Master

Kirkcaldy High School

N4/5 Chemistry

Unit 3

Chemistry in Society – part 1

Name: ______________________

Class: ______________________

Teacher: ______________________

Assessment Page

Homework

Homework title	Date	Mark/Total Mark
		/
		/
		/

Check tests

Test title	Date	Mark/Total Mark
		/

Teacher comments

Date: ________________

Metals

Learning Intentions

- To learn

Success Criteria

☐ I can describe the bonding and structure in a metal element/compound.

☐ I can explain the properties of metallic bonding and how they arise.

☐ I can determine the order of reactivity of metals.

Metallic Bonding

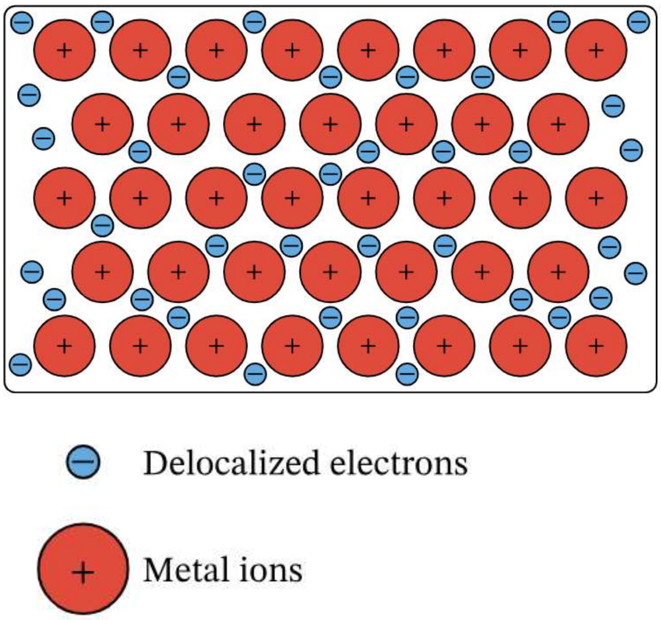
Metallic bonding is the type of bonding that occurs between **metal** atoms.

In this structure, **positive metal ions** are closely packed together in a **lattice**, but their **outer** electrons are **free to move** throughout the structure.

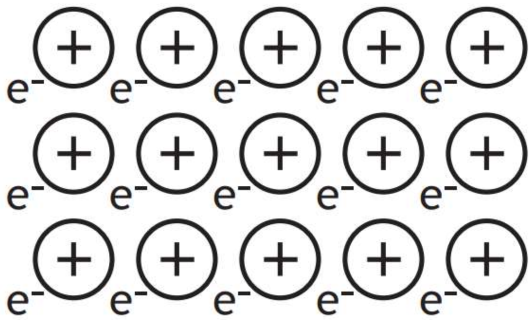
These free to move electrons are known as **delocalised electrons**.

These **delocalised electrons**, often referred to as a "sea of electrons," are **shared** among **all the atoms.** This allows metals to **conduct** electricity, be **malleable**, and have other characteristic properties.

Metallic lattice



Metallic lattice (SQA version)



Summarise the properties of metallic compounds below:

State at Room Temperature	Usually solid (except for mercury)
Melting and Boiling Points	Generally high
Electrical Conductivity	Good conductors of electricity
Thermal Conductivity	Good conductors of heat
Malleability	Can be hammered into thin sheets
Ductility	Can be drawn into wires
Luster	Shiny and reflective
Hardness	Varies; some are hard, some are soft
Type of Elements Involved	Metals
Type of Structure	Lattice with free-moving electron (delocalised)

Questions

1. **Describe** the structure of a metallic lattice.

2. **Explain** why metallic substances conduct electricity.

Reactions of metals

Metals exhibit varying degrees of reactivity with substances like oxygen, water, and acids. These reactions can be represented through chemical equations:

Metal + Oxygen → **Metal Oxide**

Metal + Water → **Metal Hydroxide** + **Hydrogen**

Metal + Dilute Acid → **Salt** + **Hydrogen**

Questions
Complete the following equations:

1. Magnesium + Oxygen → ___________________
2. Sodium + water → ___________________ + ___________________
3. Zinc + Hydrochloric acid → ___________________ + ___________________
4. Iron + Oxygen → ___________________
5. Aluminium + Sulfuric Acid → ___________________ + ___________________
6. Lithium + Water → ______________________ + ___________________
7. Potassium + ______________ → Potassium Hydroxide + Hydrogen
8. Aluminium + ___________________ → Aluminium Chloride + ______________

***Extension**: Write the balanced chemical equation for each equation.*

By examining how different metals react in these scenarios, we can determine their reactivity levels. Metals react at different **rates** with these substances, offering a practical way to arrange them in an order of reactivity.

