Chapter 17 Part 5

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Batteries

- A battery is an electrochemical cell or series of cells that produces an electric current
- There are 2 types of batteries
  - Primary
  - Secondary

Primary and Secondary Batteries

Primary Battery

- Primary batteries are single use batteries that cannot be recharged
- One example is an alkaline battery

Secondary Battery

- Secondary batteries are rechargeable
- One example is a lead acid battery

Fuel Cells

- A Fuel cell is a device that converts chemical energy into electrical energy
- Fuel cells require a continuous source of fuel and will produce electricity as long as fuel is available
- One example is a hydrogen fuel cell

Anode: $2\ H_2(g) + 2\ O^{2-}(aq) \rightarrow 2\ H_2O(l) + 4e^-$

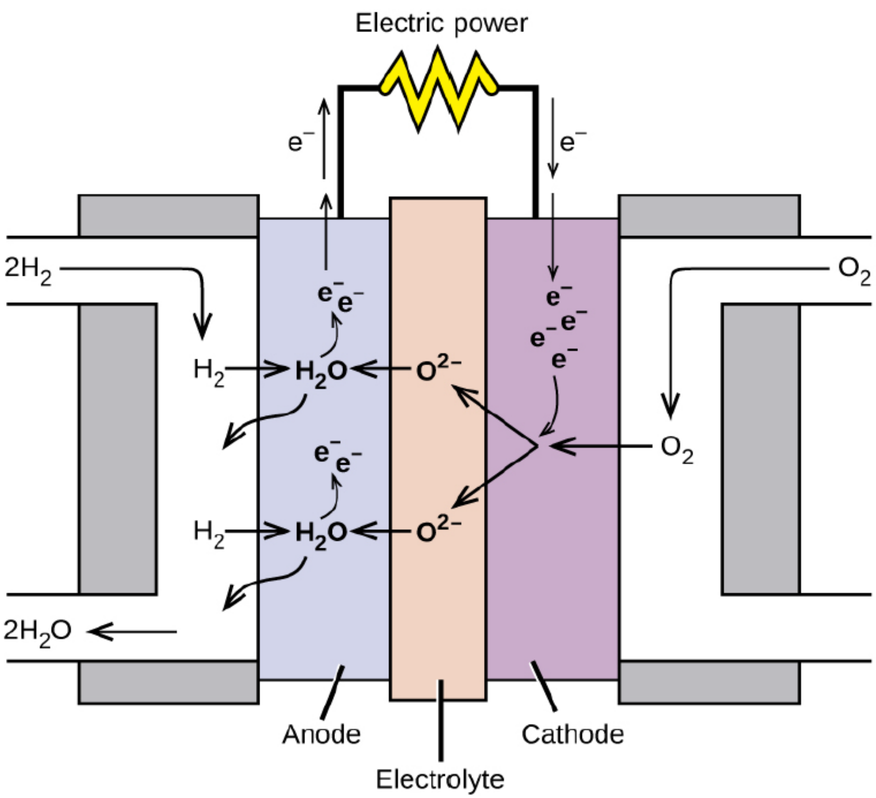
Cathode: $O_2(g) + 4e^- \rightarrow 2O^{2-}(aq)$

Overall: $2\ \overset{0}{H_2}(aq) + \overset{0}{O_2}(aq) \rightarrow 2\ \overset{+1}{H_2}\overset{-2}{O}(l)$

Hydrogen Fuel Cell

Anode: $2H_2(g) + 2O^{2-}(aq) \rightarrow 2H_2O(l)4e^-$
Cathode: $O_2(g) + 4e^- \rightarrow 2O^{2-}(aq)$

Overall: $2\ H_2(aq) + O_2(aq) \rightarrow 2\ H_2O(l)$



Types of Cells

Identify the type of cell in which there is a continuous flow of reactants and products.

