**Electrolytic Cells (ELC’s)**

**Electrolysis** – uses an external power source to cause a **non-spontaneous redox reaction** to occur.

In BOTH Electrochemical Cells (ECC’s) and Electrolytic Cells (ELC’s):

- **OXIDATION** takes place at the **ANODE** (LEOA)
- **REDUCTION** takes place at the **CATHODE** (GERC)

There are three main types of **Electrolytic Cells**:

1. Electrolysis of **Molten** Salts (no H₂O) with **Unreactive** (Inert) Electrodes
2. Electrolysis of **Aqueous** Salts (H₂O solution) with **Unreactive** (Inert) Electrodes
3. Electrolysis of **Aqueous** Salts (H₂O solution) with **Reactive** Electrodes

**Type 1 - Electrolysis of Molten Salts (no H₂O) with Unreactive (Inert) Electrodes**

All **molten** (melted) salts consist of **mobile** ions.

**Eg.**  
\[ \text{NaI}(l) \rightarrow \text{Na}^+(l) + \text{I}^-(l) \]

\[ \text{NaCl}(l) \rightarrow \text{Na}^+(l) + \text{Cl}^-(l) \]
An example of a “Type 1” ELC: Electrolysis of molten NaCl ($NaCl(l)$)

Diagram:

**Cathode Half-Reaction (Reduction of the Cation)**
Looking near the bottom of the Reduction Table, we see the half-reaction for the reduction of $Na^+$:

$Na^+ + e^- \rightarrow Na(s) \quad E^0 = -2.71 \text{ v}$

**Anode Half-Reaction (Oxidation of the Anion)**
To write the oxidation of $Cl^-$, we find $Cl^-$ on the Right Side and REVERSE the half-reaction:

$2Cl^- \rightarrow Cl_2(g) + 2e^- \quad E^0 = -1.36 \text{ v}$ (the sign on the $E^0$ is changed since the rx. is reversed.)

To find the overall redox reaction with its $E^0$, we add up the half reactions as follows:

$$
\begin{align*}
\text{Reduction at Cathode} & : & (Na^+ + e^- \rightarrow Na(s)) & \quad E^0 = -2.71 \text{ v} \\
\text{Oxidation at Anode} & : & 2Cl^- \rightarrow Cl_2(g) + 2e^- & \quad E^0 = -1.36 \text{ v} \\
\text{Overall REDOX reaction} & : & 2Na^+ + 2Cl^- \rightarrow 2Na(s) + Cl_2(g) & \quad E^0 = -4.07 \text{ v}
\end{align*}
$$

The Product at the Cathode is $Na(s)$
The Product at the Anode is $Cl_2(g)$

The industrial application of this cell is called a **Down’s Cell**.
A quick method for writing half-reactions for Type 1 Electrolytic Cells (electrolysis of molten salts with unreactive electrodes)

1. Dissociate the salt into its ions
2. Underneath the Cation, write $\text{C}^-$ (representing the Cathode)
3. Underneath the Anion, write $\text{A}^+$ (representing the Anode)
4. Draw arrow showing where each ion goes
5. The half-reactions must start with the ion (the ion must be the reactant)
6. For the Cathode Half-Reaction, write the reduction of the cation (same as on table)
7. For the Anode Half-Reaction, write the oxidation of the anion (reversed from table)

Example: For the Electrolysis of Molten Potassium Iodide (KI(l))

$\text{KI} \rightarrow \text{K}^+ + \text{I}^-$

C-        A+

Cathode Half-Reaction:

$\text{________} \rightarrow \text{________} \quad E^0 = \text{________} \text{v}$

Anode Half-Reaction:

$\text{________} \rightarrow \text{________} \quad E^0 = \text{________} \text{v}$

Overall Redox Reaction:

$\text{________} \rightarrow \text{________} \quad E^0 = \text{________} \text{v}$

Product at Cathode _______  Product at Anode _______ Minimum Voltage Needed _______v

Sketch this cell, labeling everything:
**Type 2 Electrolytic Cells – Electrolysis of Aqueous Salts and Non-reactive Electrodes**

