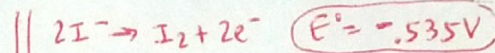
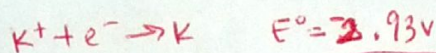


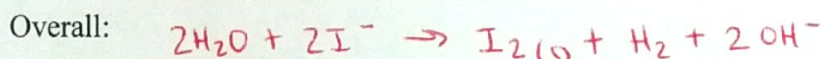
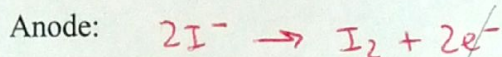
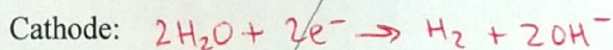
18 • Electrochemistry**ELECTROLYSIS WORKSHEET**

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 (Cl_2 and 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).
(reduction occurs at cathode)
- Write the change that water goes through at the (-) electrode. $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$
- 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).
(oxidation = anode)
- Write the change that water goes through at the (+) electrode. $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$
- We will perform this electrolysis using an aqueous solution of sodium sulfate.
Copy the reduction equation for Na^+ from the chart: $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$ $E^\circ = -2.714$ volts
- Both the Na^+ and H_2O will be near the (-) electrode. Which chemical is more likely to be reduced? H_2O
 $E^\circ = -2.714\text{V}$ vs. $E^\circ = -0.828\text{V}$
- Both the SO_4^{2-} and H_2O will be near the (+) electrode. Which chemical will be oxidized? H_2O
Why is there no equation in the chart with SO_4^{2-} on the product side? SO_4^{2-} is as oxidized as it can be; cannot be further oxidized.



11. In the electrolysis of KI(aq) $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E^\circ = -1.23V$
 Both the K^+ and H_2O will be near the (-) electrode. Which chemical is more likely to be reduced? H_2O
 Both the I^- and H_2O will be near the (+) electrode. Which chemical is more likely to be oxidized? I^-
 Write the reactions at each electrode and the overall reaction:

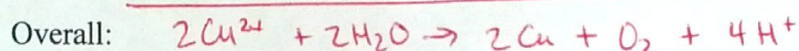
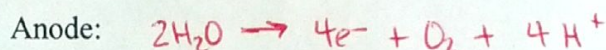
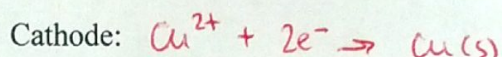


12. In the electrolysis of $CuSO_4(aq)$

Both the Cu^{2+} and H_2O will be near the (-) electrode. Which chemical will be reduced? Cu^{2+}

Both the SO_4^{2-} and H_2O will be near the (+) electrode. Which chemical will be oxidized? H_2O (see #10)

Write the reactions at each electrode and the overall reaction:



13. Silver plating occurs when electrolysis of a Ag_2SO_4 solution is used because silver metal is formed at the Cathode (cathode/anode).

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

Recall that 1 amp·sec = 1 Coulomb and 96,500 Coulombs = 1 mole e^- 's (Faraday's constant).

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

$$200s \times \frac{0.250 \text{ coulomb}}{s} \times \frac{1 \text{ mole } e^-}{96500 \text{ coulomb}} \times \frac{1 \text{ mole } Ag}{1 \text{ mole } e^-} \times \frac{107.87g}{1 \text{ mole } Ag} = 0.0559g \text{ Ag}$$

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 X^{3+} (X^+ , X^{2+} , X^{3+} , X^{4+})

$$2 \text{ hr} \times \frac{3600s}{\text{hr}} \times \frac{10.0 \text{ coulomb}}{s} \times \frac{1 \text{ mole } e^-}{96500 \text{ Coulomb}} \times \frac{1 \text{ mole } X}{? \text{ mole } e^-} = 0.250 \text{ mole } X$$

$0.746 \text{ mole } e^-$

$\frac{0.746}{x} = 0.250$
 $x = 2.984$
 $x = 3$

$x = 2.984$
 $x = 3$

15. Solutions of Ag^+ , Cu^{2+} , Fe^{3+} and Ti^{4+} are electrolyzed with a constant current until 0.10 mol of metal is deposited. Which will require the greatest length of time? Ti^{4+}

Ti^{4+} requires more e^- 's $\therefore$ it requires more time.

(You get same # of e^- 's per unit time).