3.003 - Principles of Engineering Practice

Lab 4: Sacrificial Anodes

Team eXtreme

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March 31, 2008

Due: April 1, 2008
**Background:**

Corrosion is a trillion-dollar problem worldwide that has also resulted in unacceptable human casualties. As in the case of the Mississippi River Bridge and the Minneapolis Bridge, collapse was in part due to corrosion. These collapses did not only put a sizable hole in the wallets of these states, but they also resulted in death and injury at the hands of the innocent. It has been declared a priority that similar bridges, like the Massachusetts Bridge, be kept under constant checkups to ensure that collapse would not ensue. Corrosion is a structural problem that must be taken into consideration when constructing things such as bridges.

Corrosion involves repeated chemical reactions in which the original material reacts with a chemical agent to form a new compound, or to dissolve into the chemical agent. In both cases, charge is transferred. We can divide corrosion into three categories: high-temperature gaseous (dry) corrosion (e.g. oxidation, sulfidation); aqueous (electrochemical) corrosion; and biological corrosion (attack by microbes). Aqueous corrosion of metals is controlled by the transfer of electrical current in the oxidation of the corroding species. This process is known as galvanic (anodic) corrosion. The driving force for galvanic corrosion is given by the cell potential (emf) of the corrosion reaction.

When a metal is negative (relative to another), it will act as the anode in the electrochemical cell. Corrosion in aqueous solutions is an electrochemical reaction that involves charge transfer, either oxidation (involving the loss of electrons), or reduction (involving the gain of electrons). This is a metal dissolution process, with metal atoms being converted to metal ions and going into solution. The products of the reaction depend on the pH of the aqueous solution, a measure of the hydrogen ion concentration. At the surface of the metal, neutral metal atoms are oxidized to positive metal ions in what is known as an oxidation or anodic reaction:

For example:

\[ \text{Fe} \rightarrow \text{Fe}^{2+} + 2\text{e}^- \quad \text{or} \quad \text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^- \]

The corresponding reduction reaction also depends on the nature of the environment, particularly the pH. Some typical reduction, or cathodic, reactions are shown below:

- **Hydrogen:** \[ 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2 \] (acidic solutions)
- **Oxygen:** \[ \text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O} \] (acidic solutions)
- **Oxygen (dissolved):** \[ +2\text{H}_2\text{O} + 4\text{e}^- \rightarrow \text{OH}^- \] (neutral/basic solutions)
- **Water:** \[ 2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + \text{OH}^- \] (neutral/basic solutions)

Corrosion measurements are quantified by constructing a corrosion cell (an electrochemical cell, in which the oxidation and reduction reactions generally take place at separate electrodes in the cell), which develop an electrical potential difference.
between them. The (open circuit) cell potential is a measure of the tendency of a metal to corrode. When the two electrodes are in electrical contact, an electrical circuit is formed in which electron current flows through the electrical connection between the electrodes and a corresponding ion current flows through the electrolyte between the electrodes. The current flow is a measure of the corrosion taking place at the anode.

The maximum amount of electrical energy (or work done) that can be delivered by an electrochemical cell is \( nF E \), where \( n \) is the number of moles of electrons exchanged in an electrochemical reaction, \( F \) is Faraday's constant (96,485 C /mol), and \( E \) is the cell potential (in volts) for the cell in a given state. For cell conditions in a standard state, the potential energy is \( nF E_0 \), where \( E_0 \) represents the standard-state electrochemical cell potential.

These parameters determine the relative tendencies of metal reactivity (corrosion). The corrosion rate is proportional to corrosion current. Thus, a definite way of knowing if a metal in a given environment is corroding is to perform a current measurement.

In the first part of our experiment, we determine the relative emf of different metals by measuring the current generated in an electrochemical cell. This knowledge enables us to determine whether a metal is active (anodic) or noble (cathodic) in a salt water environment. One method for mitigation of corrosion is to couple the structure we want to preserve with a metal that is more negative on the emf scale. By this way, the anode will be "sacrificed" for the cathode with respect to corrosion. This is a useful tool in the preservation of metals that comprise such things as bridges.

In the second part of our experiment, we explore different designs for corrosion protection. By impressing a voltage from the outside to balance the reaction toward a "sacrificial" anode, we can determine which metals are most effective at protecting our cathode. By using metal that is of lower cost, but not lower in the emf series, we can create a economical and practical anode that would save millions of dollars normally whisked away buy corrosion.

