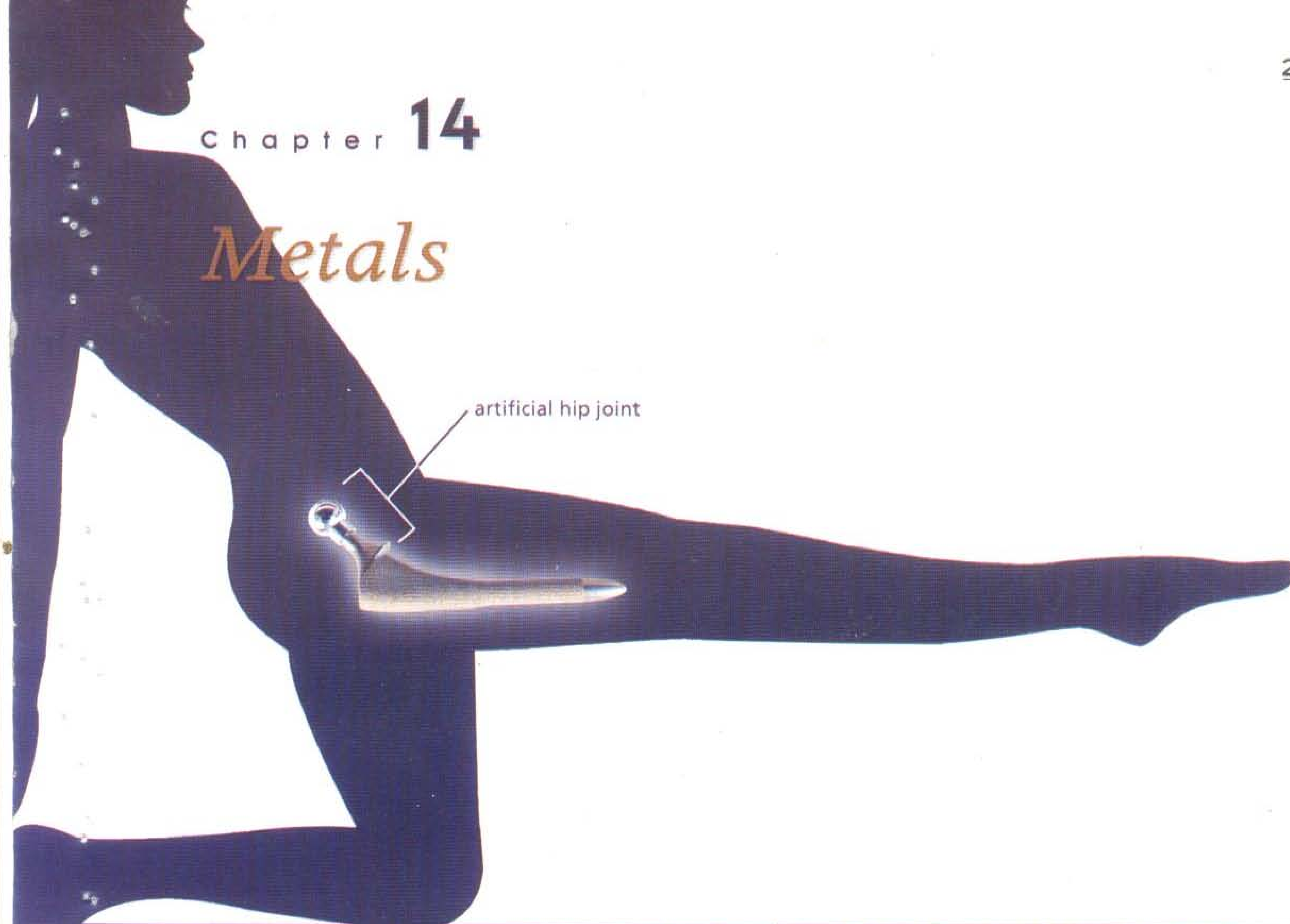
Chapter 14

Metals



The hip is the joint that joins the leg bone to the rest of the skeleton. It leaves the leg free to move. Old people often develop brittle bones and suffer hip fractures when they fall. If the hip is badly damaged, they will be unable to walk. Luckily, advances in medicine and the study of materials now allow doctors to replace a damaged hip with an artificial one made of metal.

An artificial hip must be very strong. It must not wear away easily and must also be light. Artificial joints are commonly made of metal alloys like cobalt-chromium alloys, titanium alloys and stainless steel alloys. Why are pure metals not used instead? You will understand after you learn about the physical and chemical properties of metals in this chapter.

Chapter Outline

14.1 | Metals and Alloys

The physical properties of any substance are determined by the way its particles (atoms, molecules or ions) are packed. Fig. 14.1 shows how atoms are arranged in metals.

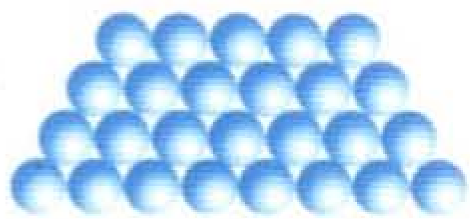


Fig. 14.1 A simple model of the structure of metals

How are the properties of metals related to their structure?

1. Atoms in a metal are packed tightly in layers and are held together by strong **metallic bonds**. Due to these strong metallic bonds, metals usually have *high densities, melting points and boiling points.*

2. In a pure metal, atoms are packed regularly in layers. All these atoms are of the same size. This makes it easy for the layers of atoms to slide over each other when force is applied (Fig. 14.2).

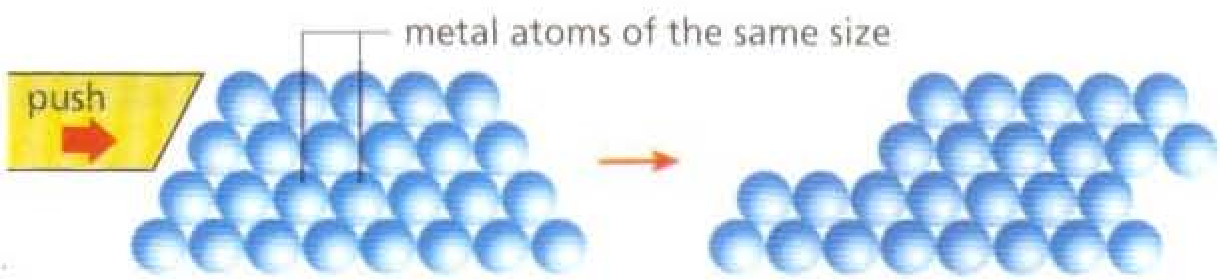


Fig. 14.2 It is easy for the atoms in a pure metal to slide over one another.

This makes metals soft, *ductile* (i.e. they can be drawn into fine wires without breaking) and *malleable* (they can be beaten into thin sheets).

3. While the atoms of a metal are tightly packed, the outermost electrons of the atoms can break away easily from the atoms. In other words, the structure of metals can also be described as positive (metal) ions surrounded by a 'sea of mobile electrons'. The mobile electrons allow metals to *conduct electricity* when they are connected to an electrical source. Heat energy is also transferred easily by the mobile electrons in the structure. This makes metals *good conductors of heat.*

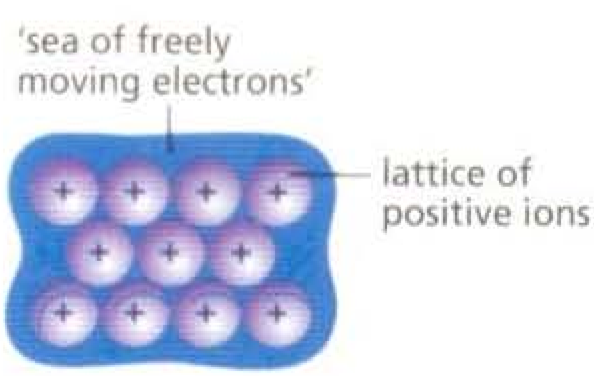


Fig. 14.3 The outermost electrons of the atoms in metals are mobile. Thus, they allow metals to conduct electricity and heat.

Pure metals have many useful properties but they are not widely used. This is because many pure metals

- are soft,
- may react with air and water and wear away easily. We say that these pure metals **corrode** easily (or have a low resistance to **corrosion**).

Most metallic substances used nowadays are **alloys**.

What are alloys?

An **alloy** is *a mixture of a metal with one or a few other elements.* For example,

- bronze is an alloy of copper and tin,
- brass is an alloy of copper and zinc,
- stainless steel is an alloy of iron, chromium, nickel and carbon.

Alloys are made by mixing the molten elements (metals or metals and carbon) in the right proportions and then allowing them to solidify. The alloys produced have more useful physical properties than the pure metals.

Why are metals often used in the form of alloys?

1. Metals can be made harder and stronger by alloying them with other elements. For example, brass is harder and stronger than its constituents, pure copper or pure zinc.

   When a pure metal is alloyed, a different element is added to the pure metal. Atoms of the added element have a different size from those of the pure metal. This breaks up the regular arrangement of atoms in the pure metal. The atoms of different sizes cannot slide over each other easily (Fig. 14.4). This makes the alloy harder and less malleable.

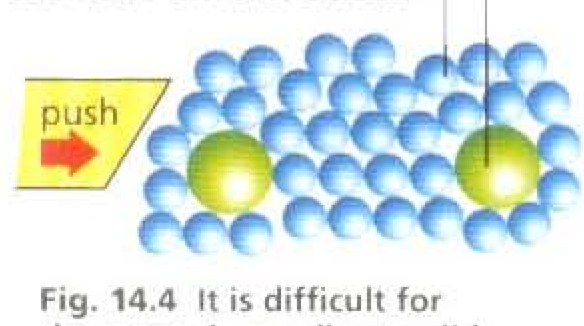


Fig. 14.4 It is difficult for the atoms in an alloy to slide over one another.

2. Alloying can also be used to improve the appearance of the metal. Pewter is an alloy of tin, antimony and copper. It is used to make ornaments and souvenirs because it looks more beautiful than pure tin.

This keychain is made of pewter, an alloy of tin.

3. Alloys are also more resistant to corrosion than pure metals. For example, pure copper corrodes easily. This is why an alloy of copper is used to make coins instead.

4. Alloying is used to lower the melting points of metals. Solder is an alloy of tin and lead. It has a lower melting point than pure tin or pure lead and can be used to join metals.

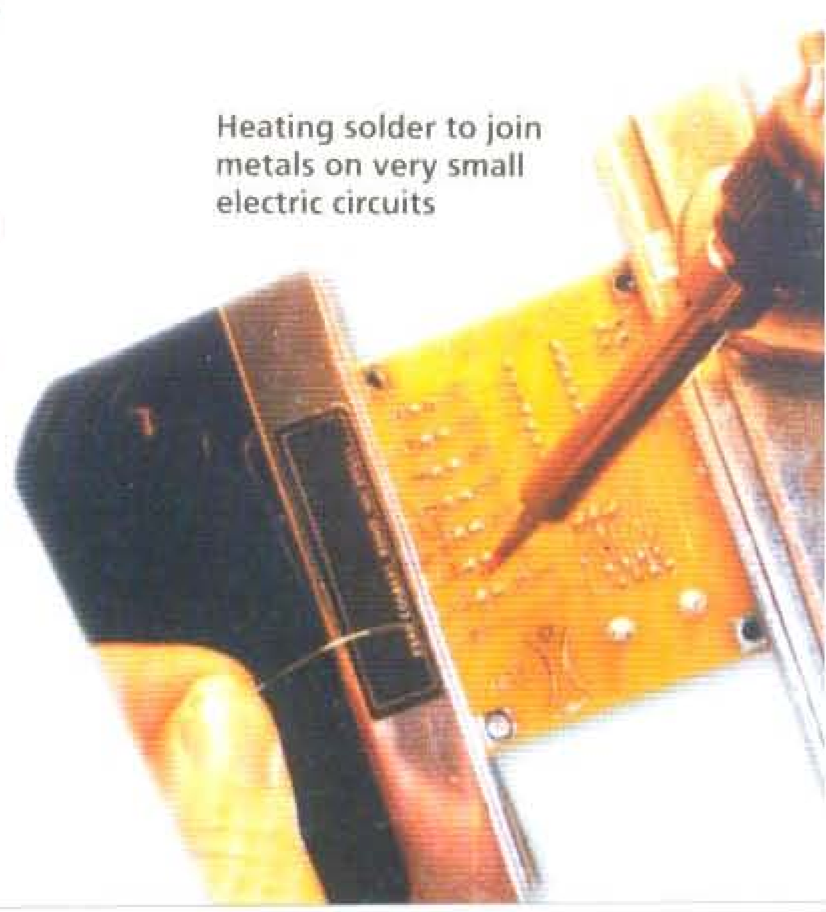
Table 14.1 shows examples of alloys, their composition, properties and uses.

Alloy	Composition*	Special properties	Some uses
Brass	copper (70%) zinc (30%)	does not corrode easily, attractive yellow colour like gold	decorative ornaments, musical instruments, coins
Stainless steel	iron (73%) chromium (18%) nickel (8%) carbon (1%)	resistant to corrosion	cutlery, utensils, medical instruments, pipes in chemical industries
Solder	tin (50%) lead (50%)	low melting point	joining metals, e.g. joining metal pipes
Pewter	tin (95%) antimony (3.5%) copper (1.5%)	bright, shiny colour like silver	decorative ornaments

*The percentages of metals used to make each alloy may vary.

Table 14.1 Examples of alloys, their compositions, properties and uses

Heating solder to join metals on very small electric circuits



Copper-nickel alloy is used to make coins. Why is pure copper not used instead?

Key ideas

1. The physical properties of metals are as follows:
   - Usually have high densities, melting points and boiling points
   - Can be bent, stretched or beaten into very thin sheets without breaking
   - Good conductors of heat and electricity
2. An alloy is a mixture of a metal with one or a few other elements.
3. There are four main reasons for making alloys:
   - To improve the strength and hardness of metals
   - To improve the appearance of metals
   - To improve the resistance of metals against corrosion
   - To lower the melting points of metals

Test Yourself 14.1

Worked Example

Which property need not be considered when choosing a metal for making coins?

