

CHEMISTRY

LEVEL 4C

(CHM 415115)

CORROSION

THEORY SUMMARY

&

REVISION QUESTIONS

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CHEMISTRY LEVEL 4C (CHM415115)

CORROSION

(CRITERION 5)

INTRODUCTION:

Corrosion is a general term referring to the **DESTRUCTIVE OXIDATION OF METALS** caused by oxidising agents in the surrounding chemical environment.

The most commonly encountered form of corrosion is *rusting* which is the term specifically applied to the destructive oxidation of iron and steel. (The term *rust* is only to be used with respect to iron and steel).

Practically all metals will undergo corrosion and even the noble metals like platinum and gold will corrode if they are in contact with very powerful oxidising agents.

In many cases, the corrosion of a metal results in the formation of the metal oxide or hydroxide. However in some cases, other corrosion products such as carbonates, sulfides and even nitrides may be formed. e.g.

METAL	POSSIBLE CORROSION PRODUCTS
IRON (Fe)	Fe_2O_3 $\text{Fe}(\text{OH})_3$ $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$
ZINC (Zn)	ZnO $\text{Zn}(\text{OH})_2$ ZnCO_3
ALUMINIUM (Al)	Al_2O_3 $\text{Al}(\text{OH})_3$
MAGNESIUM (Mg)	MgO $\text{Mg}(\text{OH})_2$ Mg_3N_2
SILVER (Ag)	Ag_2O Ag_2S

* In this study unit on corrosion, we shall be principally concerned with the corrosion of **IRON** with rust being represented as $\text{Fe}(\text{OH})_3$ although $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$ is a preferred chemical formula for rust.



Corrosion damage to cars costs Australians millions of dollars each year!

THE EFFECTS OF CORROSION:

The effects of corrosion are usually very undesirable and *often very costly*. This is because the process of corrosion changes the properties from those of a *metal* to those of *ionic compounds* which are very different from the metal's properties.

Some fairly typical properties of metals and their subsequent corrosion products are listed below.

METAL	CORRODED METAL
METALLIC PROPERTIES	IONIC PROPERTIES
strong	low strength
malleable & ductile	brittle & non-elastic
electrical conductor	electrical non-conductor
shiny	dull (non-shiny)
not water soluble	water soluble
metal's volume smaller	larger volume than the metal

- The brittleness and lower strength of metal oxides compared to the strength and malleability of the original metal has serious implications when one considers corrosion occurring to planes, bridges, buildings,.....
- Leaks in pipes and storage tanks may occur because the corrosion products dissolve away and form holes.
- Electrical faults will occur with the formation of corrosion products resulting in the electrical circuit no longer conducting electricity. Sometimes this happens with the corrosion of car battery terminals.
- Rust, for example occupies more space than the original iron metal from which it was formed. This larger volume causes blistering under paintwork and the jamming of rusty nuts and bolts.
- The oxide corrosion products are brittle and often flake off the surface of the metal progressively thinning and weakening the metal.
- The dulling of shiny metal surfaces caused by corrosion is usually an undesirable and unattractive outcome.



A severely corroded bridge suggests great danger!

ELECTROCHEMICAL NATURE OF CORROSION

Corrosion is a spontaneous ‘redox’ process in which a metal such as iron undergoes oxidation at anode sites and (usually) oxygen in the presence of water undergoes reduction at cathode sites.

As such, corrosion is an *electrochemical process* with electrons flowing from anode sites to cathode sites and ions flowing in directions as in ELECTROCHEMICAL CELLS.

i.e.

**OXIDATION OF THE METAL OCCURS AT THE ANODE
(-)**

**REDUCTION OF OXYGEN & WATER OCCURS AT THE
CATHODE (+)**

**ELECTRONS FLOW THROUGH THE METAL FROM
ANODE SITES → CATHODE SITES**

**CATIONS FLOW THROUGH THE ELECTROLYTE
TOWARDS THE CATHODE SITES**

**ANIONS FLOW THROUGH THE ELECTROLYTE
TOWARDS THE ANODE SITES**

NOTE: Corrosion does not occur *evenly* over the whole metallic surface. Destructive oxidation of the metal occurs in some regions (anodes) and the reduction of oxygen and water occurs at other sites (cathodes)

For example, on a corroding steel nail, there may be 3 or 4 anode as well as cathode sites.

It's the anode sites where the damage to the metal is being done because it's here where the metal is being oxidised and converted into ionic compounds. At cathode sites, there is no destruction to the metal occurring.

Q1. Write *oxidation* half-equations for the reactions occurring when the following metals corrode:

Zn →	Al →
Cu →	Fe →
Mg →	Cr →
Pb →	Ca →

Q2. How are the electrons transferred from the anode sites to the cathode sites?

Q3. What cations are formed during the corrosion of silver?

Q4. Why does the presence of salt water as opposed to fresh water, accelerate the rate of metallic corrosion?

Q5. Qualitatively compare the concentration of dissolved oxygen ($O_{2(aq)}$) at the ocean's surface with that at a depth of 3800 m where the wreck of “RMS Titanic” lies.

