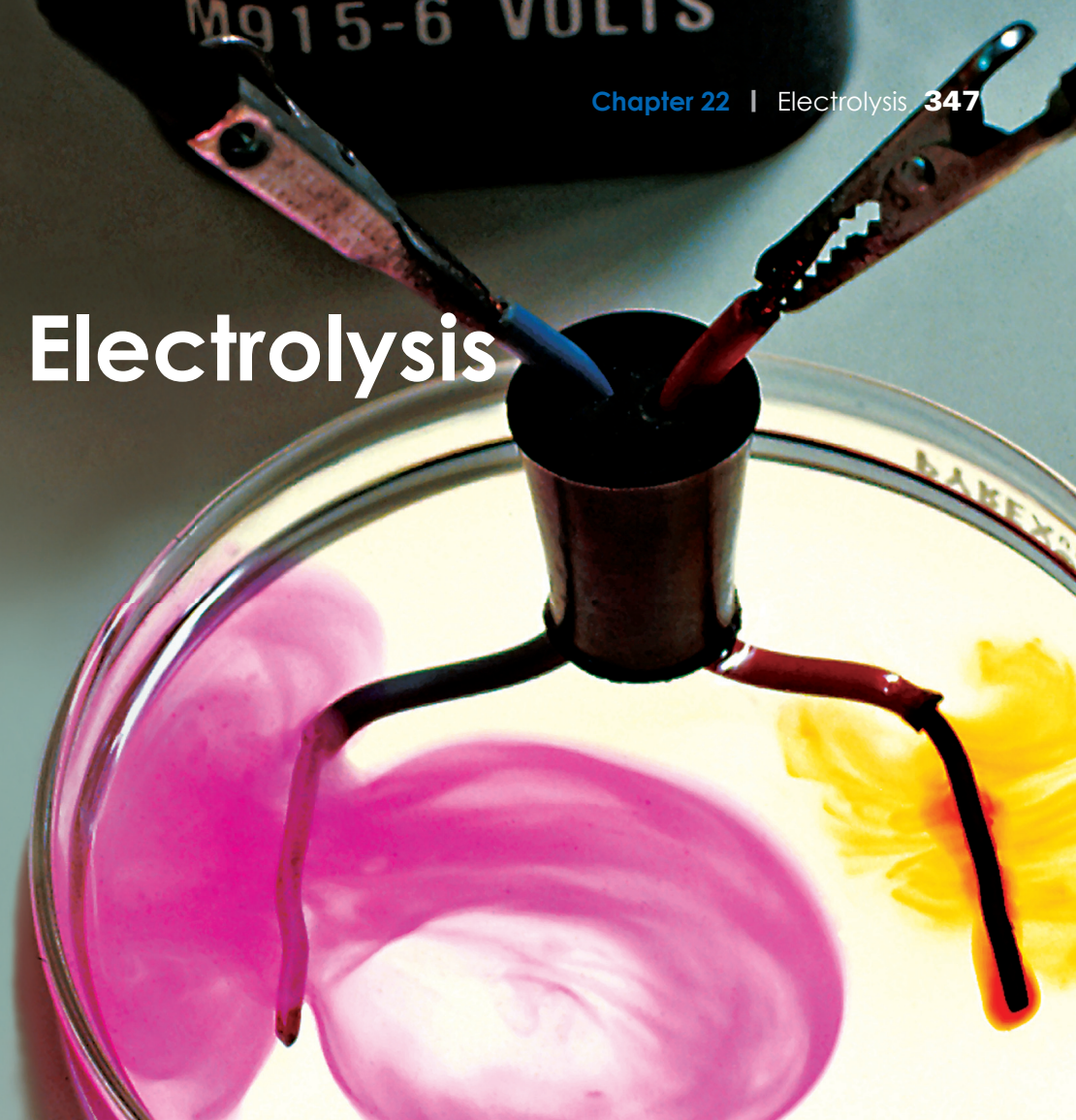


22

Electrolysis



Introduction

Electrolysis has been referred to in earlier chapters, such as the conduction of electricity by compounds (Chapter 6) and the extraction of metals (Chapter 14). In this chapter, students acquire a theoretical understanding of electrolysis. Further applications of electrolysis show the importance of the process in the modern world. The production of electrical energy from chemical cells is also investigated. Throughout the chapter, links are made to earlier ideas, including atomic structure, bonding, reactivity series and oxidation and reduction.

Chapter Opener (page 344)

1. Begin the chapter by discussing the following questions. Precise answers are not needed at this stage.

What is electrolysis? How is electrolysis applied in daily life?

Answer: Electrolysis is the decomposition of a compound by using electricity. For uses, refer to Figure 22.1 on page 345 of the Textbook.

What are the products of the electrolysis of copper(II) chloride solution using carbon electrodes?

Answer: The products are copper (at the cathode) and chlorine gas (at the anode).

What is electroplating? How can we electroplate an object in the laboratory?

Answer: Electroplating is the coating of an object with a thin layer of a metal by electrolysis. Refer to Figure 22.18 on page 356 of the Textbook.

2. Carry out an 'Inquiry Preview.'

Learning Outcomes

After completing this chapter, the students should be able to:

- ▶ describe electrolysis as the conduction of electricity by an ionic compound (an electrolyte), when molten and dissolved in water, leading to the decomposition of the electrolyte
- ▶ describe electrolysis as evidence for the existence of ions
- ▶ describe, in terms of the mobility of ions present and the electrode products, the electrolysis of molten sodium chloride, using inert electrodes
- ▶ predict the likely products of the electrolysis of a molten binary compound
- ▶ apply the idea of selective discharge based on cations, anions and concentration effects using inert electrodes
- ▶ describe the electrolysis of aqueous copper(II) sulfate with copper electrodes as a means of purifying copper
- ▶ predict the products at the electrodes of the electrolysis of aqueous electrolytes, given relevant information
- ▶ construct ionic equations for the reactions occurring at the electrodes during the electrolysis of given substances
- ▶ describe the electroplating of metals
- ▶ describe the product of electrical energy from simple electric cells

Teaching pointers

22.1 How is Electrolysis Used in Daily Life? (page 345)

Stimulation

Show the class some products of electrolysis such as those shown in Figure 22.1 on page 345 of the Textbook. Comment briefly on each product and the use of electrolysis in their production. Link these objects and questions to the applications of electrolysis in daily life and to ideas from previous chapters that will be used and developed as the chapter is being taught. End this section by getting students to carry out a class preview.

Teaching pointers

1. Electrolysis only occurs with ionic compounds. Revise the idea of an ionic compound and get the class to name some.
2. Get the class to suggest some ways our lives would be affected by the absence of electrolysis (or for people in the days before electrolysis), for example, the absence of aluminium. This is an exercise in imagination. Students could refer to life without the things in Figure 22.1.

Teaching pointers

22.2 What is Electrolysis? (page 346)

1. Spend some time discussing the terms associated with electrolysis. An electrolytic cell can also be called an electrolysis cell.
2. Graphite is commonly used as electrodes during electrolysis as it is
 - (a) cheap,
 - (b) a good conductor of electricity, and
 - (c) inert.

Link this idea with the properties of graphite that were covered in Chapter 6 of the Textbook.

- Carry out a word analysis of the term 'electrolysis' and get students to appreciate and derive its meaning. The word 'electrolysis' consists of two parts: 'electro-' (referring to electricity) and '-lysis' (meaning splitting). Electrolysis therefore is the decomposition/splitting of compounds using electricity.
- An additional Chemistry in Society on the discovery of electrolysis and the role that the English scientist, Sir Humphry Davy played in the use of electrolysis is found at the end of this chapter. You may photocopy and distribute it to students.

Skills Practice (page 347)

- Electrolysis is the decomposition/splitting of compounds using electricity.
- The cation is negative as electrons (negatively charged) flow into the electrode. The anode is positive as electrons flow out of the electrode.
 - Anions
 - Cations
- An electrolyte is an ionic compound which, when molten or in aqueous solution, conducts an electric current and is decomposed in the process.
 - Sodium chloride, copper(II) sulfate and sulfuric acid are electrolytes.