Ultimately, we will show that sacrificial anodes can be configured to protect materials that would otherwise be subjected to corrosion currents.

**Objectives:**
To understand the role of corrosion in the sustainability of societal infrastructure
To explore the electrochemical nature of aqueous corrosion
To design sacrificial anode for mitigation of corrosion of bridges

**Hypothesis:** Aluminum will be the best sacrificial anode for bridge building because of its metallic properties and its cost effectiveness.

**Materials:**
- Ag, Al, Au, Co, Cu, Fe, Ni, Pb, Pt, Zn and Brass foils
- NaCl powder
- Beakers
- Digital multimeter
Galvanic Cell Procedure:
1. Mix a 3% NaCl aqueous solution (similar to seawater) in a beaker.
2. Connect a copper (Cu) foil to one end of a multimeter and suspend the foil in the aqueous solution.
3. Connect one of the metal foils to the multimeter.
4. Suspend both the copper (Cu) foil and the metal foil in solution.
5. Apply a voltage to the system.
6. Measure and record the difference in electric potential.
7. Replace the metal foil and repeat for each type of metal foil, remembering to thoroughly clean each specimen before returning it to the box.
8. Connect an iron (Fe) foil to the multimeter.
9. Suspend the copper (Cu) and iron (Fe) foils in the multimeter.
10. Apply a voltage.
11. Measure the voltage every 30 seconds for five minutes to find the time dependency of voltage.
12. Record.
13. Measure the current every 30 seconds for five minutes to find the time dependency of current.
14. Replace the solution and measure the current after five seconds.
15. Record.
16. Repeat for accuracy.

Sacrificial Anode Procedure:
1. Connect one iron (Fe) foil and one copper (Cu) foil to a multimeter. The iron will serve as the anode and the copper will serve as the cathode.
2. Suspend these foils in 3% NaCl aqueous solution.
3. Apply a voltage to the system.
4. Record the electrical potential.
5. Connect an additional iron (Fe) foil, along with its own DC power supply to the multimeter.
6. Apply a voltage to the system.
7. Record the voltage required from the DC power supply to diminish the current flowing through the initial iron foil.
8. Repeat for accuracy.

In the first part of the sacrificial anode lab, the iron (Fe) is being corroded.
Upon edition of the second iron (Fe), both are equally corroded.

Now, with the addition of the DC power supply to the second iron (Fe), the current completely bypasses the original iron (Fe) foil. In such a way, the original iron (Fe) foil is saved from corrosion.

**Results:**

**Table 1 Galvanic Cell Potential:**

<table>
<thead>
<tr>
<th>Metal</th>
<th>Voltage (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pb</td>
<td>0.412</td>
</tr>
<tr>
<td>Pt</td>
<td>-0.281</td>
</tr>
<tr>
<td>Zn</td>
<td>0.806</td>
</tr>
<tr>
<td>Zn/Cu</td>
<td>0.032</td>
</tr>
<tr>
<td>Cu</td>
<td>0.022</td>
</tr>
<tr>
<td>Ag</td>
<td>-0.185</td>
</tr>
<tr>
<td>Al</td>
<td>0.641</td>
</tr>
<tr>
<td>Au</td>
<td>-0.270</td>
</tr>
<tr>
<td>Co</td>
<td>0.280</td>
</tr>
<tr>
<td>Fe</td>
<td>0.200</td>
</tr>
<tr>
<td>Ni</td>
<td>0.040</td>
</tr>
</tbody>
</table>

**Table 2 Timed-Trials of Galvanic Cell Potential and Current:**

<table>
<thead>
<tr>
<th>Time (sec)</th>
<th>Trial 1 Voltage (V)</th>
<th>Trial 2 Voltage (V)</th>
<th>Trial 3 Current (μA)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.082</td>
<td>0.069</td>
<td>0</td>
</tr>
<tr>
<td>30</td>
<td>0.160</td>
<td>0.182</td>
<td>66.9</td>
</tr>
<tr>
<td>60</td>
<td>0.202</td>
<td>0.221</td>
<td>77.7</td>
</tr>
<tr>
<td>90</td>
<td>0.238</td>
<td>0.243</td>
<td>85.2</td>
</tr>
<tr>
<td>120</td>
<td>0.261</td>
<td>0.258</td>
<td>89.2</td>
</tr>
<tr>
<td>150</td>
<td>0.275</td>
<td>0.267</td>
<td>91.8</td>
</tr>
</tbody>
</table>
Figure 1 Timed Cell Potentials:

