Metals (Part 1)

Learning Objectives

Properties of metals

- describe the general physical properties of metals as solids having high melting and boiling points, malleable, good conductors of heat and electricity in terms of their structure
- describe alloys as a mixture of a metal with another element, e.g. brass; stainless steel
- identify representations of metals and alloys from diagrams of structures
- explain why alloys have different physical properties to their constituent elements

Reactivity of metals

- place in order of reactivity calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to

(i) the reactions, if any, of the metals with water, steam, dilute hydrochloric acid (*and oxygen)
(ii) the reduction, if any, of their oxides by carbon and/or by hydrogen

- describe the reactivity series as related to the tendency of a metal to form its positive ion, illustrated by its reaction with

(i) the aqueous ions of the other listed metals
(ii) the oxides of the other listed metals

- deduce the order of reactivity from a given set of experimental results
- describe the action of heat on the carbonates of the listed metals and relate thermal stability to the reactivity series
- explain unreactivity of aluminium
- discuss the displacement reaction taking place in Thermit Reaction (aluminium and oxides of iron) and relate to exothermic reaction and its use in railway tracks

Extraction of metals

- describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series
- explain how reactivity of metals determines the methods of extraction
- extraction of metal from its oxides (up till 'iron' with hydrogen and 'zinc' with carbon)

Iron

- describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace
- describe steels as alloys which are a mixture of iron with carbon or other metals and how controlled use of these additives changes the properties of the iron, e.g. high carbon steels are strong but brittle whereas low carbon steels are softer and more easily shaped
- state the uses of mild steel, e.g. car bodies; machinery, and stainless steel, e.g. chemical plants; cutlery; surgical instruments
- describe the essential conditions for the corrosion (rusting) of iron as the presence of oxygen and water; prevention of rusting can be achieved by placing a barrier around the metal, e.g. painting; greasing; plastic coating; galvanising
- describe the sacrificial protection of iron by a more reactive metal in terms of the reactivity series where the more reactive metal corrodes preferentially, e.g. underwater pipes have a piece of magnesium attached to them

Recycling of metals

- describe metal ores as a finite resource and hence the need to recycle metals, e.g. recycling of iron
- discuss the social, economic and environmental issues of recycling metals

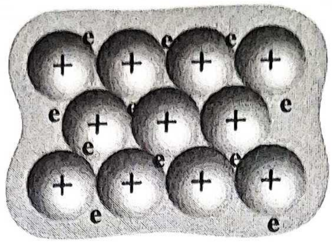
*discuss metal recycling industry in Singapore (e.g. Eco-Recycling Park at Tuas) and other parts of the world;

*discuss how Singapore promotes 3 R's;

*propose plans and programs that can support recycling of metals

1.1 METALS

- Metals are elements which **lose electrons** to form **positive ions** (except hydrogen).
- Metal ions are **closely packed** together in a regular 3-dimensional pattern or lattice, with the valence electrons from each atom **free to move** at random in the space between the metal ions, forming a **'sea' of valence electrons.**



- A metallic bond is the **electrostatic force of attraction** between positive metal ions and 'sea' of free-moving valence electrons.

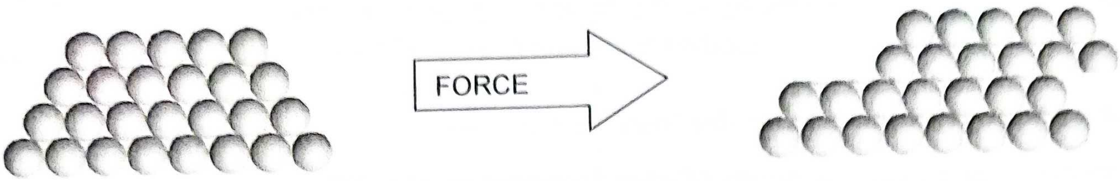
1.2 DIFFERENCES IN PROPERTIES BETWEEN METALS & NON-METALS

(A) PHYSICAL PROPERTIES

Physical Properties	Metals	Non-metals
Examples	Copper, sodium, gold	Hydrogen, carbon, oxygen
Appearance	Are usually shiny (lustrous) and can be polished	Are usually dull and cannot be polished
Melting and Boiling points	Most are solids at room temperature with high melting and boiling points (exception: mercury with melting point of –39°C, Group 1 metals, etc.)	Are usually liquids or gases at room temperature with low melting and boiling points (exception: diamond, silicon and graphite have high melting points due to their giant molecular structures)
Heat and Electrical Conductivity	Are good conductors of heat and electricity	Are poor conductors of heat and electricity (exception: graphite)
Density	Have relatively high densities	Have relatively low densities
Malleability	Malleable – can be bent and beaten into different shapes	Brittle (breaks without bending and stretching)
Ductility	Ductile – can be stretched and pulled into wires	Not ductile

Why are metals ductile and malleable?

- Atoms in pure metals are packed in <u>layers</u> in an <u>orderly</u> manner.
- Atoms in pure metal are of the same size. This allows the **orderly layers** of the **same sized atoms** to **<u>slide over each other easily</u>** when a force is applied. Hence, pure metals are **ductile** and **malleable**.



Why do metals have high melting and boiling points?

- Atoms in a metal are packed tightly in layers and are held together by strong metallic bonds.
- **<u>Large</u>** amount of energy is needed to overcome the **<u>strong</u> electrostatic forces of attraction** between <u>positive metal ions</u> and the <u>'sea' of free-moving valence electrons</u>.

Why do metals conduct electricity?

- Presence of **<u>'sea' of free-moving valence electrons</u>** helps to conduct electricity.

(B) CHEMICAL PROPERTIES

Chemical Properties	Metals	Non-metals
Formation of ions	Form **<u>positive ions</u>** by **<u>losing</u>** electrons. E.g. $Na \rightarrow Na^{+} + e^{-}$ $Mg \rightarrow Mg^{2+} + 2e^{-}$ Metals are good **reducing agents**.	Form **<u>negative ions</u>** by **<u>gaining</u>** electrons. E.g. $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$ $O_2 + 4e^{-} \rightarrow 2O^{2-}$
Formation of oxides	Form **<u>basic</u>** oxides or **<u>amphoteric</u>** oxides by burning / reacting with oxygen. E.g. $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ $2Zn(s) + O_2(g) \rightarrow 2ZnO(s)$	Form **<u>acidic</u>** oxides or **<u>neutral</u>** oxides by burning / reacting with oxygen. E.g. $S(s) + O_2(g) \rightarrow SO_2(g)$ $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$
	<u>Soluble</u> basic oxides dissolve in water to form **<u>alkalis.</u>** E.g. $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$	**<u>Soluble</u>** acidic oxides dissolve in water to form **<u>acids.</u>** E.g. $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

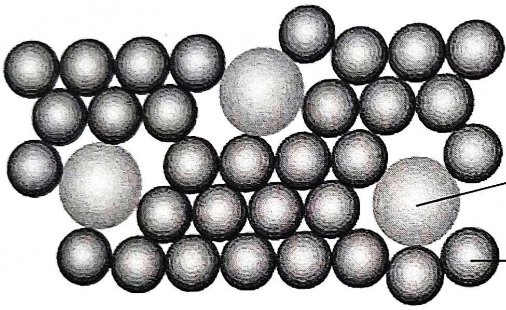
2.1 ALLOYS

- Pure metals are not commonly used as they are soft and many of them may corrode easily by reacting with oxygen and water.
- Most metals are used in the form of alloys because they are harder and stronger than pure metals.
- An alloy is a **mixture** of **a metal with one or more other elements.**
- Some alloys have special properties not found in individual metals, e.g. resistant to corrosion.

Why are alloys stronger and harder than pure metals?

- This is because in alloys, atoms of added element have a **different size** from atoms in the pure metal.
- This **disrupts** the orderly arrangement of atoms in pure metals.
- Hence, it is much **harder** for the atoms to slide over each other when a force is applied.