Notes for Teachers

The discovery of electrolysis

Electrolysis depends on a supply of electricity such as an electric cell or battery. The first electric battery was invented in 1800 and electrolysis was discovered later in the same year when the English scientists, William Nicholson (1753–1815) and Anthony Carlisle (1768–1840) decomposed water into hydrogen and oxygen.

But it was the English scientist, Humphry Davy (1778–1829) who first realised the importance of electrolysis. Using one of the first batteries, Davy discovered the metals potassium, calcium and magnesium by means of electrolysis. (Davy also suggested the use of zinc to prevent the corrosion of iron nails holding copper plates to the hulls of ships; refer to Practice question 7 in Chapter 13 on pages 206 of the Textbook.)

Davy was one of the many 18th and 19th Century scientists who developed our understanding of the relationship between matter and electricity. Other scientists included:

- Luigi Galvani (1737–1798), who showed that two different metals in certain solutions (electrolytes) can produce an electric current.
- Alessandro Volta (1745–1827), who invented the first electric battery (refer to page 360 of the Textbook).
- Michael Faraday (1791–1867), who discovered the concept of the ion, invented the terms for describing the process of electrolysis and discovered the laws of electrolysis.

On the aspects of safety in Chemistry, it is interesting to note that Davy was in the habit of smelling and tasting chemicals, which eventually killed him. You may again emphasise the danger of tasting unknown substances. (Refer also to the story of Robert Bunsen and the consequences of not using safety goggles in the laboratory on page 270 of the Textbook.)

Chemistry Inquiry (page 349)

Group Discussion

- Bubbles of brown gas (bromine) are observed at the anode. A silvery metal (lead) is deposited on the cathode.
- The electrode reactions are:

cathode: $\text{Pb}^{2+}(\text{l}) + 2\text{e}^{-} \longrightarrow \text{Pb}(\text{s})$ (reduction)

anode: $2\text{Br}^{-}(\text{l}) \longrightarrow \text{Br}_2(\text{l}) + 2\text{e}^{-}$ (oxidation)

The electrodes are electrically charged. As sodium and chlorine are produced at the electrodes during the electrolysis, we can infer that the sodium chloride must consist of charged particles, which we call ions that are attracted to the electrodes.

Skills Practice (page 349)

- Sodium ions are attracted to the cathode. Each ion takes one electron from the cathode to become a sodium ion. Chloride ions are attracted to the anode. One electron from each ion is transferred to the anode leaving chlorine atoms, which combine to form chlorine molecules Cl_2 .
- Solid sodium chloride consists of sodium and chloride ions packed closely together. The strong (electrostatic) forces of attraction between these ions hold them in place. Thus the ions are not free to move and do not conduct electricity for electrolysis to occur.
 - Chemical changes during electrolysis occur at the electrodes only. Migration of ions occurs in the molten electrolyte.
- The electrodes are electrically charged. As sodium and chlorine are produced at the electrodes during the electrolysis, we can infer that the sodium chloride must consist of charged particles, which we called ions, that are attracted to the electrodes.
- Potassium chloride:
 - K^+ and Cl^-
 - Potassium is produced at the cathode and chlorine gas is produced at the anode.
 - Cathode: $K^+(l) + e^- \longrightarrow K(l)$
Anode: $2Cl^-(l) \longrightarrow Cl_2(l) + 2e^-$

Sodium iodide:

 - Na^+ and I^-
 - Sodium is produced at the cathode and iodine gas is produced at the anode.
 - Cathode: $Na^+(l) + e^- \longrightarrow Na(l)$
Anode: $2I^-(l) \longrightarrow I_2(l) + 2e^-$

Aluminium oxide:

 - Al^{3+} and O^{2-}
 - Aluminium is produced at the cathode and oxygen gas is produced at the anode.
 - Cathode: $Al^{3+}(l) + 3e^- \longrightarrow Al(l)$
Anode: $2O^{2-}(l) \longrightarrow O_2(g) + 4e^-$

Notes for Teachers**Electrolysis of molten lead(II)bromide**

This electrolysis can be carried out using the set-up shown in Figure 22.4 on page 347 of the Textbook.

Here are some points to take note of:

- A 6 volt d.c. power supply (e.g. a battery) is suitable. Instead of an ammeter, use a light bulb (6V) to show that a current flows.
- The molten lead(II) bromide used must be of a good quality, otherwise brown fumes are given off with heat alone.
- Carry out the demonstration in a fume cupboard as the fumes produced are toxic.
- As the electrodes in the crucible are close together, it may be difficult to tell exactly where the bromine vapour comes from.
- Bubbles of brown gas (bromine) are observed at the anode; a silvery metal (lead) is deposited on the cathode. The electrode reactions are:

$$\begin{array}{ll} \text{cathode: } Pb^{2+}(l) + 2e^- \longrightarrow Pb(s) & \text{(reduction)} \\ \text{anode: } 2Br^-(l) \longrightarrow Br_2(g) + 2e^- & \text{(oxidation)} \end{array}$$
- At the end of the electrolysis, allow the crucible to cool and the electrolyte to solidify. Then tap the porcelain crucible upside-down on the bench top. The solid lead(II) bromide will fall out, revealing a shiny silvery globule of lead metal, which is produced at the cathode.

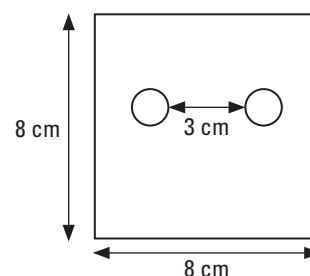
Video clips of electrolysis reactions

You may view video clips of electrolysis reactions from CD-ROMs such as *Electrochemistry* and *Chemistry Set 2000*.

Teaching pointers

22.4 What Happens During Electrolysis of Aqueous Solutions of Ionic Compounds? (page 350)

1. The rules for the selective (or preferential) discharge of ions are oversimplified. In many cases, two kinds of ions are discharged at an electrode at the same time. However, it is not necessary to be concerned about this concept at the 'O' Level. On completing this section, students should be able to predict the products at the electrodes.
2. In Experiment 22.1 of the Practical Workbook, the effect of the nature of the electrodes on the products of the electrolysis of copper(II) sulfate solution is investigated. A holder for the carbon rods in the beaker can be made from cardboard, plastic or wood, as shown on the right.
3. In Experiment 22.2 of the Practical Workbook, students test their models/understanding of the electrolysis of aqueous solutions by predicting the products formed, checking their predictions experimentally, and if necessary, modifying their models/understanding.
4. Additional Exercise 1 at the end of this chapter provides an opportunity for students to investigate products from the chlor-alkali (or chloro-alkali) industry, which uses electrolytic cells to produce chlorine, hydrogen and sodium hydroxide solution from the electrolysis of brine. The chlorine and sodium hydroxide are combined to form sodium hypochlorite, which is used in household bleach solutions. The industry began in the 1890s in Europe and the USA. The worksheet may be photocopied and distributed to the class.
5. Discuss the 'electrolysis of water' and ensure that students appreciate that this does not involve pure water but solutions that produce the elements of water (hydrogen and oxygen) during the electrolysis.



Chemistry Inquiry (page 354)

Group Discussion

1. Bubbles of a colourless gas (hydrogen) form around the cathode while bubbles of another colourless gas (oxygen) form at the anode.
 2. Cathode: $2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2(\text{g})$
Anode: $4\text{OH}^-(\text{aq}) \longrightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
 3. With both solutions, hydrogen and oxygen, i.e. the elements of water, are produced.
 4. Seawater contains a dilute solution of sodium chloride.
-

Skills Practice (page 354)

1. Aqueous solution of sodium sulfate:

(a) Sodium ions, $\text{Na}^+(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, sulfate ions, $\text{SO}_4^{2-}(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Oxygen. $4\text{OH}^-(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
Cathode: Hydrogen. $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

Aqueous solution of magnesium chloride:

(a) Magnesium ions, $\text{Mg}^{2+}(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, chloride ions, $\text{Cl}^-(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Chlorine. $2\text{Cl}^-(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$
Cathode: Hydrogen. $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

Note: Here, $\text{Cl}^-(\text{aq})$ ions are discharged in preference to $\text{OH}^-(\text{aq})$ ions. Only in (very) dilute solutions are hydroxide ions discharged in preference to chloride (and other halide) ions. This is an example of the 'concentration effect'.