FACTORS AFFECTING CORROSION RATES

The rate at which a piece of iron (or other metal) corrodes will be affected by the following variables:

- **AVAILABILITY OF OXYGEN.** For example, deep on the ocean floor there are steel ship wrecks that have not corroded significantly due to the very low level of dissolved oxygen at such depths.
- **AVAILABILITY OF WATER.** In dry desert environments where, even though O₂ gas from the atmosphere is plentiful, the lack of water means corrosion of structures such as steel railway lines is usually minimal.
- **TEMPERATURE.** The rate of corrosion is usually affected quite significantly by temperature variation. Higher tropical temperatures normally mean faster corrosion rates than those occurring in cold climates such as in Antarctica.
- **AVAILABILITY OF SALTS.** The electrochemical nature of corrosion means that there must be a salt-bridge between the anode and cathode sites. Pure water is a very poor electrolyte whereas salt water is an excellent electrolyte. Thus corrosion of a steel ship is much more likely in *salty* seawater than it would be in fresh water.
- **STRESS & METAL FATIGUE.** Where a metal has been subjected to deformation stresses and/or metal fatigue, the disrupted metallic crystal lattice is more likely to undergo corrosion. This is seen with the corrosion of nails where the head and point ends are usually more prone to rusting than other parts.
- **SURFACE AREA EXPOSED.** An increased area of contact between metal and the oxidising agent, gives faster corrosion rate; e.g. steel wool rusts much faster than a solid nail.
- **CONTACT BETWEEN DIFFERENT METALS** * (main heading - see page 11) *



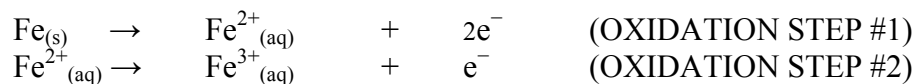
A corrugated iron roof rusting in a marine environment

THE RUSTING OF IRON

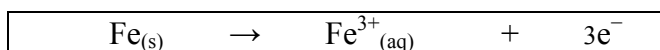
Our experiments reveal that for iron (or steel which is an iron/carbon alloy) to corrode there must be both oxygen and water present.

The oxidation of iron is a two step process with the initial formation of iron(II) (Fe^{2+} ions) and then this undergoing further oxidation to form iron(III) (Fe^{3+} ions).

i.e.

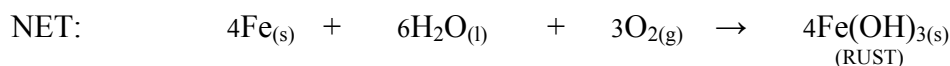
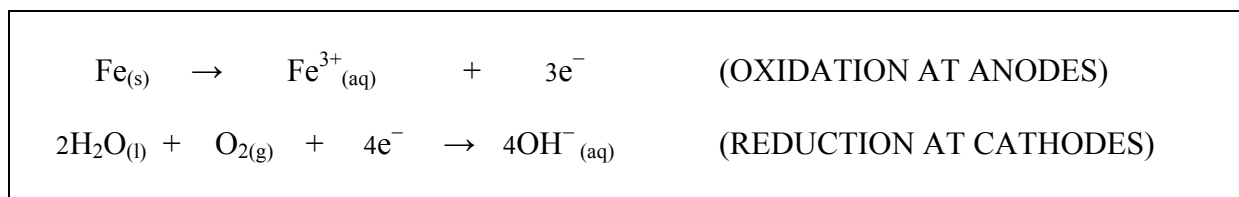


Combining gives:



For simplicity, we shall show the oxidation of iron to be a single step as shown above although we should remember that both $\text{Fe}^{2+}_{(aq)}$ and $\text{Fe}^{3+}_{(aq)}$ will be present at anode sites. Eventually, the predominant corrosion species at anode sites will be $\text{Fe}^{3+}_{(aq)}$.

THESE HALF-EQUATIONS SHOULD BE KNOWN:



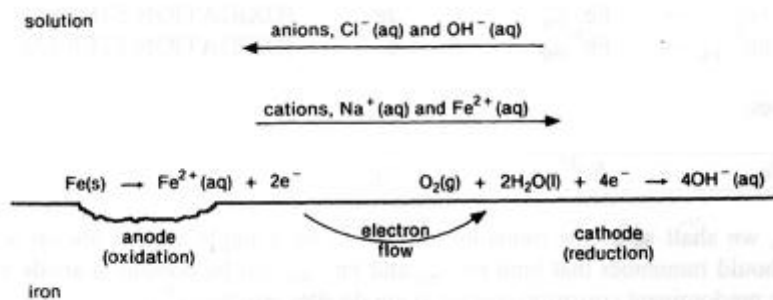
IDENTIFYING ANODE & CATHODE SITES

In experimental work when investigating the corrosion of iron, “ferroxyl” indicator is often used to locate anode and cathode sites. This is achieved by the formation of different colours. The “ferroxyl” indicator is a mixture of potassium hexacyanoiron(III) which has the chemical formula $\text{K}_3\text{Fe}(\text{CN})_6$ and the acid/base indicator phenolphthalein.