Our upcoming experiments aim to **systematically** observe and record the reactions of various metals, including Copper, Zinc, Aluminium, Iron, and Magnesium, with **oxygen**, **water**, and **dilute** hydrochloric **acid**. This will help us understand and establish the **reactivity series** of these metals.

Reaction Title: Reaction of metals with oxygen – Arculus tubes (p33 lab book)

Aim: ______________________________`____

____

Method:

Results:

Conclusion: ____

____

____

Evaluation: ____

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____

Reaction Title: Reaction of metals with acid and water (p33 lab booklet)

Aim: ____

____

Method:

Results:

Conclusion: ____

____

____

Evaluation: ____

____

____

Reaction Title: Alkali metals

Aim: ____

____

Method:

Results:

Conclusion: ____

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Evaluation: ____

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____

Order of reactivity

Summarise the reactivity series from the experiments you have undertaken below:

Your teacher will now relate these findings to page 10 of your data booklet. The diagram below has been taken from this page and shows only the metals:

Metal
lithium
potassium
calcium
sodium
magnesium
aluminium
zinc
iron
nickel
tin
lead
copper
silver
mercury
gold

Date: _______________

Extraction of metals

Learning Intentions

- To learn how to extract metals based on their reactivity.

Success Criteria

☐ I can state the definition of an ore.

☐ I can state why metals such as gold and silver are found uncombined in the earth's crust.

☐ I can determine the method of metal extraction based on its reactivity.

Extraction and Reactivity

Metals are found naturally in the earths crust but are mostly combined with other elements, forming compounds.

> Naturally occurring metal compounds are known as **ores**.

Metals found in ores aren't very useful to us as they are. We need to **extract** the metal from the ore first, and how we do that depends where the metal is in the **reactivity** series.

Heat Alone:

- **Metals:** Silver (Ag), Gold (Au), and Mercury (Hg).
- **Process:** These metals, being **less** reactive, are often found in their **pure state**. To extract them, we usually just need to **heat** them, because they don't **mix** easily with other elements.

Heating with Carbon or Carbon Monoxide:

- **Metals:** Copper (Cu), Lead (Pb), Tin (Sn), Iron (Fe), and Zinc (Zn).
- **Process:** These metals, which are **more** reactive than the first group but **less** reactive than the most reactive metals, usually combine with oxygen to form oxides. To extract the metal, we heat these oxides with **carbon** or **carbon monoxide**.

Electrolysis:

- **Metals:** Highly reactive metals, including Aluminium (Al).

- **Process:** For the **most** reactive metals, we use a method called electrolysis.

  These metals are tightly bonded to other elements in their ores, so they can't be extracted just by heating with carbon.

> **Electrolysis** is the **decomposition** of an **ionic** compound into its **elements** using **electricity**.
>
> A direct current (**DC**) electrical **supply** must be used so the **products** can be **identified.**

Summarise this information below:

Metal
lithium
potassium
calcium
sodium
magnesium
aluminium
zinc
iron
nickel
tin
lead
copper
silver
mercury
gold

<u>Remember this will be found on page 10 of your data booklet.</u>

Reaction Title: Reduction of silver oxide and copper oxide (p34/35 lab book)

Aim: ____

____

Method:

Results:

Conclusion: ____

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Reaction Title: Electrolysis of copper chloride solution (p18 lab book)

Aim: ____

____

Method:

Results:

Conclusion: ____

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Questions

1. What are naturally occurring metal compounds in the Earth's crust called?

2. State the main process used to extract highly reactive metals like Aluminium (Al)?

3. Describe why metals like Silver (Ag), Gold (Au), and Mercury (Hg) are often found uncombined in nature.

4. List two metals that are typically extracted by heating with carbon or carbon monoxide.

5. Describe what happens during the electrolysis process when extracting a metal from its ore.

6. Determine the method of extraction for the following metals:

Metal	Method of Extraction (Heat Alone, Heating with Carbon/Carbon Monoxide, Electrolysis)
Silver (Ag)	
Iron (Fe)	
Gold (Au)	
Zinc (Zn)	
Mercury (Hg)	
Aluminium (Al)	
Lead (Pb)	
Copper (Cu)	

Date: ________________

Oxidation and Reduction reactions

Learning Intentions

- To learn about oxidation and reduction for electrochemistry.

