Unit 25071

©

Demonstrate knowledge of electromotive force (e.m.f.) production

This workbook is intended for "SELF-PACED STUDY".

The assessment for this Unit is not included in this Workbook.

You will need:

Calculator – Casio fx-82

This unit was written by A. Kawana

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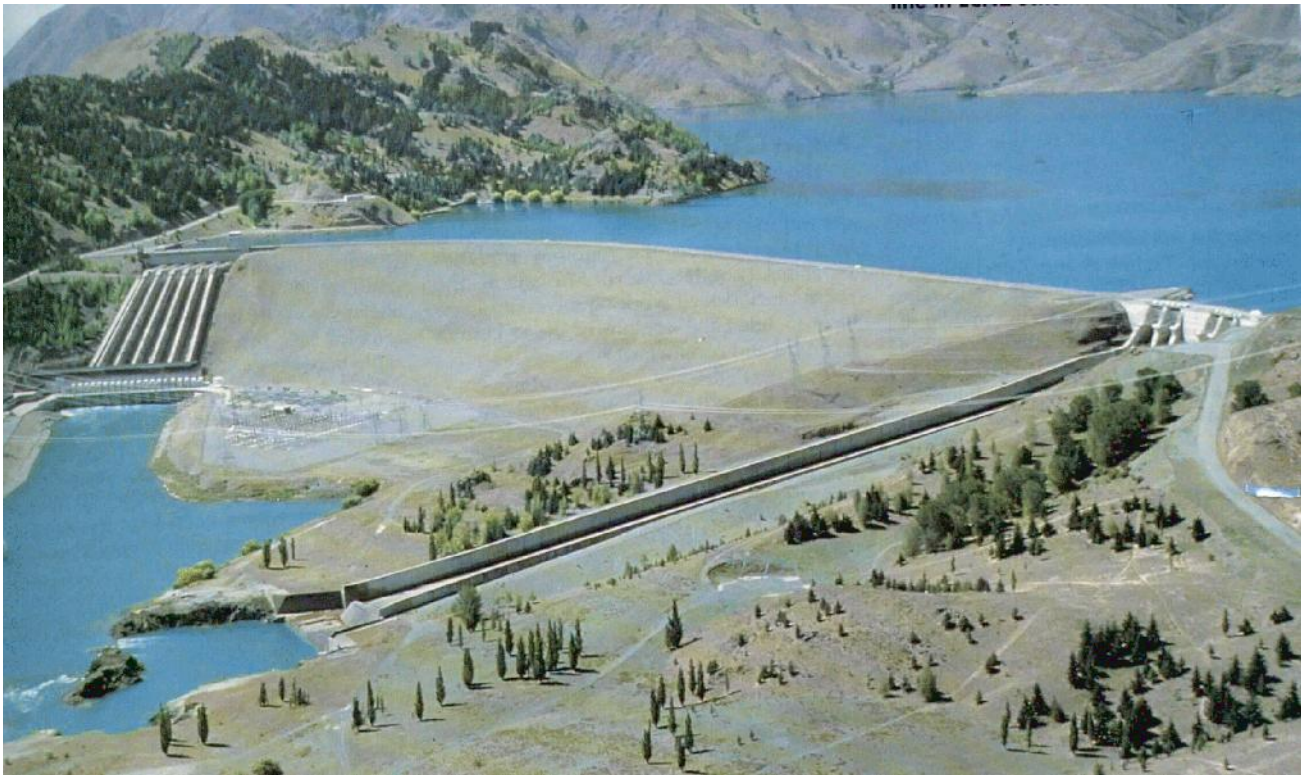
1 – Electromotive Force (EMF)

PRODUCTION OF AN EMF



Benmore Hydro Scheme - Lake Benmore on the Waitaki River - South Island NZ.

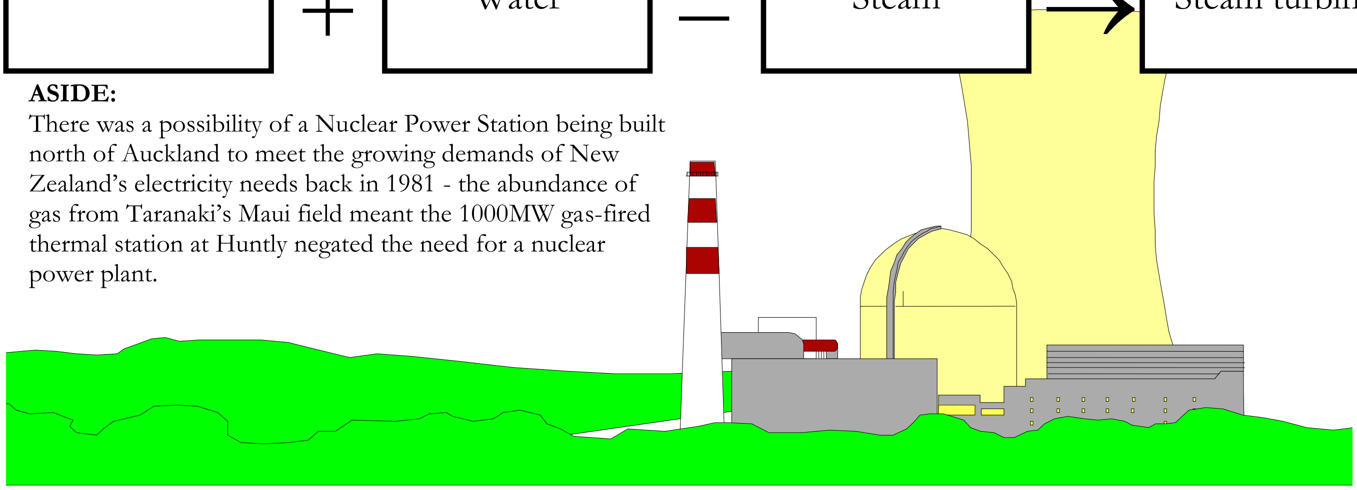
March 2002: Plans for a new billion-dollar South Island hydro-electric power scheme are moving ahead quickly, with resource consent applications to be filed within 3 months. Named Project Aqua the first electricity could be coming from the canal-based scheme within six years. Project Aqua will be based around the Waitaki River, where a series of dams already generate power on the upper reaches. The Meridian scheme envisages a diversion of the river, with two thirds of the water running down a 62km canal, going through a sequence of six power stations, then rejoining the river before it empties into the sea between Waimate and Oamaru. In combination, the six power houses would generate 570 MW or 3200 GWh of electricity.
This is as much as the massive Clyde Dam in Central Otago and enough electricity to supply 8 per cent of New Zealand's power needs.

ELECTROMAGNETIC INDUCTION

When a conductor moves in a magnetic field an emf is induced in that conductor. This will be discussed in greater detail later.

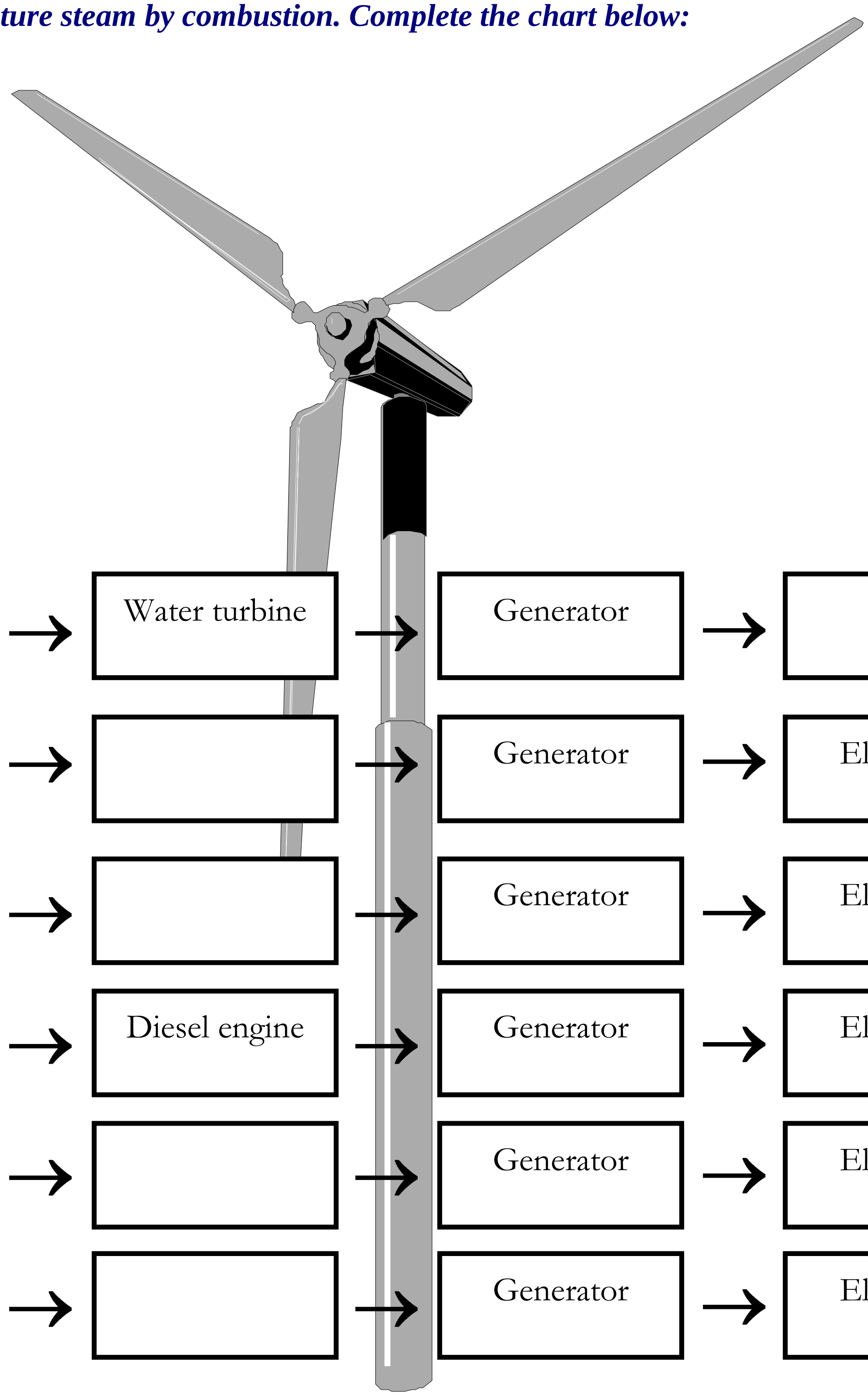
65. ***Electromagnetic generation is the most common method used to produce electrical power in large quantities. In many cases water is converted into steam - which then provides power to a steam turbine. This steam turbine is then the "prime mover" of the Electromagnetic Generator. Complete the chart below:***

Burning rubbish	+		=	Steam	→	
	+	Water	=	Steam	→	Steam turbine
Coal (fire)	+	Water	=	Steam	→	Steam turbine
	+	Water	=	Steam	→	Steam turbine
	+	Water	=	Steam	→	Steam turbine



ASIDE:
There was a possibility of a Nuclear Power Station being built north of Auckland to meet the growing demands of New Zealand's electricity needs back in 1981 - the abundance of gas from Taranaki's Maui field meant the 1000MW gas-fired thermal station at Huntly negated the need for a nuclear power plant.

