Application of redox reactions

Helpful links to read before completing this task.
The red cloth magic trick
Redcloth numbers (Teacher demonstration only)

Lesson 2 Worksheet 2 Solutions - oxidation numbers worksheets

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Applications of redox reactions in society

Redox reactions involve the simultaneous processes of oxidation and reduction, where electrons are transferred between chemical species. Oxidation is defined as the loss of electrons, while reduction is the gain of electrons. Redox reactions are fundamental chemical processes that underpin many natural phenomena and everyday technologies. They play a critical role in energy production, material protection, industrial processes and entertainment, making redox chemistry highly relevant to modern society.

During oxidation, the reductant loses electrons and its oxidation state increases, while during reduction, the oxidant gains electrons and its oxidation state decreases. These processes can be illustrated through several common examples.

One of the most familiar redox processes is the rusting of iron, which affects iron-based structures such as infrastructure, vehicles and tools. Rusting occurs when iron is oxidised in the presence of oxygen and water, leading to structural weakening and significant economic cost. The process can be described using the following half-equations:

Figure 1 – The economic impact of a reactive metal such as iron.

Oxidation (at the anode):
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
(Iron acts as the reductant and its oxidation state increases from 0 to +2)

Reduction (at the cathode):
$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
(Oxygen acts as the oxidant and its oxidation state decreases from 0 to −2)

Combining these reactions leads to the formation of iron(II) hydroxide:

$Fe(s) + O_2(g) + 2H_2O(l) \rightarrow Fe(OH)_2(s)$

Figure 2 - $Fe(OH)_3$

Iron(II) hydroxide is a greenish solid that is unstable in the presence of oxygen and water. It is further oxidised to form iron(III) hydroxide, which has the characteristic brown colour associated with rust according to this equation, $4Fe(OH)_2(s) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(s)$

An important application of redox reactions in society is the combustion of fuels, which provides energy for cooking, heating and transportation. The combustion of methane, the main component of natural gas, is a redox reaction represented by the equation below.

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Figure 3 – Combustion of natural gas (CH_4) is a redox reaction.

In this reaction, carbon is oxidised as its oxidation state increases from −4 in methane to +4 in carbon dioxide, while oxygen is reduced as its oxidation state decreases from 0 to −2. Methane acts as the **reductant**, and oxygen acts as the **oxidant**, with energy released during the reaction for practical use.

Another important application of redox reactions in society is in batteries, which power devices such as clocks, toys and computers. In batteries, chemical energy is converted into electrical energy through spontaneous redox reactions, providing a reliable source of electrical power for everyday use.

Figure 4 -Alkaline battery

In an electrochemical cell, the two electrodes have specific roles:

- The anode is where oxidation occurs and is the source of electrons.
- The cathode is where reduction occurs and is where electrons are gained.

Typical half-equations that occur in an alkaline battery are shown below:

Anode (oxidation)
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

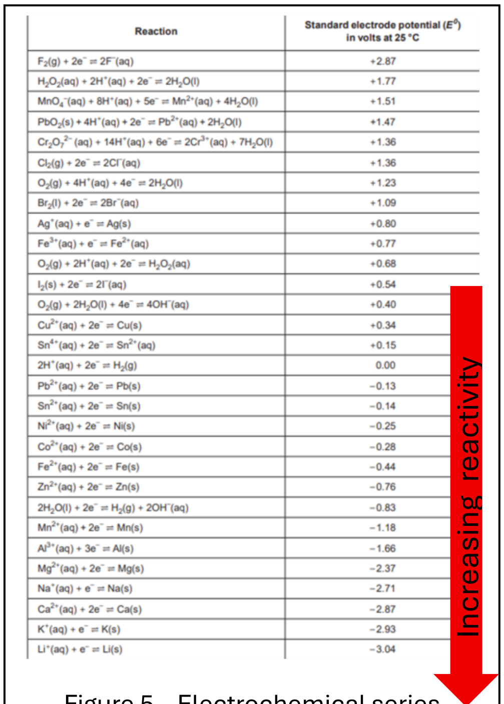
Cathode (reduction)
$2MnO_2(s) + 2H_2O(l) + 2e^- \rightarrow 2MnOOH(s) + 2OH^-(aq)$

Zinc(s) acts as the reductant and manganese dioxide(MnO_2) acts as the oxidant in this cell.
**** Take note – the reductant always reacts at the anode whilst the oxidant always acts at the cathode.***

It is the flow of electrons from the zinc anode to the manganese dioxide cathode through the external circuit that generates an electric current to power electrical devices.