Structure of an alloy



2.2 COMMON ALLOYS

Alloy	Composition	Property	Uses
Mild steel	iron, carbon	hard and strong	car bodies, machinery, buildings, ships
Stainless steel	iron(73%) chromium(18%) nickel(8%) carbon(1%)	strong and resistant to corrosion	cutlery, hospital and chemical equipment, pipes in chemical industries
Bronze	copper, tin	hard and strong	medals, swords, statues
Brass	copper(70%) zinc(30%)	does not corrode easily; yellow colour like gold	decorative ornaments, musical instruments, pins of power plug
Solder	tin(50%) lead(50%)	low melting point	joining metals
Duralumin	aluminium, copper, magnesium, manganese	light, strong and durable	aircraft parts
Cupronickel	copper, nickel	silvery appearance, unreactive	coins

3.1 REACTIVITY SERIES OF METALS

Metals can be listed in order of their chemical reactivity as given below.

The Reactivity Series of Metals	
Potassium	***Most Reactive*** ↑
Sodium	
Calcium	
Magnesium	
Aluminium	
(Carbon)	
Zinc	
Iron	
Tin	
Lead	
(Hydrogen)	
Copper	
Silver	
Gold	***Least Reactive***

A metal **'high up'** in the reactivity series:

- reacts **vigorously** and **quickly** with chemicals;
- corrodes easily;
- **readily gives up electrons** in reactions to form **positive ions.**

A metal **'low down'** in the reactivity series:

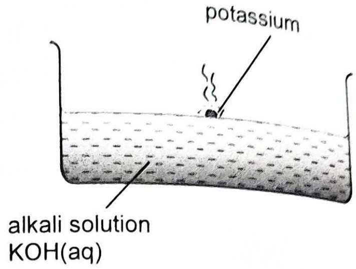
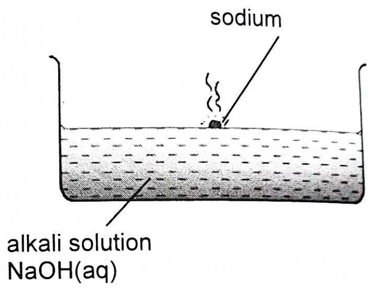
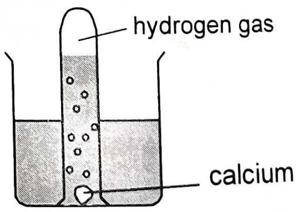
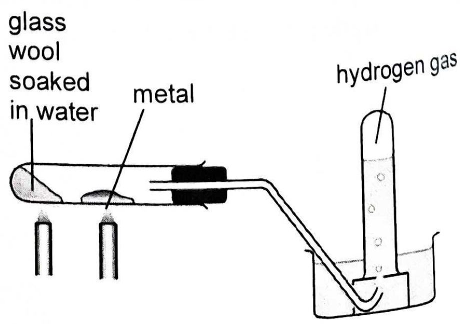
- does **not** react vigorously with chemicals;
- does **not** corrode easily;
- **does not readily give up electrons** in reactions to form **positive ions.**

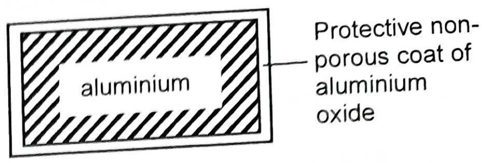
Note:

- A metal **higher up** in the reactivity series has a **greater tendency to lose electrons** to form **positive ions** than a metal lower down in the reactivity series.

- Hydrogen is not a metal. It is placed in the reactivity series as a reference – metals below hydrogen in the series do not react with acids to produce hydrogen gas.

3.2 REACTION OF METALS WITH COLD WATER OR STEAM

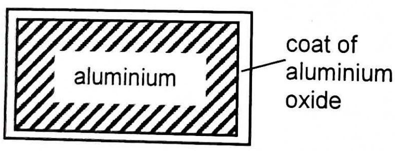
Metal	Reaction with cold water or steam	Chemical Equation (with state symbols)
Potassium (Group 1)	• Reacts **very violently** in **cold water** • Highly exothermic • **Potassium hydroxide solution** and hydrogen gas produced • Potassium **darts rapidly** on the surface of water and **melts.** • **Fizzing** occurs • The hydrogen ignites instantly. The metal is also set on fire, with sparks and a lilac flame	$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$ 
Sodium (Group 1)	• Reacts **violently** in **cold water** • **Sodium hydroxide solution** and hydrogen gas produced • Sodium **darts** rapidly on the surface of water and **melts** into silvery ball • **Fizzing** occurs • Sodium may catch fire and burn with yellow flame Note: If indicator is added: • **green** universal indicator turns **violet** • red litmus turns blue	$2Na(s)+2H_2O(l) \rightarrow 2NaOH(aq)+ H_2(g)$ 
Calcium (Group 2)	• Reacts **readily** with **cold water** • Rapid effervescence seen, colourless and odourless gas evolved • **Calcium hydroxide solution** and hydrogen gas produced • Colourless solution soon turns cloudy as calcium hydroxide is only slightly **soluble** in water	$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq)+H_2(g)$ 
Magnesium (Group 2)	• Reacts **very slowly** with **cold water** with a few bubbles of hydrogen gas produced • Hot magnesium **reacts violently** with **steam** • **Magnesium oxide** and hydrogen gas produced • Silvery/grey magnesium solid turns **white** and bright **white** glow is produced during the reaction	$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$ 

Metal	Reaction with cold water or steam	Chemical Equation (with state symbols)
***Aluminium** (Group 13)	• Aluminium metal does not seem to react with water or steam • This is because aluminium is covered with a thin layer of **<u>non-porous aluminium oxide</u>** which adheres strongly to the metal and protects it from reacting	Formation of non-porous protective layer of Al_2O_3: $$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$$ 
Zinc	• Does not react with cold water • Hot zinc **<u>reacts readily</u>** with **<u>steam</u>** • **<u>Zinc oxide</u>** and hydrogen gas produced • Grey zinc solid slowly becomes coated with a **<u>yellow</u>** solid which becomes **<u>white</u>** when cold	$Zn(s) + H_2O(g) \rightarrow ZnO(s) + H_2(g)$
Iron	• Does not react with cold water • Red hot iron **<u>reacts slowly</u>** with **<u>steam</u>** • **<u>Iron(II,III) oxide (Fe_3O_4)</u>** and hydrogen gas formed • Grey iron solid turns red hot on heating and black when cooled	$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$
Lead **Copper** **Silver** **Gold**	**<u>No reaction</u>** with **cold water or steam**	

3.3 REACTION OF METALS WITH DILUTE ACIDS

- Metals below hydrogen in the reactivity series do not react with acids to produce hydrogen gas.
- Metals above hydrogen in the reactivity series can displace hydrogen gas from dilute acids.

Reactive metal + Acid → Salt + Hydrogen

Metal	Reaction with dilute hydrochloric acid	Chemical Equation (with state symbols)
Potassium (Group 1)	• Reacts **explosively** *(Should **NOT** be carried out in school laboratory)*	**$2K(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2(g)$**
Sodium (Group 1)		**$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$**
Calcium (Group 2)	• Reacts **violently** • Very rapid effervescence seen, colourless and odourless gas evolved • Colourless calcium chloride solution and hydrogen are formed	**$Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$**
Magnesium (Group 2)	• Reacts **rapidly** • Rapid effervescence seen, colourless and odourless gas evolved • Colourless magnesium chloride solution and hydrogen are formed	**$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$**
***Aluminium** (Group 13)	• Aluminium does not seem to react at first as it is covered by a layer of non-porous **aluminium oxide** which adheres strongly to the metal  • However, hydrochloric acid will react with aluminium oxide to form **soluble aluminium chloride** and **water** • Once the oxide layer is reacted away, the reaction between the aluminium and acid is **rapid** as the metal is **reactive**	**$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$** **protective oxide layer** **$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$** **($Al$ underneath exposed once Al_2O_3 layer is reacted away)**
Zinc	• Reacts **moderately fast** • Steady effervescence seen, colourless and odourless gas evolved • Colourless zinc chloride solution and hydrogen are formed	**$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$**

Metal	Reaction with dilute hydrochloric acid	Chemical Equation (with state symbols)
Iron	• Reacts **slowly** • Slow effervescence seen, colourless and odourless gas evolved • **Pale green iron(II) chloride solution** and **hydrogen** are formed	$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
***Lead** (Group 14)	• Reacts **very slowly** and reaction stops before completion • **Insoluble** lead(II) chloride salt formed will coat onto surface of the lead metal and prevent further reaction between the acid and lead metal	$Pb(s) + 2HCl(aq) \rightarrow PbCl_2(s) + H_2(g)$
Copper	No reaction with dilute hydrochloric acid	
Silver		
Gold		

Note:

1. Most metals do react with dilute nitric acid but the products formed depend on the concentration of the acid and the nature of the metals. Oxides of nitrogen (NO, NO_2) may be formed instead of hydrogen gas.

