

the use of centrifugally cast pipe, lugs are not possible.)

The manager is interested in the answers to all these questions. He is also intensely desirous that cast-iron

pipe should continue to merit the reputation for long life acquired over the centuries. The new A21 specifications are an element in the attempt to achieve perfection.

APPENDIX

List of Committee A21 Specifications

The following specifications have been adopted by the American Standards Assn. after approval by the four sponsoring societies—American Gas Assn., American Society for Testing Materials, American Water Works Assn., and New England Water Works Assn.:

ASA Designation	AWWA Designation	Journal AWWA		Subject of Specification
		Issue	Page	
A21.2-1953	C102-53	Apr. 1953*	404	Cast Iron Pit Cast Pipe for Water or Other Liquids
A21.3-1953	C104-53	Apr. 1953*	404	Cast Iron Pit Cast Pipe for Gas
A21.4-1953				Cement Mortar Lining for Cast Iron Pipe and Fittings
A21.6-1953	C106-53	Feb. 1953	189	Cast Iron Pipe Centrifugally Cast in Metal Molds, for Water or Other Liquids
A21.7-1953	C108-53	Feb. 1953	207	Cast Iron Pipe Centrifugally Cast in Metal Molds, for Gas
A21.8-1953				Cast Iron Pipe Centrifugally Cast in Sand-lined Molds, for Water or Other Liquids
A21.9-1953	C110-52	Nov. 1952	1065	Cast Iron Pipe Centrifugally Cast in Sand-lined Molds, for Gas
A21.10-1952				Short-Body Cast Iron Fittings, 3 Inch to 12 Inch, for 250-psi Water Pressure Plus Water Hammer
A21.11-1953	C111-53	Apr. 1953	431	A Mechanical Joint for Cast Iron Pressure Pipe and Fittings

* Summary of revisions; the changes were not extensive enough to warrant republication of the complete text of the specifications in the Journal.

The above list covers the full scope of Committee A21 except for the revision of the American Recommended Practice Manual for the Computation of Strength and Thickness of Cast Iron Pipe—ASA A21.1-1939 (AWWA C101-39) and the specifications for large fittings for water. The revision of the manual is well along, but the specifications for large fittings have not yet been formulated. Much study has been given to large fittings, and a good many have been tested. The war interrupted this work, but the producers have started a new testing program, and these specifications should again begin to make progress.

Mechanism of the Internal Corrosion of Water Pipe

By Rolf Eliassen and James C. Lamb III

A paper presented on May 12, 1953, at the Annual Conference, Grand Rapids, Mich., by Rolf Eliassen, Prof. of San. Eng., and James C. Lamb III, Research Asst.; both of Massachusetts Inst. of Technology, Cambridge, Mass.

IRON occurs in nature in the form of an oxide or hydroxide, and the metal is produced through reduction of the ore. Corrosion may be considered as the reverse of this process, as iron is unstable in a corrosive environment and tends to revert to the more stable oxides.

The tendency of a metal to go into solution, or to corrode, is indicated by its electrode potential in that solution. In order to provide a standard basis for comparing the solution tendencies of the various metals, a metal ion "activity" of 1 mole per 1,000 g of water and a temperature of 25°C have been assumed in the preparation of the electromotive force series. In the abbreviated emf series presented in Table 1, the potential of hydrogen is taken as zero and all other potentials are referred to this value (1, 2). The metals are listed in order of decreasing solution tendency (increasing nobility of potential). The potential, E , of a metal under other conditions of temperature and metal ion activity may be found through the use of the Nernst equation:

$$E = E_0 - \frac{RT}{nF} \log_e (M^{++})$$

in which M^{++} is the metal ion activity

in solution (gram-moles of metal per 1,000 g of water, multiplied by the activity coefficient); R is the gas constant, in joules; T is the absolute temperature (Kelvin); n is the number of electrons participating in the reaction; E_0 is the potential, in volts, of the metal at 25°C with unit metal ion activity; and F equals 1 faraday equals 96,500 coulombs per equivalent.

At 25°C, this equation may be expressed:

$$E = E_0 - \frac{0.059}{n} \log_{10} (M^{++})$$

There has been a tendency to use the emf series as though it were inflexible, each metal displacing from solution all metals below it in the series. This condition is usually true for metals widely separated in the series, but it should not be assumed to obtain in all instances. Changes in the potentials of metals in dilute solutions, such as are normally encountered in water works practice, may result in radical changes in the order of potentials given in the emf series.

Such a change occurs, for example, in the iron-zinc system. Zinc is normally anodic (more active potential) to

iron, and is used in many piping systems as a sacrificial protective coating for iron. In fresh water with a temperature higher than 150°F, however, it becomes cathodic (more noble potential) to iron and is no longer sacrificially protective (2). The effective potential of a metal also may be varied by reactions with elements in a given environment. It is evident that the emf series, calculated for unit metal activity, cannot be applied directly to de-

ues are not given, as these vary widely with different environmental conditions (1, 2). Metals that are grouped together in a galvanic series have little tendency to form galvanic cells, while metals that are widely separated have a great tendency to form such cells when used together.

