



A study of cathodic protection of underground metal structures and cathodic protection survey of pipe lines on campus of Montana State College
by Charles C Collins

A THESIS Submitted to the Graduate Faculty in partial fulfillment of the requirements for the degree of Master of Science in Electrical Engineering
Montana State University
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Abstract:

It is estimated that the cost of underground corrosion each year in the United States is approximately \$1,000,000,000. Practically all unprotected iron structures underground will eventually waste away. It is only since 1900 that the electrolytic nature of corrosion has become understood and any general recognition given to the corrosion problem.

Underground corrosion results from the tendency of certain metals to pass into solution as positive ions. If there exists both a metallic and electrolytic path between a pair of metallic surfaces called the anode and the cathode, then a galvanic current will probably flow. The anode wastes (rusts or corrodes) away, while the cathodic surface is protected. Galvanic currents arise on underground structures because of differences in kind or condition of metallic surfaces and surrounding soils. If there is a closed electric circuit and if there is some degree of moisture present in the soil, a current will flow underground from anode to cathode.

The prevention of corrosion consists mainly of applying a water-proof coating to an underground structure, and by lowering the potential of the pipe to a point where a current will flow from ground to the pipe. The latter is called cathodic protection. It is accomplished by impressing a voltage on the ground area through a ground bed made of scrap iron or carbon rods, with the negative terminal connected to the protected metallic surface. Cathodic protection may also be achieved by the use of magnesium anodes, which are buried in the ground after being connected to the metal structure by a length of conducting wire. The magnesium being more active, will cause a current to pass from the anode to the iron structure, which will protect the latter.

Soil resistivities and pipe-to-soil potentials are useful information for locating areas where conditions favorable to corrosion exist. A sensitive voltmeter and copper-sulfate half-cell are necessary for obtaining the potential measurements. The voltage between an underground pipe and the surrounding soil is obtained by connecting- the positive terminal of the voltmeter to the half-cell and placing the half-cell in contact with the ground. The - negative terminal of the voltmeter is connected to the pipe. Roughly, a voltmeter reading of less than 0.6 volt indicates a possibility of corrosion, and a reading between 0.6 and 0.85 volt indicates some degree of protection. A reading of more than 0.85 volt is commonly accepted as an indication of full cathodic protection to the underground iron structure at that point.

A survey of the Montana -State College campus indicates most of the underground pipe is in good condition, with a few localised areas where corrosion is taking place. A complete cathodic protection system using rectifier stations to furnish a protection voltage is considered to be rather expensive in comparison to corrosion losses taking place on the campus. A system of localized protection using magnesium anodes to furnish protection voltage in areas where corrosion is known to exist is recommended as being more economical and practical.

A STUDY OF CATHODIC PROTECTION OF UNDERGROUND METAL STRUCTURES AND
CATHODIC PROTECTION SURVEY OF PIPE LINES ON CAMPUS OF MONTANA STATE COLLEGE

By

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ABSTRACT

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INTRODUCTION

Corrosion of metallic structures causes a significant waste of national resources, the cost of which is estimated at approximately \$5,000,000,000 annually in the United States. Of this, \$1,000,000,000 is the estimated loss by corrosion of underground metallic structures. Practically all unprotected iron structures in underground locations will eventually waste away.

In 1823 Sir Humphrey Davy recognized the electrochemical nature of corrosion and upon his recommendation, zinc plates were tried on English warships for the purpose of protecting the copper on the hulls. This was probably the first cathodic protection system in history and while not successful, the fault lay with the marine growths on the ship bottoms rather than with any error in Davy's thinking. However nearly 100 years were to pass before there was any general recognition of the corrosion problem. Prior to the year 1900, accelerated corrosion was noted on some buried pipe lines near electrified street-railway tracks, and by 1910 the electrolytic nature of such corrosion was understood. With the subsequent rapid growth of buried cables, pipe lines etc., American industry became more concerned about the corrosion problem and the necessity for finding a means of control. Cathodic protection was first reported to have been used in 1922, but its effectiveness was not generally recognized until the early 1930's. Since then the use of cathodic protection as a means of reducing corrosion of underground (and underwater) metal structures has become widespread.

During this time the effectiveness of waterproof coatings for the prevention of corrosion was known, and used. For example, a bituminous coating was being applied to pipes as early as 1840 in this country by a



Fig. 1. The cost of corrosion. A 6" gas main no longer serviceable because of leaks, is being replaced by new pipe at left.

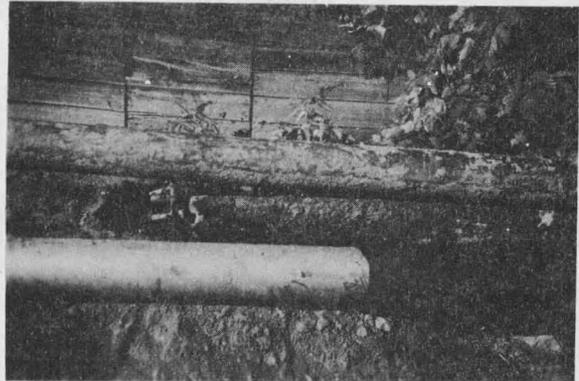


Fig. 2. Rusted pipe from project illustrated in Fig. 1, with new pipe at bottom of picture. Note waterproof wrapping (paper over asphalt) around new pipe, which is the primary means of preventing corrosion.