   E.g. $Zn(s) + 4HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

2. Dilute nitric acid can oxidise unreactive metals such as copper.

   E.g. $3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$

Summary

1. Reaction of Metal with Water/Steam

Reactive metal + **cold water** → **metal hydroxide solution (alkali)** + **hydrogen**
(K, Na, Ca)

Note: Magnesium does react with cold water but very slowly.

Heated moderately reactive metal + **steam** → **metal oxide (insoluble)** + **hydrogen**
(Mg, Zn, Fe)

Unreactive metals (Pb, Cu, Ag, Au) will **NOT** react with cold water or steam

2. Reaction of Metal with Dilute Hydrochloric Acid

Metal + **Dilute hydrochloric acid** → **metal chloride** + **hydrogen gas**
(above hydrogen in the reactivity series)

Quick Check:

1. The table shows the results of adding metals, **P**, **Q** and **R** to dilute hydrochloric acid and to water.

Metal	Dilute hydrochloric acid	Water
P	Hydrogen produced	Hydrogen produced
Q	Hydrogen produced	No Reaction
R	No reaction	No reaction

Rank the above metals from the most reactive to the least reactive.

____

2. The following shows a list of metals in order of their reactivity series.

Most Reactive ————→ **Least Reactive**
Calcium Magnesium Zinc Iron Lead Copper

(a) Which of the metals will react most quickly with cold water? _______________

Equation: _______________

(b) Name the metal which will react with steam but only react slowly with cold water.

____

(c) Name the gas produced during the reaction in (a) and (b). Describe how you would test and identify this gas.

Gas is _______________

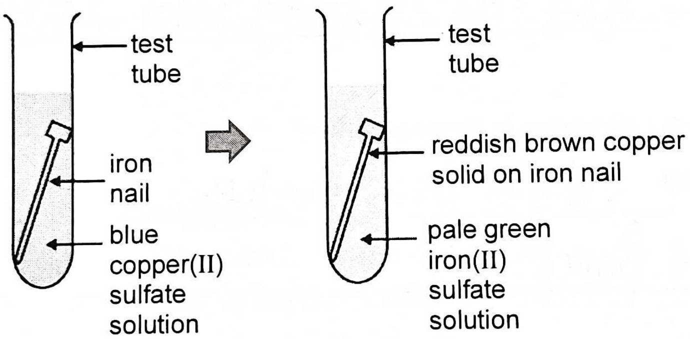
Test: _______________

4.1 DISPLACEMENT OF METALS FROM METAL SALT SOLUTIONS

A more reactive metal will displace the less reactive metal from its salt solution.

- A more reactive metal **loses electrons more readily** to form **positive** ions.
- Ions of the less reactive metal accept the electrons from the more reactive metal to form the metal atoms.
- Hence, a more reactive metal will displace the less reactive metal from its salt solution.
- The **further apart** the two metals in the reactivity series, the **easier/faster** the displacement reaction.
- Such displacement reaction is often accompanied by
  - change in **appearance(colour) of metal**,
  - change in **colour of salt solution**,
  - heat change (heat evolved).

Example:
Experiment: An iron nail is placed in copper(II) sulfate solution.



Equation: $Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$

Observations:

(1) The iron nail is coated with a **reddish brown solid**.
(2) The **blue** copper(II) sulfate solution turns **pale green** after some time.

Explanation:
- Iron, being more **reactive** than copper, **loses** electrons more readily than copper.
- Iron displaces **copper** from blue copper(II) sulfate solution. Reddish brown copper solid and pale green iron(II) sulfate solution are formed.

Ionic Equation: $Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$

Half Equations: $Fe \rightarrow Fe^{2+} + 2e^-$
oxidised

$Cu^{2+} + 2e^- \rightarrow Cu$
reduced

Qn: Explain why this is a redox reaction in terms of electrons transfer.
Fe is oxidised as Fe loses electrons to form Fe^{2+}.
$CuSO_4$ is reduced as Cu^{2+} gains electrons to form Cu.

Quick Check:

1. Write the chemical and ionic equations (including state symbols) for the following reactions (if any).

(a) zinc and silver nitrate solution

Equation: $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

Ionic Equation: $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$

→ Copper cannot displace magnesium as copper is less reactive than magnesium

(b) copper and magnesium chloride solution

Equation: $Cu(s) + MgCl_2(aq) \rightarrow$ No reaction

Ionic Equation:

(c) magnesium and iron(II) sulfate solution

Equation: $Mg(s) + FeSO_4(aq) \rightarrow Fe(s) + MgSO_4(aq)$

Ionic Equation: $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$

What would you observe in this reaction?

1) Pale green solution turns colourless.

2) Grey solid formed

(d) zinc and copper(II) sulfate solution

Equation: $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

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

What would you observe in this reaction?

1) Blue copper(II) sulfate solution turns colourless

2) Reddish brown solid formed.

2. Metal **X** (more reactive) displaces metals **Y** and **Z** from solutions of their salts. Metal **Z** displaces metal **Y** from solutions of its salt, but not metal **X** from a solution of its salt.

Place **X**, **Y** and **Z** in order of their reactivity.

~~X~~ X, Z, Y decreasing | Y, Z, X increasing

4.2 DISPLACEMENT OF METALS FROM METAL OXIDES

A more reactive metal will displace a less reactive metal from the metal oxide.

- If a more reactive metal is thoroughly mixed with the **oxide** of a less reactive metal (both in **powder form**), and **strongly heated**, the more reactive metal will displace the less reactive metal from the oxide.
- This type of reaction is called **thermit reaction**, because of the great amount of heat released.

Example: Thermit reaction between aluminium and iron(III) oxide

Equation: $2Al(s) + Fe_2O_3(s) \xrightarrow{\text{heated strongly}} Al_2O_3(s) + 2\,Fe(l)$

Reducing Agent

↳ molten iron

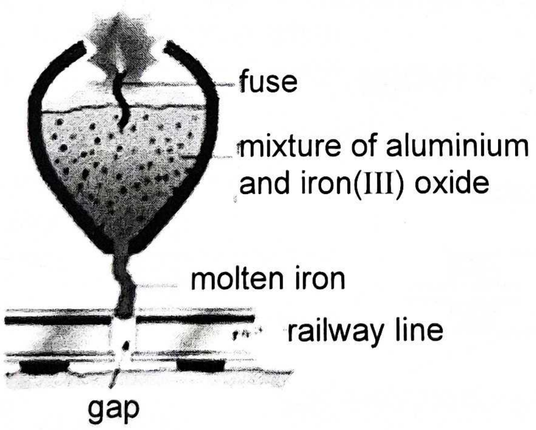
Observation: Reaction is very **exothermic** with a lot of **heat** and **light** energy given out.

Explanation:

- Aluminium, being more reactive than iron, loses electrons more readily than iron.
- Aluminium displaces iron from iron(III) oxide. Molten iron and aluminium oxide are formed.

Application:

- This reaction is used to produce iron to weld railway lines together.
- The reaction mixture is placed around the gap between railway lines and ignited with a fuse.
- The molten iron formed flows into the gap to weld the rails together.