The relationship between reactive metals and redox reactions in society

A reactive metal is a metal that readily loses electrons to form positive ions, in other words it can be easily oxidised. You can find the order of reactivity of some metals in item 1 of the VCAA Chemistry Data Booklet on page 2 as shown in fig 5. The tendency of a metal to lose electrons is indicated by its position in the electrochemical series, with metals such as potassium, magnesium, aluminium and zinc being more reactive than metals like iron and copper. Lithium is the most reactive metal positioned at the bottom of the electrochemical series, as shown in fig. 5.



Reaction	Standard electrode potential (E^0) in volts at 25 °C
$F_2(g) + 2e^- \rightleftharpoons 2F^-(aq)$	+2.87
$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l)$	+1.77
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
$PbO_2(s) + 4H^+(aq) + 2e^- \rightleftharpoons Pb^{2+}(aq) + 2H_2O(l)$	+1.47
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.36
$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
$O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O_2(aq)$	+0.68
$I_2(s) + 2e^- \rightleftharpoons 2I^-(aq)$	+0.54
$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq)$	+0.40
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Sn^{4+}(aq) + 2e^- \rightleftharpoons Sn^{2+}(aq)$	+0.15
$2H^+(aq) + 2e^- \rightleftharpoons H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$	−0.13
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14
$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s)$	−0.25
$Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$	−0.28
$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$	−0.44
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
$2H_2O(l) + 2e^- \rightleftharpoons H_2(g) + 2OH^-(aq)$	−0.83
$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s)$	−1.18
$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−1.66
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	−2.37
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	−2.71
$Ca^{2+}(aq) + 2e^- \rightleftharpoons Ca(s)$	−2.87
$K^+(aq) + e^- \rightleftharpoons K(s)$	−2.93
$Li^+(aq) + e^- \rightleftharpoons Li(s)$	−3.04

Figure 5 – Electrochemical series

Important note – Electrochemical series vs observed reactivity

The electrochemical series ranks metals according to their tendency to lose electrons under standard conditions, not how quickly they react in practice. Although lithium is the strongest reductant in the electrochemical series, sodium and potassium react more vigorously with water in [experimental observations](#).

This difference occurs because reaction behaviour is also influenced by factors such as activation energy, atomic size, ionisation energy trends down Group 1, and the presence of a surface oxide layer on lithium metal.

Students should note: the electrochemical series predicts whether a reaction is possible, not the rate at which it will occur.

Because reactive metals lose electrons easily, they are central to many useful redox reactions in everyday life, particularly in corrosion prevention and energy storage batteries.

One important societal application of reactive metals is in **rust prevention using sacrificial anodes**. Rusting is a redox process in which iron is oxidised in the presence of oxygen and water, weakening metal structures. To prevent this, a more reactive metal such as **zinc** is attached to iron. Zinc is more reactive than iron and therefore oxidises preferentially, thus protecting the iron from oxidation.

The oxidation half-equation for zinc is:

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

The electrons released flow to the iron, making it the cathode and preventing the oxidation of iron:

$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$ *(thus reducing the Fe^{2+} back to Fe(s) and hence preventing rust formation and long term damage)*

This process is widely used to protect ship hulls, pipelines and galvanised steel from corrosion.

Figure 6 – Reactive Zinc blocks placed in the hull of a ship.

Reactive metals are also essential in **batteries**, where chemical energy is converted into electrical energy through spontaneous redox reactions. As discussed earlier, in a typical **alkaline battery (fig 4)**, zinc acts as the reactive metal, readily losing electrons thus providing a flow of electrons through the external circuit.

Overall, reactive metals are fundamental to redox chemistry in society because their strong tendency to lose electrons allows redox reactions to occur in a controlled and useful way. By exploiting the reactivity of metals, society can generate electrical energy, prevent corrosion (Fig 7) and extend the lifespan of infrastructure. These applications highlight the importance of understanding reactive metals and redox reactions in both chemical and environmental contexts.

Figure 7 – galvanized corrugated steel

Water purification with metal displacement redox reactions and precipitation reactions.

