ME 212 LABORATORY EXPERIMENT #5

CORROSION OF METALS AND ALLOYS

1. OBJECTIVE:

The primary aim of this final experiment in the Materials Science course is to introduce the phenomenon of corrosion to the students along with its basic reasons and working principles and to use a simple setup to demonstrate it. Possible ways of preventing corrosion are also briefly discussed.

2. CORROSION THEORY:

Corrosion can be defined as the process of surface deterioration of metals and related materials that happens when the metal loses electrons and becomes a positively charged ion as given by the general chemical reaction below:

\[ M^0 \rightarrow M^{m+} + me^- \]  \hspace{1cm} (I)

(Metal M loses m-many electrons and becomes a positively charged ion)

The metal ions created in this way either dissolve in the surrounding electrolyte (fluid containing ions) or react with non-metallic ions in the environment to form a surface deposit. One way or the other, the properties of the material surface are in this way altered—an almost always in a negative way—and therefore corrosion is identified as one of the failure mechanisms of materials.

Although all metals are subjected to the general oxidation reaction (I), the tendencies to ionize differ from metal to metal. When two metals are exposed to an aqueous environment and release electrons according to reaction (I), an electric potential is induced the standard value of which can be measured when the metals are in 1-molar solutions of their own salts at a predetermined temperature of 25 °C. (see Figure 1):

![Figure 1. Galvanic Cell consisting of Zn and Cu electrodes (image from library.tedankara.k12.tr)](image)
The setup seen in Figure 1 is referred to as a *Galvanic Cell*. Two metal rods referred to as electrodes are each immersed in solutions of their salts, the solutions being separated by a permeable membrane to allow passage of ions. The electrodes are then connected electrically and the resulting potential difference is measured using a voltmeter. In this way the “tendency of a metal to give its electrons away with respect to another metal” is measured quantitatively. The electrode which gives its electrons away and thus becomes partly dissolved in the solution (the electrode which corrodes) is commonly referred to as the *anode*. The metal which takes the electrons given by the anode and uses them for “binding” the positively charged ions in its solution is referred to as the *cathode*. This “binding” reaction in the cathodic side can be seen as the reverse of the general reaction (I), which results in the formation of an additional layer of metal on the cathode. In this way, the cathode becomes in a way “plated”. Observe that in Figure 1, the voltage difference induced by copper (Cu) and zinc (Zn) electrodes is measured to be 1.10 V. This measured value is also a measure of the amount of electrons flowing from the anode to the cathode, thus giving an idea about the *rate of corrosion*.

The voltage differences could be measured between all possible pairs of metals in the manner just described. However, the more common practice is to use the \( \text{H}_2 \rightarrow 2\text{H}^+ + 2e^- \) (II) reaction as the reference for the anode reaction. In this way, the so-called *half-cell electrode potentials* or the *EMF series* of each metal can be determined and tabulated. If the metal is found to be receiving electrons rather than giving them away when exposed to reaction (II) on the other side of the galvanic cell, it is assigned a *positive* EMF value. If the reverse is the case, it is assigned a *negative* value. One should not forget that the obtained values are true for temperature values of 25 °C and 1 M solutions. The EMF series or half-cell potentials of some of the most important metals used in engineering are tabulated below:

Metals having high electrode potentials are referred to as “noble” metals in the sense that they do not easily give away their electrons and corrode. That’s the main reason why metals such as gold or platinum are very expensive materials. Other metals, which have low electrode potentials such as Magnesium and Zink are very prone to giving their electrons away and corroding and are thus usually called “active”.

As an additional information, it should also be known that corrosion can also occur in a single metal that is not connected to any other metal. In this case, certain areas of the metal act as anodes while other areas act as cathodes. The corrosion process in this case may for example occur due to inhomogeneities in composition or microstructure in the material.

There are some major ways one can use to prevent corrosion in components built from metals / alloys. Perhaps the most common and easiest way of preventing corrosion is through the careful selection of materials once the corrosion environment is identified. Tables like Table 1 can be used in this respect. However, due to cost problems, the “best” material for a given situation can often not be used.
One of the most widely used and effective ways of corrosion prevention is cathodic protection. In this protection technique, another metal which is more active in the given environment than the metal to be protected is electrically connected to it. The protecting metal, in this way, supplies an electron to the metal to be protected for each electron it loses, thus preventing corrosion. Over time, the protecting metal, having given almost all of its electrons completely dissolves. Thus the protecting metals used in this process are often referred to as sacrificial anodes. Zinc and Magnesium are commonly used for that purpose because of their relatively low place in Table 1. The method of galvanizing, also a way of cathodic protection, involves the coating of steel by a thin Zink layer, which in turn protects steel cathodically. The galvanizing process is mainly applied on steel ship hulls. The figure below shows cathodic protection as it is applied to underground pipelines (image from www.dpla.water.ca.gov):
3. TASKS:

Please perform these tasks under the “Results and Discussion” part of your reports.

1. Think of a corrosion problem encountered in an engineering context. Briefly discuss it and mention possible ways of preventing it. (Not more than half a page)

2. Briefly explain and discuss the following terms: Pitting, Fretting and Stress Corrosion. Give figures if necessary. (You can draw them by hand)
REFERENCES