3.2 Types of Reactions

Achievements Indicators

Upon completion of this sub-strand, students will be able to:

- ✓ Distinguish and describe different types of reactions based on chemical statements and balanced chemical equations.
- ✓ Describe the reactions of oxidation and reduction in terms of transfer of atoms and electrons.
- ✓ Study and write simple oxidation and reduction reactions involving atoms and electrons.
- ✓ Show that electrolysis of molten and aqueous salt experimental set-up involves oxidation and reduction.

Chemical Reactions

Chemical reactions are processes that will cause change in the properties of the substances involved. Most reactions are chemical changes and are *irreversible* and some are *reversible*. Chemical reactions in living things are termed biochemical including photosynthesis, respiration and digestion of food in the gut of animals.

Chemical reactions are part of our daily lives in our homes and communities.

Example

A combustion reaction

The study of chemical reactions is an integral aspect of chemistry. It equips us with the ability to understand and explain the chemical principles that involves changes.

These chemical reactions are:

1. Combustion
2. Synthesis
3. Decomposition
4. Neutralisation
5. Double Displacement
6. Precipitation
7. Oxidation-Reduction

1. Combustion

It is the chemical term for the burning of substances in oxygen to form compounds called *oxides*. Though oxygen does not burn, it is used as it supports combustion.

Metals will burn completely in oxygen to form metallic oxides. The oxides are *ionic* compounds and are *basic* in nature.

Example

Magnesium + Oxygen → Magnesium oxide

$2Mg_{(s)} + O_{2\,(g)} \rightarrow 2MgO_{(s)}$

Combustion of metals may be used to distinguish some common metals as metals burn with distinctive flame.

Non-metals burn completely in oxygen to form non-metal oxides. These oxides are *molecular* substances and are *acidic* in nature. Most are gases at room temperature.

Organic compounds are used as fuels for its high carbon content. It burns completely in oxygen to produce carbon dioxide and water. A lot of energy is released. Incomplete combustion will form harmful products such as carbon monoxide, soot (unburnt carbon) and less heat is released.

Examples

1. Carbon + Oxygen → Carbon dioxide

   $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

2. Glucose + Oxygen → Carbon dioxide + Water

   $C_6H_{12}O_{6(s)} + 3O_{2(g)} \rightarrow 6CO_{2(g)} + 6H_2O_{(g)}$

2. Synthesis

Naturally occurring elements combine chemically to form compounds. When two non-metals combine, a covalent substance is formed. However, metals combine with a non-metal to form ionic compounds.

Example 1 Combination of two non-metals.

$C_{(s)} + S_{(s)} \rightarrow CS_{2(l)}$

Carbon disulphide

Example 2 Combination of a metal and a non-metal

$Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$

Iron sulphide

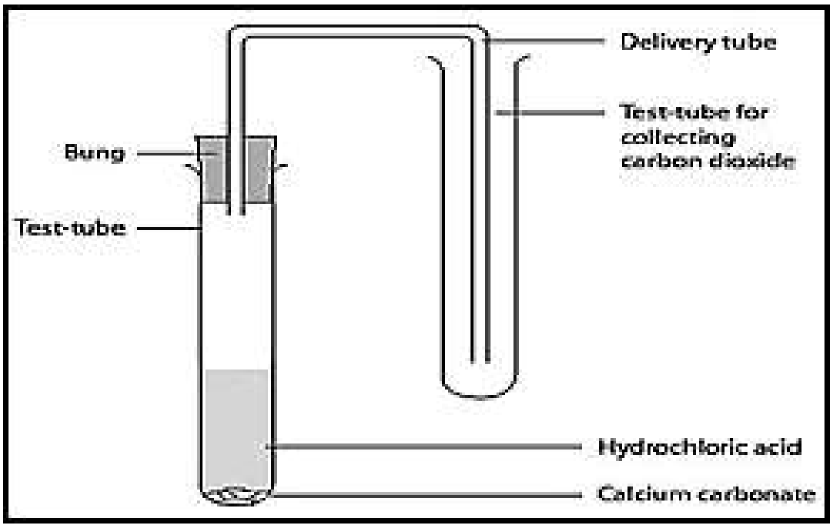
Example 3 Formation of oxides

All combustion of elements is synthesis reaction.

