Cambridge OL

Chemistry

CODE: (5070)

Chapter 10

Metals

10.1 Properties of metals

Most of the elements in the Periodic Table are metals. They have very different properties to the non-metallic elements. A comparison of their properties is shown in Table 10.1.

▼ **Table 10.1** How the properties of metals and non-metals compare

Property	Metal	Non-metal
Thermal conductivity	Good	Poor
Electrical conductivity	Good	Poor
Malleability	Good	Poor – usually soft or brittle
Ductility	Good	Poor – usually soft or brittle
Melting point	Usually high	Usually low
Boiling point	Usually high	Usually low

a Sodium burning in air/oxygen

b Iron rusts when left unprotected

You have already seen in Chapter 2, p. 11, that metals usually have similar physical properties. However, they differ in other ways. Look closely at the three photographs in Figure 10.1

c Gold is used in leaf form on this giant Buddha as it is unreactive

▲ **Figure 10.1**

▲ **Figure 10.2** Effervescence occurs when magnesium is put into acid

10.2 Metal reactions

By carrying out reactions in the laboratory with other metals and with air, water and dilute acid, it is possible to produce an order of reactivity of the metals.

With acid

When a metal reacts with dilute hydrochloric acid, hydrogen and metal chloride are produced, as shown in Figure 10.2, with effervescence and magnesium chloride as the salt.

$$\text{magnesium} + \text{hydrochloric acid} \rightarrow \text{magnesium chloride} + \text{hydrogen}$$

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

A **reactivity series** is a measure of the rate of hydrogen evolution, produced when similar reactions are performed with other metals with acid.

Key definition

The **reactivity series** is the order of the reactivity of the following elements: potassium, sodium, calcium, magnesium, aluminium, carbon, zinc, iron, hydrogen, copper, silver, gold.

With air/oxygen

Many metals react directly with oxygen to form oxides.

magnesium + oxygen → magnesium oxide

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

With cold water/steam

Reactive metals such as potassium, sodium and calcium react with cold water to produce the metal hydroxide and hydrogen gas.

sodium + water → sodium hydroxide + hydrogen

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

Moderately reactive metals like magnesium, zinc, and iron react slowly with water, but more rapidly with steam, forming metal oxide and hydrogen gas.

magnesium + steam → magnesium oxide + hydrogen

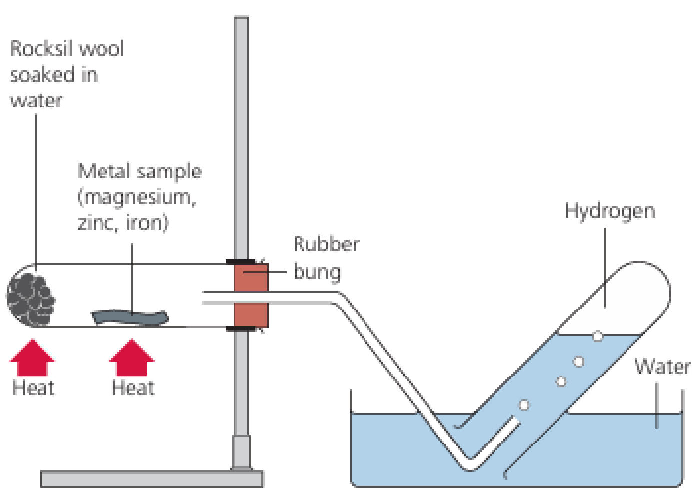
$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

This teacher demonstration requires eye protection and removal of the delivery tube before heating. Table 10.2 shows metal reactivity, with the most reactive metal first. The most reactive metal loses outer electrons to form a positive metal ion, requiring easy extraction.

a This wood-burning stove is made of iron

b Copper pots and pans

▲ **Figure 10.4**



▲ **Figure 10.3** Apparatus used to investigate how metals such as magnesium react with steam

▲ **Figure 10.5** Planes are made of an alloy which contains magnesium and aluminium