1. ***Electromagnetic generation is driven by other prime movers that do not first manufacture steam by combustion. Complete the chart below:***



Water	→	Water turbine	→	Generator	→	
Tidal movements	→		→	Generator	→	Electricity
Gas	→		→	Generator	→	Electricity
	→	Diesel engine	→	Generator	→	Electricity
Geothermal energy	→		→	Generator	→	Electricity
	→		→	Generator	→	Electricity

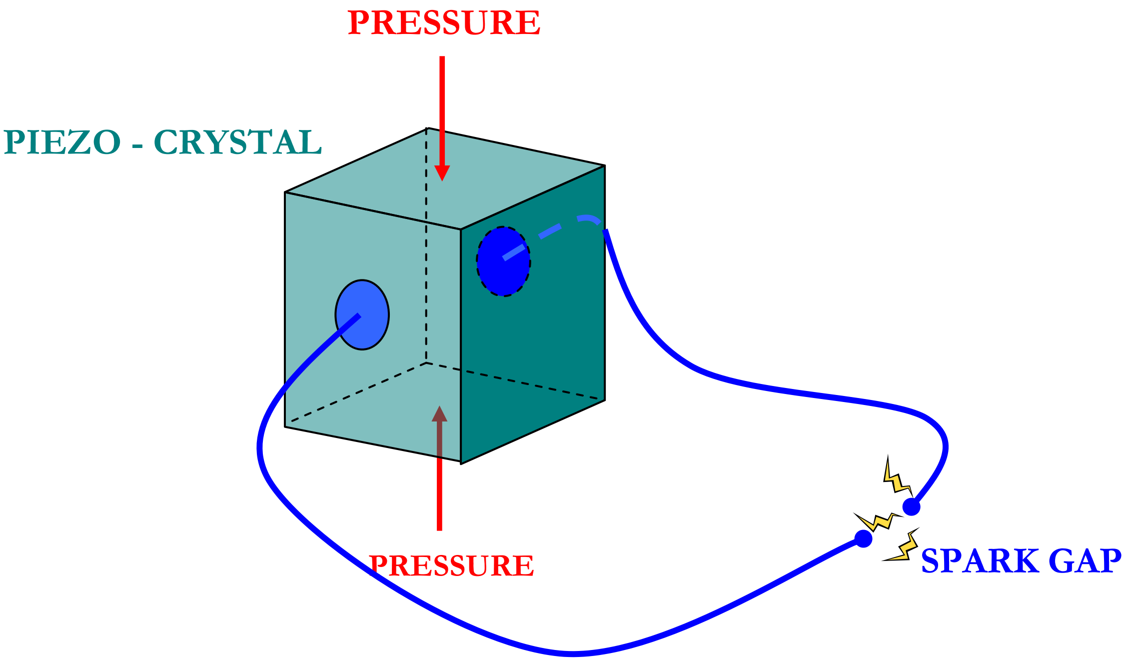
OTHER SOURCES OF ELECTRICITY

2. ***Complete the table below:***

SOURCE TYPE	COMMON EXAMPLE	COMMON APPLICATION
Chemical		
	The application of a mechanical force to opposite sides of a quartz crystal (lead zirconate) will produce an emf across the opposite faces of the crystal.	
Electrostatic Sources or Electrification by Friction	Wimshurst machine or van de Graaff generator	
	Two dissimilar metals are joined at the point where heat is applied.	
		Commonly used for battery charging in remote areas.

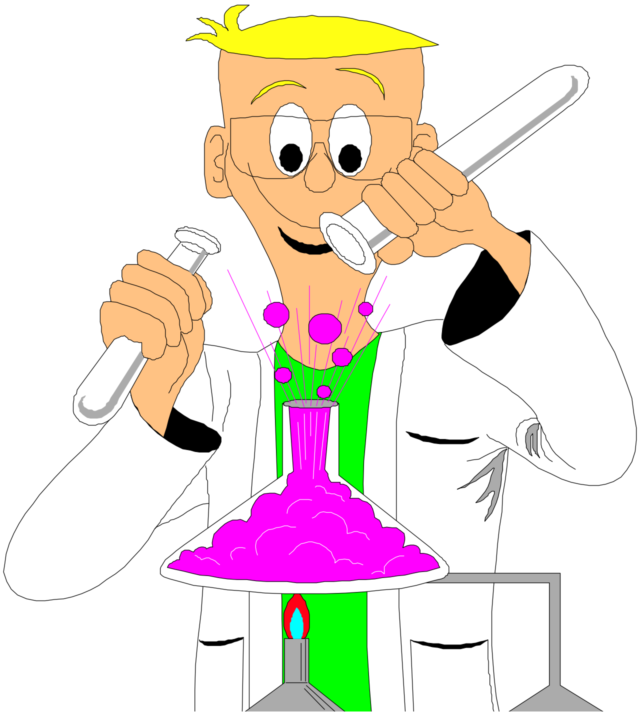
BASIC THERMOCOUPLE

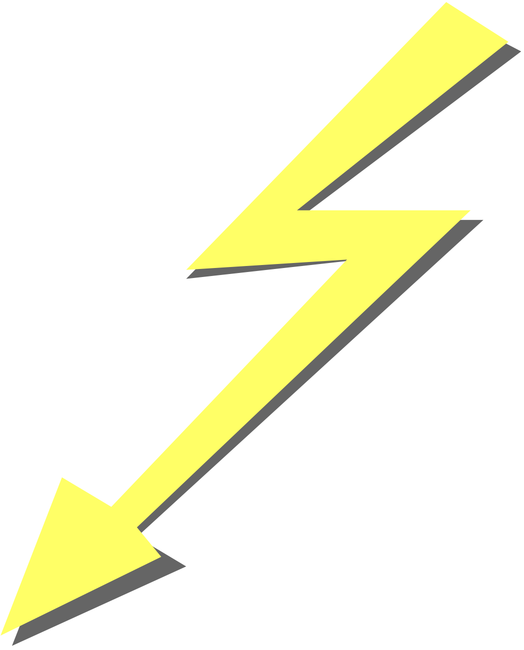
PIEZO-ELECTRIC STRIKER IGNITION SYSTEM



2 – Electrochemistry

Electrochemistry is the study of chemical reactions that produce electricity, and the chemical phenomena produced as a result of the application of electricity.





Some of these reactions and phenomena are more desirable than others.

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ELECTROLYSIS

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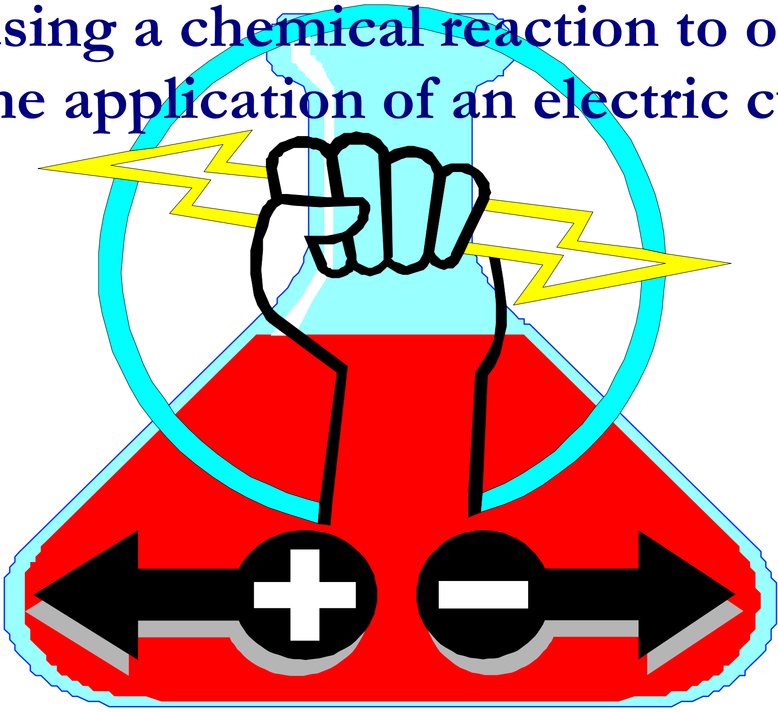
Electrolysis is the process of causing a chemical reaction to occur due the application of an electric current.