Success Criteria

☐ I can state the definition of an oxidation and reduction reaction.

☐ I can identify an oxidation and reduction reaction.

☐ I can write a balanced redox equation.

Oxidation and Reduction

We will now look at reactions involving loss or gain of **electrons**.

Oxidation is a reaction that involves a **loss** of electrons.

Ion electron equation: $Mg(s) \rightarrow Mg^{2+}(aq) + \mathbf{2e^-}$

The electrons are **products**.

Reduction is a reaction that involves the **gain** of electrons.

Ion electron equation: $F_2 + \mathbf{2e^-} \rightarrow 2F^-(aq)$

The electrons are **reactants**.

This can be remembered with: **O.I.L. R.I.G.** = **o**xidation **i**s **l**oss **r**eduction **i**s **g**ain

These reactions are shown on **page 10** of the data booklet. They are written as **reduction** reactions as shown but can be written **backwards** to show **oxidation** reactions.

Electrochemical Series (Reduction Reactions)

Metal	Reaction		
lithium	$Li^+(aq) + e^-$	$\rightarrow$	$Li(s)$
potassium	$K^+(aq) + e^-$	$\rightarrow$	$K(s)$

Questions

With the following reactions, determine whether it is an oxidation or reduction reaction.

No.	***Reaction***	***Oxidation or reduction?***
1	$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$	
2	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	
3	$Na(s) \rightarrow Na^+(aq) + e^-$	
4	$O_2(g) + 4e^- \rightarrow 2O^{2-}(aq)$	
5	$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$	
6	$H_2(g) + 2e^- \rightarrow 2H^-(aq)$	
7	$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	
8	$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	
9	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	
10	$S^{2-}(aq) \rightarrow S(s) + 2e^-$	

Now add the electrons (e^-) into the correct position on the following equations. You make use page 10 to help you.

No.	***Reaction***	***Ox. or red.?***
1	$Ca(s) \quad\quad \rightarrow Ca^{2+}(aq)$	
2	$Cl_2(g) \quad\quad \rightarrow 2Cl^-(aq)$	
3	$Li(s) \quad\quad \rightarrow Li^+(aq)$	
4	$F_2(g) \quad\quad \rightarrow 2F^-(aq)$	
5	$Sn^{2+} \quad\quad \rightarrow Sn(s)$	
6	$Cu^{2+}(aq) \quad\quad \rightarrow Cu(s)$	
7	$Mg(s) \quad\quad \rightarrow Mg^{2+}(aq)$	
8	$2Br^-(aq) \quad\quad \rightarrow Br_2(l)$	

Redox Reactions

In a **redox** reaction, **red**uction and **ox**idation take place at the same time.

Redox equations can be written from combining an oxidation and reduction ion electron equation as follows. In order to do this we must be able to cancel out the electrons.

1. To ensure that the number of electrons is the same on both sides in a redox reaction, we can cross-multiply the number of electrons involved in each half-reaction. Here's how it's done:

   2x $Ag(s) \rightarrow Ag^{+}(aq) + e^{-}$ oxidation

   = $\mathbf{2}Ag(s) \rightarrow \mathbf{2}Ag^{+}(aq) + \mathbf{2}e^{-}$

   1x $F_2(g) + 2e^{-} \rightarrow 2F^{-}(aq)$ reduction

2. Combine and cancel the electrons:

   $2Ag(s) + F_2(g) + \cancel{2e^{-}} \rightarrow 2Ag^{+}(aq) + 2F^{-}(aq) + \cancel{2e^{-}}$

3. Final redox equation (after cancelling electrons)

   $2Ag(s) + F_2(g) \rightarrow 2Ag^{+}(aq) + 2F^{-}(aq)$ redox

Attempt to write the redox equation for the following equations.

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$$

$$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$$

Questions

Write the redox reactions given the half reactions.

