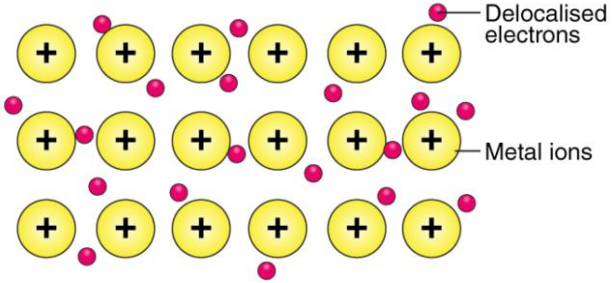
C4 – Reactivity of metals

Describe AND explain why potassium is more reactive than lithium. Potassium is **bigger** than lithium, so the **outer shell electron is further** from the nucleus, it is **easier to lose** the electron.	Complete the word equation: lithium + water -> lithium hydroxide + hydrogen	What would the charge on a potassium ion be? Why? K^{+1}, because potassium will lose one electron from it's outer shell.
Why does iodine have a low melting point? Weak intermolecular forces that don't require a lot of energy to break.	Why does graphite conduct electricity? There are delocalised electrons which can conduct electricity.	Draw a diagram to show the structure and bonding in metals. 

$500cm^3 = 0.5dm^3$

Concentration = mass ÷ volume

= 10 grams ÷ $0.5dm^3$

= 20 g/dm^3

C4 - Chemical Changes

C1 - Atomic structure and the periodic table

C2 - Bonding, structure, and the properties of matter

C3 - Quantitative chemistry

C4 - Chemical changes

C5 - Energy changes

C6 - The rate and extent of chemical change

C7 - Organic chemistry

C8 - Chemical analysis

C9 - Chemistry of the atmosphere

C10 - Using resources

Reactivity

- Reactivity series
- Oxidation and reduction
- Reactions of metals, metal oxides and metal carbonates with acids
- pH scale and neutralisation
- Strong and weak acids

Electrolysis

- Electrolysis of molten ionic compounds
- Electrolysis of aqueous ionic compounds
- Electrolysis of aluminium oxide

TRIPLE **HIGHER**

Most metals do not exist on their own...

Key definition:

Metals exist in the Earth's crust and are often combined with other elements to form ores.

A metal ore is a type of rock that contains a metal compound.

Challenge: Why is it important to extract pure metals from ores?

Define metal ores and **explain** why metals are extracted from their ores

Pure metals can be extracted from their ores so they can be used in industry and everyday life.

*Why are some metals **not** extracted from their ores?*

Metals will only be extracted from their ores if it is economically viable

Describe oxidation and reduction in terms of oxygen

IRON OXIDE ORE

In the Earth's crust, metals are bonded to other elements, usually oxygen.

Oxidation is the **gain** of oxygen by a substance:
Metal + oxygen → metal oxide

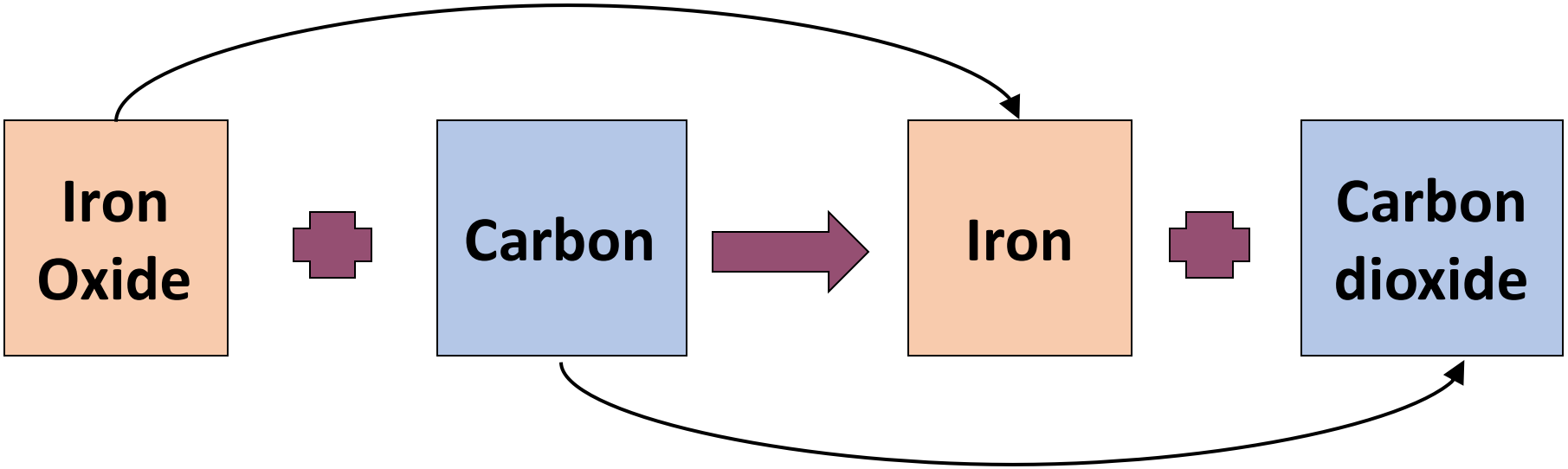
When we extract the metal from its ore, oxygen has to be removed.

Reduction is the **loss** of oxygen by a substance:
Metal oxide → metal + oxygen

Describe oxidation and reduction in terms of oxygen

Oxidation or reduction?

Iron oxide has **lost** oxygen, so it has been **reduced**



Carbon has **gained** oxygen, so it has been **oxidised**

Recognise which elements have been oxidised or reduced in chemical equations

Task: In each of the reaction below, decide whether the substance shown in bold has been oxidised or reduced:

a) aluminium + oxygen → aluminium oxide — Aluminium has been oxidised

b) calcium oxide → calcium + oxygen — Calcium has been reduced

c) iron oxide + carb — Iron has been reduced, carbon has been oxidised

d) carbon + **zinc ox** — Zinc has been reduced, carbon has been oxidised

e) alumi — Chromium has been reduced, aluminium has been oxidised

f) hydro — Copper has been reduced, hydrogen has been oxidised

g) burni — Magnesium + oxygen, so magnesium has been oxidised

Challenge: One step in the manufacture of lead is the reduction of lead oxide with carbon. Lead and carbon dioxide are the products of this reaction.

a) lead oxide + carbon → lead + carbon dioxide

b) oxygen removed

Potassium: acts as gunpowder in fireworks!

Francium: so reactive that only about 20g exists in the Earth's crust at any time!

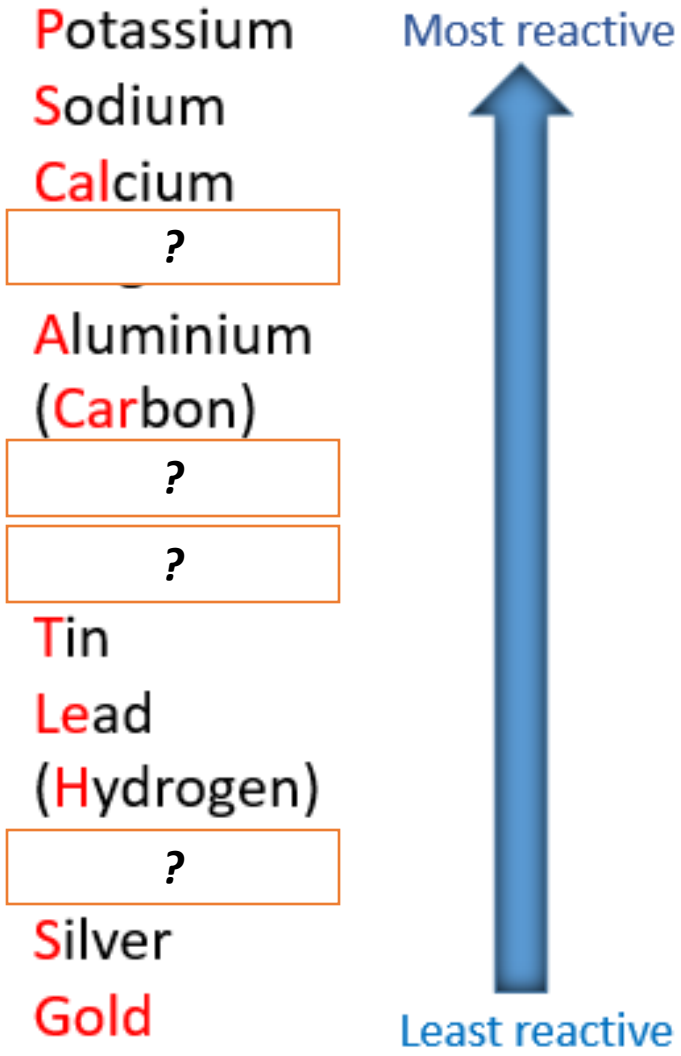
The reactivity of a metal depends on the **electronic structure**.

Metals with **different reactivities** have **different uses**.

Silver and gold: used in jewellery because it is so unreactive.

Copper: used in water pipes because it is so unreactive.

Metals can be put into order of reactivity. This is called the ***reactivity series***:



Task: Copy and complete the reactivity series into your book, using the results from your experiment

Challenge: Why is potassium more reactive than sodium?
Hint: Think about the electronic structure of both atoms!

Key definition:

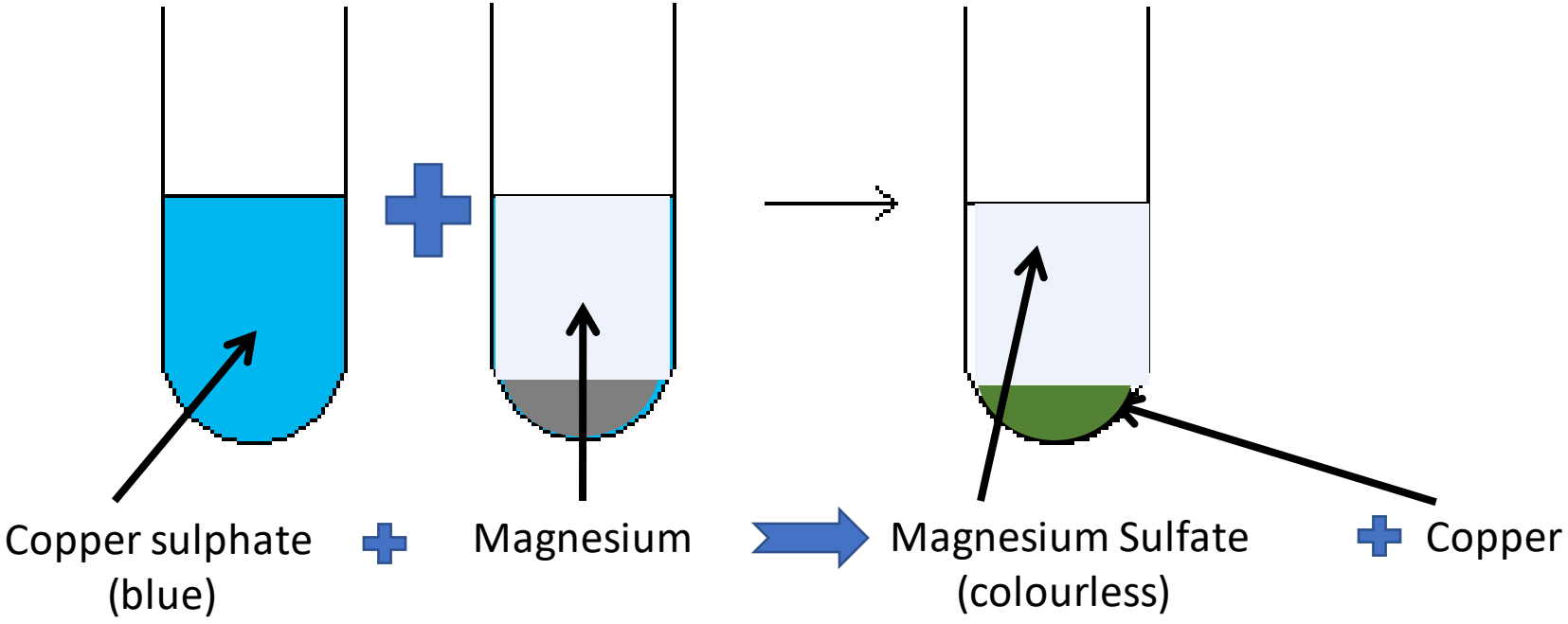
Displacement – when a more reactive metal takes the place of a less reactive metal in a compound

Single Displacement Reaction

Always use the word displace in your answers – if you say 'steals' or 'replaces' you will not get the marks.