Aqueous solution of zinc sulfate:

(a) Zinc ions, $\text{Zn}^{2+}(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, sulfate ions, $\text{SO}_4^{2-}(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Oxygen. $4\text{OH}^-(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
Cathode: Hydrogen. $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

Aqueous solution of potassium iodide:

(a) Potassium ions, $\text{K}^+(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, iodide ions, $\text{I}^-(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Iodine. $2\text{I}^-(\text{aq}) \rightarrow \text{I}_2(\text{g}) + 2\text{e}^-$
Cathode: Hydrogen. $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

Aqueous solution of sodium hydroxide:

(a) Sodium ions, $\text{Na}^+(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Oxygen. $4\text{OH}^-(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
Cathode: Hydrogen. $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$

Aqueous solution of copper(II) chloride:

(a) Copper ions, $\text{Cu}^{2+}(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, chloride ions, $\text{Cl}^-(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Chlorine. $2\text{Cl}^-(\text{aq}) \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$
Cathode: Copper. $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$

Aqueous solution of silver nitrate:

(a) Silver ions, $\text{Ag}^+(\text{aq})$, hydrogen ions, $\text{H}^+(\text{aq})$, nitrate ions, $\text{NO}_3^-(\text{aq})$, hydroxide ions, $\text{OH}^-(\text{aq})$.

(b) (c) Anode: Oxygen. $4\text{OH}^-(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
Cathode: Silver. $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$

2. The electrolysis of solutions of sodium sulfate, zinc sulfate and sodium hydroxide.

3. Cathode: Silver, $\text{Ag}(\text{s})$, is deposited as it is lower in the selective discharge series than hydrogen (which forms hydrogen gas).

Anode: The silver anode dissolves to form silver ions, $\text{Ag}^+(\text{aq})$.

4. Any aqueous solution of a compound containing a cation of hydrogen or above in the reactivity series and any anion apart from halide ions, e.g. sodium hydroxide, sulfuric acid, zinc nitrate.

Teaching pointers

22.5 What are Some Industrial Applications of Electrolysis? (page 355)

- In some copper refining plants, the silver and gold which are part of the impurities are worth more than the copper!
- Link the topic of electroplating to its use in Chapter 13 for preventing rusting by plating iron/steel with metals such as tin, chromium and silver.
- The diagram in Figure 22.18 gives the impression that electroplating is a simple process. Industrially, the electroplating solution is a complex mixture of substances. For example, in silver plating, an alkaline solution of silver ions and cyanide ions is used to get the silver to stick better. Other additives, such as ammonium thiosulfate, produce a finer surface on the final product.
- You may like to demonstrate the electroplating of one metal with another metal. Refer to the 'Notes for Teachers' on the next page for details.
- The electroplating industry is a source of environmental pollution in many countries. In Exercise 22.1 of the Theory Workbook, students get to discuss electroplating pollution issues through a role play.
- Discuss the basic idea of cathodic protection against rusting. Link and compare it with the use of sacrificial protection to protect iron and steel from rusting that was studied in Chapter 13.

(page 357)

Mystery Clue

An anode, such as graphite or scrap metal, is placed in the ground and connected to the steel part of the bridge and to a power supply. The diagram for this is similar to that in Figure 22.20 on page 357 of the Textbook. A small current from the power supply protects the steel from rusting.

7. An additional Chemistry in Society on electroless plating for electroplating non-conductors of electricity is found at the end of this chapter. You may photocopy and distribute it to students.

TWB Exercise 22.1: Role play (page 138, part B)

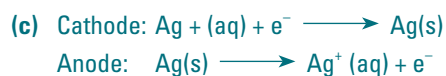
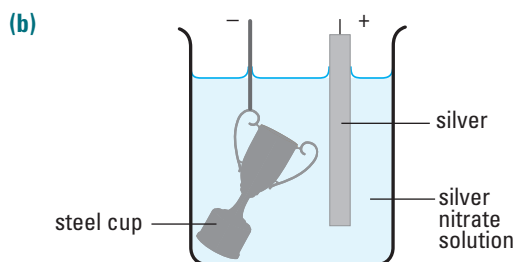
1. A role play can be used when an issue, particularly a social issue, affects different kinds of people in different ways. The issue is discussed from the point of view of each kind of person. In this activity, the discussion centres on how pollution caused by electroplating waste affects different people, such as a factory owner, workers in the factory, local residents affected by the pollution, and so on. A major purpose of a role play is to give those taking part the feeling of the role and to make decisions which represent that group of people.

The company name given in this activity is imaginary but the arguments given for the reluctance of the company to treat its wastewater are real. Some points that could be used in the role play are listed in the Theory Workbook. Here are some other points that could be considered for the various roles:

- Factories must treat their waste before discharging it.
- Treatment will increase the cost of products, making them uncompetitive with other countries.
- Treatment is not really that expensive; costs can be spread over several years.
- Get small electroplating companies to combine into larger factories where treatment can be controlled economically or group factories in one location.
- Small companies may go out of business, causing unemployment and harming the economy.
- In future, build only large factories which can afford to treat wastes.
- Our health is important.
- Let the government clean up the environment instead (i.e., the taxpayer pays!).
- The taxpayer should not have to pay.
- Pollution kills fish that we could otherwise eat.
- Provide improved drainage systems to collect wastewater into the sewerage system instead of discharging them into streams or the sea.

Skills Practice (page 357)

1. For example, copper is used in circuit wires and cables as it is a good conductor of electricity; copper is used in water pipes as it is resistant to corrosion; it is used in bronze and brass alloys as it is harder and stronger than pure copper.
2. (a) (i) The steel cup.
(ii) A sheet of silver.
(iii) An aqueous solution of a silver salt, such as silver nitrate.



3. Cover the leaf with graphite or aluminium paint. Set up an electroplating bath as in Question 2 but with the leaf as the cathode in place of the steel cup.

Notes for Teachers

Experiments on electroplating

The following information is provided for teachers and laboratory technicians on performing electroplating experiments. It is not intended for students.

(a) Copper plating:

To demonstrate copper plating, use the set-up as shown in Figure 22.18 on page 356 of the Textbook. However, use the following plating solution, which is more effective than using copper(II) sulfate solution alone: 200 g/dm³ copper(II) sulfate, CuSO₄·5H₂O, a few drops of dilute hydrochloric acid (to supply Cl⁻ ions) and 50 g/dm³ sulfuric acid. The sulfuric acid helps to increase the conductivity of the solution. It also helps to bring about a finer grain size of the copper deposit at the cathode. This solution gives a current density of 0.02 A/cm².

(b) Nickel plating a copper object:

For a simple demonstration, again use the set up shown in Figure 22.18. However, use a copper object or sheet, a nickel plate and a solution of nickel(II) sulfate together with a 6 V dc battery. To prepare an effective plating solution, use 100 g of nickel(II) sulfate, NiSO₄·6H₂O, 16 g of nickel(II) chloride, NiCl₂·6H₂O (or 18 g of sodium chloride) and 18 g of boric acid, H₃BO₃ per dm³ of aqueous solution. (About 10 g of magnesium sulfate, MgSO₄·7H₂O, might also be added.)

Nickel(II) chloride (or sodium chloride) increases the conductivity of the solution and facilitates the dissolution of the anode. Boric acid acts as a buffer to maintain a constant pH. It acts best at the pH range 4–6. Magnesium sulfate improves the conductivity of the solution, helps to produce a light and even deposit of nickel on the cathode. This solution has a pH of 4–5 and gives a current density of 0.05 A/cm².