Metal displacement reactions are redox reactions in which a more reactive metal transfers electrons to the ions of a less reactive metal, causing the less reactive metal to be deposited as a solid while the more reactive metal enters solution as ions.

These reactions are governed by the electrochemical (reactivity) series (fig 5), which ranks metals according to their tendency to lose electrons. In every displacement reaction, the more reactive metal is oxidised (loses electrons and acts as the reductant), while the less reactive metal ion is reduced (gains electrons and acts as the oxidant).

Figure 8 – copper in a solution of silver nitrate.

For example, as shown in fig 8, when copper metal (more reactive) is placed in a silver nitrate solution, copper atoms lose electrons to form Cu^{2+} ions while Ag^+ ions gain these electrons to form silver metal:

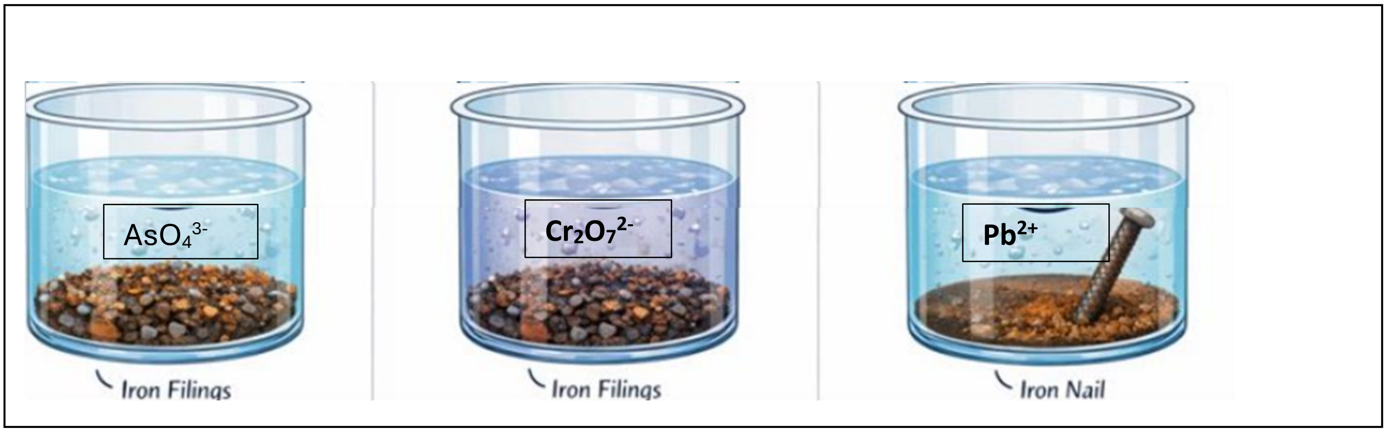
$$\mathrm{Cu}(s) + 2\mathrm{Ag}^{+}(aq) \rightarrow \mathrm{Cu}^{2+}(aq) + 2\mathrm{Ag}(s)$$

Cu is the reductant and undergoes oxidation – $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Ag is the oxidant and undergoes reduction – $Ag^+(aq) + e^- \rightarrow Ag(s)$

Looking at the electrochemical series, fig 5, you see that zinc is a more reactive metal than copper. Dissolved copper ions (Cu^{2+}) will therefore be displaced out of solution by zinc metal.

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

Redox reactions are used in water purification (fig 9) to remove toxic heavy metal ions by converting them into solid metals that can be physically separated from the water.



Reduction half-equation
$O_2(g)+2H_2O(l)+4e^- \rightarrow 4OH^-(aq)$

Oxidation half-equation
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Further oxidation takes place to oxidise Fe^{2+} to Fe^{3+}
$4Fe^{2+}(aq)+O_2(g)+4H^+(aq) \rightarrow 4Fe^{3+}(aq)+2H_2O(l)$

Final precipitation reaction to remove AsO_4^{3-}
$Fe^{3+}(aq)+AsO_4^{3-}(aq) \rightarrow FeAsO_4(s)$

Reduction half-equation
$Cr_2O_7^{2-}(aq) + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Oxidation half-equation
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Precipitation to remove Cr^{3+}

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

Reduction half-equation
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

Oxidation half-equation
$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

Metal displacement reaction to remove Pb^{2+}

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

Figure 9 – Removal of three toxic metals using redox reactions.