This can be **observed** by chemical reactions.

- If a **more reactive** metal is placed into a different metal solution it will react and **displace** the other metal.
- Magnesium is **more reactive** than copper and **displaces it**.



Calcium + lead oxide → Calcium oxide + lead

Aluminium + copper oxide → Aluminium oxide + copper

Iron + magnesium sulphate → NO REACTION

Challenge: Which substances have been oxidised and which substances have been reduced?

Will there be a reaction?

Task: For each of the reactions below, decide whether there will be a reaction or not. For those that WILL react, predict the products of the reaction.

1. Zinc oxide + calcium → calcium oxide + zinc
2. Magnesium + iron oxide → magnesium oxide + zinc
3. Zinc + tin oxide → zinc oxide + tin
4. Magnesium sulphate + Zinc → *NO REACTION*
5. Calcium + copper oxide → calcium oxide + copper
6. Magnesium + iron sulphate → magnesium sulphate + iron
7. Tin oxide + Copper → *NO REACTION*
8. Gold + copper oxide → *NO REACTION*

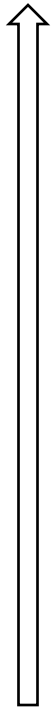
Challenge: Which substances have been oxidised and which substances have been reduced?

You Do It Alone

To **describe** how the reactivity of a metal affects how it is extracted from it's ore.

How are they extracted?

MOST REACTIVE



LEAST REACTIVE

Potassium

Sodium

Calcium

Magnesium

Aluminium

Carbon

Zinc

Iron

Copper

Silver

Gold

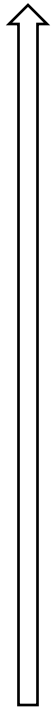
Extracted from their ores by ***electrolysis***.

Extracted from their ores by ***reduction with carbon.***

No extraction necessary – found pure in the ground.

To **describe** how the reactivity of a metal affects how it is extracted from it's ore.

MOST REACTIVE



LEAST REACTIVE

Potassium

Sodium

Calcium

Magnesium

Aluminium

Carbon

Zinc

Iron

Copper

Silver

Gold

Carbon is *more reactive* than zinc, iron and copper. What would happen if you reacted iron oxide with carbon?

Carbon would displace iron!

Extracted from their ores by ***reduction with carbon.***

To **explain** how metals are extracted by reduction using carbon.

As carbon is **more reactive** than iron, carbon will ***displace*** iron:

Iron oxide + carbon → ? + ?

Which element has been oxidised? Why?

Carbon has been oxidised because it has **gained oxygen** to form carbon dioxide.

Which element has been reduced? Why?

Iron has been reduced because iron oxide has **lost oxygen** to form pure iron.

Challenge: Write a symbol equation for the reaction above (the chemical formula for iron oxide is Fe_2O_3).

Super challenge: Balance the symbol equation!

https://www.youtube.com/watch?v=gvNuMpxqG7Q

To **explain** how metals are extracted by reduction using carbon.

Complete the word equations to show how metals are extracted from their ores:

Iron oxide + carbon →
Zinc oxide + carbon →
Copper oxide + carbon →

In each of the equations below, identify which element has been oxidised and which has been reduced.

Write symbol equations for the reactions, using the chemical formulae below:
Iron oxide (Fe2O3)
Zinc oxide (ZnO)
Copper oxide (CuO)

Challenge: Why can't magnesium and aluminium be extracted from their ore by reacting them with carbon?

To **explain** how metals are extracted by reduction using carbon.

Complete the word equations to show how metals are extracted from their ores:

a) Iron oxide + carbon → ***Iron + carbon dioxide***
b) Zinc oxide + carbon → ***Zinc + carbon dioxide***
c) Copper oxide + carbon → ***Copper + carbon dioxide***

a) Iron has been reduced (lost oxygen), carbon has been oxidised (gained oxygen)
b) Zinc has been reduced (lost oxygen), carbon has been oxidised (gained oxygen)
c) Copper has been reduced (lost oxygen), carbon has been oxidised (gained oxygen)

Write symbol equations for the reactions, using the chemical formulae below:

Iron oxide (Fe_2O_3) ***$Fe_2O_3 + C \rightarrow Fe + CO_2$***
Zinc oxide (ZnO) ***$ZnO + C \rightarrow Zn + CO_2$***
Copper oxide (CuO) ***$CuO + C \rightarrow Cu + CO_2$***

Challenge: Why can't magnesium and aluminium be extracted from their ore by reacting them with carbon?

HT ONLY!

OIL RIG!

Oxidation

Is

Loss of electrons (and gain of oxygen)

Reduction

Is

Gain of electrons (and gain of oxygen)

Ionic Equations

A sodium ion ($Na^{+)}$ loses an electron (e-):

$Na^+ - e^- \rightarrow Na$

Bromide ion loses an electron to become Br_2

2 Br^- → Br_2 + 2e^-

Each bromide ion loses 1 electron, in total 2 electrons are lost!

1. Na^+ - **e-** → Na
2. $2Cl^-$ - **2e-** → Cl_2 OR 2Cl- → Cl_2 + **2e-**
3. Mg^{2+} **+ 2e-** → **Mg**
4. Br^- **- 2e-** → **Br_2** OR 2Br- → Br_2 + **2e-**

LO: I can write ionic equations for displacement reactions

I DO

1. Identify the species that have been oxidised and reduced in this equation
2. Write their individual charges

Oxidation is loss of electrons

Reduction is gain of electrons

OIL RIG

Magnesium + Copper oxide → Copper + Magnesium oxide

$Mg + CuO \rightarrow Cu + MgO$

$Mg + Cu^{2+} O^{2-} \rightarrow Cu + Mg^{2+} O^{2-}$

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

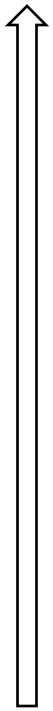
This is the ionic equation

Challenge: The magnesium is higher up the reactivity series, so will displace the copper.

LO: Describe what happens in displacement ionic equations (6)

Displacement equations

MOST REACTIVE



LEAST REACTIVE

Potassium

Sodium

Calcium

Magnesium

Aluminium

Carbon

Zinc

Iron

Copper

Silver

Gold

The more reactive metal loses electrons: **Oxidation**

The less reactive metal gains electrons: **Reduction**

LO: I can write ionic equations for displacement reactions

YOU DO

1. What are the products of this equation
2. Identify the species that have been oxidised and reduced in this equation
3. Explain how you know which species has been oxidised and reduced
4. Write an ionic equation for this reaction

Oxidation is loss of electrons
Reduction is gain of electrons
OIL RIG

Zinc + Copper oxide $\longrightarrow$

$Zn + CuO \longrightarrow$

LO: I can write ionic equations for displacement reactions

YOU DO

1. What are the products of this equation
2. Identify the species that have been oxidised and reduced in this equation

Zinc has been oxidised and the copper ion has been reduced

3. Explain how you know which species has been oxidised and reduced

Zinc has lost electrons and the copper ion has gained electrons

4. Write an ionic equation for this reaction

Oxidation is loss of electrons

Reduction is gain of electrons

OIL RIG

Zinc + Copper oxide $\longrightarrow$ Zinc oxide + Copper

$Zn + CuO \longrightarrow ZnO + Cu$

$Zn + Cu^{2+}O^{2-} \longrightarrow Zn^{2+}O^{2-} + Cu$

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

PRINT

C4 – Reactivity of metals

Describe AND explain why potassium is more reactive than lithium.	Complete the word equation: lithium + water ->	What would the charge on a potassium ion be? Why?
Why does iodine have a low melting point?	Why does graphite conduct electricity?	Draw a diagram to show the structure and bonding in metals.

Challenge – *What is the concentration of a 500cm^3 solution which contains 10 grams of hydrochloric acid?*

Most metals do not exist on their own...

Key definition:

Metals exist in the Earth's crust and are often combined with other elements to form ores.

A metal ore is a type of rock that contains a metal compound.

Challenge: Why is it important to extract pure metals from ores?

Describe oxidation and reduction in terms of oxygen

IRON OXIDE ORE

In the Earth's crust, metals are bonded to other elements, usually oxygen.

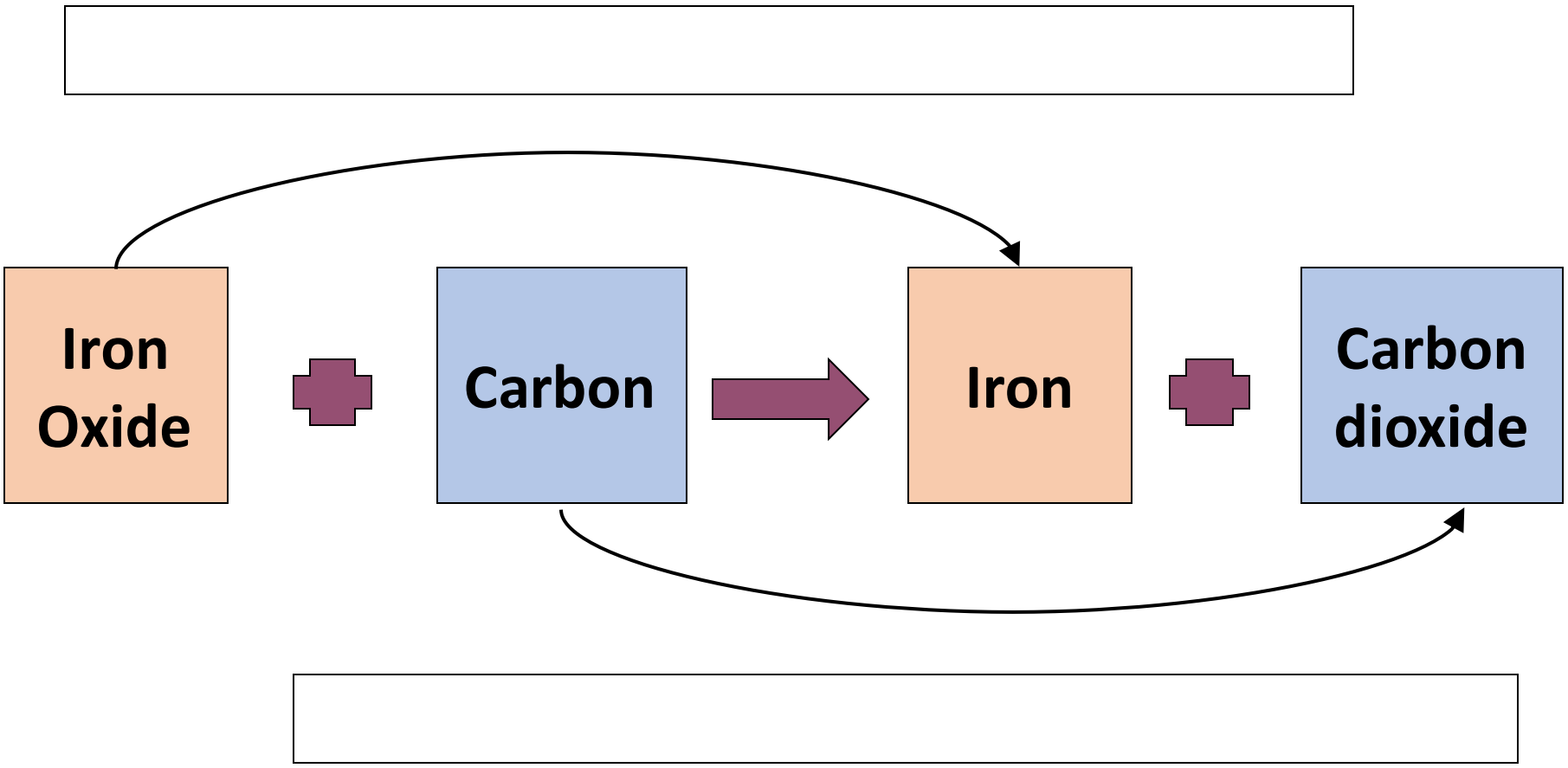
Oxidation is the **gain** of oxygen by a substance:
Metal + oxygen → metal oxide

When we extract the metal from its ore, oxygen has to be removed.

Reduction is the **loss** of oxygen by a substance:
Metal oxide → metal + oxygen

Describe oxidation and reduction in terms of oxygen

Oxidation or reduction?



