

# Stoichiometry – chemical calculations

## Additional questions

Use the data in the table below to answer the questions which follow.

Element	$A_r$
H	1
C	12
N	14
O	16
Na	23
Mg	24
Si	28
S	32
Cl	35.5
Fe	56

- Calculate the mass of:
  - 1 mole of:
    - chlorine molecules
    - iron(III) oxide.
  - 0.5 mole of:
    - magnesium nitrate
    - ammonia.
- Calculate the volume occupied, at rtp, by the following gases. (One mole of any gas occupies a volume of  $24 \text{ dm}^3$  at rtp.)
  - 12.5 moles of sulfur dioxide gas.
  - 0.15 mole of nitrogen gas.
- Calculate the number of moles of gas present in the following:
  - $36 \text{ cm}^3$  of sulfur dioxide
  - $144 \text{ dm}^3$  of hydrogen sulfide.
- Use the following experimental information to determine the empirical formula of an oxide of silicon.

Mass of crucible	18.20 g
Mass of crucible + silicon	18.48 g
Mass of crucible + oxide of silicon	18.80 g
- Calculate the empirical formula of an organic liquid containing 26.67% of carbon and 2.22% of hydrogen, with the rest being oxygen.
  - The  $M_r$  of the liquid is 90. What is its molecular formula?
- Iron is extracted from its ore, haematite, in the blast furnace. The main extraction reaction is:
$$\text{Fe}_2\text{O}_3(\text{s}) + 3\text{CO}(\text{g}) \rightarrow 2\text{Fe}(\text{s}) + 3\text{CO}_2(\text{g})$$
  - Name the reducing agent in this process.
  - Name the oxide of iron shown in the equation.
  - Explain why this is a **redox** reaction.
  - Calculate the mass of iron which will be produced from 640 tonnes of haematite.
- Consider the following information about the newly discovered element, vulcium, whose symbol is Vu. 'Vulcium is a solid at room temperature. It is easily cut by a penknife to reveal a shiny surface which tarnishes quite rapidly. It reacts violently with water, liberating a flammable gas and forms a solution with a pH of 13. When vulcium reacts with chlorine, it forms a white crystalline solid containing 29.5% chlorine.'  
( $A_r$ : Vu = 85)
  - Calculate the empirical formula of vulcium chloride.
  - To which group of the Periodic Table (p. 136) should vulcium be assigned?
  - Write a word and balanced chemical equation for the reaction between vulcium and chlorine.
  - What other information in the description supports the assignment of group you have given to vulcium?
  - What type of bonding is present in vulcium chloride?
  - Write a word and balanced chemical equation for the reaction between vulcium and water.
  - Write the formulae for:
    - vulcium sulfate
    - vulcium carbonate
    - vulcium hydroxide.Look at the Periodic Table (p. 294) to find out the real name of vulcium.
- 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure. The hydrogen gas given off was collected.
  - Write a word and balanced symbol equation for the reaction taking place.
  - Draw a diagram of an apparatus which could be used to carry out this experiment and collect the hydrogen gas.
  - How many moles of magnesium were used?
  - Using the equation you have written in your answer to a, calculate the number of moles of hydrogen and hence the volume of this gas produced.
  - Calculate the volume of a solution containing  $0.1 \text{ mol dm}^{-3}$  hydrochloric acid which would be needed to react exactly with 0.048 g of magnesium.

**Electrolysis of lead(II) bromide****Electrolysis of aluminium oxide**

Anodising

**Electrolysis of aqueous solutions**

Electrolysis of dilute sulfuric acid

The chlor-alkali industry

Tests for aqueous cations

**Electrolysis of concentrated hydrochloric acid****Electrolysis of copper(II) sulfate solution**

Purification of copper

**Electrolysis guidelines****Electroplating**

Plating plastics

**Checklist****Additional questions**

What do all the items in the photographs shown in Figure 5.1 have in common? They all involve electricity through a process known as **electrolysis**. Electrolysis is the breakdown of an ionic compound, molten or in solution, by the passage of electricity through it. The substance which is decomposed is called the **electrolyte** (Figure 5.2). An electrolyte is a substance that conducts electricity when in the molten state or in solution.



**a** This watch has a thin coating of gold over steel; the thin coating is produced by electrolysis.



**b** This picture frame has been silver plated using an electroplating process involving electrolysis.



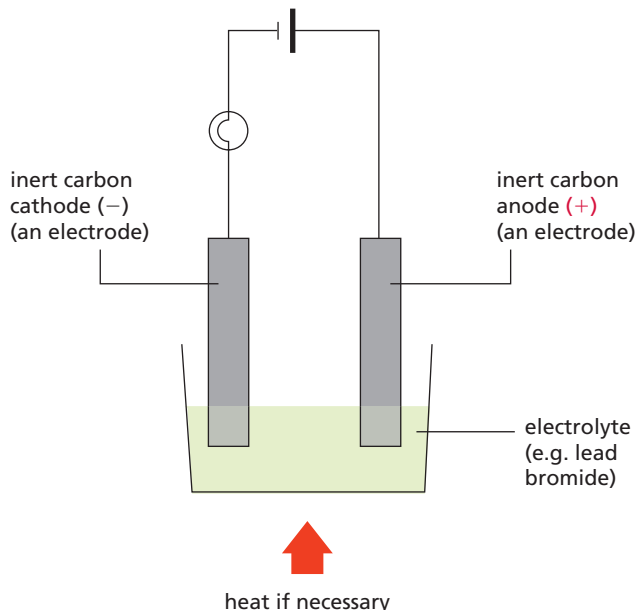
**c** Aluminium is produced by electrolysis.

**Figure 5.1**

The electricity is carried through the electrolyte by **ions**. In the molten state and in solution the ions are free to move to the appropriate electrodes due to weakened forces of attraction between them.

- Substances that do not conduct electricity when in the molten state or in solution are called **non-electrolytes**.
- Substances that conduct electricity **to a small extent** in the molten state or in solution are called **weak electrolytes**.

The electric current enters and leaves the electrolyte through **electrodes**, which are usually made of unreactive metals such as platinum or of the non-metal carbon (graphite). These are said to be **inert** electrodes because they do not react with the products of electrolysis. The names given to the two electrodes are **cathode**, the negative electrode which attracts **cations** (positively charged ions), and **anode**, the positive electrode which attracts **anions** (negatively charged ions).



**Figure 5.2** The important terms used in electrolysis.

The transfer of charge during electrolysis is by:

- the movement of electrons in the metallic or graphite electrodes
- the removal or addition of electrons from the external circuit at the electrodes
- the movement of ions in the electrolyte.

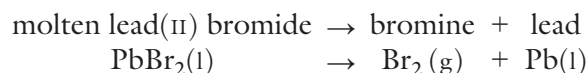
Note that the conduction which takes place in the electrodes is due to the movement of delocalised electrons (pp. 51 and 55) whereas in the electrolyte, as stated earlier, the charge carriers are ions.

Electrolysis is very important in industry. To help you to understand what is happening in the process shown in the photographs, we will first consider the electrolysis of lead(II) bromide.

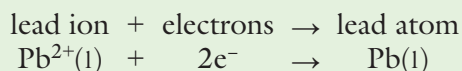
## ● Electrolysis of lead(II) bromide

Consider Figure 5.3, which shows solid lead(II) bromide ( $\text{PbBr}_2$ ) in a crucible with two carbon electrodes in contact with it. When the electrodes are first connected, the bulb does not light, because the solid compound does not allow electricity to pass through it. However, when the compound is heated until it is molten, the bulb does light. The lead(II) bromide is now behaving as an electrolyte. When this happens an orange-red vapour is seen at the anode and lead metal is produced at the cathode.