1. **Half-Reactions:**

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

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

2. **Half-Reactions:**

$2I^-(aq) \rightarrow I_2(s) + 2e^-$

$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

3. **Half-Reactions:**

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

$H_2O_2(aq) + 2e^- \rightarrow 2H_2O(l)$

4. **Half-Reactions:**

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

$Ag+(aq) + e^- \rightarrow Ag(s)$

5. **Half-Reactions:**

$Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$

$Pb^{2+}(aq) + 2e^{-} \rightarrow 3Pb(s)$

6. **Half-Reactions:**

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

$2H^{+}(aq) + 2e^{-} \rightarrow H_2(g)$

7. **Half-Reactions:**

$2Br^{-}(aq) \rightarrow Br_2(l) + 2e^{-}$

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

8. **Half-Reactions:**

$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^{-}$

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

Date: ________________

Electrochemical Cells

Learning Intentions

- To learn about electrochemical cells.

Success Criteria

- [ ] I can state the definition of an electrolyte
- [ ] I can apply my knowledge of redox reaction to electrochemical cells.
- [ ] I can investigate the link between the reactivity of metals and the voltage produced in an electrochemical cell.

Introduction to Redox Reactions in Cells and Batteries

Redox reactions, short for reduction-oxidation reactions, are fundamental to the functioning of **electrochemical cells** and **batteries**. These reactions involve the **transfer** of electrons between substances, which is key to generating **electrical** energy in batteries.

In redox reactions, one substance **loses** electrons (**oxidation**), and another **gains** electrons (**reduction**). This transfer of electrons is what powers the chemical reactions in a cell or battery.

In order to make a **simple electrochemical cell**, we must use two different substances, usually metals of different **reactivity**, and an **electrolyte**.

Electrolytes and Their Role

Electrolytes are substances that can conduct electricity when dissolved in water.

The solution becomes **conductive** because it contains **ions**, which are charged particles that are **free to move**.

Creating a Simple Cell

A simple way to create an electrochemical cell is by using two different **metals** known as **electrodes**. *Looking at the diagram below*, we take two metals, say copper and zinc, and immerse them in an **electrolyte** solution.

Each metal has a different ability to **lose** or **gain** electrons which is based on its **reactivity**. In our example, zinc tends to lose electrons more easily than copper.

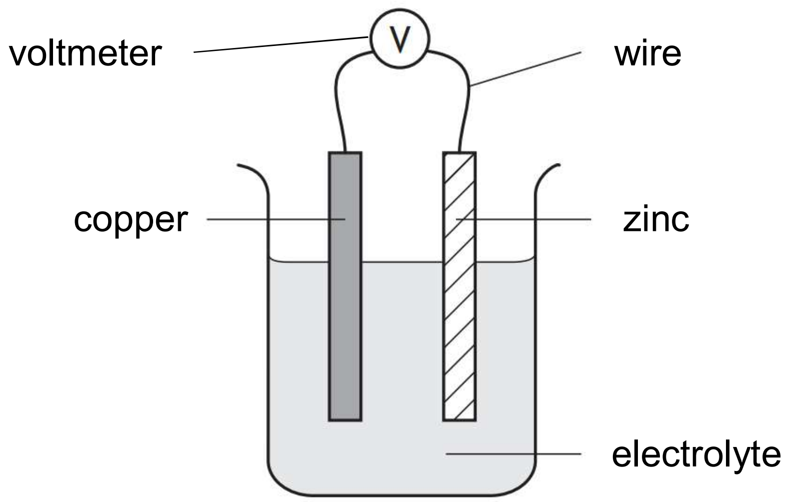
In an **electrochemical** cell, the **more** reactive metal loses the electrons through an **oxidation** reaction, the **less** reactive metal **gains** electrons through a **reduction** reaction.

When these metals are connected through a wire outside the solution, electrons will flow from the **zinc** to the **copper** through the wire. This flow of electrons is what we call **electricity**.

The **electricity** flows from the **most reactive** metal to the **least reactive** metal through the **wire**

The **ions complete** the **circuit** by moving through the **electrolyte**.

The wires can be attached to **voltmeter** to measure the voltage of the cell. The voltage is the measure of the **electrical potential** produced by the cell.



Perform the experiment on the next page and state the link between the reactivity and voltage produced in an electrochemical cell.