▼ **Table 10.2** Order of reactivity

Reactivity series	Reaction with dilute acid	Reaction with air/oxygen	Reaction with water	Ease of extraction	
Potassium (K)	Produce H_2 with decreasing vigour	Burn very brightly and vigorously	Produce H_2 with decreasing vigour with cold water	Difficult to extract	↑ Increasing reactivity of metal
Sodium (Na)					
Calcium (Ca)		Burn to form an oxide with decreasing vigour	React with steam with decreasing vigour	Easier to extract	
Magnesium (Mg)					
Aluminium (Al*)					
[Carbon (C)]					
Zinc (Zn)					
Iron (Fe)					
[Hydrogen (H)]		React slowly to form the oxide	Do not react with cold water or steam		
Copper (Cu)	Do not react with dilute acids			Found as the element (uncombined)	
Silver (Ag)		Do not react			
Gold (Au)					

* Because aluminium reacts so readily with the oxygen in the air, a protective oxide layer is formed on its surface. This often prevents any further reaction and disguises aluminium's true reactivity. This gives us the use of a light and strong metal.

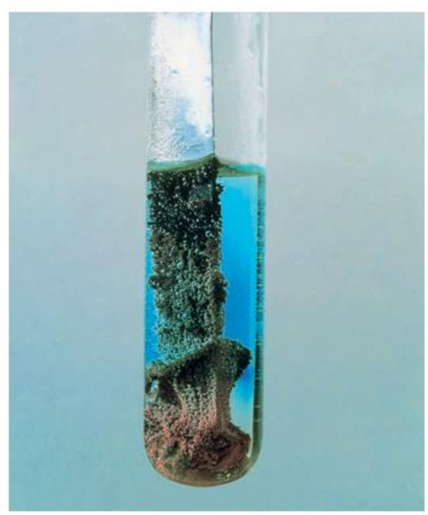
10.3 Reactivity of metals and their uses

Unreactive metals like copper are commonly used for electrical wiring due to their good conductivity and ductility, while aluminum, a reactive metal, forms a thick oxide layer to prevent further reactions.

This gives us a light, strong metal for use in:
» The manufacture of aircraft because of its low density
» The manufacture of overhead electrical cables because of its low density and good electrical conductivity
» Food containers because of its resistance to corrosion.

Displacement reactions

Metals compete in a displacement reaction, where a more reactive metal displaces a less reactive one from a solution of its salt. The reactivity series predicts which metal wins, with zinc above copper. This process causes copper (II) nitrate to lose its blue color.



▲ **Figure 10.6** Zinc displaces copper

zinc + copper(II) nitrate → zinc nitrate + copper

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

The ionic equation for this reaction is:

zinc + copper ions → zinc ions + copper

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

This redox reaction involves the transfer of two electrons from zinc metal to copper ions, resulting in zinc oxidation and copper reduction in aqueous solution, confirming metal reactivity series.

10.4 Identifying metal ions

Alkali dissolves in water, producing hydroxide ions. When added to a metal salt solution, an insoluble, often colored metal hydrogen precipitates from the solution.

If we take the example of iron (III) chloride with sodium hydroxide solution:

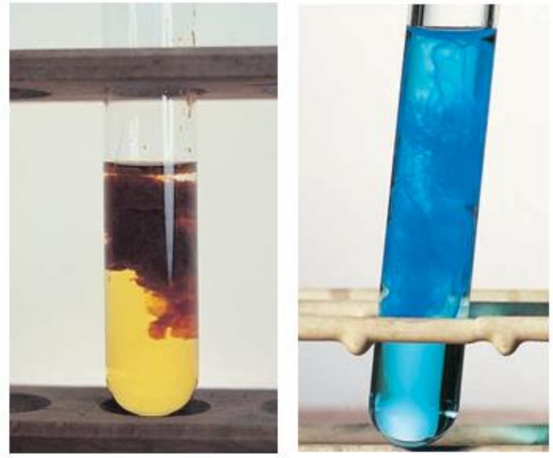
chloride + hydroxide → hydroxide + chloride
iron(III) sodium sodium iron(III)

$$FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)$$

The ionic equation for this reaction is:

