

# Calculation of Corrosion Rate from Corrosion Current (Faraday's Law)

*by*  
*James B. Bushman, P.E.*  
*Principal Corrosion Engineer*  
*Bushman & Associates, Inc*  
*Medina, Ohio USA*

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**BUSHMAN & Associates, Inc.**

CORROSION CONSULTANTS

P. O. Box 425 Medina, Ohio 44256

Phone: (330)769-3694 Fax: (330)769-2197

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Prepared by  
James B. Bushman, P.E.  
Bushman & Associates, Inc. – Corrosion Consultants  
PO Box 425 – Medina, Ohio 44258  
Phone: (330) 769-3694 -- Fax: (330) 769-2197

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Faraday's Law states the 96,486.7 Coulombs (equal to one Faraday) of charge transfer will oxidize or reduce One Gram Equivalent Weight of the material involved in the electrochemical reaction. Faraday developed the number of coulombs in what we today call a "FARADAY" by dividing Avogadro's Number ( $6.023 \times 10^{23}$  which equals the number of atoms of any specific atom whose weight equals its gram atomic weight) by the number of electrons in one coulomb (a coulomb is that amount of electrical charge equal to the charge carried by  $6.24 \times 10^{18}$  electrons):

$$\frac{6.023 \times 10^{23} \text{ Atoms} / \text{Gram Atomic Weight}}{6.24 \times 10^{18} \text{ Electrons} / \text{Coulomb}} \cong \frac{96,486.7 \text{ Coulombs}}{\text{Faraday}}$$

Faraday's Law is used to determine the corrosion rate of any species of material in "weight lost (or gained) per ampere of current flow per unit time". For example, if we want to know how many kilograms of Iron (Fe) will be corroded (oxidized) by a direct current discharge from the metal's surface into the surrounding electrolyte at a current flow of one ampere for one year, the calculation steps are as follows:

$$\frac{60 \text{ Seconds}}{\text{Minute}} \times \frac{60 \text{ Minutes}}{\text{Hour}} \times \frac{24 \text{ Hours}}{\text{Day}} \times \frac{365 \text{ Days}}{\text{Year}} \cong \frac{31,536,000 \text{ Seconds}}{\text{Year}}$$

By definition, 1 Coulomb = Ampere-Second ... and ... 1 Ampere = 1 Coulomb/Second

Therefore, if one ampere of current flows for one year, we calculate:

$$\frac{31,536,000 \text{ Seconds}}{\text{Year}} \times \frac{1 \text{ Coulomb}}{\text{Second}} \cong \frac{31,536,000 \text{ Coulombs}}{\text{Year}}$$

Since one coulomb will corrode (oxidize) one gram equivalent weight of Iron (Fe), we can divide the number of coulombs transferred in one year for a one ampere current flow by the number of coulombs that will corrode one gram equivalent weight of Iron (Fe) and:

$$\frac{31,536,000 \text{ Coulombs} / \text{Year}}{96,486 \text{ Coulombs} / \text{Gram}} \cong \frac{326.9 \text{ gram equivalent weights transferred}}{\text{year}}$$

The gram equivalent weight of Iron can be calculated as follows:

$$\begin{aligned}
 \text{GramEquivalentWeight}(Fe) &= \frac{\text{GramAtomicWeight}}{\text{No.ofElectronsTransferred} / \text{FeAtomCorroded}} \\
 &= \frac{55.85\text{Grams}(Fe)}{2(Fe \rightarrow Fe^{2+} + 2e^{-})} \\
 &= \frac{27.93\text{Grams}}{\text{GramEquivalentWeight}(Fe)}
 \end{aligned}$$

The final calculation to obtain the Faradaic consumption rate of iron (and approximately that for ordinary low carbon steels) is as follows:

$$\frac{27.93\text{Grams}(Fe)}{\text{Gram Eq.Wt.}} \times \frac{326.9\text{Gram Eq.Wt.Transferred}(Fe)}{\text{Ampere} - \text{Year}} \cong \frac{9,130\text{Grams}(Fe)}{\text{Ampere} - \text{Year}}$$

and

$$\frac{9,130\text{Grams}(Fe)}{\text{Ampere} - \text{Year}} \times \frac{1\text{Kg}}{\text{Gram}} \times \frac{2.2\text{Pounds}}{\text{Kg}} \cong \frac{20.1\text{Pounds Iron}}{\text{Ampere} - \text{Year}}$$

As shown above, 20.08 pounds of Iron will corrode (oxidize) each year for each ampere of positive (+) current flowing from the Iron (Fe) surface into the surrounding electrolyte. Using the above information, it is then a simple matter of (1) converting the density of steel to weight per unit square area per mil of thickness, (2) calculating how much metal weight will be corroded each year for the measured  $I_{\text{corr}}$ , and (3) converting this to mils of penetration per year based on the environmental conditions that existed at the time of the test.