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___

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Reaction Title: Metal/copper cell (p36 lab book)

Aim: ____

____

Method:

Results:

Metal A : Metal B	Voltage (V)
Copper : Lead	
Copper : Aluminium	
Copper : Copper	
Copper : Iron	
Copper : Magnesium	
Copper : Zinc	

Write the redox equation for the zinc and copper redox reaction:

Oxidation:

Reduction :

Redox:

Conclusion: ____

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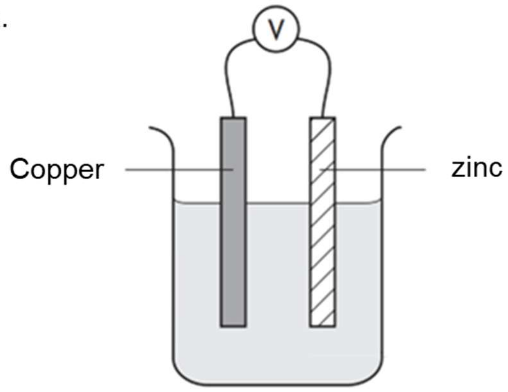
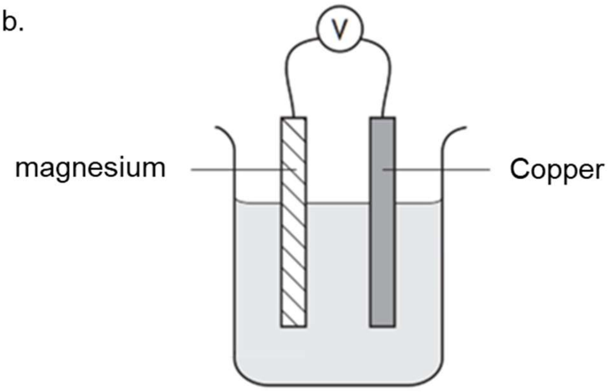
Questions

Use the following diagrams to answer the questions below.

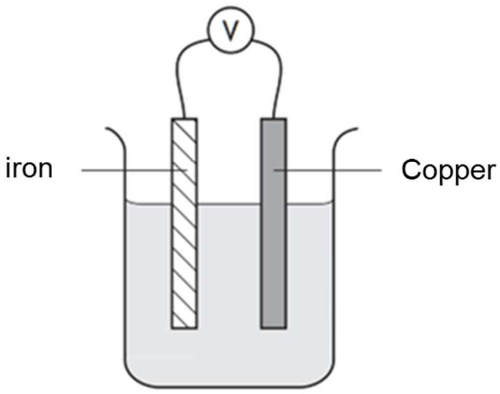
a.



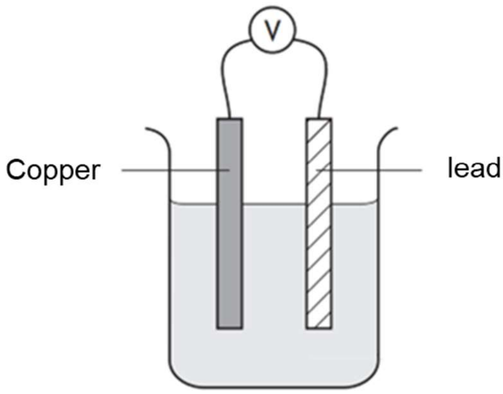
b.



c.



d.



1. Draw arrows on the diagrams above to show the flow of electrons.
2. Order the simple cells from lowest to highest voltage.

   ____________ ____________ ____________ ____________

3. State the purpose of the electrolyte in the simple cells.

4. Write the half-reactions and overall redox reaction for **b**.

Oxidation:

Reduction:

Redox:

Date: ________________

Half-cells

Learning Intentions

- To learn about making full electrochemical cells with half cells.

Success Criteria

- ☐ I can identify a full electrochemical cell.
- ☐ I can identify a salt bridge and explain its function in an electrochemical cell.
- ☐ I can write the full redox equation from half-cells.
- ☐ I can state the non- metal electrode used in a half-cell and explain why it is suitable.