The break-up (decomposition) of lead(II) bromide into its constituent elements by the passage of an electric current is called **electrolysis**.



For lead metal to be formed, or deposited, at the cathode, the lead ions must be attracted to and move towards the electrode (Figure 5.4). To produce lead metal atoms these lead ions must each collect two electrons at the cathode:



As you saw in Chapter 3 (p. 39), this process of gaining electrons is called **reduction**.

To form bromine molecules each bromide ion must first of all move to the anode and lose its extra negative charge at the anode and so form a neutral bromine atom:

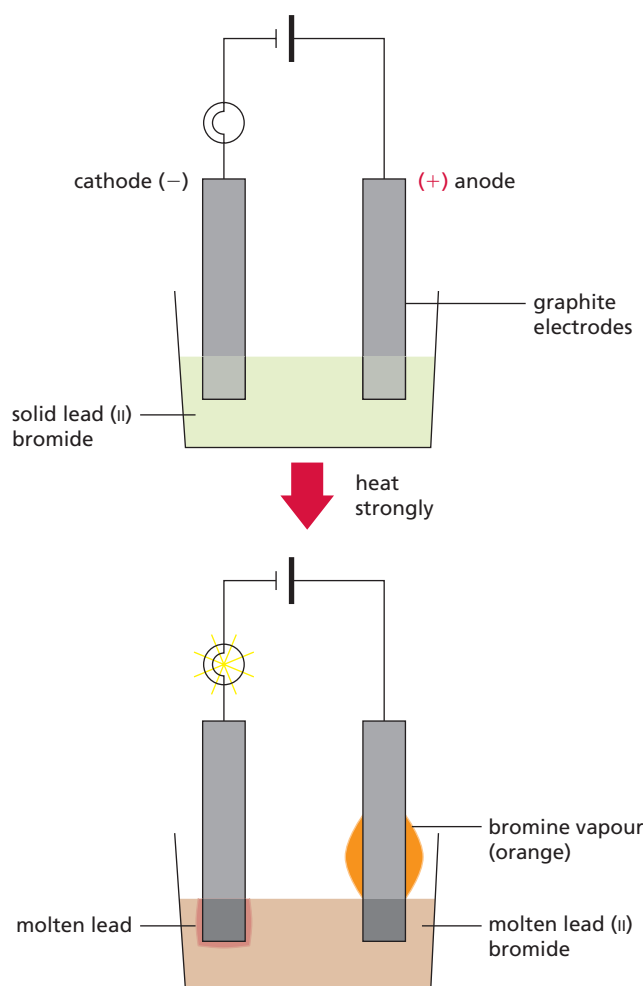


Figure 5.3 The electrolysis of molten lead(II) bromide.

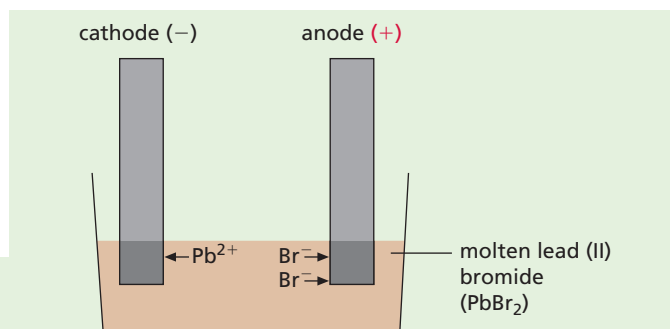
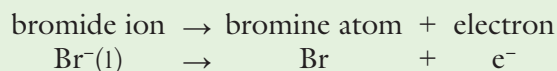
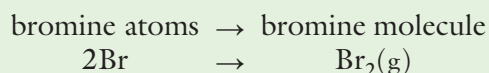


Figure 5.4 The lead ions ( $\text{Pb}^{2+}$ ) are attracted to the cathode and the bromide ions ( $\text{Br}^-$ ) are attracted to the anode.



Two bromine atoms then combine to form a bromine molecule:



As you saw in Chapter 3 (p. 39), this process of losing electrons is called **oxidation**.



**Question**

- 1 Predict the products of the electrolysis of:
- a potassium chloride
  - b lead oxide.

## ● Electrolysis of aluminium oxide

Aluminium is the most abundant metallic element in the Earth's crust. It was discovered in 1825 by Hans Christian Oersted in Copenhagen, Denmark, but was first isolated by Friedrich Wöhler in 1827. It makes up 8% of the crust and is found in the minerals bauxite (Figure 5.5), cryolite and mica, as well as in clay.

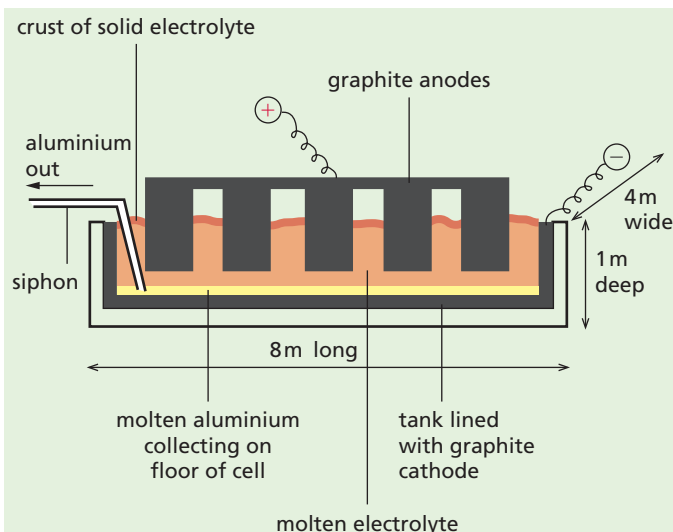
In the nineteenth century Napoleon III owned a very precious dinner service. It was said to be made of a metal more precious than gold. That metal was aluminium. The reason it was precious was that it was very rarely found as the pure metal. Aluminium is a reactive metal and as such was very difficult to extract from its ore. Reactive metals hold on tightly to the element(s) they have combined with and many are extracted from their ores by electrolysis.

Today we use aluminium in very large quantities. The annual production in the world is 19.5 million tonnes. The commercial extraction of aluminium has been made possible by two scientists, working independently of each other, who discovered a method using electrolysis. The two scientists were Charles Martin Hall (USA), who discovered the process in 1886, and the French chemist Paul Héroult, who discovered the process independently in the same year. The process they developed, often called the Hall–Héroult process, involves the electrolysis of aluminium oxide (alumina). The process involves the following stages.

- Bauxite, an impure form of aluminium oxide, is first treated with sodium hydroxide to obtain pure aluminium oxide, removing impurities such as iron(III) oxide and sand. This improves the conductivity of the molten aluminium oxide.
- The purified aluminium oxide is then dissolved in molten cryolite ( $\text{Na}_3\text{AlF}_6$ ). Cryolite, a mineral found naturally in Greenland, is used to reduce the working temperature of the Hall–Héroult cell from  $2017^\circ\text{C}$  (the melting point of pure aluminium oxide) to between  $800$  and  $1000^\circ\text{C}$ . Therefore, the cryolite provides a considerable



Figure 5.5 Bauxite mining.