3. Decomposition

Some carbonates and nitrates are decomposed by heat. Carbonates are decomposed to form carbon dioxide and the oxide of the metal.

The set up below shows the laboratory preparation of carbon dioxide by the decomposition of marble chips, $CaCO_3$.



The presence of the carbon dioxide formed can be tested by passing it through lime water.

Example

Calcium carbonate → Calcium oxide + Carbon dioxide

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)}$

Did you know?

White wash for painting decorative garden stones is calcium oxide produced in coastal Fijian villages by heating sun dried coral in earth oven, 'lovo'.

Calcium oxide was also used to form a paste to straighten/ stiffen wavy hair.

Green copper carbonate crystals will decompose to form a black solid, copper oxide (CuO). White lead carbonate is decomposed to form the yellow solid, lead oxide (PbO).

The nitrates of copper, lead and zinc will changed to liquid by gentle heating. The salt had dissolved in its water of crystallisation. When the liquid is strongly heated, a brown gas is given off and a solid deposit is also formed. The other gaseous product is oxygen.

Example

Copper nitrate → Copper oxide + Nitrogen dioxide + Oxygen

$2Cu(NO_3)_{2(s)} \rightarrow 2CuO_{(s)} + 4NO_{2\,(g)} + O_{2\,(g)}$

Blue copper nitrate crystals decomposes to form a black powdered (copper oxide), zinc nitrate crystals will decompose to form white powder (zinc oxide) and lead nitrate will decompose to form yellow lead oxide.

Exercise 3.2.1

1. For each reaction below:

    i. Write a balanced equation.
    ii. Classify the type of reaction.

    a. Burning of sulphur
    b. Burning of magnesium
    c. Formation of ammonia from nitrogen gas and hydrogen gas

2. Balance the equations given below:

    i. $Al + O_2 \rightarrow Al_2O_3$
    ii. $SO_2 + O_2 \rightarrow SO_3$
    iii. $KMnO_4 \rightarrow K_2MnO_4 + MnO_2 + O_2$

4. Neutralisation

In a neutralisation reaction, *acids react with bases to form salt and water*. The hydrogen ion of the acid will react with the hydroxide ion of the base to form water which is neutral. Substances which are acidic in nature include mineral acids, citrus juice, vinegar, gastric juice and non-metal oxides. Common bases are oxides and hydroxides of metals, ammonia and carbonates.

The reaction is also known as the **acid-base reaction**.

Example 1

Sodium hydroxide + Dilute hydrochloric acid → Sodium chloride + Water

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

$$OH^-_{(aq)} + H^+_{aq)} \rightarrow H_2O_{(l)}$$

Note: Carbonates react with dilute acids to form a salt, water and carbon dioxide.

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

Did You Know?

Rust on iron surfaces is removed by soaking the metal in lemon juice or vinegar.

5. Precipitation

It is the formation of an *insoluble* salt from the mixture of two different *clear* solutions. The *insoluble salt* formed is the *precipitate* (ppt). In earlier science classes we learnt about the formation of scum by mixing soap solution and hard water. Scum is a precipitate. Some precipitate may settle at the bottom of the test tube; others will form a suspension.

The illustration below shows the formation of the precipitate copper hydroxide from mixing copper sulphate solution with caustic soda, $NaOH_{(aq)}$

 +  → 

$Cu(OH)_{2(ppt)}$

Reaction Equation

$CuSO_{4\,(aq)} + 2NaOH_{(aq)} \rightarrow Cu\,(OH)_{2(ppt)} + Na_2SO_{4\,(aq)}$

Writing Ionic Equations

An ionic equation is used to emphasise the species involved in a reaction.

The following steps are used to determine the ionic equation from a balanced chemical equation.