Because the salts are aqueous, **water** is present. Looking at the *top and bottom shaded lines* on the reduction table, you can see that water can be oxidized or reduced:

\[
\begin{align*}
\text{NO}_3^- + 4\text{H}^+ + 3e^- & \rightarrow \text{NO}_2\text{H} + 2\text{H}_2\text{O} \quad \text{E}^0 = +0.96 \\
\text{H}_2\text{O}^+ + 2e^- & \rightarrow \text{H}_2\text{O} \quad \text{E}^0 = +0.83 \\
\frac{1}{2}\text{O}_2 + 2\text{H}^+ (10^{-7} \text{M}) + 2e^- & \rightarrow \text{H}_2\text{O} \quad \text{E}^0 = +0.82 \\
\text{NO}_3^- + 4\text{H}^+ + 2e^- & \rightarrow \text{NO}_2\text{H} + 2\text{H}_2\text{O} \quad \text{E}^0 = +0.90
\end{align*}
\]

This half-reaction, the way it’s written, is the **reduction of water** and could take place at the **CATHODE**.

The **REVERSE** of this reaction:

\[\text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2 + 2\text{H}^+ + 2e^- \quad \text{E}^0 = -0.82 \text{ v}\]

is the **oxidation of** water and could take place at the **ANODE**.

**To find the Half-Reaction at the CATHODE (reduction)**

Due to the “Overpotential Effect”, the “Arrow on the Left” tells us whether:

1. The Cation is reduced? or
2. Water is reduced?

So when **Cu**\(^{2+}\) is in solution, what will be reduced at the cathode, **Cu**\(^{2+}\) or **H**\(_2\)\(\text{O}\)? _____________

Write an equation for the half-reaction at the Cathode: _________________ \[\text{E}^0 = \] _____________

So when **Mg**\(^{2+}\) is in solution, what will be reduced at the cathode, **Mg**\(^{2+}\) or **H**\(_2\)\(\text{O}\)? _____________

Write an equation for the half-reaction at the Cathode: _________________ \[\text{E}^0 = \] _____________
To find the Half-Reaction at the ANODE (Oxidation)

Look at the “overpotential arrow on the RIGHT side of the table

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Half-Reaction at Anode</th>
<th>E°</th>
</tr>
</thead>
<tbody>
<tr>
<td>SO₄²⁻ + 2e⁻</td>
<td>SO₄²⁻</td>
<td>-0.18</td>
</tr>
<tr>
<td>Fe⁺⁺ + 2e⁻</td>
<td>Fe⁰</td>
<td>+0.87</td>
</tr>
<tr>
<td>H₂O₂ + 2H⁺ + 2e⁻</td>
<td>H₂O + 2H⁺</td>
<td>+1.78</td>
</tr>
<tr>
<td>MnO₄⁻ + 8H⁺ + 5e⁻</td>
<td>Mn²⁺ + 4H₂O</td>
<td>+1.51</td>
</tr>
<tr>
<td>Ag⁺⁺ + 2e⁻</td>
<td>Ag⁰</td>
<td>+1.50</td>
</tr>
<tr>
<td>Br₂ + 2e⁻</td>
<td>Br⁻</td>
<td>+1.08</td>
</tr>
<tr>
<td>ClO₃⁻ + 8H⁺ + 6e⁻</td>
<td>Cl⁻ + 4H₂O</td>
<td>+1.39</td>
</tr>
<tr>
<td>Cl₂ + 2e⁻</td>
<td>Cl⁻</td>
<td>+1.36</td>
</tr>
<tr>
<td>NO₃⁻ + 4H⁺ + 3e⁻</td>
<td>NO₂⁻ + 2H₂O</td>
<td>+0.96</td>
</tr>
<tr>
<td>Hg²⁺ + 2e⁻</td>
<td>Hg⁰</td>
<td>+0.85</td>
</tr>
</tbody>
</table>

Any anion ABOVE the arrow WILL NOT be oxidized in aqueous solution. The water will be oxidized instead:

\[
2 \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2(g) + 2 \text{H}^+ + 2 \text{e}^-
\]

(reverse the reaction at +0.82 v)

