CORROSION:

What's Eating at Your Metal?

By Cris Crissinger, CCS, CCCA, ASQ

This article will be published in two parts. Part I describes basic corrosion principles and galvanic reactions. Part II will discuss why some metals are more corrosion-resistant than others.

This discussion will focus on the corrosion of the common metals, such as aluminum, copper, lead, and variations of ferrous metals, including stainless steel, that are used in construction.

Most structural components, from the time of their erection, are under continual corrosive attack by the environment. How well they repel this assault is the designer's responsibility. The information presented and the recommendations suggested are not all-inclusive but are intended to make design professionals aware of potential corrosion problems and to offer some proven suggestions that will provide a foundation for solving issues.

Corrosion is like a disease. In order to prevent or control it, the designer should have a basic understanding of what corrosion is and what causes it. Of the many forms of corrosion—including stress, pitting, fretting, intergranular, and microbiological—the focal point of this article will be corrosion from galvanic action and the atmosphere.

Corrosion is a very complex process, and the ion exchange between metals and their compounds can get very messy. Numerous detailed publications have been devoted strictly to metallurgy and corrosion theory. Some readers may be well versed in corrosion principles, while others may not. The author's intent is to simplify the corrosion process and to provide information that will be beneficial to designers. For those who feel attracted to chasing positive and negative ions, the National Association of Corrosion Engineers (NACE) reference publications at the end of Part II of this article will provide plenty of exercise.

DEFINING CORROSION

The NACE defines corrosion as the deterioration of a substance (usually a metal) or its properties because of a reaction to a substance. Physics says it is the process of a metal returning to its natural oxide state. The author's mentor accurately professed that corrosion occurs when a metal is put in an environment where it should not be.

Most metals (with the exception of precious metals such as gold and platinum) do not occur naturally in their metallic form. Instead, they exist as ore, such as iron ore (ferrous oxide or iron oxide) and aluminum ore (aluminum oxide). The ore must be processed to separate it from the metal in order to make something useful such as steel. As an ore, these metals are considered to be in a stable, low-energy state. When the ore is processed into a metal, that metal is then considered to be in an unstable, high-energy state. Precious metals such as gold and platinum exist naturally in their metal state, have intrinsic electrochemical stability, are not prone to corrosion, and are referred to as noble metals.

When corrosion occurs, the corroding metal tries to revert from its high-energy, processed, and metallic state to its low-energy and more natural state, which we commonly call ore. When an ore is processed into metal, the refining process transfers its latent energy to the ore. When this energy is transferred to the ore, the ore is converted into metal. As the metal receives this energy, the metal is elevated to a higher-energy state, which means its electrochemical stability decreases. Corrosion occurs when a metal that is in a high-energy state reverts to its low-energy (ore) state.

Some metals, such as copper, form oxide films that are porous; while other metals, such as ferrous metals, form oxide films that buckle and flake off. In both cases, the underlying metal is subject to a continued oxidation process and results in metal loss. However, there can be an upside to oxide films. Most metals that form a tightly adhering layer of surface oxides are much more corrosion-resistant because the adhering

GALVANIC SERIES TABLE	
Metal	**Standard Volt Potential**
Potassium	-2.922
Magnesium	-2.340
Aluminum	-1.670
Zinc	-0.762
Chromium	-0.710
Iron	-0.440
Nickel	-0.250
Tin	-0.136
Lead	-0.126
Hydrogen	**-0.000**
Copper	+0.345
Silver	+0.800
Platinum	+1.200
Gold	+1.680

Table 1

layer—known as the passivation layer—offers significant protection by isolating the metal from the corrosive environment. Aluminum and copper are such metals. It has been said that "passivity can be considered as when a metal does not corrode when it should."

CAUSES OF CORROSION

Anyone who has been exposed to construction or construction design should be familiar with the statement, "Separate dissimilar metals to prevent galvanic action." Galvanic action caused by dissimilar metals is the basis for a significant part of the corrosion.