A. dry cell
B. wet cell
C. fuel cell
D. lead storage cell

Corrosion

- Corrosion is the degradation of metals due to an electrochemical process
- One example of corrosion is the formation of rust on iron

Ways to Prevent Corrosion

1. Painting the surface of the metal
2. Galvanization
3. Cathodic protection using sacrificial anodes

Painting the surface of the metal

- Adding a layer of paint prevents the water and oxygen necessary for rust formation from coming into contact with the iron

Rusting of Iron

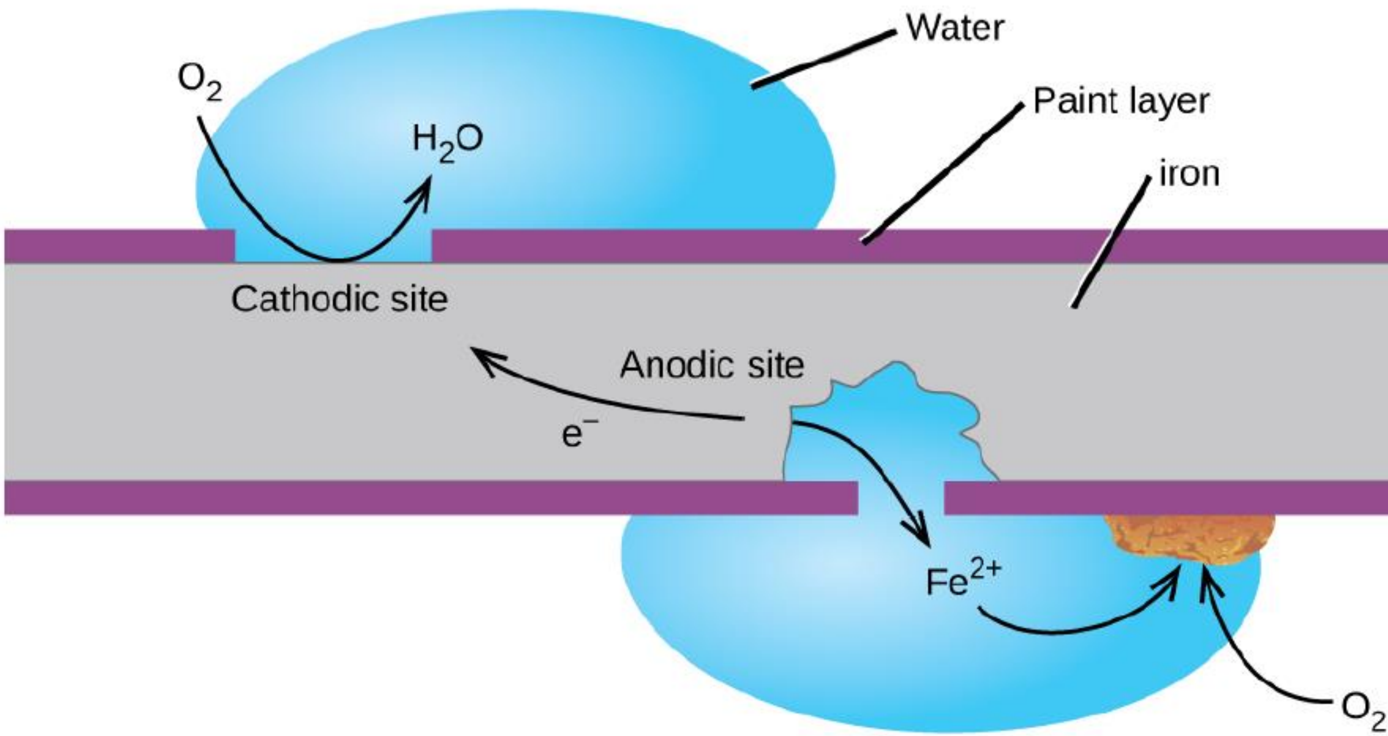


Figure 17.17 Once the paint is scratched on a painted iron surface, corrosion occurs and rust begins to form. The speed of the spontaneous reaction is increased in the presence of electrolytes, such as the sodium chloride used on roads to melt ice and snow or in salt water.

