EXPERIMENT 19

Corrosion and Electrolytic Cells

CORROSION OF IRON

Corrosion is a naturally occurring redox process that oxidizes metals to their oxides and/or sulfides. In Part A we will be focusing primarily on the corrosion of iron. Iron undergoes a two-step oxidation process as follows:

Step 1: \( \text{Fe (s)} \rightarrow \text{Fe}^{2+} + 2 \text{e}^- \)
Step 2: \( \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^- \)

The final oxidation product, iron (III), then combines with oxygen and water to form iron (III) oxide, or "rust".

In today's experiment you will determine which metals enhance the corrosion of iron and which inhibit the corrosion. You will test this by placing iron nails in contact with different metals and testing how much corrosion (if any) of the iron occurs. Since the corrosion process typically takes longer than your scheduled lab time, you will use color indicators to determine whether the iron is or is not corroding and where the oxygen is being reduced.

You will identify the regions in which iron is being oxidized (the anode) by the appearance of a blue color, caused by the reaction

\[
3\text{Fe}^{2+} (aq) + 2\text{Fe(CN)}_6^{3-} (aq) \rightarrow \text{Fe}_3[\text{Fe(CN)}_6]_2 (s)
\]

In this reaction, the \( \text{Fe}^{2+} \) ions are colorless ions in solution formed by the oxidation of iron. The \( \text{Fe(CN)}_6^{3-} \) will be added to your solution in the form of \( \text{K}_3[\text{Fe(CN)}_6] \) solid. The final product produced in the above reaction, \( \text{Fe}_3[\text{Fe(CN)}_6]_2 (s) \), is blue in color and indicates that iron was oxidized from \( \text{Fe (s)} \) to \( \text{Fe}^{2+} (aq) \).

You will be able to identify where oxygen is being reduced (the cathode) by the appearance of a pink color. Oxygen is first reduced to the -2 oxidation state in the form of hydroxide ions by the reaction

\[
\text{O}_2 (g) + 2\text{H}_2\text{O} (l) + 4\text{e}^- \rightarrow 4\text{OH}^- (aq)
\]

The \( \text{OH}^- \) ions then react with the colorless phenolphthalein indicator and turn the phenolphthalein pink, which indicates that oxygen was reduced from \( \text{O}_2 (g) \) to \( \text{OH}^- (aq) \).

To summarize, if any blue color appears, then you know iron was oxidized from \( \text{Fe} \) to \( \text{Fe}^{2+} \). If any pink color appears, then you know oxygen was reduced from \( \text{O}_2 \) to \( \text{OH}^- \). Unfortunately, these are the only two indicators we have in solution. Other oxidation reactions are possible; however, there is no indicator available to detect their presence. By making careful observations of all the nail systems today, you should be able to deduce what other reactions may have occurred even if there was an absence of one of the colors.
ELECTROLYTIC CELLS

Electrolytic cells use batteries or electricity to power reactions that are not normally spontaneous. One use of electrolytic cells is for electroplating, the process of coating the surface of a metal object with a layer of a different metal through electrochemical means. In Part B of the experiment, an electrolytic cell will be used to plate nickel onto copper. The electrolytic cell will be set up as shown below.

In this experiment the anode is a strip of nickel metal and the cathode is a strip of copper metal. The cathode is hooked up to the negative pole of the power supply and the anode is hooked up to the positive pole of the power supply. As the cell operates the electrons are forced through the external circuit from the anode to cathode by the power supply. The following equations describe the reactions in the cell:

Anode: \( \text{Ni} (s) \rightarrow \text{Ni}^{2+} (aq) + 2e^- \)

Cathode: \( \text{Ni}^{2+} (aq) + 2e^- \rightarrow \text{Ni} (s) \)

Nickel ions form at the anode, migrate through the solution, and deposit at the cathode as nickel solid. As the cell runs, the nickel anode dissolves into nickel ions and electrons and nickel plates onto the copper cathode.

In order to calculate the expected amount of nickel metal plated onto the copper, Faraday’s law must be used. Faraday’s Law establishes a relationship between the quantity of electricity (in coulombs, C) and the quantity of electrons (in moles). Faraday’s Law states that \( 96,485 \text{ C} = 1 \text{ mole } e^- \). The number of coulombs provided by an electrolytic cell is determined by its amperage and the number of seconds it operates: amperes x seconds = coulombs.

For example, if an electrolytic cell that plates chromium was operated for 2.0 minutes at 0.25 amperes, the coulombs = \( 120 \text{ s} \times 0.25 \text{ amperes} = 30. \text{ coulombs} \). Then, Faraday’s Law can be used to determine the mass of chromium plated as follows:

\[
30.\text{C} \times \frac{1 \text{ mol } e^-}{96,485\text{C}} \times \frac{1 \text{ mol Cr}}{3 \text{ mol e}^-} \times \frac{52.0 \text{ g Cr}}{\text{mol Cr}} = 0.0054 \text{g Cr}
\]

The amount of actual metal plated compared to the theoretical (stoichiometric) mass plated that is calculated by calculating the current efficiency, as shown by the following equation:

\[
\text{Current Efficiency} = \frac{\text{Actual Mass}}{\text{Theoretical Mass}} \times 100\%
\]
PROCEDURE

1. Students will work individually for this experiment. Except for the laboratory handout, remove all books, purses, and such items from the laboratory bench top, and placed them in the storage area by the front door. For laboratory experiments you should be wearing closed-toe shoes. Tie back long hair, and do not wear long, dangling jewelry or clothes with loose and baggy sleeves. Open you lab locker. Put on your safety goggles, your lab coat, and gloves.