Task: In each of the reaction below, decide whether the substance shown in bold has been oxidised or reduced:

a) **aluminium** + oxygen → aluminium oxide

b) **calcium oxide** → calcium + oxygen

c) **iron oxide + carbon** → iron + carbon dioxide

d) **carbon** + **zinc oxide** → carbon dioxide + zinc

e) **aluminium** + **chromium oxide** → aluminium oxide + chromium

f) **hydrogen** + **copper oxide** → copper + water

g) burning magnesium in air

Challenge: One step in the manufacture of lead is the reduction of lead oxide with carbon. Lead and carbon dioxide are the products of this reaction.

a) lead oxide + carbon → lead + carbon dioxide

b) oxygen removed

Metals can be put into order of reactivity. This is called the ***reactivity series***:

Potassium
Sodium
Calcium
Magnesium
Aluminium
(Carbon)
Zinc
Iron
Tin
Lead
(Hydrogen)
Copper
Silver
Gold

Most reactive

Least reactive

Key definition:

Displacement – when a more reactive metal takes the place of a less reactive metal in a compound

Single Displacement Reaction

Will there be a reaction?

Task: For each of the reactions below, decide whether there will be a reaction or not. For those that WILL react, predict the products of the reaction.

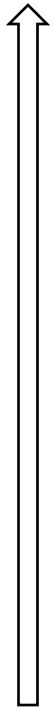
1. Zinc oxide + calcium →
2. Magnesium + iron oxide →
3. Zinc + tin oxide →
4. Magnesium sulphate + Zinc →
5. Calcium + copper oxide →
6. Magnesium + iron sulphate →
7. Tin oxide + Copper →
8. Gold + copper oxide →

Challenge: Which substances have been oxidised and which substances have been reduced?

You Do It Alone

How are they extracted?

MOST REACTIVE



LEAST REACTIVE

Potassium

Sodium

Calcium

Magnesium

Aluminium

Carbon

Zinc

Iron

Copper

Silver

Gold

Extracted from their ores by ***electrolysis***.

Extracted from their ores by ***reduction with carbon.***

No extraction necessary – found pure in the ground.

As carbon is **more reactive** than iron, carbon will ***displace*** iron:

Iron oxide + carbon →

Which element has been oxidised? Why?

Which element has been reduced? Why?

Complete the word equations to show how metals are extracted from their ores:

Iron oxide + carbon →
Zinc oxide + carbon →
Copper oxide + carbon →

In each of the equations below, identify which element has been oxidised and which has been reduced.

To **explain** how metals are extracted by reduction using carbon.

Complete the word equations to show how metals are extracted from their ores:

Iron oxide + carbon →
Zinc oxide + carbon →
Copper oxide + carbon →

In each of the equations below, identify which element has been oxidised and which has been reduced.

Write balanced symbol equations for the reactions. To do this, you will need to work out the chemical formulae for iron oxide, zinc oxide and copper oxide:
Fe forms a 3+ ion
Zinc forms a 2+ ion
Copper forms a 2+ ion

Challenge: Why can't magnesium and aluminium be extracted from their ore by reacting them with carbon?

HT ONLY!

OIL RIG!

Oxidation

Is

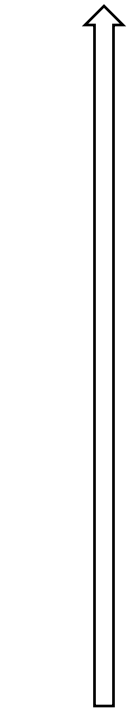
Loss of electrons (and gain of oxygen)

Reduction

Is

Gain of electrons (and gain of oxygen)

Displacement equations



Potassium

Sodium

Calcium

Magnesium

Aluminium

Carbon

Zinc

Iron

Copper

Silver

Gold

The more reactive metal loses electrons:

Oxidation

The less reactive metal gains electrons:

Reduction

Ionic Equations

A sodium ion ($Na^{+)}$ gains an electron (e-):

$Na^+ + e^- \rightarrow Na$

Bromide ion loses an electron to become Br_2

$Br^- \rightarrow Br_2 + e^-$

Each bromide ion loses 1 electron, in total 2 electrons are lost!

1. $Na^+ +$ _____ $\rightarrow Na$
2. $2Cl^-$ - _____ $\rightarrow Cl_2$ OR 2Cl- $\rightarrow Cl_2 +$ ____
3. $Mg^{2+} \rightarrow$
4. $Br^- \rightarrow$ OR 2Br- $\rightarrow Br_2 +$ ____

LO: I can write ionic equations for displacement reactions

I DO

1. Identify the species that have been oxidised and reduced in this equation
2. Write their individual charges

Oxidation is loss of electrons

Reduction is gain of electrons

OIL RIG

Magnesium + Copper oxide → Copper + Magnesium oxide

$Mg + CuO \rightarrow Cu + MgO$

$Mg + Cu^{2+} O^{2-} \rightarrow Cu + Mg^{2+} O^{2-}$

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

This is the ionic equation

Challenge: The magnesium is higher up the reactivity series, so will displace the copper.

LO: I can write ionic equations for displacement reactions

YOU DO

1. What are the products of this equation
2. Identify the species that have been oxidised and reduced in this equation
3. Explain how you know which species has been oxidised and reduced
4. Write an ionic equation for this reaction

Oxidation is loss of electrons
Reduction is gain of electrons
OIL RIG

Zinc + Copper oxide $\longrightarrow$

$Zn + CuO \longrightarrow$

Metal Oxides

- Metals react with oxygen in the air to produce metal oxides
- Oxidation and reduction involve the addition or removal of oxygen from a substance and are called **redox** reactions
- Oxidation is a reaction in which:
    - **Oxygen** is **added** to an element or a compound

- Reduction is a reaction in which:
    - **Oxygen** is **removed** from an element or a compound

- A common example is the reaction with red-brown copper metal to produce black copper oxide:

$$2Cu + O_2 \rightarrow 2CuO$$

- In this reaction copper metal has been **oxidised** since oxygen has been **added** to it
- Another example is the reaction of zinc oxide with carbon:

$$ZnO + C \rightarrow Zn + CO$$

- In this reaction the zinc oxide has been **reduced** since it has **lost** oxygen. The carbon atom has been oxidised since it has gained oxygen.

REDOX are simultaneous reactions as they occur at the same time in the same reaction.

The Reactivity Series of Metals

- Based on how they react with other substances, a **reactivity series** can be produced in which the metals are placed in order of their reactivity
- Metal atoms form **positive ions** by loss of electrons when they react with other substances
- The **tendency** of a metal to lose electrons is a measure of how reactive the metal is
- A metal that is high up on the series loses electrons easily and is thus **more reactive** than one which is lower down on the series
- Note that although carbon and hydrogen are nonmetals, they are included in the series as they are useful in **extracting metals** from their oxides by reduction processes

Metals Reacting with Water

Reactivity with water

- Some metals react with water
- Metals **above hydrogen** in the reactivity series will react with water, but the reaction may be very slow
- Metals that react with cold water form a metal hydroxide and hydrogen gas:

metal + water → metal hydroxide + hydrogen

- For example calcium:

$$Ca + 2H_2O \rightarrow Ca(OH)_2 + H_2$$

calcium + water → calcium hydroxide + hydrogen

Metals Reacting with Acids

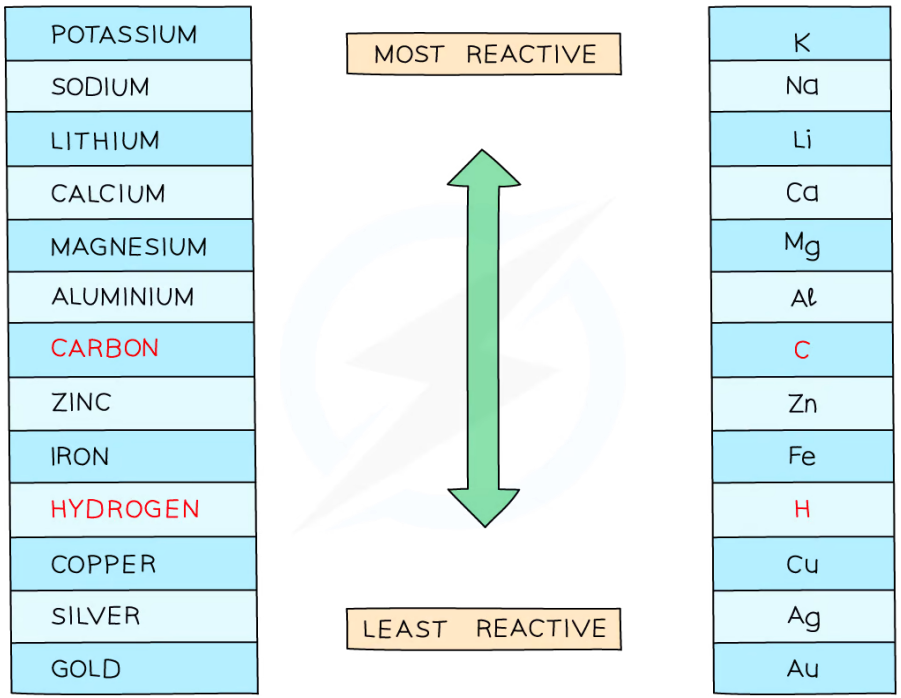
- Most metals react with dilute acids such as HCl
- Only the ones below hydrogen in the reactivity series will not react with acids
- When acids and metals react, the hydrogen atom in the acid is replaced by the metal atom to produce a salt and hydrogen gas:

metal + acid → metal salt + hydrogen

For example iron: $Fe + 2HCl \rightarrow FeCl_2 + H_2$

iron + hydrochloric acid → iron(II)chloride + hydrogen

- In both these types of reactions (water and acids) the metals are becoming positive ions
- The reactivity of the metals is related to their tendency to become an ion
- The **more reactive** the metal the **more easily** it becomes an ion (by losing electrons)



Metal	Reaction with water	Reaction with acid
Most reactive		
Potassium	Reacts violently	Reacts violently
Sodium	Reacts quickly	Reacts violently
Calcium	Reacts less strongly	Reacts vigorously
Magnesium		Reacts vigorously
Zinc		Reacts less strongly
Iron		Reacts less strongly
Hydrogen		
Copper		
Least reactive		

Displacement Reactions

- The reactivity of metals decreases going down the reactivity series.
- This means that a more reactive metal will **displace** a less reactive metal from its compounds
- Two examples are:
  - Reacting a metal with a metal oxide (by heating)
  - Reacting a metal with an aqueous solution of a metal compound

- For example it is possible to reduce copper(II) oxide by heating it with zinc.
- The reducing agent in the reaction is zinc:

$$Zn + CuO \rightarrow ZnO + Cu$$

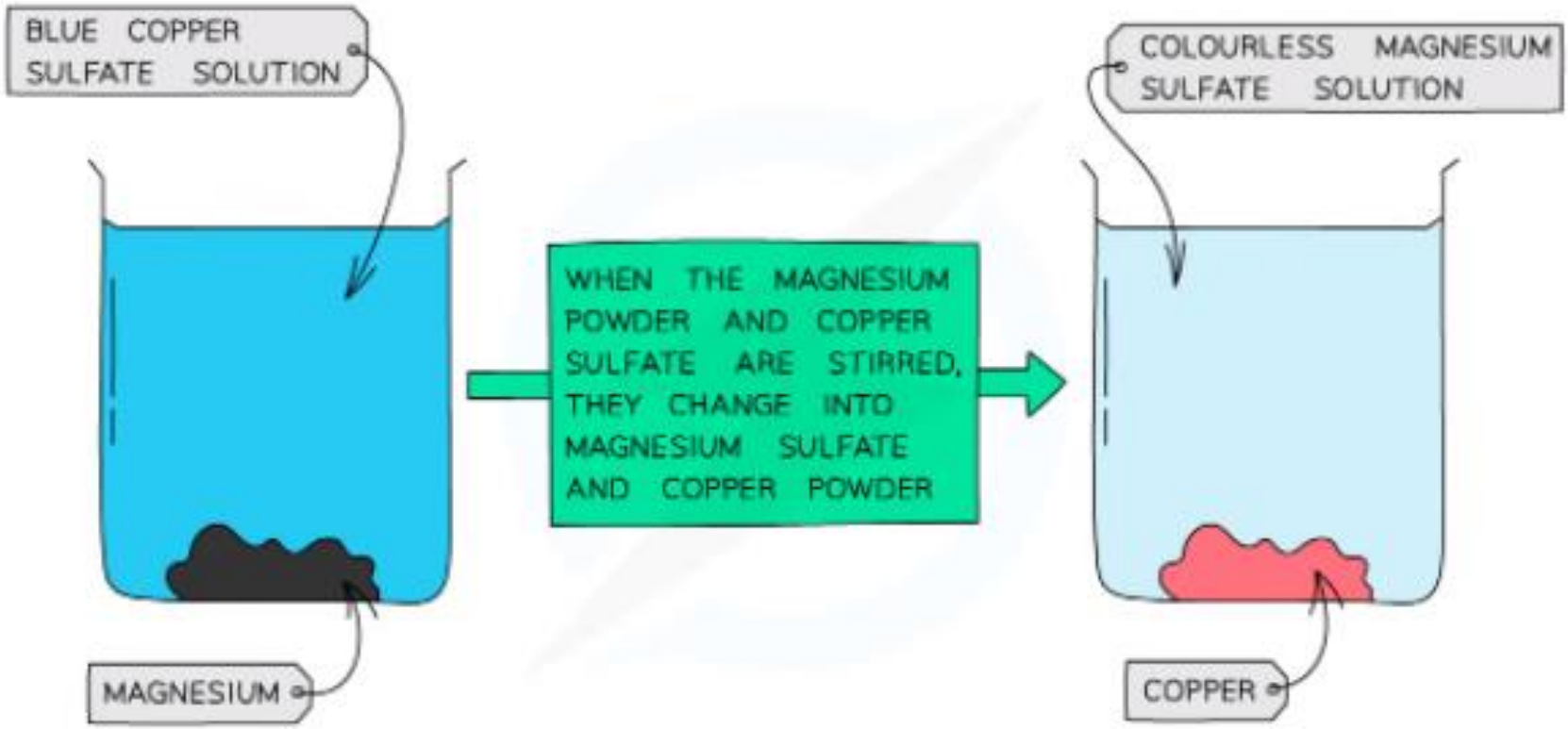
zinc + copper oxide → zinc oxide + copper

