3.7 CORROSION OF METALS

Corrosion is an electrochemical process that occurs when metals are oxidised by substances in their environment.

Anode (metal oxidises): Metal → (Metal ions) $^{x+}$ + x e-

Cathode: Four common, but not exhaustive, reduction processes are given below:

(i)	$2H^+ + 2e^- \rightarrow H_2$	metal exposed to moist, acidic conditions
(ii)	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	metal exposed to moisture and oxygen (very common)
(iii)	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	metal exposed to moisture but a low concentration of oxygen, e.g. shovel blade in the soil
(iv)	$(\text{Metal ions})^{x+} + x\ e^- \rightarrow \text{Metal}$	metal in contact with another metal with a greater reduction potential

Preventing corrosion is often about limiting the contact of the metal with these four processes.

Question 3.32

List the following metals in order of increasing tendency to be oxidised.

Fe, Zn, Au, Mg, Na, Ag, Al, Pb, Cu

When some metals oxidise, they form a tough, protective oxide coating that greatly limits the ability of the oxidising agent to come in contact with the metal. Aluminium is a common example. Even though Al will oxidise rapidly, the oxide coating protects the underlying Al from further corrosion. Iron, however forms a oxide coating that is easily penetrated by O_2 and H_2O and subsequently the oxide layer does not protect the Fe from further corrosion.

Rust – A Special Example of Corrosion

The most common and important example of corrosion is the oxidation of iron. This is caused by the action of oxygen and water vapour in the air.

The formation of rust involves a series of reactions.

- *The initial oxidation of the iron.*

$$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^- \quad \text{(anodic reaction)}$$
$$O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)} \quad \text{(cathodic reaction)}$$

- *The further oxidation of the $Fe(OH)_{2(s)}$ formed.*

$$4Fe(OH)_{2(s)} + 2H_2O_{(l)} + O_{2(g)} \rightarrow 4Fe(OH)_{3(s)}$$

pale green — red brown

- *The partial dehydration of the $Fe(OH)_{3(s)}$ to rust. One possible reaction is shown.*

$$2Fe(OH)_3 \rightarrow Fe_2O_3.H_2O_{(s)} + 2H_2O_{(l)}$$

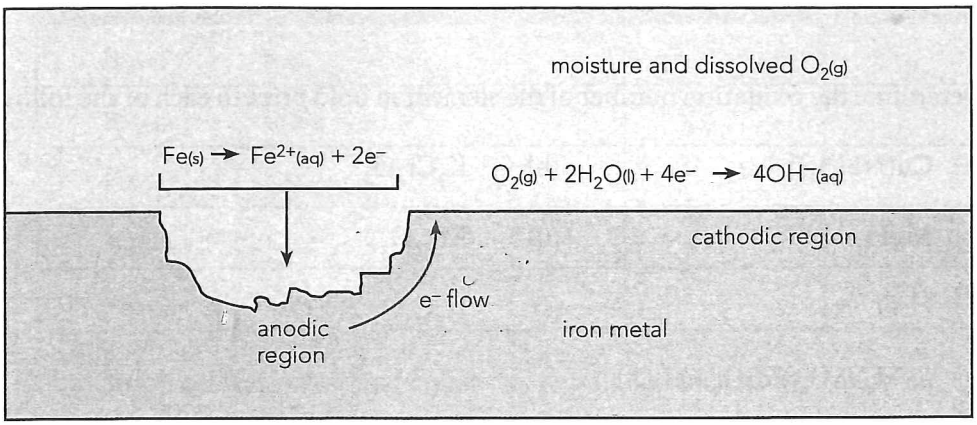

Figure 3.12 *Corrosion of iron by the action of oxygen and water.*

This is an electrochemical process. Areas of stress in the iron become anodic while areas of high oxygen concentration become cathodic. Electrons flow through the metal from the anode to the cathode. The moisture acts as the electrolyte.

Corrosion prevention

Prevention of corrosion is of tremendous economic importance (future millionaires – here is an area where you can make money). To prevent corrosion we must inhibit the reactions discussed on the previous page. This can be done by:

- excluding air and/or water from the metal surface by
  - protecting the iron surface with paint, grease, plastic
  - plating the iron surface with metals such as chromium or metallic tin
- using a sacrificial anode (more reactive metal) by
  - galvanising iron with zinc – zinc corrodes in preference
  - attaching magnesium and aluminium to ships
- using cathodic protection so that jetties and pipelines are rendered negative by a low voltage dc current. The anodes may be made of scrap iron (e.g. old engine blocks) or titanium coated inert electrodes.

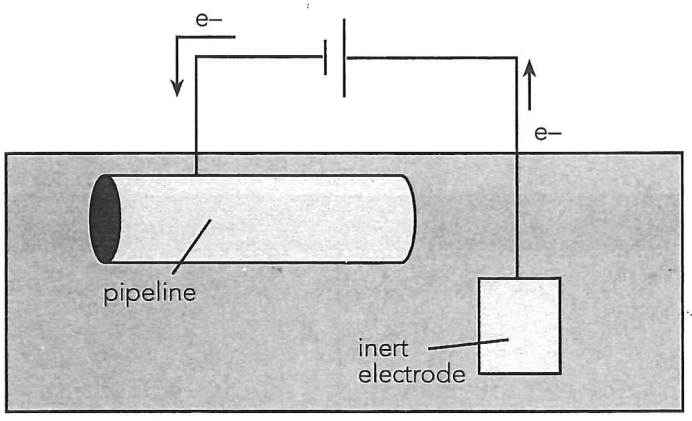

Figure 3.13 *Cathodic protection. The pipeline is made negative to prevent its oxidation. Instead, H_2O is reduced at the pipeline, and oxidised at the inert electrode.*