Name KEY

Electrochemistry

ELECTROLYSIS WORKSHEET

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 oxidizing 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 is 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. $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

6. In an electrolytic 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 oxidized (oxidized/reduced). The (+) electrode in electrolysis is called the anode (cathode/anode).

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

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

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

$E° = -2.714V \rightarrow E° = -0.828V$

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 so H_2O must be oxidized.

$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

$4H_2O + 4e^- \rightarrow 2H_2 + 4OH^-$

$6H_2O + 4e^- \rightarrow O_2 + 4H^+ + 4e^- + 2H_2 + 4OH^-$

$-4H_2O$

$2H_2O \rightarrow O_2 + 2H_2$

11. In the electrolysis of KI(aq) → $E^\circ = -0.828V$
Both the K^+ ($E^\circ = -2.93V$) and H_2O will be near the (-) electrode. Which chemical is more likely to be reduced? H_2O
Both the I^- ($E^\circ = +0.535V$) and H_2O ($E^\circ = +1.23V$) will be near the (+) electrode. Which chemical is more likely to be oxidized? I^-
Write the reactions at each electrode and the overall reaction:

Cathode: $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Anode: $2I^- \rightarrow I_2 + 2e^-$

Overall: $2H_2O + 2I^- \rightarrow I_2 + H_2 + 2OH^-$

12. In the electrolysis of $CuSO_4$(aq) ($E^\circ = +0.33V$) → $E^\circ = -0.820V$
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: $Cu^{2+} + 2e^- \rightarrow Cu^\circ$

Anode: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

Overall: $2Cu^{2+} + 2H_2O \rightarrow O_2 + 4H^+ + 2Cu$

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?

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

$$= \boxed{0.056\,g\,Ag}$$

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+})

$$10.0\,amp \cdot 2.00\,h \cdot \frac{60\,min}{1\,h} \cdot \frac{60\,s}{1\,min} \cdot \frac{1\,C}{1\,amp} \cdot \frac{1\,mol\,e^-}{96{,}500\,C} = 0.746\,mol\,e^-$$

$$\frac{0.746\,mol\,e^-}{0.250\,mol} = \sim 3\,mol\,e^- \quad \boxed{X^{3+}}$$

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+}