

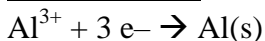
APPLICATIONS OF ELECTROLYSIS

1. *Extraction of reactive metals.*

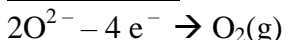
Electrolysis is used to extract reactive metals from their ores. Examples of reactive metals are potassium (K), sodium (Na), calcium (Ca), magnesium (Mg) and aluminium (Al).

E.g. Aluminium from alumina (aluminium oxide – Al_2O_3)

At the Cathode



At the Anode



2. *Electroplating*

Electroplating involves covering one metal with a protective coat of another metal. In practice, the outer layer is made of a less reactive metal. There are several reasons for electroplating metal objects, including:

- To enhance the appeal of the plated article – the electroplated article is more decorative than the metal underneath;
- To protect the covered metal from corrosion;
- To avoid using expensive metals for the object

In electroplating, the object to be covered is made at the cathode and the pure plating metal is made at the anode. The solution needs to contain ions of the plating metal. During plating, ions pass into solution from the pure anode and are discharged as a thin layer on the cathode.

Chromium plating –

Plating solution: usually chromic (VI) acid, to which chromium (III) sulphate and chromium (III) fluorosilicate are added

Anode: Pure chromium

Cathode: The metal/object to be plated. In decorative chrome plating, this is first given an undercoat of copper and nickel

Nickel Plating –

Plating solution: Nickel (II) sulphate or nickel (II) chloride solution, to which boric acid and a chemical wetting agent are added.

Anode: Nickel

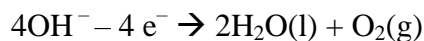
Cathode: The metal to be coated

3. *Anodizing*

When exposed to the air, aluminium combines with oxygen to form aluminium oxide. The oxide forms an even coat and seals the surface, thus protecting the metal from further corrosion. Electrolysis is used to make this protective layer thicker and tougher; the process is called anodizing.

Aluminium is made at the anode of an electrolytic cell that contains dilute sulphuric acid or dilute chromic (VI) acid as the electrolyte. (Any electrolyte that releases oxygen gas at the anode can be used.)

The reaction at the anode is:



The liberated oxygen combines with the aluminium anode, coating it with oxide.

Another advantage of anodizing is that the protective layer can be made to absorb dyes which then are permanently fixed by treatment with boiling water.

N.B. Anodizing of iron is known as rusting.

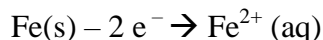
Rusting

Corrosion of metals can sometimes be attractive, such as the patina on brass objects or the green colour of a copper roof. Corrosion can be protective, as in the case of aluminium; however, the corrosion of iron (rusting) is very destructive and costly.

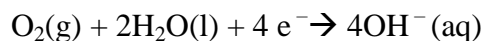
Rust is hydrated iron (III) oxide – $\text{Fe}_2\text{O}_3 \cdot x\text{H}_2\text{O}$. Rust does not stick to the iron underneath; instead it flakes off and a fresh surface is exposed for further attack. It is this continuous process that weakens the metal.

As much as 10 – 25% of the iron and steel manufactured each year is lost through rusting. To replace it costs huge sums of money. Rusting can also lead to loss of life; examples are when railway line rust through and trains become derailed or when boilers and machinery malfunction in factories. The rusting of iron is significant since iron and steel are widely used in construction (steel is an alloy of iron).

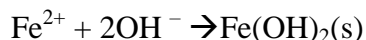
Rusting is an oxidation process. Iron loses electrons according to the equation:



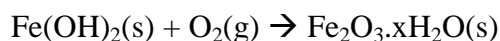
This reaction occurs in parts of the iron that are not exposed to air. In parts of the iron where there is a good supply of oxygen and water, the following reaction occurs:



The Fe^{2+} ions react with the OH^{-} ions to form iron (II) hydroxide.



Oxygen from the air oxidizes the iron (II) hydroxide to hydrated iron (III) oxide, which is rust:



How Can Rusting Be Controlled?

- **Coupling:**

Iron corrodes less readily when it is coupled with a more reactive metal. The iron and the more reactive metal, such as magnesium or zinc, must be in contact with each other, and with an electrolyte. Such combinations of two metals (a couple) and an electrolyte are known as electrolytic cells.

The more reactive metal in the couple loses electrons (is oxidized) in preference to the iron. This is known as sacrificial protection; the magnesium or zinc corrodes in preference to the iron, and the iron is protected.

Sometimes, the iron is completely covered with a coat of zinc. This process is called galvanising. The zinc is oxidized and protects the iron from corrosion. Even if the surface becomes scratched, the zinc is preferentially oxidized to the iron.

There are other methods which can be used to try to prevent or control rusting:

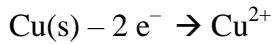
- **Tin Plating** – “Tins” are actually cans of steel coated with layers of tin. However, because tin is less reactive than iron, the iron corrodes when a tin is scratched. Tin-coated cans can be safely used for food storage since tin salts are non-toxic. The small amounts of tin that may dissolve out into foods do not harm the consumer.
- **Painting the object**
- **Coating the object with organic plastics, rubbers and greases**
- **Dipping the object into dilute solutions of oxidizing agents such as sodium nitrite, potassium dichromate (VI) and sulphuric acid.** These result in the formation of a tough oxide layer.

4. *Electrorefining of Metals*

A common metal which is refined is copper. Copper is a good conductor of electricity, and if it has impurities in it, the conduction is hampered. The principle is that the anode is made of the impure metal, e.g. Impure copper, and the cathode is made of the pure metal, e.g. Cu. The electrolyte will contain Cu^{2+} ions, e.g. from CuSO_4 solution.

At the anode:

The impure Cu anode dissolves, forming Cu^{2+} ions; the impurities fall within the electrolyte.



At the cathode:

The Cu^{2+} ions are discharged and Cu metal is deposited at the cathode.