Mixture	Products	Equation for Reaction
Magnesium and Iron (II) Sulfate	Magnesium Sulfate and Iron	$Mg + FeSO_4 \longrightarrow MgSO_4 + Fe$
Zinc and Sodium Chloride	No Reaction as Sodium is above Zinc	—
Lead and Silver Nitrate	Lead (II) Nitrate and Silver	$Pb + 2AgNO_3 \longrightarrow Pb(NO_3)_2 + 2Ag$
Copper and Calcium Chloride	No Reaction as Calcium is more reactive than Copper	—
Iron and Copper (II) Sulfate	Iron (II) Sulfate and Copper	$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

- The reactivity between two metals can be compared using **displacement reactions** in salt solutions of one of the metals
- This is easily seen as the more reactive metal slowly **disappears** from the solution, **displacing** the less reactive metal
- For example, magnesium is a reactive metal and can displace copper from a copper sulfate solution:

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

- The blue color of the $CuSO_4$ solution **fades** as colorless magnesium sulfate solution is formed.
- Copper coats the surface of the magnesium and also forms solid metal which falls to the bottom of the beaker



- The Earth's crust contains metals and metal compounds such as gold, copper, iron oxide and aluminium oxide
- Useful metals are often **chemically combined** with other substances forming **ores**
- A metal ore is a rock that contains enough of the metal to make it worthwhile extracting
- The position of the metal on the reactivity series determines the method of extraction
- Higher placed metals (above carbon) have to be extracted using **electrolysis** as they are too reactive and cannot be reduced by carbon
- Lower placed metals can be extracted by heating with carbon which **reduces** them

METAL	ABBREVIATION
MOST REACTIVE	
POTASSIUM	EXTRACTED BY ELECTROLYSIS OF THE MOLTEN CHLORIDE OR MOLTEN OXIDE LARGE AMOUNTS OF ELECTRICITY REQUIRED SO EXPENSIVE PROCESS
SODIUM	
LITHIUM	
CALCIUM	
MAGNESIUM	
ALUMINIUM	
CARBON	
ZINC	EXTRACTED BY HEATING WITH A REDUCING AGENT SUCH AS CARBON OR CARBON MONOXIDE IN A BLAST FURNACE CHEAP PROCESS AS CARBON IS CHEAP AND CAN BE SOURCE OF HEAT AS WELL
IRON	
HYDROGEN	
COPPER	
SILVER	FOUND AS PURE ELEMENTS
GOLD	
LEAST REACTIVE	

Oxidation & Reduction in Terms of Electrons

- Displacement reactions can be analysed in terms redox reactions by studying the **transfer** of **electrons**
- For the example of magnesium and copper sulfate, a balanced equation can be written in terms of the **ions** involved:

$$Mg\ (s) + Cu^{2+}\ (aq) + SO_4^{2-}\ (aq) \rightarrow Mg^{2+}\ (aq) + SO_4^{2-}\ (aq) + Cu\ (s)$$

- The sulfate ions, SO_4^{2-}, appear on both sides of the equation unchanged hence they are spectator ions and do not participate in the chemistry of the reaction so can be omitted:

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

- This equation is an example of a balanced **ionic equation** which can be further split into two **half equations** illustrating oxidation and reduction individually:

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

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

- The magnesium atoms are thus **oxidised** as they **lose** electrons
- The copper ions are thus **reduced** as they **gain** electrons

Worked Example

Zinc displaces copper from a solution of copper(II)sulfate. Using ionic equations, determine which species undergoes oxidation and which species undergoes reduction.

Answer

Step	Working out
1. Write out the full equation	$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$
2. Write out the ionic equation	$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + Cu(s)$
3. Species oxidised	By analysing the ionic equation, it becomes clear that zinc has become oxidised as it has **lost** electrons: • $Zn(s) \longrightarrow Zn^{2+}(aq)$
4. Species reduced	Copper has been reduced as it has **gained** electons: • $Cu^{2+}(aq) \longrightarrow Cu(s)$

C4 – Reactivity of acids

What is a displacement reaction? When a more reactive metal takes the place of a less reactive metal.	Complete the word equation: magnesium + iron oxide → magnesium oxide + iron	What is a metal ore? A rock containing a metal compound.
Which has been oxidised and which has been reduced? carbon + zinc oxide → carbon Carbon has been oxidised, zinc has been reduced.	Write the half equation for the reaction below: $Mg + Zn^{2+} \rightarrow Zn + Mg^{2+}$	What is oxidation and reduction in terms of electrons? OILRIG! Oxidation is loss, reduction is gain

$300cm^3 = 0.3dm^3$

Concentration = mass ÷ volume

= 6 grams ÷ $0.3dm^3$

= 20 g/dm^3

C4 - Chemical Changes

C1 - Atomic structure and the periodic table

C2 - Bonding, structure, and the properties of matter

C3 - Quantitative chemistry

C4 - Chemical changes

C5 - Energy changes

C6 - The rate and extent of chemical change

C7 - Organic chemistry

C8 - Chemical analysis

C9 - Chemistry of the atmosphere

C10 - Using resources

Reactivity

- Reactivity series
- Oxidation and reduction
- Reactions of metals, metal oxides and metal carbonates with acids
- pH scale and neutralisation
- Strong and weak acids

Electrolysis

- Electrolysis of molten ionic compounds
- Electrolysis of aqueous ionic compounds
- Electrolysis of aluminium oxide

TRIPLE **HIGHER**

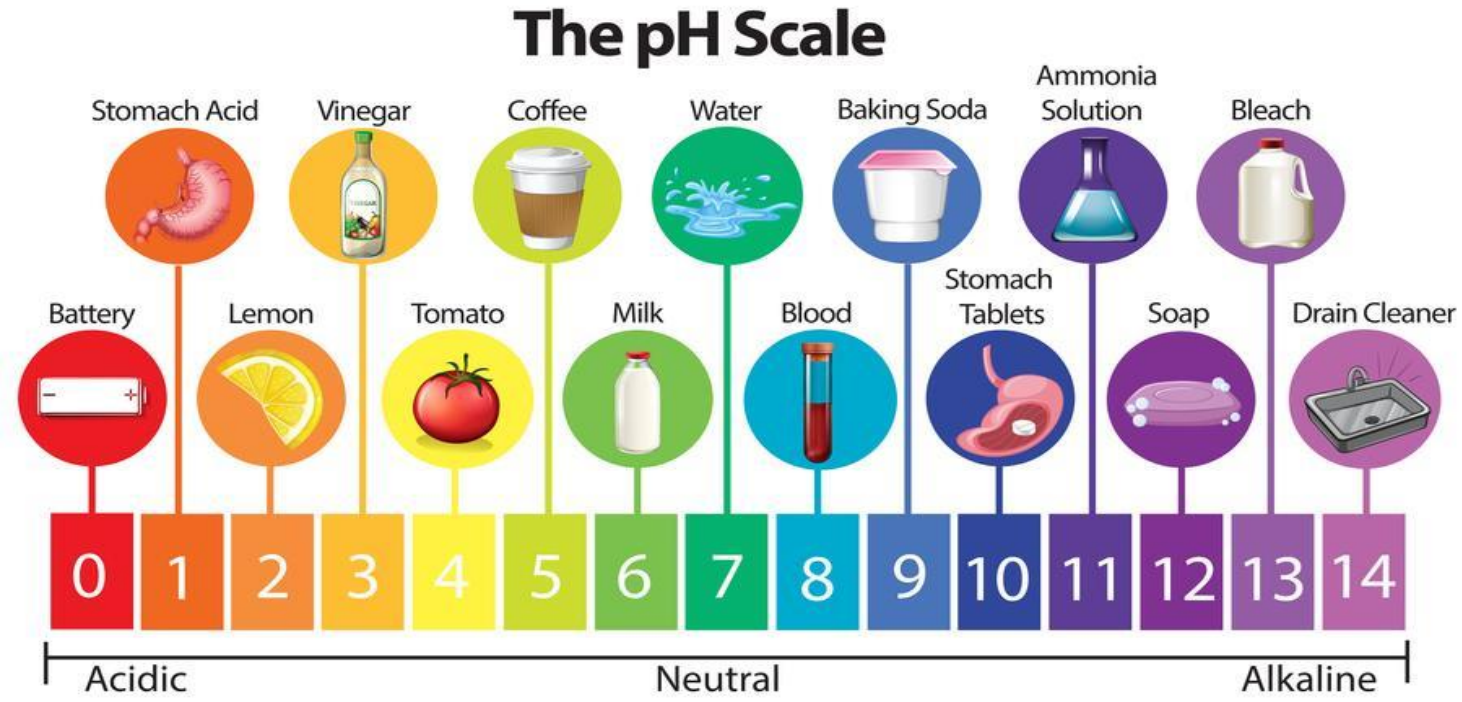
Describe and **explain** what the pH scale shows

- pH **1-6**
- Go **red/orange** in universal indicator
- Usually taste **sour**

- pH **7 ONLY**
- Goes **green** universal indicator

- pH **8-14**
- Go **blue/purple** in universal indicator
- Usually '**soapy**' to touch

ACID **NEUTRAL** **ALKALI**

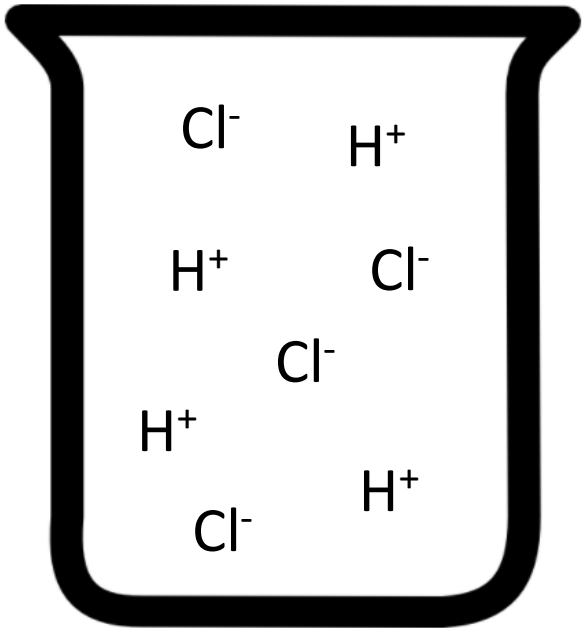


What is an acid?

What is an acid? What acids have you heard of before?