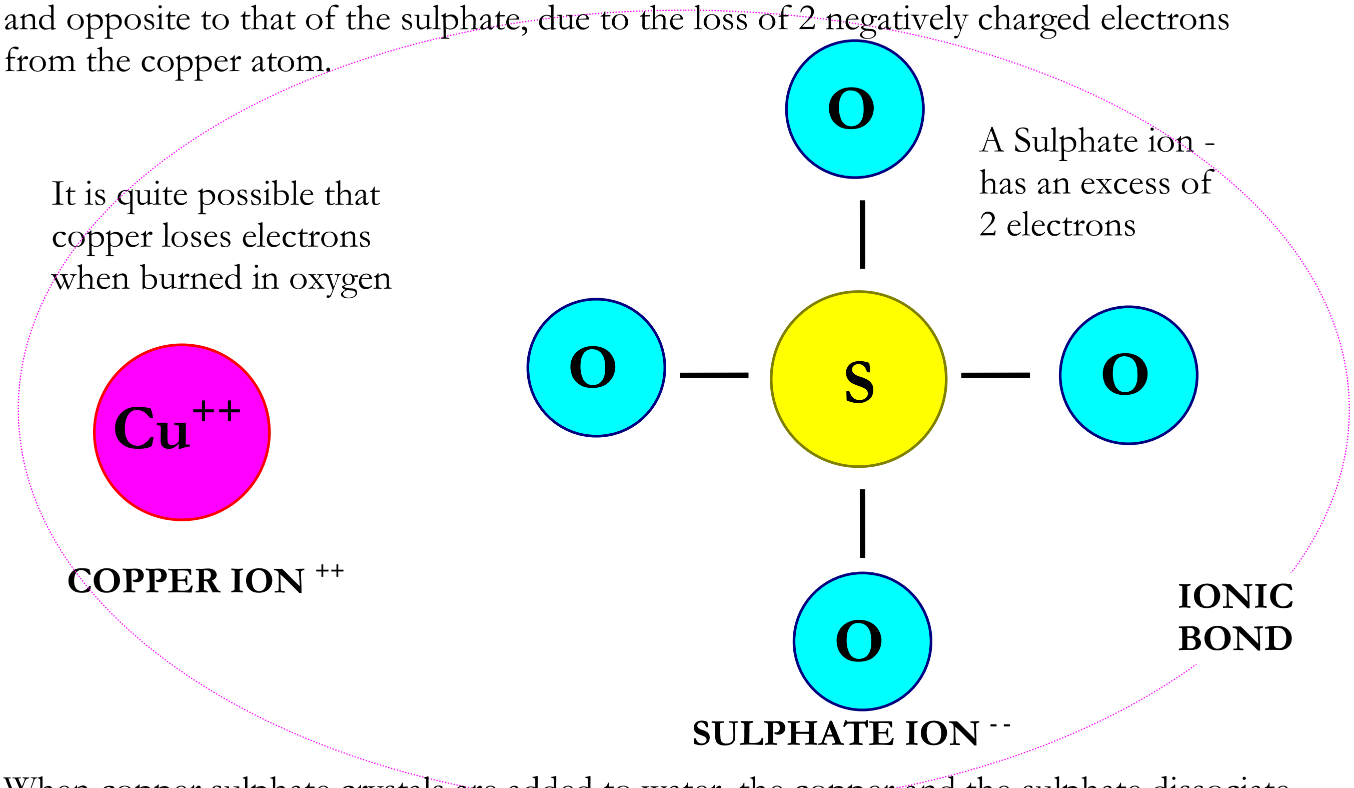
An “ion” is the name given to a molecule (or atom) that has either more or less electrons than would normally be associated with that molecule - e.g. if an atom of copper were to lose an electron, it would be missing a negatively charged electron (all electrons have a negative charge) and would therefore have a nett positive charge - it would be a positively charged copper ion (Cu^+).

If on the other hand an atom, that had perhaps 7 valence electrons (e.g. fluorine) attracted a free electron from somewhere else, the combined charge of all of the electrons would now be greater than that of the protons (an atoms electrons normally have a charge that is equal and opposite to the charge of the protons) - and the fluorine would no longer be a fluorine atom, it would be a negatively charged fluorine ion (F^-).

Unlike charges attract - this is the fundamental law of electrostatics. This would suggest then that a positively charged ion would attract a negatively charged ion - and this is certainly the case. The bond that is created by oppositely charged ions is an ionic bond.

An example:
$CuSO_4$ (copper sulphate) in crystal form consists of copper ions^{++} and sulphate ions^{--}.

These two ions are ionically bonded due to the copper ions having a charge that is equal and opposite to that of the sulphate, due to the loss of 2 negatively charged electrons from the copper atom.


It is quite possible that copper loses electrons when burned in oxygen

A Sulphate ion - has an excess of 2 electrons

COPPER ION $^{++}$

SULPHATE ION $^{--}$

IONIC BOND

When copper sulphate crystals are added to water, the copper and the sulphate dissociate (break their ionic bond) to form copper ions^{++} and sulphate ions^{--}.

Because these ions are electrically charged, the copper sulphate solution is able to carry an electric current and is thus known as an electrolyte.

When a potential difference (measured in volts) is applied to two electrodes in the copper sulphate solution, the copper ions^{++} move toward the negative electrode (cathode), and the sulphate ions^{--} move toward the positive electrode (anode).

As the cathode had an excess of free electrons (because they are connected to the negative terminal of an external supply), two electrons can pass from the cathode to each copper ion that comes into contact with the cathode. This transforms those copper ions back into copper atoms, which will remain deposited on the cathode.

The negatively charged sulphate ions migrate toward the anode and discharge their excess electrons thus completing an electric circuit.

An electric current passing through an electrolytic solution is associated with the movement of ions.

The terminals from which the current flows are called electrodes, the positive being the anode and the negative being the cathode.

The ions that move or migrate toward the anode are anions, while those that migrate toward the cathode are cations.

It was stated on the previous Page, that copper ions^{++} combining with electrons at the cathode were deposited on the cathode - this would mean that the copper cathode would have more copper atoms on it as a result of this electric current flow (and thus it must increase in weight).

The amount of chemical change produced by an electric current is proportional to the quantity of electricity.

Faraday's Laws of Electrolysis

1. *The mass of metal deposited is proportional to the current flowing and its duration*

2. *The mass of metal deposited by a quantity of electric charge is proportional to the electrochemical equivalent of the metal.*

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Theory

In 1834, Michael Faraday, on the basis of extensive experimentation, concluded that the extent of an electrolytic chemical change was proportional to the quantity of electric current flowing.

The amounts of different substances liberated by a given quantity of electricity are proportional to their chemical equivalent mass where:

$$\text{chemical equivalent mass} = \frac{\text{relative atomic mass}}{\text{valency}}$$

1 coulomb of charge requires 6.25×10^{18} electrons to pass a given point.

1 coulomb = 1 ampere × 1 second

6.023×10^{23} atoms of any element constitutes 1 mole of that element.

relative atomic mass - every atom has a mass. This mass is measured in grams, and is the weight of one mole (6.023×10^{23} atoms) of that element or compound.

The "relative atomic mass" (atomic weight) of copper is 63.54

valency - the number of electrons in the outermost energy level.

The transfer of one copper ion from the electrolyte to one copper atom on the cathode in the "copper sulphate experiment" electrolytic reaction requires that two electrons move from the cathode into the electrolyte and then to the anode and back to the supply.

5. ***How many coulombs of energy are supplied if 1A flows for 20min?***

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6. ***How many free electrons must pass a given point to equate to the number of coulombs in Answer 5?***

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7. ***In the "copper sulphate experiment", how many atoms of copper would the movement of the number of electrons in your answer to Question 6 facilitate?***

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8. ***How much would this much copper weigh?***

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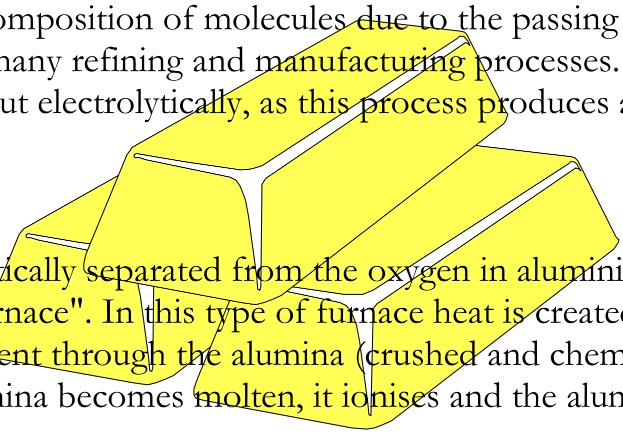
INDUSTRIAL APPLICATIONS OF ELECTROLYSIS

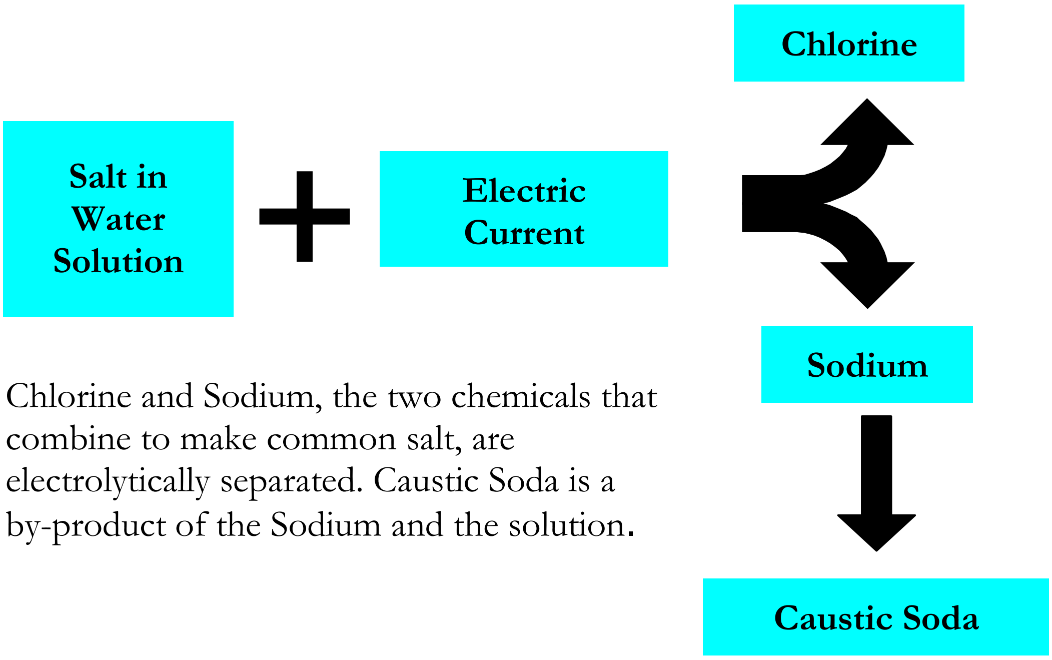
One undesirable effect of naturally occurring electrolysis is **corrosion** - but we will discuss this later.

