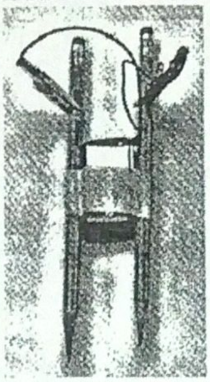
Culver City H.S. • AP Chemistry

Name KEY
Period ___ Date ___/___/___

# 18 • Electrochemistry

## ELECTROLYSIS LABETTE

**Introduction:**



Electricity can be used to cause a chemical change. We will try this using a pencil electrolysis apparatus. Notice that it is two pencils taped together with a 9-volt battery. The battery top has wires that connect to the graphite of each pencil. Remember, $C_{(graphite)}$ is a conductor due to the extended pi bond (chicken-wire shaped) above and below the $sp^2$ hybridized carbon atoms. The carbon has another benefit in that they are relatively inert and will not react with the solutions we will use.

Pencil "lead" itself is interesting stuff and you can read more about it here:
`http://pubs.acs.org/cen/whatstuff/stuff/7942sci4.html`

1. Get a clean Petri dish and put in some distilled $H_2O$. Place the tips of your pencil apparatus into the distilled water and make observations. No change

2. You probably notice that nothing happens because distilled $H_2O$ is a very poor conductor of electricity. Add a pinch of $Na_2SO_4$ to your water to increase the conductivity. Also, add a few drops of universal indicator. (Don't use the pencil apparatus to mix the solution.) You should see ~~a green~~ orange color if your water is neutral.

   Place the tips of your pencil apparatus into the distilled water and make observations.
   Purple color at one electrode ; slight pink at other electrode

3. At one pencil tip, $H^+$ ions are being formed. At the other pencil tip, $OH^-$ ions are being formed. We need to explain this observation. Let's consider the negative electrode first.
   Water surrounds the negative electrode and $Na^+$ ($Na^+$ | $SO_4^{2-}$) ions will be attracted to the negative electrode. The negative electrode is negative because it has too many (too many | too few) electrons. It will force a chemical to gain (gain | lose) electrons which is called reduction (oxidation | reduction).

| | Standard Reduction Potential | E° (volts) | E° oxidation | |
|---|---|---|---|---|
| (reverse) oxidation of $H_2O$ → | $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ | +1.23 | −1.23 V | > possible oxidation |
| | $SO_4^{2-}(aq) + 4\,H^+(aq) + 2e^- \rightarrow SO_2(g) + 2\,H_2O$ | +0.20 | | |
| reduction of $H_2O$ → | $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ | -0.828 | | > possible reduction |
| | $Na^+(aq) + e^- \rightarrow Na(s)$ | -2.714 | | |

This chart shows various reduction reactions with the most likely to be reduced higher in the chart. So, for our two choices of reductions at the negative electrode, write the chemical reaction that occurs:

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

The color at the negative electrode is purple due to the presence of $OH^-$ ions. Since reduction is occurring at the negative electrode, this is the cathode (anode | cathode).

4. At the p[illegible] electrode, $H_2O$ is oxidized. The S in $SO_4^{2-}$ (+6) is already in the most oxidized state it can be in, so [illegible]es not change. Use the chart to find the oxidation of $H_2O$. Since this is a chart of reductions, you must look at the equations backwards. Find the reaction in which the $H_2O$ is oxidized.

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

5. In this situation, water is both oxidized and reduced. Write the two equations and add them together. You will need to double one of the reactions to cancel the e⁻'s. Write the net reaction.