Metal Activity

- Metals which readily give up electrons (are oxidized easily) are said to be more active
- Accordingly, active metals have very low reduction potentials

Metal Activity

Using the reduction potential data, rank the following elements in order from most active to least active.

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s) \quad E° = +0.800$$

$$Fe^{2+}(aq) + 2e^{-} \rightarrow Fe(s) \quad E° = -0.440$$

$$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g) \quad E° = 0.000$$

$$Mg^{2+}(aq) + 2e^{-} \rightarrow Mg(s) \quad E° = -2.356$$

Metal Activity

Identify the strongest reducing agent

A. Cu^{+}
B. Cu
C. Al^{3+}
D. Al

Half-Reaction	E°_{red} (V)
$Cu^{+}(aq) + e^{-} \rightarrow Cu(s)$	0.52
$Fe^{2+}(aq) + 2\ e^{-} \rightarrow Fe(s)$	– 0.44
$Al^{3+}(aq) + 3\ e^{-} \rightarrow Al(s)$	– 1.66

Galvanization

- Iron can be electroplated with zinc. This means that electric current can be used to form thin metal coating on zinc.
- Zinc is more easily oxidized than iron because zinc has a lower reduction potential (meaning that zinc is a more active metal)

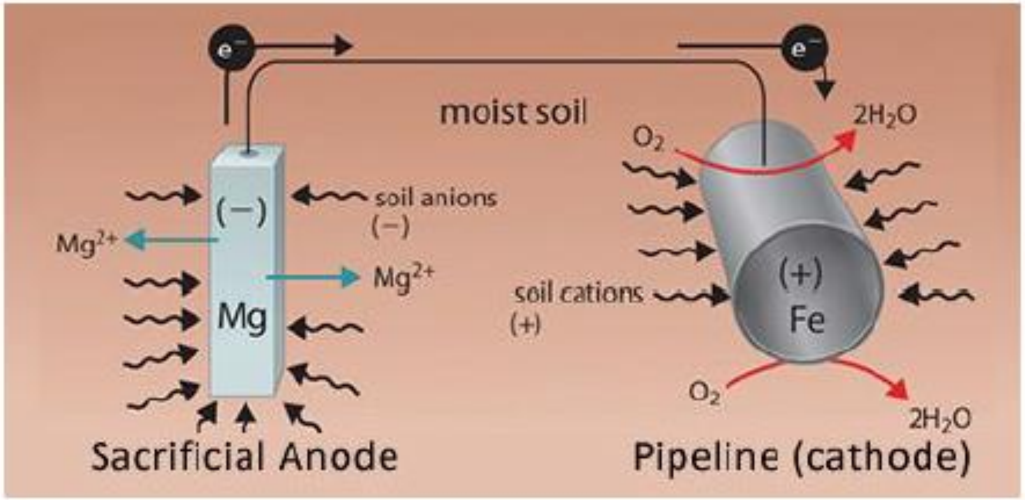
$$Fe^{2+}(aq) + 2\ e^{-} \rightarrow Fe(s) \qquad E^{\circ}_{cell} = -0.447$$

$$Zn^{2+}(aq) + 2\ e^{-} \rightarrow Zn(s) \qquad E^{\circ}_{cell} = -0.762$$

- Even if the zinc surface is scratched, the zinc will still oxidize before the iron