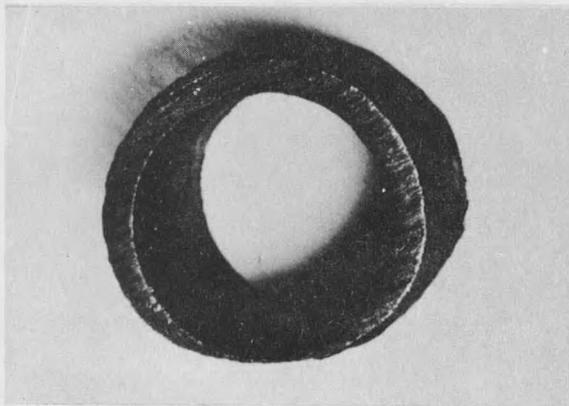


Fig. 3. Section from 1" gas line showing advanced corrosion (rusting) along lower side of pipe.

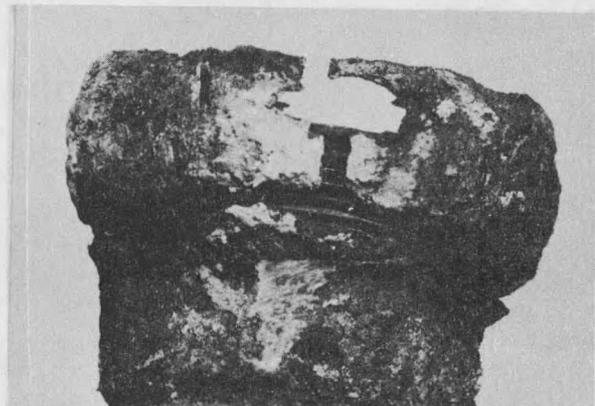


Fig. 4. 1 1/2" union corroded entirely through. Pipe and cut-off valve of same line no longer serviceable.

gentleman of the name of Angus Smith.¹ While coatings are still the primary means of extending the life of underground piping, it remained for the advent of cathodic protection to close the gap in the protection given by coatings, by providing a means for stopping corrosion at any point where a break existed in the coating. The two methods applied together, give the best economic means known today for the prevention of pipe line leaks, the inevitable result of rusting and pitting of underground metal surfaces.

¹Speller, F. N., Corrosion Causes and Prevention, New York; McGraw-Hill, 1935, p. 566.

THE ELECTROCHEMICAL NATURE OF CORROSION

From a practical standpoint the corrosion engineer is not so much concerned with the exact chemistry of corrosive action, as he is with the more immediate one of corrosion control. But to obtain an understanding of the electrical means for corrosion control it is necessary to set up a certain minimum of electrochemical theory.

If we recognize corrosion as an electrochemical phenomena, that is, by a corrosive area as being one where an electric current is flowing away from a metallic surface; then it would appear that the elimination of this corrosive tendency could be accomplished by the application of a reverse potential. This is precisely what a cathodic protection system does. Simply stated, it is an electrical circuit for applying a positive potential to the ground area adjacent to buried metal structures (or negative potential to the metal), so that any tendency for galvanic currents to flow from metal to soil will be smothered out by this reverse difference of potential applied between soil and metal. At least some degree of galvanic currents are almost certain to flow between or along bare metal surfaces underground, as the result of electrolytic action between the metal and contacting soil. As we shall see later, the metallic surfaces become electrodes in the electrolytic soil solutions, and the electrical phenomena is exactly the same as that in a wet cell.

The electrical nature of corrosion has been well established, at least as far as underground surfaces are concerned. In fact it is difficult to draw an exact dividing line between electrochemical and straight chemical reactions. The tarnishing by oxidation of metallic surfaces in air for example, has been connected with electrochemical processes, largely by the

work of C. Wagner in the early 1930's.¹ The theory developed by Wagner led to an equation to explain the tarnishing rate, and while there are observed deviations from the equation, Wagner did establish the fact that the passage of metal ions through the oxidation layer played a significant role.

A number of investigators have performed experiments which tend to establish the electrochemical nature of any corroding action occurring on a metallic surface immersed in a chemical solution. In this connection it is interesting to note that one theory suggested to explain the rusting of iron in soft water states that the iron becomes covered with a layer of magnetite (Fe_3O_4) which in turn is covered with a layer of hydrated ferric oxide $\text{Fe}(\text{OH})_3$.² But the magnetite and iron constitute a galvanic couple with iron being the anodic surface, that is; the iron is slowly carried away by a current flowing from the iron to the magnetite layer.

In order to understand the nature of the protection problem, some of the electrochemical aspects of underground corrosion will be explained. In a sense almost any chemical action can be regarded as electrochemical, since there is a transfer or displacement of electrons. However in this paper, an electrochemical action will be considered as a chemical action resulting from the flow of an electric current between spatially separated anodic and cathodic areas.