1. Write the balanced equation for the reaction.

Example

$$Mg(NO_3)_{2(aq)} + Na_2CO_{3\,(aq)} \rightarrow MgCO_{3(s)} + 2NaNO_{3\,(aq)}$$

2. Write all the ions present in aqueous solutions in the reaction separately. Ions contained in solids, liquids and gases will not be separated.

$$Mg^{2+}{}_{(aq)} + 2NO_3^-{}_{(aq)} + 2Na^+{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow MgCO_{3(s)} + 2Na^+{}_{(aq)} + 2NO_3^-{}_{(aq)}$$

3. Cancel all the ions that appear on both sides of the equation, *spectator ions*. These ions are not involved in the reaction being emphasised.

$$Mg^{2+}{}_{aq)} + \cancel{2NO_3^-}{}_{(aq)} + \cancel{2Na^+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow MgCO_{3(s)} + \cancel{2Na^+}{}_{(aq)} + \cancel{2NO_3^-}{}_{(aq)}$$

4. Write the balanced ionic equation from the remaining species in Step 3.

$$Mg^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow MgCO_{3(s)}$$

The balanced ionic equations shows that the ions involved in forming the precipitate are magnesium, Mg^{2+} and carbonate ions, CO_3^{2-}.

Exercise 3.2.2

1. For each reaction below:
   i. Write a balanced chemical equation.
   ii. Classify the type of reaction and give a reason for your choice.

      a. Precipitation of silver chloride by reacting barium chloride with silver nitrate.
      b. Formation of solid barium sulphate by reacting barium chloride with dilute sulphuric acid.
      c. Release of carbon dioxide by reacting sodium carbonate with dilute sulphuric acid.
      d. Copper metal formed as zinc granules is placed into a test tube containing copper sulphate solution.
2. Balance the equations given below:
   1. $MgCO_3 + HNO_3 \rightarrow Mg(NO_3)_2 + CO_2 + H_2O$
   2. $MgO + CH_3COOH \rightarrow Mg(CH_3COO)_2 + H_2O$
   3. $NH_4OH + Al_2(SO_4)_3 \rightarrow Al(OH)_3 + (NH_4)_2SO_4$
   4. $K_2O + H_2SO_4 \rightarrow K_2SO_4 + H_2O$

6. Double Displacement

When two different salt solutions react forming a clear solution. The resultant salts formed are both soluble in water. It is termed double displacement as the anions are exchanged between the two cations.

Example

Barium chloride + Sodium nitrate → Barium nitrate + Sodium sulphate

$BaCl_{2(aq)}$ + $NaNO_{3(aq)}$ → $Ba\ (NO_3)_{2(aq)}$ + $Na_2SO_{4(aq)}$

7. Oxidation-Reduction

Oxidation is the gain of oxygen. For example, combustion and corrosion reactions are oxidation reactions. Reduction is the loss of oxygen. Oxidation and reduction reactions occur simultaneously. As a substance is reduced, the other reactant will be oxidised.

Example 1

In the extraction of metal from metal oxides using carbon, the metal oxide is reduced to the metal and carbon is oxidised to carbon dioxide.

$C_{(s)}$ + $2CuO_{(s)}$ → $2Cu_{(s)}$ + $CO_{2(g)}$

Example 2

Iron metal is produced in the Blast Furnace by the reduction of iron (III) oxide by carbon monoxide. The carbon monoxide is oxidised to carbon dioxide.

$Fe_2O_{3(s)}$ + $3CO_{(g)}$ → $2Fe_{(s)}$ + $3CO_{2(g)}$

The collective term for oxidation-reduction reaction is **redox**. Some of these changes may not involve the movement of oxygen, so the two reactions are also defined by the transfer of hydrogen and electrons.

Oxidation is the loss of hydrogen or electrons. Reduction is the gain of hydrogen or electrons.

Example 1: Oxidation reaction

Ionisation of metal atoms to positive metal ions (cations) is an **oxidation** reaction as electrons are *lost* from its valence shell.