Half-cells

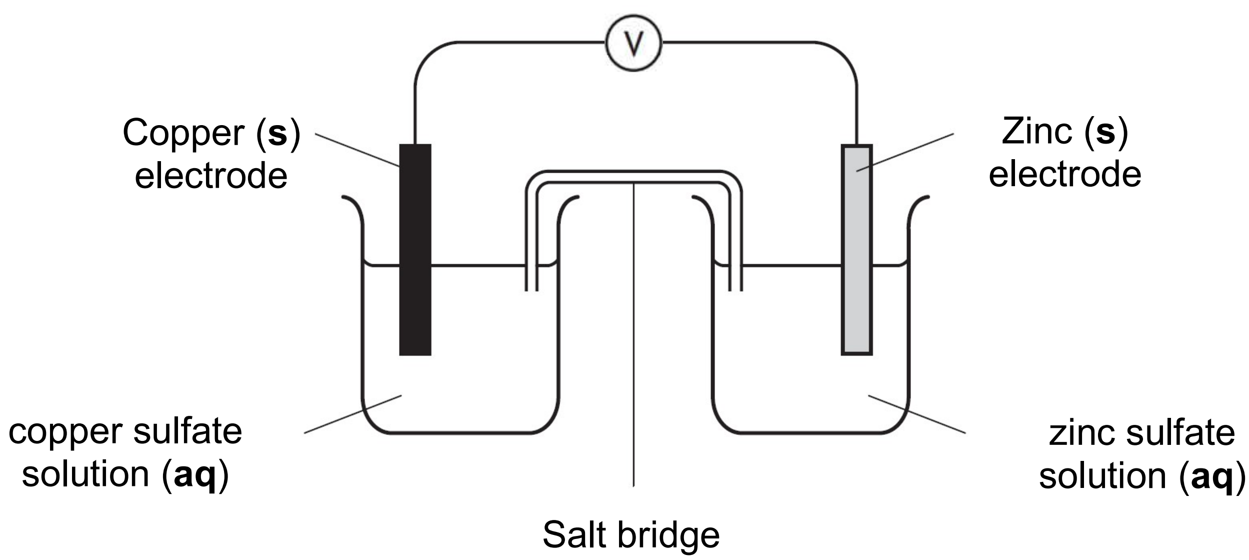
A half-cell is one part of an electrochemical cell. It consists of a **metal** sitting in a solution of its **own ions**.

For instance, a copper half-cell would be a copper metal placed in a **copper sulfate** solution. When you connect two such half-cells, say a copper half-cell and a zinc half-cell, you get a **full** electrochemical cell.

To connect the two half-cells together, the cells are linked by a **salt bridge**.

> The **salt bridge** in a full electrochemical cell **completes** the **circuit** as it allows the **ions** to **move freely** between the two **half** cells.

The voltage of this can be measured the same way as a simple cell.



Reaction Title: Half cells (p36 lab book)

Aim: ____

____

Method:

Results:

Metal A : Metal B	Voltage (V)
Copper : Lead	
Copper : Aluminium	
Copper : Copper	
Copper : Iron	
Copper : Magnesium	
Copper : Zinc	

Conclusion: ____

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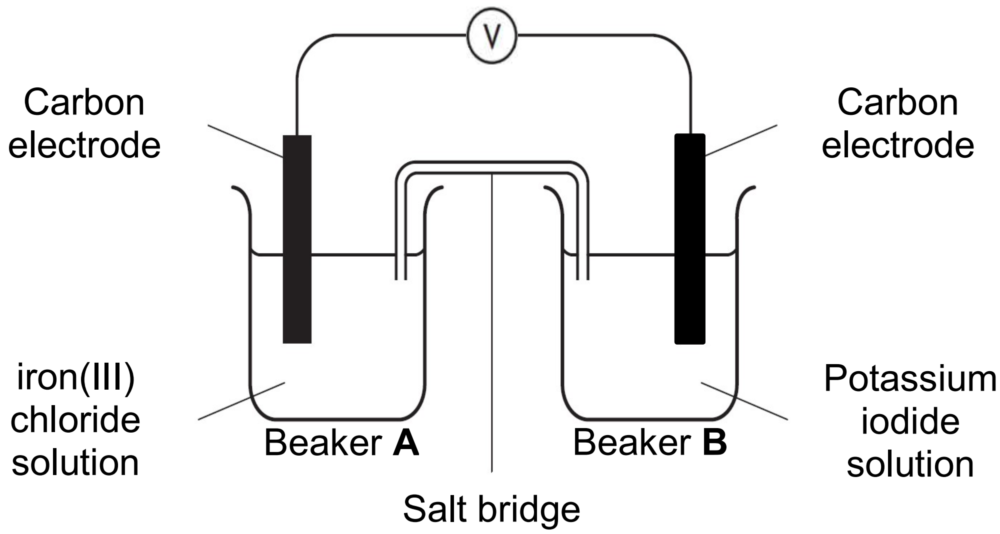
____

Half-cells with non-metals

Half-cells can be made with **conductive** non-metals as the **electrode**.