Oxidation: $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$   $E° = -1.23V$

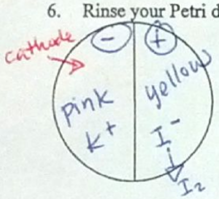
Reduction: $2\times(2H_2O + 2e^- \rightarrow H_2 + 2OH^-)$   $E° = -.828V$

Net Reaction: $2\cancel{6}H_2O \rightarrow O_2 + \underbrace{4H^+ + 4OH^-}_{4H_2O} + 2H_2$   $E° = -2.058V$

Remove the pencil apparatus and mix the solution. Observations: $\boxed{2H_2O \rightarrow O_2 + 2H_2}$ → Color goes back to orange

Explain: $H^+ + OH^-$ ions mix to form $H_2O$

6. Rinse your Petri dish and again add distilled water. Using a spatula, add a "pinch" of KI to the water and dissolve using a clean glass stirring rod. Add a few drops of phenolphthalein indicator solution. Place a strip of index card down the center of the dish to separate the dish into two areas. (It doesn't have to be at all perfect.)




Now, touch the pencil apparatus to the solution, one tip on each side of the card. Hold it still and make observations. one side is pink; the other side is yellow.

In this solution, you have $K^+$, $I^-$, and $H_2O$. It is now your task to determine the two possible reactions at the cathode. Write the two possible reductions that occur. Use the reduction potential chart to determine which of the two will actually occur.

| Reduction | E° | Which will occur? |
|---|---|---|
| $K^+ + e^- \rightarrow K$ | -2.93 | NO |
| (*) $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ | -.828 | Yes (pink = basic) |

larger E° so larger reduction potential.

| 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 |
| $SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow SO_2(g) + 2H_2O$ | +0.20 |
| $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (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 |

7. It is a little more tricky to determine which oxidation will occur at the anode[illegible] In this solution, you have $K^+$, $I^-$, and $H_2O$. Determine the two possible oxid[illegible] You will consider the reactions on the chart below written in **reverse**. Use the **reduction** potential chart to determine which of the two will actually occur by **changing the sign** of the reduction potentials. You are now listing the **oxidation potentials**. The higher the E° value, the more likely the chemical is to be oxidized.

| Oxidation | E° | Which will occur? |
|---|---|---|
| $2I^- \rightarrow I_2 + 2e^-$ | -.535 | Yes |
| $2H_2O \rightarrow 4e^- + 4H^+ + O_2$ | +1.23 | NO |

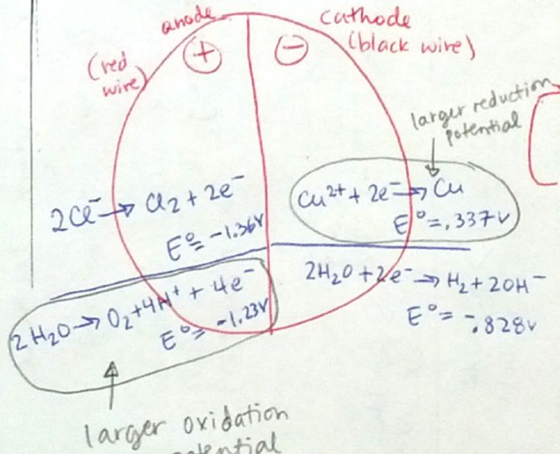
larger oxidation potential (E°)

| 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 |
| $SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow SO_2(g) + 2H_2O$ | +0.20 |
| $2H^+(aq) + 2e^- \rightarrow H_2(g)$ (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 |

8. Link your equations with the observations.

| Electrode: | What is observed? | Explain (write the equation) |
|---|---|---|
| anode | yellow color = $I_2$ | $2I^- \rightarrow I_2 + 2e^-$ |
| cathode | pink color = basic solution $OH^-$ | $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ |

9. What do you predict would happen with a solution of $CuCl_2$? What reactions would occur? What would you observe? Rinse your Petri dish and try it. $Cu^{2+}$, $Cl^-$, $H_2O$




larger reduction potential

$Cl_2 + 2e^- \rightarrow 2Cl^-$   $E° = 1.36V$

$Cu^{2+} + 2e^- \rightarrow Cu$   $E° = .337V$

larger E° is the anode

$2Cu^{2+} + \cancel{4e^-} \rightarrow 2Cu$

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

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

overall rxn.