4.3 REDUCTION OF METAL OXIDES WITH NON-METALS

(i) Reduction with Carbon

- Carbon can remove oxygen from the oxides of some metals and **reduce** the oxides to metals.

metal oxide + carbon $\xrightarrow{\text{heat}}$ metal + carbon dioxide

Carbon reduces oxide, not the metal

Example: **$2CuO(s)$ + $C(s)$ → $2Cu(s)$ + $CO_2(g)$**

- The **more reactive** the metal, the **harder it is for carbon to take oxygen away** from its oxides.
- Oxides of metals that are above **zinc** in the reactivity series are often not reduced by carbon.
- This is an important reaction in chemical industry as metals that are below aluminium in the reactivity series are often extracted from their ores by reduction of metal oxide with carbon.

(ii) Reduction with Hydrogen

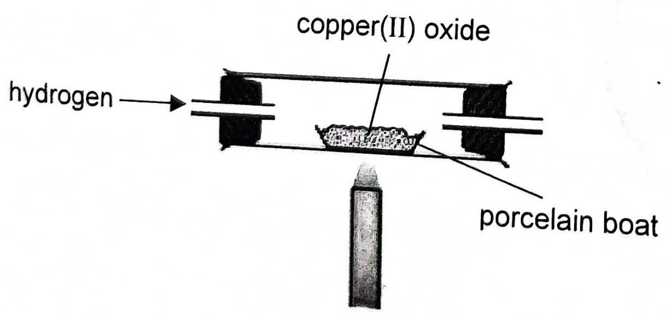
- Hydrogen can also be used for reducing some metal oxides to their metals.

metal oxide + hydrogen $\xrightarrow{\text{heat}}$ metal + water

- The **more reactive** the metal, the **harder it is for hydrogen to take oxygen away** from its oxides.
- Oxides of metals that are above **iron** in the reactivity series are often not reduced by hydrogen.

Example: Reduction of copper(II) oxide by hydrogen.

Equation: **$CuO(s)$ + $H_2(g)$ → $Cu(s)$ + $H_2O(g)$**



Observation: The **black** copper(II) oxide solid turns **red-brown/ pink**.

Note:

- At the end of experiment, hydrogen is still passed into the combustion tube until the apparatus cools down.
- This is to prevent the **air** from entering the hot combustion tube as **oxygen** will react with hot copper to form back copper(II) oxide.

5 EFFECT OF HEAT ON METAL CARBONATES

- Most carbonates decompose when heated, producing **metal oxide** and **carbon dioxide**.

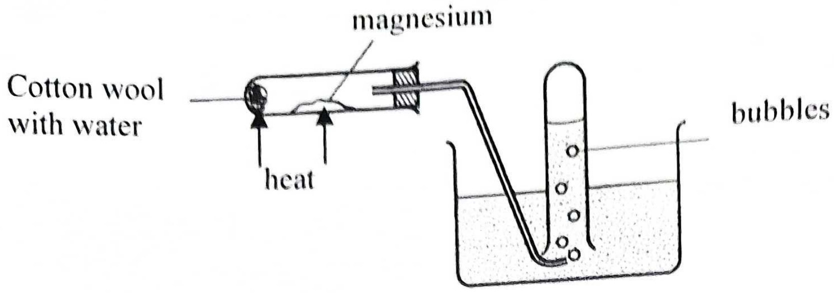
Metal carbonate	Condition	Equation
K_2CO_3 Na_2CO_3	Not decomposed by heat i.e. carbonate stable to heat	
$CaCO_3$	Very strong heating	$CaCO_3 \rightarrow CaO + CO_2$
$MgCO_3$ $ZnCO_3$ $FeCO_3$ $PbCO_3$	Moderate heating	$MgCO_3 \rightarrow MgO + CO_2$ $ZnCO_3 \rightarrow ZnO + CO_2$ $FeCO_3 \rightarrow FeO + CO_2$ $PbCO_3 \rightarrow PbO + CO_2$
$CuCO_3$ Ag_2CO_3	Gentle heating Note: • silver carbonate decomposes to give silver, oxygen and carbon dioxide. • This explains why silver carbonate must be kept in a cool place	$CuCO_3 \rightarrow CuO + CO_2$ *green* *black* $2Ag_2CO_3 \rightarrow 4Ag + O_2 + 2CO_2$

Note: $Al_2(CO_3)_3$ and $Fe_2(CO_3)_3$ do not exist.

- The **more reactive** the **metal**, the **more thermally stable** is the **carbonate**; i.e. the metal carbonate is less easily decomposed.
- A greater amount of energy (higher temperature) is required to decompose the compound of a more reactive metal.
- However, **potassium carbonate** and **sodium carbonate do not decompose** even upon heating at high temperature.

Self Check Exercise

1 The diagram below shows the reaction involving magnesium.



What are the products for the reaction above?

A Magnesium hydroxide and hydrogen
B Magnesium hydroxide, hydrogen and steam
C Magnesium oxide and hydrogen
D Magnesium oxide, hydrogen and steam

(C)

2 Experiments were carried out to determine the positions of metals **X**, **Y** and **Z** in the reactivity series. The table shows the results.

most unreactive most reactive

Test	Metal **X**	Metal **Y**	Metal **Z**
Reaction between metal and hydrochloric acid	Effervescence observed	No visible observation	Effervescence observed
Heating the metal oxide with carbon	Metal oxide reduced	Metal oxide reduced	Metal oxide not reduced

Place the three metals in order of decreasing reactivity.

A **X, Z, Y**
B **Y, X, Z**
C **Z, X, Y**
D **X, Y, Z**

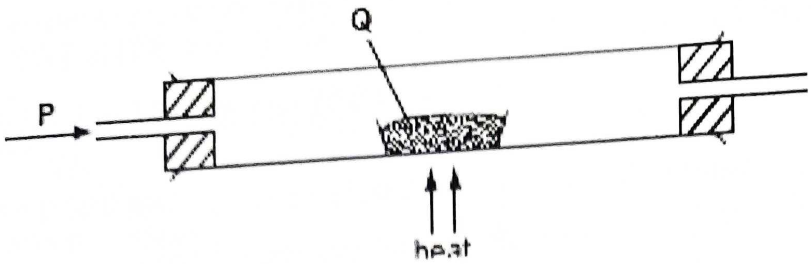
(C)

3 Element **U** displaces element **V** from the aqueous nitrate of **V**.
Element **W** reacts with cold water to give hydrogen.
Element **U** gives hydrogen only when heated with steam.
What could elements **U**, **V** and **W** be?

	U	V	W
A	calcium	copper	silver
B	copper	zinc	sodium
C	silver	copper	calcium
D	zinc	copper	calcium

(D)

4 In the apparatus shown, gas **P** is passed over solid **Q**



No reaction occurs if **P** and **Q** are

	P	**Q**
A	hydrogen	lead(II) oxide
B	hydrogen	calcium oxide
C	oxygen	carbon
D	oxygen	sulfur

(B)

5 Which carbonate takes the shortest time to decompose?

A K_2CO_3
B $CaCO_3$
C $ZnCO_3$ → more reactive than Cu
D $CuCO_3$

(D)

<table>
<tr><th>Qn</th><th>Ans</th><th>Explanation</th></tr>
<tr><td>1</td><td>C</td><td>Hot magnesium reacts with steam to produce magnesium oxide and hydrogen.
$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$</td></tr>
<tr><td>2</td><td>C</td><td>
<table>
<tr><th>Test</th><th>Metal X</th><th>Metal Y</th><th>Metal Z</th></tr>
<tr><td>Reaction between metal and hydrochloric acid</td><td>Effervescence observed (shows that X is above hydrogen in the reactivity series)</td><td>No visible observation (shows that X is below hydrogen in the reactivity series, thus metal Y is the least reactive among the 3 metals)</td><td>Effervescence observed (shows that Z is above hydrogen in the reactivity series)</td></tr>
<tr><td>Heating the metal oxide with carbon</td><td>Metal oxide reduced</td><td>Metal oxide reduced</td><td>Metal oxide not reduced (shows that metal Z is the most reactive metal as carbon is unable to take the oxygen away from Z oxide)</td></tr>
</table>
</td></tr>
</table>

3	D	Element **U** displaces element **V** from the aqueous nitrate of **V** → U is more reactive than V. Element **W** reacts with cold water to give hydrogen → W is a highly reactive metal Element **U** gives hydrogen only when heated with steam → U is a moderately reactive metal.