A Chemical reactivity
B Electrical conductivity
C Hardness
D Melting point

Thought Process

If the metal is too reactive, coins made from it will corrode easily. The chosen metal has to be hard so that the coins do not change shape. It must also have a high melting point so that the coins are not melted down easily. All metals conduct electricity, so this property need not be considered.

Answer

B

Questions

1. Tungsten is used to make the filament inside a bulb because it conducts electricity and has a high melting point. Explain how the structure of metals results in these properties.

2. Draw simple diagrams to show the difference in the arrangement of atoms between copper and brass.

3. 'The properties of alloys are usually different from those of the elements they contain.' Explain why this is so for any two properties.

14.2 | The Reactivity Series

Metals not only have many common physical properties, they also undergo many similar chemical reactions. Some of these reactions are as follows:

1. Form positive ions by the loss of electrons, e.g.
   $Fe \longrightarrow Fe^{2+} + 2e^-$
2. Form ionic compounds, e.g. metal chlorides and metal oxides.
3. Usually react with dilute hydrochloric acid or sulphuric acid to give hydrogen and a salt.
4. React with oxygen to form basic oxides or amphoteric oxides.

However, one metal may react more or less vigorously with a substance than another metal. The metal that reacts more vigorously is said to be more **reactive** than the other metal.

It is useful to arrange metals according to how reactive they are. As such, scientists have worked out a list called the **reactivity series**. In the reactivity series, *metals are listed from the most reactive to the least reactive.*

How is the order of reactivity determined?
Let us determine the reactivity of metals by looking at how they react with water, steam and hydrochloric acid in the laboratory.

The Reaction of Metals with Water

Which metals react with cold water?
Some metals react with cold water to form the metal hydroxide and hydrogen gas.

metal + water $\longrightarrow$ metal hydroxide + hydrogen

Fig. 14.5 and Fig. 14.6 show what happens when pieces of different metals are dropped into a beaker of cold water.

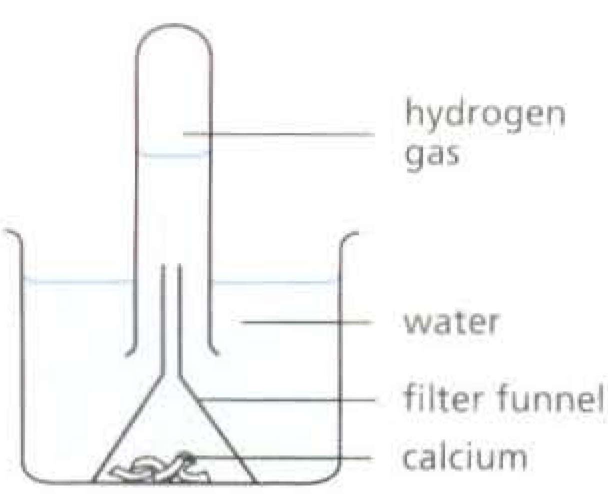


Fig. 14.5 Bubbles of gas are produced. This indicates that calcium reacts readily with cold water.

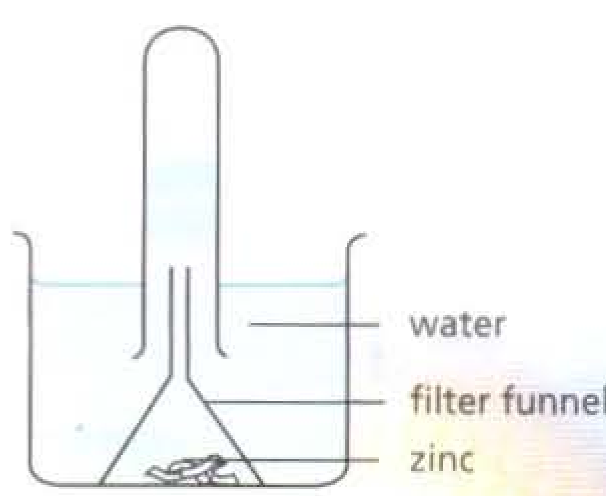


Fig. 14.6 No gas bubbles are produced. Zinc does not react with cold water.

This is a metal part of a ship that sank over a hundred years ago. Why has it worn away?

Table 14.2 gives the observations and chemical equations for the reactions of different metals with cold water.

Metal(s)	Observations	Equation
Potassium	Reacts very violently to form potassium hydroxide and hydrogen gas. Enough heat is produced to ignite the hydrogen gas produced. The hydrogen gas burns with a lilac flame.	$2K(s) + 2H_2O(l) \longrightarrow 2KOH(aq) + H_2(g)$
Sodium	Reacts violently to form sodium hydroxide and hydrogen gas. The hydrogen gas formed may catch fire and burn with a yellow flame.	$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$
Calcium	Reacts readily to form calcium hydroxide and hydrogen gas.	$Ca(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(aq) + H_2(g)$
Magnesium	Reacts very slowly to form magnesium hydroxide and hydrogen gas. A test tube of hydrogen gas is produced only after a few days.	$Mg(s) + 2H_2O(l) \longrightarrow Mg(OH)_2(s) + H_2(g)$
Zinc Iron* Lead Copper Silver	No reaction occurs.	

*Iron reacts with water very slowly in the presence of air, in a process called rusting. See section 14.5.

Table 14.2 Reactions of metals with cold water

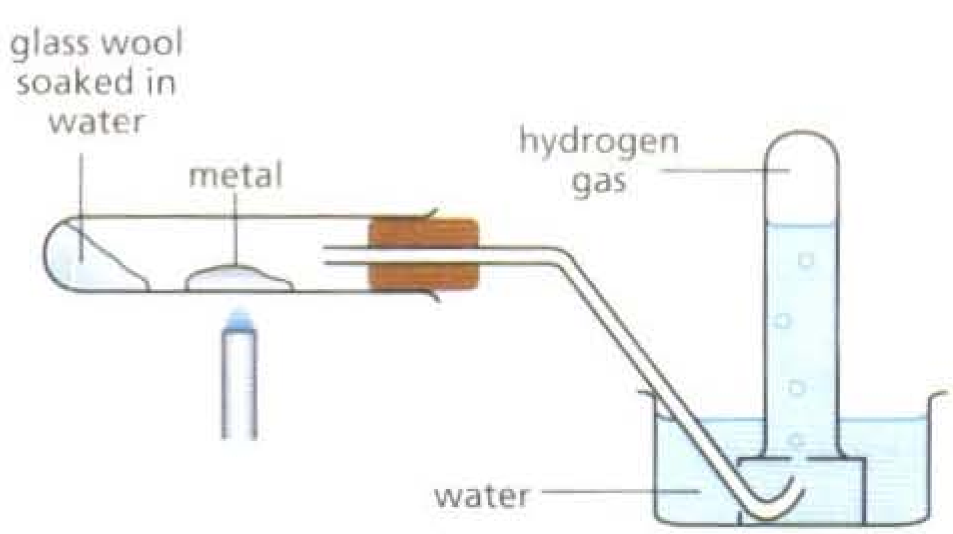
As you can see from how vigorous the reactions are, the most reactive metals are potassium, followed by sodium and then calcium.

Link

If metals such as potassium, sodium and calcium react with cold water, we can expect them to have even more violent reactions with hot water or steam. Find out why a chemical reaction is affected by temperature in chapter 18.

Which metals only react with steam?

Zinc and iron do not react with cold water but they do react with steam. Fig. 14.7 shows the apparatus used for reacting a metal with steam.



To prevent water from being sucked back into the hot test tube, the delivery tube must be removed from the water in the trough before turning off the Bunsen flame.

Fig. 14.7 Reacting a metal with steam

The metal is strongly heated until it is very hot. The glass wool soaked in water is also heated to generate a flow of steam over the hot metal.

Metals react with steam to form the metal oxide and hydrogen gas.

metal + steam $\longrightarrow$ metal oxide + hydrogen

Table 14.3 shows the observations and chemical equations for the reactions of some metals with steam.

Metal(s)	Observations	Equation
Magnesium	Hot magnesium reacts violently with steam to form magnesium oxide (a white powder) and hydrogen gas. A bright white glow is produced during the reaction.	$Mg(s) + H_2O(g) \longrightarrow MgO(s) + H_2(g)$
Zinc	Hot zinc reacts readily with steam to produce zinc oxide and hydrogen gas. Zinc oxide is yellow when hot and white when cold.	$Zn(s) + H_2O(g) \longrightarrow ZnO(s) + H_2(g)$
Iron	Red-hot iron reacts slowly with steam to form iron oxide and hydrogen gas. The iron must be heated constantly in order for the reaction to proceed.	$3Fe(s) + 4H_2O(g) \longrightarrow Fe_3O_4(s) + 4H_2(g)$
Lead Copper Silver	No reaction occurs.	

Table 14.3 Reactions of some metals with steam

From the experimental results shown in Tables 14.2 and 14.3, we can conclude that

- potassium, sodium and calcium are highly reactive metals.
- magnesium, zinc and iron are fairly reactive metals.
- lead, copper and silver are unreactive metals.
- reactive metals tend to react to form compounds, while unreactive metals tend to remain as metals.
- the reactivity of metals is in the order shown in Fig. 14.8.

Quick Check

Explain why coins made of zinc look dull after some time but silver coins remain bright and shiny for years.

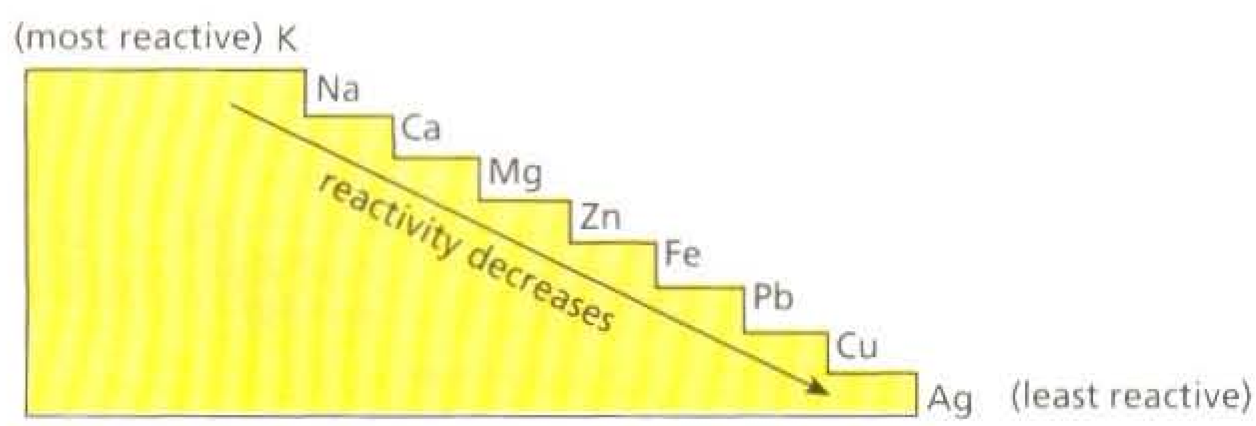


Fig. 14.8 Order of reactivity of metals

Link

Do metals react with dilute sulphuric acid in the same way they react with dilute hydrochloric acid? Recall what you have learnt in chapter 11.

The Reaction of Metals with Dilute Hydrochloric Acid

Many metals react with dilute acids to produce a salt and hydrogen.

$$\text{metal} + \text{dilute acid} \longrightarrow \text{salt} + \text{hydrogen}$$

When a metal reacts with dilute hydrochloric acid, the products are the metal chloride and hydrogen gas.

$$\text{metal} + \text{dilute hydrochloric acid} \longrightarrow \text{metal chloride} + \text{hydrogen}$$

The reactions of different metals with dilute hydrochloric acid also indicate how reactive the metals are.

Table 14.4 summarises the reactions of metals with dilute hydrochloric acid.

Metal(s)	Observations	Equation
Potassium Sodium	Explosive reaction. These reactions should not be carried out in the school laboratory.	$2K(s) + 2HCl(aq) \longrightarrow 2KCl(aq) + H_2(g)$ $2Na(s) + 2HCl(aq) \longrightarrow 2NaCl(aq) + H_2(g)$
Calcium	Reacts violently to give hydrogen gas.	$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$
Magnesium	Reacts rapidly to give hydrogen gas.	$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$
Zinc	Reacts moderately fast to give hydrogen gas.	$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$
Iron	Reacts slowly to give hydrogen gas.	$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$
Lead Copper Silver	No reaction occurs.	

Table 14.4 Reactions of metals with dilute hydrochloric acid

The experimental results shown in Table 14.4 confirm that the reactivity of metals is in the order shown by the reactivity series below (Fig. 14.9).

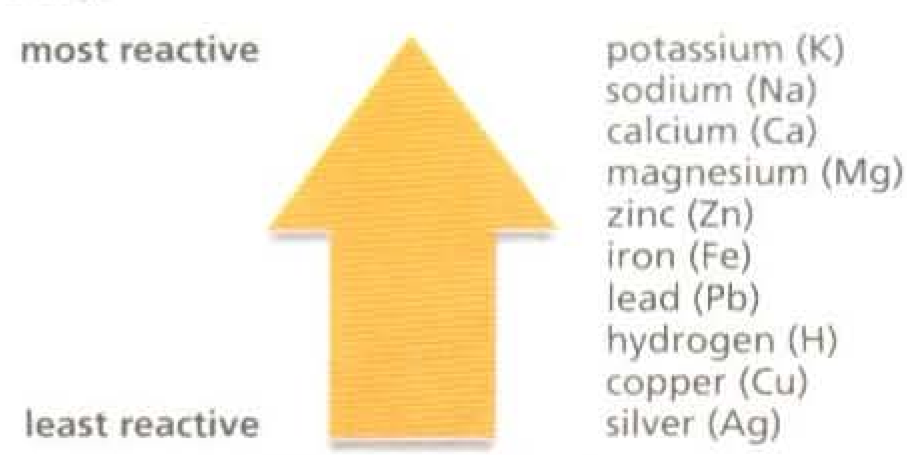


Fig.14.9 Reactivity series

The reactivity series also includes non-metals such as hydrogen. Hydrogen is placed between lead and copper. Only metals that are more reactive than hydrogen react with dilute acids to produce hydrogen gas.

The Reduction of Metal Oxides with Carbon

We have compared the reactivity of metals by studying their reactions with water and hydrochloric acid. We can also compare the reactivity of different metals by studying how easily metal oxides decompose. The more reactive a metal is, the more difficult it is to decompose its oxides — reduce the oxide to the metal.

