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

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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 (6.023×10^{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 6.24×10^{18} electrons):

$$\frac{6.023\times10^{23}\,Atoms/Gram\,Atomic\,Weight}{6.24\times10^{18}\,Electrons/Coulomb} \cong \frac{96{,}486.7\,Coulombs}{Faraday}$$

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:

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

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

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

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

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:

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

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

$$GramEquivalentWeight(Fe) = \frac{GramAtomicWeight}{No.ofElectronsTransferred\,/\,FeAtomCorroded}$$

$$= \frac{55.85 Grams(Fe)}{2\left(Fe \rightarrow Fe^{++} + 2e^{-}\right)}$$

$$= \frac{27.93 Grams}{GramEquivalentWeight\left(Fe\right)}$$

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

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

and

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

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_{corr}, and (3) converting this to mils of penetration per year based on the environmental conditions that existed at the time of the test.