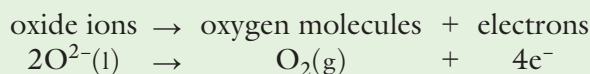
**Figure 5.6** The Hall-Héroult cell is used in industry to extract aluminium by electrolysis.

saving in the energy requirements of the process. In recent years it has become necessary to manufacture the cryolite.

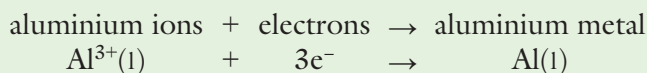
- The molten mixture is then electrolysed in a cell similar to that shown in Figure 5.6.

The anodes of this process are blocks of graphite which are lowered into the molten mixture from above. The cathode is the graphite lining of the steel vessel containing the cell.

Aluminium oxide is an ionic compound. When it is melted the ions become mobile, as the strong electrostatic forces of attraction between them are broken by the input of heat energy. During electrolysis the negatively charged oxide ions are attracted to the anode (the positive electrode), where they lose electrons (oxidation).

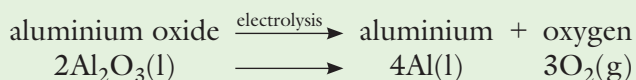


The positive aluminium ions are attracted to the cathode (the negative electrode). They gain electrons to form molten aluminium metal (reduction).

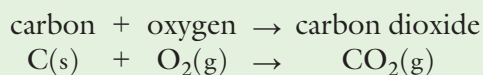


A handy way of remembering it is **OIL RIG** (Oxidation Is Loss, Reduction Is Gain of electrons).

The overall reaction which takes place in the cell is:



The molten aluminium collects at the bottom of the cell and it is siphoned out at regular intervals. No problems arise with other metals being deposited, since the cryolite is largely 'unaffected' by the flow of electricity. Problems do arise, however, with the graphite anodes. At the working temperature of the cell, the oxygen liberated reacts with the graphite anodes, producing carbon dioxide.



The anodes burn away and have to be replaced on a regular basis.

The electrolysis of aluminium oxide is a continuous process in which vast amounts of electricity are used. Approximately 15 kWh of electricity are used to produce 1 kg of aluminium. In order to make the process an economic one, a cheap form of electricity is required. Hydroelectric power (HEP) is usually used for this process. The plant shown in Figure 5.7 uses an HEP scheme to provide some of the electrical energy required for this process. Further details about HEP are given in Chapter 6, p. 94.



**Figure 5.7** An aluminium smelting plant.



Using cheap electrical energy has allowed aluminium to be produced in such large quantities that it is the second most widely used metal after iron. It is used in the manufacture of electrical cables because of its low density, chemical inertness and good electrical conductivity. Owing to the first two of these properties, it is also used for making cars, bikes, cooking foil and food containers as well as in alloys (Chapter 10, p. 167) such as duralumin, which is used in the manufacture of aeroplane bodies (Figure 5.8). Worldwide production of aluminium now exceeds 40 million tonnes each year.



Figure 5.8 Aluminium is used in the manufacture of aeroplane bodies.

Environmental problems associated with the location of aluminium plants are concerned with:

- the effects of the extracted impurities, which form a red mud (Figure 5.9)
- the fine cryolite dust, which is emitted through very tall chimneys so as not to affect the surrounding area



Figure 5.9 Bauxite pollution of the Amazon in Brazil.

- the claimed link between environmental aluminium and a degenerative brain disorder called Alzheimer's disease – it is thought that aluminium is a major influence on the early onset of this disease. However, the evidence is still inconclusive.

### Questions

- 1 Produce a flow chart to summarise the processes involved in the extraction of aluminium metal from bauxite.
- 2 Explain why the mixture of gases formed at the anode contains oxygen, carbon dioxide and traces of fluorine.

### Anodising

This is a process in which the surface coating of oxide on aluminium ( $\text{Al}_2\text{O}_3$ ) is made thicker. In this process the aluminium object is made the anode in a cell in which the electrolyte is dilute sulfuric acid. During the electrolysis process, oxygen is produced at the anode and combines with the aluminium. The oxide layer on the surface of the aluminium therefore increases. Dyes can be mixed with the electrolyte and so the new thicker coating of oxide is colourful and also decorative (Figure 5.10).



**Figure 5.10** The oxide layer on the surface of these aluminium cups has been thickened, and dyes added to obtain the vibrant colours.

### Question

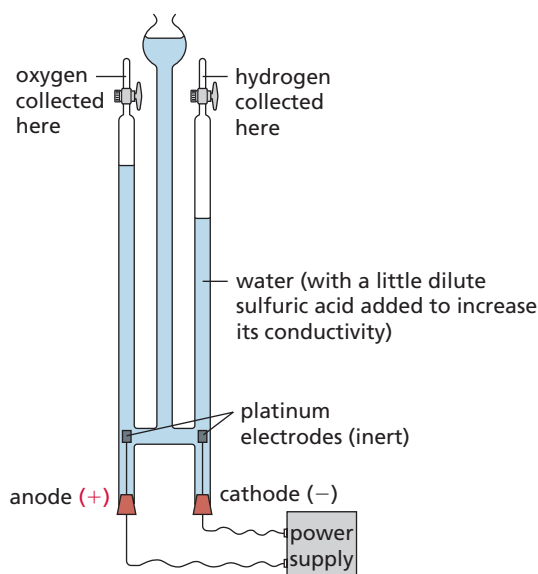
- 1 A student carries out the electrolysis of molten lead(II) iodide in a fume cupboard.
  - a Draw a diagram to show a suitable apparatus the student could use to carry out this experiment.
  - b Write anode and cathode reactions to represent the processes taking place during the electrolysis.
  - c Why does this experiment need to be carried out in a fume cupboard?
  - d Find uses for the anode product of this cell.

## ● Electrolysis of aqueous solutions

Other industrial processes involve the electrolysis of aqueous solutions. To help you to understand what is happening in these processes, we will first consider the electrolysis of dilute sulfuric acid.

### Electrolysis of dilute sulfuric acid

Pure water is a very poor conductor of electricity because there are so few ions in it. However, it can be made to decompose if an electric current is passed through it in a Hofmann voltameter, as in Figure 5.11.

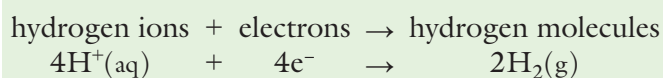


**Figure 5.11** A Hofmann voltameter used to electrolyse water.

To enable water to conduct electricity better, some dilute sulfuric acid (or sodium hydroxide solution) is added. When the power is turned on and an electric current flows through this solution, gases can be seen to be produced at the two electrodes and they are collected in the side arms of the apparatus. After about 20 minutes, roughly twice as much gas is produced at the cathode as at the anode.

The gas collected at the cathode burns with a squeaky pop, showing it to be hydrogen gas (Table 16.6, p. 263).

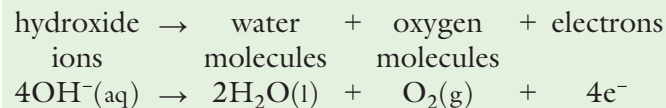
For hydrogen to be collected in this way, the positively charged hydrogen ions must have moved to the cathode.



If during this process the water molecules lose  $\text{H}^+(\text{aq})$ , then the remaining portion must be hydroxide ions,  $\text{OH}^-(\text{aq})$ . These ions are attracted to the anode.

The gas collected at the anode relights a glowing splint, showing it to be oxygen (Table 16.6, p. 263).