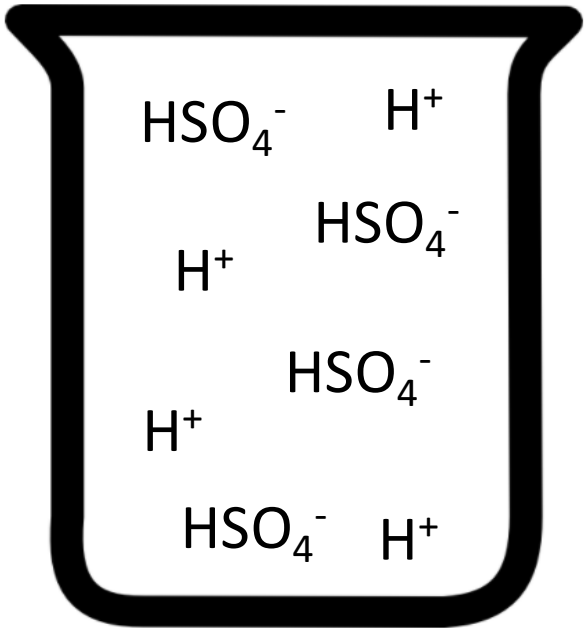
Hydrochloric acid

HCl



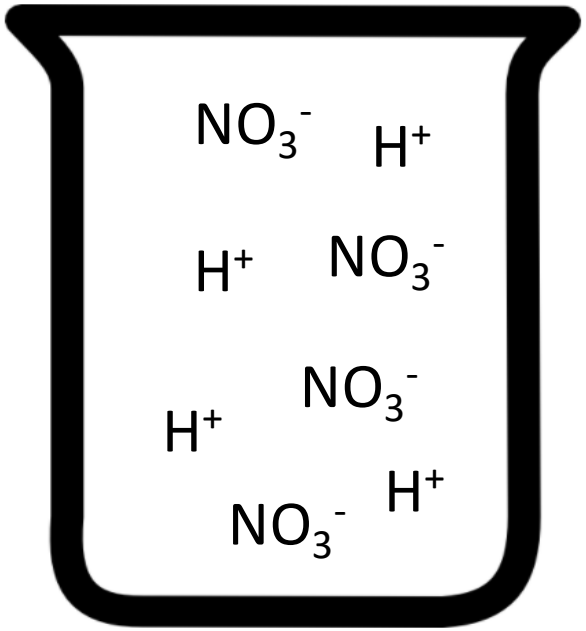
Sulphuric acid

H_2SO_4

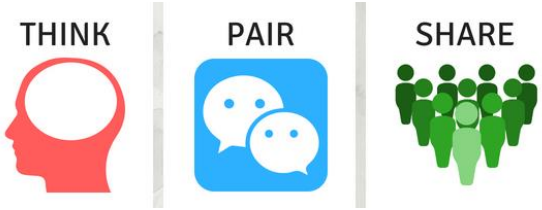


Nitric acid

HNO_3



Describe the reactions between metals and acids

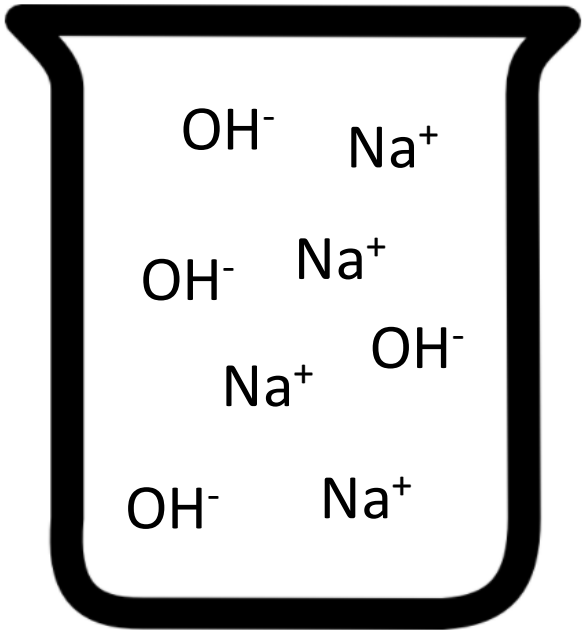


What is an alkali?

An alkali is a solution that contains OH- (hydroxide) ions.

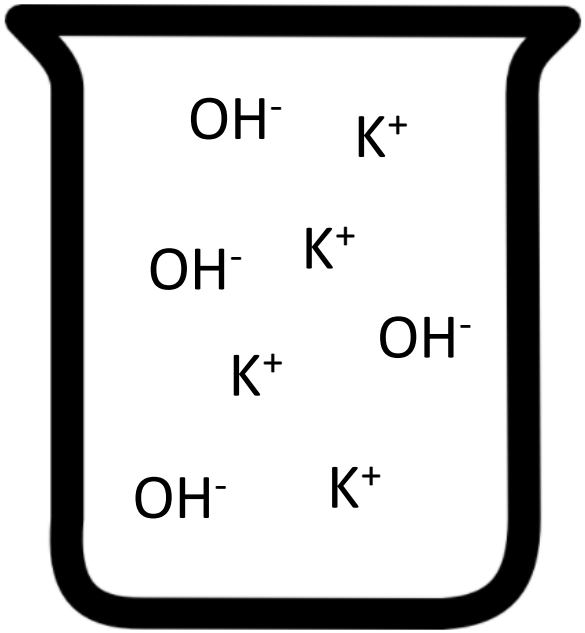
Sodium hydroxide

NaOH



Potassium hydroxide

KOH



Explain the difference between acids and alkalis

Key definition:

An acid is a substance that contains **H+ ions** in an aqueous solution.
The three common acids are:
Hydrochloric acid (HCl)
Sulphuric acid (H_2SO_4)
Nitric acid (HNO_3)

Key definition:

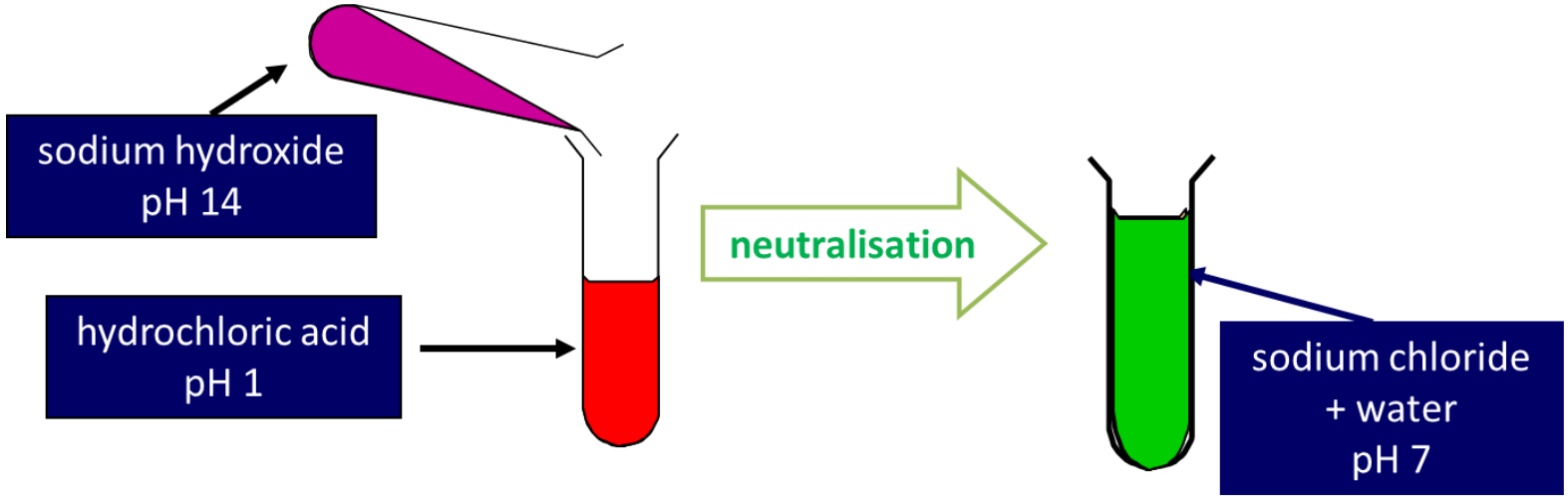
An alkali is a substance that contains **OH- ions** in an aqueous solution.
Some examples are:
Sodium hydroxide (NaOH)
Potassium hydroxide (KOH)

Describe what happens in a neutralisation reaction.

What is neutralisation?

Key definition*:*

***Neutralisation** is when an **acid** reacts with a **base**, to form a **neutral** solution. A base that can dissolve in water is called an alkali.*



Challenge – Write a word equation for the reaction above.

What is neutralisation?

There are three types of compounds that are bases:

1. Metal oxide (e.g. MgO)
2. Metal hydroxide (e.g. $Mg(OH)_2$)
3. Metal carbonates (e.g. $MgCO_3$)

You need to know how to write word equations for the reactions of each of these with acids.

Neutralisation can be represented by the **ionic equation:**

$H^+ (aq) + OH^- (aq) \rightarrow H_2O (aq)$

LEARN THESE!!!

metal + water → metal hydroxide + hydrogen

metal + acid → salt + hydrogen

NEUTRALISATION:

metal oxide/ hydroxide + acid → salt + water

metal carbonate + acid → salt + water + carbon dioxide

Construct word equations for neutralisation reactions.

Neutralisation word equations

1. Metal oxide (e.g. MgO)
2. Metal hydroxide (e.g. $Mg(OH)_2$)

General word equation:

acid + metal oxide/metal hydroxide → salt + water

3. Metal carbonates (e.g. $MgCO_3$)

General word equation:

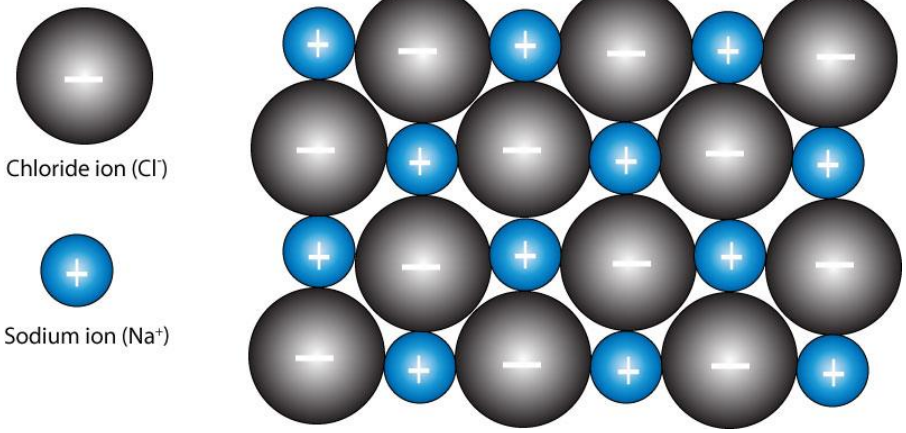
acid + metal carbonate → salt + water + carbon dioxide

Neutralisation word equations

Recap!

In chemistry, a salt is a **compound made up of two ions** (charged particles).

e.g. sodium chloride = table salt!

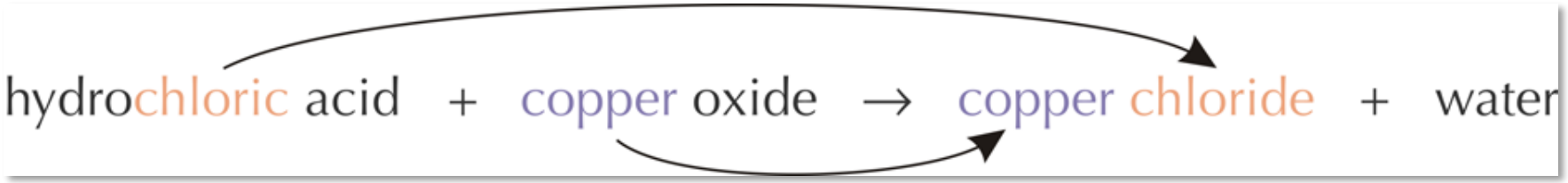


hydrochloric acid + copper oxide → copper chloride + water

ACID **BASE** **SALT** **WATER**

Construct word equations for neutralisation reactions.

How do we name the salts formed in neutralisation reactions?



ACID **BASE** **SALT** **WATER**

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid = nitrate
Sulphuric acid = sulphate
Hydrochloric acid = chloride

Construct word equations for neutralisation reactions.

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid = nitrate
Sulphuric acid = sulphate
Hydrochloric acid = chloride

magnesium oxide + hydrochloric acid → magnesium chloride + water

BASE **ACID** **SALT** **WATER**

Construct word equations for neutralisation reactions.