The dissociation or decomposition of molecules due to the passing of an electric current (electrolysis) is used in many refining and manufacturing processes. The **refining** of many metals is carried out electrolytically, as this process produces a very pure end product.



Aluminium is electrolytically separated from the oxygen in aluminium oxide (alumina) using an "electrolytic furnace". In this type of furnace heat is created by passing large amounts of electric current through the alumina (crushed and chemically refined bauxite). When the alumina becomes molten, it ionises and the aluminium metal is deposited electrolytically.

Magnesium is the fourth most abundant metal on earth, and is quite commonly extracted from **seawater** by electrolysis. One cubic metre of seawater contains more than one kilogram of magnesium, but requires 15kW of electricity in the electrolysis process.



Chlorine and Sodium, the two chemicals that combine to make common salt, are electrolytically separated. Caustic Soda is a by-product of the Sodium and the solution.

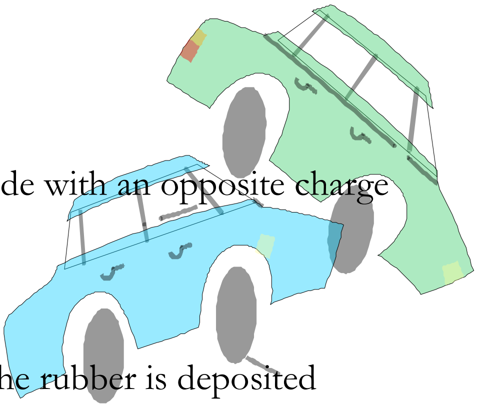
ELECTROPLATING

Electroplating is another industrial application of electrolysis. Electroplating is the name given to the process of depositing layers of precious metals on cheaper base metals, and of depositing layers of metals or alloys on other metals for strengthening, conductivity, corrosion resistance or just good looks.

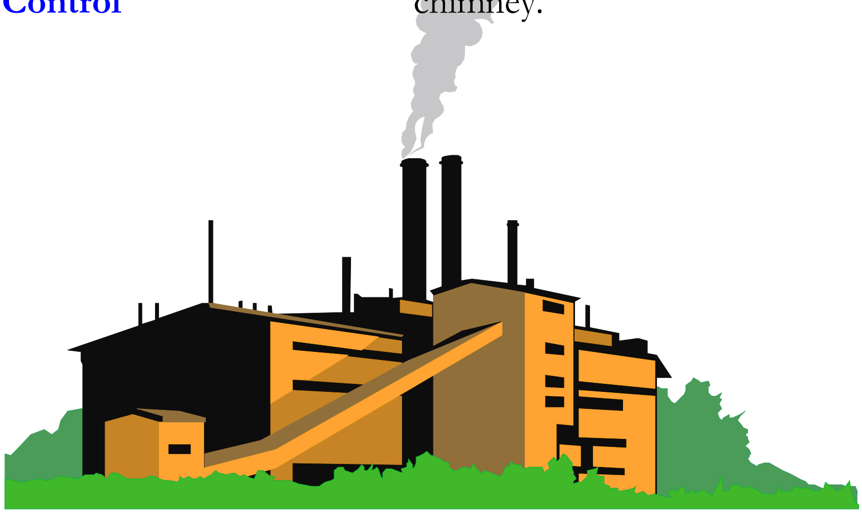
The object to be electroplated becomes the cathode, and is immersed in a bath of a solution of the metal with which it is to be coated. Just like in the copper sulphate experiment, the metal from the solution is deposited on the cathode when a d.c. current flows from anode to cathode which causes the solution to ionise.

Cadmium, chromium, copper, gold, nickel, silver and tin are metals commonly used in electroplating.

OTHER ELECTROLYTIC INDUSTRIAL PROCESSES:



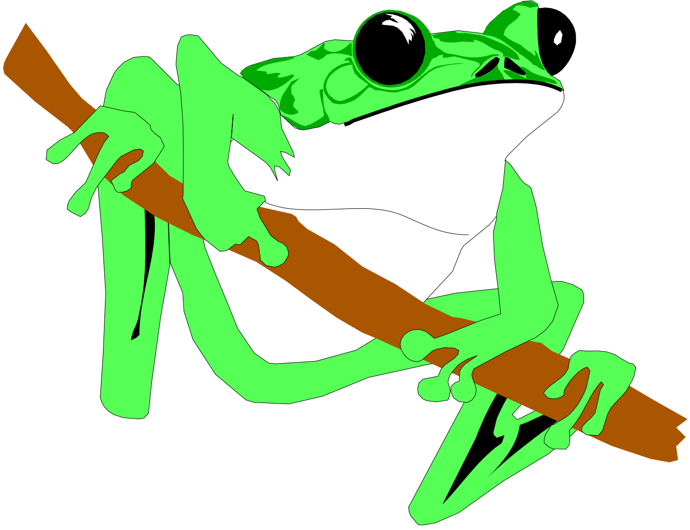
Painting (car parts etc.)	The car part becomes an electrode with an opposite charge to that supplied to the paint.
Rubber Moulding	The mould is an electrode and the rubber is deposited evenly over the entire mould (the thickness of the rubber is dependant on the amount of current flow and the duration - Faradays Law)
Pollution Emission Control	Pollutants are deposited on an electrode inside the chimney.



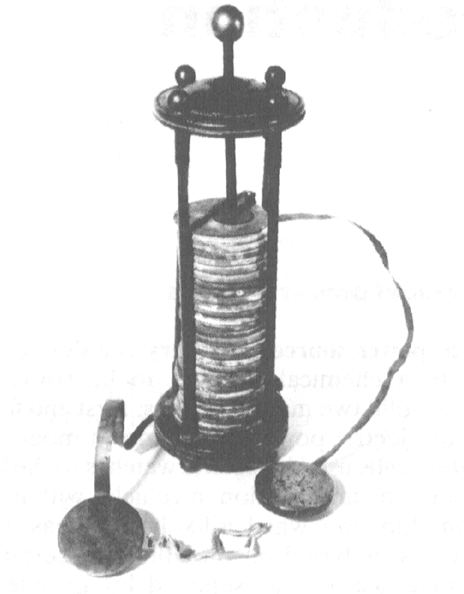
THE PRIMARY CELL

(NON-RECHARGEABLE)

Luigi Galvani (an Italian anatomy professor) first observed the electrical effect of dissimilar metals. Galvani observed that dissected frogs legs suspended from an iron balcony by a copper wire would convulse every time they came in contact with the iron.



Shortly after Galvani's discovery, Allesandro Volta (a physics professor) proved that by touching one end of the nerve in a frogs leg with a piece of copper, and the other end of the same nerve with a piece of iron, the muscle spasm would occur.



This discovery led to the construction of the "**Voltaic Pile**", which consisted of a disk of copper separated from a disk of zinc by a cloth soaked in brine (a salt solution). When the frog's leg came in contact with both the copper and the zinc this spasm occurred. When the copper and zinc discs were stacked one on top of each other (copper-zinc-copper-zinc-copper-zinc etc., and hence the "pile") the spasm was greater.

The magnitude of the potential difference between two dissimilar metals (measured in volts), will determine the maximum voltage available from that type of cell.

9. *Complete the following table:*

The Electromotive Force Series of Metals (referenced to hydrogen)

Metal	Electrode Potential in Volts
Magnesium	-2.37
Aluminium	
Zinc	
Chromium	-0.71
Iron	-0.44
Cadmium	-0.40
Tin	-0.14
Lead	
Hydrogen	0.00
Copper	
Silver	
Platinum	+1.20
Gold	+1.42

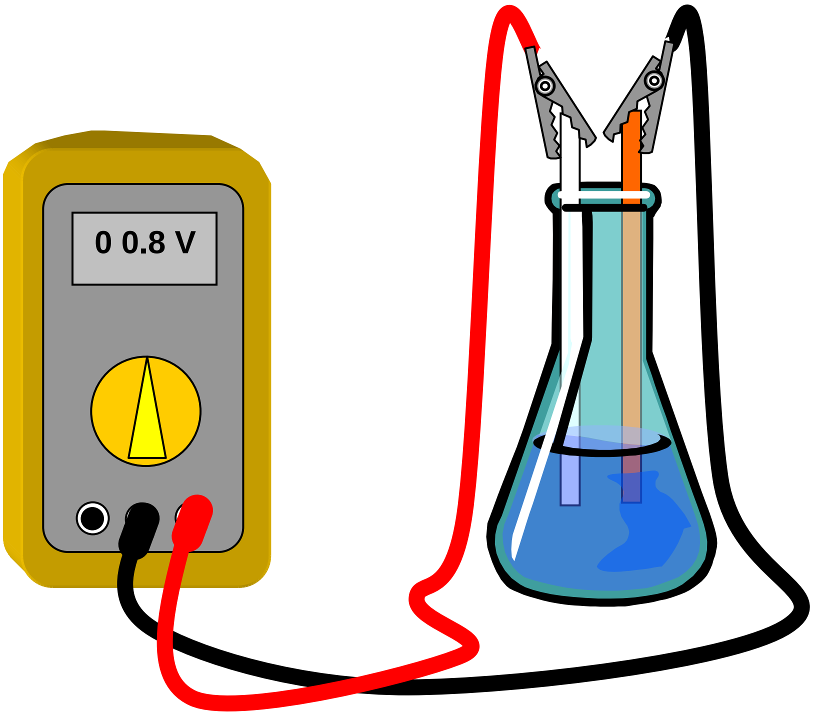
We can assume then, that an electric cell formed by the junction of **magnesium** and **platinum** in an electrolyte would have a potential of **+1.20V - -2.37V = 3.57V**.