This gas is produced in the following way.



This experiment was first carried out by Sir Humphry Davy. It confirmed that the formula for water was  $\text{H}_2\text{O}$ .

It should be noted that in the electrolysis of dilute sulfuric acid, platinum (an inert electrode) may be replaced by carbon (graphite). The only difference to occur is that as well as oxygen being produced at the anode, a little carbon dioxide will also be formed.

## The chlor-alkali industry

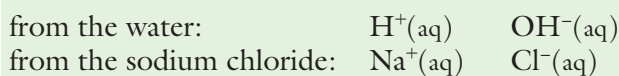
The electrolysis of saturated sodium chloride solution (brine) is the basis of a major industry. In countries where rock salt (sodium chloride) is found underground it is mined. In other countries it can be obtained by evaporation of sea water in large shallow lakes. Three very important substances are produced

in this electrolysis process – chlorine, sodium hydroxide and hydrogen. The electrolytic process is a very expensive one, requiring vast amounts of electricity. The process is economical only because all three products have a large number of uses (Figure 5.12).

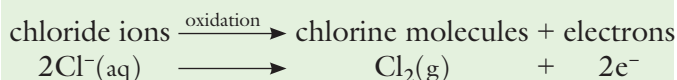
There are two well-established methods for electrolysis of brine, the **diaphragm cell** and the **mercury cell**. However, recent developments in electrolysis technology, by chemical engineers, have produced the **membrane cell** (Figure 5.13). This method is now preferred to the other two because it produces a purer product, it causes less pollution and it is cheaper to run.

The brine is first purified to remove calcium, strontium and magnesium compounds by a process of ion exchange (see Chapter 11, p. 189).

The membrane cell is used continuously, with fresh brine flowing into the cell as the process breaks up the brine. The cell has been designed to ensure that the products do not mix. The ions in this concentrated sodium chloride solution are:

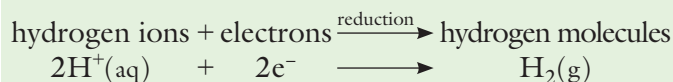


When the current flows, the chloride ions,  $\text{Cl}^-(\text{aq})$ , are attracted to the anode. Chlorine gas is produced by the electrode process.



This leaves a high concentration of sodium ions,  $\text{Na}^+(\text{aq})$ , around the anode.

The hydrogen ions,  $\text{H}^+(\text{aq})$ , are attracted to the cathode and hydrogen gas is produced.



This leaves a high concentration of hydroxide ions,  $\text{OH}^-(\text{aq})$ , around the cathode. The sodium ions,  $\text{Na}^+(\text{aq})$ , are drawn through the membrane, where they combine with the  $\text{OH}^-(\text{aq})$  to form sodium hydroxide,  $\text{NaOH}$ , solution. The annual production worldwide is now in excess of 60 million tonnes.



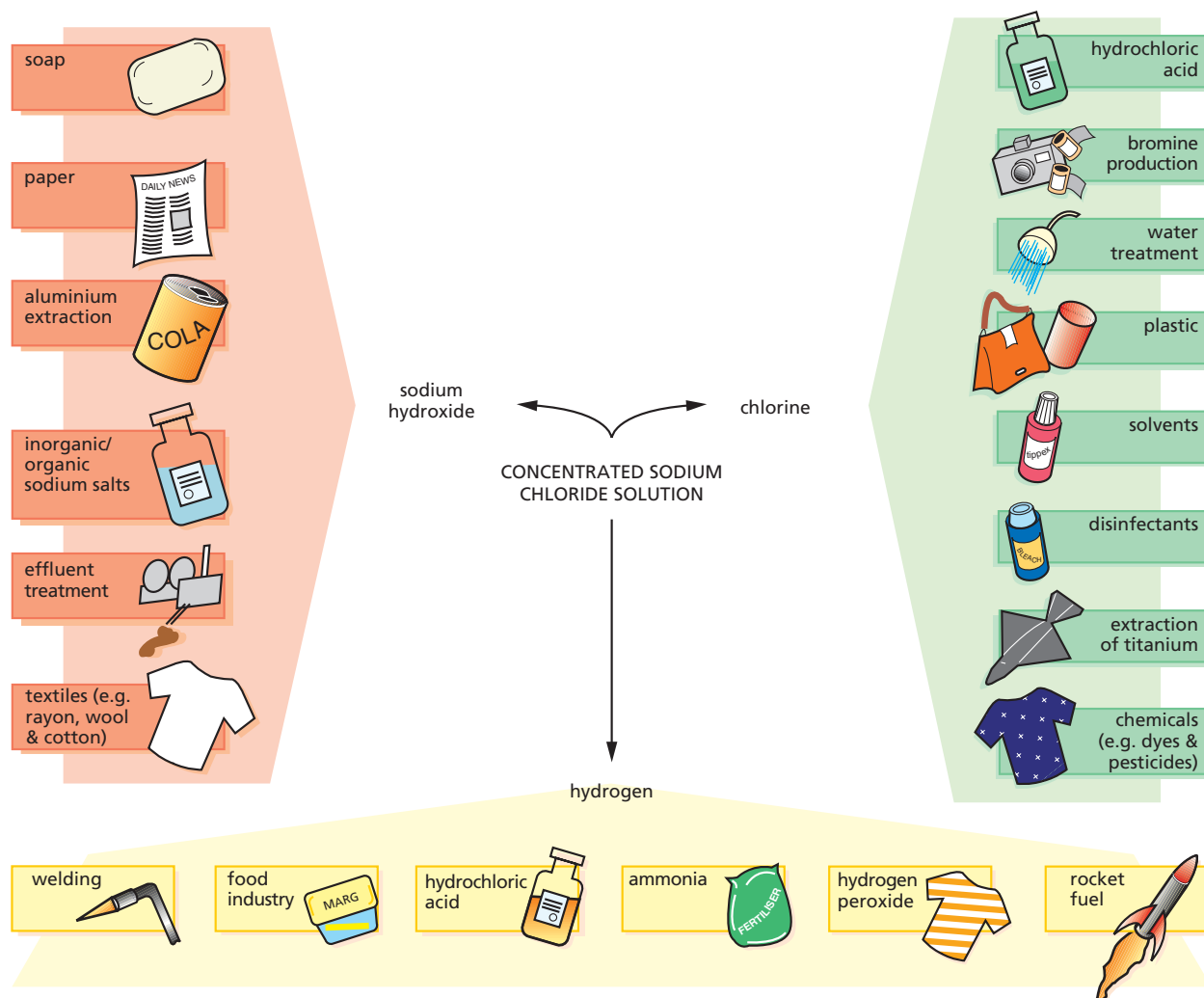


Figure 5.12 The chlor-alkali industry.