When a mixture of copper(II) oxide and carbon is heated (Fig. 14.10), copper(II) oxide is reduced to copper and carbon is oxidised to carbon dioxide. The equation for the reaction is

$$\underset{\text{(black solid)}}{\text{copper(II) oxide}} + \text{carbon} \xrightarrow{\text{heat}} \underset{\text{(reddish-brown solid)}}{\text{copper}} + \text{carbon dioxide}$$

$$2CuO(s) + C(s) \xrightarrow{\text{heat}} 2Cu(s) + CO_2(g)$$

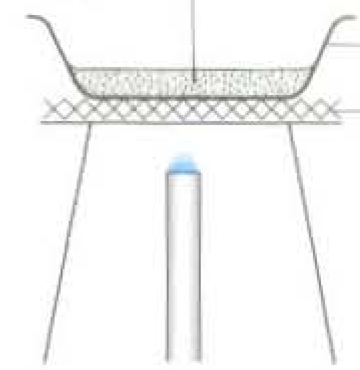


Fig. 14.10 Apparatus used to reduce copper(II) oxide with carbon

If the experiment is repeated with magnesium oxide, no reaction takes place. The reactions of some metal oxides with carbon are shown in Table 14.5.

Metal oxide	Reaction
Potassium oxide (K_2O) Sodium oxide (Na_2O) Calcium oxide (CaO) Magnesium oxide (MgO)	Oxides are not reduced by carbon.
Zinc oxide (ZnO) Iron(II) oxide (FeO) Lead(II) oxide (PbO) Copper(II) oxide (CuO)	Oxides are reduced by carbon.
Silver oxide (Ag_2O)	Oxide is reduced by heating.

Table 14.5 Reduction of metal oxides with carbon

If a metal is below copper in the reactivity series, its oxide will decompose simply by heating, without the need for a reducing agent such as carbon. For example, silver oxide decomposes into silver and oxygen when it is heated to 200 °C as shown in this equation:

$$\text{silver oxide} \xrightarrow{\text{heat}} \text{silver} + \text{oxygen}$$

$$2Ag_2O(s) \xrightarrow{\text{heat}} 4Ag(s) + O_2(g)$$

What is the importance of this reaction in industry?
Metals need to be extracted from their ores, which are compounds found naturally. These ores are mainly made up of metal oxides. In industry, metals that are below magnesium in the reactivity series are often extracted from their ores by reduction with carbon.

Zinc and lead are extracted by reduction with carbon, as represented by the equations below:

zinc oxide + carbon ⟶ zinc + carbon monoxide

$$ZnO(s) + C(s) \longrightarrow Zn(s) + CO(g)$$

lead(II) oxide + carbon ⟶ lead + carbon monoxide

$$PbO(s) + C(s) \longrightarrow Pb(s) + CO(g)$$

Iron is also extracted from its ore by reduction with carbon. We shall learn the details of this process in section 14.3.

The oxides of metals that are above zinc in the reactivity series are not reduced by carbon. These oxides are so stable that they can only be reduced by passing electricity through them.

The Reduction of Metal Oxides with Hydrogen

Besides carbon, hydrogen can also be used for reducing metal oxides to metals. Fig. 14.11 shows the apparatus used for the reduction of a metal oxide to the metal using hydrogen.

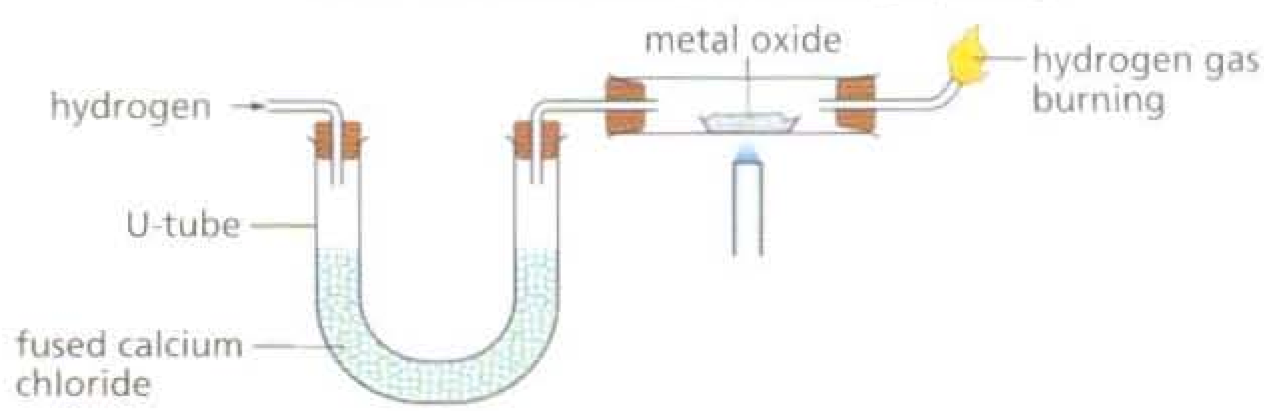


Fig. 14.11 The reduction of a metal oxide with a stream of dry hydrogen gas

Hydrogen gas is passed over the metal oxide. It acts as a reducing agent. Hydrogen reduces oxides of some metals, such as iron, copper and silver, to the metals as shown by the equation:

metal oxide + hydrogen ⟶ metal + steam

Oxides of reactive metals such as potassium, sodium, calcium, magnesium and zinc, are not reduced by hydrogen. Table 14.6 summarises the reactions of metal oxides with hydrogen.

Metal oxide	Reaction with hydrogen
Potassium oxide (K_2O) Sodium oxide (Na_2O) Calcium oxide (CaO) Magnesium oxide (MgO) Zinc oxide (ZnO)	Heated metal oxides are not reduced.
Iron(II) oxide (FeO) Lead(II) oxide (PbO) Copper(II) oxide (CuO) Silver oxide (Ag_2O)	Heated metal oxides are reduced.

Table 14.6 The reduction of metal oxides with hydrogen

Key ideas

1. A summary of the reactions of metals with water and dilute acids:

Metal(s)	Reaction with water or steam	Reaction with dilute acid
potassium sodium	react with cold water and steam	explosive reaction
calcium	reacts with cold water and steam	violent reaction
magnesium zinc	react with steam	moderately fast reaction
iron	reacts with steam	slow reaction
lead copper silver	no reaction with water or steam	no reaction

2. Metals are listed in the reactivity series from the most reactive to the least reactive.

Reactivity series

most reactive
K
Na
Ca
Mg
Zn
Fe
Pb
(H)
Cu
Ag
least reactive

3. Oxides of metals above zinc cannot be reduced by heating with carbon.
4. Oxides of metals above iron cannot be reduced by heating with hydrogen.

Test Yourself 14.2

Worked Example

Aluminium (Al) is above zinc but below magnesium in the reactivity series. Predict the reaction between aluminium and hydrochloric acid.

Thought Process

Order of reactivity: magnesium > aluminium > zinc. Since magnesium and zinc react rapidly with hydrochloric acid to form their metal chlorides and hydrogen, aluminium will undergo the same reaction.

Answer

Aluminium will react rapidly with hydrochloric acid to form aluminium chloride ($AlCl_3$) and hydrogen.

Questions

1. Manganese (Mn) is between magnesium and iron in the reactivity series. It reacts with steam to form manganese(II) oxide. Write the equation for the reaction.

2. Why are hot water tanks made of copper and not steel?

3. a) *X* is a metal which does not react with cold water but reacts with steam to form a white oxide. Which **two** metals could be *X*?
   b) The reaction between *X* and dilute hydrochloric acid is moderately fast at room temperature. From this observation, deduce the identity of *X*.

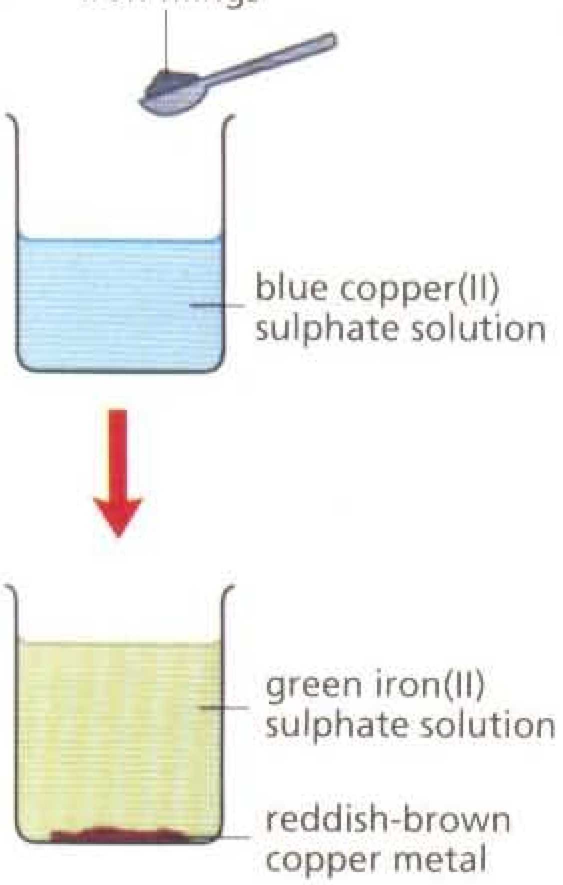


Fig. 14.12 Iron displaces copper from copper(II) sulphate solution.

The Displacement Reactions of Metals

More reactive metals can displace less reactive metals from their salt solutions. For example, solid iron displaces copper ions from a solution of copper(II) sulphate (Fig. 14.12). Due to the action of iron, copper metal is precipitated out of the solution as a pink or a reddish-brown solid.

The reaction can be represented by the equation below:

$$\text{iron} + \text{copper(II) sulphate} \longrightarrow \text{iron(II) sulphate} + \text{copper}$$

$$Fe(s) + CuSO_4(aq) \longrightarrow FeSO_4(aq) + Cu(s)$$

The ionic equation is

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

In other words, atoms of the more reactive metal become ions and form compounds while ions of the less reactive metal change back to atoms. The following experiment investigates more of these displacement reactions.

Chem-Aid

1. Copper often appears pink when it is freshly formed (displaced).
2. Iron(II) sulphate is prepared using excess acid so that it does not break down easily.

Experiment 1

To investigate the displacement of metals from their solutions by another metal.

Procedure

1. Solutions of the following salts are prepared in separate test tubes.
   - Copper(II) sulphate
   - Lead(II) nitrate
   - Zinc sulphate
   - Iron(II) sulphate
   - Magnesium sulphate

2. A clean strip of copper is placed into each solution. What happens then is recorded.

3. The experiment is repeated with clean strips of lead, zinc, magnesium and iron nails using fresh salt solutions.

Table 14.7 shows some of the observations for the experiment.

Metal \ Salt solution	Magnesium sulphate	Zinc sulphate	Copper(II) sulphate	Lead(II) nitrate	Iron(II) sulphate
Magnesium		Solution remains colourless. Grey deposit of zinc formed on magnesium. $Mg(s) + ZnSO_4(aq) \longrightarrow Zn(s) + MgSO_4(aq)$	Blue solution turns colourless. Reddish-brown deposit of copper formed on magnesium. $Mg(s) + CuSO_4(aq) \longrightarrow Cu(s) + MgSO_4(aq)$	Solution remains colourless. Grey deposit of lead formed on magnesium. $Mg(s) + Pb(NO_3)_2(aq) \longrightarrow Pb(s) + Mg(NO_3)_2(aq)$	Pale green solution turns colourless. Grey deposit of iron formed on magnesium. $Mg(s) + FeSO_4(aq) \longrightarrow Fe(s) + MgSO_4(aq)$
Zinc	No reaction.		Blue solution turns colourless. Reddish-brown deposit of copper formed on zinc. $Zn(s) + CuSO_4(aq) \longrightarrow Cu(s) + ZnSO_4(aq)$	Solution remains colourless. Grey deposit of lead formed on zinc. $Zn(s) + Pb(NO_3)_2(aq) \longrightarrow Pb(s) + Zn(NO_3)_2(aq)$	Pale green solution turns colourless. Grey deposit of iron formed on zinc. $Zn(s) + FeSO_4(aq) \longrightarrow Fe(s) + ZnSO_4(aq)$
Copper	No reaction.	No reaction.		No reaction.	No reaction.
Lead	No reaction.	No reaction.	Blue solution turns colourless. Reddish-brown deposit of copper formed on lead. $Pb(s) + CuSO_4(aq) \longrightarrow Cu(s) + PbSO_4(aq)$		No reaction
Iron	No reaction.	No reaction.	Blue solution turns pale green. Reddish-brown deposit of copper formed on iron. $Fe(s) + CuSO_4(aq) \longrightarrow Cu(s) + FeSO_4(aq)$	Colourless solution turns pale green. Grey deposit of lead formed on iron. $Fe(s) + Pb(NO_3)_2(aq) \longrightarrow Pb(s) + Fe(NO_3)_2(aq)$	

Table 14.7 Observations for some displacement reactions

From the results in Table 14.7, you can see that a *metal higher up in the reactivity series will displace a metal that is lower in the series from its salt solution.*

Metal displacement reactions are redox reactions

In a metal displacement reaction, the more reactive metal is oxidised while the less reactive metal is reduced. Thus, a displacement reaction is a **redox reaction**.

Link

Displacement reactions also take place between non-metals, for example, between the halogens in Group VII of the Periodic Table. Find out more in chapter 16.

For example, in the reaction between iron and copper(II) sulphate, the more reactive metal, iron, reduces the copper(II) ions to copper. At the same time, iron itself is oxidised to the iron(II) ions. The ionic equation for the reaction is

Copper(II) ions are reduced to become copper atoms.

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

Iron atoms are oxidised to become iron(II) ions.

A similar redox reaction occurs when a piece of magnesium is dipped into a copper(II) salt solution to displace copper. The more reactive metal, magnesium, reduces the copper(II) ions to copper. Magnesium itself is oxidised to the magnesium ions.

Copper(II) ions are reduced to become copper atoms.

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$

Magnesium atoms are oxidised to become magnesium ions.