Notice in fig 9, above, how iron metal is used to purify water. Iron in the form of ions (Fe^{3+}) is far less toxic and easier to remove from water than heavy metals such as lead, arsenic or chromium. As a result, iron is particularly effective in water purification because it enables the simultaneous removal of multiple contaminants. There are three main reasons why iron acts as a natural and efficient purifying agent:

- **Low toxicity and natural abundance**. Iron, in the form of Fe^{3+} ions, is essential for living organisms and are naturally present in soils, rocks, and groundwater. While very high concentrations of Fe^{3+} can cause staining or alter the taste of water, they are far less toxic than metals such as lead, arsenic, or chromium, which are hazardous even at very low concentrations.
- **Formation of insoluble compounds**. Fe^{3+} ions readily form insoluble solids such as iron(III) hydroxide, $Fe(OH)_3$, a brown substance that does not remain dissolved in water:

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

  These solids can be easily removed through sedimentation or filtration using sand filters.

- **Removal of other toxic metals**. Freshly formed $Fe(OH)_3$ has a large surface area and acts as an effective collector, capturing and binding toxic species such as arsenate, chromate, and phosphate. This allows iron to simultaneously remove multiple contaminants, enhancing its effectiveness in water purification

The following links are essential reading and at times , revision of prior knowledge. You will go over these with your teacher. Complete all the relevant exercises linked to each online page.

Reactivity of metals activity

Reactive metals

Comparing the reactivity of metals of group 1 and 2 metals

Comparing the reactivity of metals using a lemon battery.

Rusting experiment

1. Consider the two unbalanced combustion reactions shown below.
   i. $H_2(g) + O_2(g) \rightarrow H_2O(l)$
   ii. $C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

   For each reaction identify the oxidant and reductant. Justify your answer with reference to oxidation states.

   Reaction i.

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   ____ *2 marks*

   Reaction ii.

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   ____ *2 marks*

2. A student was asked to use a lemon to construct a battery. Copper and zinc metals were used by the student.
   Figure 10 shows the outcome of the student's efforts. Using the electrochemical series shown in fig. 5, above, answer the following questions.

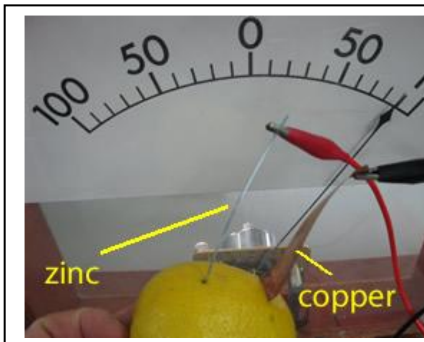

Figure 10 – A lemon battery

   a. Identify the oxidant and justify your answer

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   ____ *2 marks*

   b. Identify the reductant and justify your answer

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   ____ *2 marks*

   c. Which metal represents the:

   Anode ____ *1 mark*

   Cathode ____ *1 mark*

3. Iron is used in construction because it is strong, relatively cheap and easily accessible as iron ore. Iron, however, is a very reactive metal and tends to rust, adding billions of dollars to the cost of maintaining ships, buildings, bridges and infrastructure generally. The student setup four test tubes.
Test tube A – iron nail wrapped with copper wire submerged in fresh water.
Test tube B – iron nail wrapped with zinc wire submerged in fresh water.
Test tube C – iron nail submerged in fresh water.
Test tube D – iron nail submerged in salty water.
Below is an image of the results fig 11.