Just as a cold offers subtle symptoms at its onset, such as a runny nose, slight cough, lump in one's throat, chills, and fever, corrosion also offers clues at its beginning that include discoloration, pitting, flaking, distortion, changes in size or shape, and invisible cracks.

HOW CORROSION WORKS

Until designers start putting buildings in a vacuum or until the laws of electrochemical properties of metals are repealed, there will be corrosion. The four conditions that must exist for corrosion to occur:

- There must be an anode (the corroding component) and a cathode (the protected component).
- There must be an electrical potential between the anode and the cathode such as that created by dissimilar metals.
- The anode and cathode must be immersed in an electrolyte, which is an electrically conductive fluid.
- The anode and cathode must be connected by a path of low resistance.

GALVANIC REACTION

Galvanic action is probably the most common type of corrosion that affects metals and is based on the grouping of metals together on a scale called the Galvanic Series Table. *Table 1* shows the position of some common metals in the electromotive series.

A metal's position relative to the other metals on the scale is based on its electrical potential. The strength of this electrical potential, when compared to other metals in the series, determines the probability of a reaction. The position of a metal on the scale also determines the severity of that reaction when the metals are exposed to a common and suitable environment. The common and suitable environment chosen to determine the galvanic scale is seawater because of its abundance and consistency.

The metals that are the most negative tend to be the most reactive and are considered anodic. The metals that are the most positive are the least reactive and are considered to be cathodic. This is why dissimilar metals are those that have significantly different electrical potential (large separation). For example, based on the galvanic table, a metal should be protected by a metal that is higher on the list (more negative electrical potential than a metal below it). The farther apart the metals are on the scale, the higher the probability for corrosion. The voltage difference creates the current flow and increases the rate of corrosion; and the greater the difference, the faster the corrosion. Hydrogen is considered to be neutral.

The Galvanic Series Table reveals several useful bits of information:

1. Metals high in the series are active and less corrosion-resistant.
2. A metal will replace another metal in solution if the metal in solution is lower in the series.
3. Whether a metal will be corroded by an acid. Metals above hydrogen are usually attacked by acids.
4. An indication of the bond strength between combined cells based on electromechanical potential. The larger (less negative) the combined voltage potential in metals is, the stronger the bond will be. Thus, the -2.432-volt bond between aluminum and zinc (-1.670 + -0.762 = -2.432 volts) is stronger than the -1.325-volt bond between aluminum and copper (-1.670 + +0.345 = -1.325 volts).

ELECTROLYTES

Water in some form is probably the most common electrolyte. It can originate from tap water, weather, oceans, lakes, cleaning solutions, etc. However, all water is not pure. The pH of a solution is one of the most important factors that can determine the aggressiveness of water.

The pH scale in *Table 2* shows the acidity (red end) and the alkalinity (blue end) of

Classification	pH	Color	Common Item
	0		Battery Acid
	1		Gastric Acid
	2		Lemon Juice
More Acidic	3		Vinegar
	4		
	5		Coffee
	6		Milk
Neutral	7		Distilled Water
	8		Baking Soda
	9		Sea Water
	10		
More Alkaline	11		Milk of Magnesia
	12		Ammonia
	13		Bleach
	14		Lye