On the other hand, if a piece of copper foil is dipped into magnesium sulphate solution, no reaction occurs.

no change

$$Cu(s) + Mg^{2+}(aq) \longrightarrow Cu(s) + Mg^{2+}(aq)$$

no change

You can see from the above ionic equations that a metal higher up in the reactivity series has a greater tendency to form positive ions. This is why a more reactive metal can displace a less reactive metal from its salt solution. Copper is less reactive than magnesium. It has a lower tendency to form ions compared to magnesium. Thus, it cannot displace magnesium from its salt solution.

Reaction Between a Metal and the Oxide of Another Metal

A more reactive metal has a higher tendency to form its positive ions compared to a less reactive metal. This is why a more reactive metal can reduce the oxide of a less reactive metal. An example is the reaction between zinc and copper(II) oxide. Fig. 14.13 shows the apparatus used to study the reaction.

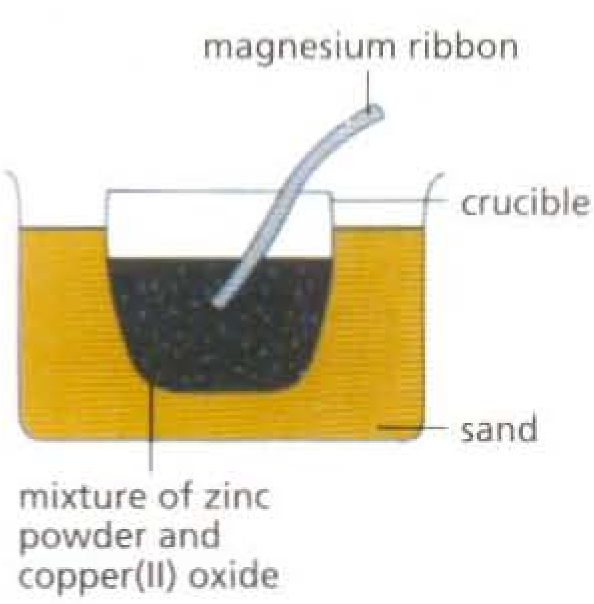


Fig. 14.13 Reaction between zinc and copper(II) oxide

The magnesium ribbon acts as a fuse. When the magnesium ribbon is ignited, it provides enough energy to start the reaction between zinc and copper(II) oxide to form zinc oxide and copper.

Since zinc is more reactive than copper, it will form positive ions more readily. Zn^{2+} ions react with O^{2-} ions from copper(II) oxide to form zinc oxide. In the process, copper(II) oxide is reduced to copper.

zinc + copper(II) oxide ⟶ zinc oxide + copper

$$Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$$

The more reactive the metal is, the more readily it forms compounds. On the other hand, unreactive metals tend to stay uncombined. Therefore, if the experiment shown in Fig. 14.13 is repeated using a mixture of zinc oxide and copper powder, no reaction will occur.

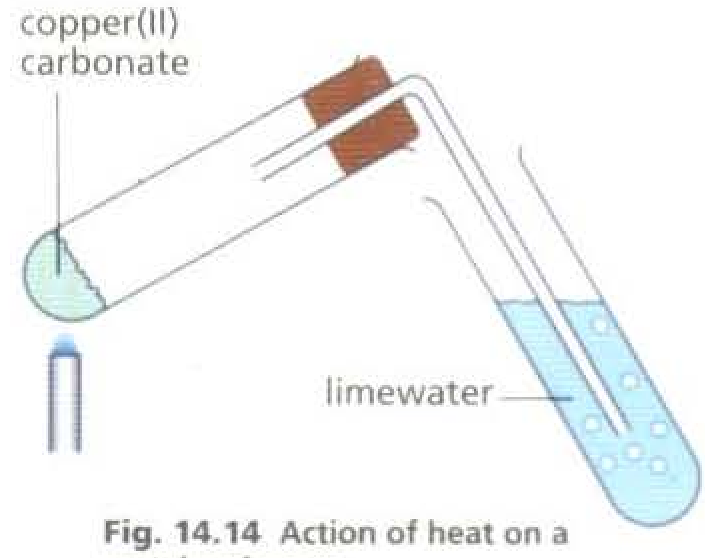


Fig. 14.14 Action of heat on a metal carbonate

Action of Heat on Metal Carbonates

Some compounds are more difficult to decompose by heat than others. This means that these compounds are more stable to heat than others. The thermal stability of metal carbonates can be tested by heating them in a dry test tube (Fig. 14.14). Table 14.8 summarises the action of heat on some metal carbonates.

Metal carbonate	Observation
Potassium carbonate, K_2CO_3 Sodium carbonate, Na_2CO_3	unaffected by heat
Calcium carbonate, $CaCO_3$ Magnesium carbonate, $MgCO_3$ Zinc carbonate, $ZnCO_3$ Iron(II) carbonate, $FeCO_3$ Lead(II) carbonate, $PbCO_3$ Copper(II) carbonate, $CuCO_3$	decompose into metal oxide and carbon dioxide on heating
Silver carbonate, Ag_2CO_3	decomposes into silver and carbon dioxide on heating

Table 14.8 Action of heat on metal carbonates

You can see that the thermal stability of the metal carbonates is related to the position of the metal in the reactivity series. *The more reactive the metal is, the more difficult it is to decompose its compounds.* Therefore, sodium and potassium carbonates are not affected by heat.

The carbonates of metals below sodium in the reactivity series decompose to form the oxides of the metals and carbon dioxide. In the case of silver carbonate, the silver oxide produced is thermally unstable. It further decomposes to form silver.

Using the Reactivity Series

The reactivity series is useful for

1. predicting the behaviour of a metal from its position in the reactivity series,
2. predicting the position of an unfamiliar metal in the reactivity series from a given set of experimental results.

Example 1

Tin is below iron but above lead in the reactivity series. Predict the reaction between tin(II) oxide, SnO, and

a) carbon.
b) magnesium.

Solution:

a) Order of reactivity: iron > tin > lead. Since the oxides of both iron and lead can be reduced by carbon, tin(II) oxide will undergo the same reaction. Tin(II) oxide will be reduced by carbon on heating to form tin and carbon monoxide.

b) Magnesium is more reactive than tin. It has a higher tendency to form its positive ions (and compounds). Thus, it will reduce tin(II) oxide to tin and form magnesium oxide at the same time. Tin(II) oxide will react with magnesium to form magnesium oxide and tin.

Example 2

When a piece of chromium is placed in zinc sulphate solution, no reaction occurs. When a mixture of chromium powder and iron(III) oxide is heated strongly, a reaction takes place. Deduce the position of chromium in the reactivity series.

Solution:

Since chromium does not react with zinc sulphate solution, it must be less reactive than zinc, i.e. below zinc in the reactivity series. Since it reacts with iron(III) oxide, it must be more reactive than iron. Thus, chromium is below zinc but above iron in the reactivity series.

Knowing the reactivity series also helps us to choose suitable methods for extracting metals and to prevent iron from rusting.

Key ideas

A metal that is higher up in the reactivity series has a greater tendency to form positive ions than a metal lower in the series.

- A metal that is higher up in the reactivity series will displace a metal that is lower in the series from the salt solution or oxide of the less reactive metal.
- A more reactive metal will form compounds more readily than a less reactive metal.
- Carbonates of metals higher up in the reactivity series are more thermally stable than carbonates of metals lower in the reactivity series.

Test Yourself 14.3

Questions

1. The following experiments were carried out on the metals *W*, *X*, *Y* and *Z*.
   - *Z* displaces *W* from the salt solution of the metal *W*.
   - *W* displaces *X* from the salt solution of the metal *X*.
   - *X* displaces *Y* from the salt solution of the metal *Y*.

   What is the order of reactivity of these metals, placing the most reactive metal first?

   A $W > X > Y > Z$
   B $X > W > Y > Z$
   C $Y > W > X > Z$
   D $Z > W > X > Y$

2. Describe the colour change in copper(II) carbonate when it is heated strongly. What happens when dilute sulphuric acid is added to the heated salt? Write balanced equations for the reactions.

14.3 | Extracting Metals

Only a small number of metals (the very unreactive ones), such as gold and platinum, occur freely in nature as uncombined metals. Most metals react with other elements to form **ores**. Obtaining metals from their ores generally involves three major stages:

1. Concentrating the ore
2. Extracting crude metal from the ore
3. Refining crude metal to obtain the pure metal (see chapter 15)

Fig. 14.15 The stages involved in obtaining metals from their ores

Concentrating metal ores

An **ore** is *a compound of the metal* (usually the oxides, sulphides, chlorides or carbonates) *mixed with large amounts of earth and rock*. Earth and rock are removed before metal is extracted from the ore. This results in a metal ore that contains little waste material.

Extracting metal from ores

There are two main methods for extracting metals from their ores:

1. Reducing the metal compound (ore) to the metal using carbon
2. Using electricity to decompose the molten metal compound (ore) to the metal

The position of a metal in the reactivity series determines the method used for its extraction. The methods used for extracting some common metals are summarised below.

Chem-Aid

Non-metals such as hydrogen and carbon are also listed in the reactivity series. Carbon is placed above zinc and below aluminium. Therefore, carbon can reduce the compounds of zinc and compounds of metals below zinc in the reactivity series.

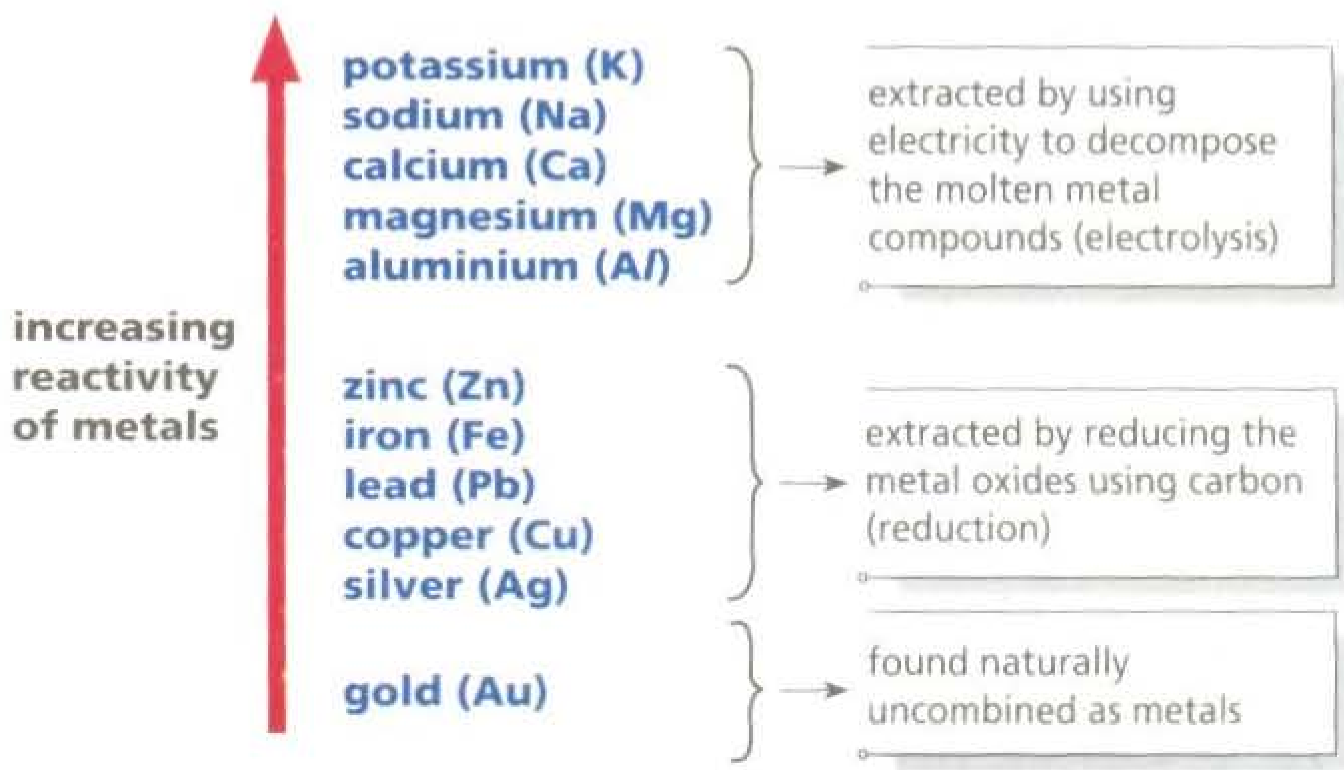


Fig. 14.16 Methods of extracting metals based on their reactivity

Try it Out

The five most common metals in the Earth's crust are aluminium, iron, magnesium, manganese and titanium. Find out the names of the ores from which magnesium, manganese and titanium are extracted.

In general, the more reactive the metal is, the harder it is to extract the metal from its ore. Reactive metals such as sodium, potassium, calcium, magnesium and aluminium cannot be extracted by reduction with carbon. The compounds of these metals are very difficult to split up. These metals are extracted using electricity, in a process called electrolysis (see chapter 15).

The metals placed in the middle of the reactivity series, such as zinc and iron, are not so reactive. They are readily extracted by reducing their oxides with carbon. The metals lowest in the reactivity series, such as gold, can be found in nature as uncombined elements.

Extracting Iron from Haematile

The main ore of iron is **haematite**. Haematite contains iron(III) oxide mixed with impurities such as sand and clay. Iron is extracted from haematite in a **blast furnace** (Fig. 14.17). Haematite, coke (which is mainly carbon) and limestone (calcium carbonate) are added at the top of the blast furnace. Blasts of hot air are blown into the furnace near the bottom.

