Electrochemistry

Section 109 – Corrosion

109-1 Which member of each pair of metals is more likely to corrode (oxidize)?

(a) Mg or Ca
(b) Au or Hg
(c) Fe or Zn
(d) Ag or Pt

Solution

The metal most likely to corrode has the smaller (less positive or more negative) standard reduction potential: (a) Ca; (b) Hg; (c) Zn; and (d) Ag.

109-2 Consider the following metals: Ag, Au, Mg, Ni, and Zn. Which of these metals could be used as a sacrificial anode in the cathodic protection of an underground steel storage tank? Steel is an alloy composed mostly of iron, so use –0.447 V as the standard reduction potential for steel.

Solution

For cathodic protection, the overall reaction must be spontaneous. Thus, the standard cell potential of the anode must be less than the standard reduction potential of steel/iron (the metal to be protected), which, from the problem, has $E^{\circ}_{\text{cathode}} = -0.447$ V. With standard reductions in parentheses: Mg (–1.185 V), and Zn (–0.7618 V) could be used as a sacrificial anode, while Ag (+0.7996 V), Au (+1.629 V), and Ni (–0.257 V) could not.

109-3 Aluminum ($E^{\circ}_{\text{Al}^{3+}/\text{Al}} = -2.07$ V) is more easily oxidized than iron ($E^{\circ}_{\text{Fe}^{3+}/\text{Fe}} = -0.477$ V), and yet when both are exposed to the environment, untreated aluminum has very good corrosion resistance while the corrosion resistance of untreated iron is poor. What might explain this observation?

Solution

Untreated aluminum and untreated iron exposed to the environment both form an oxide layer. Aluminum oxide bonds tightly to the underlying aluminum, protecting it from further corrosion. The iron(III) oxide (rust) flakes off the underlying iron, exposing fresh metal that then corrodes. Thus as time passes, essentially all the aluminum remains unoxidized, while the fraction of the iron that remains unoxidized get progressively smaller.

109-4 If a sample of iron and a sample of zinc come into contact, the zinc corrodes but the iron does not. If a sample of iron comes into contact with a sample of copper, the iron corrodes but the copper does not. Explain this phenomenon.

Solution

Both examples involve cathodic protection. The (sacrificial) anode is the metal that corrodes (oxidizes or reacts). In the case of iron (–0.447 V) and zinc (–0.7618 V), zinc has a more negative standard reduction potential and so serves as the anode. In the case of iron and copper (0.34 V), iron has the smaller standard reduction potential and so corrodes (serves as the anode).

109-5 Suppose you have three different metals, A, B, and C. When metals A and B come into contact, B corrodes and A does not corrode. When metals A and C come into contact, A corrodes and C does not corrode. Based on this information, which metal corrodes and which metal does not corrode when B and C come into contact?

Solution

The metal with the smaller standard reduction potential from each pair acts as a sacrificial anode and corrodes, "protecting" the other metal from corrosion. Since B corrodes and protects A when they are in contact, $E_B^\circ < E_A^\circ$. Likewise, A protects C from corrosion, so $E_A^\circ < E_C^\circ$. Combining these gives $E_B^\circ < E_A^\circ < E_C^\circ$. This means that B corrodes and protects C (which does not corrode) when B and C come into contact.

109-6 Why would a sacrificial anode made of lithium metal be a bad choice?

Solution

While the reduction potential of lithium would make it capable of protecting the other metals, this high potential is also indicative of how reactive lithium is: It would have a spontaneous reaction with most substances. This means that the lithium would react quickly with other substances, even those that would not oxidize the metal it is attempting to protect. Reactivity like this means the sacrificial anode would be depleted rapidly and need to be replaced frequently. (Optional additional reason: fire hazard in the presence of water.)