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid = nitrate
Sulphuric acid = sulphate
Hydrochloric acid = chloride

sodium hydroxide + nitric acid → sodium nitrate + water

BASE **ACID** **SALT** **WATER**

Construct word equations for neutralisation reactions.

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid = nitrate
Sulphuric acid = sulphate
Hydrochloric acid = chloride

aluminium oxide + hydrochloric acid → aluminium chloride + water

Construct word equations for neutralisation reactions.

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid = nitrate
Sulphuric acid = sulphate
Hydrochloric acid = chloride

magnesium hydroxide + sulphuric acid → magnesium sulphate + water

Construct word equations for neutralisation reactions.

The first part of the salt comes from the metal in the base.

The second part of the salt comes from the acid used:
Nitric acid (HNO_3) = nitrate
Sulphuric acid (H_2SO_4) = sulphate
Hydrochloric acid (HCl) = chloride

calcium carbonate + sulphuric acid → calcium sulphate + water + carbon dioxide

BASE **ACID**

Construct word equations for neutralisation reactions.

Task: Complete the word equations!

Copper oxide (s) + nitric acid (aq) → copper nitrate (s) + ……

Iron hydroxide (s) + hydrochloric acid (aq) → ……+ ……

Calcium carbonate (s) + …… → calcium nitrate (s) + hydrogen (?) + ……

…… + hydrochloric acid (aq) → magnesium chloride (s) + ……

…… + …… → lead sulphate (s) + ……

Zinc carbonate (?) + …… → zinc sulphate (s) + …… + ……

Challenge: *Write the symbol equations!*

Construct word equations for neutralisation reactions.

Self-assessment:

Copper oxide (s) + nitric acid (aq) → copper nitrate (s) + Hydrogen (g)

Iron hydroxide (s) + hydrochloric acid (aq) → Iron chloride (s) + Hydrogen (g)

Calcium carbonate (s) + Nitric acid (aq) → calcium nitrate (s) + hydrogen (g) + Carbon dioxide (g)

Magnesium oxide/ hydroxide (s) + hydrochloric acid (aq) → magnesium chloride (s) + Hydrogen (g)

Lead oxide/ hydroxide (s) + Sulfuric acid (aq) → lead sulphate (s) + Hydrogen (g)

Zinc carbonate (s) + Sulfuric acid (aq) → zinc sulphate (s) + Water (l) + Carbon dioxide (g)

Exam practice

This question is about acids and bases.

(a) Which ion is found in all acids? Tick **one** box.

Cl^- ☐ H^+ ☐ Na^+ ☐ OH^- ☐

(1)

(b) Zinc nitrate can be produced by reacting an acid and a metal oxide.

Name the acid and the metal oxide used to produce zinc nitrate.

Acid _____

Metal oxide _____

(2)

(c) In an equation, zinc nitrate is written as $Zn(NO_3)_2(aq)$. What does (aq) mean?

Dissolved in water ☐

Insoluble ☐

Not all reacted ☐

Reactant ☐

(1)

(d) The pH of a solution is 8. Some hydrochloric acid is added to the solution.

Suggest the pH of the solution after mixing.

pH = _____

(1)

Exam practice

The reaction between hydrochloric acid and sodium hydroxide is a neutralisation reaction.

The reaction produces a salt and one other product.

(d) Complete the word equation for the reaction.

hydrochloric acid + sodium hydroxide → _____ + _____

(2)

(e) Universal indicator is used to measure the pH of solutions.

Hydrochloric acid is pH 1

Sodium hydroxide is pH 13

Draw **one** line from the pH to the colour of universal indicator in a solution with that pH.

pH	Colour of universal indicator
	green
1	orange
	purple
13	red
	yellow

(2)

PRINT

C4 – Reactivity of acids

What is a displacement reaction?	Complete the word equation: Magnesium + iron oxide →	What is a metal ore?
Which has been oxidised and which has been reduced? carbon + zinc oxide → carbon dioxide + zinc	Write the half equation for the reaction below: Mg + ZnO → Zn + MgO	What is oxidation and reduction in terms of electrons?

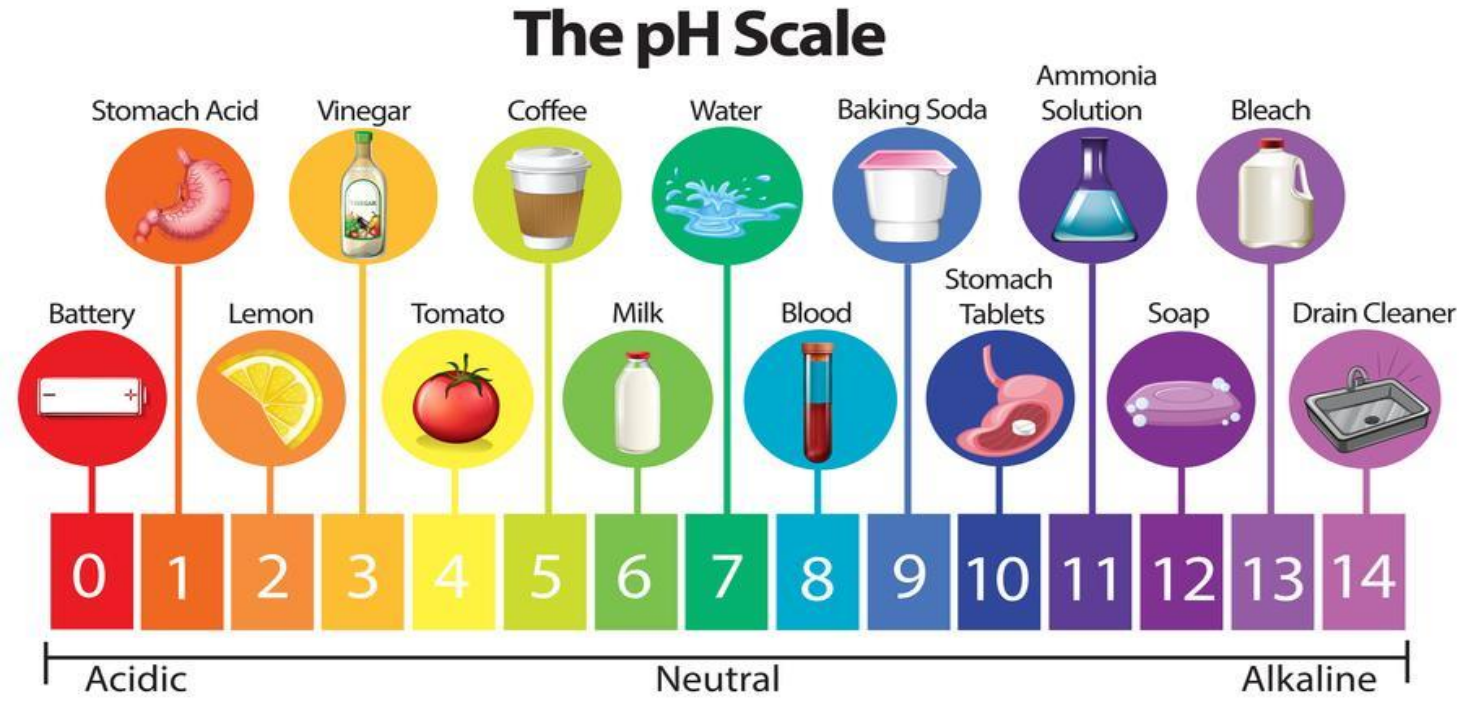
Challenge – *What is the concentration of a 300cm^3 solution which contains 6 grams of hydrochloric acid?*

- pH **1-6**
- Go **red/orange** in universal indicator
- Usually taste **sour**

- pH **7 ONLY**
- Goes **green** universal indicator

- pH **8-14**
- Go **blue/purple** in universal indicator
- Usually '**soapy**' to touch

ACID **NEUTRAL** **ALKALI**



Key definition:

An acid is a substance that contains **H+ ions** in an aqueous solution.
The three common acids are:
Hydrochloric acid (HCl)
Sulphuric acid (H_2SO_4)
Nitric acid (HNO_3)

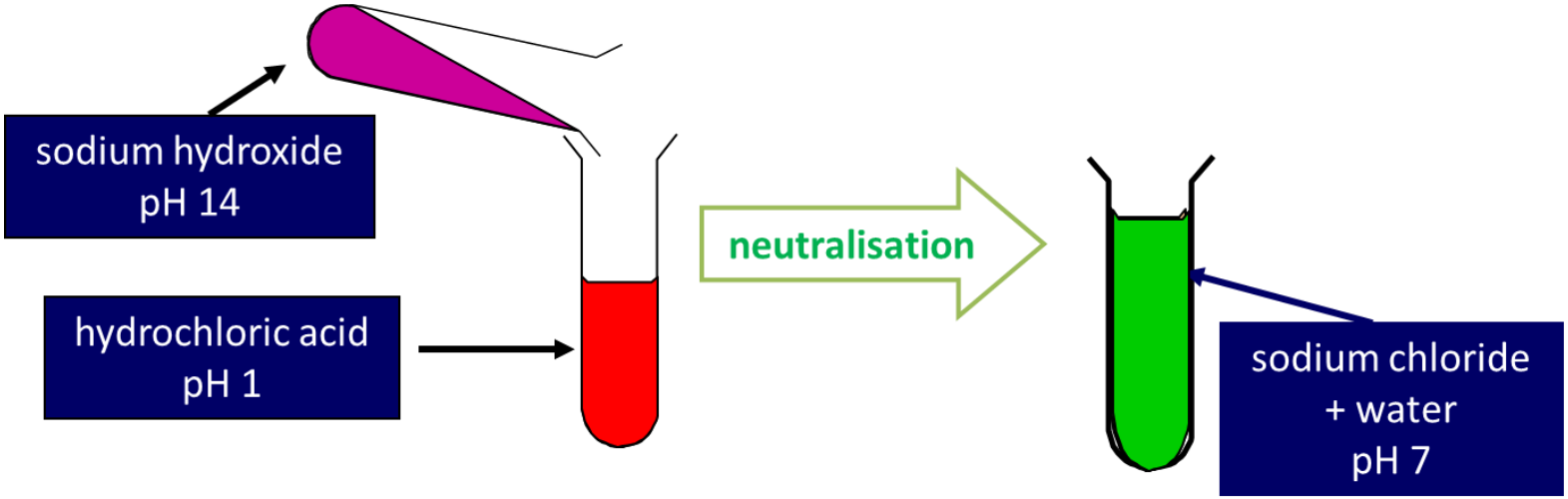
Key definition:

An alkali is a substance that contains **OH- ions** in an aqueous solution.
Some examples are:
Sodium hydroxide (NaOH)
Potassium hydroxide (KOH)

What is neutralisation?

Key definition*:*

***Neutralisation** is when an **acid** reacts with a **base**, to form a **neutral** solution. A base that can dissolve in water is called an alkali.*



Challenge – Write a word equation for the reaction above.

What is neutralisation?

There are three types of compounds that are bases:

1. Metal oxide (e.g. MgO)
2. Metal hydroxide (e.g. $Mg(OH)_2$)
3. Metal carbonates (e.g. $MgCO_3$)

You need to know how to write word equations for the reactions of each of these with acids.

Neutralisation can be represented by the **ionic equation:**

$H^+ (aq) + OH^- (aq) \rightarrow H_2O (aq)$

LEARN THESE!!!

metal + water → metal hydroxide + hydrogen

metal + acid → salt + hydrogen

NEUTRALISATION:

metal oxide/ hydroxide + acid → salt + water

metal carbonate + acid → salt + water + carbon dioxide

Construct word equations for neutralisation reactions.

Neutralisation word equations

1. Metal oxide (e.g. MgO)
2. Metal hydroxide (e.g. $Mg(OH)_2$)

General word equation:

acid + metal oxide/metal hydroxide → salt + water

3. Metal carbonates (e.g. $MgCO_3$)

General word equation:

acid + metal carbonate → salt + water + carbon dioxide

Construct word equations for neutralisation reactions.

Task: Complete the word equations!

Copper oxide (s) + nitric acid (aq) → copper nitrate (s) + ……

Iron hydroxide (s) + hydrochloric acid (aq) → ……+ ……

Calcium carbonate (s) + …… → calcium nitrate (s) + hydrogen (?) + ……

…… + hydrochloric acid (aq) → magnesium chloride (s) + ……

…… + …… → lead sulphate (s) + ……

Zinc carbonate (?) + …… → zinc sulphate (s) + …… + ……

Challenge: *Write the symbol equations!*

Exam practice

This question is about acids and bases.