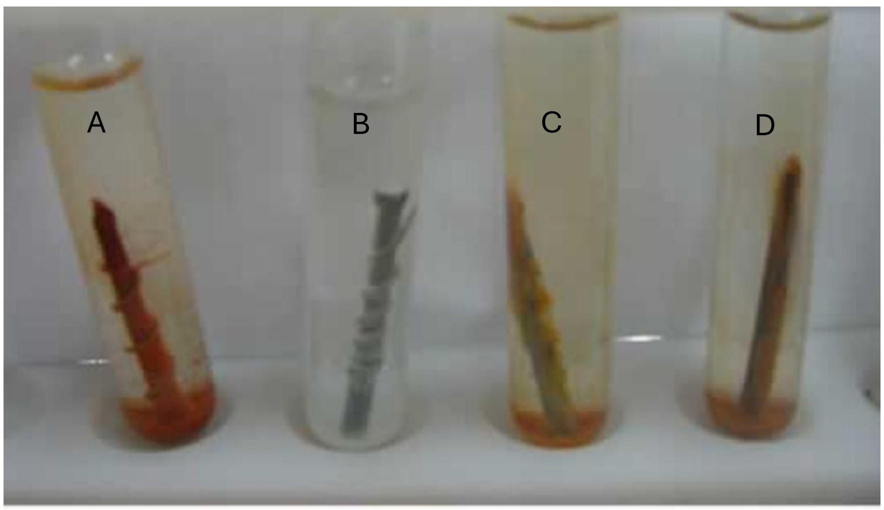

Figure 11 – Results from a student's rusting experiment.

a. Using the electrochemical series shown in figure 5 , above, give the species gaining electrons (oxidant) and the species losing electrons (reductant) in each of the four test tubes.

Test Tube A	Oxidant ____	*1 mark*
	Reductant ______________________________________	*1 mark*
Test Tube B	Oxidant ____	*1 mark*
	Reductant ______________________________________	*1 mark*
Test Tube C	Oxidant ____	*1 mark*
	Reductant ______________________________________	*1 mark*
Test Tube D	Oxidant ____	*1 mark*
	Reductant ______________________________________	*1 mark*

b. Using your knowledge of redox reactions, oxidation numbers and the electrochemical series pictured in fig 5, explain the difference in rusting taking place between test tubes A and B.

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____*4 marks*

c. Using your knowledge of redox chemistry and the conductivity of salt solutions, explain the difference in rusting between test tubes C and D.

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____ *3 marks*

d. A student suggests that covering the nail in test tube D completely with a thin layer of gold metal would make the nail rust proof even if scratched.
Do you agree with the student? Justify your answer with reference to electrochemical series shown in fig 5.

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____ *3 marks*

e. Identify five controlled variables in the student's practical investigation.

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____ *1 mark*

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f. Consider the results shown in figure 10. They represent a qualitative investigation.

i. Explain the difference between the terms "Qualitative" and "Quantitative" and suggest a possible format to visually represent the results as a graph.

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____ *2 marks*

ii. The experimental report calls for the results to represented as a graph of some sort. Suggest a suitable format (line graph, bar graph, pie chart etc) to visually represent the results and justify your reasoning.

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____ *2 marks*

iii. Write a procedure that can be followed by another student to change this investigation from a qualitative to a quantitative investigation.

4 marks

4. A simplistic, graphical representation of a battery is shown below in fig 12. Two sections are shown, section A and section B. A wire connects both ends of the battery to allow for electron flow.

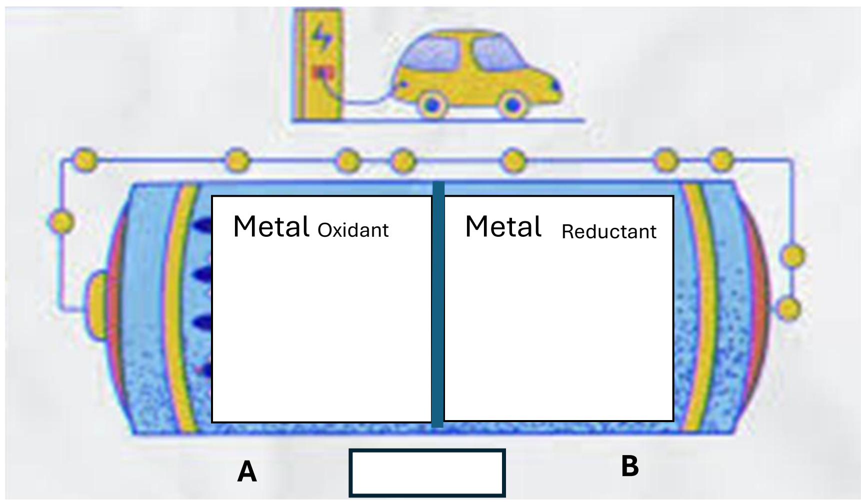

Figure 12 -graphical representation of a battery

a. The battery contains two metals lithium (Li) and Copper (Cu). Each metal is housed in a separate section of the battery and are kept apart.

i. Write the metals Li and Cu in their respective sections (A or B) in the box provided and indicate with an arrow in the box provided the direction of electron flow. *2 marks*

ii. Identify terminals X and Y as either the anode or cathode.