	U	V	W
A	calcium (calcium reacts with cold water to give hydrogen)	copper	silver (not a highly reactive metal)
B	copper (copper does not react with steam, copper is less reactive than zinc)	zinc	sodium
C	silver (silver does not react with steam)	copper	calcium
D	zinc	copper	calcium

4	B	A: hydrogen is able to reduce lead(II) oxide B: hydrogen is unable to reduce calcium oxide C: oxygen reacts with carbon to give carbon dioxide D: sulfur reacts with oxygen to give sulfur dioxide
5	D	The more reactive the metal, the more thermally stable is the carbonate. Since the reactivity of K> Ca> Zn> Cu, the thermal stability of K_2CO_3> $CaCO_3$ > $ZnCO_3$ > $CuCO_3$. K_2CO_3 does not decompose. Time taken for the carbonates to decompose decreases in this order. K_2CO_3> $CaCO_3$ > $ZnCO_3$ > $CuCO_3$. Therefore, $CuCO_3$ decomposes the fastest.

Metals (Part 2)

6.1 THE EXTRACTION OF METALS

- Most metals are found in the ground as **compounds** in the form of **ores.**
- These ores are solids that miners dig from earth to extract the metals.
- Ores are **compounds** of metals (usually in the form of oxides, sulfides, chlorides and carbonates) mixed with large amount of impurities such as earth and rock.
- Extraction of metal from its ore involves:
  (1) getting rid of the impurities to obtain concentrated form of the mineral.
  (2) obtaining metal from the mineral. The method to extract the metal depends on the **position of the metal in the reactivity series**.
- There are two main methods for extracting metals:
  (1) **Electrolysis**
  (2) **Reduction of metal oxide using reducing agents** such as carbon etc.

METHODS OF EXTRACTING METALS

	Metal	Method to extract metals
↑ Increasing reactivity of metals	Potassium Sodium Calcium Magnesium Aluminium	**Very reactive metals** • Always exist as compounds with **strong chemical bonds**. • **Electrolysis** is used to break down the **molten ionic compounds** to obtain the metals. • Electrolysis is an expensive method to extract metals as large amount of electricity is used.
	Zinc Iron Tin Lead Copper Silver	**Moderately reactive metals** • Metals usually found as oxides or sulfides. • Usually extracted by **reducing their oxides with carbon.** • The sulfides are first heated in air to form oxides. • The oxides are then heated with carbon to obtain the metal. E.g. Extraction of zinc from its ore containing zinc sulfide $2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$ $2ZnO + C \rightarrow 2Zn + CO_2$
	Gold Platinum	**Unreactive metals** • Gold and platinum are found naturally **uncombined** in the ground as **elements**.

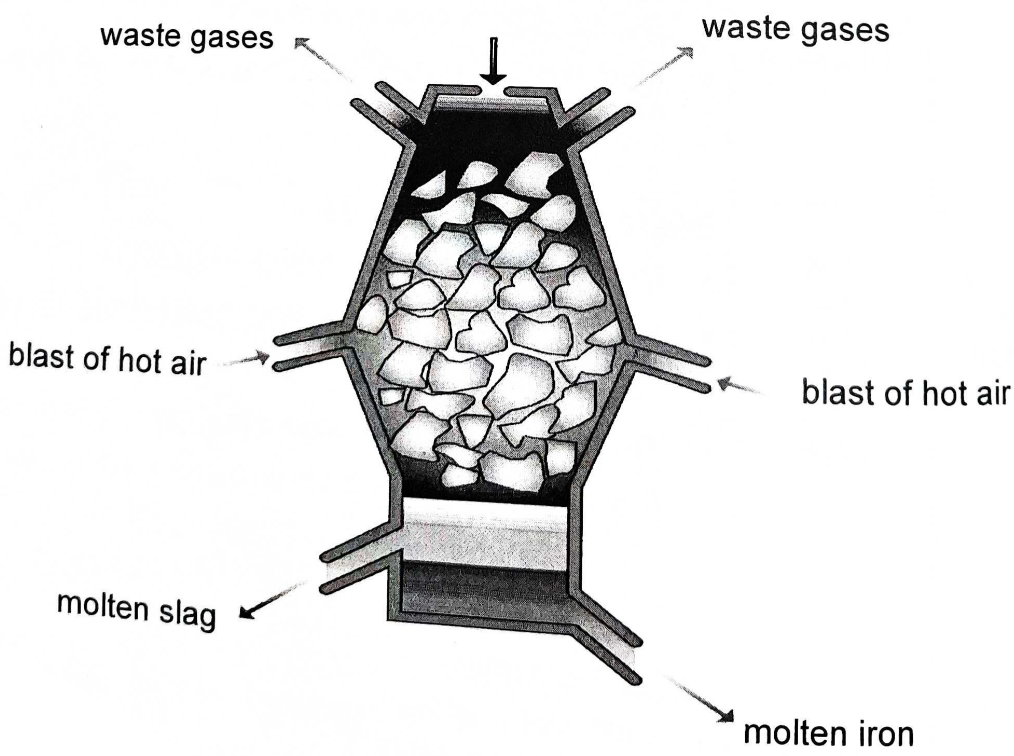
6.2 THE EXTRACTION OF IRON

- The main ore of iron is **haematite**.
- Haematite contains mainly **iron(III) oxide (Fe_2O_3)** mixed with impurities such as sand and clay (mainly silicon dioxide), phosphorus and sulfur.
- Iron is extracted from haematite in the **blast furnace**.
- The 3 raw materials introduced into the top of the furnace are:

Common Name	Chemical Name	Formula
(a) haematite	**iron(III) oxide**	Fe_2O_3
(b) coke	**carbon**	**C**
(c) limestone	**calcium carbonate**	$CaCO_3$

- **Hot air** is pushed into the furnace **near the bottom**.

1. haematite
2. coke
3. limestone



The following chemical reactions take place in the blast furnace.

(a) <u>The production of carbon dioxide</u>

(i) The coke burns in air to produce **carbon dioxide** and a lot of **heat**. The heat from the combustion heats the furnace.

Equation: $C(s) + O_2(g) \rightarrow CO_2(g)$ ΔH = **negative**

(ii) The heat produced in (a)(i) helps to **decompose** the limestone to produce **carbon dioxide** and **calcium oxide (quicklime)**

Equation: $CaCO_3(s) \rightarrow CO_2(g) + CaO(s)$
limestone quicklime

Note: The purpose of the reaction in (a)(ii) is to obtain calcium oxide which is used to remove impurities as shown in (d) below.

(b) <u>The production of carbon monoxide</u>

- As the carbon dioxide rises up the furnace, it reacts with more **coke** to produce **carbon monoxide**.

Equation: $CO_2(g) + C(s) \rightarrow 2CO(g)$ ΔH = **positive**
excess coke reducing agent

(c) <u>The reduction of haematite</u>

- The carbon monoxide reacts with the iron(III) oxide to produce iron.

+3 -6
Equation: $3CO(g) + Fe_2O_3(s) \rightarrow 2Fe(l) + 3CO_2(g)$

- In this reaction, iron(III) oxide loses oxygen, we say that **iron(III) oxide** is **reduced.**
- The carbon monoxide acts as a **reducing agent** as carbon monoxide **gains** oxygen from iron(III) oxide and **reduces** iron(III) oxide to iron.
- The iron formed is **molten** and runs to bottom of the furnace.

(d) <u>The removal of impurities</u>

- The main impurities in iron ores are sand and clay which consist primarily of silicon dioxide. The formula of silicon dioxide is SiO_2.
- Silicon dioxide is an **acidic** oxide.
- The limestone added first breaks down in presence of heat to form **calcium oxide** (**basic** oxide). - see (a)(ii)

Calcium oxide is able to react with silicon dioxide to produce molten **slag** which is mainly **calcium silicate**.