Corrosion Cells

Many kinds of electrolytic cells are capable of causing the corrosion of metal, but those most commonly en-

TABLE 1

Electromotive Force Series*

Electrode Reaction	Standard Electrode Potential (E ₀), 25°C
Ca = Ca ⁺⁺ + 2e ⁻	-2.87
Na = Na ⁺ + e ⁻	-2.712
Mg = Mg ⁺⁺ + 2e ⁻	-2.34
Al = Al ⁺⁺⁺ + 3e ⁻	-1.67
Zn = Zn ⁺⁺ + 2e ⁻	-0.762
Cr = Cr ⁺⁺⁺ + 3e ⁻	-0.52
Fe = Fe ⁺⁺ + 2e ⁻	-0.440
Ni = Ni ⁺⁺ + 2e ⁻	-0.250
Sn = Sn ⁺⁺ + 2e ⁻	-0.136
Pb = Pb ⁺⁺ + 2e ⁻	-0.126
H ₂ = 2H ⁺ + 2e ⁻	0.000
Cu = Cu ⁺⁺ + 2e ⁻	0.345
Ag = Ag ⁺ + e ⁻	0.800
Pt = Pt ⁺⁺ + 2e ⁻	1.2

*Signs of potential employed by the American Chemical Society are opposite to those of this table.

termine the potential of a metal under the conditions normally found in water works practice.

A galvanic series is one in which metals are arranged in their usual electro-potential sequence in a given environment. Therefore, although there is only one emf series, there may be many galvanic series, each applying to a given set of environmental conditions. In a galvanic series, like the abbreviated example in Table 2, actual potential val-

TABLE 2

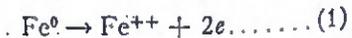
Galvanic Series of Metals and Alloys

Corroded end (anodic or least noble)	Magnesium Zinc Aluminum 2S* Steel or iron Cast iron Lead Tin Brasses Copper Bronzes Chromium-iron (passive)
↑	Silver
Protected end (cathodic or most noble)	Graphite Gold Platinum

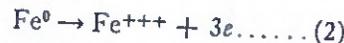
* "Commercial pure" aluminum.

countered in practice may be divided into a few basic types: impressed-current cells, galvanic or differential-metal cells, concentration cells, differential-aeration cells, differential-temperature cells, and differential-stress cells.

Impressed-current cells. It is well known that, in an impressed-current cell (Fig. 1), the current flow between the electrodes results in a loss of metal from the anode in accordance with one of the following reactions, depending upon the current density:



or:



The rate at which the metal passes into solution is directly related to the rate of current flow between the electrodes. Faraday's law states that the passage of 1 faraday of electricity (95,500 amp-sec or coulombs) through an electrolytic solution produces one chemical equivalent of some chemical change at each electrode. Therefore, the passage of 1 faraday of electricity between the

This reaction results in the formation of atomic hydrogen, which tends to plate out on the metal. The layer of atomic hydrogen on this electrode tends to interfere with the transfer of electrons from the metal to the solution and, therefore, must be removed in order for the reaction to proceed. This may be accomplished through the formation of hydrogen gas:

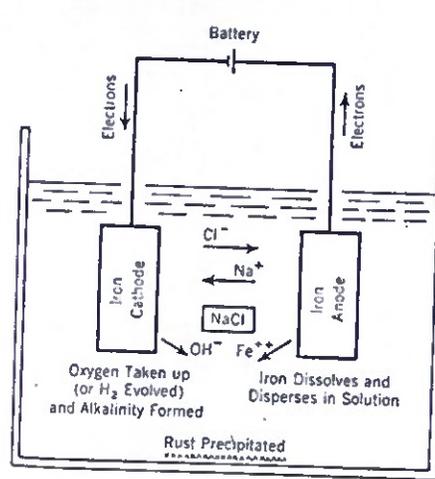


Fig. 1. Impressed-Current Cell

The current flow between the electrodes results in a loss of metal from the anode.

electrodes in Fig. 1 must cause one chemical equivalent of iron to enter solution from the anode.

One chemical equivalent of change is also produced at the cathode for each faraday of electricity passing through the solution. The primary reaction at this electrode involves the removal of electrons from the metal through neutralization of hydrogen ions in solution:

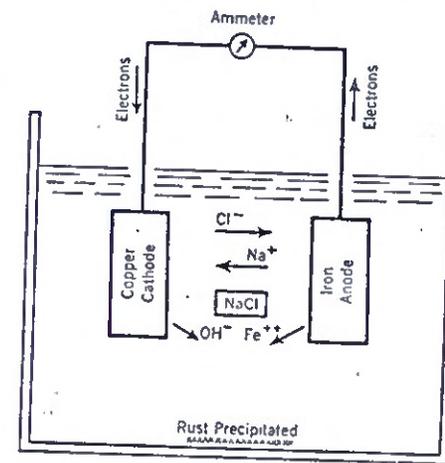
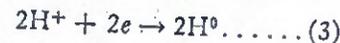
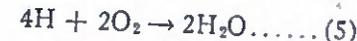


Fig. 2. Galvanic Cell

The reactions at the electrodes are identical with those in the impressed-current cell, but the current flow is due to the difference in the electrode potentials rather than to an external source.

or through reaction between the atomic hydrogen and chemicals in solution in the water. The most common reaction of this type involves the atomic hydrogen and oxygen in solution in the water:



Other possible types of reactions between hydrogen and the oxygen in so-

lution have been suggested (1-4). The neutralization of hydrogen ions in the solution at the cathode results in the production of alkaline conditions in the vicinity of this electrode. The metal ions or salts entering solution from the anode react with alkalinity and oxygen in solution to form iron oxides, known as rust. These oxides may be formed in the vicinity of the anode or cathode, or in a location remote from either electrode.

