Dougherty Valley HS AP Chemistry
Electrochemistry – Electrolysis

WORKSHEET #5 KEY

Name: **Date:** **Period:** **Seat #:**

Show all work

Standard Reduction Potential	E° (volts)
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.535
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.337
$SO_4^{2-}(aq) + 4\,H^+(aq) + 2e^- \rightarrow SO_2(g) + 2\,H_2O$	+0.20
$2\,H^+(aq) + 2\,e^- \rightarrow H_2(g)$ (reference electrode)	0.00
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.828
$Na^+(aq) + e^- \rightarrow Na(s)$	-2.714
$K^+(aq) + e^- \rightarrow K(s)$	-2.93

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 **OXDIZING 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. **$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-$**

6. In an electrochemical cell, the (+) electrode is positive because is has **TOO FEW** (too many/too few) electrons. Chemicals that come into contact with the (+) electrode will **LOSE** (gain/lose) electrons and be **OXDIZED** (oxidized/reduced). The (+) electrode in electrolysis is called the **ANODE** (cathode/anode).

7. Write the change that water goes through at the (+) electrode. **$2H_2O(l) \rightarrow 4e^- + 4H^+(aq) + O_2(g)$**

8. Add these two reactions together (make certain the electrons cancel) and write the overall reaction for the electrolysis of water.
$2H_2O(l) \rightarrow H_2(g) + O_2(g)$

9. We will perform this electrolysis using an aqueous solution of sodium sulfate. Both the Na^+ and H_2O will be near the (–) electrode. Which chemical is more likely to be reduced? **H_2O,** $\boldsymbol{E^\circ_{Na^+} = -2.714\,V, E^\circ_{H_2o} = -0.828, more + is\ reduced}$

10. Both the SO_4^{2-} and H_2O will be near the (+) electrode. Which chemical will be oxidized? **H_2O**
* SO_4^{2-} cannot be oxidized therefore H_2O must be

<table>
<tr><td colspan="2">[11] In the electrolysis of KI(aq)
Both the K^+ and H_2O will be near the (–) electrode. Which chemical is more likely to be reduced? H_2O, look at chart above, E°
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:</td></tr>
<tr><td>Cathode
$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$</td><td>Anode
$2I^- \rightarrow I_2 + 2e^-$</td></tr>
<tr><td colspan="2">Overall
$2H_2O + 2I^- \rightarrow H_2 + 2OH^- + I_2$</td></tr>
</table>

[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**
Write the reactions at each electrode and the overall reaction:

Cathode	Anode
$2Cu^{2+} + 4e^- \rightarrow 2Cu$	$2H_2O(l) \rightarrow 4e^- + 4H^+(aq) + O_2(g)$

Overall

$$2Cu^{2+} + 2H_2O \rightarrow 2Cu + 4H^+(aq) + O_2(g)$$

[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: **$Ag^+ + e^- \rightarrow Ag$** (reduction)

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° will be deposited?

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