iron(III) ions + hydroxide ions → iron(III) hydroxide

$$Fe^{3+}(aq) \quad + \quad 3OH^-(aq) \quad \rightarrow \quad Fe(OH)_3(s)$$



a Iron(III) hydroxide is precipitated **b** Copper(II) hydroxide is precipitated

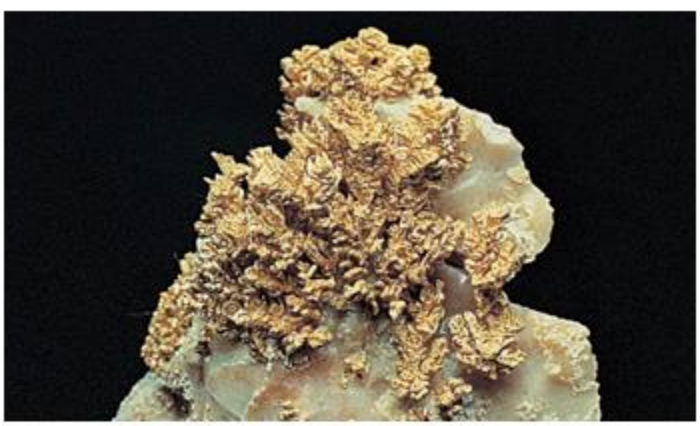
▲ **Figure 10.8**

▼ **Table 10.3** The effect of adding sodium hydroxide solution to solutions containing various metal ions

Metal ion present in solution	Effect of adding sodium hydroxide solution	
	A few drops	**An excess**
Aluminium, Al^{3+}	White precipitate of aluminium hydroxide	Precipitate is soluble in excess, giving a colourless solution
Calcium, Ca^{2+}	White precipitate of calcium hydroxide	Precipitate is insoluble in excess
Chromium(III), Cr^{3+}	Green precipitate of chromium(III) hydroxide	Precipitate is insoluble in excess
Copper(II), Cu^{2+}	Light blue precipitate of copper(II) hydroxide	Precipitate is insoluble in excess
Iron(II), Fe^{2+}	Green precipitate of iron(II) hydroxide	Precipitate is insoluble in excess, turns brown near the surface on standing
Iron(III), Fe^{3+}	Red-brown precipitate of iron(III) hydroxide	Precipitate is insoluble in excess
Zinc, Zn^{2+}	White precipitate of zinc hydroxide	Precipitate is soluble in excess, giving a colourless solution

10.5 Extraction of metals

Metals have been used since prehistoric times, with primitive iron tools excavated from meteorite rock. Around 2500 BC, iron became more widely used, with people learning to extract it from its ores using charcoal reduction. Most metals are too reactive to exist on their own in the Earth's crust and occur naturally in rocks as compounds in ores. Some metals, like gold and silver, exist in a native form as free metals.



▲ **Figure 10.10** Gold crystals

Heavy machinery crushes large ore lumps, with some ores already concentrated. Copper pyrites are less concentrated, requiring concentration before extraction. The method depends on the metal's position in the reactivity series. Hematite contains over 80% Fe_2O_3.

▼ **Table 10.4** Some common ores

Metal	Name of ore	Chemical name of compound in ore	Formula	Usual method of extraction
Aluminium	Bauxite	Aluminium oxide	$Al_2O_3.2H_2O$	Electrolysis of oxide dissolved in molten cryolite
Copper	Copper pyrites	Copper iron sulfide	$CuFeS_2$	The sulfide ore is roasted in air
Iron	Hematite	Iron(III) oxide	Fe_2O_3	Heat oxide with carbon
Sodium	Rock salt	Sodium chloride	NaCl	Electrolysis of molten sodium chloride
Zinc	Zinc blende	Zinc sulfide	ZnS	Sulfide is roasted in air and the oxide produced is heated with carbon

▲ **Figure 10.9** Chalcopyrite, an ore of copper (top) and galena, an ore of lead (bottom)

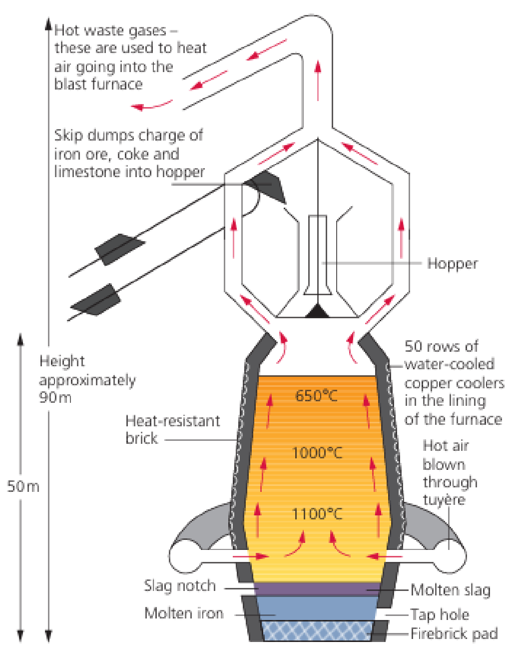
Reactive metals like sodium, like sodium chloride, are difficult to extract due to their strong bonding with other elements. Electrolysis is used to separate these ions and isolate sodium metal, but it is expensive and often located in hydroelectric power-rich regions to keep costs low.