Platinum would be known as the "positive electrode" and the magnesium would be the "negative electrode".

THE VOLTAIC CELL

It didn't take much to rearrange this Voltaic Pile so that the copper and the zinc were immersed in the brine, and so long as the copper and the zinc did not come in contact with each other this eliminated the need for the brine-soaked cloth.

Further experimentation led to the replacement of the brine with diluted sulphuric acid, as this produced a greater potential difference (0.8V) between the copper and the zinc electrodes. This became known as the "Voltaic Cell".



Zinc and Copper electrodes in H_2SO_4

A voltaic cell can only transform chemical energy into electrical energy. It is because of this a voltaic cell will only produce electrical energy until it has used up the chemical ingredients - more ingredients must be added to maintain this electrochemical process; and for this reason it is a **primary cell**.

10. What is the maximum emf that should have been available from Volta's Zinc and Copper cell?

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THE LECLANCHÉ CELL

A battery is actually the term for more than one cell connected in series, but more often than not, we refer to a single cell (like the one you buy in the corner store) as a battery.

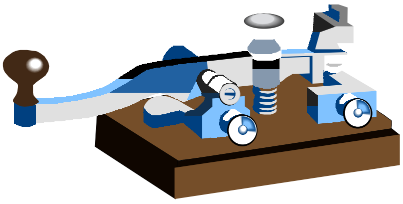
One of the most common batteries in use today is the Leclanché cell (popularly called a dry-cell* battery or a flashlight battery). It was invented in France in the 1860's, and is the battery that finds common use in flashlights, radios, electrical test meters, and other low powered portable equipment that does not have rechargeable batteries.

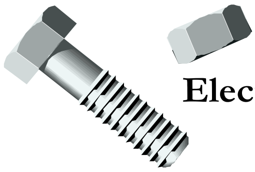
*Known as a "dry-cell" because of the powder electrolyte - Leclanché also invented a wet-cell with a liquid electrolyte.

EARLY HISTORY OF THE PRIMARY CELL

Primary cells found initial commercial use in the following areas:

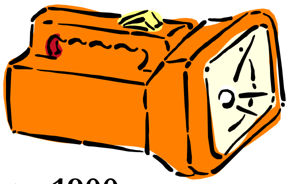
Telegraph - 1830





Electroplating -1840

Electric Bell - 1870



Flashlight - 1900

As people gained more knowledge about the electromagnetic phenomenon that occurred when a conductor was moved in a magnetic field (from 1870 onwards), a new type of battery was to be developed - one that could be replenished by supplying an electric current, produced by a d.c. generator, in the opposite direction to the direction of normal current flow from the cell.

THE SECONDARY CELL

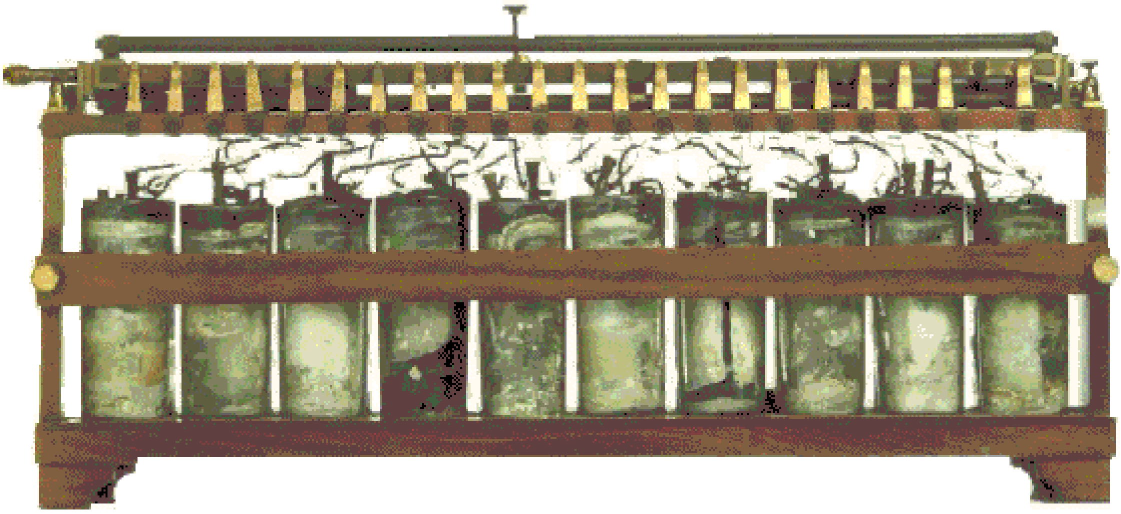
The secondary cell was the next step in the evolution of the battery. A battery that could be replenished by supplying it with a d.c. current in the opposite direction to that of the discharging current.

This type of cell is also known as a "storage cell"

THE LEAD-ACID BATTERY

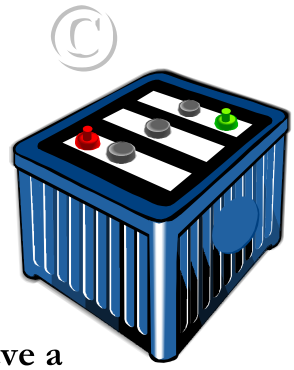
Another Frenchman Gaston Planté invented the Lead-acid battery, probably the most widely used storage battery, as early as 1860.

The electrodes in Plantés original lead acid cell were two coiled strips of lead separated by a piece of cloth. The diagram below illustrates 10 of these cells - a far cry from the modern version of the lead-acid battery under the hood of your car.



The lead acid battery has been so popular due to its low cost, the availability of the raw materials (lead and sulphur), the ease of manufacture and this type of batteries lifespan. The lifespan of a typical lead-acid car battery is about 4 years, and more than 1/3 of all lead today goes into the manufacture of lead-acid batteries.

THE CONSTRUCTION OF THE MODERN LEAD-ACID BATTERY



The modern lead-acid battery (sometimes known as a lead-lead-acid battery) as the name would suggest, still requires both the positive and the negative electrode be made out of lead (lead in different forms). The electrolyte is sulphuric acid diluted in water.

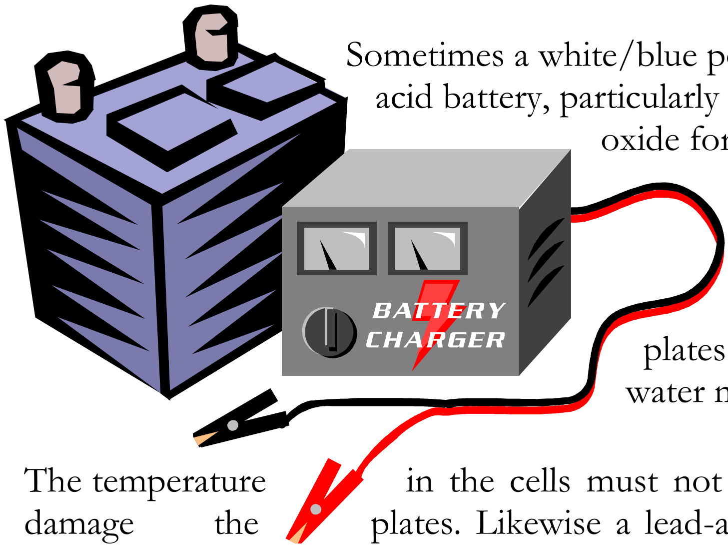
The construction of the lead-acid battery is well covered by **"Electrotechnology.." on Pages 58 – 62. You are required to have a comprehensive understanding of the lead-acid battery to pass this unit.**

LEAD-ACID BATTERY MAINTENANCE

GENERAL

Lead-acid batteries (and other storage batteries) can store extremely large quantities of energy - this energy can be released in an instant. It is important that you treat these batteries with respect. Short-circuiting a fully charged "gassing" lead-acid battery can cause an explosion that will destroy the battery and injure the person working on it. If you must use an over sized spanner - it should have an insulated handle.

A lead acid battery should be charged immediately after discharge, and fully charged then drained of electrolyte, before being stored for any extended lengths of time.



Sometimes a white/blue powder may form on the top of a lead-acid battery, particularly around the terminals. This powder is an oxide formed by electrolysis, and should be removed using a damp cloth.

The level of the electrolyte must be kept to approximately 10mm above the plates - use distilled water or rainwater, tap water may shorten the life of the battery.

The temperature in the cells must not get too high (i.e., ≤ 40°C) as this will damage the plates. Likewise a lead-acid battery should never be allowed to get too cold as this will shorten the life expectancy of the battery (don't leave a battery on a cold concrete floor).

Lead-acid battery cells will "gas" while charging, for this reason the lids or caps should be removed if they are removable (or according to manufacturers instruction).

The gases given off by a charging lead-acid battery are Oxygen and Hydrogen (oxygen helps fires burn and hydrogen likes to explode) so battery charging should only be carried out in a well-ventilated room.

11. ***It is obvious then that there are many dangers associated with the use and charging of a lead acid battery. List some of the dangers and precautions below:***

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12. ***How can you tell the state of the electrolyte in a lead-acid cell?***

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NICKEL CADMIUM CELLS

First patented by Waldemar Jugner of Sweden in 1899, nickel-cadmium cells are also secondary cells - they can be recharged. Nickel-cadmium batteries are Nickel-alkaline batteries, and are also referred to as “Jugner cells”, “NiCads” and “Cadmium-nickel oxide batteries”. NiCads are fully sealed, maintenance free rechargeable long-life batteries.