Any anion BELOW the arrow WILL be oxidized in aqueous solution:

\[
\begin{align*}
2 \text{Cl}^- & \rightarrow \text{Cl}_2 + 2\text{e}^- \\
2 \text{Br}^- & \rightarrow \text{Br}_2 + 2\text{e}^-
\end{align*}
\]

So when SO₄²⁻ is in solution, what will be oxidized at the ANODE, SO₄²⁻ or H₂O? _________

Write an equation for the half-reaction at the **Anode**: ________________________ E°=____

So when I⁻ is in solution, what will be oxidized at the ANODE, I⁻ or H₂O? _________

Write an equation for the half-reaction at the **Anode**: ________________________ E°=____

A method for finding the Half-Reactions (Oxidation at the Anode) and (Reduction at the Cathode) in this type of cell is:

1. Dissociate the salt into it’s ions
2. Write “H₂O” between the two ions.
3. Underneath the Cation, write(C₂) (representing the Cathode)
4. Underneath the Anion, write(A⁺) (representing the Anode)
5. Using the rules about the “Overpotential Arrows” determine what is reduced at the Cathode (the cation or water) and what is oxidized at the Anode (the anion or water).
6. Complete the half-reactions at the Cathode and the Anode (with their E°’s)
7. Add half-reactions to get the overall redox reaction if you are asked to.

Let’s do an example: A major industrial process is the electrolysis of brine (NaCl \(_{(aq)}\)):

\[
\begin{align*}
\text{Cathode:} & \quad 2 \text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2(g) + 2 \text{OH}^- \quad E^\circ = -0.41 \text{ v} \\
\text{Anode:} & \quad 2 \text{Cl}^- \rightarrow \text{Cl}_2(g) + 2e^- \quad E^\circ = -1.36 \text{ v}
\end{align*}
\]

Now, we must look at the Reduction Table:
For the Cathode, look at the “overpotential arrow” on the LEFT side. Notice that the Na\(^+\) is BELOW the arrow. This means Na\(^+\) will NOT be reduced. The water will be reduced instead. The half-reaction for the reduction of water is at -1.41 v

\[
\text{Cathode Half-Reaction:} \quad 2 \text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2(g) + 2 \text{OH}^- \quad E^\circ = -0.41 \text{ v}
\]

For the Anode, look at the “overpotential arrow” on the RIGHT side. Notice that the Cl\(^-\) is BELOW the arrow. This means Cl\(^-\) WILL be oxidized (rather than H\(_2\)O). So…

\[
\text{Anode Half-Reaction:} \quad 2 \text{Cl}^- \rightarrow \text{Cl}_2(g) + 2e^- \quad E^\circ = -1.36 \text{ v}
\]

**Overall Redox Reaction:**

\[
2 \text{Cl}^- + 2 \text{H}_2\text{O} \rightarrow \text{Cl}_2(g) + \text{H}_2(g) + 2 \text{OH}^- \quad E^\circ = -1.77 \text{ v}
\]

*The Products at the Cathode* would be H\(_2\)(g) + 2 OH\(^-\) (the pH near the cathode would increase)
*The Product at the Anode* would be Cl\(_2\)(g)
There would be bubbles observed at both electrodes! Draw them into the diagram above!

The **minimum voltage required** to carry this reaction out would be 1.77 v (just enough to overcome the -1.77 v E\(^\circ\))

Now, try to complete the examples on the next page….
For the electrolysis of aqueous CuCl₂ using platinum (inert) electrodes. Find:

The half-reaction at the Cathode: ____________________________ \( E^0 = \) ________

The half-reaction at the Anode: ____________________________ \( E^0 = \) ________

The overall redox reaction: _______________________________ \( E^0 = \) ________

Product(s) at the Cathode: _______________  Product(s) at the Anode _______________

The minimum voltage required: ________________ \( \text{v} \)

For the electrolysis of Na₂SO₄(aq) using carbon (inert) electrodes. Find:

The half-reaction at the Cathode: ____________________________ \( E^0 = \) ________

The half-reaction at the Anode: ____________________________ \( E^0 = \) ________

The overall redox reaction: _______________________________ \( E^0 = \) ________

Product(s) at the Cathode: _______________  Product(s) at the Anode _______________