Basically, natural corrosion results from the tendency of refined metals to return to the compounds in which they exist in nature. After a metal such as iron, has been reduced and refined from its natural ores, it will have a tendency to rejoin the elements from which it was taken. An

¹Wiley, H. A., Corrosion Handbook, ed. by H. H. Uhlig, New York; Wiley & Sons, 1948, p. 16.

²Ibid., Evans, U. R., p. 7.

iron sample if placed in an electrolytic solution, will tend to pass into that solution as positive ions. This tendency is termed the solution pressure or potential, and it may be measured in an electrolyte containing its own salts, with reference to the hydrogen electrode. This forms the basis for the electromotive series, or sometimes called the electrode-potential series. The series actually represents the relative tendency of the respective metallic atoms to give up one or more valence electrons to form positive ions.

An electrolyte may be defined as a solute furnishing free ions in solution. Faraday investigated the conduction of electric currents in electrolytic solutions, theorized on the possibility of small particles being the "carriers," and even gave them the name ions. But it remained for Arrhenius to clarify the actual dissociation of the compound (always an electrovalent compound) into electrically charged particles, or ions. Although not all of the electrolyte may dissociate, there will be at least a partial dissociation (in a water solution) and the resulting ions will enter into the mechanics of corrosion.

As before stated any metal with corrosive properties will, when placed in an electrolytic solution, give off positive ions. These ions in turn will join negative oxygen or other ions, already present by reason of the dissociation properties of the solution. Hydrogen ions are displaced and a balance is soon reached in solution pressures and no further ions will leave the metal. Then if another piece of metal, or electrode, is placed in the solution and an outside electric potential impressed, with the positive connection to the first electrode, the transfer of ions will then be resumed and will continue until that electrode is entirely transferred to the solution, or the electric current stopped.

Another way to set up a current in an electrolytic solution is to use two dissimilar metals for electrodes. Most any two metals will do; it is only necessary that they exhibit different solution pressures by reason of their positions in the electromotive series. Then whichever electrode passes current into the solution will waste away. The process is illustrated in Fig. 5.

In the wet cell shown, zinc and copper electrodes are placed in a hydrochloric acid solution. The zinc, having a higher solution pressure, (greater negative voltage from electromotive series) will transfer positive ions into the solution. An excess of electrons will be left on the zinc plate, which may be said to be an electronegative metal when in this state. These elec-

trons are available for an outside circuit and will flow to the other electrode if a suitable connection is made between the two. At the copper plate, the electrons will join positive hydrogen ions, resulting in a layer of neutral hydrogen gas. As the hydrogen accumulates on the copper plate (polarization) additional resistance is thereby introduced into the electric circuit, and the magnitude of the current will decrease.

Inside the solution of HCl of the above cell, the hydrochloric acid dissociates. The negative Cl ions, sometimes called anions, migrate toward the zinc electrode, or anode, by virtue of the electric field through the solution. Conversely, the positive hydrogen ions, or cations, migrate toward the copper plate, or cathode. Note that the electrons flow from the

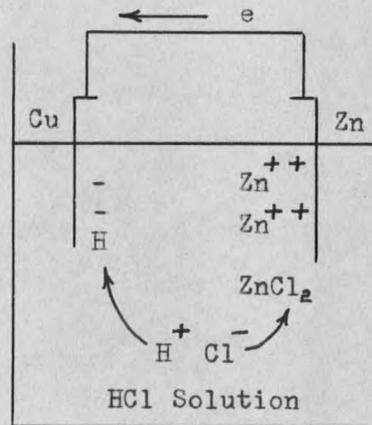


Fig. 5. Electrolytic action of copper and zinc plates immersed in an HCl solution.

zinc to copper plates in the outside circuit. The conventional current flow is in the opposite direction. An anode may be defined as the electrode which passes current into the electrolytic solution. A metal surface is called anodic when it is passing into solution by the release of positive ions. A pair of metal surfaces in an electrolyte (two dissimilar metals, or two dissimilar electrolytes must be involved) is called a galvanic cell. Electrolysis is the term given to conduction of electricity through an electrolytic solution, accompanied by a transfer of ions.

The common dry cell offers another illustration of electrochemical action. The electrolyte in this case is ammonium chloride (actually a wet solution), the zinc shell is the anode, and the carbon rod in the center is the cathode. Positive zinc ions tend to go into solution, but the action will be self limiting unless an outside electrical connection is established between the electrodes; in which case the excess of electrons on the zinc can flow over to the carbon rod where they will join NH_4^+ ions.

The above and previous illustrations are given to bring out basic corrosion theory. The actual transfer of electrons and chemical changes in many electrolytic cells become quite complex, and will show a wide variety of reactions according to the chemical constituents of the cell, current density, temperature, and ionic concentration. However it is not the purpose of this paper to give an extended treatment of electrochemical theory, but rather to apply certain fundamental facts to an understanding of corrosion processes that occur on underground metal structures. Once the nature of the corrosive action is understood, a means for controlling it will be evident.

From the theory discussed so far, the following facts are set out as being particularly important in corrosive processes. An anodic surface