The potential of a metal in water may be changed through the application of impressed currents, causing an increase or decrease in the rate of corrosion of the metal, depending upon the direction and magnitude of the applied emf (2). This phenomenon can be observed in the external corrosion of pipelines by stray currents and in the cathodic protection of pipeline exteriors and other structures by applied emf.

The corrosion of iron in water usually occurs in the absence of impressed currents. Whitney (5) showed that the corrosion of iron in water is electrochemical in nature, and Evans and Hoar (6) demonstrated that there is a direct relationship between the current flowing between anodes and cathodes on a metal surface and the weight of metal lost from that surface.

Galvanic cells. The emf and galvanic series show that, when different metals are introduced into a given environment, they have different tendencies to corrode, and display different potentials. Figure 2 illustrates such a situation, with iron and copper immersed in the same aqueous solution. Examination of a galvanic series (Table 2) reveals that copper may be expected to display more noble (cathodic) potential than iron. If the two metals are connected electrically, the difference

in potentials between the iron and copper results in a flow of current between the electrodes. This may be demonstrated by connecting the two electrodes through an ammeter, as shown, and measuring the galvanic current flow.

In the cell shown in Fig. 2, the iron has the greater tendency to pass into solution, leaving electrons at the metal surface. These electrons pass through the metallic portion of the circuit to the copper electrode and are transferred

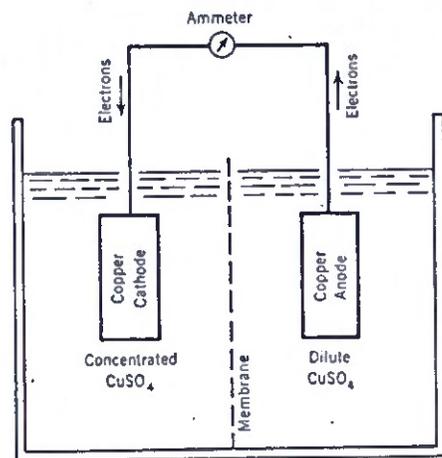


Fig. 3. Concentration Cell

In this type of corrosion cell, the difference in electrode potentials results from the difference in metal ion concentrations in the solutions.

from the metal to the solution by neutralization of hydrogen ions in the solution. The atomic hydrogen is removed from the copper through the formation of hydrogen gas or reaction with oxygen in solution. In this cell, the iron serves as the anode (electrode at which oxidation takes place), while the copper functions as the cathode (electrode at which reduction takes place). The current flow that, accompanies the loss

of metal from the anode is generated by the difference in potentials between the electrodes in the galvanic cell rather than by an external source of emf as in the impressed-current cell. Therefore, the galvanic cell in Fig. 2 can be considered—and, in fact, is—nothing more or less than a battery. The reactions taking place at each electrode are identical with those described above for the impressed-current cell.

Concentration cells. From the Nernst equation previously given, it can be seen that the potential of a metal in a solution varies with the concentration of its ions in that solution. As the concentration of metal ions in solution increases, it becomes more difficult for additional metal to enter the solution as ions. Therefore, a metal can be expected to display a more noble potential in the presence of a high metal ion concentration. Figure 3 shows two identical copper electrodes immersed in solutions of different concentrations of copper sulfate. As these electrodes are exposed to different concentrations of copper ions, they have different potentials. The electrode in the more concentrated solution has a more noble potential and acts as the cathode in the corrosion cell, while the other electrode functions as the anode.

The potential of a metal in solution is affected by the concentration of ions other than metal ions. For example, a concentration cell similar to the one in Fig. 3 can be prepared by using two sodium chloride solutions of different strength. This effect can also be observed with metals immersed in solutions of different substances (acids, bases, salts, or gases).

Differential-aeration cells. A differential-aeration cell is a special type of concentration cell which is of particular

significance in the water works field. The rate of the chemical reaction between the oxygen in solution and the hydrogen atoms on the metal depends upon the rate of access of oxygen to the metal surface. Therefore, if one electrode of a corrosion cell receives oxygen at a higher rate than the other, the polarizing layer of hydrogen on this electrode is removed more rapidly. The rapid removal of hydrogen facilitates the transfer of electrons from the metal to the solution at that point, and this area of the metal becomes the cathode of the corrosion cell. The electrode that receives less oxygen becomes the anode of the differential-aeration cell. This electrode suffers an increase in corrosion rate owing to the increased rate of depolarization at the cathode. The significance of this type of corrosion cell in water works practice will be discussed more fully later.

Differential-temperature and differential-stress cells. Differences in temperatures and stresses of metals may result in the formation of corrosion cells. These types of cells may be important in initiating corrosion, but their effects are usually obscured by other, more significant kinds of corrosion cells. Therefore, on a long-term basis, differential-temperature and differential-stress cells are of relatively little importance, under the conditions usually found in water works.

Formation of Corrosion Cells

For the sake of clarity, the electrodes in the cells considered in the preceding section were assumed to be individual pieces of metal. The corrosion of a single submerged metallic object results from the formation of anodes and cathodes on the surface of the metal. Figure 4 shows, schematically, a typical

corrosion cell on the surface of a submerged sheet of iron. The anode and cathode of this cell are short circuited by the body of metal. It will be noted that the cell and the electrode reactions shown in Fig. 4 do not differ basically from those which have been discussed previously. The surface of a large section of metal might be covered by many such corrosion cells. The presence of local anodes and cathodes on the surfaces of corroding metals has been demonstrated, and it has been shown that

gen acts as the cathode in the corrosion cell, while the portion shielded from oxygen becomes the anode.