Notes:

- To obtain good results for electroplating experiments in the laboratory, the plating solutions should be prepared from reagent grade chemicals and distilled water. The solution should be filtered to eliminate any insoluble particulate impurities.
- In nickel plating, the pH of the solution may be adjusted by the addition of 10% sulfuric acid or nickel(II) carbonate. (The pH of the solution should be monitored closely with a pH meter.)
- Stirring of the solution and moving the cathode during electroplating can produce a better deposit.

Teaching pointers

22.6 How do Simple Electric Cells Work? (page 358)

- To introduce this topic, show the class several kinds of electrical cells used in daily life. Point out that these are rather complex cells and that in this section, it is just the basic principles of electric cells that will be studied.
- Note that the type of electric cells studied in this section can also be called chemical cells. See 'Notes for Teachers' on the next page.
- Refer to the electrodes in electric cells as 'positive' and 'negative'. Although the terms anode and cathode can be used, they should be avoided because the convention is opposite to that used in electrolysis. This will prevent confusion.
- Demonstrate a simple chemical cell. This could include
 - an orange/lemon connected to a digital clock. Refer to Figure 22.21 on page 358 of the Textbook. (See also 'Notes for Teachers' on the next page)
 - the cell in Figure 22.22 on page 358 of the Textbook connected to a light bulb or small digital clock (instead of a voltmeter).

- Strictly speaking, the voltage of a cell depends on the order of metals in the electrochemical series, which is based on voltage (e.m.f.) measurements, rather than the reactivity series, which is based on chemical reactions. Both series, however, show the relative tendencies of metals to form ions. But the order of some metals, such as calcium, differs in the two series because of the different ways of deriving the series.
- In Experiment 22.3 of the Practical Workbook, ensure students are aware that on electrical meters, the red terminal is the positive terminal while the black terminal is the negative terminal.
- An additional exercise on the Babylon Electric Cell is found at the end of this chapter. This can be used in conjunction with the Chemistry in Society on electric cells in history. You may photocopy and distribute the worksheet for the exercise to students.
- The chapter can be rounded off by getting students to play a card game on several industrial processes studied in this and previous chapters, which can be found at the end of this chapter. (See the notes on page 357.)

Skills Practice (page 359)

- No, as sugar solution is not an electrolyte.
 - Sodium is above hydrogen in the reactivity/selective discharge series and sodium ions are harder to discharge than hydrogen ions.
- The metal at the top and the metal at the bottom. For example, caesium and silver or gold.
- The further away the metal is from magnesium in the reactivity series, the larger the voltage.
 - Aqueous sodium chloride

(c) (i)

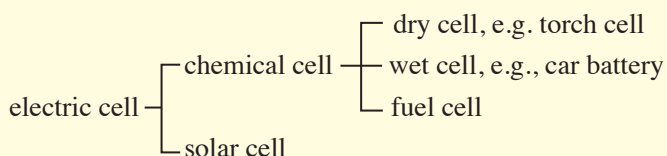
Cell number	Metal
1	Silver
2	Copper
3	Iron
4	Zinc
5	Magnesium

(ii) Gold

Notes for Teachers

Terminology: Types of electric cells

Electric (or electrical) cell is the general term for all kinds of devices that produce electricity as follows:



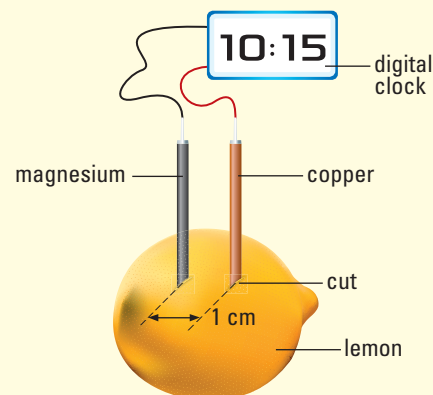
In chemical cells, chemical energy is converted into electrical energy. Fuel cells are investigated in Chapter 24.

Luigi Galvani and electric current

Luigi Galvani (1737–1798), an Italian doctor, discovered that two different metals produce an electric current when they make contact with solutions (electrolytes) in animal tissue, such as in the muscles of frogs' legs. The terms galvanic cell and galvanometer are named in honour of him. Simple chemical cells are also called galvanic cells.

A lemon cell

Make two narrow cuts in a lemon about 1 cm apart. Clean a strip of copper (6 cm × 2 cm) and a length of magnesium ribbon (10–15 cm) with sandpaper. Fold the magnesium ribbon twice. Insert the metals into the cuts and connect them to a small digital clock.



Chemistry in Society

Electric Cells in History (page 360)

Exercise

- Electrodes – iron rod and copper cylinder. Electrolyte – probably salt solution or vinegar.
 - Electrodes – Zinc and silver. Electrolyte – salt water.
- His interest in science could have made him curious about the components as to whether they could be part of some ancient device. This prompted him to examine the components further until he discovered that the components were for an electric cell.
 - The study of science need not be confined to be only done by scientists. Anybody can contribute to the development of scientific knowledge.
- He found the prior scientific knowledge important as he could build on the knowledge and discover the electric cell earlier.

Additional Exercise 2: The Babylon Electric Cell

These notes refer to the 'Chemistry in Society' on page 360 of the textbook and to Additional Exercise 2.

Constructing the historical cells

Demonstrate the two historical cells referred to in the textbook. These are

- Volta's battery, and
- The 2500-year old cell from Baghdad.

For Volta's battery, use strips of zinc and copper (instead of the original silver that Volta used) separated by filter paper or cloth soaked in salt water. Measure the voltage produced then see if the battery will light up a light bulb. (See also Exercise 22.2B of the Theory Workbook and the activity for 'Extension' on page 365 of the Textbook.)

To replicate the Baghdad cell, place an iron rod in a copper can filled with salt solution or vinegar. The iron rod must not touch the iron can. Measure the voltage produced.

Video

Show students a video clip of a zinc/copper battery similar to Volta's battery on a CD-ROM (e.g. *Chemistry Set 2000*).

Part B: Testing the archaeologist's hypothesis

Construct the cell as described above. To test the cell, connect the copper and iron electrodes to an electroplating cell (such as the one shown in Figure 22.18 on page 356 of the Textbook). If there is no observable plating, increase the voltage by connecting the cells in series.

Notes on the ancient electric cells

Replicas of the original Baghdad cells have been made that produce voltages ranging from 0.8 to nearly two volts. Connected in series, a set of batteries could produce a much higher voltage, though no wires have ever been found that would prove this had been the case.

It is uncertain exactly what the cells were used for. One theory holds that several cells could have been linked together to generate a higher voltage for electroplating gold to a silver surface. Experiments with several Baghdad-type cells have shown this to be possible.

It is also known that the Greeks and Romans used certain species of electric fish in the treatment of pain. They would stand on a live electric eel until their pained feet went numb. Perhaps the cells were used as a source of electricity for the relief of pain

IT Link

There are many Internet sites providing information on the Baghdad cells. Here are examples of some of these websites:

<http://news.bbc.co.uk/2/hi/science/nature/2804257.stm>

<http://www.unmuseum.org/bbattery.htm>

<http://fusionanomaly.net/batteries.html>

Additional Activity 3: Industrial Process Game

This card game is designed to reinforce important industrial processes studied in earlier chapters. The following processes are included in the game:

- Manufacture of steel (Textbook Chapter 14)
- Recycling of aluminium (Textbook Chapter 14)
- Haber Process (Textbook Chapter 17)
- Household bleach (sodium hypochlorite) production (Activity 20.1 of the Theory Workbook)

The aim of this game is to select appropriate cards to build up the flow diagram of the processes. The game would serve as an interesting and instructional conclusion to this chapter. It can be used now or at any other convenient time.

A game board and instructions on how the game could be played are given with the activity. The set of diagram cards needed for the game boards is provided at the end of this chapter. You may photocopy these cards onto (coloured) cardboard, cut out the individual cards and made them into sets for the class to use. You may want to include a few extra copies of some of these cards to include in the game. All cards should be the same size to fit the game board for this activity.

Solving the **Mystery** (page 361)

Why did the Kinzua Bridge Collapse?