$Na \rightarrow Na^+ + e^-$

$Mg \rightarrow Mg^{2+} + 2e^-$

$Al \rightarrow Al^{3+} + 3e^-$

Note: Electron(s) lost will appear on the right side of the equation.

Example 2: Reduction reaction

Non-metals *gaining* electron(s) to form stable negative ions (anions) are **reduction** reactions.

$Cl_2 + 2e^- \rightarrow 2Cl^-$

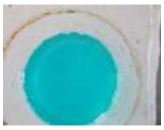
$O_2 + 4e^- \rightarrow 2O^{2-}$

Note: Electron(s) gained will appear on the left side of the equation.

Example 3

Displacement reactions of metals are redox reactions. The more active metal is oxidised to its ions and the less active metals is reduced from its ionic form (in aqueous solution) to the metal.

The figure below shows iron filings placed in a beaker of copper sulphate solution reduces the copper ions (light blue solution) to copper metal (reddish brown)



Copper sulphate solution

Iron filings

Mixture of Iron and copper sulphate solution

Reddish brown deposits formed

Oxidation Reaction: $Fe \rightarrow Fe^{2+} + 2e^-$

Reduction Reaction: $Cu^{2+} + 2e^- \rightarrow Cu$

Example 4

Displacement reactions of active metals with dilute acid are redox reactions. The more active metal is oxidised to its ions and the hydrogen ions in acid is reduced to hydrogen gas (bubbles evolved). (*Refer to Chapter 2*)

Oxidation Reaction: $Mg \rightarrow Mg^{2+} + 2e^-$

Reduction Reaction: $2H^+ + 2e \rightarrow H_2$

Exercise 3.2.3

1. Balance the equations given below and identify the type of reaction shown by each equation:

   i. $Zn + AgNO_3 \rightarrow Zn(NO_3)_2 + Ag$
   ii. $FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$
   iii. $H_2O \rightarrow H_2 + O_2$
   iv. $Zn + HCl \rightarrow ZnCl_2 + H_2$
   v. $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$
   vi. $HBr + NaOH \rightarrow NaBr + H_2O$

2. Identify the following equations as either oxidation or reduction.

   i. $2Cl^- \rightarrow Cl_2 + 2e^-$
   ii. $I_2 + 2e^- \rightarrow 2I^-$
   iii. $Mg \rightarrow Mg^{2+} + 2e^-$
   iv. $Zn^{2+} + 2e^- \rightarrow Zn$

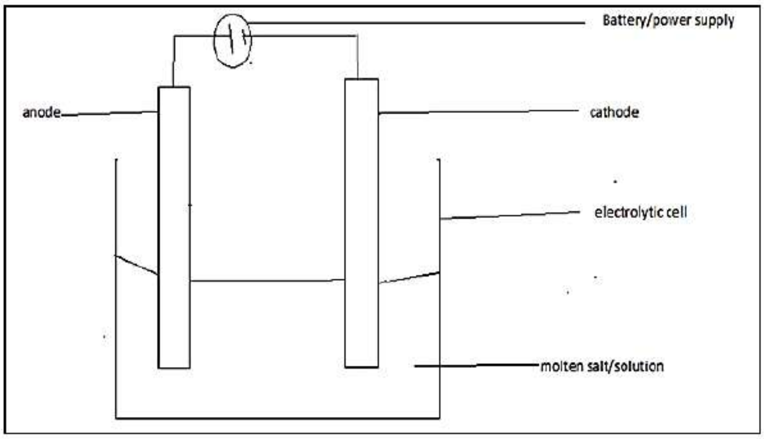
3. The reaction of lead oxide with carbon forms lead metal and carbon dioxide.

   i. Write a balanced chemical equation to represent the reaction above.
   ii. From the equation, determine which reactant is oxidised and which is reduced.
   iii. Explain why the reaction between lead oxide and carbon is called a redox reaction.

Electrolysis

Redox is commercially used in a process called electrolysis. An English chemist, Humphry Davy in the 1800s found out that electricity could be used to make reactions. Electrolysis is the decomposition of an electrolyte by passing an electric current through it. An electrolyte is a molten salt or solution that conducts electricity.