<table>
<thead>
<tr>
<th>Time (sec)</th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Voltage (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>0.280</td>
<td>0.274</td>
<td>93.8</td>
</tr>
<tr>
<td>50</td>
<td>0.291</td>
<td>0.280</td>
<td>95.7</td>
</tr>
<tr>
<td>100</td>
<td>0.301</td>
<td>0.282</td>
<td>96.6</td>
</tr>
<tr>
<td>150</td>
<td>0.308</td>
<td>0.285</td>
<td>97.5</td>
</tr>
<tr>
<td>200</td>
<td>0.314</td>
<td>0.287</td>
<td>98.3</td>
</tr>
<tr>
<td>250</td>
<td>0.318</td>
<td>0.287</td>
<td>98.9</td>
</tr>
<tr>
<td>300</td>
<td>0.322</td>
<td>0.290</td>
<td>99.4</td>
</tr>
</tbody>
</table>

Figure 2 Timed Cell Current:
Sacrificial Anode voltage: -0.063 V with Fe

Table 3 Galvanic Cell Current:

<table>
<thead>
<tr>
<th>Trial</th>
<th>Current (μA)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>87</td>
</tr>
<tr>
<td>2</td>
<td>100</td>
</tr>
<tr>
<td>3</td>
<td>108</td>
</tr>
</tbody>
</table>

**Analysis:**

When we tested the materials for different galvanic cell potentials, we found that the potentials were positive or negative depending on the sign of the standard electrode potentials given in the lab. Cu/Zn (brass) was interesting because it is an alloy of two metals. The cell potential was between those of copper and zinc. Because we didn’t know the percentage composition of or if other metals were present in the brass, the exact relationship between cell potentials cannot be determined.

In the time trials of galvanic cell potential and current, we found that the cell reaction proceeded towards equilibrium. The plots of time vs. voltage/current displayed data that appeared to follow logarithmic curves. Because galvanic cells have a voltage according to the Nernst equation, displayed below, the fact that the voltage and current were logarithmic matched our predictions.

\[
E = E^0 - \frac{kT}{n e} \ln Q.
\]

**Nernst Equation:**
Where $E^0$ is the standard electrode potential, $k$ is Boltzmann’s constant, $T$ is temperature in Kelvin, $n$ is number of moles, $e$ is the number of electrons transferred in the cell, and $Q$ is the reaction quotient, $E$ gives the potential of the cell at a given point.

The voltage that we found for the sacrificial anode was -0.063V. Therefore, iron will not act as a good sacrificial anode for iron, but if a material had a galvanic potential less than the opposite of that voltage, it would successfully protect the object from corroding.

When we placed iron in the galvanic cell and ran current for five second intervals, we found that the current increased with the number of repetitions. The current increased with time because metal ions were placed in solution, allowing current to flow through the cell easier. This fit expectations, which was why we used short, five second intervals.

**Conclusion:**

In this lab we sought to determine the difference in electric potentials between electrode reactions, as well as the time dependency of voltage and current in the oxidation reaction between iron and copper. Using the knowledge thus gained, we set up a system including a sacrificial anode with the aim of designing such a system to be used in the real world to protect bridges from corrosion.

We hooked various electrodes up to a voltmeter, giving them five seconds to equilibrate before measuring their potential difference. We then used just iron and copper, and measured first voltage, then current, with respect to time.

To design a sacrificial anode, we included a third electrode in our Fe/Cu galvanic system, also of iron. By applying various voltages across the sacrificial anode and the copper cathode, we tried to minimize the galvanic voltage from the protected anode.

Both the current and potential difference between the iron and copper electrodes increased at a diminishing rate as time passed. After six minutes it was unclear whether or not they had reached a state of equilibrium or would continue to increase for far longer. Thus, these measurements are time dependent, most likely due to the accumulation of metal oxides on the electrodes negatively affecting their conductivity.

Our experiments with the sacrificial anode indicated that a voltage of -0.063 V was necessary to avoid corrosion on the protected anode. Therefore, any material with galvanic potential less than 0.063 less than that of iron would function. This list includes zinc, aluminum, and others.

There are several ways to avoid corrosion on bridges. We could build them out of non-metallic materials which do not corrode, like plastics or pykrete, a frozen blend of water and sawdust with high crush strength and low thermal conductivity (such an option would only be feasible in cold climates). Modern materials are being created with incredible strength and durability, at the price of cost-efficiency; it is always possible that new manufacturing techniques will render such materials feasible in the future.