Extraction of iron

Iron is extracted from its oxides, hematite and magnetite, in a blast furnace. The 50m high steel tower is lined with heat-resistant bricks and loaded with iron ore, carbon, and limestone. A blast of hot air illuminates the 'charge'.

carbon + oxygen → carbon dioxide

$$C(s) + O_2(g) \rightarrow CO_2(g)$$



▲ **Figure 10.12** Cross-section of a blast furnace

▲ **Figure 10.11** A blast furnace

Several chemical reactions then follow.

» The limestone begins to decompose:

calcium carbonate → calcium oxide + carbon dioxide

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

» The carbon dioxide gas produced reacts with more hot coke higher up in the furnace, producing carbon monoxide in an endothermic reaction.

$$\text{carbon dioxide} + \text{coke} \rightarrow \text{carbon monoxide}$$

$$CO_2(g) + C(s) \rightarrow 2CO(g)$$

» Carbon monoxide is a reducing agent (Chapter 3, p. 36). It rises up the furnace and reduces the iron(III) oxide ore. This takes place at a temperature of around 700°C:

$$\text{iron(III) oxide} + \text{carbon monoxide} \rightarrow \text{iron} + \text{carbon dioxide}$$

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

» Calcium oxide is a base and this reacts with acidic impurities such as silicon(IV) oxide in the iron, to form slag which is mainly calcium silicate

$$\text{calcium oxide} + \text{silicon(IV) oxide} \rightarrow \text{calcium silicate}$$

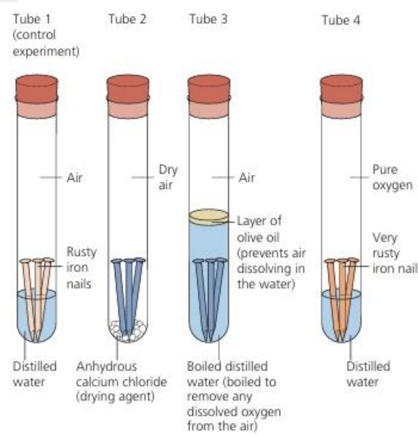
$$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(s)$$

Key definition
Rusting of iron and steel to form hydrated iron(II) oxide requires both water and oxygen.

The molten iron and slag are poured into a furnace, with slag floating on top due to its less dense nature. The molten iron and slag are periodically tapped off, while waste gases, mainly nitrogen and carbon oxides, escape to reduce energy costs. Slag is used by builders and road builders for foundations.

10.6 Metal corrosion

Rusting is a significant issue affecting iron and steel structures, costing over $2.5 trillion annually. Rust, an orange-red powder, is primarily composed of hydrated iron (III) oxide. Water and oxygen are essential for iron to rust, with salt promoting this process.



▲ **Figure 10.17** Rusting experiment with nails

Rust prevention

To prevent iron rusting, it is necessary to stop oxygen (from the air) and water coming into contact with it. There are several ways of doing this.

Painting

Ships, lorries, cars, bridges and many other iron and steel structures are painted to prevent rusting (Figure 10.18). However, if the paint is scratched, the iron beneath it will start to rust (Figure 10.19) and corrosion can then spread under the paintwork which is still sound. This is why it is essential that the paint is kept in good condition and checked regularly.

▲ **Figure 10.18** Painting keeps the air and water away from the steel used to build a ship

Oiling/greasing

Oil is commonly used to protect iron and steel in machinery moving parts from air or moisture, but the protective film must be renewed regularly.

Coating with plastic

The exteriors of refrigerators, freezers and many other items are coated with plastic, such as PVC, to prevent the steel structure rusting (Figure 10.20).

▲ **Figure 10.19** A brand new car is protected against corrosion (top). However, if the paintwork is damaged, then rusting will result

▲ **Figure 10.20** A coating of plastic stops metal objects coming into contact with oxygen or water

Galvanising

Steel beams and waste collection bins are galvanized by dipping them in molten zinc. This process slowly corrodes the zinc layer, protecting the iron and preventing it from being scratched away.