- Hexacyanoiron(III) ions in the presence of $\text{Fe}^{2+}_{(aq)}$ ions turn a **deep blue** colour (called Prussian Blue). Thus, blue regions involve sites where $\text{Fe}^{2+}_{(aq)}$ has formed and are therefore anode sites. Similarly, rusty brown regions involve sites where $\text{Fe}^{3+}_{(aq)}$ has formed and are also anode sites.
- Phenolphthalein indicator turns pink in regions with a $\text{pH} > 9$, i.e. regions where the $[\text{OH}^{-}_{(aq)}]$ is high. Reviewing the half-equations above, we see that hydroxyl ions are formed at the cathode. This means that **pink** areas are very likely to be cathode sites!

DIAGRAMMATIC SUMMARY

The diagram below shows just one anode site and one cathode site on the surface of a piece of iron which is in a salty and oxygen rich environment such as may occur in a marine location.



NOTES:

- Although the diagram shows only $\text{Fe}^{2+}(\text{aq})$ ions being formed (at the anode), there will also be $\text{Fe}^{3+}(\text{aq})$ formed as well.
- The anode site is where the actual destruction of the metal occurs but there is no metal destruction at the cathode site.
- The electrons flow **through the metal** from anode → cathode.
- The saltwater environment provides the necessary ions (Na^+ and Cl^-) for the ‘salt-bridge’ to complete the electrochemical cell’s circuit.
- The $\text{Fe}^{2+}(\text{aq})$ and $\text{Na}^+(\text{aq})$ cations flow away from the anode towards the cathode site.
- The $\text{OH}^-(\text{aq})$ and $\text{Cl}^-(\text{aq})$ anions flow away from the cathode towards the anode site.
- Precipitates of insoluble $\text{Fe}(\text{OH})_{2(\text{s})}$ and $\text{Fe}(\text{OH})_{3(\text{s})}$ are likely to form when these moving ions meet up between the anode and cathode sites.
- Careful inspection of the diagram above should convince you that it could be possible to have two anode sites adjacent to two cathode sites or vice-versa.



Steel bolts destroyed by rust

DIFFERENTIAL AERATION

As already mentioned, there can be a number of anode and cathode sites existing on the same metallic surface.

How can we possibly predict where an anode or a cathode site may form?

We've already said that an anode site is more likely to occur where the metal's crystalline lattice has been deformed or damaged by stress and/or metal fatigue. This is likely to occur during hammering, bending, stretching,.....

Oxygen undergoes reduction in the presence of water at cathode sites. Consequently, a cathode site is more likely to exist where the availability of oxygen is greatest. This gives rise to DIFFERENTIAL AERATION.

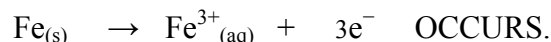
Thus, in regions surrounding the metal surface where the concentration of oxygen differs, the cathode site is likely to be where oxygen's availability is highest and the anode site is likely to be where oxygen's availability is lowest. There is a potential difference between regions of differing oxygen concentration, $[O_2]$ and oxygen undergoes reduction where it is at a higher concentration.

i.e.

HIGHER $[O_2]$ RESULTS IN CATHODE SITES WHERE:

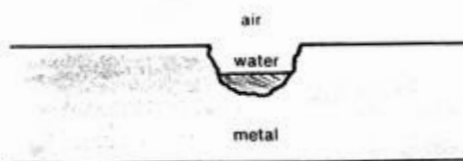


LOWER $[O_2]$ RESULTS IN ANODE SITES WHERE:



Thus, unless other factors are significant, the rusting of iron will occur preferentially where the concentration of oxygen is **lower** so long as there is a region of higher oxygen concentration close by.

Q6. The diagram below shows a piece of iron or steel which has a small depression in its surface that has partly filled with water. "Pitting" of the metal has occurred!



- (i) Around the region of the depression, where is the $[O_2]$ greatest? Explain.
- (ii) Around the region of the depression, where is the $[O_2]$ least? Explain.
- (iii) Where will the reduction of oxygen occur? Is this an anode or cathode site?
- (iv) Where will the oxidation of iron occur? Is this an anode or cathode site?
- (v) What happens to the depth of the depression as corrosion proceeds? Explain.
- (vi) Label the diagram above with all relevant features and show the direction of electron flow as well as the movement of ions.

Q7. The diagram (below right) shows a 'coach screw' made of steel that was used to secure a wooden section of a jetty in a marine (seawater) environment. When the bolt was removed for inspection the internal section revealed severe corrosion although the head of the bolt appears in good order.

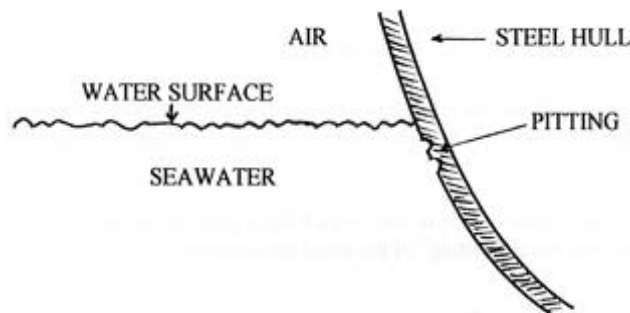
- (i) Explain the reason for this severe deterioration of the bolt parts that were not so directly exposed to the air.
- (ii) Why was the head of the bolt not so significantly affected by corrosion?
- (iii) Why does an occurrence like this present potential hazards for structures subject to possible corrosion?