Infer

The iron bolts should have been protected against rusting by painting.

Connect

Temperature is the main factor. For effects of speed of reaction in daily life, refer to Chapter 20.

Further Thought

Use more modern methods for protection against rusting such as sacrificial protection or cathodic protection (followed up by regular checking and replacement of the bolts when necessary).



22 Chapter Review



Self-Management

Misconception Analysis (page 362)

1. **True** Electrolysis occurs because ions from compounds move towards the electrodes in an electrolytic cell.
2. **False** An anion has a negative charge and is attracted to the positively-charged anode.
3. **True** When ions are discharged at electrodes, they gain or lose electrons. Thus, reactions at electrode are redox reactions.
4. **False** 'The electrolysis of water' refers only to solutions that produce hydrogen and oxygen during electrolysis, such as dilute sulfuric acid and sodium sulfate solution.
5. **True** Only hydrogen ions or ions of metals low in the reactivity series are discharged. This is the idea behind the preferential discharge of ions.
6. **False** Electrodes are inert when they conduct electricity but do not react. For example, carbon electrodes are inert electrodes when used for the electrolysis of sodium chloride solution. But in the electrolysis of copper(II) sulfate with copper electrodes, the copper anode forms copper(II) ions and dissolves and so is not an inert electrode.
7. **True** Thus a simple cell with magnesium and copper as electrodes will produce a higher voltage than a cell with zinc and copper electrodes.

Practice

Structured Questions (pages 363 – 364)

- oxygen; hydrogen; chlorine; potassium
 - First process:
 Anode: $4\text{OH}^-(\text{aq}) \longrightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^-$
 Cathode: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$

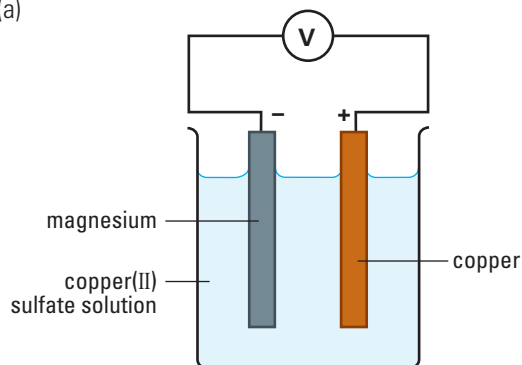
 Second process:
 Anode: $2\text{Cl}^-(\text{aq}) \longrightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$
 Cathode: $2\text{H}^+(\text{aq}) + 2\text{e}^- \longrightarrow \text{H}_2(\text{g})$

 Third process:
 Anode: $2\text{Cl}^-(\text{l}) \longrightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$
 Cathode: $\text{K}^+(\text{l}) + \text{e}^- \longrightarrow \text{K}(\text{l})$
- Molten sodium chloride conducts electricity as the ions are free to move. Solid sodium chloride does not conduct electricity, therefore electrolysis does not occur.
 - This is to lower the melting point of sodium chloride.
 - Cathode
 - $\text{Na}^+(\text{l}) + \text{e}^- \longrightarrow \text{Na}(\text{l})$
 - Chlorine gas
 - $2\text{Cl}^-(\text{l}) \longrightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$
 - Chlorine gas is toxic (poisonous). Molten sodium will react violently if it comes into contact with water or moisture.
- SO_4^{2-}
 - By combining the two half equations and simplifying, the overall equation becomes:
 $2\text{H}_2\text{O}(\text{l}) \longrightarrow 2\text{H}_2(\text{g}) + \text{O}_2(\text{g})$
 Thus the volume of hydrogen gas (formed at the cathode) is twice as large as oxygen gas (formed at the anode).
 - Anode: iodine; cathode: lead
 - Anode: oxygen; cathode: hydrogen
 - Electroplating is used on steel objects to protect them from rusting.
- Carbon/Graphite rods
 - Syringe **A** collects hydrogen gas while syringe B collects oxygen gas.
 - The gas in **B** (oxygen gas) will relight a glowing splint.
 - Electrode **Q**
 - Oxidation occurs at the anode (electrode **R**) as anions lose electrons at this electrode.

Compound	Possible Identity
A	Sodium or potassium bromide
B	Sodium or potassium chloride
C	Copper(II) sulfate
D	Sodium or potassium iodide

- Calcium hydroxide
 - $\text{Mg}^{2+}(\text{aq}) + 2\text{H}^-(\text{aq}) \longrightarrow \text{Mg}(\text{OH})_2(\text{s})$
 - In molten magnesium chloride (unlike the solid), the ions are free to move. Thus electrolysis can occur.
 - $2\text{Cl}^-(\text{l}) \longrightarrow 2\text{Cl}^-(\text{g}) + 2\text{e}^-$
 - $\text{Mg}^{2+}(\text{l}) + 2\text{e}^- \longrightarrow \text{Mg}(\text{l})$
 - Aluminium
- Aqueous copper(II) sulfate solution (unlike the solid) contains ions that are free to move, therefore it can conduct electricity.
 - Bubbles of a colourless gas (oxygen) are produced.
 - $4\text{OH}^-(\text{aq}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + 4\text{e}^-$
 - Copper
 - E.g. For plating, the sheet of copper used for plating is the anode in the electroplating cell. / To replace the copper(II) ions in the solution that are deposited on the object being plated.
 - $\text{Cu}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq}) \longrightarrow \text{Cu}(\text{OH})_2(\text{s})$
 - E.g. To remove substances that may discolour water or be harmful to living organisms before being discharged into rivers or the sea.

8. (a)



- The magnesium (negative) strip is connected to the black terminal of the voltmeter as the black terminal is the negative terminal.
 - $\text{Mg}(\text{s}) \longrightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$
 - $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \longrightarrow \text{Cu}(\text{s})$

Note: The $\text{Cu}^{2+}(\text{aq})$ ions come from the copper(II) sulfate solution.
 - No
 - The magnesium reacts directly with the copper(II) sulfate solution and becomes covered with a layer of copper. As the two electrodes are now the same, no current flows.
 - Use an electrolyte that does not react with magnesium. An example is sodium chloride solution.
- Copper is the positive terminal of the battery as it is lower than calcium in the reactivity series.
 - Replace the copper with silver.
 - It can be used in a spacecraft landing on a hot planet such as Mercury or Venus.

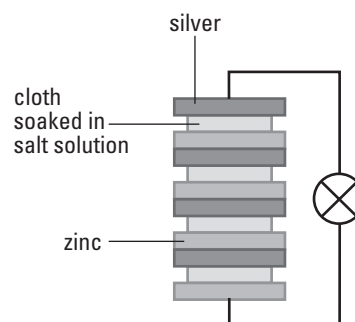
Free Response Questions (pages 363–364)

- Silver, copper, hydrogen (from the water) and lead.
- Responses to this question may include the following points:
Description of the experimental set-up:
 - Pure copper as cathode and impure copper as anode.
 - Copper(II) sulfate solution as electrolyte.
 Process at anode:
 - Copper forms copper(II) ions.
 - $\text{Cu(s)} \longrightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-}$
 - Impurities fall to the bottom of the beaker.
 Process at cathode:
 - Copper(II) ions change to copper atoms.
 - $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \longrightarrow \text{Cu(s)}$
 - Copper deposited.
- Responses to this question may include the following points:
 - Connect the magnesium strip to the negative terminal of a voltmeter and the copper strip to the positive terminal.
 - Place the two metal strips into the beaker of sodium chloride solution. Note the voltage obtained.
 - Repeat the experiment with the silver connected to the negative terminal of the voltmeter in place of copper. The voltage obtained will be larger than with the magnesium-copper pair as magnesium and silver are further apart in the reactivity series than magnesium and copper.

- Repeat with other metal combinations, such as zinc-copper and zinc-silver. Smaller voltages are obtained (as zinc is lower than magnesium in the series) but the zinc-silver pair produces a larger voltage than the zinc-copper pair as zinc and silver are further apart in the series.