Basic oxide   acidic oxide

Equation: $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$

- The slag **floats** on top of the molten iron.
- The slag and iron are tapped off **separately** from the blast furnace.
- The iron extracted from the blast furnace is known as **pig iron**.
- Main use of slag: **to make road surfaces**

Summary

Extraction of iron from iron ore (haematite) in the blast furnace

To remove oxygen from iron(III) oxide	To remove impurities (acidic oxides) in iron ore
1. Combustion Reaction (Exothermic) $C(s) + O_2(g) \rightarrow CO_2(g)$ 2. Formation of CO using excess coke $CO_2(g) + C(s) \rightarrow 2CO(g)$ 3. Reduction of iron(III) oxide $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(l) + 3CO_2(g)$	1. Decomposition (Endothermic) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ (basic oxide) 2. Neutralisation $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$ (sand) (acidic oxide)

(e) The removal of waste gases

- Hot waste gases containing mainly **nitrogen** (from hot air), **carbon dioxide** and unreacted **carbon monoxide**, escape through the top of the furnace.

Quick Check:

1 In the blast furnace, which reaction produces most of the carbon monoxide used to extract iron?

A Burning coke in air.
B Reacting coke with carbon dioxide.
C Reacting iron oxide with coke.
D Decomposition of limestone.

(B) ✓

2 Which statement about the production of iron from haematite in a blast furnace is correct?

A The iron(III) oxide is reduced by carbon dioxide.
B Limestone is added to remove basic impurities. ✗
C Slag floats on molten iron at the furnace base. ✗
D Decomposition of limestone is exothermic.

(D) ✗ C

3 Which reaction is **not** a step in the production of iron from haematite in the blast furnace?

A Carbon(coke) burning in air to produce carbon dioxide

B Carbon reacting with carbon dioxide to produce carbon monoxide

C Iron(III) oxide reacting with carbon monoxide to form iron

D Iron reacting with limestone to produce slag

(D)

6.3 TYPES OF STEEL AND THEIR USES

- The **pig iron** obtained from the blast furnace is **not very useful** as it still contains impurities which make iron weak and brittle.
- Most of these impurities need to be removed before iron is converted to steel.
- Steel is an **alloy of iron** and **carbon** with small amount of **other elements**. Steel is stronger and tougher than iron.
- Different types of steel with **different physical properties** are made by
  - varying the amount of **carbon**
  - adding different **elements**

Type of steel	mild steel (low carbon steel)	high carbon steel	stainless steel
Composition	iron 0.25% carbon	iron 0.45-1.5% carbon	iron, carbon, chromium and nickel
Special properties	hard, strong and **malleable**	strong but **brittle**	durable and highly resistant to **corrosion**
Uses	car bodies and machinery	tips of drill bits knives, hammers, chisels, cutting & boring tools	cutlery, surgical and dental instruments

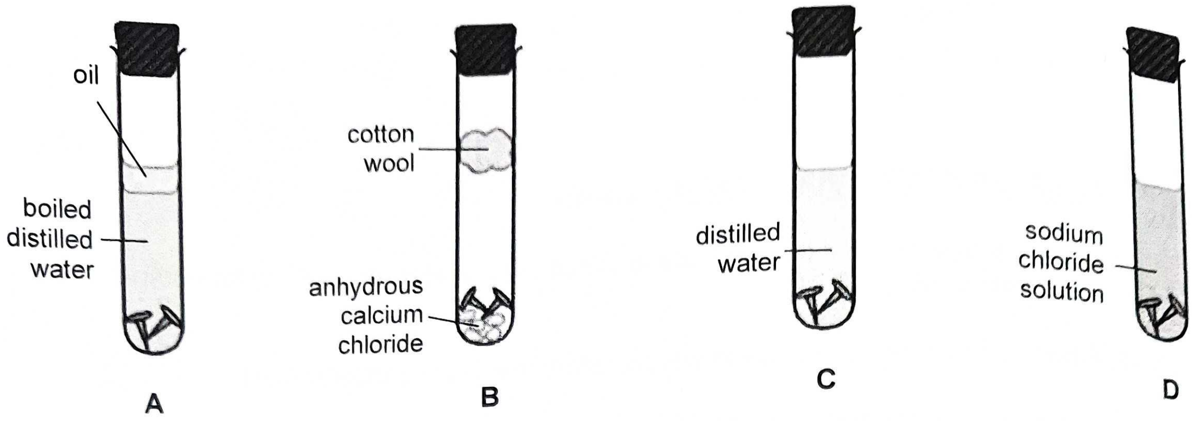
7.1 RUSTING

- Rust is a **reddish-brown** substance that forms slowly on the surface of an object made of **iron** when the object is exposed to **air** and **moisture** for some time. It consists mainly the compound **hydrated iron(III) oxide**.
- Rusting is the slow **oxidation** of iron to form **hydrated iron((III) oxide.**
- Overall chemical equation for rusting:

$$4Fe + 3O_2 + 2x\,H_2O \rightarrow 2Fe_2O_3\,xH_2O$$

Conditions for Rusting

- The following experiment shows that water and oxygen are essential conditions for rusting.



Results after one week

Test tube	Condition(s)	Result
A	**Water is present.** Boiled distilled water has no oxygen and oil stops oxygen in air from entering.	No rusting
B	**Air (oxygen) is present.** Anhydrous calcium chloride, a drying agent removes moisture in the test-tube.	No rusting
C	**Air (oxygen) and water are present.**	Nails rusted
D	**Air (oxygen) and water are present.** Sodium chloride is also present.	Nails rusted severely

Conclusion

- Both **air (oxygen)** and **water** are needed for rusting to occur.
- Presence of sodium chloride **increases** the speed of rusting.

Note:

- Besides sodium chloride, **acidic substances such as sulfur dioxide and carbon dioxide also speed up the rusting process.**
- **Iron objects near the sea and in industrial areas corrode very rapidly** because of the presence of sodium chloride and other acidic pollutants such as sulfur dioxide and nitrogen dioxide that speed up the rusting process.

7.2 RUST PREVENTION

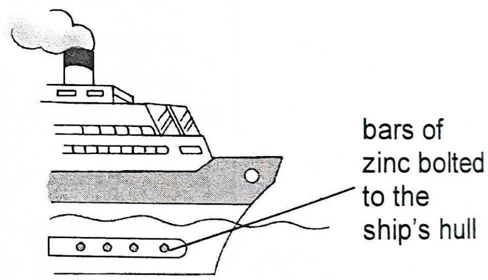
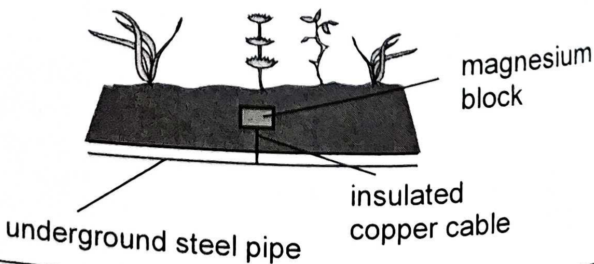
- Rust is very brittle and flaky.
- When iron corrodes, the rusted surface of metal flakes away. This produces new surface to corrode. Eventually all the metal will rust and flake away.
- There are three main methods of rust prevention:

  (1) **Using a protective layer**
  - Coat the iron surface with a protective layer (e.g. paint and grease) that acts as a barrier to air and water

  (2) **Using a sacrificial metal**
  - Protect iron against rusting by using a more reactive metal, e.g. zinc