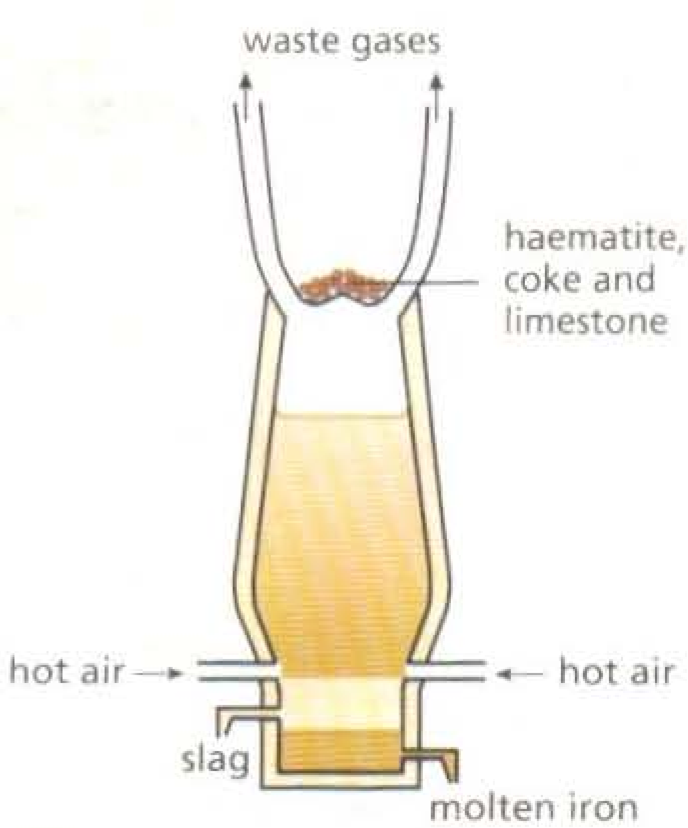


Fig. 14.17 A simple diagram of a blast furnace

Fig. 14.18 is a simple flow chart of what goes on in a blast furnace.

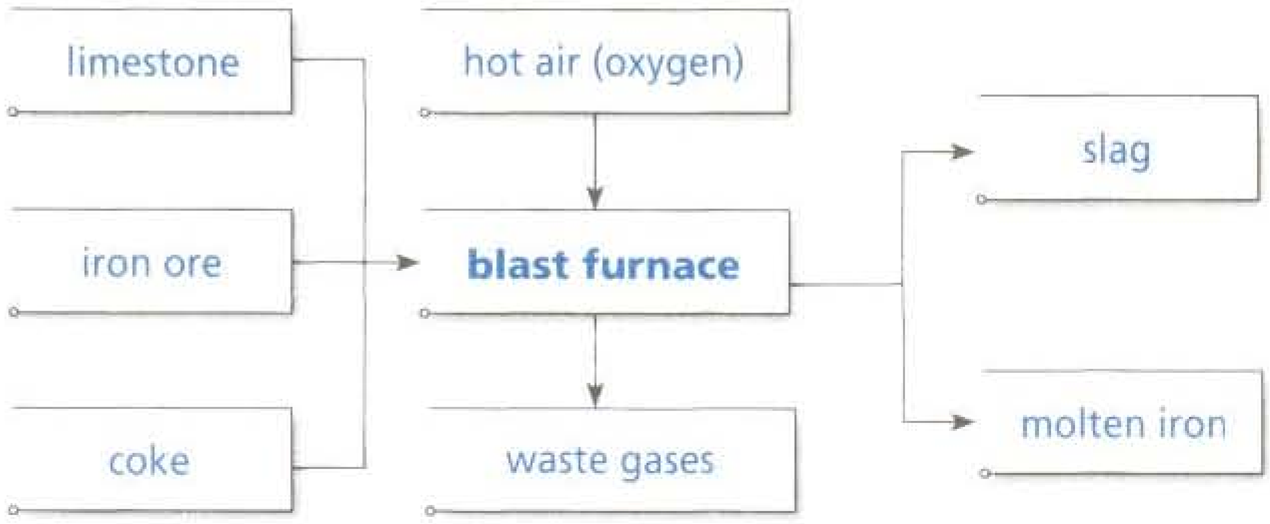


Fig. 14.18 A simple flow chart for extracting iron in a blast furnace

What chemical reactions take place in the blast furnace?
In the blast furnace, a series of chemical reactions takes place.

1. **Carbon dioxide is produced.**
   The carbon in coke burns in a blast of hot air to produce carbon dioxide. This reaction produces a lot of heat.

   $$\text{carbon} + \text{oxygen} \xrightarrow{\text{heat}} \text{carbon dioxide}$$

   $$C(s) + O_2(g) \xrightarrow{\text{heat}} CO_2(g)$$

   The limestone (calcium carbonate) is decomposed by heat to produce carbon dioxide and calcium oxide (quicklime).

   $$\text{calcium carbonate} \xrightarrow{\text{heat}} \text{calcium oxide} + \text{carbon dioxide}$$

   $$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

2. **Carbon monoxide is produced.**
   As the carbon dioxide rises up the furnace, it reacts with more coke to form carbon monoxide.

   $$\text{carbon} + \text{carbon dioxide} \xrightarrow{\text{heat}} \text{carbon monoxide}$$

   $$C(s) + CO_2(g) \xrightarrow{\text{heat}} 2CO(g)$$

3. **Haematite is reduced to iron.**
   The carbon monoxide reduces the iron(III) oxide in haematite to iron.

   $$\text{iron(III) oxide} + \text{carbon monoxide} \xrightarrow{\text{heat}} \text{molten iron} + \text{carbon dioxide}$$

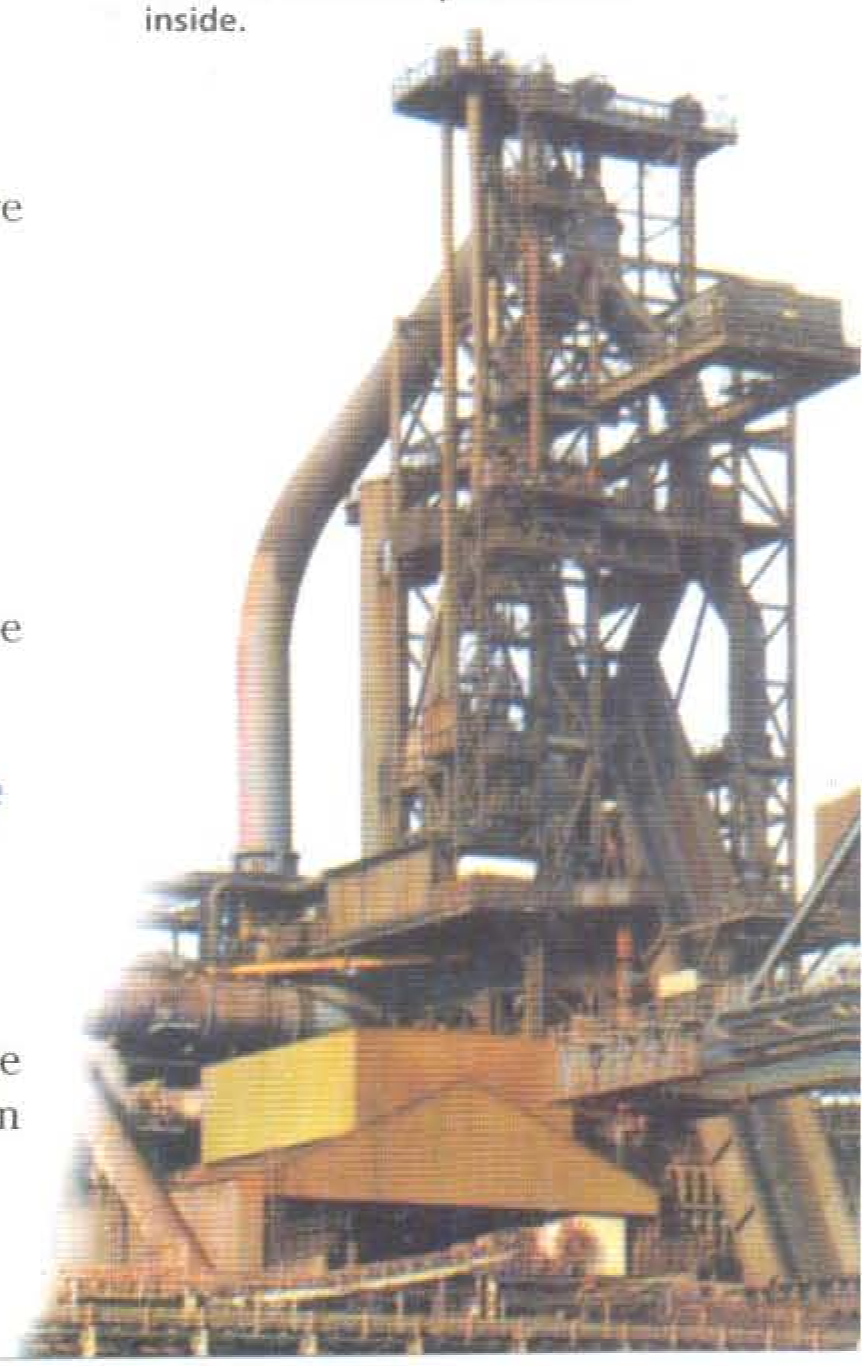
   $$Fe_2O_3(s) + 3CO(g) \xrightarrow{\text{heat}} 2Fe(l) + 3CO_2(g)$$

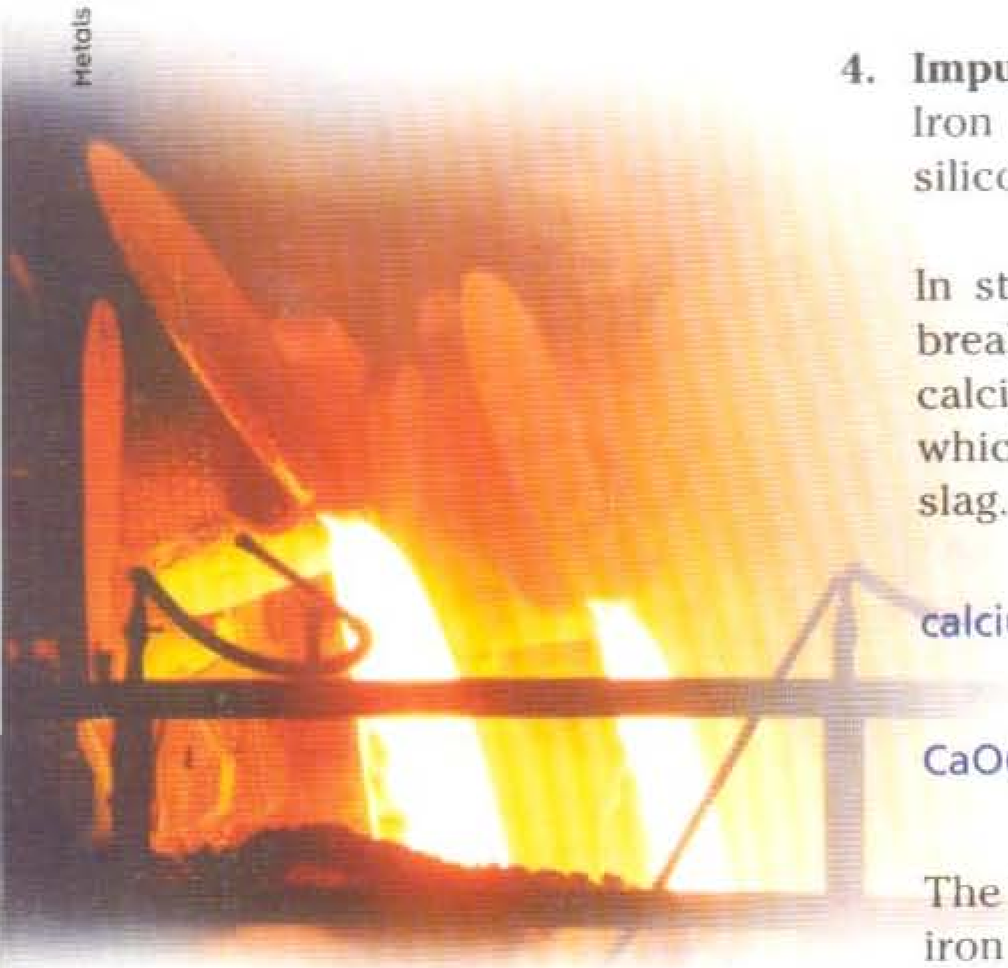
   The iron formed is molten and runs to the bottom of the furnace. Hot waste gases containing carbon monoxide, carbon dioxide and nitrogen escape through the top of the furnace.

Chem-Aid

Upon heating, calcium carbonate breaks down into calcium oxide and carbon dioxide. This process is called thermal decomposition.

A blast furnace can be 50 – 70 m high. It is constructed from steel and lined with fireproof bricks inside.




Dumping blast furnace slag

4. **Impurities are removed.**
Iron ore contains the impurities sand and clay, which are silicon oxides. Limestone is added to remove these impurities.

In step 1, you saw that the limestone (calcium carbonate) breaks down to form calcium oxide and carbon dioxide. The calcium oxide, a basic oxide, reacts with silicon(IV) oxide which is acidic, and with other impurities, to form a molten slag.

$$\text{calcium oxide} + \text{silicon(IV) oxide} \xrightarrow{\text{heat}} \underset{\text{(slag)}}{\text{calcium silicate}}$$

$$CaO(s) + SiO_2(s) \xrightarrow{\text{heat}} CaSiO_3(l)$$

The lighter slag floats on top of the molten iron. The slag and iron are tapped off separately at the bottom of the furnace. Solidified slag is mainly used for road surfacing.

Key ideas

1. The method used to extract a metal from its ore depends on the position of the metal in the reactivity series.
   - Metals higher up in the series need to be extracted using electricity.
   - Metals lower in the series can be extracted by reduction with carbon.
2. Iron is extracted from haematite (Fe_2O_3) in the blast furnace by reduction with coke (carbon). Limestone (calcium carbonate) is added to remove impurities.

Test Yourself 14.4

Worked Example

Which substance is a reducing agent in the blast furnace for the production of iron?

A Air
B Calcium carbonate
C Carbon monoxide
D Silicon(IV) oxide

Thought Process

A reducing agent removes oxygen from another substance. Carbon monoxide removes oxygen from iron(III) oxide (in haematite) to form iron.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

Answer

C

Questions

1. The positions of barium and tin in the reactivity series are shown below.

   barium
   calcium
   iron
   tin
   copper

   increasing reactivity

   Suggest the method that would be most suitable for extracting these metals from their respective ores
   a) Barium
   b) Tin

2. Name the following substances present in the blast furnace during the extraction of iron from haematite.
   a) **Four** ionic compounds
   b) **Two** reducing agents

14.4 | The Uses of Iron and Steel

The iron that is extracted from the blast furnace is known as cast iron (or pig iron). Most of it goes into producing steel.