Figure 5f shows a galvanic cell caused by connecting pipes made of different metals. Copper and brass are cathodic to iron, and, therefore, the iron pipe corrodes while the copper or brass pipe acts as the cathode in the corrosion cell. Figure 5g shows conditions under which an electrolyte concentration cell may be encountered in the external corrosion of pipelines passing through soils

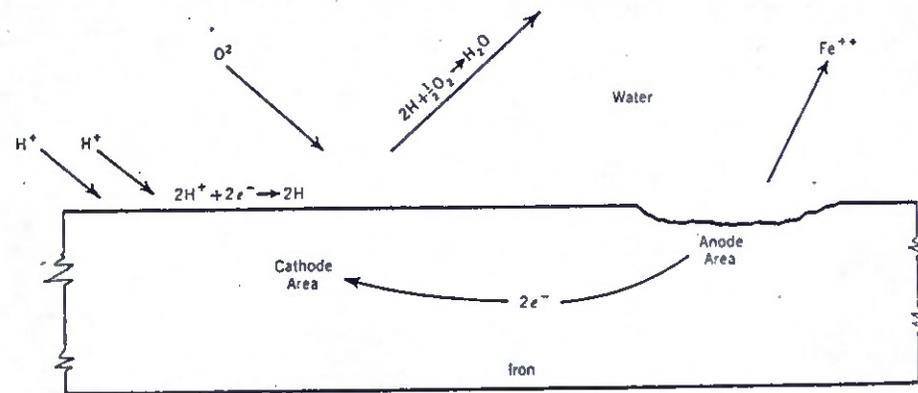


Fig. 4. Corrosion Cell on Surface of Submerged Iron

The surface of a large section of metal may be covered by many such cells, the electrode reactions of which are similar to those discussed previously.

the amount of current flowing between the electrodes and the weight of metal lost can be definitely correlated (6, 7).

The formation of a corrosion cell on the surface of a metal may be due to differences in metal composition, electrolyte concentration cells, or differential-aeration, -temperature, or -stress cells. Figure 5 shows a few typical examples of how such cells may be formed in practice. The first five (Fig. 5a-5e) are examples of differential-aeration cells. In each instance, the portion of the metal that has freer access to oxy-

gen or waters that vary widely in chemical characteristics.

Finally, Fig. 5h illustrates a common example of corrosion by impressed current. If the water main may become part of the return circuit to a power plant furnishing electricity for a streetcar system, the portion of the pipe where the current passes from pipe to ground suffers an increase in corrosion rate. The corrosion rate of that portion of pipe at which no interchange of current takes place between pipe and ground is not affected, and the internal

corrosion of the pipeline is likewise unaffected.

The types of cells capable of causing corrosion have been discussed individually, but it must be stressed that two or more varieties may tend to occur

metal ions would indicate a tendency toward the formation of a corrosion cell in which the area of metal under the tubercle would act as the cathode. The differential-aeration cell formed by the action of the tubercle in shielding a por-

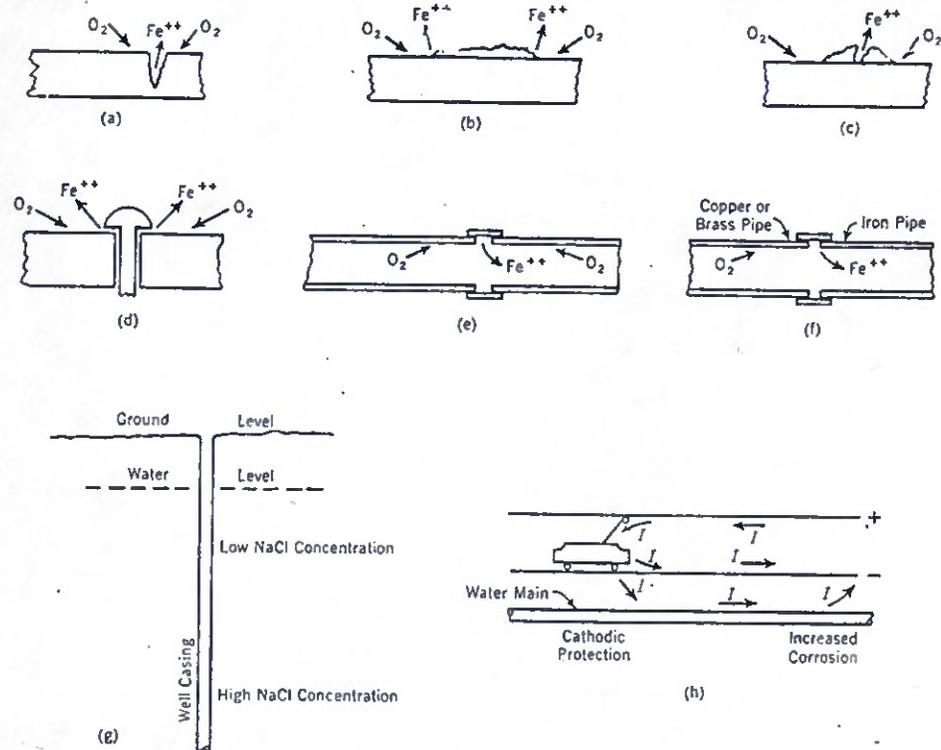


Fig. 5. Typical Corrosion Cells Found in Practice

Key: (a) pit or depression; (b) mill scale or corrosion products; (c) biological growth; (d) rivet or bolt; (e) pipe coupling (same metals); (f) pipe coupling (dissimilar metals); (g) well casing in water with varying salt content; (h) stray-current corrosion. The first five are examples of differential-aeration cells; the remaining three illustrate, respectively, galvanic, concentration, and impressed-current cells.