The minimum voltage required: ________________ \( \text{v} \)

For the electrolysis of CuSO₄(aq) using inert electrodes. Find:

The half-reaction at the Cathode: ____________________________ \( E^0 = \) ________

The half-reaction at the Anode: ____________________________ \( E^0 = \) ________

The overall redox reaction: _______________________________ \( E^0 = \) ________

Product(s) at the Cathode: _______________  Product(s) at the Anode _______________

The minimum voltage required: ________________ \( \text{v} \)
**Type 3 Electrolytic Cells – Electrolysis of Aqueous Salts with Reactive Electrodes**

In this type of cell, the electrodes are normal **metals**, not inert ones like platinum or carbon.

Something that’s important:

The metal that the **Cathode** is made from is **NOT reduced**. That’s because **metals cannot gain electrons** and become negative metal ions (no such ion as Fe\(^{2-}\)!)  

The **Cathode** metal is **NOT oxidized**. That’s because **oxidation does not take place at the cathode**!  

To summarize: **THE CATHODE METAL NEVER REACTS!!**

The Cathode only supplies the **SURFACE** for the **reduction of the Cation** (+ ion) or for the **reduction of water** (depending on whether the cation is above or below the left overpotential arrow.) Here’s an example:

An **aqueous** solution containing the Cu\(^{2+}\) ion is electrolyzed. The cathode is made of **Iron**. Write the equation for the Half-Reaction taking place at the **Cathode**:

The Iron will not react! *(It only provides the surface—so there is no half-rx. involving Fe!)*

The Cu\(^{2+}\) ion is ABOVE the overpotential arrow on the left, so Cu\(^{2+}\) will be reduced!

The Half-Reaction at the Cathode would be:  

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}(s) \quad \text{E}^0 = +0.34 \text{ v} \]

When this happens, the Iron Cathode will become coated with copper, and will turn reddish in colour. (But remember, the iron itself does NOT react!)

Here’s one for you:
An **aqueous** solution containing the Mg\(^{2+}\) ion is electrolyzed. The cathode is made of **Nickel**. Write the equation for the Half-Reaction taking place at the **Cathode**:

_____________________________________________________________________

What would you observe at the cathode? __________________________

_____________________________________________________________________

_____________________________________________________________________

**Chemistry 12—Notes on Electrolytic Cells**

Page 48
In a Type 3 cell, there are 3 possibilities for Oxidation at the ANODE:

1. The anion in the solution is oxidized
2. Water is oxidized
3. The Anode Metal is oxidized

The one with the Highest Oxidation Potential (The LOWEST one on the RIGHT of the table) will be the one that is Oxidized. (Treat water as if it were at the right overpotential arrow.)

\[
\begin{align*}
\text{BrO}_3^- + 6\text{H}^+ + 5\text{e}^- & \rightarrow \frac{1}{2}\text{Br}_2 + 3\text{H}_2\text{O} \quad E^{\circ} = +1.48 \\
\text{ClO}_4^- + 8\text{H}^+ + 6\text{e}^- & \rightarrow 3\text{Cl}^- + 4\text{H}_2\text{O} \quad E^{\circ} = +1.39 \\
\text{Cl}_2\text{O}_3^{4-} + 2\text{e}^- & \rightarrow 2\text{Cl}^- + 7\text{H}_2\text{O} \quad E^{\circ} = +1.36 \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} \quad E^{\circ} = +1.23
\end{align*}
\]

Let's do an example:

An aqueous solution containing the Cl\(^-\) ion is electrolyzed. The anode is made of silver. Write the equation for the Half-Reaction taking place at the Anode:

\[
\begin{align*}
\text{BrO}_3^- + 6\text{H}^+ + 5\text{e}^- & \rightarrow \frac{1}{2}\text{Br}_2 + 3\text{H}_2\text{O} \quad E^{\circ} = +1.48 \\
\text{ClO}_4^- + 8\text{H}^+ + 6\text{e}^- & \rightarrow 3\text{Cl}^- + 4\text{H}_2\text{O} \quad E^{\circ} = +1.39 \\
\text{Cl}_2\text{O}_3^{4-} + 2\text{e}^- & \rightarrow 2\text{Cl}^- + 7\text{H}_2\text{O} \quad E^{\circ} = +1.36 \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} \quad E^{\circ} = +1.23
\end{align*}
\]