Sacrificial protection

Zinc bars are used to attach to ship hulls and oil rigs, as it reacts favorably with iron, causing corrosion and more easily forming positive ions.

$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$$

Zinc bars protect iron structures from rusting, while magnesium blocks protect pipes made of iron and steel. As the metal corrodes, it loses electrons to the iron, ensuring its protection.

▲ **Figure 10.21** The Burnley Singing Ringing Tree is a sculpture made from galvanised tubes

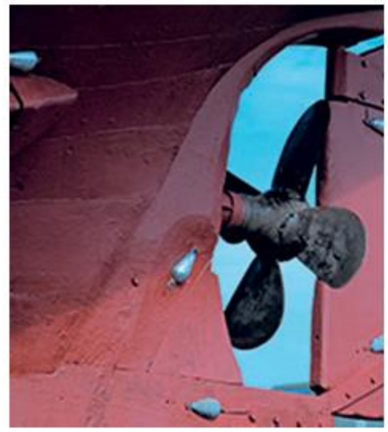
Corrosion

Rusting is the most common form of corrosion, primarily affecting iron and steel. Corrosion occurs when metals are chemically attacked by oxygen, water, or other substances. Higher metals corrode more rapidly, while others like magnesium, calcium, and aluminum are covered by oxide. Copper, for example, turns brown due to copper(II) oxide formation. Verdigris, a green color, forms on copper roofs and pipes. Gold and platinum are unreactive and do not corrode.

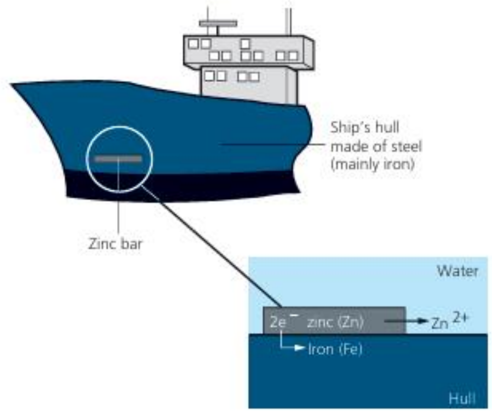
▲ **Figure 10.23** Verdigris soon forms on the surface of copper materials in exposed environments

10.7 Alloys

Alloys are mixtures of metals with other elements, producing metallic substances with more useful properties than pure metals. Examples include brass made from copper and zinc, which is harder and more corrosion resistant. Steel, a mixture of iron and non-metal carbon, is the most important alloy, containing iron, carbon, and other metals like nickel and chromium.



a Bars of zinc on the hull of a ship



b The zinc is sacrificed to protect the steel. Electrons released from the dissolving zinc cause reduction to occur at the surface of the hull

▲ **Figure 10.22** Sacrificial protection

Key definition

Alloys can be harder or stronger than the pure metals and are more useful.

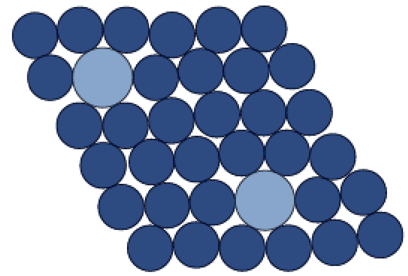
a Bronze is often used in sculptures

b A polarised light micrograph of brass showing the distinct grain structure of this alloy

▲ **Figure 10.24**

Alloys to order

Alloys are designed by metallurgists to suit various uses, with thousands of alloys available. Many alloys are harder or stronger than pure metals, as they mix different-sized atoms, causing the metal to lose its malleability and ductility. This results in alloys with less repeating structures, allowing layers to slide over each other.



▲ **Figure 10.25** Alloy structure. The dark circles represent atoms of a metal; the pale circles are the larger atoms of a different metal added to make the alloy. The different size of these atoms gives the alloy different physical properties from those of the pure metal

▼ **Table 10.5** Uses of common alloys

Alloy	Composition	Use
Brass	65% copper, 35% zinc	Jewellery, machine bearings, electrical connections, door furniture
Bronze	90% copper, 10% tin	Castings, machine parts
Cupro-nickel	30% copper, 70% nickel	Turbine blades
Cupro-nickel	75% copper, 25% nickel	Coinage metal
Duralumin	95% aluminium, 4% copper, 1% magnesium, manganese and iron	Aircraft construction, bicycle parts
Magnalium	70% aluminium, 30% magnesium	Aircraft construction
Pewter	30% lead, 70% tin, a small amount of antimony	Plates, ornaments and drinking mugs
Solder	70% lead, 30% tin	Connecting electrical wiring

Revision questions

9. Nov/2021/Paper_21/No.7

This question is about metals and metal compounds.