Table 2

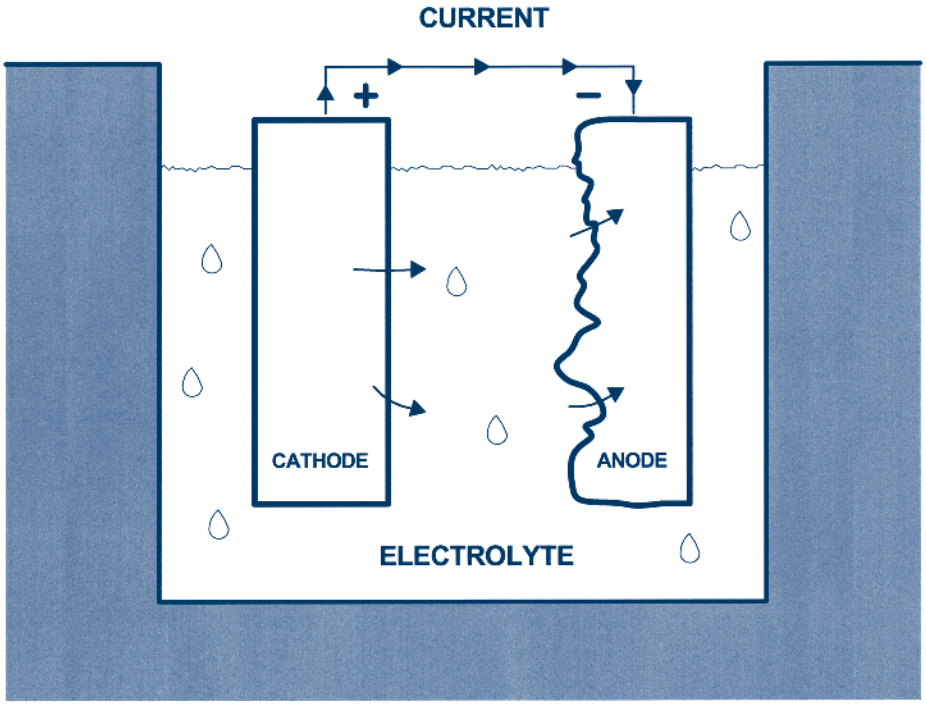

Figure 1 – Galvanic reaction, separated metal.

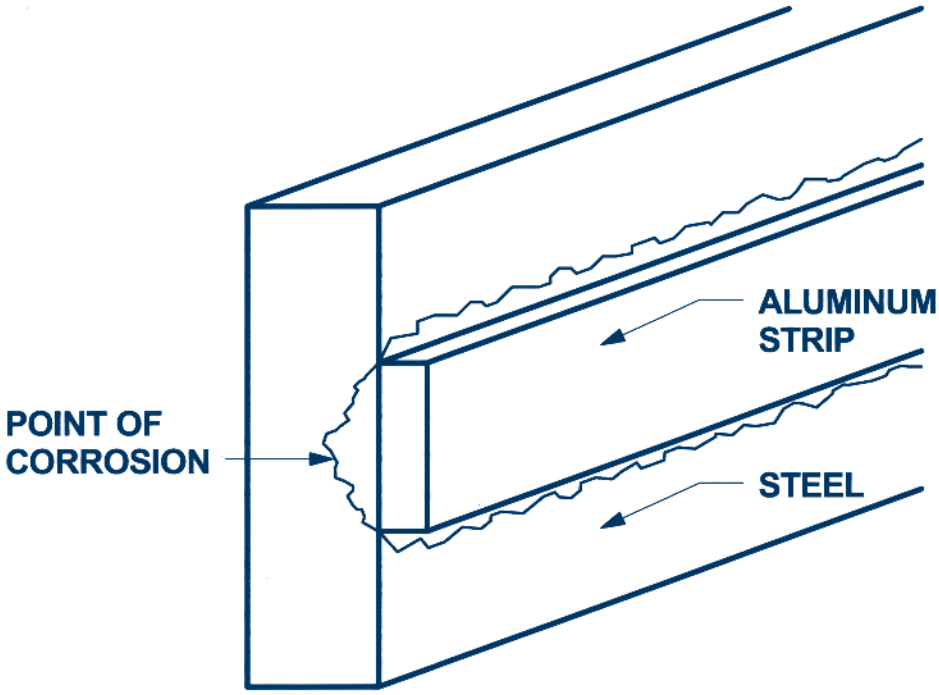

Figure 2 – Galvanic reaction, contacting metal.

common solutions for comparison. The pH is measured on a logarithmic scale beginning at the top from 0 to 14, with 7 being neutral. This means that an electrolyte with a 6 pH is 10 times more acidic than a 7 pH, and a 5 pH is 100 times more acidic than a 7 pH. Small changes make a big difference! Electrolytes with a pH at the extreme ends of the scale tend to be more corrosive than those toward the center, and those at the acidic end tend to be the most corrosive.