Q8. A steel fishing boat is operating in a marine environment there appears to be no obvious sign of corrosion *above* the waterline.

However, when the boat was removed from the water for inspection, there was a significant level of deep pitting (corrosion hollows) just below the waterline.

- (i) Explain why most corrosion had occurred just *below* the waterline rather than above the waterline.
- (ii) What is this effect known as?
- (iii) On the diagram below, label all relevant sections indicating the electrochemical nature of the corrosion process occurring in this case



SHIP WRECK, CHUUK LAGOON, MICRONESIA

DIFFERENT METALS IN CONTACT

When two different metals are in electrical contact in a moist, oxygen rich environment, corrosion of one of the metals is highly likely to occur and the rate of corrosion is likely to be significantly accelerated. Situations where this may occur are in:

- Plumbing connections between steel and copper piping.
- Aluminium boats in contact with steel attachments such as anchors.
- Motor car engines with a cast iron engine block and an aluminium head.
- “The Statue of Liberty” where external copper sheeting is attached to an internal steel framework.
- Battery terminals made of lead are connected to steel car body components.
- Joining two identical pieces of metal by welding or soldering.

Although this corrosion outcome is normally undesirable and costly, there are definite occasions when we deliberately bring two different metals into contact and ‘sacrifice’ one metal in order to protect the other. (See later under “Corrosion Protection”)

Q9. The photograph (right) illustrates a major corrosion problem in a plumbing situation where a steel water pipe has been joined to a copper water pipe.

The steel pipe has corroded dramatically but it appears that the copper pipe is unaffected.



(i) How might plumbers connect these two different metals and yet not have an electrical contact between them? Explain.

(ii) Why is the rusty steel pipe now a larger overall diameter than the original steel pipe?

IMPORTANT SUMMARY STATEMENTS:

WHEN TWO DIFFERENT METALS ARE IN CONTACT IN AN ELECTROLYTE SOLUTION, THE MORE REACTIVE METAL (i.e. THE MORE POWERFUL REDUCER) WILL BECOME THE ANODE AND CORRODE.

THE LESS REACTIVE METAL WILL BECOME THE CATHODE AND NOT CORRODE BUT BE THE SURFACE ON WHICH OXYGEN IS REDUCED.

THIS COMBINATION OF TWO DIFFERENT METALS FORMS AN ELECTROCHEMICAL CELL AND THE POTENTIAL DIFFERENCE (E.M.F.) PRODUCED MAKES THE MORE REACTIVE METAL CORRODE EVEN FASTER THAN IT WOULD IF IT WERE ON ITS OWN!

THE LESS REACTIVE METAL IS PROTECTED.

Q10. In the following combinations, two different metals have been brought into electrical contact in an electrolyte. Use your E.C.S. to determine which metal will become the anode and corrode, and which metal will become the cathode and be protected. In each case write the half-equations occurring on that metal's surface. The first line has been completed for you.

METALS IN CONTACT	ANODE (CORRODING METAL)	CATHODE (PROTECTED METAL)
zinc & copper	zinc $\text{Zn}_{(s)} \rightarrow \text{Zn}^{2+}_{(aq)} + 2e^{-}$	copper $\text{O}_{2(g)} + 2\text{H}_2\text{O}_{(l)} + 4e^{-} \rightarrow 4\text{OH}^{-}_{(aq)}$
zinc & magnesium		
aluminium & iron		
silver & copper		
nickel & cadmium		
tin & lead		
chromium & manganese		
zinc & iron (see diagram below)		

Q11. A marine biologist drops and breaks a mercury thermometer in the bottom of an aluminium dinghy. The nature of the spillage made it extremely difficult to retrieve all the droplets of Hg metal from the bottom of the boat.

Quite soon after this accident, the boat developed small holes in its hull.

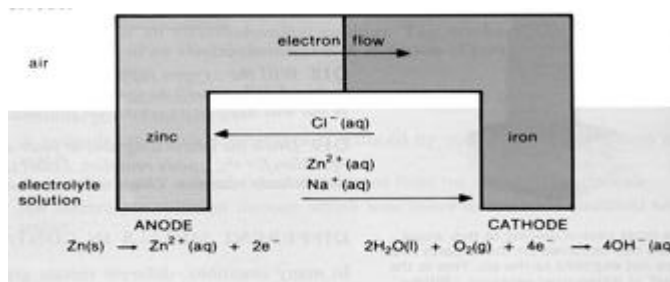
(i) Why did the boat corrode? Give two reasons.

(ii) Why do airline operators forbid passengers taking any mercury containing materials on flights?

The diagram below represents the outcome of having the two metals, zinc and iron in electrical contact in a saltwater environment.

The zinc becomes the anode and undergoes oxidation and corrodes.

The electrons flow towards the iron which becomes a cathode and is protected. Iron is the surface on which oxygen in the presence of water is reduced.