Extension (page 365)**Making an Electric Battery**

To make this battery, refer to the following diagram:



A smaller voltage is obtained if zinc and lead are used as the electrodes instead of zinc and silver.

Additional Teaching Material

Chemistry in **Society**

The Discovery of Electrolysis



Electrolysis depends on a supply of electricity such as an electric cell or battery. The first electric battery was invented in 1800. Electrolysis was discovered later the same year. But it was the English scientist, Humphry Davy who first realised the importance of electrolysis. The story below tells part of Davy's life and work.

	<p>1 Humphry Davy was born in England in 1778. He went to school but did not work hard. He was often naughty at school.</p>		<p>4 Davy also gave lectures to the public. Hundreds of people would come to hear him talk and do experiments about Chemistry.</p>
	<p>2 When Davy was aged 15, his father died. Davy left school to help his mother and her four other children. He began work by helping a doctor. At the age of 19, he began to study Chemistry by himself. He did experiments in the attic of his home.</p>		<p>5 Davy was interested in many activities. He helped to start the London Zoo. He loved fishing and enjoyed poetry. He wrote many poems himself.</p>
	<p>3 In 1801, Davy went to work in London and soon became famous. Using one of the first batteries, he discovered the metals potassium, calcium and magnesium by means of electrolysis. Davy also suggested the use of zinc to prevent the corrosion of iron nails holding copper plates to the hulls of ships (see Chapter 13).</p>		<p>6 Davy used to smell and taste new chemicals. This dangerous practice often made him sick and eventually killed him!</p>

Chemistry in **Society**

Electroless plating

Electroplating Non-conductors of Electricity

Non-conducting materials, such as wood and plastic, can also be electroplated. A simple way of doing this is to first paint or spray the object with a conducting paint such as aluminium or graphite. The electroplating is then carried out on this surface.

A more advanced way is called electroless plating. This is used to plate plastic circuit boards (Figure 20.18). A circuit is first etched into the plastic, creating many grooves. The plastic board is dipped into a bath containing a metal compound and then treated with

chemicals to reduce the metal ions in the grooves to the metal. These metallic grooves are then plated with copper to produce the electrical circuit.

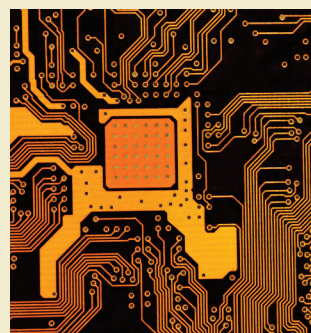


Figure 1 A circuit board

Additional Teaching Material

Additional Exercise 1: Products of the Salt Industry

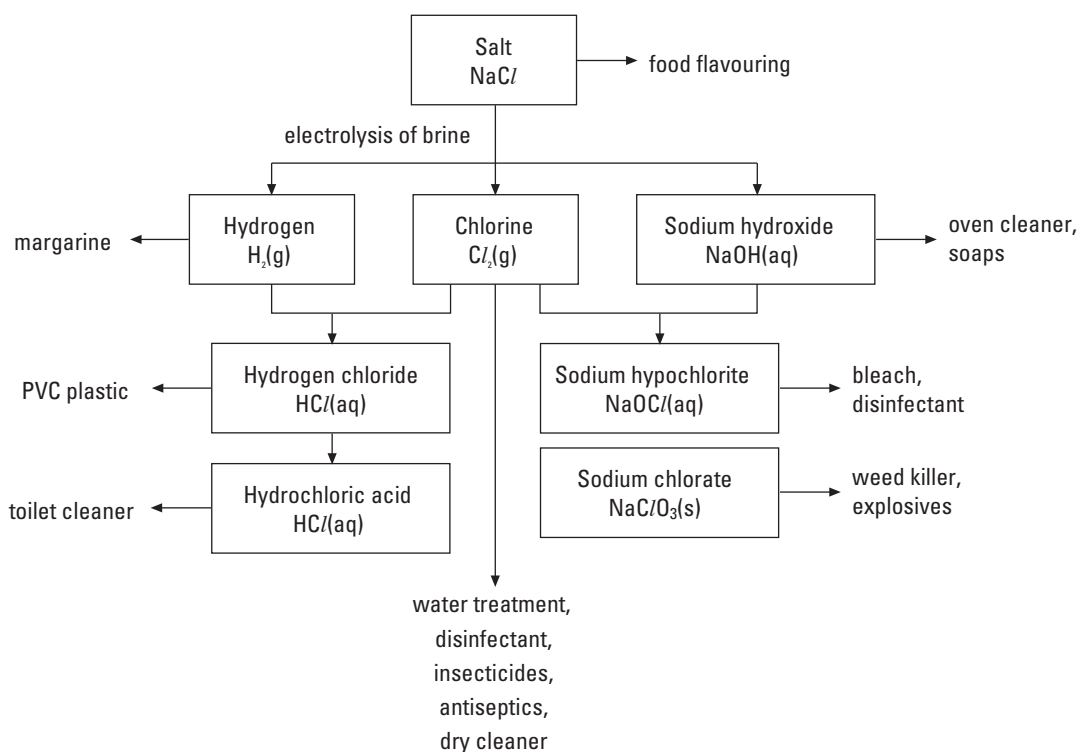
Objective

- To appreciate the variety of uses of the products made from the electrolysis of brine (concentrated sodium chloride solution) in daily life

Key Competency

ICS: Management of information [applying chemistry in daily life]

There are many products in our homes that are obtained from the salt, sodium chloride. The flow chart below shows the uses of chemicals produced by the electrolysis of brine (concentrated salt solution).



Sodium hydroxide is used to make soaps and detergents.



Liquid hydrogen is a fuel for rockets.



Chlorine is used to sterilise water in swimming pools.



Look around your home and make a list of uses of salt and chemicals from the electrolysis of brine. To help you, refer to the flow chart on the previous page.

Complete the table below. One use has been done for you.

Use in the home	Chemical from the salt industry
Chlorine of tap water	Chlorine gas

Additional Teaching Material

Additional Exercise 2: The Babylon Electric Cell

Objective

- ▶ To appreciate that simple electric cells may have been used in ancient times

Key Competency

ICS: management of information

A. An ancient electric cell

Babylon is an ancient city situated south of Baghdad, in present-day Iraq (Figure 1). At the site of ancient Babylon, archaeologists discovered what they thought was an electric cell made about 2500 years ago. The cell consisted of an iron rod inside a copper cylinder as shown in Figure 2. The cylinder probably contained an electrolyte solution such as salt solution or vinegar.

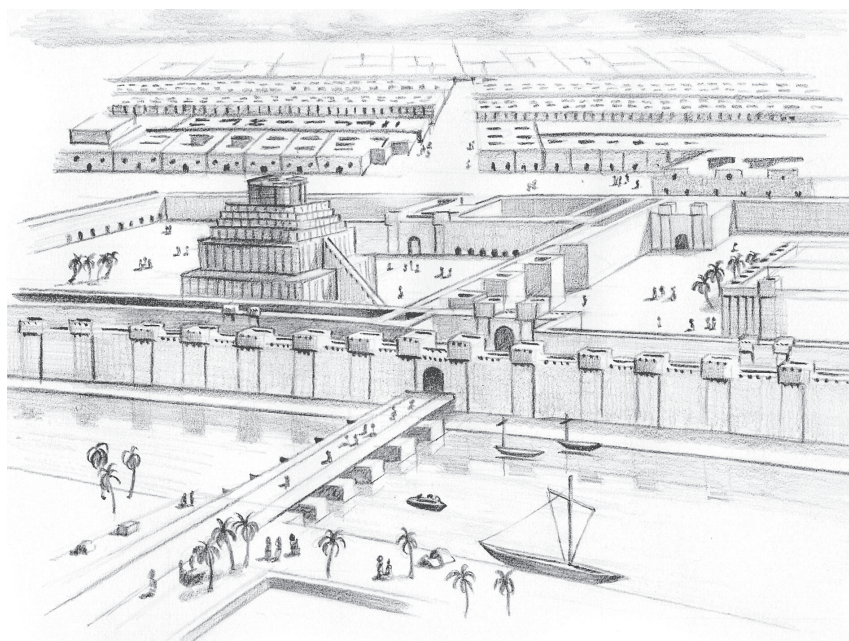


Figure 1 Babylon as it may have looked 2500 years ago

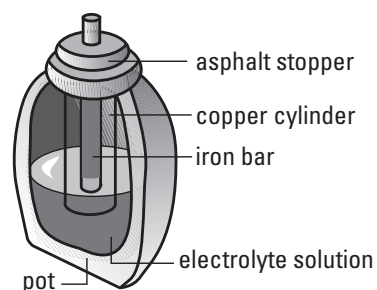


Figure 2 A model of the electric cell found at the site of ancient Babylon

Archaeologists are not sure what the cells were used for but have hypothesised that electric current from the cells may have allowed jewellers to electroplate metals and to make copper shine like gold or silver.