NiCads are known as “recombining cells”, this means the dissociated active material produced due to overcharging these types of cells will recombine, and no nett change will occur in the cell as a result of overcharging (although overheating may be a serious problem).

Disadvantages:

Cadmium is several times the cost of lead, cell construction cost is greater than lead-acid, overall cost for same amount of energy stored may be 10 times that of lead-acid. Health risks associated with the use of cadmium.

The “memory effect”:

Another disadvantage of the NiCad battery is the memory effect. This occurs when the battery is constantly recharged when it has not been completely discharged. To avoid this effect, NiCad batteries should be fully discharged before recharging begins (this procedure is quite often noticeable when routine maintenance discharging is being carried out on emergency lighting). If this is not done the problem will manifest itself in the cell losing voltage too quickly.

Advantages:

Long cycle-life, continuous overcharge capability, relatively high rates of charge and discharge, almost constant discharge voltage, low temperature capability, low maintenance, reliable, used in: emergency lighting, engine starting, cordless electrical appliances, aircraft and satellites, extremely long shelf-life in any state of charge, higher initial purchase price but usually prove cost effective when compared with primary cells.

Nickel cadmium button cells

Complete the following questions:

13. Name a situation where naturally occurring electrolysis is a problem:

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14. Name four desirable practical applications of electrolysis:

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15. What is a “recombination cell”?

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16. What is the difference between a “primary cell” and a “secondary cell”?

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17. What is meant by polarisation of a cell?

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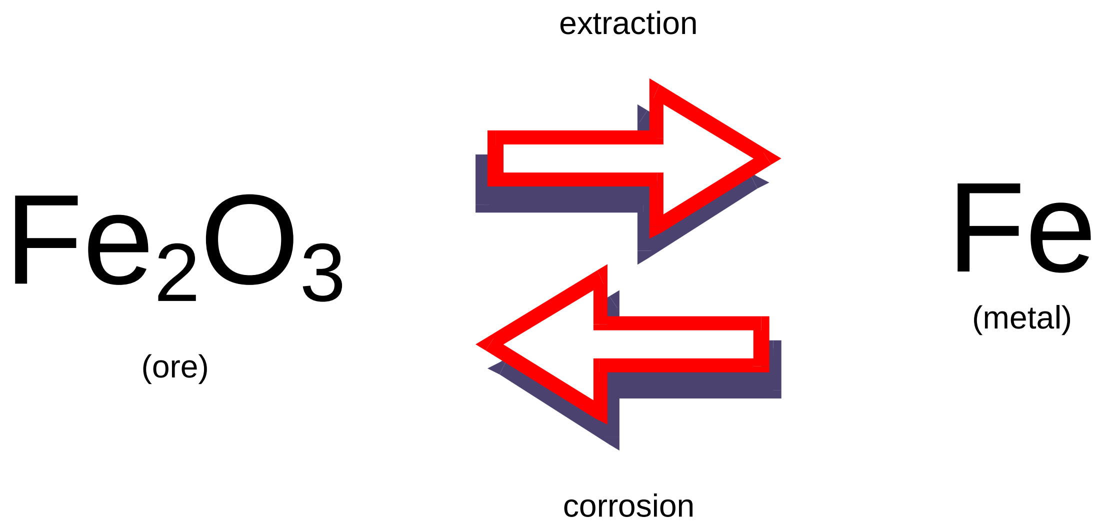
18. ***Complete the TABLE below to identify the characteristics and typical applications of cells in common use:***

CELL - TYPE	PHYSICAL SIZE (L, M or S)	NOMINAL CELL VOLTAGE	PRIMARY OR SECONDARY	FEATURES AND APPLICATIONS
LEAD-ACID				
NICKEL-IRON (alkaline)	*M,L*	*1.2*	*S*	*Welded steel casing, mechanically & electrically robust, will not deteriorate if left discharged, long life, high cost – Engine starting, standby power*
NICKEL-CADMIUM		*1.25*		
CARBON-ZINC	*S,M*			
ALKALINE	*S*		*P*	
MERCURY				
SILVER OXIDE				
ZINC-AIR				*Extremely long life cell, constant output voltage – Hearing aids, paging devices*
LITHIUM	*S*			

CORROSION

Oxygen in the air, or dissolved in water will attack most metals. The evidence that proves this can be seen in the fact that no metals occur naturally in their metallic state*, but tend to be found combined with oxygen or sulphur in their ores.



The reaction that takes place to reduce metals is known as oxidation, and this process takes place at the metals surface.

This means that the reaction can be controlled, and reduced, by modifying conditions at the metals surface.

There are two main mechanisms of corrosion -

"**dry corrosion**" due to the reaction of metals with gases#; and

"**wet corrosion**" which is by far the most common and the subject we discuss here:

*Copper, Gold and Silver can occur naturally in their metallic state.

#This is also known as tarnishing, and tends to increase at greater temperatures. Sometimes dry corrosion will leave the metal covered in a layer of oxide, which will prevent any further tarnishing so long as that oxide layer is not punctured.

WET CORROSION

Wet corrosion is an electrochemical reaction, and is responsible for the rusting of iron roofs, ships hulls, bridges, the authors car and buildings located in industrial areas. This type of corrosion happens quite readily at ambient (normal) temperatures.

There are certain requirements for wet corrosion to occur:

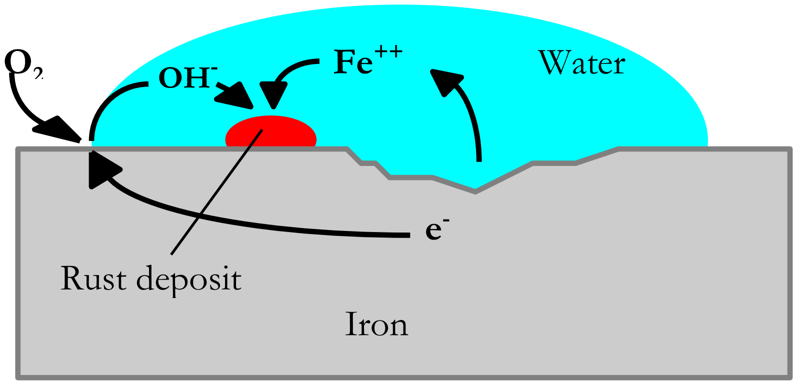
1. there must be an anode and a cathode;
2. there must be an electrical potential between the anode and the cathode;
3. there must be a metallic path electrically connecting the anode and cathode;
4. the anode and cathode must be immersed in an electrically conductive electrolyte.

When a metal corrodes, atoms of the metal are lost from the surface as cations (positive ions), leaving behind a number of electrons.

The electrons left behind by the cations migrate through the metal to a location where the iron is exposed to oxygen being dissolved in water.

A hydroxide ion (OH^-) forms out from the reaction between the oxygen, the water, and the free electrons.

The Iron ions++ are attracted to the Hydroxide ions- forming iron hydroxide that in turn reacts with the Oxygen dissolved in the water and forms Fe_2O_3 (rust).



The place the cations leave from is the anode, and the area the electrons are removed from the iron at is the cathode.

DISSIMILAR METAL CORROSION

Whenever there is a contact between different metals, and there is a common electrolyte (which could just be water or moisture), those dissimilar metals will form an electric cell. The more active metal will become the anode (this is where the current enters the electrolyte), the other will form the cathode. The corrosion takes place in the anode, this is known as **galvanic** corrosion.

The table below is based on the galvanic corrosion of dissimilar metals in sea water:

ACTIVE (anodic)
1. Magnesium
2. Zinc
3. Aluminium alloys (Al-Mg-Si)
4. Aluminium[#1]
5. Mild Steel (Fe - low carbon)
6. Cast Iron (Fe - high carbon)
7. Lead
8. Tin
9. Naval Brass (Cu-Zn-Sn)
10.Muntz Metal ((60Cu-40Zn)
11.Yellow Brass (70Cu-30Zn)
12.Copper
13.Bronzes (Cu-Sn-Al-Mn)
14.Cupronickel (70Cu-30Ni)
15.Nickel
16.Cupronickel (Monel 70Ni-30Cu)
17.Inconel (76Ni-16Cr-7Fe)
18.Hastelloy (58Ni-17Mo-6Fe-14Cr-5W)
19.Silver
20.Stainless Steel
21.Titanium
22.Platinum
23.Gold
NOBLE (cathodic)[#2]

The use of metal combinations that are found in the same area of the Galvanic Series will reduce the likelihood of galvanic corrosion becoming a major problem.

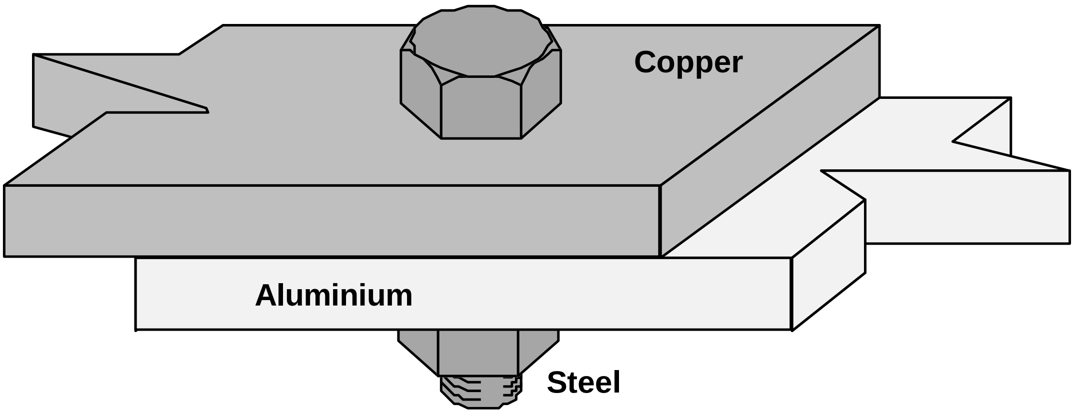
The combined use of these similarly positioned metals, would reduce the electrical potential between the two and thereby reduce the electric current flow (remember Faraday's Law - the metal removed was proportional to the current flow).

