

Name: \_\_\_\_\_

Date: \_\_\_\_\_

Period: \_\_\_\_\_

Seat #: \_\_\_\_\_

Show all work

| Standard Reduction Potential   | E° (volts) |
|--|------------|
| $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$  | +1.36      |
| $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$                      | +1.23      |
| $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$  | +0.80      |
| $\text{I}_2(\text{s}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$  | +0.535     |
| $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$  | +0.337     |
| $\text{SO}_4^{2-}(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{SO}_2(\text{g}) + 2\text{H}_2\text{O}$ | +0.20      |
| $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$ (reference electrode)                                | 0.00       |
| $2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-(\text{aq})$                     | -0.828     |
| $\text{Na}^+(\text{aq}) + \text{e}^- \rightarrow \text{Na}(\text{s})$  | -2.714     |
| $\text{K}^+(\text{aq}) + \text{e}^- \rightarrow \text{K}(\text{s})$  | -2.93      |

- All of the equations in the chart above are written as **REDUCTIONS** (oxidations/reductions).
- The chemicals at the upper left ( $\text{Cl}_2$  and  $\text{O}_2$ ) are the most likely to be **REDUCED** (oxidized/reduced) and therefore the best **OXIDIZING AGENTS** (oxidizing agents/reducing agents).
- 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).
- 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).
- Write the change that water goes through at the (-) electrode.  **$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g}) + 2\text{OH}^-$**
- In an electrochemical 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).
- Write the change that water goes through at the (+) electrode.  **$2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{e}^- + 4\text{H}^+(\text{aq}) + \text{O}_2(\text{g})$**
- Add these two reactions together (make certain the electrons cancel) and write the overall reaction for the electrolysis of water.  **$2\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2(\text{g}) + \text{O}_2(\text{g})$**
- We will perform this electrolysis using an aqueous solution of sodium sulfate. Both the  $\text{Na}^+$  and  $\text{H}_2\text{O}$  will be near the (-) electrode. Which chemical is more likely to be reduced?  **$\text{H}_2\text{O}$ ,  $E_{\text{Na}^+}^\circ = -2.714\text{ V}$ ,  $E_{\text{H}_2\text{O}}^\circ = -0.828$ , more + is reduced**
- Both the  $\text{SO}_4^{2-}$  and  $\text{H}_2\text{O}$  will be near the (+) electrode. Which chemical will be oxidized?  **$\text{H}_2\text{O}$**   
\*  $\text{SO}_4^{2-}$  cannot be oxidized therefore  $\text{H}_2\text{O}$  must be

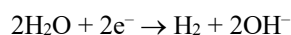
[11] In the electrolysis of  $\text{KI}(\text{aq})$

Both the  $\text{K}^+$  and  $\text{H}_2\text{O}$  will be near the (-) electrode. Which chemical is more likely to be reduced?  **$\text{H}_2\text{O}$ , look at chart above,  $E^\circ$**

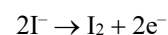
Both the  $\text{I}^-$  and  $\text{H}_2\text{O}$  will be near the (+) electrode. Which chemical is more likely to be oxidized?  **$\text{I}^-$**

Write the reactions at each electrode and the overall reaction:

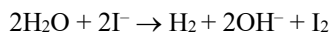
Cathode



Anode



Overall



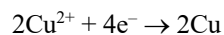
[12] In the electrolysis of  $\text{CuSO}_4(\text{aq})$

Both the  $\text{Cu}^{2+}$  and  $\text{H}_2\text{O}$  will be near the (-) electrode. Which chemical will be reduced?  $\text{Cu}^{2+}$

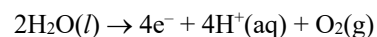
Both the  $\text{SO}_4^{2-}$  and  $\text{H}_2\text{O}$  will be near the (+) electrode. Which chemical will be oxidized?  $\text{H}_2\text{O}$

Write the reactions at each electrode and the overall reaction:

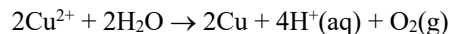
Cathode



Anode



Overall



[13] Silver plating occurs when electrolysis of a  $\text{Ag}_2\text{SO}_4$  solution is used because silver metal is formed at the **CATHODE** (cathode/anode).

This is the (-) (+/-) electrode. The reaction at this electrode is:  $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$  (reduction)

Recall that  $1 \text{ amp} \cdot \text{sec} = 1 \text{ Coulomb}$  and  $96,500 \text{ Coulombs} = 1 \text{ mole } \text{e}^-$ 's (Faraday's constant).

If a cell is run for 200. seconds with a current of 0.250 amps, how many grams of  $\text{Ag}^0$  will be deposited?

$$200 \text{ s} \times 0.250 \text{ amps} \times \frac{1 \text{ C}}{1 \text{ amp} \cdot \text{s}} \times \frac{1 \text{ mol } \text{e}^-}{96,500 \text{ C}} \times \frac{1 \text{ mol } \text{Ag}}{1 \text{ mol } \text{e}^-} \times \frac{107.87 \text{ g } \text{Ag}}{1 \text{ mol } \text{Ag}} = \mathbf{0.056 \text{ g } \text{Ag}}$$