The one with the highest oxidation potential. (The lowest on the RIGHT side) is Ag (the anode)

So the Half-Reaction at the Anode is the oxidation of Ag: \(\text{Ag(s)} \rightarrow \text{Ag}^+ + \text{e}^- \quad E^{\circ} = -0.80\text{v}\)
Here’s an example question. Given the following cell:

Identify the Anode ____________________ the Cathode ____________________

Write an equation for the half-rx. at the Anode ____________________ \( E^0 = \) ______

Write an equation for the half-rx. at the Cathode ____________________ \( E^0 = \) ______

Write an equation for the overall redox rx. ____________________ \( E^0 = \) ______

Product at Anode _____ Product at Cathode _____ Minimum voltage necessary _______

**Electrorefining**

In Electrorefining, an impure metal is refined by making it the ANODE. The pure metal is the CATHODE.

Eg.) An ANODE made of IMPURE COPPER could have something like the following make-up:
- Mostly copper
- Some Ag or Au (above Cu on the reduction table)
- Some Zn (below Cu on the reduction table)
- Some “dirt” —any non-metal impurities

The CATHODE would be made of PURE COPPER

The SOLUTION is an aqueous solution of a compound containing \( \text{Cu}^{2+} \) (eg CuSO\(_4\))

Remember, for Electrorefining: IMPURE ANODE ⇒ PURE CATHODE (PC)
Use your Reduction Table to help you explain all the information shown in the diagram:

First, we look at the **ANODE**:

- **Au** and **Ag** will NOT oxidize as long as **Cu** is present. As the **Cu** oxidizes around it, they fall down, with the dirt, to become part of the Anode Mud (sludge).

- **Zn** will oxidize first (highest Ox. Potential), putting **Zn^{2+}** ions into the solution.

- Once **Zn** is gone, **Cu^{2+}** (the next highest Ox. Potential) will start oxidizing:
  \[ \text{Cu} \rightarrow \text{Cu}^{2+} + 2e^- \]

- The solution is **CuSO_4(aq)**. So there are lots of **Cu^{2+}** ions there to begin with.

- The solution contains only **Cu^{2+}**, **SO_4^{2-}** and **Zn^{2+}** ions. Metals above **Cu** (Au & Ag) have fallen to the Anode mud (not oxidized into ions!)
Now, let’s look at the **CATHODE**

Coming from the Anode

An Excess of e’s are forced onto the Cathode by the external power supply. The Cathode thus becomes Negative.

Only \( \text{Cu}^{2+} \) will be reduced at the Cathode. The \( \text{SO}_4^{2-} \) ion is negative so will not go to the negative cathode. The \( \text{Zn}^{2+} \) ions have a Lower Reduction Potential than \( \text{Cu}^{2+} \), so they would not be reduced until all the \( \text{Cu}^{2+} \) is gone. But that NEVER happens. If \( \text{Cu}^{2+} \) runs low, the anode is replaced with another one and more \( \text{Cu}^{2+} \) is produced. So ions like \( \text{Zn}^{2+} \) will stay in the solution without ever being reduced.

The Half-Reaction at the Cathode is:

\[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

**So, in summary. In Electorefining:**

- The **Impure** Metal is the **Anode**
- The **Pure** Metal is the **Cathode**
- The **Electrolyte** (solution) contains the cation of the **metal to be purified**

**Electroplating**

An application of an electrolytic cell is Electroplating. In **Electroplating**:

- The **object to be Plated** is the **Cathode** (attached to the Negative Terminal of the Battery)
- The **Electrolyte** must contain the **Cation of the Metal** to be Plated on the Object
- The **Best Anode** is the **Metal to be Plated** onto the Object
Example: We wish to plate an iron ring with copper.

- The **Iron** ring (made by Dwarfs?) is made the **Cathode** (connected to the **negative**)  
- The **Anode** is **Copper** (to keep supplying Cu$^{2+}$ ions as it is oxidized)  
- The **Electrolyte** is aqueous CuSO$_4$ (supplies Cu$^{2+}$ to start with. SO$_4^{2-}$ does not react here.)