  (3) **Using alloys**
  - Use of rust-resistant alloys, e.g. stainless steel

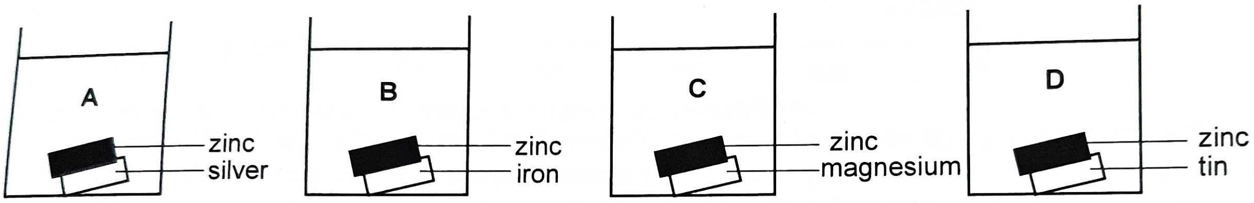
Types of rust prevention	Method	Where it is used	Remarks
1. Using a protective layer	**Painting**	Large iron and steel objects such as cars, ships, bridges, steel pipes etc.	If the protective layer is removed, rusting occurs as oxygen and water can reach the surface of iron and steel.
	Oiling & greasing	Tools and machine parts	
	Plastic coating	Clothes hanger, kitchenware such as draining racks	
	Tin plating	Food cans	• If the tin layer is damaged or scratched, oxygen and water can reach the iron, iron will rust very quickly. • This is because **iron, being more reactive than tin, loses electrons more readily than tin.** Iron will corrode very quickly.
	Chrome plating	Water taps, car bumpers, bicycle handle bar	• Iron can be coated with chromium using electrolysis. • This gives an attractive appearance as well as rust protection. • Chromium is resistant to corrosion as it has a protective **coating of hard and non-porous chromium(III) oxide, Cr_2O_3** on its surface.

Type of rust prevention	Method	Where it is used	Remarks
	Zinc plating (Galvanising)	Roof of houses, dustbins, water buckets etc.	• Zinc has a fairly low melting point, easily applied by dipping iron or steel in molten zinc. • If zinc layer is damaged, iron **will not rust**. • **Zinc, being more reactive than iron, loses electrons more readily** than iron. • Hence **zinc will corrode in place of iron.** This type of rust prevention is called **sacrificial protection**.
2. Using a sacrificial metal (Sacrificial Protection)	**Attaching metal blocks such as zinc or magnesium**	Steel hulls of ships	• Blocks / bars of zinc are attached to the ship's hulls. • Sea water and air can reach the steel but **zinc being more reactive than iron** in steel, **loses electrons more readily than iron.** • **Zinc will corrode in place of iron** and protect iron from rusting. 
		Underground steel pipes	• Magnesium, being **more** reactive than iron in steel pipe, **loses** electrons more readily than iron. • Hence **magnesium will corrode in place of iron** and protect iron from rusting. 

Type of rust prevention	Method	Where it is used	Remarks
3. Using rust resistant alloys	**Using stainless steel**	Cutlery, surgical and dental instruments	Stainless steel is an alloy of iron and carbon with quite a large amount of **chromium** and **nickel**.

Quick Check:

Each beaker contains two strips of metal fastened together and immersed in dilute hydrochloric acid. All the strips are of the same size.



After five minutes, which beaker will contain the least zinc ions in the solution? Explain your answer.

....

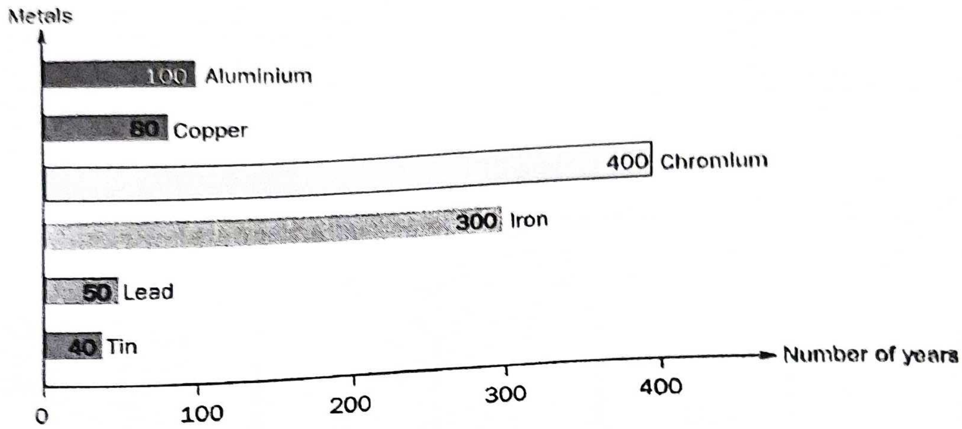
....

....

8.1 RECYCLING

- Metals are **finite** resources. This means that the amounts of the various metal ores in the earth are **limited**.

- The figure shows the estimated time for the reserves of some metals to run out.

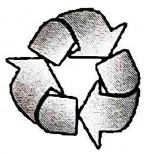


- The chart shows that mining of some metals from the Earth's surface will last longer than others. There is enough iron available to last for centuries but serious shortages of metals such as copper and tin could occur before the end of this century.

- Hence it is important to make the metals last longer by either recycling them or replacing the metals with other materials.

- **Recycling** metals means **reusing** them by **collecting** and **melting down** scrap metal to **make blocks of clean fresh metal** which can be then used to make new objects.

8.2 ADVANTAGES OF RECYCLING



1. **Natural Resource Conservation**

   Recycling metals allows us to conserve the finite metal resources on the Earth and make them last longer.

2. **Energy Conservation**

   Recycling scrap metal requires significantly less energy than manufacturing new metals from their metal ores.

   Example:
   Recycling aluminium saves 95% of the cost of extracting new metal as expensive electrolysis is not needed to decompose its ore. Hence, aluminium soft drink cans are separated from rubbish for recycling.

3. **Environmental Benefits**

- Substantial **reduction in greenhouse gas emissions** through practice of recycling metals as less fossil fuels are burned.
- Recycling metals **cuts down on mining**. Mining of metal ores generates a large amount of waste materials. If the waste is not disposed carefully, the waste may leach into soil and nearby water bodies, polluting land and water. Mined land cannot support plant and animal life.
- Recycling metals **reduces waste disposal problem** - e.g. less piles of rusting cars / less landfill waste sites needed.

8.3 ISSUES OF RECYCLING

1. **Economic Issues of recycling**

Recycling metals can be an **extremely expensive operation.**

For cheap metals, recycling can be more costly than extracting the metals. This is because of the cost incurred in sorting, transporting and cleaning the scrap metal.

2. **Social Issues of recycling**

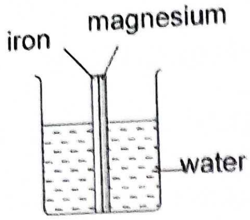
It takes time and effort for communities to adopt recycling as a lifestyle, so it might not be immediately effective.

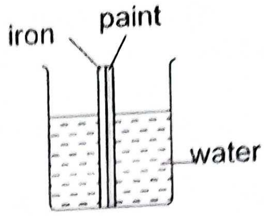
3. **Environmental Issues**

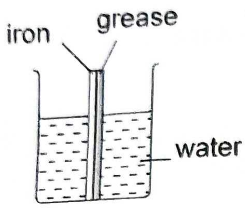
Recycling can cause pollution problems.

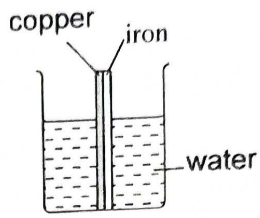
Example:
Car batteries contain a lot of lead. When old batteries are smelted down to extract and recycle lead. The smelting process can send a lot of toxic lead fumes into the air, resulting in polluting the environment.

4 Four beakers were set up as shown and left for a week. Each piece of iron was protected on one side by a different coating. In which beaker is the iron least likely to rust?