(a) Which ion is found in all acids? Tick **one** box.

Cl^- ☐ H^+ ☐ Na^+ ☐ OH^- ☐

(1)

(b) Zinc nitrate can be produced by reacting an acid and a metal oxide.

Name the acid and the metal oxide used to produce zinc nitrate.

Acid _____

Metal oxide _____

(2)

(c) In an equation, zinc nitrate is written as $Zn(NO_3)_2(aq)$. What does (aq) mean?

Dissolved in water ☐

Insoluble ☐

Not all reacted ☐

Reactant ☐

(1)

(d) The pH of a solution is 8. Some hydrochloric acid is added to the solution.

Suggest the pH of the solution after mixing.

pH = _____

(1)

Exam practice

The reaction between hydrochloric acid and sodium hydroxide is a neutralisation reaction.

The reaction produces a salt and one other product.

(d) Complete the word equation for the reaction.

hydrochloric acid + sodium hydroxide → _____ + _____

(2)

(e) Universal indicator is used to measure the pH of solutions.

Hydrochloric acid is pH 1

Sodium hydroxide is pH 13

Draw **one** line from the pH to the colour of universal indicator in a solution with that pH.

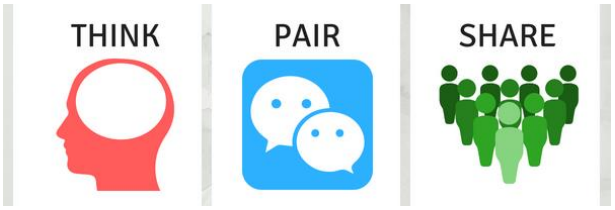
pH	Colour of universal indicator
	green
1	orange
	purple
13	red
	yellow

(2)

C4 – Soluble salts RP

What is a neutralisation reaction? When an acid reacts with an alkali to form a neutral salt and water.	Complete the word equation: magnesium oxide + sulphuric acid → magnesium oxide + sulphuric acid → magnesium sulphate + water	What particle is found in all acids? What particle is found in all alkalis? All acids contain H+, all alkalis contain OH-.
What is the ionic equation for a neutralisation reaction. $H^+ + OH^- \rightarrow H_2O$	Complete the word equation: lithium carbonate + hydrochloric acid → lithium carbonate + hydrochloric acid → lithium chloride + water + carbon dioxide	What does the word soluble mean? If a substance is soluble, it can dissolve in water.

Challenge – *Write balanced chemical equations for the reactions above!*



How can you make **copper sulphate** crystals?

Required practical:
"Making soluble salts"

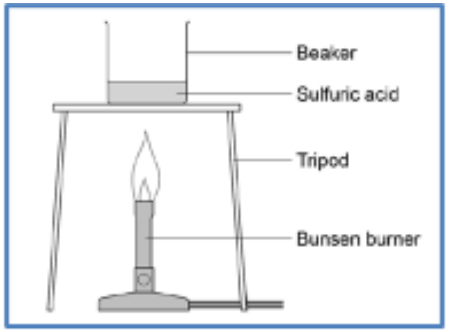
copper oxide + ACID? → copper sulphate + OTHER PRODUCT?

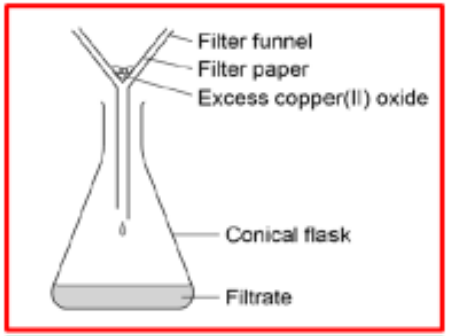
$CuO + H_2SO_4 \rightarrow \mathbf{CuSO_4} + \mathbf{H_2O}$

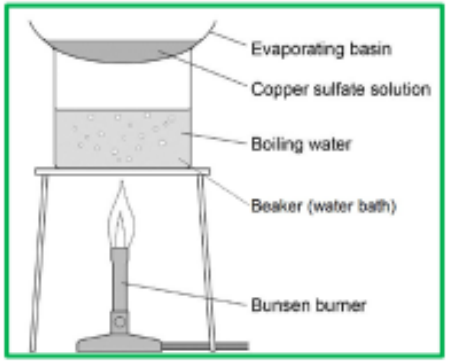
Required practical: *"Making soluble salts"*

https://www.youtube.com/watch?v=qIOMlwBoe_4

Task: Watch the video of the practical and complete the method sheet.





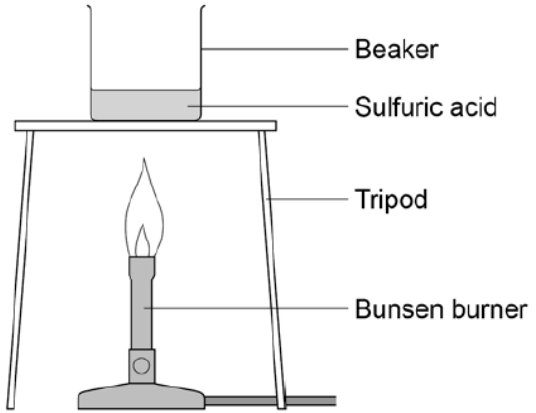


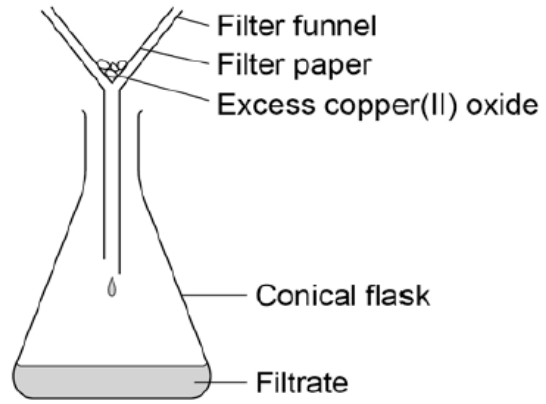
Making soluble salts method:

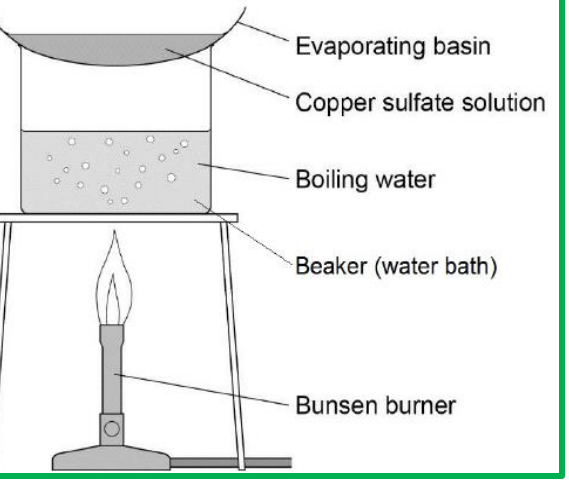
Overall reaction:

→

REMEMBER: This method is for the formation of **COPPER SULPHATE** crystals – you would need different reactants for other crystals!







Heat 40cm^3 of **sulphuric acid** until nearly boiling.

↓

Add **excess copper oxide** and stir, until no more reacts. The solution will turn blue.

↓

Once cool, **filter** out the **unreacted copper oxide**. You will be left with a pure blue copper sulphate solution.

↓

Place the **copper sulphate solution** in an evaporating dish and heat gently until **half** of the solution remains.

↓

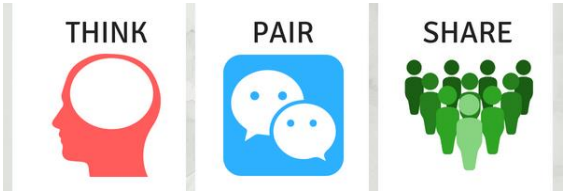
Leave the evaporating dish in a cool place until **crystals are formed**.

↓

Gently **pat dry** the crystals using filter paper.

Overall reaction:

copper oxide + sulphuric acid → copper sulphate + water



Key questions…

1. Why does the metal oxide have to be added IN EXCESS?

The metal oxide must be added in excess to make sure that the **reaction is complete** (and that there is no sulphuric acid left over).

2. What safety precautions must you consider?

- Wear safety glasses
- Take care when handling the acid (as it is corrosive)
- Make sure the sulphuric acid does not boil
- Do not touch the boiling beaker

Challenge:

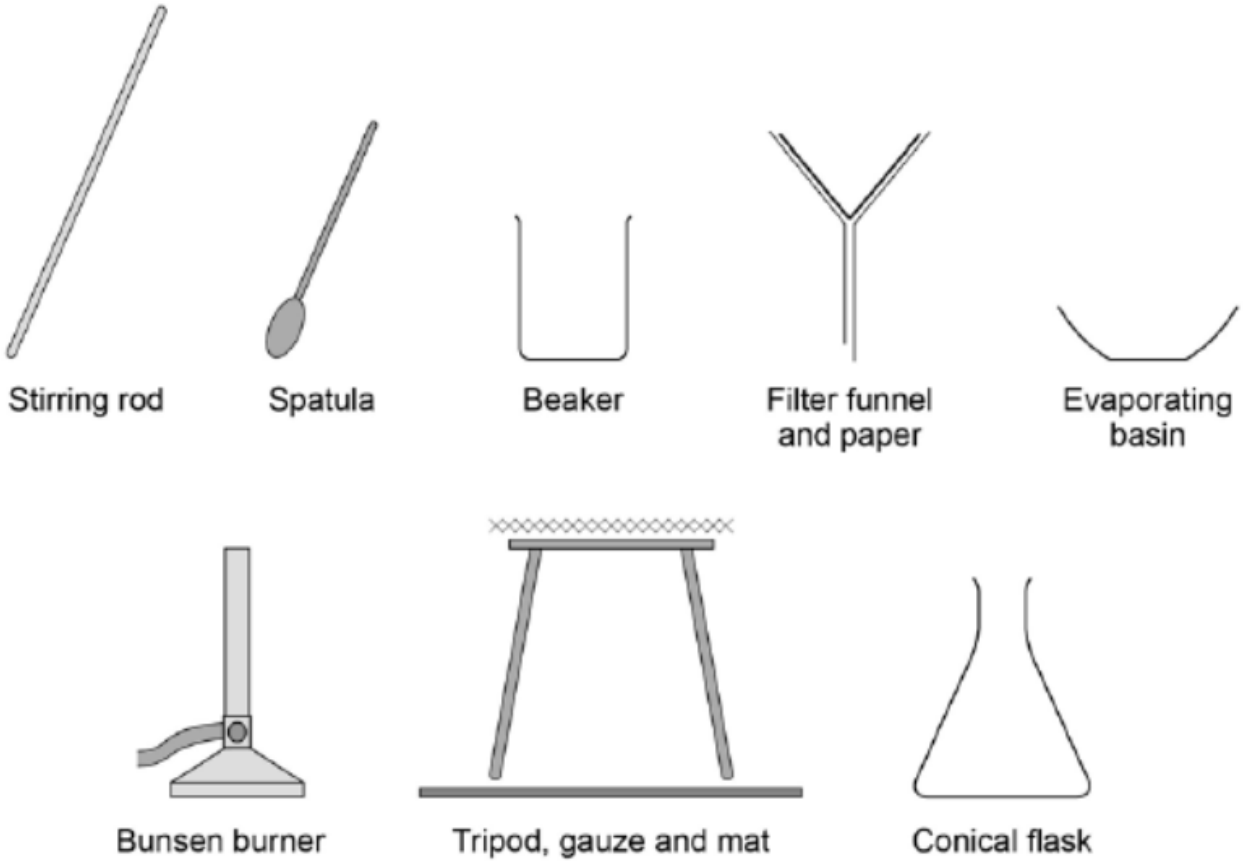
What effect would the rate of evaporation have on the size of the crystals?

Exam practice

This question is about making copper salts.

(Total 6 marks)

The figure below shows the apparatus given to a student.



Outline a safe plan the student could use to make pure, dry, crystals of the soluble salt copper sulfate from the insoluble metal oxide and dilute acid.

What is this asking us to do?

Safety precautions?

Outline a safe plan that the student could use to make pure, dry, crystals of the soluble salt copper sulphate from an insoluble metal oxide and dilute acid.

*How do we make sure the product is **pure** and **dry?***

What reactants do we need to use?

Self assessment

Level 3 (5–6 marks):
A coherent method is described with relevant detail, which demonstrates a broad understanding of the relevant scientific techniques, procedures and safety precautions. The steps in the method are logically ordered with the dependent and control variables correctly identified. The method would lead to the production of valid results.