Here is the diagram of the set-up:

**Electrowinning**

The name given to the **reduction of ores** to produce **metals** in industry.

Eg.)  
\[
\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}_{(s)}
\]

Comes from Zinc ores like ZnS, ZnCl$_2$ etc.

Zinc Metal
**Aluminum Production** (see p. 246 in SW)

Looking at the reduction table:

We see that Al\(^{3+}\) is BELOW the left overpotential arrow. Therefore, Al\(^{3+}\) **CANNOT** be reduced from an aqueous (water) solution.

In order to reduce Al\(^{3+}\) to its metal Al, you must electrolyze a **MOLTEN** salt of aluminum.

The main ore used to produce Aluminum is called **Bauxite** – Hydrated Aluminum Oxide or \(\text{Al}_2\text{O}_3\cdot3\text{H}_2\text{O}\)

When this is heated, the water is forced off: \(\text{Al}_2\text{O}_3\cdot3\text{H}_2\text{O} \rightarrow \text{Al}_2\text{O}_3 + 3\text{H}_2\text{O}\)

The melting point of **Alumina** (\(\text{Al}_2\text{O}_3\)) is **much too high** to melt it economically. (mp. = 2072 °C)

So Alumina (\(\text{Al}_2\text{O}_3\)) is mixed with **Cryolite** (\(\text{Na}_3\text{AlF}_6\))

The melting point of this mixture is \(~1000\) °C.

Some of the ions present in this mixture of molten salts are: \(\text{Al}^{3+}, \text{O}^{2-}, \text{F}^-\) and \(\text{Na}^+\)

The Half-Reaction at the Cathode is the **Reduction of Al\(^{3+}\)** ions to Al metal:

**Cathode** Half-Reaction: \(\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}\)

(At these high temperatures both \(\text{Al}^{3+}\) & Al are liquids)

The Half-Reaction at the Anode is the Oxidation of oxide ions using Carbon Electrodes:

**Anode** Half-Reaction: \(\text{C} + 2\text{O}^{2-} \rightarrow \text{CO}_2 + 4\text{e}^-\)

Some texts state more simply: \(2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-\)
This process is carried out by Alcan in Kitimat, B.C. It uses **10 million amps** of electricity. Alcan has their own power generating plant at Kemano B.C.

Because of high temperature and great amounts of electrical energy used, production of Al is expensive! Recycling Al used a lot less energy!

**Corrosion**

Corrosion is Undesirable Oxidation of a metal (usually Fe)

**Oxygen** from the air is reduced at the **Cathode**

The Fe becomes the **Anode** (is oxidized)

**Moisture** can provide a solution for these processes to take place in. O₂ from the air dissolves in the water droplet. The dissolved O₂ is reduced at the Cathode and the iron surface becomes the Anode. The series of reactions that take place are as follows:

1. \( \text{Fe} \rightarrow \text{Fe}^{2+} + 2e^- \) (iron is oxidized to Fe²⁺)
2. \( \frac{1}{2} \text{O}_2 + \text{H}_2\text{O} + 2e^- \rightarrow 2\text{OH}^- \) (O₂ picks up e’s from the Fe and is reduced to OH⁻)
3. \( \text{Fe}^{2+} + 2\text{OH}^- \rightarrow \text{Fe(OH)}_{2(s)} \) (the Fe²⁺ from Rx.1 and the OH⁻ from Rx.2 form the ppt. (Fe(OH)₂(s))
4. \( \text{Fe(OH)}_{2} \) Oxidized by O₂ and H₂O \( \rightarrow \text{Fe(OH)}_{3(s)} \)
5. Some \( \text{Fe(OH)}_{3(s)} \) Reacts with H₂O to produce \( \text{Fe}_2\text{O}_3\cdot3\text{H}_2\text{O}(s) \) (Hydrated Iron(III) Oxide)

A mixture of Fe(OH)₃(s) and Fe₂O₃·3H₂O(s) is called RUST

**Preventing Corrosion**

✔ Keep the Fe (or steel) surface protected from moisture and oxygen—Paint or other Coatings

✔ Keep the Fe away from metals which are higher on the Reduction Table. eg.) When Fe and Cu are in close proximity and the conditions are right—moisture is present and there are some dissolved salts—a type of Electrochemical Cell can form in which Cu becomes the Cathode and Fe becomes the Anode (and Fe is oxidized)

✔ Use what’s called **Cathodic Protection**
**In Cathodic Protection:**

A metal **lower** on the Reduction Table (higher oxidation potential) is attached to or brought near to the Fe.