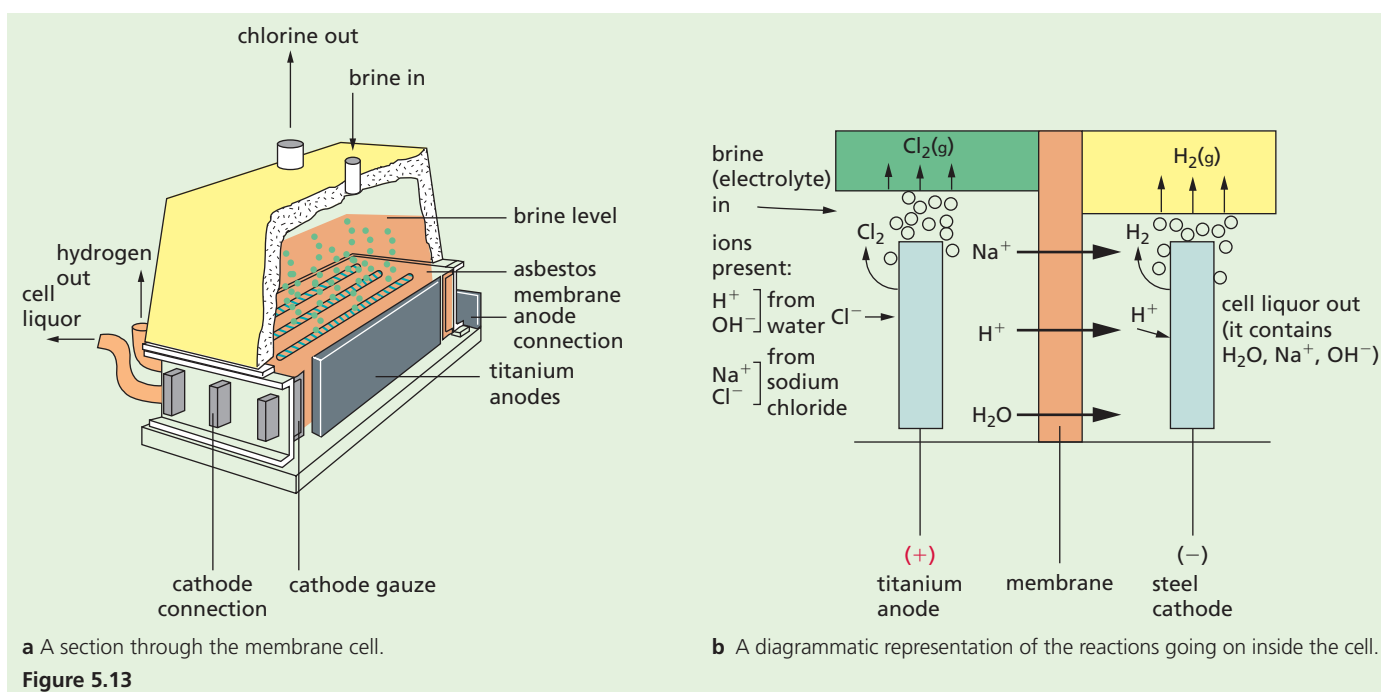


Figure 5.13

## Tests for aqueous cations

Aqueous sodium hydroxide can be used to identify salts of  $\text{Al}^{3+}$ ,  $\text{Ca}^{2+}$ ,  $\text{Cr}^{3+}$ ,  $\text{Cu}^{2+}$ ,  $\text{Fe}^{2+}$ ,  $\text{Fe}^{3+}$  and  $\text{Zn}^{2+}$ . The colour of the precipitate and its behaviour in excess sodium hydroxide solution will help identify the metal present. In the case of ammonium salts, ammonia gas is produced on warming.

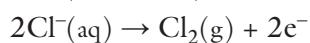
Table 5.1

Cation solution	Effect of aqueous sodium hydroxide
Aluminium ( $\text{Al}^{3+}$ )	White precipitate, soluble in excess giving a colourless solution
Ammonium ( $\text{NH}_4^+$ )	Ammonia produced on warming
Calcium ( $\text{Ca}^{2+}$ )	White precipitate, insoluble in excess
Chromium ( $\text{Cr}^{3+}$ )	Green precipitate, soluble in excess
Copper ( $\text{Cu}^{2+}$ )	Light blue precipitate, insoluble in excess
Iron(II) ( $\text{Fe}^{2+}$ )	Green precipitate, insoluble in excess
Iron(III) ( $\text{Fe}^{3+}$ )	Red-brown precipitate, insoluble in excess
Zinc ( $\text{Zn}^{2+}$ )	White precipitate, soluble in excess, giving a colourless solution

If the nitrate ion is present in solution, then upon addition of aqueous sodium hydroxide and a little aluminium foil, ammonia gas is produced on warming.

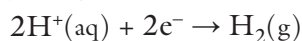
## Electrolysis of concentrated hydrochloric acid

When concentrated hydrochloric acid is electrolysed, two gases are produced. Both chloride ions and hydroxide ions would be attracted to the anode. The chloride ions are at a much higher concentration than the hydroxide ions and go on to produce chlorine gas by losing electrons (oxidation).



The chlorine gas can be tested by placing a piece of damp blue litmus paper near the anode. The paper will first turn red and then it will be bleached (turned white).

The only positively charged ion is the hydrogen ion. These are attracted to the cathode where they produce hydrogen gas by the gain of electrons (reduction).



Hydrogen gas, if collected from the cathode, will give a 'squeaky pop' when mixed with air and ignited.

## Questions

- Suggest a reason for only 'roughly' twice as much hydrogen gas being produced at the cathode as oxygen gas at the anode in the electrolysis of water.
- Account for the following observations which were made when concentrated sodium chloride solution, to which a little universal indicator had been added, was electrolysed in the laboratory in a Hofmann voltameter.
  - The universal indicator initially turns red in the region of the anode, but as the electrolysis proceeds it loses its colour.
  - The universal indicator turns blue in the region of the cathode.
- Why is it important to remove compounds of calcium, strontium and magnesium before brine is electrolysed?
- The uses of sodium hydroxide can be separated on a percentage basis as follows:
 

Neutralisation	5%
Paper manufacture	5%
Oil refining	5%
Soap/detergents	5%
Manufacture of rayon and acetate fibres	16%
Manufacture of chemicals	30%
Miscellaneous uses	34%

 Use a graph-plotting program to create a 'pie' chart of these data.

## Electrolysis of copper(II) sulfate solution

Copper(II) sulfate solution ( $\text{CuSO}_4(\text{aq})$ ) may be electrolysed using inert graphite electrodes in a cell similar to that shown in Figure 5.14. When the solution is electrolysed, oxygen gas and copper metal are formed at the anode and cathode respectively. Four ions are present in solution:

from the water:  $\text{H}^+(\text{aq})$   $\text{OH}^-(\text{aq})$   
 from the copper(II) sulfate:  $\text{Cu}^{2+}(\text{aq})$   $\text{SO}_4^{2-}(\text{aq})$

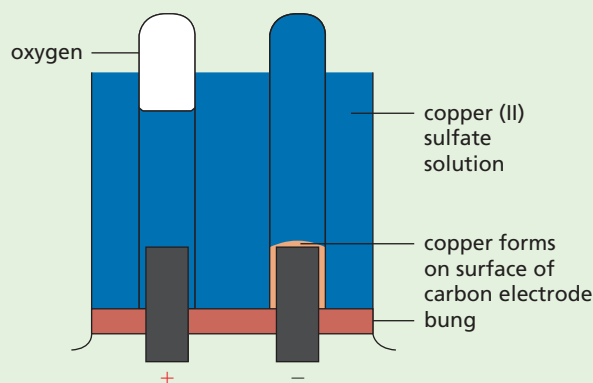
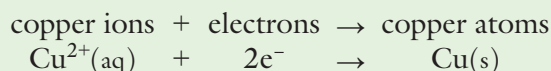
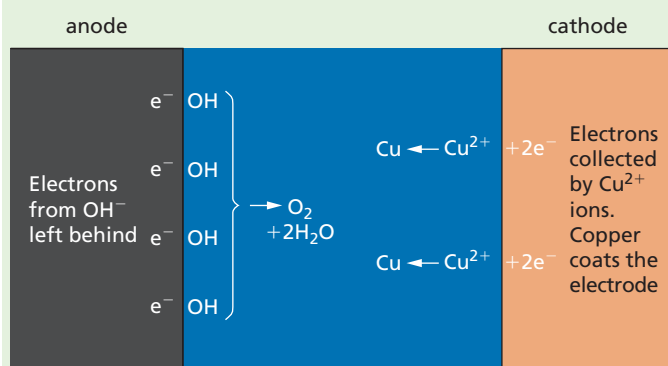
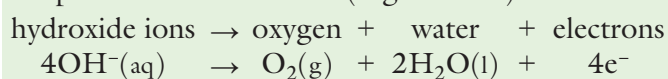


Figure 5.14 The electrolysis of copper(II) sulfate solution using inert electrodes.