What is steel?
Steel is an alloy of iron with carbon and/or other metals. There are many types of steel. Depending on the amount of carbon and other metals added, and the type of metals added to iron, each type of steel has its own special properties and uses.

How is steel made from cast iron?
Steel-making involves oxidation, followed by alloying. The impurities in iron are removed by passing pure oxygen through molten cast iron. Carbon and small amounts of other metals are then added to produce different types of steel with special properties (Fig. 14.19).

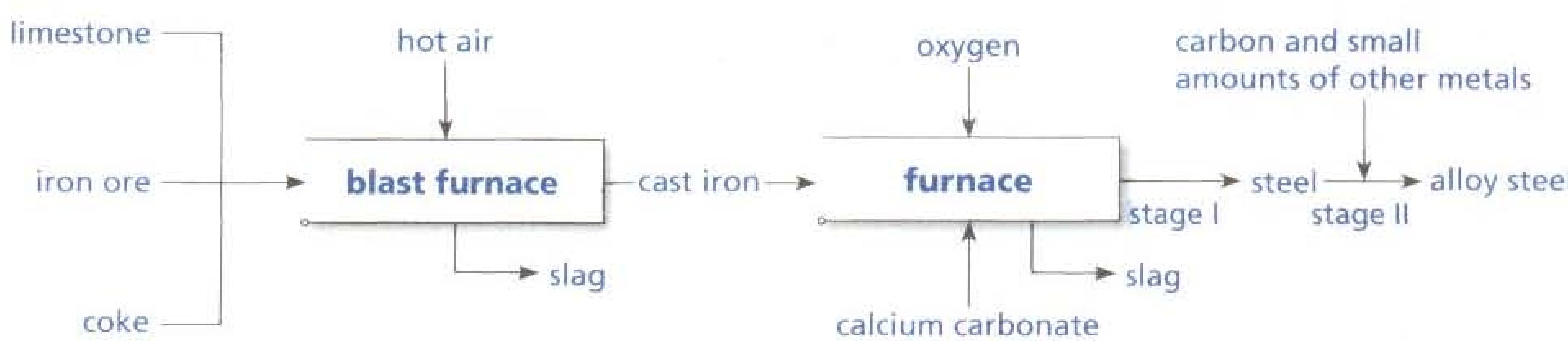


Fig. 14.19 Making steel from iron ore

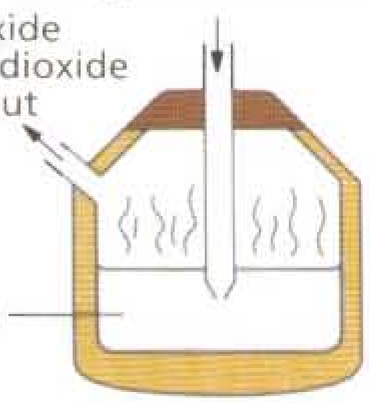


Fig. 14.20 Purifying cast iron

Quick Check

Most metals can be recycled. State three reasons why the recycling of metals is important in modern society.

1. **Removing impurities by oxidation**
   The common impurities in cast iron are non-metals such as carbon, sulphur, phosphorus and silicon. To remove these impurities, cast iron is first melted and poured into a large container (Fig. 14.20). Pure oxygen at high pressure is then blown through the molten iron to oxidise the impurities. The impurities form gases, such as sulphur dioxide and carbon dioxide, that are blown out of the furnace.

   Calcium carbonate is later added to the furnace. At high temperatures, calcium carbonate decomposes to form calcium oxide which reacts with silicon(IV) oxide and phosphorus oxide to form slag.

   The slag can be poured off. The molten metal left behind is pure iron, also known as wrought iron. Wrought iron is soft and bends easily. It is used for making garden gates and chains.

2. **Adding other elements to make the various types of steel**
   Carbon and other metals are now stirred into the molten iron in the correct amounts to make steel with special properties.

What are the uses of steel?

Although there are many different types of steel, it is convenient to group them into two main categories — carbon steels and alloy steels.

Carbon steels

These are general purpose steels which contain mainly iron and carbon. Mild steel (low carbon steel) contains 99.5% iron and up to 0.25% carbon. It is strong and malleable. It is used to make car bodies and machinery.

Car bodies are made of mild steel.

High carbon steel contains 0.45 – 1.5% carbon. It is strong but brittle. High carbon steel is used to make knives, hammers, chisels, saws and other cutting and boring tools.

Alloy steels

Alloy steels consist of iron and carbon together with one or more of the following metals: manganese (Mn), nickel (Ni), chromium (Cr), tungsten (W) and vanadium (V). The reason for adding these metals is to change the properties of the steel.

For example, in manganese steel used for making springs and drills, manganese is added to increase its strength and hardness. Nickel and chromium are added to increase resistance to corrosion.

Stainless steel is an alloy of iron, chromium, nickel and a little carbon. It does not rust and is, therefore, used extensively for making cutlery and surgical instruments. It is also used in chemical plants because it is extremely durable and is resistant to corrosion even when heated.

Stainless steel is used to make surgical tools because it does not rust.

Key ideas

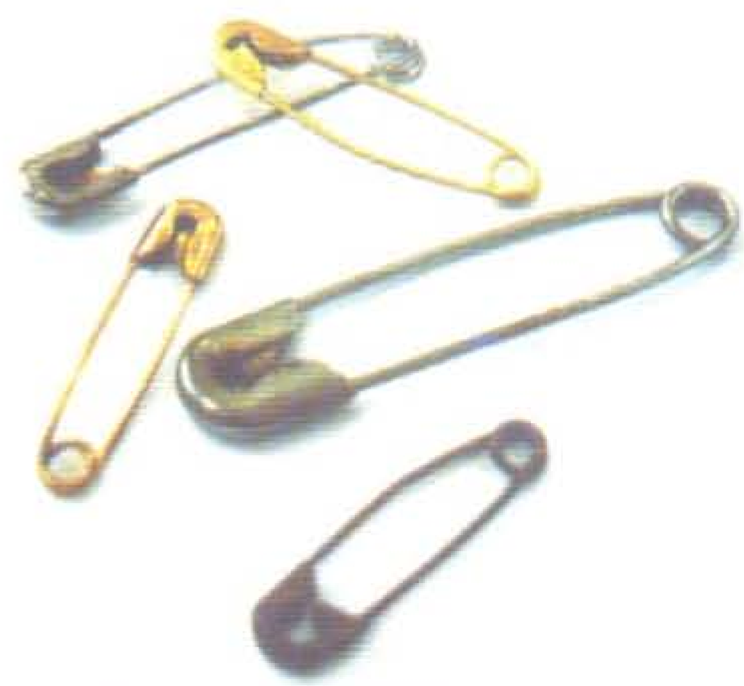
1. Iron extracted from the blast furnace is mainly used to produce steel.
2. Steel is an alloy of iron with carbon and/or other metals.
3. Adding different amounts of carbon or metals, and adding different metals to iron produce alloys with different properties.
4. High carbon steels are strong and brittle; low carbon steels are softer and more easily shaped.
5. The uses of steel are given in the table below.

Type of steel	Uses	Special properties
mild steel	car bodies and machinery	hard, strong and malleable
high carbon steel	cutting and boring tools	strong but brittle
stainless steel	equipment in chemical plants, cutlery and surgical instruments	resistant to corrosion

14.5 Rusting

What is rusting?

It was mentioned in section 14.1 that metals may react with air and water, and corrode. When an object made of iron is exposed to moist air for some time, a reddish-brown substance slowly forms on the surface of the metal. This substance is called **rust** and has the chemical name *hydrated iron(III) oxide*.



Have you ever left safety pins made of iron or steel by the sink and forgotten about them? What do you notice on the pins after a few days?

The process that produces rust is known as **rusting** or the **corrosion of iron**. It is the slow oxidation of iron to form hydrated iron(III) oxide (rust). The equation for rusting is

$$\text{iron} + \text{oxygen} + \text{water} \longrightarrow \underset{\text{(rust)}}{\text{hydrated iron(III) oxide}}$$

$$4Fe(s) + 3O_2(g) + 2xH_2O(l) \longrightarrow 2Fe_2O_3.xH_2O(s)$$

(The x in the chemical formula of hydrated iron(III) oxide indicates the number of water molecules present in the compound.)

What conditions are essential for rusting?

The experiment on page 250 investigates the essential conditions for rusting.

Experiment 2

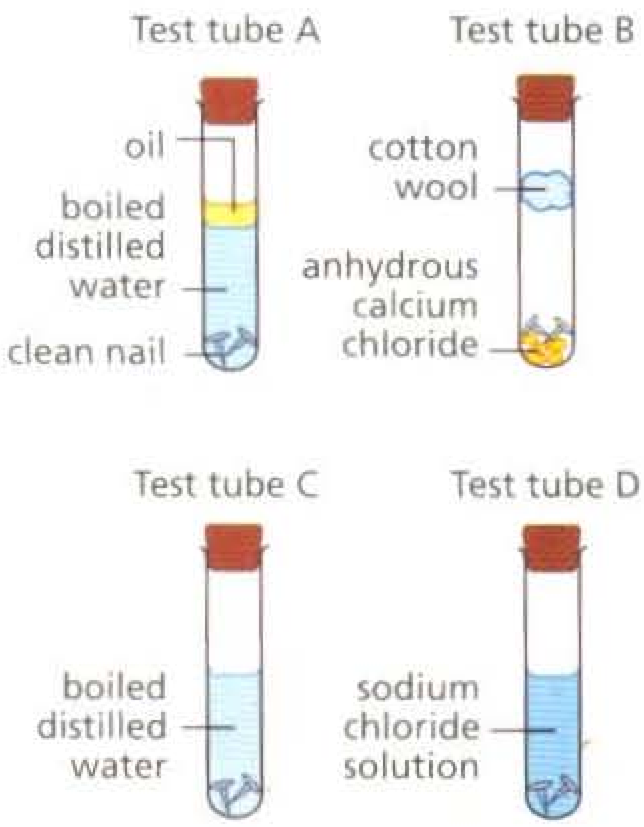


Fig. 14.21 Studying the conditions necessary for rusting

To investigate the conditions necessary for rusting.

Procedure

1. The iron nails are cleaned with sandpaper to remove any rust present.
2. The iron nails are then placed in four different test tubes as shown in Fig. 14.21.
3. After one week, the iron nails are examined and their appearance is noted.

The results of the experiment are summarised in Table 14.9.

Test tube	Conditions		Observation
	Is water present?	Is air present?	
A	yes	no	no rusting
B	no	yes	no rusting
C	yes	yes	nails rusted
D	yes	yes	nails rusted heavily

Table 14.9 Results of experiment 2

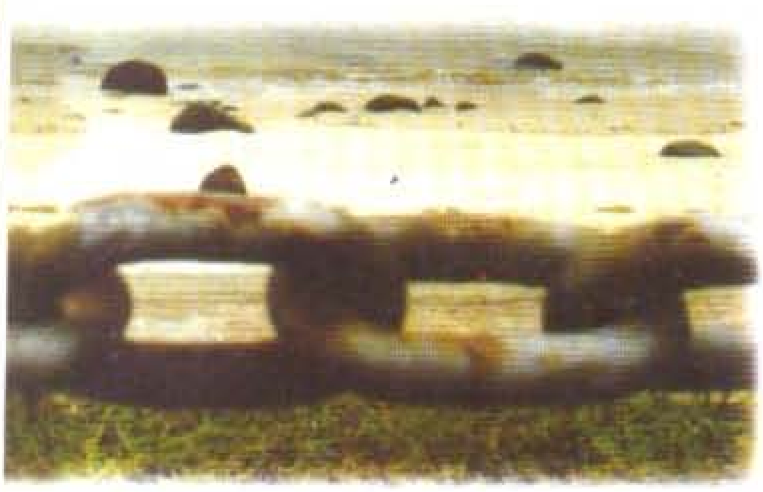
Science Skills

Scientists have shown that it is oxygen in the air that is essential for rusting. Think of ways to modify experiment 2 to verify this fact.

The results obtained show two important facts about rusting:

1. Both air and water are needed for rusting to occur.
2. The presence of sodium chloride increases the speed of rusting.

Besides sodium chloride, acidic substances such as sulphur dioxide and carbon dioxide also speed up the rusting process. Thus, iron objects near the sea and in industrial areas corrode more rapidly because of the presence of salt and other pollutants in the air.



This iron chain rusts rapidly because it is exposed to salt in the air.

Rust Prevention

Rust is very brittle and flaky. Thus, when iron corrodes, the rusted surface of the metal flakes away. This produces a new metal surface to corrode. Eventually, all of the metal will rust and flake away. There are three general methods of rust prevention — using a protective layer, using a sacrificial metal and using alloys.

How does using a protective layer prevent rusting?

In order to prevent iron from rusting, it has to be kept away from water and oxygen. One general method of 'rust proofing' is to coat the metal object with a protective layer of substance. This includes coating the iron/steel surface with paint or grease, covering it with plastic, electroplating it or dip plating it.

Both electroplating and dip plating involve galvanising, i.e. coating iron/steel with a layer of another metal. Electroplating uses electricity to do this. Dip plating is the process of dipping iron into molten zinc or tin. A thin film of zinc or tin then covers the iron/steel. This layer prevents water and air from coming into contact with the iron or steel surface.

Fig 14.22 Rust prevention using a sacrificial metal

Using a sacrificial metal

Fixing bars of zinc to a ship's hull prevents the ship's steel body from rusting (Fig. 14.22). Heavy blocks of magnesium attached to underground pipes made of iron protects the pipes from rusting (Fig. 14.23). These metals attached are more reactive than iron and will corrode instead of iron. As long as magnesium or zinc is present, iron will not rust. This is called sacrificial protection because the magnesium or zinc is 'sacrificed' to protect the iron or steel.

