FACTFILE: GCSE CHEMISTRY: UNIT 2.2

Rewarding Learning

Redox, Rusting and Iron

Learning Outcomes

Students should be able to:

2.2.1 recognise oxidation and reduction in terms of loss or gain of oxygen or hydrogen and identify in a reaction or symbol equation which species is oxidised and which is reduced (link to suitable industrial processes covered in this specification);

2.2.2 **recognise oxidation and reduction in terms of loss or gain of electrons, and identify in a symbol equation, ionic equation or half equation which species is oxidised or reduced (link to suitable industrial processes covered in this specification)**;

2.2.3 investigate experimentally rusting as a reaction of iron with water and air producing hydrated iron(III) oxide;

2.2.4 demonstrate knowledge and understanding of the methods used to prevent iron from rusting, including barrier methods such as painting, oiling, plastic coating and suitable metal coating/plating (galvanising) and explain sacrificial protection of iron related to the reactivity series;

2.2.5 describe the extraction of iron from haematite, including:

- the production of the reducing agent;
- the reduction of haematite; and
- the removal of acidic impurities; and

2.2.6 demonstrate knowledge and understanding that iron is used in bridges and structures due to its strength.

Redox reactions: Oxidation and Reduction

A redox reaction is a reaction in which both oxidation and reduction occur. Oxidation and reduction can be defined in terms of change in oxygen or hydrogen content or in terms of electrons.

	Oxidation	Reduction
In terms of oxygen	Gain of oxygen	Loss of oxygen
In terms of hydrogen	Loss of hydrogen	Gain of hydrogen
In terms of electrons	Loss of electrons	Gain of electrons

Examples involving the gain or loss of oxygen

These examples involve the change in oxygen content. Look for the elements or compounds which gain or lose oxygen. Make sure you give the definition of oxidation and/or reduction in terms of change in oxygen content.

$S + O_2 \rightarrow SO_2$

In this example:
Sulfur gains oxygen and the gain of oxygen is oxidation

$ZnO + C \rightarrow Zn + CO$

In this example:
A – zinc oxide loses oxygen and the loss of oxygen is reduction
B – carbon gains oxygen and the gain of oxygen is oxidation

$Mg + H_2O \rightarrow MgO + H_2$

In this example:
A – water loses oxygen and the loss of oxygen is reduction
B – magnesium gains oxygen and the gain of oxygen is oxidation

Examples involving the gain or loss of hydrogen

These examples involve the change in hydrogen content. Look for the elements or compounds which gain or lose hydrogen. Make sure you give the definition of oxidation and/or reduction in terms of change in hydrogen content

$N_2 + 3H_2 \rightarrow 2NH_3$

In this example:
Nitrogen gains hydrogen and the gain of hydrogen is reduction

$Cl_2 + H_2S \rightarrow 2HCl + S$

In this example:
A – hydrogen sulfide loses hydrogen and the loss of hydrogen is oxidation
B – chlorine gains hydrogen and the gain of hydrogen is reduction

Examples involving gain or loss of Electrons

Example 1: Oxidation and reduction can also be considered in terms of electron transfer from one atom or ion to another. For example, consider the following reaction:

Zinc reacts with copper(II) sulfate solution forming zinc sulfate and copper.

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

The ionic equation is:

$$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$$

The half equations are:

$$Zn \rightarrow Zn^{2+} + 2e^-$$

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

- Zinc loses electrons and loss of electrons is oxidation so zinc is oxidised
- Copper(II) ions gain electrons and gain of electrons is reduction so copper(II) ions are reduced

Example 2 – When chlorine water is added to a solution containing bromide ions:

$$2Br^- + Cl_2 \rightarrow Br_2 + 2Cl^- \quad \text{[ionic equation]}$$

The bromide ions *LOSE* electrons and therefore are *OXIDISED* to bromine atoms as loss of electrons is oxidation.
The chlorine atoms *GAIN* electrons and therefore are *REDUCED* to chloride ions as gain of electrons is reduction.