Cathodic Protection

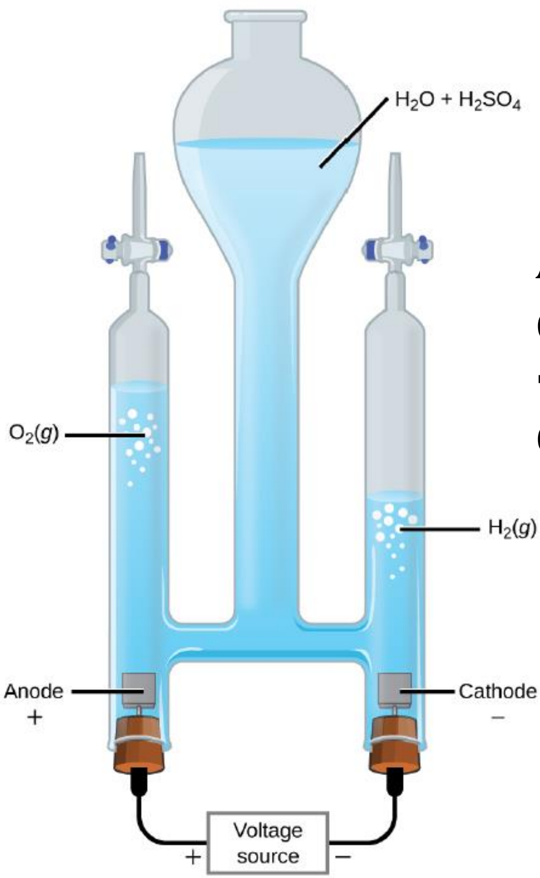
- If a wire is used to connect a metal to a more active metal (a metal with a lower reduction potential), the more active metal will corrode (oxidize) first.
- Since the more active metal is corroded in order to protect the less active metal, the more active metal is called a sacrificial anode.



Electrolysis

- Electrolysis is a process where batteries can be used to cause nonspontaneous reactions
- These reactions take place because batteries are used to funnel electrons into the more active metal to force its reduction
- Since they are nonspontaneous, they will have a negative E°_{cell} value.
- The amount of voltage required by the battery in order to cause this nonspontaneous process to occur is $-E^{\circ}_{cell}$

Electrolysis of Water



$$\begin{array}{lll}
\text{Anode:} & 2\,H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^- & E^\circ_{H_2O/O_2} = -1.23\text{ V} \\
\text{Cathode:} & 2\,H_2O(l) + 2e^- \rightarrow H_2(g) + 2\,OH^-(aq) & E^\circ_{H_2O/H_2} = -0.83\text{V} \\
\hline
\text{Overall:} & 2\,H_2O(l) \rightarrow 2\,H_2(g) + O_2(g) & E^\circ_{cell} = -2.06\text{ V}
\end{array}$$

Electrolysis

Select the best description of electrolysis.

A. A flow of electrons is used to cause a nonspontaneous chemical reaction to occur.
B. A flow of electrons is used increase the speed of a spontaneous chemical reaction.
C. A spontaneous chemical reaction is used to create electricity.
D. A spontaneous chemical reaction is used to create a flow of electrons.

Quantitative Aspects of Electrolysis

Using Faraday's constant

$$1 \text{ mol e}^- = 96{,}485 \text{ C}$$

Also

$$\text{Charge(C)} = \text{current}\left(\frac{\text{C}}{\text{s}}\right) \times \text{time(s)}$$

Here, current is in amperes which is equivalent to charge per second

To determine the number of moles of electrons involved in an electrolysis reaction

$$\text{number of mol e}^- = \text{current}\left(\frac{\text{C}}{\text{s}}\right) \times \text{time(s)} \times \frac{1 \text{ mol e}^-}{96{,}485 \text{ C}}$$

Quantitative Aspects of Electrolysis

The electrodeposition of copper can be used to determine the copper content of a sample. The sample is dissolved to produce $Cu^{2+}(aq)$, which is electrolyzed. At the cathode, the reduction half-reaction is $Cu^{2+}(aq) + 2\ e^{-} \rightarrow Cu(s)$. What mass of copper can be deposited in 1.00 hour by a current of 1.62 A (C/s)?

Quantitative Aspects of Electrolysis

If 12.3 g of Cu is deposited at the cathode of an electrolytic cell after 5.50 h, what was the current used?