X ____________________ Y ___________________ 1 mark

iii. Justify your answer to question ii above using half equations.

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____*2 marks*

b. Compared to lithium, sodium is easier and cleaner to extract. Both lithium and sodium give away one electron each and yet lithium is the metal of choice when making small batteries for electrical devices. Using your knowledge of the periodic table and the electrochemistry you have encountered, give a reason for this.

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____*3 marks*

c. Explain why lithium is a safer group 1 element to use in batteries than sodium or potassium. In your explanation include a balanced chemcial equation that highlites the danger of these elements.

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____*2 marks*

Traditionally, iron has been produced in a **blast furnace** by reacting iron ore, primarily iron(III) oxide (Fe_2O_3), with carbon. This process produces molten iron but also generates large quantities of carbon dioxide, a greenhouse gas, as shown in the reaction below.

$2Fe_2O_3(l) + 3C(s) \rightarrow 4Fe(l) + 3CO_2(g)$

As a result of increasing concern over carbon emissions, new innovations aim to produce **green steel**, which significantly reduces the amount of carbon dioxide released during steelmaking. One emerging method involves using **ammonia (NH_3)** as a reducing agent to convert iron(III) oxide into iron metal.

$2NH_3(g) + Fe_2O_3(s) \rightarrow 2Fe(l) + 3H_2O(g) + N_2(g)$

In this process, ammonia is oxidised while iron(III) oxide is reduced to iron, eliminating the direct production of carbon dioxide.

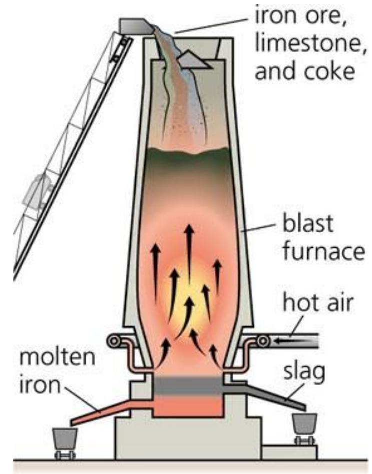

Figure 13 – Blast furnace

Ammonia used in this method can be produced using **hydrogen generated by the electrolysis of water**, a process that can be powered by renewable electricity.

$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

The hydrogen produced is then reacted with nitrogen in the **Haber process** to form ammonia.

$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$

The Haber process operates at high temperatures of approximately **450 °C** and high pressures, making it an energy-intensive process. If the required energy is supplied by fossil fuels, the overall sustainability of ammonia production is reduced. However, when renewable energy sources are used to generate hydrogen, the process becomes significantly more sustainable.

Ammonia is easier to store and transport than hydrogen, making it a practical energy carrier that can be integrated into existing infrastructure. This makes ammonia-based steel production a promising and economically viable pathway toward reducing carbon emissions in the steel industry.

1. In the traditional blast furnace reaction
   $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$
   identify the oxidant and the reductant and justify your answer using oxidation states.

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   ____ *4 marks*

2. Write the oxidation half-equation for carbon in the blast furnace process.

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   ____ *2 marks*

3. In the ammonia based reduction of iron(III) oxide,
   $2NH_3 + Fe_2O_3 \rightarrow 2Fe + 3H_2O + N_2$
   Identify the oxidant and reductant in this reaction and give the balanced chemical equations for the oxidation and reduction half-reactions taking place in this process.

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   ____ *4 marks*

4. Consider the production of iron via the use of ammonia

$2NH_3 + Fe_2O_3 \rightarrow 2Fe + 3H_2O + N_2$

Using the items 26 and 27 of the VCAA Chemistry Data Booklet identify:

a. one United Nations Sustainability Development Goal addressed by this method and justify your choice.

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____ *2 marks*

b. one Green Chemistry Principle and explain the conditions under which this principle can be met by this process.

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____ *2 marks*

5. Ammonia is described as easier to store and transport than hydrogen.
a. Explain why in terms of **intermolecular bonding and physical state**. In the space provided, on the next page, draw a diagram to assist with your explanation.