Examples of Redox reactions

Rusting

When Iron rusts it forms an oxide by gaining oxygen from air and the iron is oxidised. Water is also needed. Rust is hydrated iron(III) oxide

iron + oxygen + water ⟶ hydrated iron(III) oxide

Combustion of fuels

Combustion is the reaction with oxygen forming an oxide and releasing energy. This means that when a fuel, for example methane burns, an oxidation reaction is taking place.

methane + oxygen ⟶ carbon dioxide + water

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$$

Displacement reactions

A more reactive metal will displace a less reactive metal from its compound.

zinc + iron(II) chloride ⟶ zinc chloride + iron

$$Zn + FeCl_2 \longrightarrow ZnCl_2 + Fe$$

The zinc is changing from an atom to an ion and the iron is changing from an ion to an atom, there is a transfer of electrons. This ionic equation shows what is happening to the zinc and the iron:

$Zn + Fe^{2+} \longrightarrow Zn^{2+} + Fe$ [ionic equation]

The Zn atom becomes an ion and the Fe^{2+} ion becomes an atom.

If asked to explain why the reaction between zinc and iron(II) chloride is a redox reaction in terms of electrons the following would be expected:

- Zinc atoms lose electrons or this can be represented by the half equation $Zn \longrightarrow Zn^{2+} + 2e^-$.
- Loss of electrons is oxidation.
- Iron(II) ions gain electrons or this can be represented by the equation $Fe^{2+} + 2e^- \longrightarrow Fe$.
- Gain of electrons is reduction.
- A redox reaction is one in which both oxidation and reduction are occurring simultaneously.

Practice Questions

1. In each of the following reactions, say what is being oxidised and what is being reduced:

   (a) magnesium + water ⟶ magnesium oxide + hydrogen

   ____

   ____ [2]

   (b) copper(II) oxide + hydrogen ⟶ copper + water

   ____

   ____ [2]

   (c) $CH_4 + Cl_2 \longrightarrow C + 4HCl$

   ____

   ____ [2]

2. In the following questions decide which substance is being oxidised and which is reduced. Give reasons for your decision.

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

____

____ [2]

(b) $CuO + H_2 \rightarrow Cu + H_2O$

____

____ [2]

(c) $H_2S + Cl_2 \rightarrow 2HCl + S$

____

____ [2]

(d) $PbO + C \rightarrow Pb + CO$

____

____ [2]

(e) $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$

____

____ [2]

(f) $SO_2 + 2Mg \rightarrow 2MgO + S$

____

____ [2]

3. In each of the following reactions indicate whether the first substance in each equation has been oxidised or reduced, giving a reason.

(a) $H_2O + Mg \longrightarrow MgO + H_2$

____

____ [2]

(b) $2I^- + Cl_2 \longrightarrow I_2 + 2Cl^-$

____

____ [2]

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

____

____ [2]

(d) $2H_2S + SO_2 \longrightarrow 2H_2O + 3S$

____

____ [2]

(e) $4H_2O_2 + PbS \longrightarrow PbSO_4 + 4H_2O$

____

____ [2]

(f) $Sn^{2+} + Hg^{2+} \longrightarrow Sn^{4+} + Hg$

____

____ [2]

(g) $2Fe^{3+} + S^{2-} \longrightarrow 2Fe^{2+} + S$

____

____ [2]

Rusting

When a metal reacts with substances around it, such as water or air, it corrodes. Many transition metals corrode slowly, if at all, but iron and steel corrode quickly. Iron, usually as steel, has many uses because of its strength and relatively low cost. Unfortunately, iron and steel react in the atmosphere to produce reddish brown rust. The rust eventually flakes off the surface to reveal a fresh surface available for more rusting. Understanding how rust is formed and how to prevent or reduce the rusting process is of great economic importance. Rusting is an oxidation reaction where iron reacts with oxygen in the presence of water to form rust. The full chemical name of rust is hydrated iron(III) oxide, $Fe_2O_3.xH_2O$. The xH_2O is variable in rust depending on the amount of water present.

iron + oxygen + water $\longrightarrow$ rust

NOTE: Rusting is speeded up by the presence of acid (for example dissolved carbon dioxide or sulfur dioxide), or salt.

Problems with rusting

Iron or steel is the main metal used in large constructions (e.g. bridges, machinery, and scaffolding) and so the rusting of these things has economic and environmental implications.