Electrolysis is carried out in an electrolytic cell, as in Figure below.



The components of an electrolytic cell are:

1. **Electrolyte** – molten or solutions of ionic compounds. The mobile/free ions are the carriers of electric current. Examples include: $NaCl_{(l)}$, $NaCl_{(aq)}$, $H_2O_{(l)}$, $MgCl_{2(aq)}$, $CuSO_{4(aq)}$.
2. **Batteries/Direct Current, DC power supply** – source of current, creates or discharge ions in the electrolyte. The electrode potential should be large enough to drive the reactions.
3. **Electrodes** – connects batteries/DC power supply to electrolyte. The two types are **anode** (positively charged) and **cathode** (negatively charged). Electrodes are usually inert or unreactive and a conductor of electricity. A common electrode is carbon (graphite) as it is inert and a conductor. Less reactive metals such as copper, iron and zinc, are used in electroplating.

   The electrolytic cell is a complete circuit. Electrons move from the anode to the cathode. As electrons moves away from the anode, the anode becomes positive and as electrons are deposited on the cathode, the cathode becomes negative.

Electrolysis of a salt solution

The degree of electrolysis of an aqueous solution depends on the concentration, ions present and nature of the electrodes.

Water could be oxidised or reduced. Its presence complicates the electrolysis of aqueous solutions.

Anode Reaction (Oxidation)

The anions that are easily oxidised instead of water include chloride ion (Cl^-), bromide ion (Br^-) and iodide ion (I^-). Polyatomic ions are not discharged. An example of such ions include sulphate ion (SO_4^{2-}). Instead, oxygen gas is evolved due to oxidation of water.

Cathode Reaction (Reduction)

A solution that contains cation of an element below aluminium in the Activity Series will be reduced instead of water. Cations of salt solutions of very active elements will not be reduced as water is reduced.

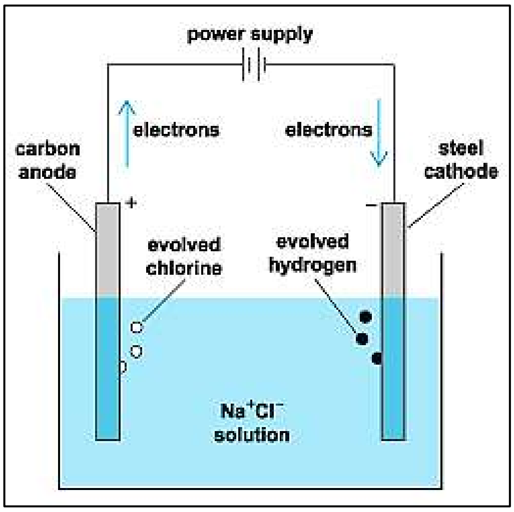
The figure below shows the electrolysis of brine, concentrated $NaCl_{(aq)}$; it is an important industrial application as it produces much needed chlorine. The salt solution contains the electrolyte, sodium ions, chloride ions and water.

Water is reduced at the cathode instead of sodium ions as it was **easier** to reduce; at the anode, chloride ion was easier to oxidise than water.

Anode Reaction

$2Cl^-_{(aq)} \rightarrow Cl_{2\,(g)} + 2e^-$

Greenish yellow gas



Source: http://look4chemistry.blogspot.com

Cathode Reaction

$2H_2O_{(l)} + 2e^- \rightarrow H_{2(g)} + 2OH^-_{(aq)}$

Note: *Hydroxide ion formed (OH^-) is picked by sodium ion in the solution producing sodium hydroxide, NaOH)*

Electrolysis of a molten salt

Electrolysis of a molten salt is less complicated than electrolysis of ionic solutions as it involves only two species. In solutions, the presence of water may increases the number of species that may be oxidised or reduced.

The power supply is connected to the electrodes. The positive electrode (anode) attracts the negatively charged ions (anions) in the electrolyte and the negative electrode (cathode) attracts the positively charged ions (cations) in the electrolyte.