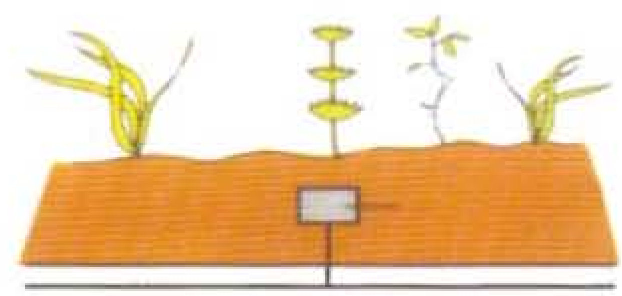


Fig 14.23 Sacrificial protection of underground pipes

Using alloys

The best known rust-resistant alloy of iron is stainless steel. Stainless steel contains nickel and chromium. On exposure to air and moisture, a very hard coating of chromium(III) oxide, Cr_2O_3, forms on the surface of stainless steel, protecting it from further corrosion.

Table 14.10 summarises the common methods of rust prevention.

Method	Where it is used	Comment
painting	large objects like cars, ships and bridges	If the paint on the metal surface is scratched, rusting will take place under the painted surface.
oiling or greasing	tools and machine parts	The protective film of oil or grease gathers dust and must be renewed.
plastic coating	kitchenware such as draining racks	If the plastic layer is torn, the iron starts to rust.
galvanising (zinc-plating)	water buckets, dustbins, 'zinc' roofs, kitchen sinks	The metal does not rust even if the zinc layer is damaged. (This is because zinc is more reactive than iron. So zinc corrodes instead of iron.)
tin-plating	food cans	If the tin layer is scratched, the iron beneath it rusts.
chrome-plating	taps, kettles, bicycle handle bars	This gives a bright shiny finish as well as rust protection.
metal block of zinc or magnesium	underground pipes, ships, legs of steel piers	Magnesium and zinc corrode in place of iron because they are more reactive metals.
stainless steel	cutlery, surgical instruments, pipes in chemical plants	Stainless steel contains chromium and nickel, which do not rust.

Table 14.10 Various methods of rust prevention

Key ideas

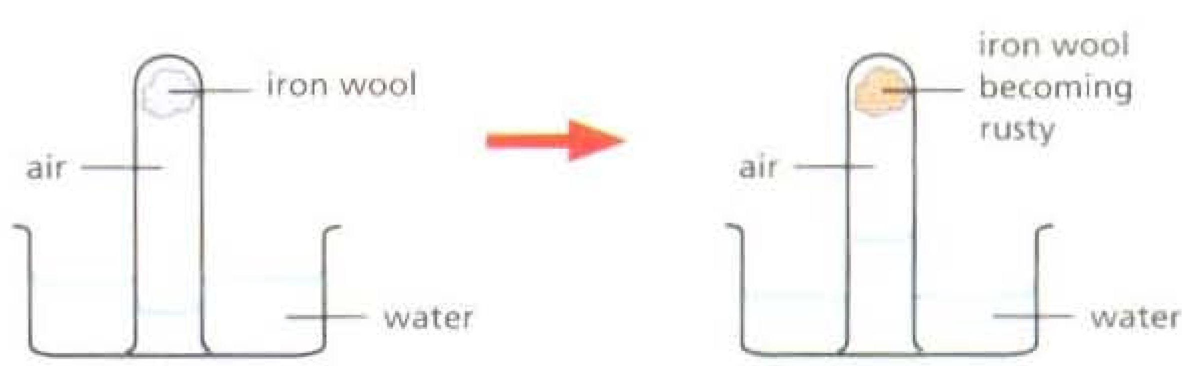
1. Rusting is the oxidation of iron to form hydrated iron(III) oxide. It only occurs in the presence of both oxygen and water.
2. Rusting can be prevented by placing a protective layer over the metal. This includes painting, greasing, applying a plastic coating over the metal surface or galvanising.
3. Placing a protective layer over iron prevents rusting because the protective layer stops oxygen and water from coming into contact with the metal.
4. Rusting can be prevented by sacrificial protection, i.e. attaching a more reactive metal to the iron or steel object. Rusting is prevented because oxygen reacts with the more reactive metal.

Test Yourself 14.5

Questions

1. Describe how rusting is prevented on
   a) the moving parts of a machine.
   b) food cans.
   c) kitchen sink taps.
   d) the parts of a steel pier submerged in water.

2. The diagram shows a simple experiment on rusting. What information can be drawn from the experiment?



At the beginning of the experiment   After one week

14.6 | Recycling Metals

Metals are **finite resources**. This means that *the amounts of the various metals in the Earth are limited.* Over the centuries, we have used up much of the Earth's metal resources. With the increasing demand for metals, our natural resources will not last much longer. Fig. 14.24 shows approximately how long it will take for the reserves of some metals to run out.

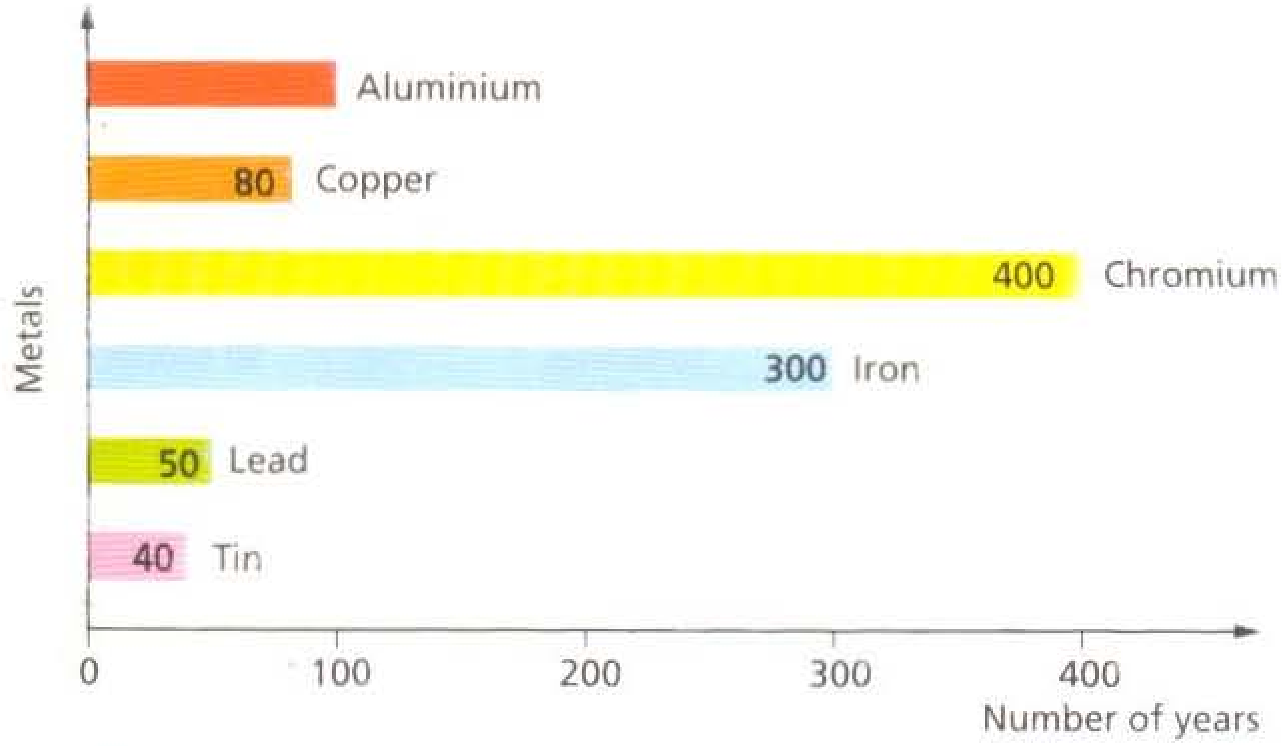


Fig. 14.24 Approximate time for the reserves of some metals to run out

Some metals such as chromium and iron are so abundant that it is difficult to believe they can ever run out. On the other hand, metals such as lead and tin are in very short supply. The reserves of these metals could run out in your lifetime! The world's reserves of raw metals may last longer if

- we find new ore deposits using advanced technology like satellites.
- we find substitutes to replace metals so that we can use the existing metals more sparingly. A good example is the replacement of some metal parts in cars with plastic or ceramic parts. Optical fibres made of special glass are replacing the metal wires used in communications instruments.
- we recycle metals. Metals are readily recycled — old metal objects can be crushed and melted for reuse (Fig. 14.25).

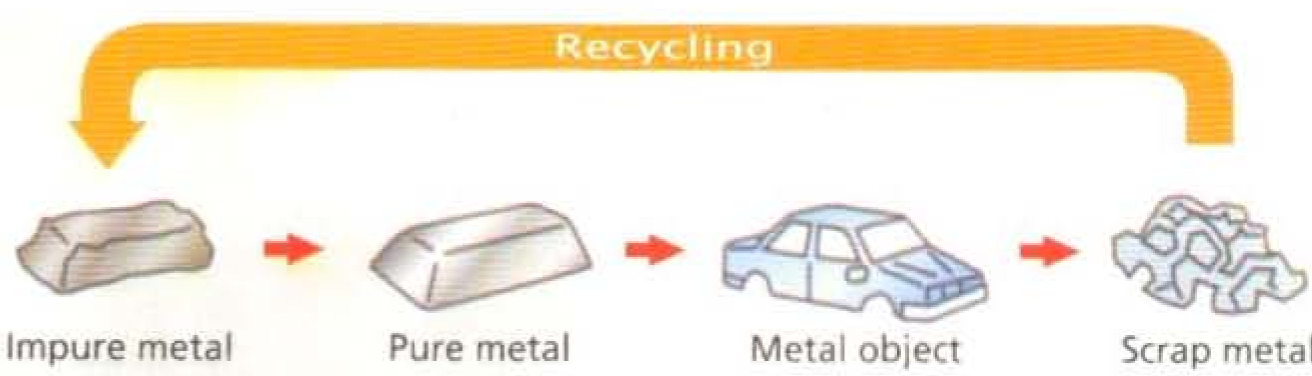


Fig. 14.25 Recycling metals for reuse

Let us now discuss the issues surrounding recycling.

Environmental impact of metal extraction

Before metals can be extracted from their ores, the land is mined for ores. Mined land is usually unsightly and it cannot support plant and animal life.

An enormous amount of waste material is also generated from the extraction process as shown in Fig. 14.26.

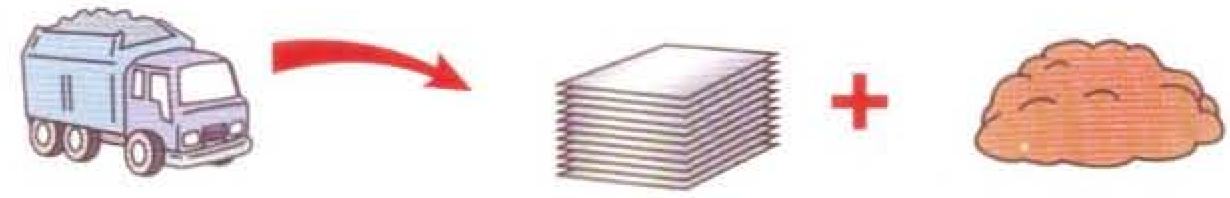


Fig. 14.26 About 90% of a metal ore ends up as waste after the extraction process.

About 90% of the metal ore used for metal extraction turns out to be waste and needs to be disposed of after the metal is extracted. If not disposed of carefully, the waste may leach into soil and nearby water bodies, polluting land and water. To dispose of the waste, huge landfill sites need to be dug.

The smelting of ores also causes more air pollution compared to any other industrial process. In addition, the extraction of metals from their ores requires a continuous supply of energy. The energy is usually generated from burning fossil fuels, which are already in scarce supply.

Advantages of recycling

If metals are recycled and reused, there will be less need to dig for metal ores. With a decrease in mining operations, land will be free for other uses such as agriculture. Air and water pollution will also be greatly reduced. Fewer landfills will be required to dispose of both used metal objects and waste material from metal extraction. This will help save on the cost of building landfill sites. By recycling metals instead of extracting them from ores, we are also conserving our limited fossil fuel reserves.

A type of mining known as open-pit mining leaves very large holes in the ground.

How are some metals commonly recycled?

Lead, iron and aluminium are mainly recovered by scrap metal recycling. Lead is recovered from car batteries. When a car battery no longer works, the lead inside it is reused. A large fraction of iron and steel produced today is also recycled from scrap metal. Aluminium is recycled mainly from drink cans and food containers.

Economic issues of recycling

One problem with recycling metals is that recycling can also be extremely costly. There is the cost of transporting the scrap metal to the processing plant. The different types of metals must be separated before they can be recycled. Additional costs are incurred to sort and clean the scrap metal.

Some metal-producing businesses may decide that the costs are too high and that recycling is not worthwhile. This is especially if the metal to be recycled is not an expensive nor a very valuable one.

Social issues of recycling

We need energy, clean air and water in order to survive. Compared to extracting metals from ores, recycling does not produce as much waste as that may endanger human health. With the increasing human population worldwide, more land will be needed to grow food, rear livestock, and build homes, factories, offices and highways. Building new mines reduces the land available for these other important uses.

Eventually, resources of metal would still be used up. Thus, it would be wise to start developing metal recycling programs and processes that are cost-effective and environmentally friendly now.

Although there are obvious advantages for recycling metals, it will take effort and time for communities and businesses to practise recycling as a way of life. Everyone will need to realise that each of us plays a vital role in conserving our natural resources.

Recycling bins — a common sight along the streets of Singapore

Key ideas

1. Metals are finite resources and need to be conserved.
2. Recycling metals
   - helps us to conserve natural resources, such as land and fossil fuels,
   - helps us to reduce the environmental problems related to extracting metals from metal ores,
   - may save money spent on building landfill sites.

Test Yourself 11.1

Worked Example

Why is it easier to recycle lead than to recycle aluminium?