- **Unsafe** – Rust is weaker than the metal iron.
- **Expensive** – Iron is relatively easy to extract from its ore and so it is a cheap metal. However rusty iron *cannot be recycled* so it needs to be *replaced* when it rusts, so it is expensive to repair.
- **Unsightly** – Tonnes of rusty iron get thrown onto landfill sites as it cannot be recycled. Therefore, it affects the environment.

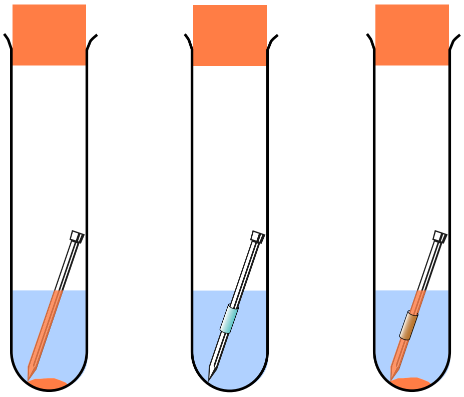
Preventing Rust

Most methods of rust prevention involve covering the iron with a material that acts as a **barrier**, stopping air (oxygen) and moisture from coming in contact with the iron.

	Method of preventing rust	Suitable for?
Barrier Methods	Painting	GATES, CARS, FENCING
	Oiling or greasing	MOVING ENGINE PARTS, HINGES, BICYCLE CHAINS
	Plastic Coating	WIRES AND CHAINS
	Chromium plating	DECORATIVE ITEMS, e.g. TAPS
	Galvanising *(covering with a layer of zinc)*	GATES, BUCKETS, NAILS MANY APPLICATIONS
	Sacrificial protection	HULL OF SHIPS

Sacrificial protection involves the use of another metal that reacts before the iron does.

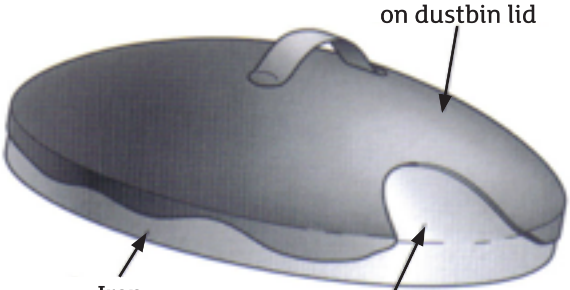
This other metal, therefore, must be higher in the reactivity series than iron is.



Ships' hulls and marine oil rigs are made of steel but they are used in an environment where rusting is going to be a serious problem. Attaching bars of zinc or magnesium to hulls or supporting rig legs greatly reduces the rust formation. In each case the more reactive metal reacts first. It is 'sacrificed' to protect the iron.

The zinc or magnesium blocks can be easily replaced when they have corroded away.

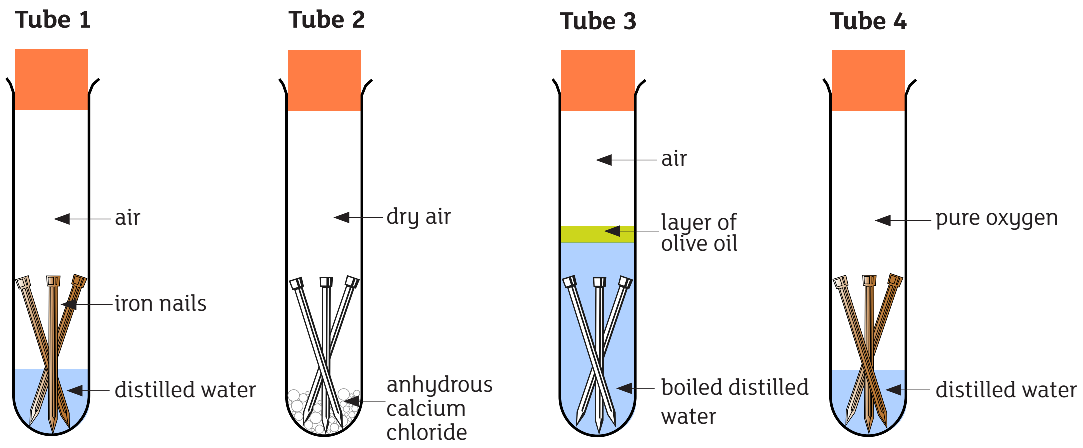
On a galvanised dustbin, zinc reacts before the iron, even where the zinc coating is scratched or broken.