CORROSION PROTECTION

There are several different methods for protecting a metallic structure from corrosion. These methods broadly divide into five or so main groups:

- applying coatings like paint and polymers (plastics)
- alloying (using mixtures of metals rather than pure metals)
- noble metallic coatings (e.g. tin plating steel cans)
- sacrificial coatings and blocks (e.g. galvanising) = “CATHODIC PROTECTION”
- impressed current (applied EMF) = “CATHODIC PROTECTION”

1. PAINTING

A coat of paint or polymer film provides a non-electrically conducting *barrier* between the metal and the surrounding environment where oxidising agents may be present.

So long as the paint coverage is 100% complete, protection will be provided. Scratches or any other damage to the paint's surface will result in corrosion occurring in those damaged areas.

As soon as corrosion has commenced under a paint surface, the larger volume of corrosion product formed (metal hydroxide) results in swelling and blistering which breaks the paint coverage and further corrosion occurs. This blistering sign of corrosion is often seen with rusting at the bottom of car doors; (see page 3).

The protective quality of some paints is enhanced by adding special chemicals to the paint. These additives such as phosphate and chromate compounds are called **INHIBITORS** and slow down and tend to prevent the corrosion reactions from occurring.

Other paints contain anti-corrosive agents such as zinc dust.



The 52 800 tonnes of steel in the Sydney Harbour Bridge are protected against corrosion by painting. Repainting is the full-time task of an engineer, three foremen and thirty maintenance workers. The cost of corrosion prevention is over \$1 700 000 a year.

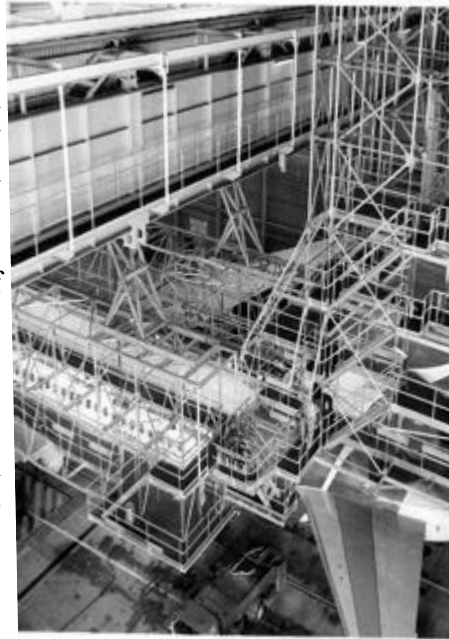
2. ALLOYING

Mixtures of metals (alloys) are often much more resistant to corrosion than the pure metals themselves.

Example 1. "18-8 stainless steel" is 18% chromium (Cr) and 8% nickel (Ni). This has a very low tendency to rust even when exposed to saltwater for prolonged periods.

The protection from corrosion is achieved by the formation of a thin but highly stable film of chromium(III) oxide (Cr_2O_3) which acts as a protective layer for the metal beneath.

Example 2. Pure aluminium metal corrodes quite readily in saltwater. However, by alloying aluminium with silicon (Si) and magnesium (Mg), its resistance to corrosion in the marine environment is enhanced. These alloys are used for boats, scuba cylinders, buildings (see right),.....



3. NOBLE METAL COATINGS

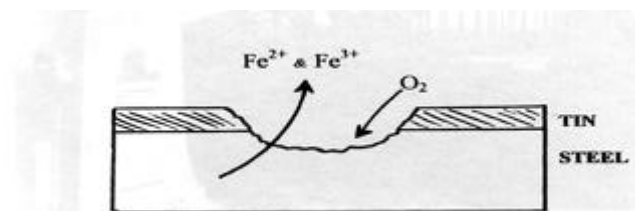
Noble metals such as tin, silver or gold are generally low in reactivity as these metals are weak reducing agents. Being very unreactive, they tend not to corrode in normal environments.

Steel food cans are usually protected from corrosion by a total but very thin coating of tin metal and hence we talk about 'tin cans' although the main metal present is steel.

Noble metal coatings provide corrosion protection in much the same way as a coat of paint does except that paints in contact with food may cause toxic contaminants to affect the food quality.

If the noble metal coating gets scratched or damaged, then the more reactive metal underneath will corrode rapidly due to the electrochemical cell created; (see page 11).

Noble metal coatings only provide protection so long as the coverage of the metal is 100% complete.



A tin coating protects the steel in the same way that paint does, by providing a barrier to oxygen and water contacting the steel. If the barrier is broken (as shown), the more reactive metal (Fe) will corrode

4. SACRIFICIAL COATINGS & SACRIFICIAL ANODES

As already discussed (on page 11) when two different metals are in electrical contact in an electrolyte solution, the more reactive metal (the more powerful reducing agent) will become the anode of an electrochemical cell and corrode whereas the less reactive metal becomes the cathode and does not corrode.

This becomes the basis of a corrosion protection method where a more reactive metal is sacrificed in order to prevent a less reactive metal from being damaged by corrosion.

For protecting iron or steel, the sacrificial metal commonly used is **ZINC** and there are two main methods of having the two metals in contact.