Questions

1. A simple electric cell consists of three main parts.

(a) Name these parts.

(b) Name these parts in the Babylon cell.

2. (a) Which of the two metals used as electrodes is higher in the reactivity series?

(b) Therefore, which metal is the negative electrode in the cell?

3. Write an ionic equation for the reaction at the iron electrode.

4. In the cell, bubbles of gas are formed at the copper electrode.

(a) Suggest what is the gas formed.

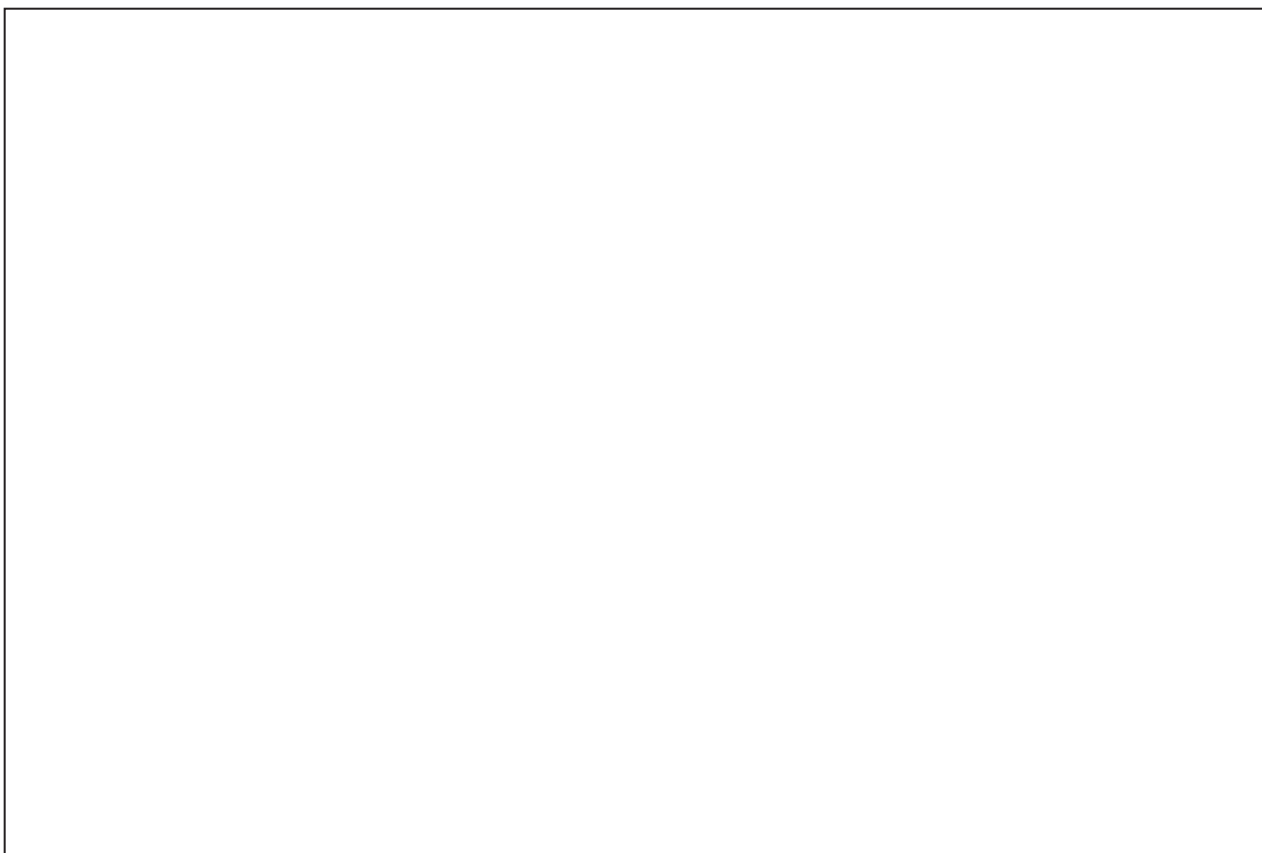
(b) Write an ionic equation for the formation of this gas.

5. In which direction do electrons move in the external circuit, i.e. through the connecting wires?

6. (a) Estimate the voltage of the cell discovered in Babylon.

(b) Suggest one metal that could have been used instead of copper to increase the voltage.

7. Draw and label a diagram using a beaker, the same electrolyte and metal electrodes as in the cell discovered in Babylon.



B. Testing the archaeologists' hypothesis

The archaeologists believed that the cell was used to electroplate metals. Briefly describe an experiment you could do to test the hypothesis that electroplating is possible with a simple iron-copper cell.

Note:

If you want a larger voltage to carry out the electrolysis, several cells can be connected in series. However, by doing this, you are assuming that people knew about series circuits 2500 years ago. Is this a reasonable assumption?

Additional Teaching Material



Additional Exercise 3: Industrial Process Game

Objective

- ▶ To complete a flow chart for an industrial process

Key Competencies

- CIT:** curiosity and creativity
ICS: communicating effectively [*learning through games*]

Game contents

1 game board

1 set of diagram cards for each of the following industrial processes from this theme:

- Manufacture of steel
- Recycling aluminium
- Haber Process
- Household bleach (sodium hypochlorite solution)
(Your teacher will give you these cards.)

To play the game

1. Form groups of 2 to 4 players.
2. Each player chooses one of the four industrial processes. Write the name of the process below the game board.
3. Place the diagram cards face down on a table.
4. Decide who will play first.
5. Each player, in turn, turns up one card.
 - (a) If this card is part of the process chosen by this player, the card is placed anywhere on the game board. As more cards are collected, they are placed in the correct order to make the flow chart.
 - (b) If the card is not suitable, it is returned to the same place on the table, face down.

Note: If you see a card that is part of your process, but is not your turn, try to remember the position of the card on the table. Then you can pick it up when your turn comes.

6. The winner is the first person to complete the flow chart correctly. The game may continue until the other players complete their boards.

Note: The teacher can decide if the completed flow chart is correct. If a player thinks another person has a wrong card, check with the teacher. If the card is wrong, it is immediately returned to the table.