(a) Silver is a transition element. Potassium is a metal in Group I of the Periodic Table.

State two differences in the physical properties of silver and potassium.

1 ....

2 ....
[2]

(b) An ion of silver has the symbol

$${}^{109}_{47}\mathrm{Ag}^{+}$$

Deduce the number of protons, neutrons and electrons in this ion.

number of protons ....

number of neutrons ....

number of electrons ....
[3]

(c) Potassium reacts with water to form a gas which 'pops' with a lighted splint.

Complete the equation for this reaction.

...... K + ...H_2O $\rightarrow$ 2KOH + ... [1]

(d) When zinc carbonate is warmed in a closed container, an equilibrium mixture is formed.

$$ZnCO_3(s) \rightleftharpoons ZnO(s) + CO_2(g)$$

The forward reaction is endothermic.

(i) Describe and explain the effect, if any, on the position of equilibrium when the temperature is decreased.

(ii) Describe and explain the effect, if any, on the position of equilibrium when the concentration of carbon dioxide is increased.

20. Jun/2021/Paper_21/No.4

Copper(II) chloride, copper(II) iodide and copper(II) carbonate are ionic compounds.

(a) Predict two physical properties, other than electrical conductivity, of copper(II) chloride.

1. ....

2. ....
[2]

(b) Copper is a transition element.

Suggest **one** property of copper(II) chloride that is characteristic of a compound of a transition element.

.... [1]

(c) Copper reacts with chlorine to make copper(II) chloride.

$$Cu + Cl_2 \rightarrow CuCl_2$$

Copper(II) chloride contains Cu^{2+} and Cl^- ions.

Explain, in terms of the movement of electrons, how $CuCl_2$ is formed from copper atoms and chlorine molecules.

(d) Copper(II) iodide decomposes to make iodine and copper(I) iodide.

The ionic equation for this reaction is shown.

$$2Cu^{2+} + 4I^- \rightarrow 2CuI + I_2$$

(i) Use the information to explain that oxidation takes place.

....

....

(ii) Use the information to explain that reduction takes place.

....

....

(e) A sample of copper(II) carbonate is heated strongly.

Name the products of this reaction.

21. Jun/2021/Paper_21/No.5

Iron pyrite, FeS_2, is an ore of iron.

When heated in air, FeS_2 produces both iron(III) oxide and sulfur dioxide.

(a) Balance the equation shown.

.... FeS_2 + ... O_2 $\rightarrow$ .... Fe_2O_3 + ... SO_2 [1]

(b) Describe one environmental problem caused by sulfur dioxide in the air.

.... [1]

(c) Describe how sulfur dioxide is converted into sulfuric acid in the contact process.

(d) State one **other** use of sulfur dioxide.

.... [1]

(e) Iron(III) oxide, coke, limestone and hot air are heated together in a blast furnace to make molten iron.

Describe the function in the blast furnace of:

(i) coke

....

.... [1]

(ii) limestone

....

.... [1]

(iii) hot air.

11. Nov/2021/Paper_21/No.3

This question is about copper and copper compounds.

(a) Copper is a metal.

Explain why copper conducts electricity.

.... [1]

(b) Describe a test for copper(II) ions.

test ....

observations ....

[2]

(c) Aqueous copper(II) sulfate is electrolysed using graphite electrodes.

(i) Describe what is observed during this electrolysis:

- at the positive electrode

....

- at the negative electrode

....

- in the electrolyte.

....

[3]

(ii) Graphite conducts electricity.

Give one other reason why graphite electrodes are used in electrolysis.

[1]

(d) Aqueous copper(II) sulfate reacts with magnesium.

$$CuSO_4 + Mg \rightarrow Cu + MgSO_4$$

Construct the ionic equation, including state symbols, for this reaction.