METHOD 1. Applying Zn metal as a thin coating over the whole surface of iron. This is called **GALVANISING**. The zinc metal coating is very slowly oxidised and the surface layer of ZnO (and ZnCO₃) formed protects the iron underneath. As long as Zn metal is present, the Fe is protected even if the galvanising is scratched and damaged.

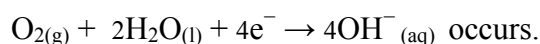
Apart from using zinc as a sacrificial coating, aluminium and magnesium can also be used to protect iron/steel.

METHOD 2. Attaching Zn metal blocks (or Al or Mg blocks) to the iron/steel. This method is used where it would be impractical to coat the whole surface as discussed in method 1.

The zinc blocks become anodes and are oxidised whilst the steel is a cathode and is protected. The blocks will need replacing eventually!

This method is used to protect large ships and railway lines plus many other large steel structures. These blocks are called **SACRIFICIAL ANODES**.

In all such cases involving sacrificial coatings or anode blocks, the iron being the less reactive metal is made to be a cathode where the reduction half-reaction:



This technique of protecting a metal by making it the cathode is called **CATHODIC PROTECTION**.

The picture (right) shows a steel structure protected by aluminium alloy sacrificial anodes. These anodes will need replacing quite regularly but this proves to be a lot less expensive than replacing the structure.



5. IMPRESSED CURRENT (APPLIED E.M.F.)

Destructive oxidation of a metal occurs at an **ANODE** site!

Any technique that makes a metal a **CATHODE SITE** will ensure that reduction occurs on that cathode surface and thus ensures the metal is protected from corrosion.

Apart from attaching sacrificial anode metals, another way of protecting a metal is to make it a cathode by supplying electrons to it by application of an EMF from an external electrical source. This technique is used for protecting bridges, piers, pipelines, underground fuel tanks,.....etc....

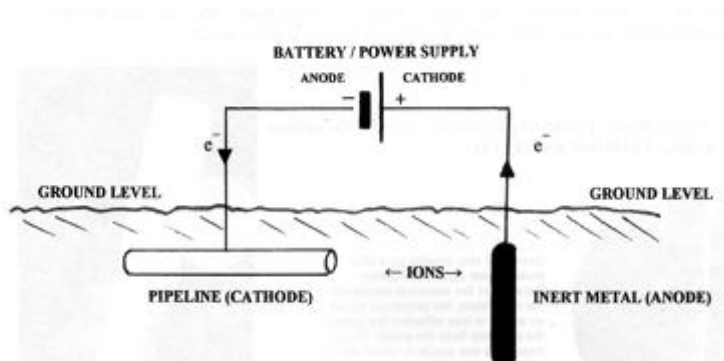
Using an impressed current to protect metal structures is another form of **CATHODIC PROTECTION**.

The **negative terminal** of a battery or D.C. power supply is connected to the metal to be protected. (Remember, electrons flow from anode → cathode). This makes the metal to be protected a cathode!

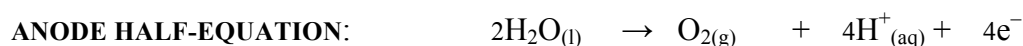
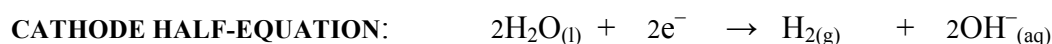
The anode is usually an inert electrode which is corrosion resistant and thus, rather than the anode being oxidised, water is oxidised.

Consider an underground steel pipeline passing through moist soil where sufficient oxygen is present to result in severe corrosion to the pipe.

An impressed current (or applied EMF) technique would prove an ideal method of protecting the pipe; i.e.



The two half equations are:



CORROSION OF ALUMINIUM

Although sometimes aluminium metal is used as a sacrificial metal for the protection of iron and steel, there are many instances where aluminium structures themselves need to be protected.

Q12. Aluminium is used extensively as a construction metal. What are two important properties of Al that make it particularly desirable for construction purposes?

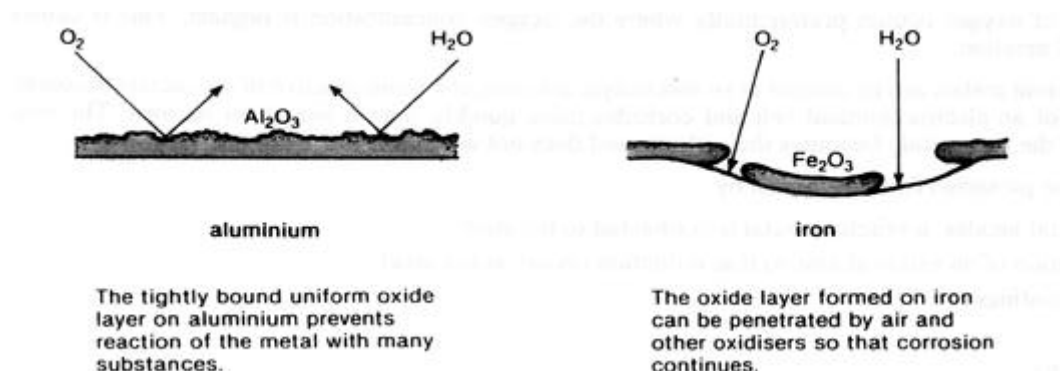
Aluminium has an apparently high resistance to atmospheric corrosion, giving the impression that Al is a comparatively unreactive metal.