Aluminium is higher in the reactivity series than iron but aluminium is not suitable for use in sacrificial protection – Aluminium does not actually corrode faster than iron because a thin layer of aluminium oxide forms on the surface, and so prevents further oxidation!

Questions

1. In the experiment below, the iron nails in tubes 1 & 4 rust.



(a) What is the reason for having calcium chloride in tube 2?

____

____ [1]

(b) Why do the iron nails in tube 2 not rust?

____

____ [2]

(c) In tube 3 what was removed from the water when it was boiled?

____ [1]

(d) Why do the nails in tube 3 not rust?

____

____ [1]

(e) What conditions does this experiment show are needed for iron to rust?

____

____ [2]

2. When iron reacts with oxygen in damp air it forms rust.

   (a) What is the full chemical name for rust?

   ____ [1]

   (b) Explain clearly **why** putting zinc blocks on the sides of ships helps to prevent rusting.

   ____

   ____ [2]

3. A student is very proud of his new motorbike. He is determined to protect it from rusting.

   (a) The manufacturers have galvanised the iron spokes. What is meant by the term **galvanised**?

   ____

   ____ [1]

   (b) How does galvanising protect the iron spokes from rusting?

   ____

   ____ [1]

4. Bridges are made of iron and is painted regularly to prevent it from rusting.

   (a) Give **two** reasons why the bridge must be prevented from rusting.

   ____

   ____ [1]

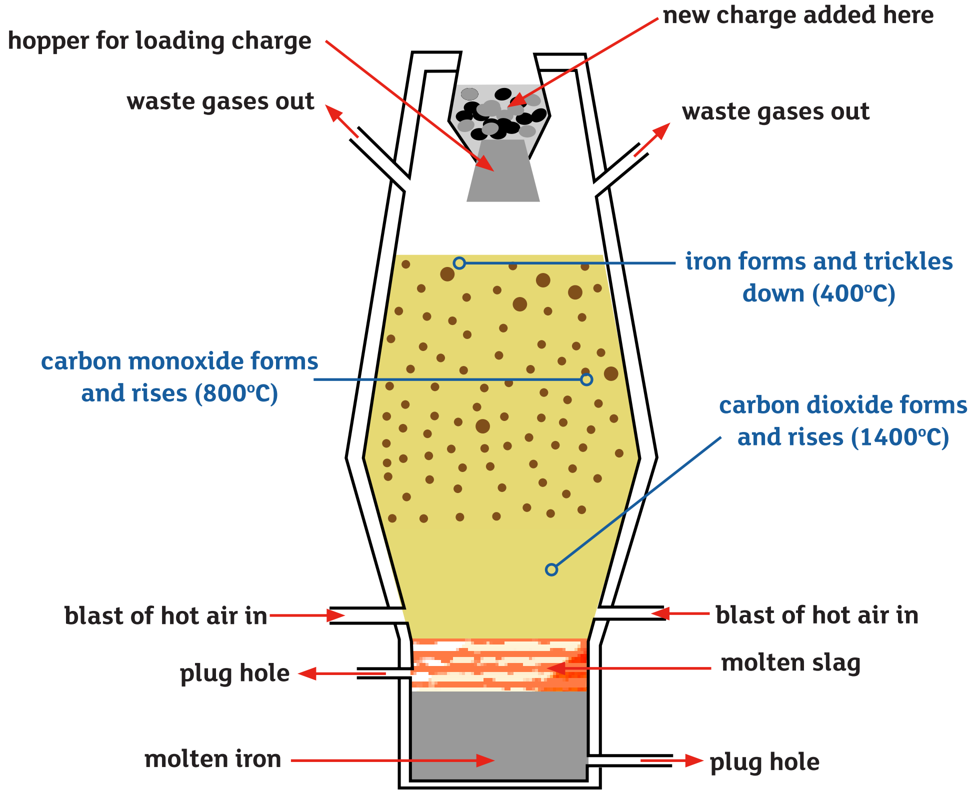
   (b) Rusting and combustion are examples of the same **type** of reaction. What is this type of reaction called?