<table>
<thead>
<tr>
<th>Reducing Agent</th>
<th>Oxidized Product</th>
<th>Oxidation Potential (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe²⁺ + 2e⁻</td>
<td>Fe⁰</td>
<td>-0.45</td>
</tr>
<tr>
<td>Ag⁺ + 2e⁻</td>
<td>Ag⁰</td>
<td>-0.80</td>
</tr>
<tr>
<td>Cr³⁺ + 3e⁻</td>
<td>Cr⁰</td>
<td>-0.74</td>
</tr>
<tr>
<td>Zn²⁺ + 2e⁻</td>
<td>Zn⁰</td>
<td>-0.76</td>
</tr>
<tr>
<td>Te⁶⁺ + 2H⁺</td>
<td>Te⁰</td>
<td>-0.79</td>
</tr>
<tr>
<td>Fe⁰ + 2H⁺</td>
<td>H₂⁰</td>
<td>-0.83</td>
</tr>
<tr>
<td>Mn²⁺ + 2e⁻</td>
<td>Mn⁰</td>
<td>-1.19</td>
</tr>
<tr>
<td>Al³⁺ + 3e⁻</td>
<td>Al⁰</td>
<td>-1.66</td>
</tr>
<tr>
<td>Mg²⁺ + 2e⁻</td>
<td>Mg⁰</td>
<td>-2.37</td>
</tr>
<tr>
<td>Na⁺ + e⁻</td>
<td>Na⁺</td>
<td>-2.71</td>
</tr>
<tr>
<td>Ca²⁺ + 2e⁻</td>
<td>Ca⁰</td>
<td>-2.87</td>
</tr>
<tr>
<td>Sr²⁺ + 2e⁻</td>
<td>Sr⁰</td>
<td>-2.99</td>
</tr>
<tr>
<td>Ba²⁺ + 2e⁻</td>
<td>Ba⁰</td>
<td>-3.18</td>
</tr>
<tr>
<td>K⁺ + e⁻</td>
<td>K⁺</td>
<td>-3.93</td>
</tr>
<tr>
<td>Rb⁺ + e⁻</td>
<td>Rb⁺</td>
<td>-2.98</td>
</tr>
<tr>
<td>Cs⁺ + e⁻</td>
<td>Cs⁺</td>
<td>-3.03</td>
</tr>
<tr>
<td>Li⁺ + e⁻</td>
<td>Li⁺</td>
<td>-3.04</td>
</tr>
</tbody>
</table>

Any of these metals would work to Cathodically protect Iron.

Metals below Mg would cathodically protects Fe, BUT they are highly reactive with water to form explosive H₂(g)!

An example would be Mg

In the presence of O₂ or any other oxidizing agent, Mg (with an oxidation potential = +2.37 v) Will give up e’s more readily than Fe (oxidation potential = +0.45 v). This saves the Fe from being oxidized!

**Some common examples of Cathodic Protection:**

- Plates of Mg are attached to the hulls of ships. When Mg is oxidized it is replaced.
- Galvanized Steel is Steel (mainly Fe) coated or mixed with Zn (oxidation potential = +0.76v)
  
  \[
  \text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^- \quad \text{a strong coating of ZnO is formed, protecting the metal from further oxidation.}
  \]

- Stainless Steel contains Fe with other metals like Cr etc.

**Impressed Current**

A voltage is applied with an external power supply to keep the Fe slightly negative. Thus, the Fe surface has excess e’s so it doesn’t have to oxidize in order to supply e’s to O₂.

**The End of Chem 12 Notes!** *(And you thought we’d never get there!)*