This is explained by aluminium forming a strongly bonded protective oxide layer. This layer is so thin that the lustre of the metallic aluminium is normally visible beneath.

Q13. Write the chemical equation for the reaction between aluminium and oxygen.

The oxide layer forms on the surface of aluminium metal is a **continuous impermeable protective layer** unlike that which forms on iron. This protects aluminium from undergoing oxidation with many corrosive oxidising agents. We call it a **PASSIVE LAYER**.

Conversely, the rusty (Fe_2O_3) oxide layer on steel is porous and not strongly bonded to the surface. (It brushes off quite readily!)



However, the protective/passive oxide layer on aluminium can be destroyed by:

- chloride ions (hence **pure** aluminium corrodes readily in seawater)
- mercury and mercury(II) compounds

The corrosion resistance of aluminium can be enhanced by several methods. The two most commonly encountered methods are:

METHOD 1. Alloying with silicon and magnesium

METHOD 2. Thickening the oxide layer electrolytically in a process called **ANODISING**.

Q14. If aluminium metal is made the anode of an electrolytic cell, what will be the anode half-reaction?

THEORY SUMMARY

Corrosion is the oxidation of metals by substances in the surroundings.

Rusting is the corrosion of iron by oxygen in a moist environment. Rust can be represented by the formula $\text{Fe}_2\text{O}_3 \cdot \text{H}_2\text{O}$.

Corrosion is an electrochemical process:

- ▶ Oxidation occurs at an anode.
- ▶ Electrons are transferred through a metallic path to the cathode.
- ▶ At the cathode, electrons are accepted by a substance which is reduced.
- ▶ Electrical neutrality is maintained by movement of ions in an electrolyte solution.

Rusting of iron requires the presence of

- ▶ oxygen, as the oxidiser
- ▶ water and dissolved salts as the electrolyte.

Oxidation of iron occurs preferentially where the metal lattice has been deformed.

Reduction of oxygen occurs preferentially where the oxygen concentration is highest. This is called differential aeration.

When different metals are in contact in an electrolyte solution, the more reactive of the metals becomes the anode of an electrochemical cell and corrodes more quickly than it would on its own. The less reactive of the two metals becomes the cathode and does not corrode.

Steel can be protected from corrosion by

- ▶ sacrificial anodes; a reactive metal is connected to the steel
- ▶ application of an external emf so that reduction occurs at the steel
- ▶ metal coatings
- ▶ paints
- ▶ inhibitors
- ▶ alloying.

Protective coatings on steel may be either of the following types:

- ▶ Noble coatings, such as tin, cadmium and chromium, which are less reactive than iron. At a break in a noble coating, the iron corrodes rapidly.
- ▶ Sacrificial coatings, such as zinc and zinc-aluminium alloys, which are more reactive than iron. At a break in a sacrificial coating, protection of the steel is maintained.

Galvanised steel has a sacrificial coating of zinc.

Metal coatings are applied either by dipping or electroplating.

Stainless steels are corrosion-resistant because of a protective layer of chromium(III) oxide.

Aluminium is corrosion-resistant because of a tightly bound protective oxide layer. Corrosion of iron continues because its oxide layer is porous.

The protective oxide layer on aluminium and on most stainless steels is destroyed by chloride ions.

The oxide layer on aluminium can be thickened by an electrolytic process called anodising.

Aluminium and zinc are corroded by hydroxide ions.

NAME _____ LINE _____

CORROSION OF METALS TEST

TOTAL = 36 marks

Q1. Suggest four(4) factors that will **increase** the rate of corrosion of a piece of iron.

- (i) _____
- (ii) _____
- (iii) _____
- (iv) _____

(4 marks)

Q2. Suggest four(4) different ways of protecting an iron pipe from corrosion.

- (i) _____
- (ii) _____
- (iii) _____
- (iv) _____

(4 marks)

Q3(a). When steel (Fe) undergoes corrosion in a moist environment, give the:

- (i) oxidation half-equation. _____
- (ii) reduction half-equation. _____

(2 marks)

Q3(b). A steel hulled fishing boat was found to have suffered severe corrosion in the form of "pitting" in regions just **below** the water line. By referring to your answer to part (a) above, explain why the corrosion was greatest in these areas and not in areas in direct contact with the air.

(6 marks)

Q4. Six pieces of iron are each coated with a different metal to demonstrate the effects of metal coatings as means of preventing corrosion. For each coating listed, indicate whether the protection is due to a **noble** or **sacrificial** coating.

- | | |
|---------------------|---------------|
| (a) lead | (b) zinc..... |
| (c) magnesium | (d) gold..... |
| (e) cadmium | (f) tin..... |

Q5. Three identical iron nails are prepared as follows and then placed in a moist aerated environment:

Nail #1 = iron with **no other metal contact**

Nail #2 = iron wrapped with **aluminium wire**

Nail #3 = iron wrapped with **copper wire**

(a) Which nail will show the **greatest** extent of corrosion? Explain why.