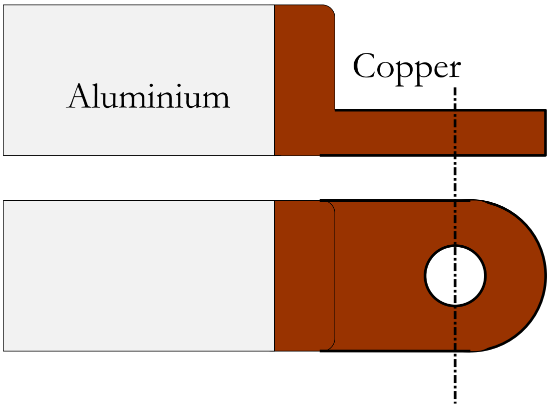
The materials nearer the top of this list are more likely to deteriorate when installed near the materials lower on the list - and emersed in sea water.

[#1] Aluminium (like lead and zinc) tend to corrode very rapidly, but in doing so form a fine layer of oxide on the surface of the metal that reduces any further corrosion.

[#2] Noble metals - when the new kings and queens recalled all of the coins from the reign of the previous monarch for recoining with the new queen/kings head, they found that very little of these valuable metals had been lost due to corrosion.

19. ***Outline the corrosion that would take place in this example, including which metal would be worst effected and why:***





Bimetal lugs, like those shown here, and special conducting pastes are used to prevent electrolytic corrosion due to dissimilar metal junctions - in electrical applications.

©

Corrosion Prevention

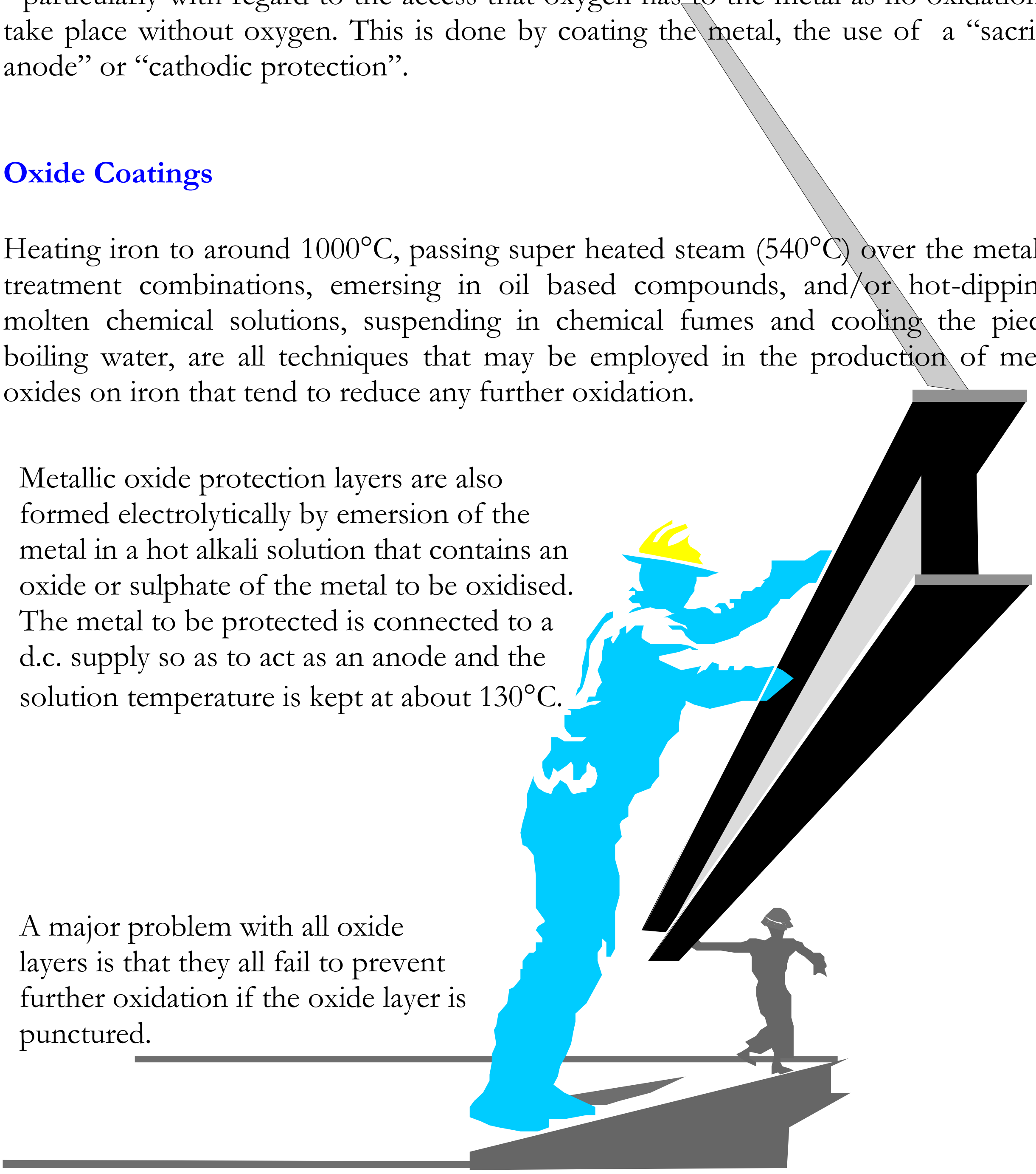
We mentioned earlier that corrosion occurs at the metals surface. It stands to reason then that we can halt or reduce corrosion by altering the conditions at the metals surface - particularly with regard to the access that oxygen has to the metal as no oxidation can take place without oxygen. This is done by coating the metal, the use of a "sacrificial anode" or "cathodic protection".

Oxide Coatings

Heating iron to around 1000°C, passing super heated steam (540°C) over the metal, gas treatment combinations, emersing in oil based compounds, and/or hot-dipping in molten chemical solutions, suspending in chemical fumes and cooling the piece in boiling water, are all techniques that may be employed in the production of metallic oxides on iron that tend to reduce any further oxidation.

Metallic oxide protection layers are also formed electrolytically by emersion of the metal in a hot alkali solution that contains an oxide or sulphate of the metal to be oxidised. The metal to be protected is connected to a d.c. supply so as to act as an anode and the solution temperature is kept at about 130°C.

A major problem with all oxide layers is that they all fail to prevent further oxidation if the oxide layer is punctured.



This unit was written by A. Kawana

METALLIC PROTECTIVE COATINGS

ELECTRO-POSITIVE COATING

A good practical example of an electro-positive metallic protective coating is steel protected by a coating of zinc. Zinc will oxidise readily and a layer of zinc oxide will form on the surface of the zinc that will prevent any further corrosion.

Zinc is commonly used due to the availability, comparatively low cost and ease of application.

Steel protected by a continuous layer of non-porous zinc will withstand normal atmospheric conditions well - even if the surface is punctured, the zinc will tend to corrode before the steel it protects does. The zinc layer once punctured however will tend to disappear as it will act as an anode in a steel - zinc voltaic cell, and thus provide no further protection to the steel.

Galvanic Series

ACTIVE (anodic)
1. Magnesium
2. Zinc
3. Aluminium alloys (Al-Mg-Si)
4. Aluminium
5. Mild Steel (Fe - low carbon)
6. Cast Iron (Fe - high carbon)
7. Lead

Zinc can be applied to steel by 4 methods:

1. Hot-dipping;
2. Hot-melding ("Sheradising" for intricate parts);
3. Electrolytic plating; and
4. Painting

ELECTRO-NEGATIVE COATING

Tin, lead and copper are attacked only slowly by usual atmospheric influences, and may be used to protect a base metal from the atmosphere. Not only do they corrode only slowly, but the result of the corrosion of these metals produces a protective layer that protects against further corrosion also.

The protective layer of either an electro-negative or electro-positive metallic protective coating will only work properly if the coating is continuous and unpunctured.

THE USE OF SEALANTS AND PAINTING TO PREVENT CORROSION

Companies like "3M" manufacture amalgamating rubber tapes and epoxy sealants, the purpose of which being, the tape/epoxy combination impedes the corrosion of the metal it covers by reducing the access of oxygen to the metals surface. "DENSO" tape may also fill this oxygen exclusion function.

Epoxy moulds used in underground cable jointing also provide corrosion protection by halting the access of oxygen to the metals inside the join.

A particularly economic method of reducing corrosion is painting.

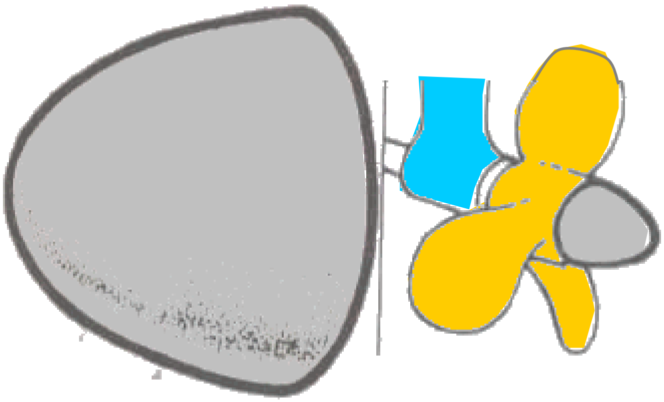
Most structural steelwork to be left exposed in the finished structure is protected by paint. These paints may be zinc, synthetic rubber, oil, coal, bitumen, epoxy or polyurethane based.

There is an increasing popularity in paints that can be used for both the under coating and the finishing coat.