$\text{H}^+(\text{aq})$  and  $\text{Cu}^{2+}(\text{aq})$  ions are both attracted to the cathode, the  $\text{Cu}^{2+}$  ions accepting electrons more readily than the  $\text{H}^+$  ions (preferential discharge). Copper metal is therefore deposited at the cathode (Figure 5.14).



$\text{OH}^-(\text{aq})$  and  $\text{SO}_4^{2-}(\text{aq})$  ions are both attracted to the anode. The  $\text{OH}^-$  ions release electrons more easily than the  $\text{SO}_4^{2-}$  ions, so oxygen gas and water are produced at the anode (Figure 5.15).

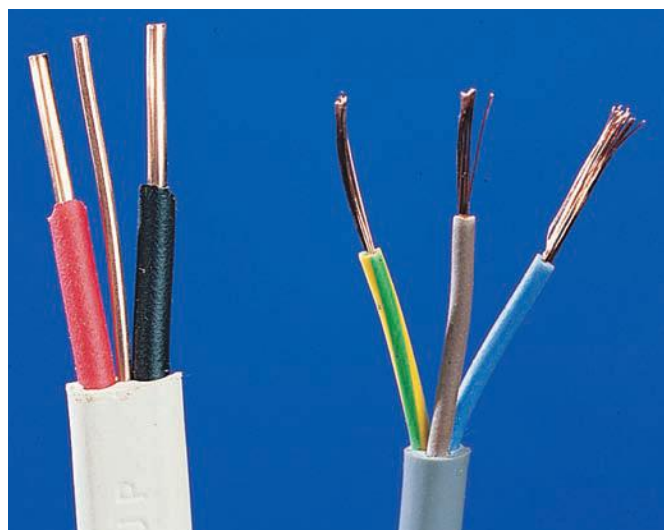


**Figure 5.15** Oxygen is given off at the anode and copper is deposited at the cathode.

## Purification of copper

Because copper is a very good conductor of electricity, it is used for electrical wiring and cables (Figure 5.16). Pure copper is also used in the manufacture of cooking utensils owing to its high thermal conductivity, a property of its metallic structure (Chapter 3, p. 55).

However, even small amounts of impurities cut down this conductivity quite noticeably whether in fine wires or larger cables. The metal must be 99.99% pure to be used in this way. To ensure this level of purity, the newly extracted copper has to be purified by electrolysis.



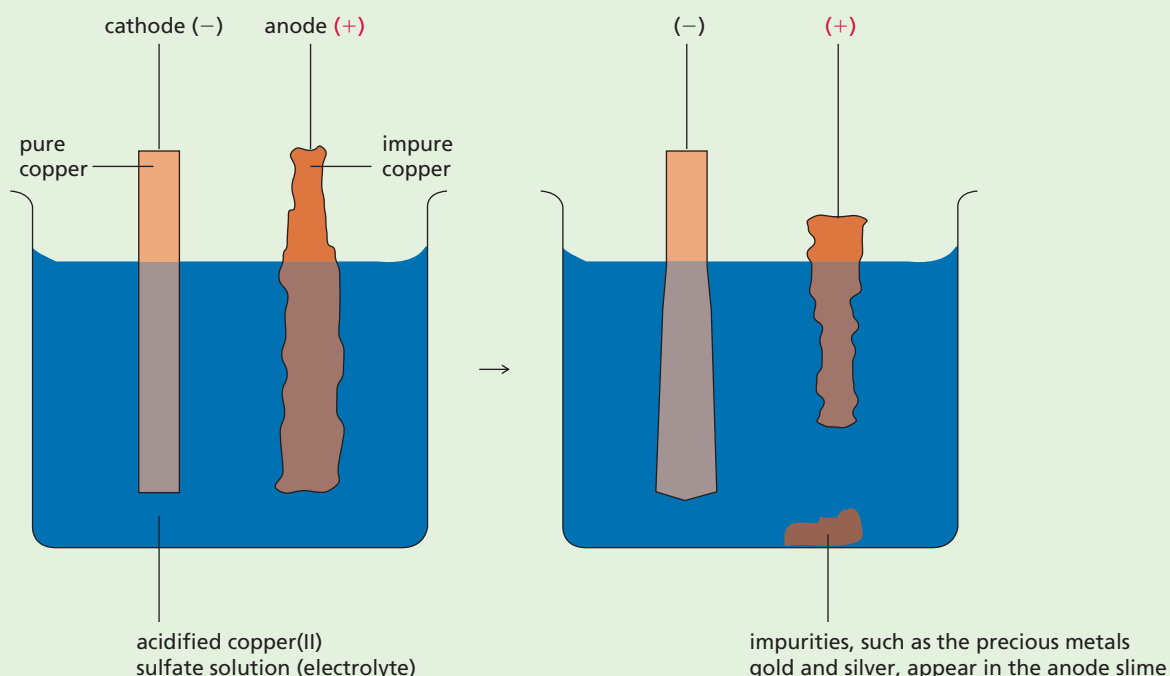
**Figure 5.16a** The copper used in electrical wiring has to be very pure.



**Figure 5.16b** Due to the high density of copper and its cost, steel-cored aluminium cables are used for electrical transmission through national grids.

The impure copper is used as the anode and is typically 1 m square, 35–50 mm thick and 330 kg in weight. The cathode is a 1 mm thick sheet and weighs about 5 kg; it is made from very pure copper. Because copper is itself involved in the electrolytic process, the copper cathode is known as an ‘active’ electrode. The electrolyte is a solution of copper(II) sulfate ( $0.3 \text{ mol dm}^{-3}$ ) acidified with a  $2 \text{ mol dm}^{-3}$  solution of sulfuric acid to help the solution conduct electricity (Figure 5.17).





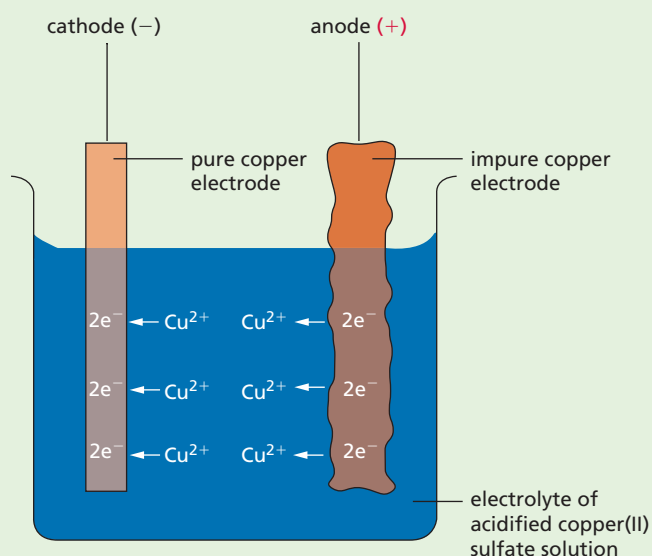
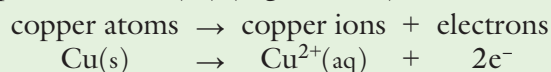
**Figure 5.17** Copper purification process.