simultaneously. For example, the concentration of metal ions under a tubercle (Fig. 5c) could be expected to be higher than in the water surrounding the tubercle. Therefore, consideration of the relative concentrations of

the metal from oxygen tends, however, to result in the formation of a cell in which the area of metal outside the tubercle acts as the cathode and the area under the tubercle serves as the anode.

Rate of Corrosion

It has been pointed out that the rate of corrosion of a metal in aqueous solution is a function of the rate of current flow between anodes and cathodes on the surface of the metal. The reactions taking place in a corrosion cell have been presented in the preceding sections: It is evident that, in order for corrosion to take place (Reaction 1), the electrons left on the metal by this reaction must be removed (Reactions 3-5). Otherwise, a counter, or back, emf will be built up on the metal, and the corrosion will be stifled. Therefore, in a corrosion cell, the anode and cathode reactions must proceed at the same rate. Any reduction in the rate of the reaction at either electrode must result in a corresponding reduction in the rate of current flow between the electrodes and, hence, in the rate of corrosion of the metal. In practice, the rate of corrosion may be controlled by reducing the rate of the anode reaction, the cathode reaction, or both.

The removal of the polarizing layer of hydrogen from the cathode may be accomplished through hydrogen evolution (Reaction 4) or through oxygen depolarization (Reaction 5). It has been shown by past investigators that, in neutral or mildly alkaline waters, corrosion with evolution of hydrogen at the cathode is negligible, although the relative positions of iron and hydrogen in the emf series might lead to the supposition that the corrosion of iron should be accompanied by continuous evolution of hydrogen. The explanation is that resistance or "over-voltage" must be surmounted before hydrogen can be evolved as a gas from a metal surface. In neutral or mildly alkaline waters, the overvoltage is too

high to allow the escape of appreciable amounts of hydrogen gas from iron cathodes. As the pH of a water decreases below a value of approximately 4.0, the hydrogen ion concentration is increased to such an extent that hydrogen evolution is greatly facilitated, and the formation of this gas becomes a major cathodic reaction.

Because the waters encountered in municipal and industrial systems are generally neutral or slightly alkaline, the amount of corrosion due to hydrogen evolution is of minor significance—approximately 3 per cent, according to Speller (2). Therefore, the primary cathodic reaction in these systems must be the removal of hydrogen by reaction with oxygen in solution in the water (Reaction 5). The anode reaction rate is much more rapid than the depolarization reaction rate at the cathode (2). As both reactions must proceed at the same rate, it is evident that the rate of corrosion of iron in natural waters is controlled by the rate of oxygen supply to the cathodes of the corrosion cells. The latter rate is influenced by the concentration of oxygen in solution, the rate of flow and turbulence of the water, and the presence of films of oxides or inhibitors on the metal surface. Within the range of oxygen concentrations normally found in municipal and industrial waters, the rate of corrosion increases directly with the rise in oxygen concentration.

Uhlig (8) has stated that: "All iron and steel, regardless of minor variations of composition or structure, corrode at the same rate in natural waters." This fact is of great significance to the water works industry, which uses many types of iron and steel pipes for various reasons, some realistic and others not.

The formation of films on the surfaces of metals has a significant effect on their corroding characteristics. The films formed on metals may include layers of hydrogen or oxygen, films formed by chemical reaction between corrosion products and elements of the corrosive environment, and films formed by inhibiting agents added to the corrosive medium. Film formation may occur preferentially at either anodes or cathodes of corrosion cells, or over the entire surface of the metal. Films on the cathodes of corrosion cells may interfere with the transfer of electrons from the metal to the solution. The formation of these films may result from the electrodeposition of positively charged ions or colloidal particles on the cathode, from sorption of chemicals on the metal surface, or from the reaction of some chemical in solution with the alkalinity produced at this electrode, resulting in the formation of an insoluble precipitate on the metal.

Anode films may be formed through electrodeposition of negatively charged ions or colloidal particles on this electrode, through sorption of chemicals on the metal surface, or through the production of an insoluble precipitate by reaction between the iron entering solution and chemicals in solution. The formation of an anode film may result in the effective separation of the metal from the water at the anode, stifling the anode reaction.

Under certain conditions, film formation on a metal may be accompanied by a radical change in the potential of that metal to a more noble value. When a metal undergoes a sharp reduction in corrosion rate, accompanied by a shift in potential to a more noble value, it is said to be "passive."