Electrolyte temperature tends to have an effect on corrosion. Corrosion seems to increase in tropical seawater where water surface temperatures are often above 85°F. Water temperatures above 140°F have produced a noticeable increase of corrosion in hot-water heaters. Conversely, the exploration of steel ships sunk in deep, cold ocean water shows corrosion to be extremely slow.

If two dissimilar metals are immersed or otherwise exposed to an electrolyte, they may or may not corrode by themselves if left separated in the electrolyte. But when they contact each other and a galvanic (electric) couple is created, the corrosion rates of the individual metals change. The corrosion rates change because one of the metals becomes the anode and corrodes faster than it would by itself. The other metal becomes the cathode and corrodes slower than when by itself. The difference in electric potential causes a current to flow from the anode to the cathode. This electrical circuit causes the anode to lose some of the metal in the form of ions. This process slowly consumes the anode and effectively strengthens the cathode against corrosion. The metal removal is the corrosion. As the electrons flow across the low-resistance, metallic path to the cathode, the electrical potential of the cathode is reduced enough to stop the formation of metal ions. When the metal is not in a positively charged ionized form, it is much less reactive. This is why the cathode develops an inherent protection from metal loss to corrosion. *Figure 1* is a simple cell showing galvanic reaction of separated metals and how a current leaves a metal (the anode) and enters an electrolyte. The electrolyte conducts the current to the metal surface (the cathode) that receives the current. A wire connecting the anode to the cathode completes the circuit, allowing the current to return to the anode. The anode is the portion of a metal that corrodes. Pit ladders and off-shore oil platforms are examples. *Figure 2* shows similar conditions, except that the metals contact each other, eliminating the connector wire.

Oxygen is frequently involved in the corrosion process because most metals in their ore state are oxides. Copper, on the other hand, is typically found as copper sulfate, meaning that in the presence of sulfur, copper will readily react with sulfur to return to its natural state of copper sulfate. Among steel, aluminum, and copper, copper is the most cathodic or noble, and is typically less affected by galvanic reactions. However, copper does not get a free ride regarding corrosion. When used in contact with more noble metals, such as nickel, Type 316 and 304 stainless-steel alloys, or titanium, it will corrode by galvanic action just as easily as steel or aluminum.

If the electrical potential of a metal is the most important factor in determining the potential for galvanic action, then the cath-

ode-to-anode size ratio is probably the second most determining factor. A large mass of the more reactive metal tends to lessen the potential for a reaction. For instance, similar masses of stainless steel and structural steel will produce a galvanic reaction. However, when stainless steel fasteners (a smaller and less reactive mass) are in contact with structural steel (a larger and more reactive mass), the reaction is negligible. Stainless steel fasteners with aluminum windows (a larger and more reactive mass) produce similar results. Conversely, aluminum screws (a small and reactive mass) in a steel or stainless-steel assembly (large and less reactive mass) can produce severe corrosion. Early American metalsmiths understood this relationship very well as shown in *Photo 1* of an old copper pot with iron handles. Even though slight traces of rust bloom are visible on the iron handle, the mass ratio of the iron handle to that of the copper pot is large enough to prevent a severe

Photo 1 – Old copper pot and steel handle.