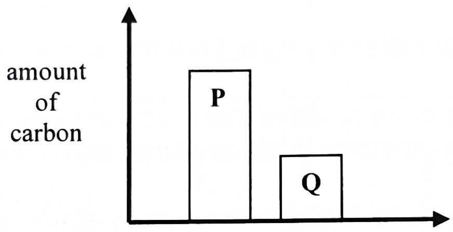
A


B


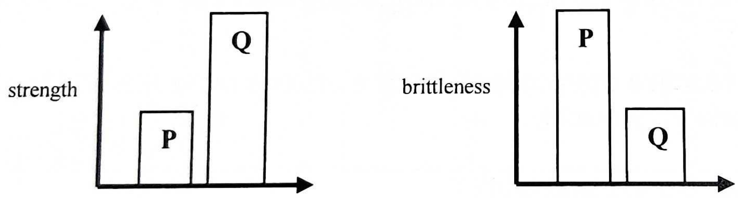
C


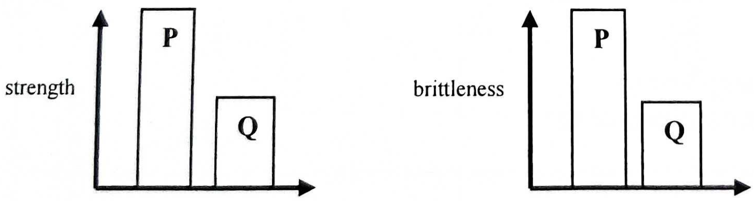
D


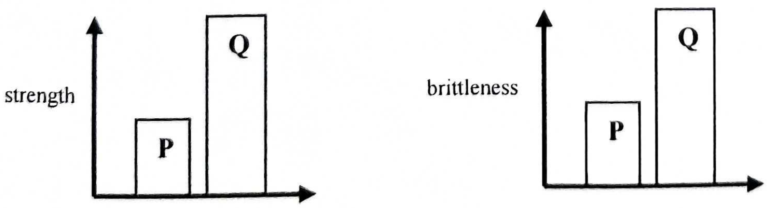
5 The diagram compares the amount of carbon in two types of steel, P and Q.

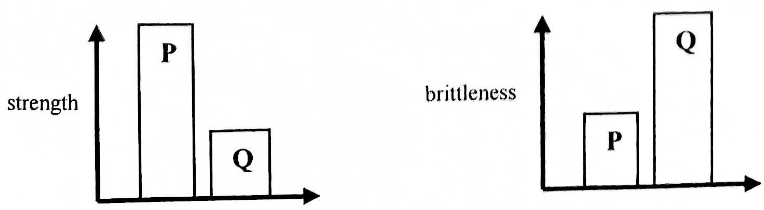


Which two diagrams correctly compare the strength and brittleness of P and Q?

A


B


C


D


Self Check Exercise

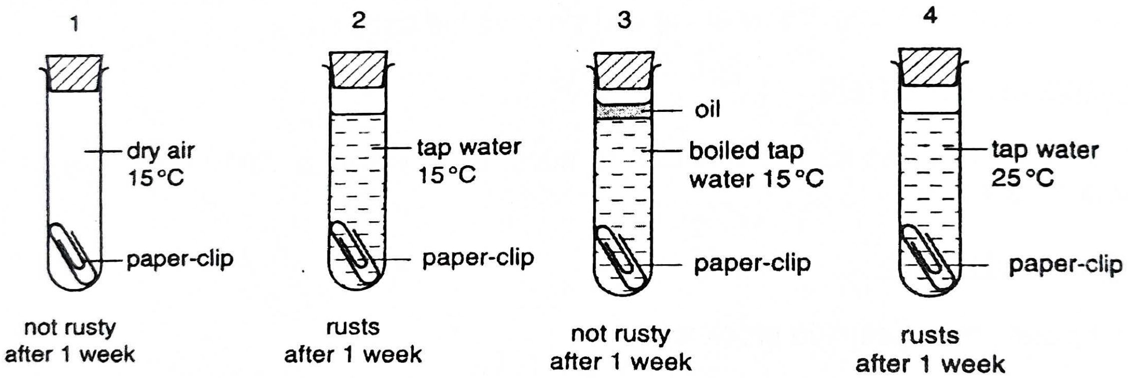
1 In the extraction of iron in the blast furnace, which substance reacts with haematite to produce molten iron?

A Calcium oxide
B Carbon monoxide
C Carbon dioxide
D Silicon dioxide

2 Molten iron and molten slag can be tapped off separately at the bottom of the blast furnace. Which property allows the two to be separated this way?

A One of them is ionic compound while the other is covalent.
B Both substances have different melting and boiling points.
C They are immiscible liquids.
D Molten slag has a lower density than molten iron.

3 Four experiments on rusting are shown.



Which experiment(s) can be used to show that water is needed for iron to rust?

A 1 only
B 1 and 2
C 2 and 4
D 3 only

Qn	Ans	Explanation
1	B	Reactions in blast furnace A: Reaction of calcium oxide: $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$ B: Reaction of carbon monoxide: $3CO(g) + Fe_2O_3(s) \rightarrow 2Fe(l) + 3CO_2(g)$ C: Reaction of carbon dioxide: $CO_2(g) + C(s) \rightarrow 2CO(g)$ D: Reaction of silicon dioxide: $SiO_2(s) + CaO(s) \rightarrow CaSiO_3(l)$
2	C	As molten iron and slag are immiscible liquids, they do not mix. Hence, the lower density slag floats on top of the molten iron, and they can be tapped off separately.
3	B	A: 1 only – shows that iron does not rust in the absence of water but no comparison to whether iron will rust in water. B: 1 and 2 – 1 shows that iron does not rust in the absence of water (control experiment), 2 shows that iron rust in the presence of water. C: 2 and 4 – both showed rusting but does not prove that water is a factor required for rusting. D: 3 only – shows that water is present but iron still did not rust.
4	A	A: Magnesium, being more reactive than iron, loses electrons more readily than iron. Hence magnesium will corrode in place of iron and protect iron from rusting. B and C: Paint and grease protects the iron by acting as a barrier to air and water. In this example, the paint and grease did not cover the entire iron. Thus, the iron will rust. *(However, even in the event that the paint and grease covers the iron, when the layer is scratched or damaged, rusting still occurs as oxygen and water can reach the surface of iron.)* D: Iron being more reactive than copper, loses electrons more readily than copper. Thus, iron will corrode very quickly.
5	B	P – high carbon steel (strong and brittle) Q – low carbon steel (hard, strong and malleable) Therefore, strength of P > strength of Q Brittleness of P > brittleness of Q

The Chemical Properties of Metals: the reactivity series

**Al is resistant to corrosion because of non-porous Al_2O_3 on its surface.

	K	Na	Ca	Mg	Al**	Zn	Fe	Pb	(H)	Cu	Ag/Au
Readiness to form positive ions	**Increasing readiness to lose electrons** and form positive ions ←										
Reducing agents (readiness to give away electrons)	**Increasing strength** of reducing agents ←										
With water	React with cold water to form **metal hydroxide** and **hydrogen**. $2Na + 2H_2O \rightarrow 2NaOH + H_2$			Do not react with cold water (except Mg). React with steam (except Al) with **increasing difficulty** to form **metal oxide** and **hydrogen**. → $Zn + H_2O \rightarrow ZnO + H_2$				No reaction	—	No reaction	
With dilute $HCl(aq)/H_2SO_4(aq)$	React explosively		Liberate hydrogen with **increasing strength** ← metal + acid → salt + H_2					Insoluble salt formed with dilute HCl/H_2SO_4	—	No reaction	
Reduction of oxides with H_2	Not reduced						Reduced with **increasing ease**: $CuO + H_2 \rightarrow Cu + H_2O$ →				
Reduction of oxides with C	Not reduced					Reduced with **increasing ease**: $2CuO + C \rightarrow 2Cu + CO_2$ →					
Displacement reaction	Metals higher in the reactivity series displace metals lower in the series from its salt solutions/oxide. $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$ $2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(l)$										
Properties of carbonates	1. soluble in water. 2. on heating, will not decompose.		Insoluble in water; on heating, generally decompose to **oxide & carbon dioxide** with **increasing ease** → $CaCO_3(s)$ (white) → $CaO(s)$ (white) + $CO_2(g)$ (very high temp); $ZnCO_3(s)$ (white) → $ZnO(s)$ (yellow(hot); white(cold)) + $CO_2(g)$; $CuCO_3(s)$ (green) → $CuO(s)$ (black) + $CO_2(g)$							** $2Ag_2CO_3(s) \rightarrow 4Ag(s) + 2CO_2(g) + O_2(g)$	

reducing agent

reduction

reduction

metal → cannot decompose, reaction compound unstable