When the current flows, the copper moves from the impure anode to the pure cathode. Any impurities fall to the bottom of the cell and collect below the anode in the form of a slime. This slime is rich in precious metals and the recovery of these metals is an important aspect of the economics of the process. The electrolysis proceeds for about three weeks until the anodes are reduced to about 10% of their original size and the cathodes weigh between 100 and 120 kg. A potential of 0.25 V and a current density of  $200 \text{ A m}^{-2}$  are usually used.

The ions present in the solution are:

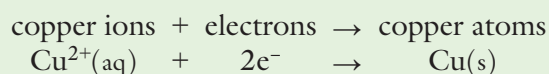
from the water:  $\text{H}^+(\text{aq})$   $\text{OH}^-(\text{aq})$   
 from the copper(II) sulfate:  $\text{Cu}^{2+}(\text{aq})$   $\text{SO}_4^{2-}(\text{aq})$

During the process the impure anode loses mass because the copper atoms lose electrons and become copper ions,  $\text{Cu}^{2+}(\text{aq})$  (Figure 5.18).



**Figure 5.18** The movement of ions in the purification of copper by electrolysis.

The electrons released at the anode travel around the external circuit to the cathode. There the electrons are passed on to the copper ions,  $\text{Cu}^{2+}(\text{aq})$ , from the copper(II) sulfate solution and the copper is deposited or copper plated on to the cathode.



The annual production of copper worldwide is in excess of 16 million tonnes. However, a large amount of the copper we need is obtained by recycling. This way of obtaining copper is increasing in importance as it becomes more difficult and expensive to locate and extract the copper ore.

## ● Electrolysis guidelines

The following points may help you work out the products of electrolysis in unfamiliar situations. They will also help you remember what happens at each electrode.

- Non-metals are produced at the anode whereas metals and hydrogen gas are produced at the cathode.
- At the inert anode, chlorine, bromine and iodine (the halogens) are produced in preference to oxygen.
- At the inert cathode, hydrogen is produced in preference to metals unless unreactive metals such as copper and nickel are present.

### Questions

- 1 Why do you think it is advantageous to use inert electrodes in the electrolysis processes?
- 2 Predict the products of electrolysis of a solution of copper(II) sulfate if carbon electrodes are used instead of those made from copper as referred to in the purification of copper section.
- 3 Predict the products of the electrolysis of concentrated hydrochloric acid using platinum electrodes.
- 4 Using your knowledge of electrolysis, predict the likely products of the electrolysis of copper(II) chloride solution, using platinum electrodes. Write electrode equations for the formation of these products.
- 5 Suggest reasons why:
  - a copper and aluminium are used in electrical cables
  - b plastics and ceramics are used to manufacture handles for metallic cooking utensils.

## ● Electroplating

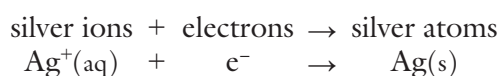
Electroplating is the process involving electrolysis to plate, or coat, one metal with another or a plastic with a metal. Often the purpose of electroplating is to give a protective coating to the metal beneath. For example, bath taps are chromium plated to prevent corrosion, and at the same time are given a shiny, more attractive finish (Figure 5.19).

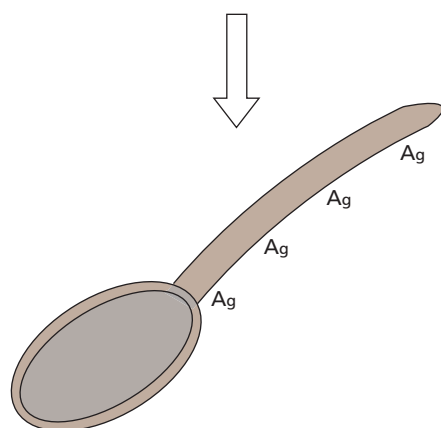
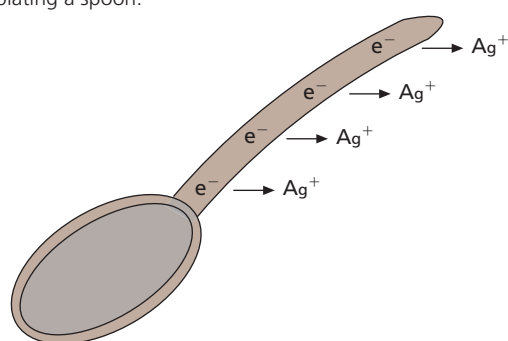
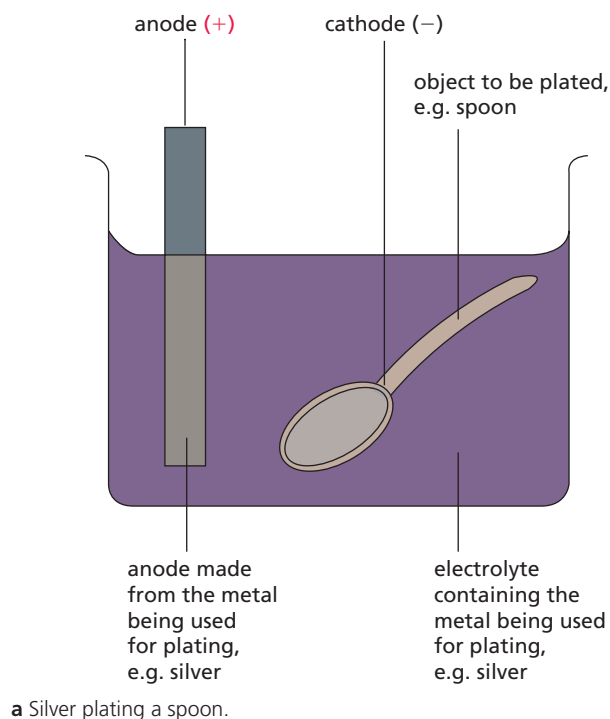


Figure 5.19 This tap has been chromium plated.

The electroplating process is carried out in a cell such as the one shown in Figure 5.20a (p. 84). This is often known as the 'plating bath' and it contains a suitable electrolyte, usually a solution of a metal salt.

For silver plating the electrolyte is a solution of a silver salt. The article to be plated is made the cathode in the cell so that the metal ions move to it when the current is switched on. The cathode reaction in this process is:





b Explaining silver plating.

Figure 5.20

## Plating plastics

Nowadays it is not only metals that are electroplated. Plastics have been developed that are able to conduct electricity. For example, the plastic poly(pyrrole) can be electroplated in the same way as the metals we have discussed earlier (Figure 5.21).



Figure 5.21 This plastic has been coated with copper by electrolysis.



Figure 5.22 This leaf has been electroplated.



## Questions

- 1 The leaf in Figure 5.22 has been copper plated. Suggest a suitable method for copper plating the leaf.
- 2 Explain why copper(II) chloride solution would not be used as an electrolyte in the electrolyte cell used for copper plating.
- 3 Write equations which represent the discharge at the cathode of the following ions:
  - a  $K^+$
  - b  $Pb^{2+}$
  - c  $Al^{3+}$
 and at the anode of:
  - d  $Br^-$
  - e  $O^{2-}$
  - f  $F^-$

## Checklist

After studying Chapter 5 you should know and understand the following terms.