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TABLE 10—GUIDELINES FOR SELECTION OF FASTENERS BASED ON GALVANIC ACTION

Base Metal ↓ / Fastener Metal →	Zinc & Galvanized Steel	Aluminum & Aluminum Alloys	Steel & Cast Iron	Brasses, Copper, Bronzes, Monel	Martensitic Stainless Type 410	Austenitic Stainless Types 302/304, 303, 305
Zinc and Galvanized Steel	A	B	B	C	C	C
Aluminum and Aluminum Alloys	A	A	B	C	Not Recommended	B
Steel and Cast Iron	AD	A	A	C	C	B
Terne (Lead-Tin) Plated Steel Sheets	ADE	AE	AE	C	C	B
Brasses, Copper, Bronzes, Monel	ADE	AE	AE	A	A	B
Ferritic Stainless Steel (Type 430)	ADE	AE	AE	A	A	A
Austenitic Stainless Steel (Type 302/304)	ADE	AE	AE	AE	A	A

Key:

A. The corrosion of the base metal is not increased by the fastener.

B. The corrosion of the base metal is marginally increased by the fastener.

C. The corrosion of the base metal may be markedly increased by the fastener material.

D. The plating on the fastener is rapidly consumed, leaving the bare fastener metal.

E. The corrosion of the fastener is increased by the base metal.

NOTE:
Surface treatment and environment can change activity.

Table from Designer Handbook – "Stainless Steel Fasteners – A Systematic Approach to Their Selection" published by Special Steel Industry of North America

Table 3

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galvanic reaction. Also, notice the slight green passivation layer that has formed on the copper where the iron handle is connected to the pot. This is the protective layer that prevents further corrosion of the copper.

Since using the wrong fastener can cause corrosion, *Table 3* shows expected reactions with various metal combinations.

The galvanic corrosion process has led to using metals in the galvanic series for cathodic protection. Cathodic protection uses galvanic reaction principles in a positive way to prevent corrosion of a desired component. There is still corrosion, but the corrosion is controlled so that a predetermined metal is sacrificed. Galvanic protection is generally used when metal components are buried in earth or immersed in water, such as ships or offshore drilling rigs. Magnesium, zinc, and aluminum are metals often coupled to steel. For example, electrically connecting buried steel pipe with zinc, which is more anodic, allows the zinc to corrode before the steel. It is offered up as a sacrificial anode and limits the corrosion of the steel pipe until all of the zinc is consumed.

Impressed cathodic protection is usually a modified procedure. For instance, for a ship, anodes are strategically positioned on a ship's hull and emit an electrical current that suppresses current flow from the anodic areas of the hull. This converts the ship's hull to a cathode, removing/reducing any damage from corrosion—hence the term "impressed-current cathodic protection." The system continuously monitors and controls the impressed current output to provide optimum corrosion protection.

ATMOSPHERIC CORROSION OF STEEL

Salts are probably the most common components that affect electrolytes. Several

Photo 3 – Salt on vehicle.

Photo 2 – Salt on a highway.

Photo 4 – Corroded fire escape.

common sources of salts found in electrolytes are deicing salts and coastal salts. Road salts (most commonly sodium chloride) can become airborne on both wet and dry highways; and heavy, fast-moving traffic can send salt mist airborne like an aerosol spray. *Photos 2* and *3* show a mixture of deicing salt and sand on the highway and salt residue on a vehicle.

Prevailing winds and thermals can carry these airborne salts hundreds of feet and deposit them on unprotected structures. Once deposited, they remain dormant until moisture from rain, melting snow, fog, and condensation create an electrolyte or carry the salts to the structure's interior where, unobserved, they perform their destruction. The height that airborne salts can be carried up the building façade will vary depending on building shape and orientation, topography, landscaping, prevailing winds, etc. Regardless of the distance traveled, they can cause damage to improperly protected materials. *Photo 4* shows a corroded steel fire escape along a busy metropolitan street.