(e) A 2.25 g sample of an oxide of copper contains 0.250 g of oxygen.

Deduce the empirical formula of this oxide of copper.

(f) There are several commonly used alloys of copper.

What is the meaning of the term *alloy*?

12. Nov/2021/Paper_21/No.7

Aluminium is extracted by the electrolysis of molten aluminium oxide.

(a) (i) Explain why aluminium is extracted by electrolysis and not by reduction with carbon.

.... [1]

(ii) The electrolyte is a mixture of aluminium oxide and cryolite.

Explain the purpose of the cryolite.

....

.... [1]

(iii) At the positive electrode (anode) oxide ions are converted to oxygen.

Construct the equation for this reaction.

.... [1]

(b) Aluminium can also be produced on a small scale by reacting aluminium oxide with magnesium.

$$Al_2O_3 + 3Mg \rightarrow 2Al + 3MgO$$

(i) Use this equation to explain why the Al_2O_3 is reduced.

....

.... [1]

(ii) Calculate the maximum mass of aluminium formed when 25.5 g of aluminium oxide reacts with excess magnesium.

(c) Aluminium is a metal.

Use your knowledge of the structure of metals to explain why aluminium is malleable.

....

....

.... [2]

(d) When aluminium is heated in chlorine, aluminium chloride is formed.

The reaction is exothermic.

Explain, in terms of bond making and bond breaking, why this reaction is exothermic.

13. Nov/2021/Paper_22/No.7

Iron is extracted from iron ore in a blast furnace using limestone and coke (carbon).

(a) Name a common ore of iron.

.... [1]

(b) The coke burns to form carbon dioxide.

This reaction is exothermic.

Explain, in terms of bond making and bond breaking, why this reaction is exothermic.

....

....

.... [2]

(c) Carbon dioxide reacts with hot coke to form carbon monoxide.

The carbon monoxide reduces the iron(III) oxide in the iron ore.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

(i) Use this equation to explain why the Fe_2O_3 is reduced.

....

.... [1]

(ii) Calculate the maximum mass of iron formed when 12.5g of iron(III) oxide react with excess carbon monoxide.

(e) Iron is a metal.

Describe metallic bonding.

(d) Silicon dioxide is an impurity in the iron ore.

Explain how the addition of limestone helps remove silicon dioxide from the blast furnace.

23. Jun/2020/Paper_21/No.2

Part of the reactivity series is shown.

calcium	more reactive
aluminium	↓
manganese	↓
zinc	less reactive

(a) Predict the names of the products formed when manganese reacts with dilute hydrochloric acid.

....

.... [1]

(b) A sample of manganese(II) carbonate, $MnCO_3$, is heated strongly.

Construct the equation for this reaction.

.... [1]

(c) Powdered manganese is added to aqueous zinc sulfate to form aqueous manganese(II) sulfate, $MnSO_4$.

Construct an ionic equation, with state symbols, for this reaction.

.... [2]

(d) Zinc powder, a reducing agent, is added to acidified aqueous potassium manganate(VII).

Describe the colour change during this reaction.

(e) Aluminium is extracted by the electrolysis of aluminium oxide dissolved in molten cryolite.

(i) Write the electrode equation for the formation of aluminium atoms at the cathode.

.... [1]

(ii) Write the electrode equation for the formation of oxygen molecules at the anode.

.... [1]

(f) State one advantage of recycling aluminium.

24. Jun/2020/Paper_22/No.4

Part of the reactivity series is shown.

magnesium	more reactive
aluminium	↓
zinc	
chromium	
iron	less reactive

(a) Predict the names of the products formed when chromium reacts with dilute hydrochloric acid.

....

.... [1]

(b) Powdered zinc is added to aqueous chromium(III) ions, $Cr^{3+}(aq)$.

Construct an ionic equation, with state symbols, for this reaction.

.... [2]

(c) Explain why aluminium does **not** react with water.

(d) Hydrogen peroxide, an oxidising agent, is added to aqueous potassium iodide in a test-tube.

Describe the colour change seen in the test-tube.

.... [1]

(e) Chromium is extracted by the reaction of aluminium with chromium(III) oxide, Cr_2O_3.

(i) Write the equation for this reaction.

.... [1]

(ii) Suggest a **compound** that can reduce chromium(III) oxide to chromium metal.

.... [1]

(f) State one advantage of recycling metals.

[1]