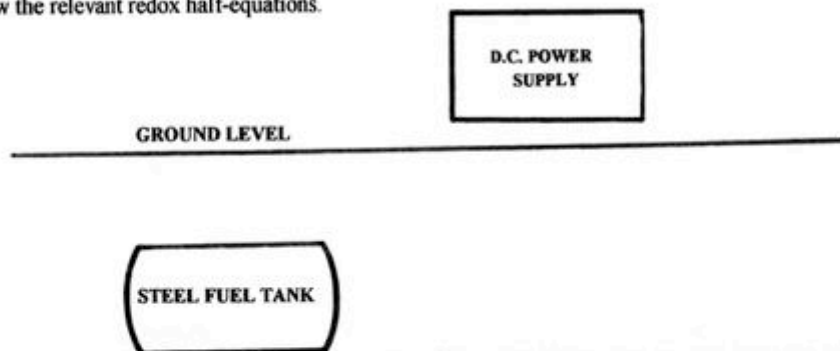
(5 marks)

(b) Which nail will show the **least** extent of corrosion? Explain why.

(5 marks)

Q6. Draw and label a diagram that shows how a D.C. electrical power supply could be used to prevent or at least decrease the extent of corrosion of a steel underground tank used for storing diesel fuel.

Show the relevant redox half-equations.



(5 marks)

(ii) Redox half-equations occurring

OXIDATION 1/2 EQUATION:

REDUCTION 1/2 EQUATION:

(2 marks)

NAME: (ANSWERS) LINE:
CORROSION OF METALS
TEST QUESTIONS

TOTAL = 36 marks

Q1. Suggest four(4) factors that will increase the rate of corrosion of a piece of iron.
 (i) connect the Fe to a less reactive metal eg Cu
 (ii) connect the Fe to this more reactive metal of a battery
 (iii) place it in salty water
 (iv) increase the surface area ... etc

Q2. Suggest four(4) different ways of protecting an iron pipe from corrosion.
 (i) sacrificial anode (eg Zn block)
 (ii) noble coating (eg Sn)
 (iii) applied IMF (eg connected to pipe)
 (iv) painting etc

Q3(a). When steel (Fe) undergoes corrosion in a moist environment, give the:
 (i) oxidation half-equation. $Fe \rightarrow Fe^{2+} + 2e^{-}$
 (ii) reduction half-equation. $2H_2O + O_2 + 4e^{-} \rightarrow 4OH^{-}$

Q3(b). A steel hulled fishing boat was found to have suffered severe corrosion in the form of "pitting" in regions just below the water line. By referring to your answer to part (a) above, explain why the corrosion was greatest in these areas and not in areas in direct contact with the air.

The corrosion is greatest where the concentration of O_2 is lower
 & differential aeration has occurred
 A PD exists between the two areas of different O_2 conc
 and O_2 is reduced where the $[O_2]$ is greatest & above water line
 $\therefore Fe$ is oxidised where $[O_2]$ is lowest & below water line

Q4. Six pieces of iron are each coated with a different metal to demonstrate the effects of metal coatings as means of preventing corrosion. For each coating listed, indicate whether the protection is due to a noble or sacrificial coating.
 (a) lead N
 (b) zinc S
 (c) magnesium S
 (d) gold N
 (e) cadmium N

(3 marks)

Q5. Three identical iron nails are prepared as follows and then placed in a moist aerated environment:
 Nail #1 = iron with no other metal contact
 Nail #2 = iron wrapped with aluminium wire
 Nail #3 = iron wrapped with copper wire

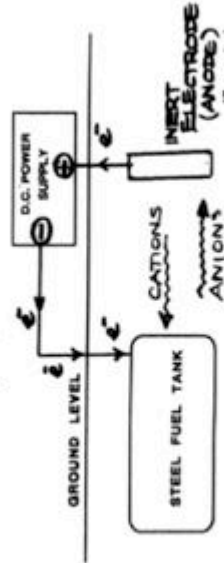
(a) Which nail will show the greatest extent of corrosion? Explain why.

Nail #3
 Cu is less reactive than Fe
 \therefore the more reactive metal Fe corrodes faster than it would on its own (#1)
 This is because the two metals in contact set up an electrochemical cell in which the EMF produced accelerates the corrosion process. $\therefore Fe = \text{ANODE} \text{ \& } \text{OXIDISES}$
 $Cu = \text{CATHODE}$ (5 marks)

(b) Which nail will show the least extent of corrosion? Explain why.

Nail #2
 Al is more reactive than Fe
 The aluminium metal being more reactive corrodes preferentially and thus makes the Fe nail the "protected" cathode site.
 \therefore aluminium is a sacrificial anode and the Fe is likely to not corrode at all

Q6. Draw and label a diagram that shows how a D.C. electrical power supply could be used to prevent or at least decrease the extent of corrosion of a steel underground tank used for storing diesel fuel. Show the relevant redox half-equations.



(i) Redox half-equations occurring:
 OXIDATION 1/2 EQUATION: $2H_2O \rightarrow O_2 + 4H^{+} + 4e^{-}$
 REDUCTION 1/2 EQUATION: $2H_2O + 2e^{-} \rightarrow H_2 + 2OH^{-}$ (5 marks)

(2 marks)