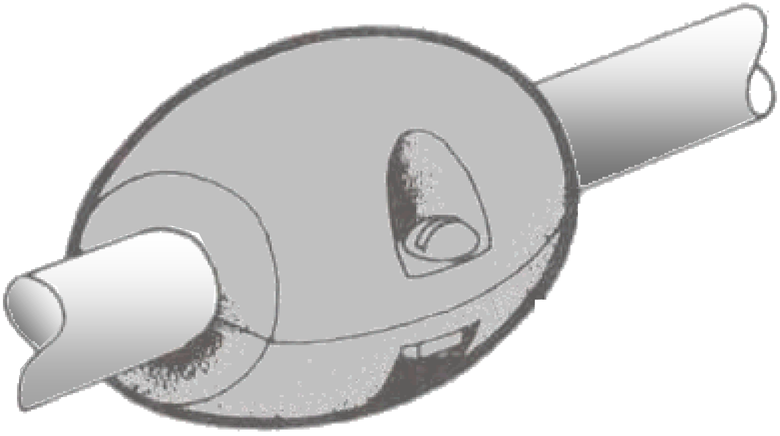
It is very important however, that any paint used for undercoating will not stimulate the corrosion of the metal it is supposed to protect. It is also important that this first layer of paint be non-conductive.

Have a good talk to the man at the paint shop.

SACRIFICIAL ANODES

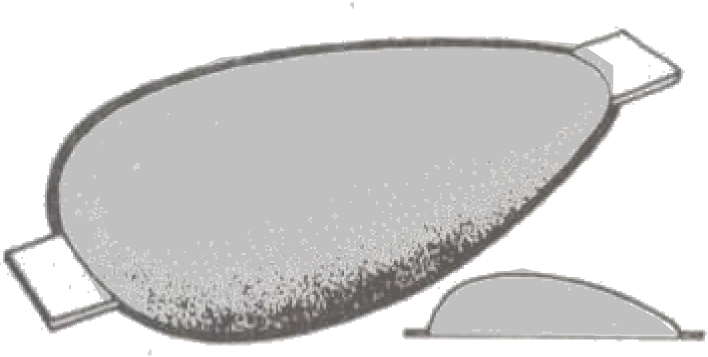


An electrolytic cell is deliberately created on boats with exposed metal that may be subject to corrosion.



Sacrificial zinc anodes are mounted near the areas that are most susceptible to corrosion.

The anode is attacked but the propeller, and hull or any steering fixings are protected.



These anodes will eventually be destroyed but are inexpensive to replace.

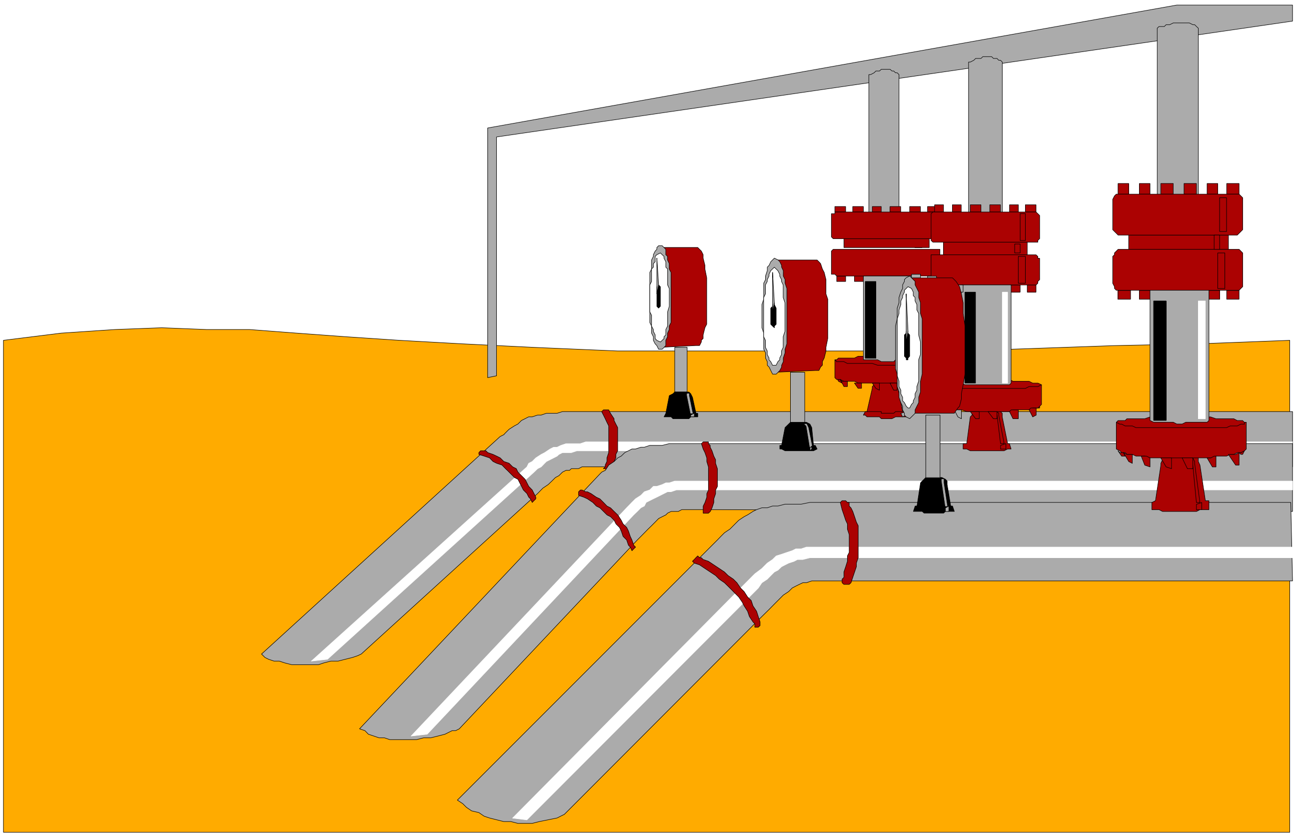
CATHODIC PROTECTION

Cathodic protection is used extensively in the protection of pipelines from corrosion caused by oxygen in the pipelines immediate surroundings.

If the anode that is destroyed in electrolytic corrosion. It stands to reason then, that if we make the metal we wish to protect the cathode, it should not corrode.

This is achieved by simply forcing a direct current to flow to all of the metalwork to be protected from the surrounding electrolyte (soil, water, seawater etc.).

This cathodic protection is usually achieved by burying anodes supplied from the positive terminal of a d.c. supply, along the length of the pipeline which is supplied from the negative terminal of the same d.c. supply. The anode voltage should be at least 0.5V greater than the cathode.



Complete the following questions:

20. ***What problems (in context with this element of learning) can stray electric currents cause, and what can create these stray currents?***

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21. ***Aluminium is very high on the galvanic table. How come (electrolytically) aluminium is used to make lightweight boats?***

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22. ***There are 3 requirements for wet corrosion (electrolytic corrosion); list them below:***

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23. ***There is another form of corrosion that is not electrolytic; what is the name given to this type of corrosion, and when is it most likely to occur?***

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24. ***Which of these elements is a requirement for corrosion?***

Hydrogen ***Sulphur*** ***Oxygen***

25. What is "galvanic corrosion" and how can this be minimised in an electrical situation?

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26. What are two requirements for an oxide layer to protect the metal it coats?

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27. Why is zinc used extensively as a metallic protective coat?

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28. What is a "sacrificial anode"?

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29. Explain "cathodic protection":

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NOTES

NOTES

ANSWERS

4. Electroplating, car part painting, pollution emission control, battery charging and rubber moulding.

6. $6.25\times10^{18} \times 1200 = 7.5\times10^{21}$ electrons

8. $3.75\times10^{21} \div 6.023\times10^{23} \times 63.54g = 395.6mg$

10. $+ 0.34V - - 0.76V = 1.1V$

12. Use a hydrometer – the higher the S.G. the stronger the Sulphuric Acid the less discharged the battery is

14. Electroplating, car part painting, pollution emission control, battery charging and rubber moulding

16. Unlike the primary cell a secondary cell is rechargeable by supplying a current in the opposite direction to the discharge current

20. They can cause electrolytic corrosion - caused by inadequate earthing, inadequate bonding (particularly traction supply systems).

22. Anode, Cathode, electrolyte

24. Oxygen

26. It must be continuous and non-porous

28. A piece of metal from the "anodic" end of the Galvanic Series that is placed adjacent to the metal it is intended to protect. This could be a block of zinc protecting a piece of steel - the anode will be destroyed in preference to the corrosion of the other metal.

$X_L = 2\pi f L$ $MMF = S.\phi$

FORMULAE

$U_m = I.N$

$X_L = 2\pi f L$ $\mu = \mu_0 \mu_r$ $P = V.I.\cos\vartheta$

$P = \sqrt{3}.V_{Line}.I_{Line}.\cos\vartheta$

$X_C = \dfrac{1}{2\pi f C}$ $f_0 = \dfrac{1}{2\pi}\sqrt{\dfrac{1}{LC}}$ $L = \dfrac{\mu_0 \mu_r N^2 a}{l}$

$\Phi = B.a$ $SHC_{H^2O} = 4180$ $\varepsilon_0 = 8.854 \times 10^{-12} F/m$

$C = \dfrac{\varepsilon_0 \varepsilon_r A(n-1)}{d}$

$E = B.l.v.\sin\vartheta$ $\mu_0 = 4\pi \times 10^{-7} H/m$

$\dfrac{R_1}{R_2} = \dfrac{1+\alpha.t_1}{1+\alpha.t_2}$

$S = \dfrac{l}{\mu_0 \mu_r a}$ $\rho_{Cu} = 1.72 \times 10^{-8} \Omega m$

$f = \dfrac{B^2 a}{2\mu_0}$

$S = {}^{O}\!/_{H}$ $C = {}^{A}\!/_{H}$ $T = {}^{O}\!/_{A}$

$R = \dfrac{\rho.l}{a}$

$W = {}^{1}\!/_{2} C.V^2$

$H = \dfrac{I.N.l}{a}$ $v = V(1 - e^{\frac{-t}{TC}})$

$Q = CV$

$R = R_0(1 + \alpha_0.t)$

$Z^2 = R^2 + (X_L - X_C)^2$

$accl._{Gravity} = 9.81 m/s/s$