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____ *6 marks*

Diagram for question 5a.

b. Currently hydrogen is produced on an industrial scale via a process known as steam reformation. The reaction is shown below.

$$CH_4(g) + H_2O(g) \rightarrow 3H_2(g) + CO(g)$$

i. Describe why the production of hydrogen in this manner is considered nonrenewable does not follow any of the Green Chemistry Principles listed in item 28 part ii, of the DataBooklet.

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____ *2 marks*

ii. Describe a possible way that green hydrogen could be manufactured on an industrial scale. Include relevant half equations in your explanation.

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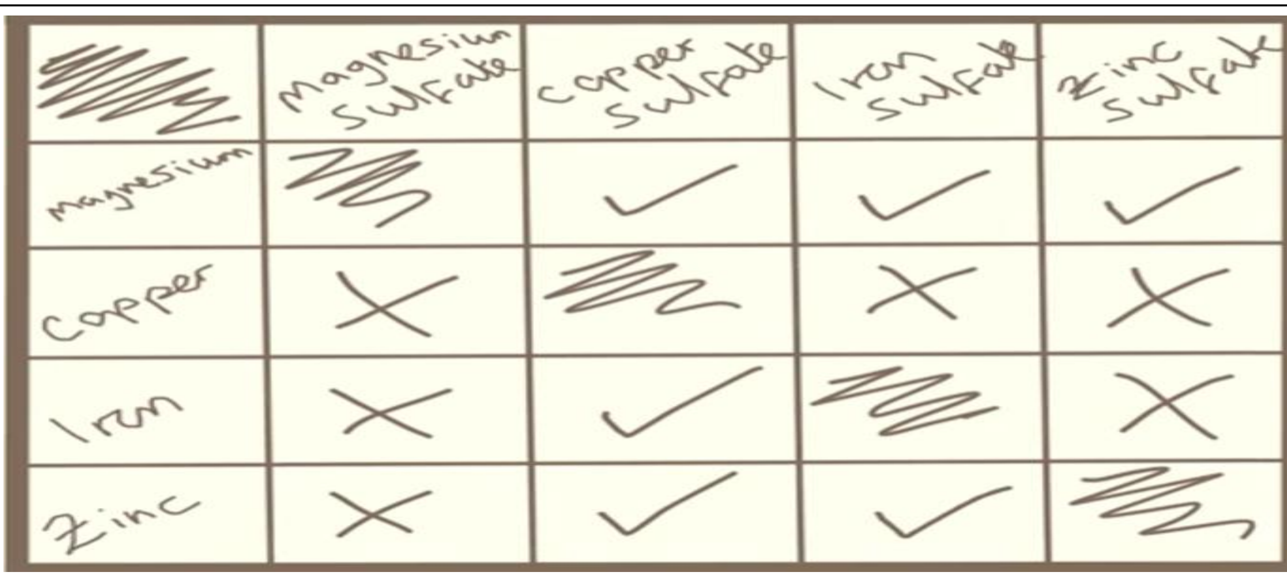
____ *2 marks*

Metal displacement investigation

Figure 14 – A student's experimental results of a metal displacement experiment.



	magnesium sulfate	copper sulfate	iron sulfate	zinc sulfate
magnesium		✓	✓	✓
copper	✗		✗	✗
iron	✗	✓		✗
zinc	✗	✓	✓	

Figure 15 – The student's table of results

6. *When a more reactive metal is placed in a solution containing ions of a less reactive metal, it will donate electrons and displace the less reactive metal from the solution.*

a. Using the information above and the table of the student's results (fig 15) derive the order of reactivity of the four metals tested from lowest to highest. *1 mark*

Low reactivity ,__________, ___________, ___________, ___________, High reactivity

b. If a bridge was to be built from iron what metal/s can be used to protect the bridge from rusting? Justify your answer with reference to the results.

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____ *2 marks*

c. One of the reactions is circled in red in fig. 14. A reaction should not have occurred, between Cu(s) and $CuSO_4$(aq). The student labelled the results table as having no reaction, but the results show a distinct discolouration of the copper metal indicating a reaction has taken place. Three possible scenarios to explain this random error were given.

i. The tile used was not properly washed after using $AgNO_3$ solution.

ii The tile used was not properly washed after using $Pb(NO_3)_2$ solution.

iii. The tile used was not properly washed after using $Ni(NO_3)_2$ solution.

Identify the likely error and justify your choice using the E^o series in fig 5 and a balanced chemical equation.

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____ *4 marks*

Solutions