   ____ [1]

The extraction of iron in the Blast Furnace

Iron is extracted from its ore, called haematite (iron(III) oxide) in a blast furnace. The raw materials are loaded in at the top of the furnace. This mixture is known as the charge. The charge consists of:

- coke (mainly carbon);
- limestone (calcium carbonate);
- iron ore (haematite).



Hot air is blasted into the bottom of the furnace.

There are 3 main processes:

1. The formation of carbon monoxide (the reducing agent)

The coke (mainly carbon) burns in the presence of hot air to form carbon dioxide, which further reactions with more coke to produce carbon monoxide gas, which is the reducing agent (NB the reducing agent gains oxygen). The equations for this step are given below:

carbon + oxygen → carbon dioxide

$$C + O_2 \rightarrow CO_2$$

carbon dioxide + carbon → carbon monoxide

$$CO_2 + C \rightarrow 2CO$$

2. The reduction of haematite (iron oxide) to iron

The carbon monoxide reacts with the iron(III) oxide because carbon is more reactive than iron. It takes the oxygen from the iron. As the temperature of the furnace is very high, the iron is in molten form and therefore sinks to the bottom where it can be tapped off.

carbon monoxide + iron(III) oxide ⟶ carbon dioxide + iron

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

3. The removal of acidic impurities

The limestone, added as part of the charge, is there to remove acidic impurities like silicon dioxide (sand). At the high temperature within the furnace, the limestone thermally decomposes to form calcium oxide and carbon dioxide:

calcium carbonate ⟶ calcium oxide + carbon dioxide

$$CaCO_3 \rightarrow CaO + CO_2$$

The calcium oxide formed above then goes on to react with silicon dioxide to form calcium silicate, otherwise referred to as “slag”, which is less dense than iron so it floats:

calcium oxide + silicon dioxide ⟶ calcium silicate

$$CaO + SiO_2 \rightarrow CaSiO_3$$

The waste material known as slag is less dense than the iron, so it floats on the molten iron. It is removed and allowed to cool before being sold on for uses such as road building.

Practice Questions

1. Iron is extracted from its ore in the blast furnace.

(a) Name the **three** substances which are present in the charge added to the blast furnace.

1. ____

2. ____

3. ____ [3]

(b) Write **two** balanced symbol equations to show how acidic impurities are removed from the blast furnace.

1. ____

2. ____ [4]

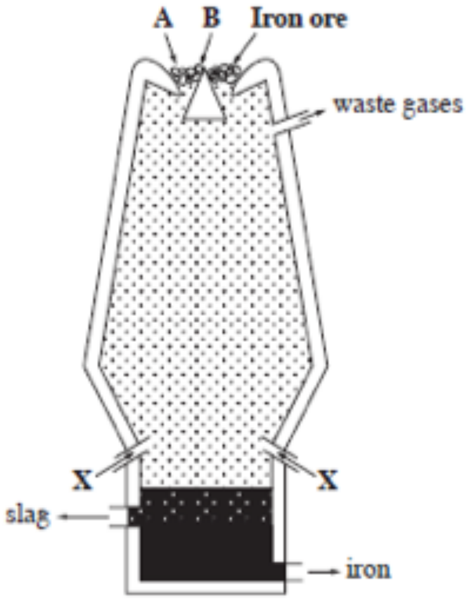
(c) Name the reducing agent in the blast furnace.

____ [1]

(d) How is the iron removed from the blast furnace?

____ [1]

2. Iron is obtained in the blast furnace from its ore, iron(III) oxide.



(a) Complete the table below about the raw materials, A and B, which are added at the top of the Blast Furnace. [2]

Substance	Common name	Chemical name
A		carbon
B		calcium carbonate

(b) Name the substance X that is blasted in lower down the Blast Furnace.

____ [1]

(c) Carbon monoxide is the reducing agent in the Blast Furnace. Write a balanced symbol equation to show how carbon dioxide produced in the Blast Furnace is converted into carbon monoxide.

____ [3]

(d) Use a balanced symbol equation to show how carbon monoxide reduces iron(III) oxide to iron in the Blast Furnace.

____ [3]