At the anode, the anion will lose electron(s) and oxidised to the respective element. At the cathode, the cation will gain electron(s) and reduced to the respective metal.

Anode Reaction

$B^- \rightarrow B + e^-$

Cathode Reaction

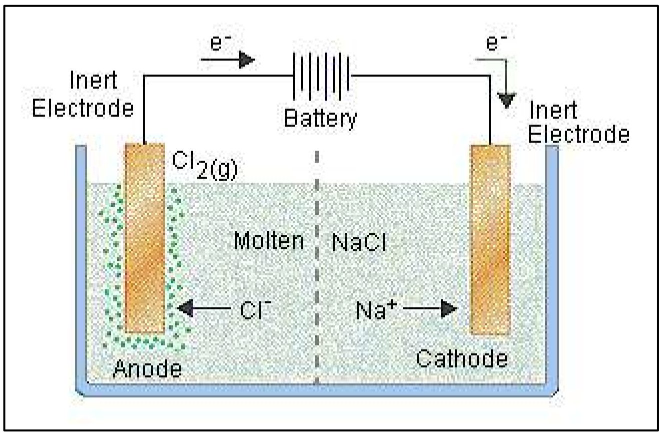
$A^+ + e^- \rightarrow A$

The net effect is the decomposition of the ionic compound forming two separate elements.

Example **Electrolysis of molten NaCl**

Anode Reaction

$2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$



Source: http://chemed.chem.purdue.edu

Cathode Reaction

$Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$

Sodium ions (Na^+) gains electron and is reduced to sodium metal at the cathode. Chloride ions moves to the anode and is oxidised by losing electrons to form chlorine gas. The anode is separated from the cathode by a diaphragm. It prevents the greenish yellow chlorine gas produced at the anode from reacting with the liquid silver sodium metal formed at the cathode. This is because the reaction is explosive.

Other Examples

1. **Electrolysis of molten Lead Bromide Using Carbon Electrodes**

Lead bromide melts at 373°C to form molten lead bromide ($PbBr_2$), which is made up of mobile lead ions and bromide ions. During electrolysis, lead ions are attracted to the cathode and are reduced to silver lead metal and bromide ions are attracted to the anode forming bromine, a red coloured gas.

Cathode Reaction

$Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(l)}$

Silver liquid

Anode Reaction

$2Br^-_{(aq)} \rightarrow Br_{2(g)} + 2e^-$

red gas

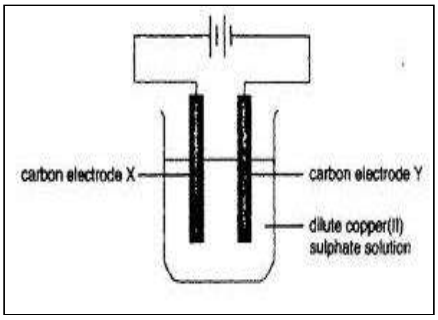
2. **Electrolysis of Copper Sulphate solution Using Carbon**

Copper sulphate solution contains the electrolytes; copper ions, sulphate ions and water.

Copper ions are attracted to the cathode and are reduced to reddish brown copper metal. Sulphate ions are attracted to the anode. However, water is oxidised as sulphate ions cannot be oxidised. It releases oxygen gas, so a colourless gas is formed at the anode.

Anode Reaction

$2H_2O_{(l)} \rightarrow 4H^+_{(aq)} + O_{2(g)} + 4e^-$



Source: www.gcestudybuddy.com

Cathode Reaction

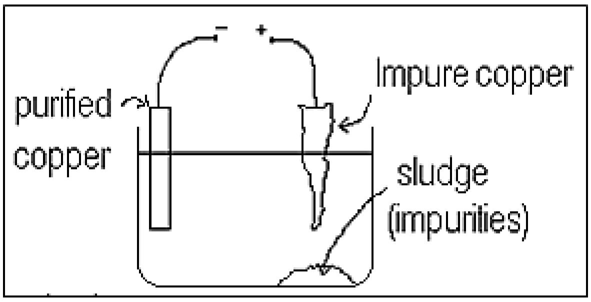
$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$

3. Electrolysis of Copper sulphate solution using copper metal electrodes

This set up is used in the *industrial production of copper metal.* In the figure below, impure copper is the anode and pure copper is the cathode.