Sodium chloride, calcium chloride, and magnesium chloride are salts common to coastal environments that may also be found in deicing salts and may affect corrosion in multiple ways. For instance, since salt is considered to be hygroscopic (able to absorb moisture in the air) and does not have to wait for rain to begin its destructive path, it can simply absorb whatever moisture (including humidity) is in the surrounding air. A salt's chemistry can breach the protective oxide barrier that some metals, such as aluminum and zinc, form when exposed. Salts also act as a catalyst to an electrolyte, which increases the electrolyte's conductivity. As previously mentioned, during galvanic action, the rate of corrosion is directly proportional to its conductivity, and the atmosphere near the seashore is more con-

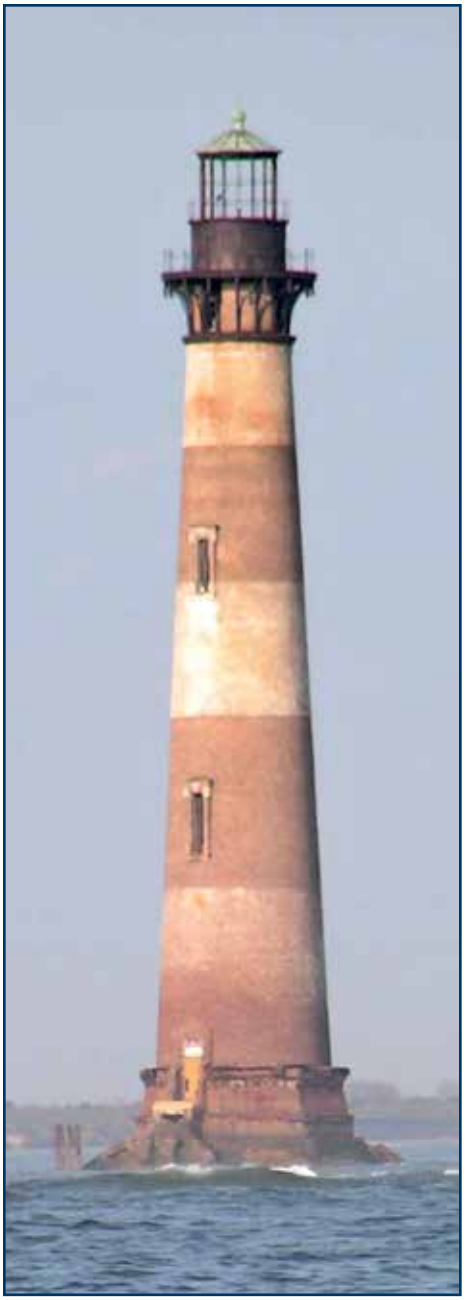

Photo 5 – Corroded lighthouse.

ductive because of the associated salt and is, therefore, more corrosive than inland. Similarly, condensate near the seashore is more brackish and thus is more conductive and a better electrolyte than an inland condensate. Steel typically corrodes or rusts rapidly in moist environments, and the rate of corrosion accelerates significantly if that moisture is salty, as shown in *Photo 5,* a coastal lighthouse.

Although this author is not aware of definitive tests or experiments to confirm or refute the supposition, it is possible that mixing sand with deicing salt, as is done in some localities, increases the airborne tendencies of the salt. This is because the salt particles can cling to the sand and are not absorbed by the concrete and are available to become airborne.

As previously discussed, ferrous oxide forms on the surface of iron or steel during the initial corrosion process and acts as a passivating layer. However, in the presence of water, oxygen, sulfur dioxide, and carbon dioxide solutions, the protective ferrous oxide layer becomes nonprotective rust, due to the formation of magnetic iron oxide called magnetite. Once this protective layer flakes off at the surface, fresh metal is exposed to corrosion. This process continues until either all of the iron or all of the oxygen, water, carbon dioxide, or sulfur dioxide in the system is removed or consumed.

Smokestacks often have a metal liner that can corrode when the furnace is shut down and the stack cools. As the stack cools, the surface temperature can fall below dew point, causing condensation to form on the surface. The condensation can react with carbon and sulfur deposits that may have accumulated on the stack lining from burning fossil fuels. This reaction can produce corrosive sulfurous acid that will drain down the lining and start a corrosion process with unprotected metal that is susceptible to corrosion.

Look for Part II of this article in an upcoming issue of Interface.

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