PART A – CORROSION

2. Heat 100 mL of deionized water to boiling in a 250-mL beaker using your hot plate. While the water is heating, prepare 5 sandpapered nails as follows (see figure below):

- Keep one nail intact
- Score a nail with a metal file in several places
- Pierce a small piece of zinc with a nail
- Pierce a small piece of copper with a nail
- Pierce a small piece of copper with a nail, wrap one piece of copper wire around the head of a nail, and wrap the other end around a piece of zinc

![Diagram of nails](image)

3. Once the water has begun to boil, take the beaker off the hot plate and stir in 0.7 g of agar in four portions. Don't add the agar all at once, it won't dissolve well. After the agar dissolves, stir in 0.5 g of NaCl, 0.07 g of K₃[Fe(CN)₆] and 15 drops of phenolphthalein solution. Stir until all of the solids dissolve.

4. Cool the gel-indicator solution until it is just lukewarm or begins to thicken. Arrange your 5 nails in 5 large test tubes.

5. Pour the cooled agar mixture over the nails so as to cover them, except the plain nail; leave part of it out of the agar. Watch the tubes over a period of 60 minutes. Note the appearance of any colors around the nail assemblies. Record your observations in your Data Table. Once you have recorded your observations, scrape the agar and nails into their appropriate waste containers.

6. For each of the nail set-ups, you will need to determine in which set-up the iron is corroding and which systems there is a metal other than iron corroding. Determine the reactions at both the anode and cathode for each nail.

7. Clean and wipe dry your laboratory work area and all apparatus. When you have completed your lab report have the instructor inspect your working area. Once your working area has been checked your lab report can then be turned in to the instructor.
PART B – ELECTROLYTIC CELL – DEMONSTRATION

8. Immerse a copper strip in 6 M hydrochloric acid for about 3 minutes, then rinse it with deionized water and dry it. Weigh the copper strip and record its mass in your Data Table.

9. Set up the electrolytic cell, attaching the copper strip to the negative pole of the power supply.

10. With a stopwatch ready, turn on the power supply and begin timing. Record the current being passed through the electrolytic cell in your Data Table.

11. After the cell has run for about 10 minutes, turn off the power and record the exact amount of time the cell operated in your Data Table.

12. Remove the copper strip and dry it. Determine the mass of the copper strip and record it in your Data Table. Determine the experimental mass of nickel plated on the copper strip.

13. Calculate the theoretical mass of nickel plated and the current efficiency of the cell.
## EXPERIMENT 19 LAB REPORT

Name: ___________________________________________  Student Lab Score: ______________

Date/Lab Start Time: _____________________________  Lab Station Number: ______________

### DATA TABLE

#### PART A

<table>
<thead>
<tr>
<th>Nail</th>
<th>Color</th>
<th>Anode Metal</th>
<th>Cathode Metal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Plain Nail</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Scored Nail</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Nail + Zn</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Nail + Cu</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Nail + Cu + Zn</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Draw pictures of the 5 different nails after the one-hour time period in the five boxes

<table>
<thead>
<tr>
<th>Plain Nail</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Scored Nail</td>
<td></td>
</tr>
<tr>
<td>Nail + Zn</td>
<td></td>
</tr>
<tr>
<td>Nail + Cu</td>
<td></td>
</tr>
<tr>
<td>Nail + Cu + Zn</td>
<td></td>
</tr>
</tbody>
</table>
## PART B

<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial Mass of Cathode</td>
<td>.</td>
<td>g</td>
</tr>
<tr>
<td>Current</td>
<td>.</td>
<td>A</td>
</tr>
<tr>
<td>Time the Cell Operated</td>
<td>.</td>
<td>s</td>
</tr>
<tr>
<td>Final Mass of Cathode</td>
<td>.</td>
<td>g</td>
</tr>
<tr>
<td>1 Experimental Mass of Plated Nickel</td>
<td>.</td>
<td>g</td>
</tr>
<tr>
<td>2 Theoretical Mass of Plated Nickel</td>
<td>.</td>
<td>g</td>
</tr>
<tr>
<td>3 Current Efficiency</td>
<td>.</td>
<td>%</td>
</tr>
</tbody>
</table>

## CALCULATIONS

1.

2.

3.
1. From Part A, write the oxidation half-reaction and the reduction half-reaction for each of the nails.

Plain Nail:
Oxidation: _____________________________  Reduction: _____________________________

Scored Nail:
Oxidation: _____________________________  Reduction: _____________________________

Nail + Zn:
Oxidation: _____________________________  Reduction: _____________________________

Nail + Cu:
Oxidation: _____________________________  Reduction: _____________________________

Nail + Cu + Zn:
Oxidation: _____________________________  Reduction: _____________________________

2. Which metal prevented the iron from corroding because it underwent corrosion instead?
__________________________________________________________________________________

3. Why did the metal you answered for question 2 prevent the iron from corroding by undergoing corrosion itself, but the other metal did not?
__________________________________________________________________________________
__________________________________________________________________________________
__________________________________________________________________________________
__________________________________________________________________________________

4. A “tin can” is an iron can that has been coated with tin. How does it prevent the iron from corroding? Why should you never buy food products in tin coated iron cans if they have a sharp dent?
__________________________________________________________________________________
__________________________________________________________________________________
__________________________________________________________________________________
5. An electrolytic cell that plates zinc from a zinc ion solution operates for 3 minutes 35 seconds at 325 milliamperes.

(a) Calculate the theoretical mass, in milligrams, of zinc deposited.

(b) The actual mass of zinc deposited was 20.5 mg. Calculate the current efficiency.

6. An electrolytic cell plates platinum from a platinum (IV) sulfate solution.

(a) Calculate the current needed to electroplate 40.0 g of platinum in 3.00 hours.

(b) Calculate the current needed in (a) if the current efficiency is 93.0%