Figure 6 shows the effect of the pH of the water on the corrosion rate of mild steel in the absence of the formation of protective coatings (9). The rate is essentially constant in the range of pH 4-10 at 22°C. In this range, the predominant cathodic reaction is oxygen depolarization, and the corrosion rate is determined by the rate of supply of oxygen to the metal surface. As the pH of the water is decreased below 4, hydrogen evolution becomes a major cathodic reaction. The increased corrosion rate at low pH values can be attributed to increased hydrogen ion concentration in the solution and the dissolving action of the water on any protective layers of corrosion products on the metal. As the pH is increased above 10, a protective layer of corrosion products is formed on the metal, or the surface reaction rate of the iron becomes lower, resulting in passivity of the metal and a decrease in the corrosion rate. The chemicals present in natural waters may have a great effect upon the shapes of the curves in Fig. 6, through reaction with corrosion products and formation of protective coatings.

Polarization

In a corrosion cell, the driving force for the electrical current flow between the electrodes is the difference in potential between the anode and the cathode of the cell. From Kirchoff's second law:

$$\sum emf = \sum IR$$

or:

$$E_c - E_a = IR_c + IR_m$$

in which R_c is the resistance of the electrolytic portion of the circuit; R_m is the resistance of the metallic portion of the

circuit; E_c is the effective (polarized) potential of the cathode; E_a is the effective (polarized) potential of the anode; and I is the current flow between the electrodes.

The potentials of anodes and cathodes in corrosion cells may vary with the current flow through the cells. This effect is known as polarization and is comparable to friction in a mechani-

tential difference between the electrodes in a corrosion cell through which current is flowing is usually less than the difference in the open-circuit potentials, or potentials of the electrodes with no current flowing. An increase in the degree of polarization of either electrode results in a decrease in the potential difference between the electrodes (or driving force for the corrosion proc-

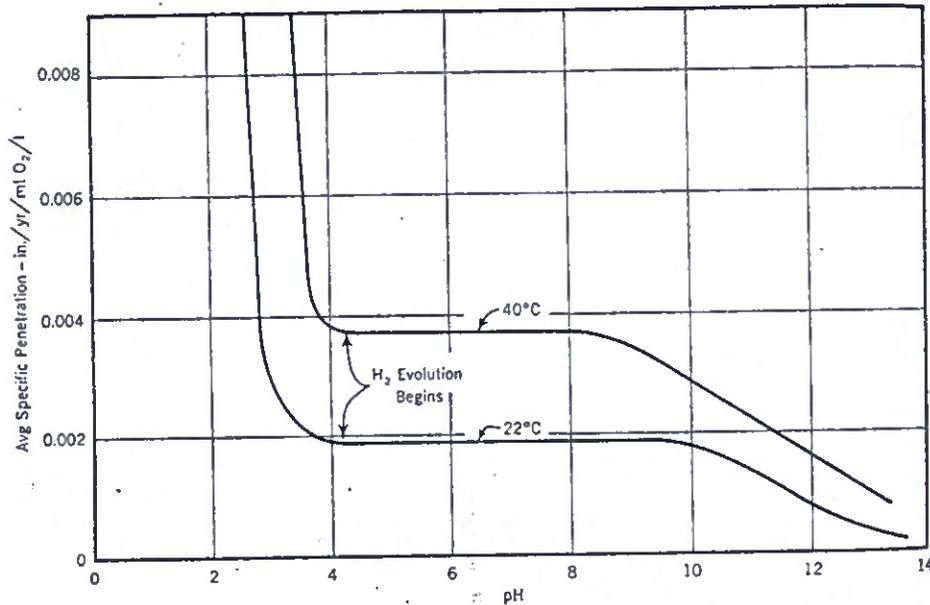


Fig. 6. Effect of pH on Corrosion of Steel

In the pH range 4-10, the corrosion rate depends on the rate of oxygen supplied to the metal surface. Above pH 10, the formation of a layer of corrosion products leads to passivity. Below pH 4, hydrogen evolution becomes a major cathodic reaction.

cal or hydraulic system. As the current flow through a corrosion cell increases, the potentials of the anode and cathode tend to approach a common value. Polarization of both electrodes would result in a shift in the anode potential in the cathodic direction and a shift in the cathode potential in the anodic direction. Therefore, the po-

less) and, thus, in a decrease in the rate of current flow in the cell.

The polarizing characteristics of separate anodes and cathodes under a given set of conditions may be determined by varying the rate of current flow between the electrodes (by varying R_m) and measuring the potential of each electrode against that of a stand-

ard reference cell. In Fig. 7 is shown a typical pair of polarization curves, indicating the potentials of the anode and cathode of a corrosion cell at various values of current flow between the electrodes. The points at which the curves intercept the ordinate axis represent the open-circuit potentials of the electrodes (no current flowing in the corrosion

trodes against a "remote" reference cell, or a reference cell located midway between the electrodes. The broken lines in Fig. 7 are those which would be obtained through the use of probe electrodes, located in the immediate vicinity of the respective corrosion cell electrodes. These curves are separated by a distance equal to $IR_c + IR_m$.