The most common **non-metal electrode** is **carbon** in the form of **graphite**.

Graphite can be used as an **electrode** as it contains **delocalised** electrons and will not **react** with the solutions.



Beaker **A:** $Fe^{3+}(aq)$ + e^- → $Fe^{2+}(aq)$

Beaker **B:** $2I^-(aq)$ → $I_2(l)$ + $2e^-$

Questions

1. Draw arrows on the diagram above to show the flow of electrons.
2. Explain why is carbon a suitable material to use as an electrode.

3. State the function of the salt bridge.

4. Identify the oxidation and reduction reactions above.
5. Write the overall redox reaction below:

Redox:

Reaction Title: Fruit Batteries (p36 lab booklet)

Aim: ____

____

Method:

Results:

Conclusion: ____

____

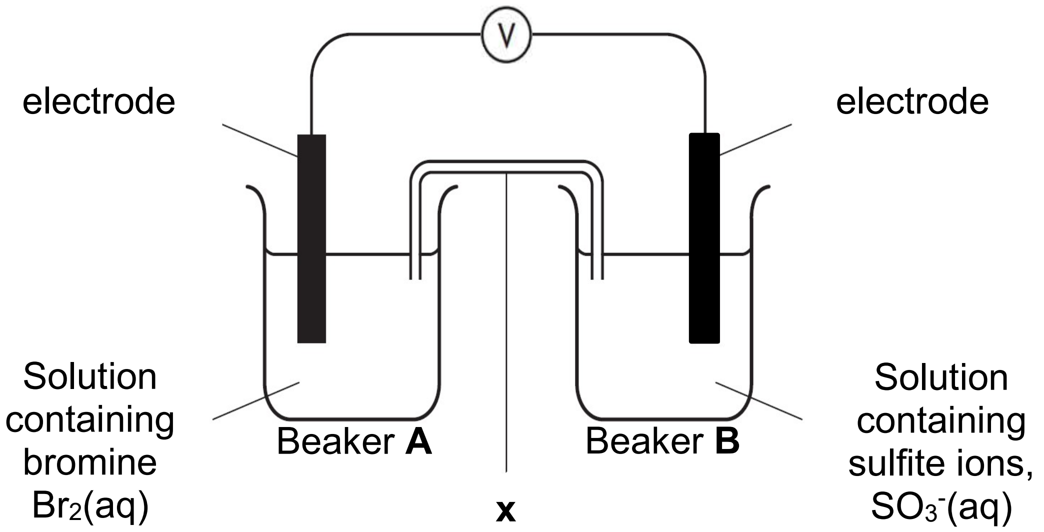
____

Evaluation: ____

____

____

Use the following diagrams to answer the questions below.



Beaker **A**: $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$

Beaker **B**: $SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

1. Draw arrows on the diagrams above to show the flow of electrons.
2. Label the oxidation and reduction reactions above.
3. Suggest a suitable non-metal electrode for the above diagram

____

4. State two reasons that this non-metal electrode can be used in electrochemical cells.

Reason 1:____

Reason 2:____

5. State the purpose of the electrolyte in the simple cells.

6. Name the piece of apparatus labelled X.

____

7. Write the overall redox reaction below:

Redox:

Date: ________________

Metals – Summary

Use the space below to summarise key points before doing past paper questions and extension work.

Extension questions:

Chemcord purple books (N5): page 101-123

SCHOLAR

Topic	2015	2016	2017	2018	2019
Metals	MC – 15,16,18 S2 – 9a, 9bi,9c	MC – 4,15 S2 – 10	MC – 15,16 S2 – 4b, 8, 10a	MC – 17,18,19,23 S2 – 11	MC – 18,19 S2 – 8d,10

MC = multiple choice section, S2 = section 2, the written section.