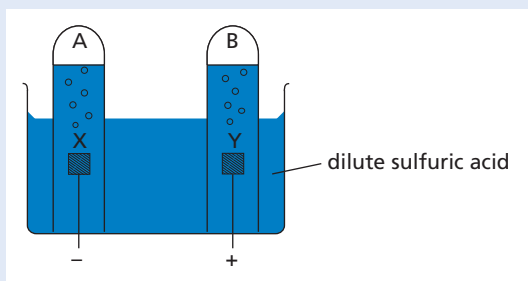
- **Anions** Negative ions; these are attracted to the anode.
- **Anode** The positive electrode. It is positively charged because electrons are drawn away from it.
- **Cathode** The negative electrode. It is negatively charged because an excess of electrons move towards it.
- **Cations** Positive ions; these are attracted to the cathode.
- **Electrode** A point where the electric current enters and leaves the electrolytic cell. An inert electrode is usually made of platinum or carbon and does not react with the electrolyte or the substances produced at the electrodes themselves.
- **Electrolysis** A process in which a chemical reaction is caused by the passage of an electric current.
- **Electrolyte** A substance which will carry electric current only when it is molten or dissolved.
- **Electroplating** The process of depositing metals from solution in the form of a layer on other surfaces such as metal or plastic.
- **Hall-Héroult cell** The electrolysis cell in which aluminium is extracted from purified bauxite dissolved in molten cryolite at  $900^\circ\text{C}$ . This cell has both a graphite anode and a graphite cathode.
- **Inert electrode** These are electrodes that do not react with the products of electrolysis, e.g. carbon, platinum.
- **Membrane cell** An electrolytic cell used for the production of sodium hydroxide, hydrogen and chlorine from brine in which the anode and cathode are separated by a membrane.
- **Oxidation** Takes place at the anode and involves a negative ion losing electrons.
- **Reduction** Takes place at the cathode and involves a positive ion gaining electrons.



# Electricity and chemistry

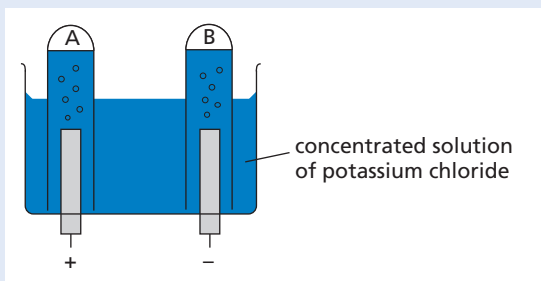
## Additional questions

1



This is a diagram of an experiment in which electricity was passed through a mixture of distilled water containing a little dilute sulfuric acid.

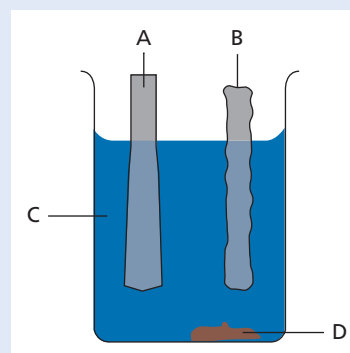
- Name the gas that collects at **A**.
  - Name the gas that collects at **B**.
  - If  $100\text{ cm}^3$  of gas collects in **A** how much would there be in **B**?
  - Name the metal usually used for **X** and **Y**.
  - X** is called the \_\_\_\_\_.
  - Y** is called the \_\_\_\_\_.
  - Write down the formulae of the three ions present in the solution.
  - Write down the equations for the reactions that take place at both **X** and **Y** (or describe the changes that take place if you cannot write the equations).
- 2 The apparatus shown in the diagram below was used to investigate the gases produced when a concentrated solution of potassium chloride was electrolysed.



- Name a non-metal suitable for use as electrodes.
- Name the gas collected in **A** and the gas collected in **B**.
- Describe how you would test the gases collected.
- The volume of gas collected in **B** was slightly less than that collected in **A**. The teacher said the volumes should have been equal but gave

a simple explanation of the 'missing' gas in **B**. What was the explanation? (Assume that the apparatus was working perfectly).

- Write down the equations which describe the production of the gases at the electrodes in **A** and **B**.
  - If the concentrated solution of potassium chloride was now replaced by dilute sodium hydroxide what gases would be produced at **A** and **B**?
    - In what ratio would you expect these gases to be produced?
- 3 Explain the meaning of each of the following terms. Use a suitable example, in each case, to help with your explanation.
- |                 |              |
|-----------------|--------------|
| a Anode.        | e Anion.     |
| b Cathode.      | f Cation.    |
| c Electrolysis. | g Oxidation. |
| d Electrolyte.  | h Reduction. |
- 4 Copper is purified by electrolysis, as in the example shown below.



- Name the materials used for the electrodes **A** and **B**.
- Name the electrolyte **C** and substance **D**.
- Why is substance **D** of economic importance in respect of this process?
- Write equations for the reactions which take place at the cathode and anode during this process.
- Draw a labelled diagram to show the cell after electrolysis has taken place.
- Why has electrolyte **C** to be acidified with the dilute sulfuric acid?
- Why does copper have to be 99.99% pure for use in electrical cables?

- 5 Copy and complete the table below, which shows the results of the electrolysis of four substances using inert electrodes.

Electrolyte	Product at anode (positive electrode)	Product at cathode (negative electrode)
Molten aluminium oxide		Aluminium
Concentrated sodium chloride solution	Chlorine	
Molten lithium chloride		
Silver nitrate solution		Silver

- State what you understand by 'inert electrodes'.
  - Explain why the lithium chloride solution becomes progressively more alkaline during electrolysis.
  - Explain why solid lithium chloride is a non-conductor of electricity, whereas molten lithium chloride and lithium chloride solution are good conductors of electricity.
  - During the electrolysis of molten aluminium chloride ( $\text{AlCl}_3$ ) the carbon anodes are burned away. Explain why this should happen and write balanced chemical equations for the reactions that take place.
- 6 A pupil carried out an experiment in a fume cupboard to find out how electricity affected different substances. Some of the substances were in aqueous solution, others were in the molten state. Carbon electrodes were used in each experiment and she wrote down her results in a table with these headings.

Substance	What was formed at the cathode (-)	What was formed at the anode (+)

Make a table like the one shown and fill it in with what you think happened for each of the substances below.

- Molten lead iodide.
- Sugar solution.
- Silver nitrate solution.
- Copper(II) sulfate solution.
- Molten sodium bromide.
- Ethanol solution.

- 7 Sodium hydroxide is made by the electrolysis of brine.

- Draw and label a simplified diagram of the cell used in this process. Make certain that you have labelled on the diagram:
    - the electrolyte
    - the material of the electrodes
    - the material of the membrane.
  - Write equations for the reactions which take place at the cathode and anode. State clearly whether a reaction is oxidation or reduction.
  - Give two large-scale uses of the products of this electrolytic process.
  - Comment on the following statement: 'This electrolytic process is a very expensive one'.
  - Both the membrane cell and the older mercury cell make sodium hydroxide of high purity. Explain why the membrane cell is now the preferred way of making sodium hydroxide.
- 8 Electroplating is an important industrial process.
- Explain what electroplating is.
  - Why are certain metals electroplated?
  - Give two examples of the use of electroplating.