Game board

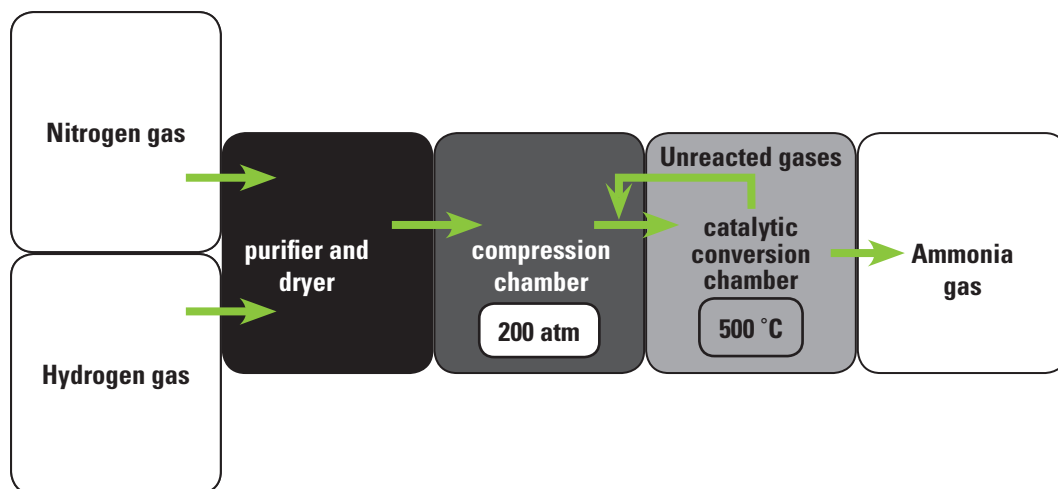
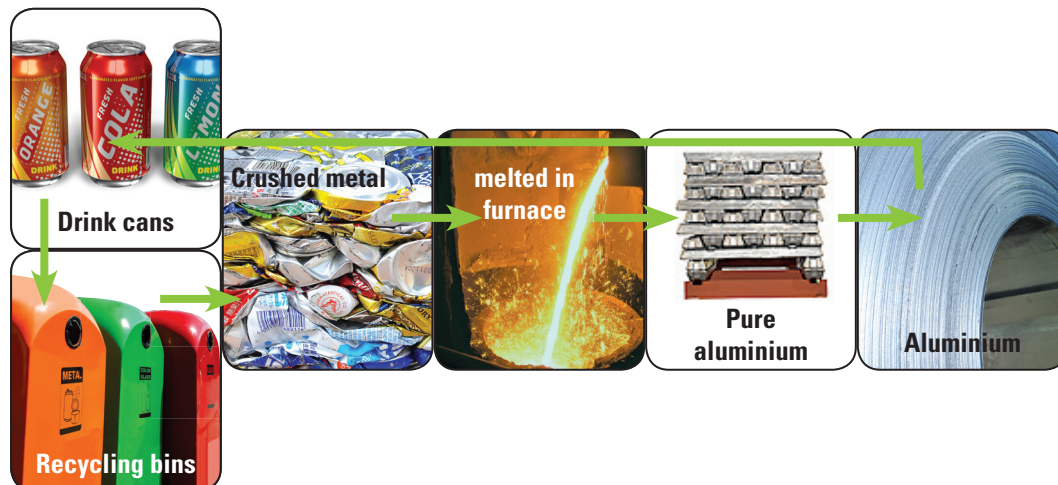
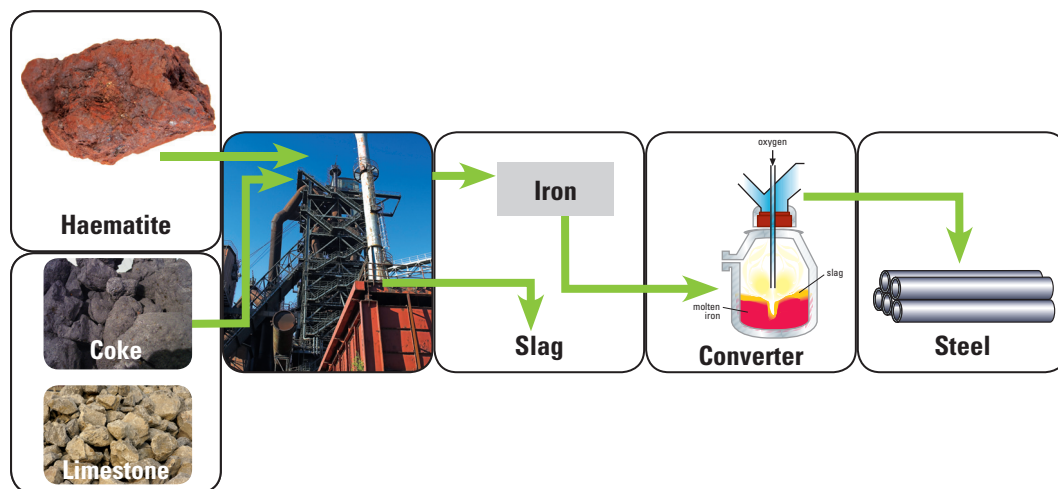
Industrial Process Game

Industrial process: _____

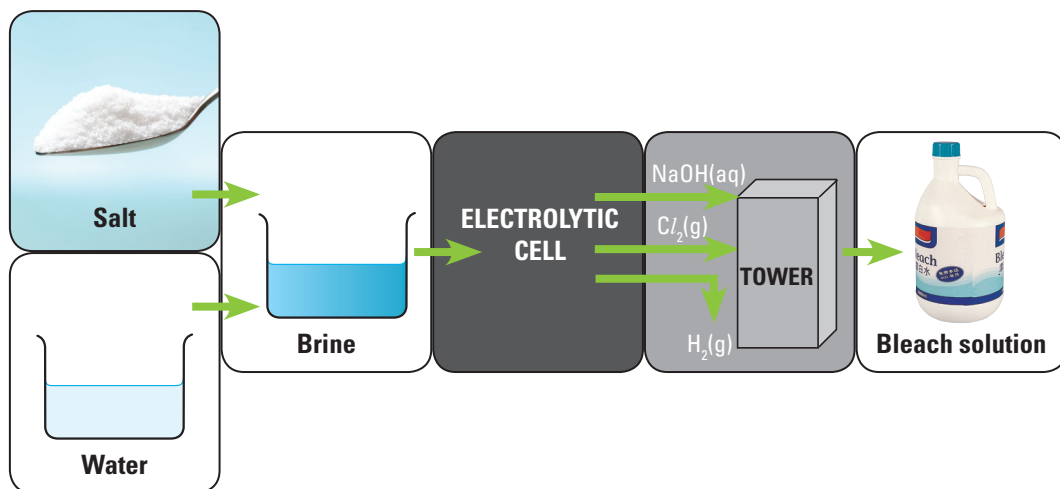
Cards for Additional Exercise 3

Completed flow diagrams of the following processes:








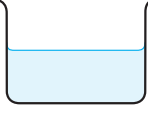
Manufacture of steel



Household bleach (sodium hypochlorite) production



Cards for Additional Exercise 3

<p>Unreacted gases</p> <p>catalytic conversion chamber</p> <p>500 °C</p>	 <p>Bleach solution</p>	 <p>Aluminium</p>	
 <p>Drink cans</p>	<p>Iron</p> <p>Slag</p>	<p>Nitrogen gas</p>	 <p>melted in furnace</p>
 <p>Crushed metal</p>	<p>purifier and dryer</p>	<p>NaOH(aq)</p> <p>Cl₂(g)</p> <p>TOWER</p> <p>H₂(g)</p>	 <p>Coke</p>  <p>Limestone</p>
<p>Hydrogen gas</p>	 <p>Haematite</p>	<p>Ammonia gas</p>	 <p>Salt</p>
 <p>Pure aluminium</p>	 <p>Brine</p>	 <p>molten iron</p> <p>slag</p> <p>oxygen</p> <p>Converter</p>	<p>ELECTROLYTIC CELL</p>
 <p>Steel</p>	<p>compression chamber</p> <p>200 atm</p>	 <p>Water</p>	 <p>Recycling bins</p>

Answers

Additional Exercise 2:

A. An ancient electric cell

- (a) Any two different metals (electrodes) and an electrolyte solution.

(b) Metal electrodes — iron and copper.
Electrolyte — vinegar (dilute ethanoic acid).
- (a) Iron is higher than copper in the reactivity series.

(b) Iron is the negative electrode.
- (a) $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$
- (a) Hydrogen gas, H_2

(b) $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$
- From the iron electrode to the copper electrode.
- (a) [Actual value under standard conditions = 0.78 volts.]

(b) A metal lower in the reactivity series such as silver or gold.

B. Testing the archaeologists' hypothesis

E.g. To replicate the cell, place an iron rod and a copper rod in a beaker filled with salt solution or vinegar.

The iron rod must not touch the copper rod.
To test the cell, connect the copper and iron electrodes to an

electroplating cell. If there is no observable plating, increase the voltage by connecting the cells in series.

Blank