Cathode Reaction

$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$



Source:http://faculty.chem.queensu.ca/

Anode Reaction

$Cu \rightarrow Cu^{2+}_{(aq)} + 2e^-$

As the electricity is switched on, copper ions are attracted to the cathode and sulphate ions are attracted to the anode. At the cathode, copper ions are reduced to copper metal (reddish brown) deposited on the surface of the pure copper electrode. At the anode, copper is oxidised to copper ions and goes back to the copper sulphate solution. These ions replace the copper ions reduced at the cathode.

Did you know?

Impure gold present in the copper anode forms a sludge/deposit below the anode. This process is called electroplating. It is used to either protect a metal surface from corroding by coating it with a more active or less active metals or to cover a cheaper metal with an expensive metal. E.g. Tin coated iron cans, galvanised iron, gold/silver coated iron chain. The metal to be plated or coated is the anode; the metal to be used as coat is the cathode.

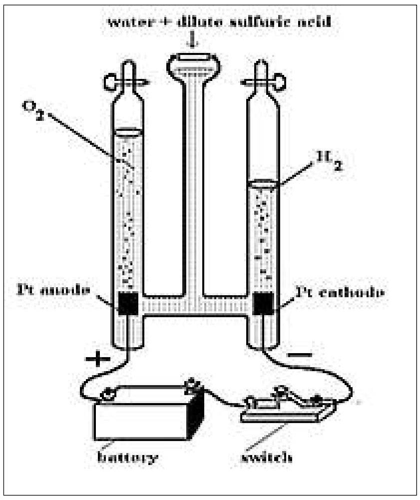
4. Electrolysis of water using carbon electrodes

Hydrogen gas is used as rocket fuel as it burns explosively (self-combust) in oxygen releasing a lot of energy. The product formed is water. The *Energy Industries* around are researching how it can be used as fuel for vehicles, especially buses as its product, water is environment friendly. The risk is the explosive nature of the reaction.

Electrolysis of water is used industrially to produce hydrogen gas. Water reduced at the cathode produces hydrogen gas and water oxidised at the anode forms hydrogen ions.

Anode Reaction

$2H_2O \rightarrow 4H^+_{(aq)} + O_{2\,(g)} + 4e$



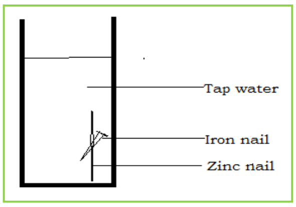
Source: www.uq.edu.au

Cathode Reaction

$2H_2O + 2e^- \rightarrow H_{2\,(g)} + 2OH^-$

Exercise 3.2.4

1. Why is graphite used as an electrode?
2. Why does an open can rust faster than an unopened can?
3. Is water an electrolyte? Give a reason for your answer.
4. Suggest why the anode of an electrolytic cell is positive.
5. Suggest why the cathode of an electrolytic cell is negative.
6. How does electroplating prevent corrosion?
7. Study the set up given below. The iron nail is touching the zinc nail as they are tied together.



i) Will the iron corrode? Give a reason for your answer.
ii) The zinc nail was removed. State an observation that you will make after a few days.

8. Suppose you are given the following materials:

A silver table spoon, copper electrodes, electrical wires, 1L beaker, 500 mL 1.0 moles/litre copper sulphate solution, AC/DC Transformer

i. Draw a diagram of the electrolytic cell you would construct using the given materials in order to plate the silver spoon with copper.
ii. Why was the silver spoon placed as the cathode?
iii. Will the copper sulphate solution decolourise? Give a reason for your answer.

9. Draw a concept map on the sub-strand “Types of Reactions”.