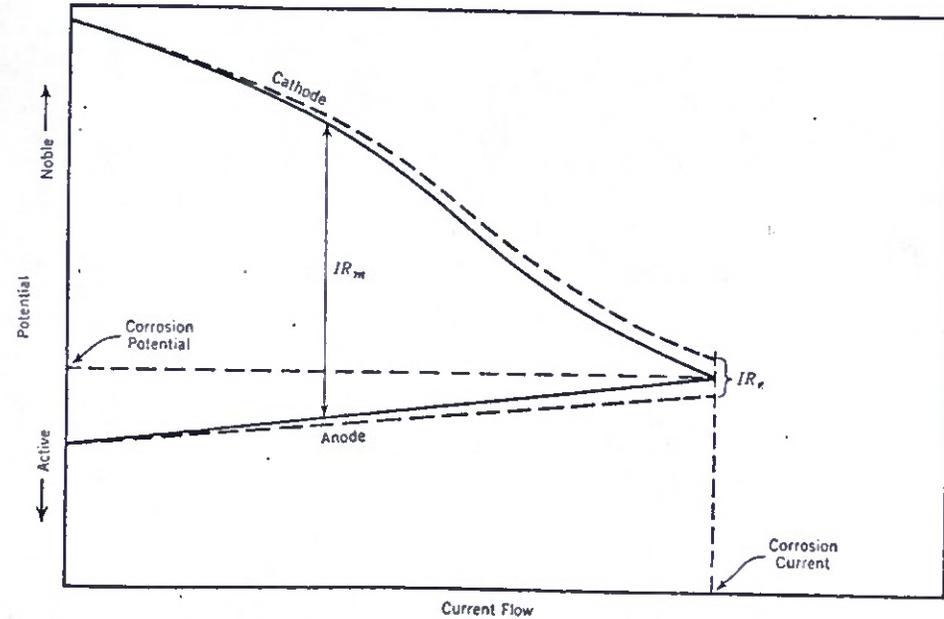


Fig. 7. Typical Polarization Curves

The solid lines represent the curves that would be obtained if the electrode potentials were measured against a "remote" reference cell; the broken lines indicate the curves resulting from the use of "probe" electrodes, located close to the respective corrosion cell electrodes.

cell). The intersection of the curves gives the corrosion potential and the corrosion current with the electrodes short-circuited, which is usually the situation in practice. The curves represented by the solid lines are separated by a distance equal to IR_m . These are the curves that would be obtained by measuring the potentials of the elec-

It has been stated that the rate of corrosion may be controlled by the rate of the cathode reaction, the anode reaction, or both. Altogether, four types of control are possible in corrosion cells. These are illustrated in Fig. 8. Figure 8a represents the limiting case of "resistance control," in which the current flow between the electrodes has no ef-

fect on the potential of either of them; the current flow in the corrosion cell is limited by the IR_c and IR_m of the system, and the potential of each electrode remains constant at the value of the open-circuit potential for that electrode. This situation would never occur in water works practice. Figure 8b represents the case in which the current flow in the corrosion cell is limited by

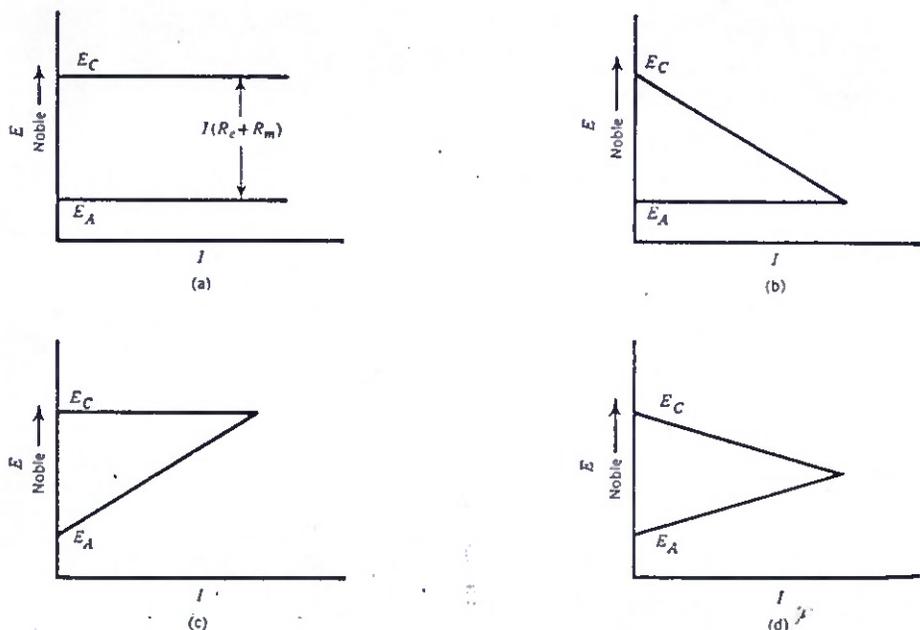


Fig. 8. Types of Corrosion Control

Key: (a) resistance control; (b) cathodic control; (c) anodic control; (d) mixed control.

the polarization (change in potential with current flow) of the cathode. In this instance, the corrosion is said to be under "cathodic control." Figures 8c and 8d show polarization curves that might be obtained when the rate of corrosion is controlled by polarization of the anode ("anodic control") or of both electrodes ("mixed control").

It has been pointed out that, in aqueous solutions with no inhibitors present, the anode reaction rate in the corrosion of iron is very much faster than the depolarization reaction rate at the cathode. In neutral or slightly alkaline waters, the rate of corrosion is almost completely governed by the rate of supply of oxygen to the cathode. Therefore, in most municipal and industrial

water systems, the corrosion is under primarily cathodic control, as illustrated in Fig. 8b. The rate of corrosion is governed by the degree of polarization of the cathode.

