 4\,H^+$ **$E^\circ_{net} =$**

**Demonstration #3: The electrolysis of $SnCl_2$ (aq)**

a) **Setup:** Two paper clips are attached to a power source (9V) and are placed into $SnCl_2$ (aq). What ions are in the sol'n? _____
b) In the diagram, label what is observed at the anode and cathode.
c) After writing the half reactions below, go back and label the diagram with the substances observed at the anode and cathode.
d) Given the overall reaction below, fill in the half reactions, balance electrons, fill in $E^\circ$ values and calculate $E^\circ_{net}$.




**Ox:** $E^\circ_{ox} =$
**Red:** $E^\circ_{red} =$
**Overall:** $Sn^{2+}\,(aq) \rightarrow Sn\,(s) + Sn^{4+}\,(aq)$ **$E^\circ_{net} =$**

then $Sn^{4+} + 4\,Cl^- \rightarrow SnCl_4\,(s)$

**Post Lab Questions on Reading:** Read article, "Back Burner: Friedrich Wöhler's Lost Aluminum"

1) Wöhler's preparation of Al metal in 1827:
   a) Write out the reaction that Wöhler did when he first prepared Al metal in 1827.

   b) What was the drawback to this process?

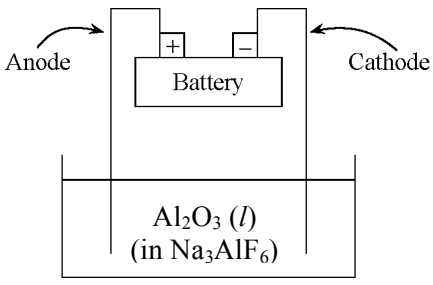
2) Davy's electrolytic method of separating a compound into its component elements:
   a) What metals could Davy produce by the electrolysis of salts?

   b) Why was the electrolytic method seemingly impractical for the preparation of Al in 1827, but possible for metals such as potassium?

3) Charles Hall's electrolytic process for the preparation of Al: (Made aluminum significantly cheaper.) .
   a) Describe Charles Hall's electrolytic process for the production of aluminum metal in 1886.

   b) Why did this method make it possible to directly make aluminum from its ore by electrolysis?

4) The basic reaction for the electrolysis of Al is shown below.
   a) Write out the half reactions (split $Al_2O_3$ into ions and use ions in half reactions).
   b) In diagram, label where Al (s) and $O_2$ (g) are produced.




**$Al_2O_3$ (s) → Al (s) + $O_2$ (g)**

Ox:

Red: