XVI) Corrosion of Iron

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What must be present in order for iron to rust?

$H_2O + O_2$

"rusting"

Iron corroding is actually a redox reaction. What do you think happens to the Fe metal?

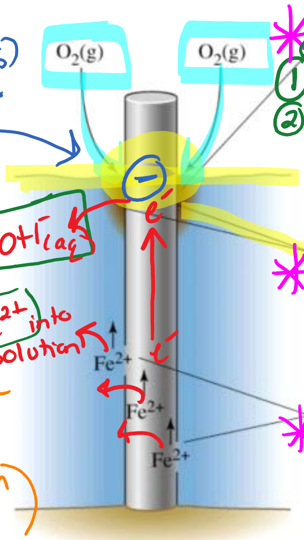
$Fe_{(s)} \longrightarrow Fe^{2+} + 2e^-$ oxidizes to form the iron ion. (Anode)

Where do the resulting electrons go? The e^- will flow to the cathode for a reduction rxn. (reduce O_2)

$O_{2(g)} + 2H_2O + 4e^- \longrightarrow 4OH^-$

So what serves as the anode? The cathode?

$Fe_{(s)}$ serves as BOTH anode AND cathode



"cathode" end of Fe(s) ⊖ charged.

$OH^-_{(aq)}$

e^-

Fe^{2+} into solution

dissolve in H_2O, this helps to maintain +/- neutral charge. (also OH^- in solution)

Rust formation:

① $Fe^{2+}(aq) + 2\,OH^-(aq) \longrightarrow Fe(OH)_2(s)$

② $4\,Fe(OH)_2(s) + O_2(g) + 2\,H_2O(l) \longrightarrow 4\,Fe(OH)_3(s)$ ③ decomposes into:

$FeO(OH)$ H_2O

④ $2\times FeO(OH)$ combine

$Fe_2O_{3(s)} + H_2O_{(l)}$ RUST

(also hydrated form: $Fe_2O_3 \cdot xH_2O$)

Cathode: (reduction rxn)

$O_2(g) + 2\,H_2O(l) + \boxed{4\,e^-} \longrightarrow 4\,OH^-(aq)$

Anode: (oxidation)

$Fe(s) \longrightarrow Fe^{2+}(aq) + 2\,e^-$

Anode Half-Reaction: $Fe \Rightarrow Fe^{2+} + 2e^-$

Cathode Half-Reaction: $O_2 + 2H_2O + 4e^- \Rightarrow 4OH^-$

The Fe^{2+} created at the anode follows the electrons to the cathode and dissolve in the water present. This helps maintain neutrality as the negatively charged OH^- is produced at the cathode. In fact, Fe^{2+} combines with the hydroxide ion, producing solid iron (II) hydroxide.

$$Fe^{2+} + 2OH^- \Rightarrow Fe(OH)_2$$

← ion combining in solution is what starts the rust formation process.

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steps (1) - (4) formation of rust.

The iron (II) hydroxide then reacts with oxygen and water to make $Fe(OH)_3$. Then $Fe(OH)_3$ decomposes to $FeO(OH) + H_2O$, and then two FeO(OH) molecules collide to make $Fe_2O_3 + H_2O$. Rust is actually Fe_2O_3 as well as its hydrated form, $Fe_2O_3 \bullet xH_2O$. This accounts for the different colours in rust.

Corrosion cannot occur in dry air or in oxygen depleted water (deep water). Why?

• "dry air" - no H_2O present (O_2 reduction not occur)
• oxygen depleted water - no $O_{2(g)}$ ∴ no reduction rxn.

Protection from Corrosion

There are two protection types: Physical (A) and Electrochemical (B)

(A) **Physical Protection**

What does physical protection from corrosion mean?

There is a physical layer on top of iron, prevent exposure to $H_2O_{(l)} + O_{2(g)}$ (no reduction rxn.)

What are some different types of physical protection?

• treated coating
• paint
• galvanized
• plastic covering
• grease
• chrome plating (Sn, Zn)

(B) **Electrochemical Protection**

What does electrochemical protection from corrosion mean?

e^- are provided for O_2 to reduce by some other source => not Fe.

1. cathodic Protection
2. Electric current supply.

There are two types of electrochemical protection:

1. Cathodic Protection

Why do people put zinc strips on the iron hull of boats?

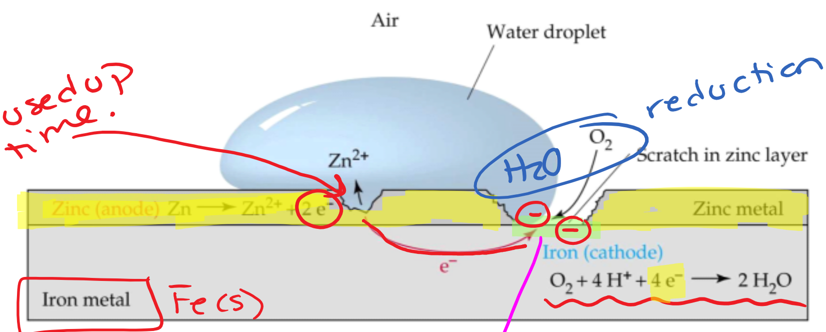
Zinc oxidizes more readily than iron (check your table). Therefore, if zinc is available, it will oxidize to Zn^{2+} before any Fe will oxidize. Attaching zinc strips to iron will prevent iron from rusting as the zinc will oxidize first, and the electrons from the zinc will conduct through the iron to the site of reduction (where the water and oxygen is). The zinc is called a **sacrificial anode** (it sacrifices itself for the iron). Zinc provides Fe with electrons and Fe acts as the cathode to conduct electrons to O_2 and H_2O. The the zinc does not have to completely cover the Fe.

The preferred rxn will oxidize 1st.

b/c when Fe + Zn are both present, Zn will always oxidize first.

(oxidation) Anode Half-Reaction: $Zn \Rightarrow Zn^{2+} + \boxed{2e^-}$ → $Fe_{(s)}$ → $O_2 + H_2O$
(reduction) Cathode Half-Reaction: $O_2 + 2H_2O + 4e^- \Rightarrow 4OH^-$

The oxidation of the Zn will decrease its mass. Regularly replacing the sacrificial anode (Zn strips) will ensure that the Fe will not corrode.



2. Supplying an electric current to the iron

Some ships will supply a low voltage electric current (a stream of electrons) to the iron hull. This prevents the iron hull from having to oxidize to supply electrons to oxygen and water, thereby preventing corrosion. (of $Fe_{(s)}$)

→ e^- provided to $O_2 + H_2O$ for reduction

Assignment 15: Corrosion Exercises

1. Can Zn be used as a sacrificial anode for Al? Explain.
2. Would it be smart to use a tin based paint on the bottom of an aluminum boat? Why or why not?
3. The Statue of Liberty consists of an iron frame covered with copper. Discuss the reasons for the rapid corrosion in this structure.
4. Why hasn't the Titanic corroded?
5. To prevent an environmental disaster, how could you stop an underground iron septic tank from rusting through without actually having to dig it up?
6. Why does most car rust start in or around the wheel wells?