Distribution of Corrosion

The factors affecting the overall rate of corrosion of iron in water have been

discussed. Other factors of equal significance are those which influence the distribution of corrosion occurring; the corrosion may take place uniformly over the entire metal surface or be localized, with the entire loss of metal confined to a relatively small area. The life of a metallic structure in water is determined, to a large extent, by the rate of penetration of the structure by corrosion. Localized attack may lead to rapid failure of the structure, although a relatively small proportion of the metal has been lost to solution.

As the rate of corrosion of iron in water is usually controlled by the rate of the cathode reactions, the formation of a protective coating that only partially shields the anode of a corrosion cell from the corrosive medium may have little or no effect on the rate at which metal passes into solution. In these circumstances, incomplete shielding of the anode may lead to a reduction in the effective anode area and to an increased rate of penetration of the metal. This results in more rapid failure of the structure.

Chemical deposits, as well as deposits of biological organisms on the surface of a metal, may result in the formation of tubercles. Tubercles restrict the access of oxygen to the surface of the metal, thereby establishing differential-aeration cells. The metal under a tubercle undergoes pitting attack, while the metal surrounding the tubercle serves as the cathode of the corrosion cell. The anaerobic conditions within tubercles favor the growth of the sulfate-splitting type of bacteria, even though oxygen may be present in the water outside the tubercle. These organisms are capable of utilizing cathodic hydrogen, thus causing microbiological depolarization, and they pro-

duce hydrogen sulfide, which reacts with the ferrous metal to form ferrous sulfide (2).

The formation of protective coatings on the cathode of a corrosion cell under cathodic control results in an overall reduction in the rate of metal loss without increasing the rate of penetration of the metal, because the effective area of the cathode is reduced. The formation of protective films on the cathode tends to restrict the access of oxygen to this electrode. Simultaneous removal of corrosion products from the anode of the corrosion cell would result in the formation of a differential-aeration cell, with the shielded cathode becoming the anode of the new cell, while the area that was originally anodic but now has freer access to oxygen would serve as the cathode. It is evident that such conditions favor continuous reversals of the potentials of anodes and cathodes on the metal surface. This situation results in a uniform attack of the metal surface over a period of time.

In solutions of low conductivity, the anodes and cathodes are limited in size by the resistance of the electrolytic portion of the corrosion cell. The surface of iron in a water of low conductivity is covered by a large number of small, evenly distributed anodes and cathodes, resulting in a relatively uniform attack on the metal surface. In a water of high conductivity, the cathodic area serving a given anode may be much larger, owing to the lower internal resistance of the corrosion cell. Under such conditions, there are fewer corrosion cells on a given area of metal, and localization of corrosion is increased (2).

Stray electrical currents cause localization of corrosion at areas where the

current leaves the pipe. Localized corrosion may also occur near the water-line in storage tanks.

Summary

The mechanism of the internal corrosion of water pipes is of a very complex nature. The rate and distribution of pipeline corrosion are affected by many variables, a few of which have been discussed briefly.

The evaluation of existing corrosion inhibitors and the development of more effective means for reducing pipeline corrosion are problems of prime significance to the water works profession. The solution of these problems can be accomplished only after a thorough understanding of the mechanism of corrosion and corrosion inhibition has been obtained through extensive research programs. In the light of such knowledge, it should be possible to design laboratory and field tests to evaluate accurately the existing methods of reducing losses through corrosion, and to develop more efficient methods for controlling internal pipeline corrosion.

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Findings and Recommendations on Underground Waste Disposal

Task Group Report

A task group report presented on May 14, 1953, at the Annual Conference, Grand Rapids, Mich., by Norman F. Billings, Chairman, Task Group E4.C—Underground Waste Disposal and Control; Hydrogeologist, Michigan Water Resources Com., Lansing, Mich. Other members of the task group were: L. E. Blakeley, A. G. Fiedler, T. R. Lathrop, T. J. Powers, E. J. Schaefer, S. W. Wells, and C. V. Youngquist.

THE assignment of this task group has been to study the subject of industrial-waste disposal in or on the ground and assess its effects upon present or future ground water supply. A report was made in 1952 on findings up to that time, with special consideration of the Michigan situation (1). The task group felt that data then available did not present a complete picture of the nationwide ground water pollution situation, and further investigation was proposed. In deliberating methods for collecting additional data, there was a question whether anything would be gained by further refinement of the picture and whether the principles of ground water pollution control were not so simple and basic that they would apply regardless of further findings.

Following this line of thought, those facts which appear fundamental were selected and their practical significance was interpreted. This analysis suggested recommendations that seem not only eminently logical but also adequate for translation into statutory control measures. The task group has had recourse to studies conducted by the

ground water pollution committee of the Michigan Section, AWWA, two members of which are included in the task group. Results of the task group analysis are herewith presented under the categories of findings, conclusions, and recommendations.

Findings

1. Although ground water pollution by industrial-waste disposal is reported as relatively minor in many states, and even nonexistent in some, it is, nevertheless, nationwide in distribution. Important areas of such pollution are found in the Far West, the South, the East, the glaciated midcontinent region, and some central states.

2. Pollution is reported to have resulted from a number of wastes, including natural and artificial brines, phenolic compounds, cleaning fluids, salts of heavy metals, picric acid, gasoline and other petroleum products, pickling liquors, fruit and vegetable cannery wastes, milk wheys, and spent cooling water.

3. Ground water pollution resulting from industrial wastes is generally a