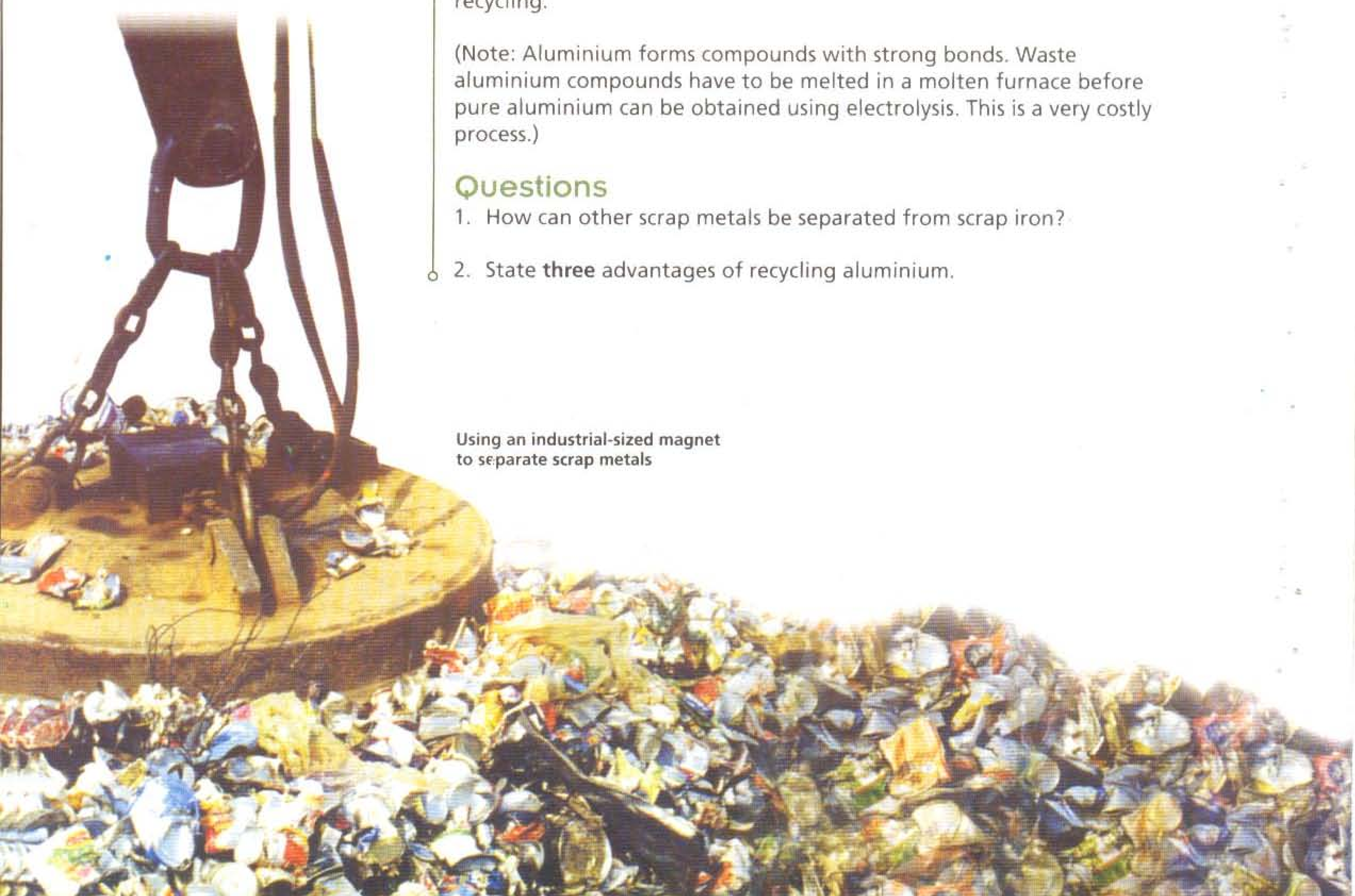
Answer

Lead is below aluminium in the reactivity series. Lead compounds can be easily reduced to lead using carbon, unlike aluminium. Lead also has a lower melting point than aluminium. Hence, it is easier to separate out lead from its mixtures with other substances for recycling.

(Note: Aluminium forms compounds with strong bonds. Waste aluminium compounds have to be melted in a molten furnace before pure aluminium can be obtained using electrolysis. This is a very costly process.)

Questions

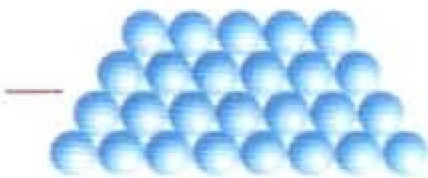
1. How can other scrap metals be separated from scrap iron?
2. State **three** advantages of recycling aluminium.

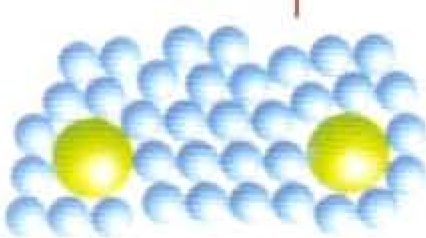


Using an industrial-sized magnet to separate scrap metals

Concept Map

Metals

- structure


- are mixed with other element(s) to form **Alloys**
  - structure


- are ordered from most eactive to least reactive in / are obtained as follows

Reactivity series
- potassium
- sodium
- calcium
- magnesium
- zinc
- iron
- lead
- copper
- silver
- gold

from ores using electricity (potassium, sodium, calcium, magnesium)

from ores by reduction with carbon (zinc, iron, lead, copper)

found uncombined (silver, gold)

the higher up the metal, the greater its tendency to form → **Positive ions**

- can be conserved by → **Recycling**

e.g. to extract iron

Coke (C) + **Limestone** ($CaCO_3$) + **Haematite** (Fe_2O_3 + impurities)

Reactions in blast furnace

- $C(s) + O_2(g) \xrightarrow{heat} CO_2(g)$
- $CaCO_3(s) \xrightarrow{heat} CaO(s) + CO_2(g)$
- $C(s) + CO_2(g) \xrightarrow{heat} 2CO(g)$
- $Fe_2O_3(s) + 3CO(g) \xrightarrow{heat} 2Fe(l) + 3CO_2(g)$
- $CaO(s) + SiO_2(s) \xrightarrow{heat} CaSiO_3(l)$

→ waste gases + molten iron + slag
(N_2, CO, CO_2)

most metals react

- **With water** → metal hydroxide + hydrogen
- **With steam** → metal oxide + hydrogen
- **With hydrochloric acid** → salt + hydrogen

iron requires both oxygen and water for **Rusting**

prevented by

- Greasing
- Painting
- Plastic coating
- Sacrificial protection

Exercise 14

Foundation

1. Which alloy contains copper as the main constituent?

   A Brass B Pewter
   C Solder D Stainless steel

2. Which metal does not react with dilute hydrochloric acid?

   A Iron B Sodium
   C Zinc D Copper

3. What happens when clean iron filings are placed in copper(II) sulphate solution?

   A Copper(II) ions are oxidised.
   B Iron is oxidised.
   C Iron(III) sulphate is formed.
   D There is no reaction.

4. How is iron extracted from iron(III) oxide in the blast furnace?

   A Oxidise iron(III) oxide with carbon dioxide.
   B Oxidise iron(III) oxide with oxygen.
   C Reduce iron(III) oxide with carbon monoxide.
   D Reduce iron(III) oxide with limestone.

5. Which substance is used to remove impurities from iron ore in the blast furnace?

   A Carbon B Carbon monoxide
   C Limestone D Silica

6. Which pair of metals will slow down rusting when they are in contact with steel?

   A Magnesium and copper.
   B Magnesium and zinc.
   C Zinc and copper.
   D Zinc and lead.

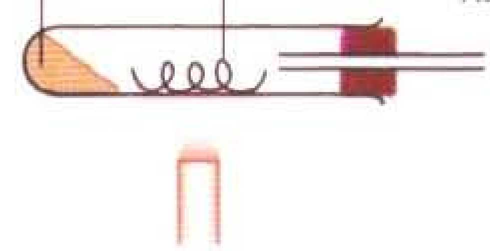
7. a) Place the metals calcium, potassium and zinc in order of chemical reactivity towards water, stating the most reactive metal first.

   b) Write the chemical equation for the reaction between the most reactive metal in (a) and water.

8. a) Steel is an alloy of iron. Explain what you understand by this statement.

   b) Draw diagrams of the structures of iron and steel to show their differences.

9. Magnesium reacts with steam to form magnesium oxide and a gas.



   a) Name the gas formed during the reaction.

   b) Complete the above diagram to show how the gas could be collected.

   c) Describe a test for the gas and state the result you would expect to obtain.

   d) This experiment should not be carried out using potassium. Why?

10. a) Explain what is meant by the term 'recycling'.

    b) Why are copper and lead easy to recycle?

    c) Name the resources that are saved by recycling.

Challenge

1. Metal *X* reacts with water to give a solution *Y*. Solution *Y* gives a white precipitate when carbon dioxide is passed through it. What is metal *X*?

   A Calcium B Lead
   C Nickel D Potassium

2. Which statement about the rusting of iron is not correct?

   A Rusting is a redox process.
   B Rusting is accelerated by the presence of carbon dioxide.
   C Rusting requires both oxygen and water to be present.
   D The composition of rust is $Fe_3O_4.xH_2O$.

3. Four metals, A, B, C and D, are tested with water and with dilute hydrochloric acid. The table below shows the results of the experiment.

Metal	Reaction with water	Reaction with steam	Reaction with HC*l* (aq)
A	×	√	√
B	×	×	√
C	√	√	√
D	×	×	×

a) Place the metals A, B, C and D in order, stating the most reactive first.

b) Between which two metals should hydrogen be placed in the series in (a)?

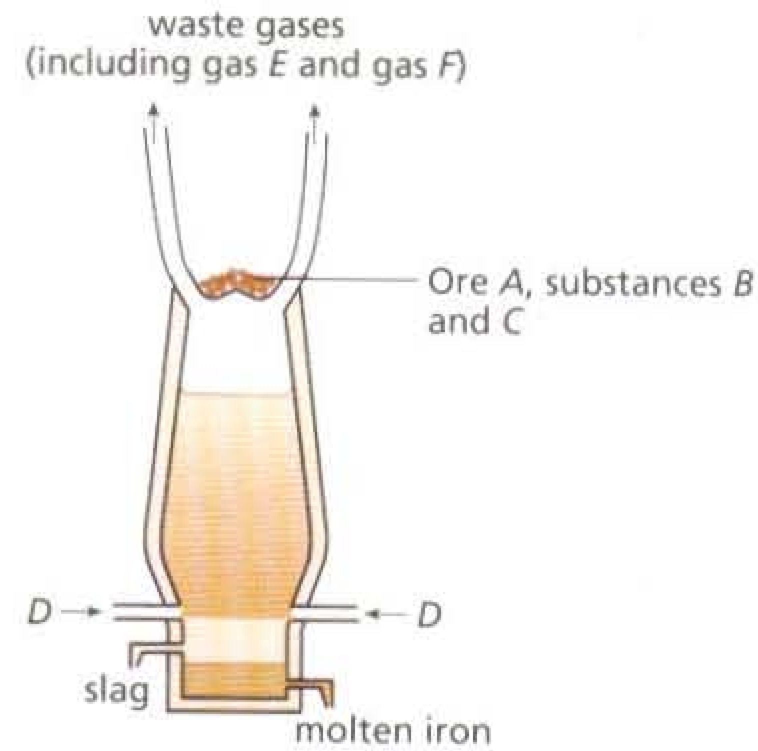
c) Predict the method used for the extraction of metal C.

4. a) Nickel is below iron but above lead in the reactivity series. Aqueous solutions of nickel compounds are green. Predict the observations and write the chemical equation for the reaction that takes place when
   i) nickel is added to dilute hydrochloric acid.
   ii) nickel is heated in steam.

b) Predict how you would expect nickel to react when
   i) it is heated with magnesium oxide.
   ii) a piece of nickel is added to copper(II) nitrate solution.

c) i) From the reactions in (b)(i) and (b)(ii), state which of these metals, nickel, magnesium and copper, has the highest tendency to form its positive ions?
   ii) Based on their tendency to form positive ions, deduce the positions of magnesium and copper with respect to nickel in the reactivity series.

5. The diagram below shows the manufacture of iron from ore A in a blast furnace.



Substance B burns in the furnace to produce gas E, which is reduced by more B to give gas F. This gas reduces the iron ore to iron.

Substance C decomposes to form carbon dioxide and a white solid G, which reacts with silica to form a molten slag containing calcium silicate. Molten iron and slag settle to the base of the furnace.

a) Identify substances A – G.

b) Write the equation for the reaction between G and silica to form calcium silicate.

c) Select one redox reaction that occurs in the blast furnace. Write the equation for this reaction and identify the reducing agent used.

d) If the haematite is wet, the waste gases also contain hydrogen.
   i) Explain, with an equation, how hydrogen is formed.
   ii) Why is the formation of hydrogen in the furnace dangerous?

e) Iron is often alloyed to reduce rusting. Name one element used for this purpose.

6. You are asked to use simple apparatus to investigate the effect of sulphur dioxide on the corrosion of copper, magnesium and zinc.

a) Why must a control be set up for the experiment?

b) How do you know if corrosion has occurred?

c) Predict the metal that will corrode most easily. Explain your answer.

Chemistry Today

You may have bought rechargeable batteries before. Have you ever read the small print on these batteries? It says 'Please dispose of these wisely due to their cadmium content.' Cadmium, a metal, is a hazardous waste because it is toxic.

The nickel-cadmium battery (commonly abbreviated NiCd or NiCad) is a popular rechargeable battery used in toys, cordless phones and power tools.

Most NiCd batteries are sealed inside these appliances. When the appliances are worn out, they are thrown away. Many people do not realise that by throwing away these appliances, they are contributing to toxic waste. NiCd batteries end up in landfill sites where their casings gradually deteriorate. The cadmium then leaches into the soil and eventually, the water supply. If the batteries happen to be incinerated with rubbish, the ash produced is also a hazardous waste.

Recharging nickel-cadmium batteries

Over 1.5 billion NiCd batteries are sold worldwide annually, so the problem caused by disposal is enormous. However, many manufacturers are unwilling to recycle the batteries because the costs are very high and there is very little profit to be made.

NiCd batteries are highly efficient and have a very long shelf life. Thus, scientists are working to invent alternatives that work as well as NiCd batteries but are non-toxic. The nickel metal hydride cell and lithium ion batteries are potential alternatives. Both these types of cells do not contribute to toxic waste. In Europe, the use of cadmium in electrical equipment was banned in December 2004.

CRITICAL THINKING

You have been asked to reduce the toxicity of batteries. You have thought of four possible solutions:

- Redesigning batteries to reduce the toxic parts
- Replacing batteries with less toxic constituents
- Reducing the number of batteries thrown away by extending battery life
- Recycling

Discuss the arguments for and against each solution.