Level 2 (3–4 marks):
The bulk of a method is described with mostly relevant detail, which demonstrates a reasonable understanding of the relevant scientific techniques, procedures and safety precautions. The method may not be in a completely logical sequence and may be missing some detail.

Level 1 (1–2 marks):
Simple statements are made which demonstrate some understanding of some of the relevant scientific techniques, procedures and safety precautions. The response may lack a logical structure and would not lead to the production of valid results.

0 marks:
No relevant content

Indicative content

Named chemicals

- copper oxide
- sulfuric acid
- copper sulfate

Correct use of apparatus

- stirring rod
- spatula
- beaker
- filter funnel and filter paper
- evaporating basin
- Bunsen burner
- tripod and gauze
- bench mat
- conical flask

Method

- add (excess) copper oxide to sulfuric acid
- heat the mixture
- filter the mixture
- method to evaporate some of the water from the filtrate eg using a water bath or evaporating to half volume
- leave solution (to cool and) to form crystals
- remove and dry crystals

Safety

- wearing of safety glasses / goggles
- care with use of sulfuric acid as corrosive
- warming not boiling mixture of copper oxide and sulfuric acid
- hold beaker containing warm mixture with tongs whilst filtering

4.4.2.3 Soluble salts

Content
Soluble salts can be made from acids by reacting them with solid insoluble substances, such as metals, metal oxides, hydroxides or carbonates. The solid is added to the acid until no more reacts and the excess solid is filtered off to produce a solution of the salt. Salt solutions can be crystallised to produce solid salts. Students should be able to describe how to make pure, dry samples of named soluble salts from information provided.

RED	AMBER	GREEN

Required practical 1: preparation of a pure, dry sample of a soluble salt from an insoluble oxide or carbonate using a Bunsen burner to heat dilute acid and a water bath or electric heater to evaporate the solution.

C4 – Electrolysis

A solution of hydrochloric acid contains 3.2 g of hydrogen chloride in 50 cm^3. Calculate the concentration

$50cm^3 = 0.05dm^3$
Concentration
$= 3.2 \div 0.05$
$= 64g/dm^3$

magnesium + sulphuric acid -> magnesium sulphate + hydrogen

lithium oxide + hydrochloric acid -> lithium chloride + water

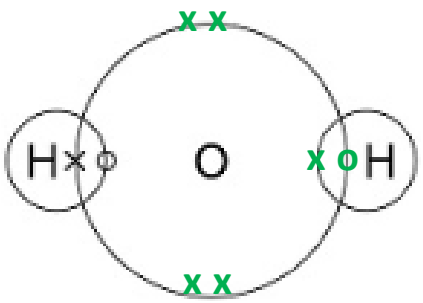
zinc carbonate + nitric acid -> zinc nitrate + water + carbon dioxide

magnesium + iron oxide -> magnesium oxide + iron

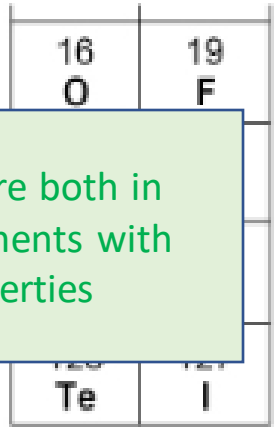
A student reacted zinc oxide powder with hydrochloric acid to produce zinc chloride solution. Complete the equation for the reaction by balancing the equation AND writing the state symbols

$ZnO\ (s) + 2\ HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2O\ (l)$

Complete the dot and cross diagram for water:



Why did Mendeleev swap the positions of te



So that they were both in groups with elements with similar properties

Why is it important to add copper oxide in excess when making copper sulphate?

To ensure that all of the sulphuric acid reacts.

Challenge – Calculate the mass of $CaSO_3$ produced when 7.00 g of calcium oxide reacts with an excess of sulfur dioxide:

$CaO + SO_2 \longrightarrow CaSO_3$

C4 – Electrolysis

Challenge – Calculate the mass of $CaSO_3$ produced when 7.00 g of calcium oxide reacts with an excess of sulfur dioxide: $CaO + SO_2 \longrightarrow CaSO_3$

$(M_r\ CaO =)\ 56$

$(M_r\ CaSO_3 =)\ 120$

$$\frac{7}{56} \times 120$$

$= 15(.0\ g)$

DECODE IT NOW

Word:

Electrolysis
(tier 3)

Define it:

Process of using electricity to separate ions in molten or aqueous ionic compounds.

Write a sentence of your own that uses the word **electrolysis.**

Write your own definition of the word **electrolysis.**

Which subjects or topics will this word be relevant to?

Digging Deeper:

Electrolysis is a method that is used to extract metals from **ores**. Aluminium is extracted from aluminium oxide (found in the ground) to make products like aluminium cans. We use about ***180 billion*** aluminium cans every year worldwide!

Use it:

Metals are extracted from their ores by electrolysis.

Deconstruct it (Root word):

From **electro** meaning 'electricity' and **lysis** meaning 'break down'.

Metals **above** carbon must be extracted from their ores by using **electrolysis.**

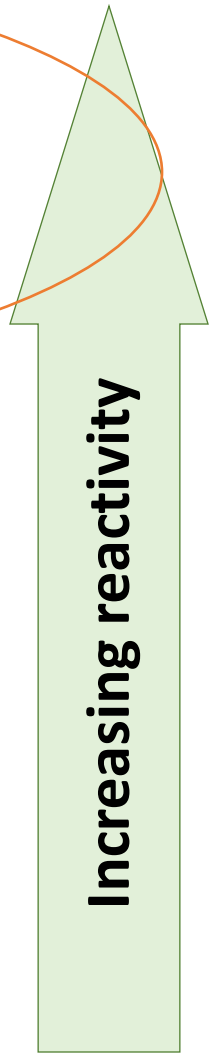
Metals **below** carbon can be extracted from their ores by **reduction** using **carbon.**

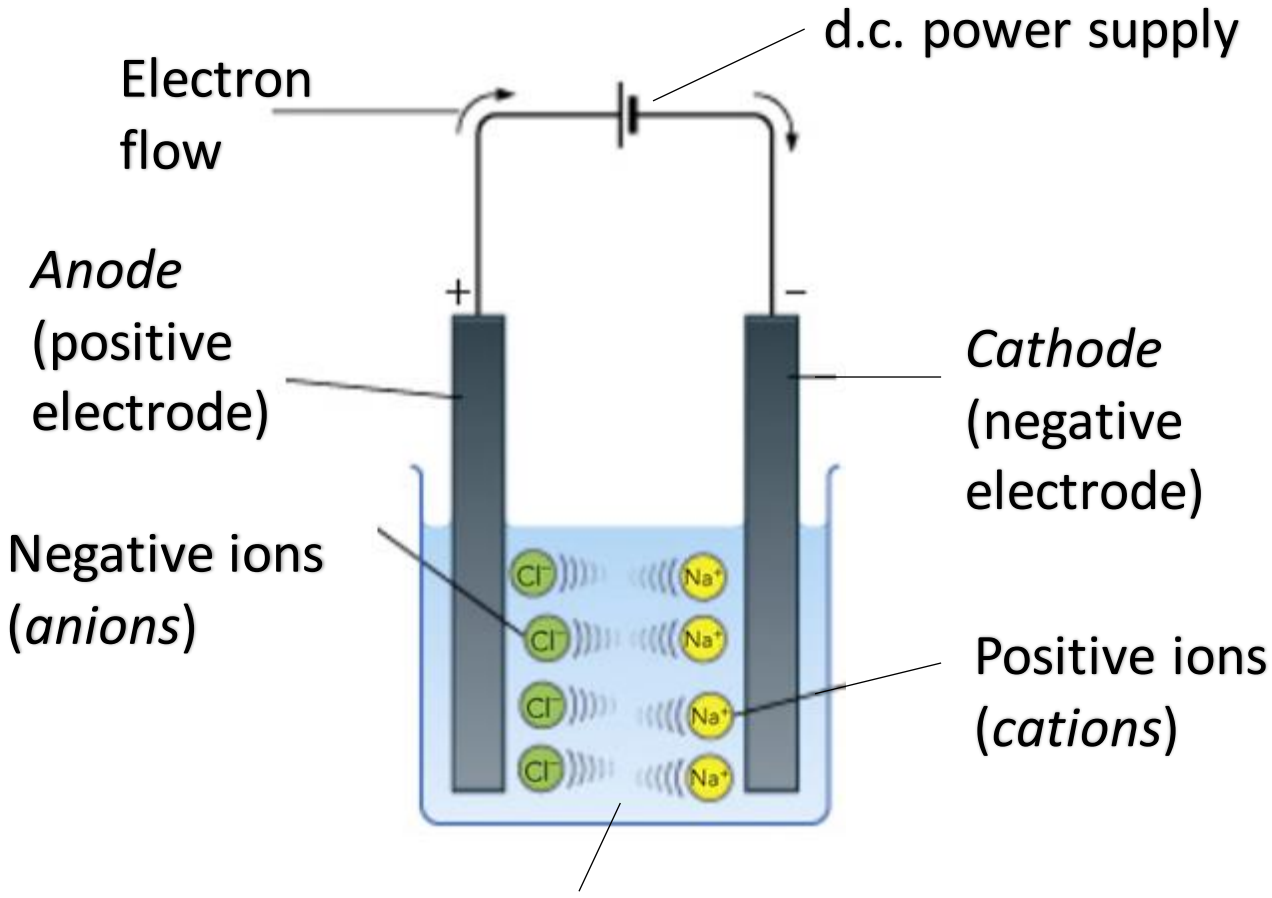
They occur naturally and don't need to be extracted.

Potassium
Sodium
Calcium
Magnesium
Aluminium
CARBON

Zinc
Iron
Lead
HYDROGEN

Copper
Silver
Gold
Platinum





POSITIVE (IS)

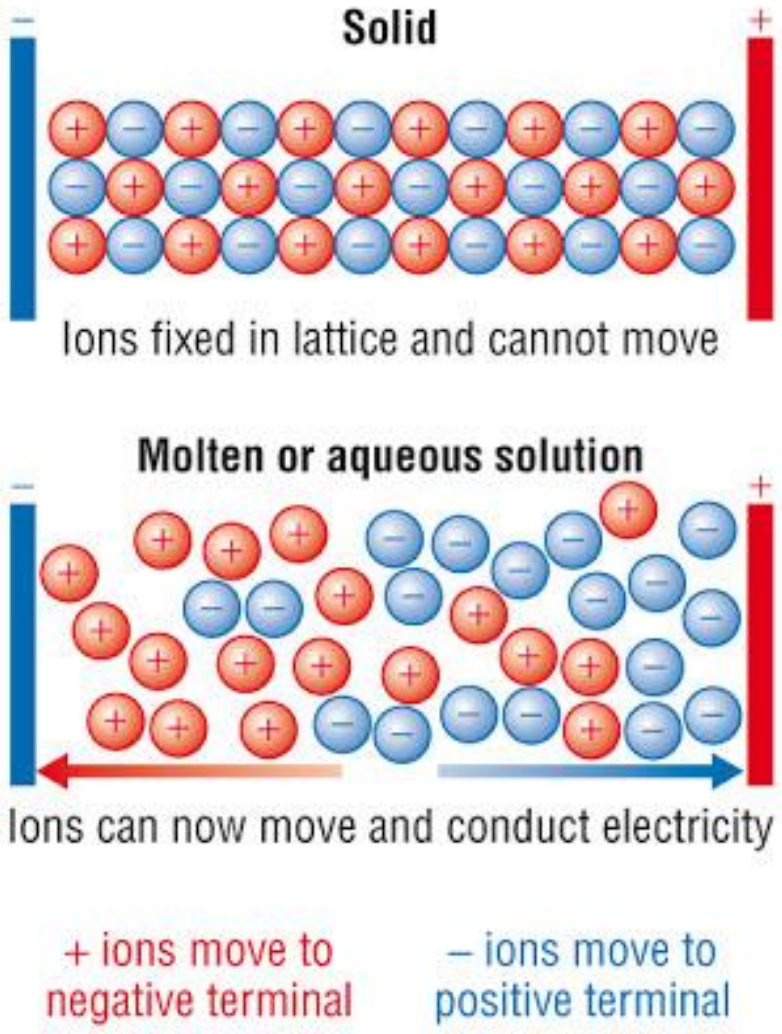
ANODE

NEGATIVE

IS

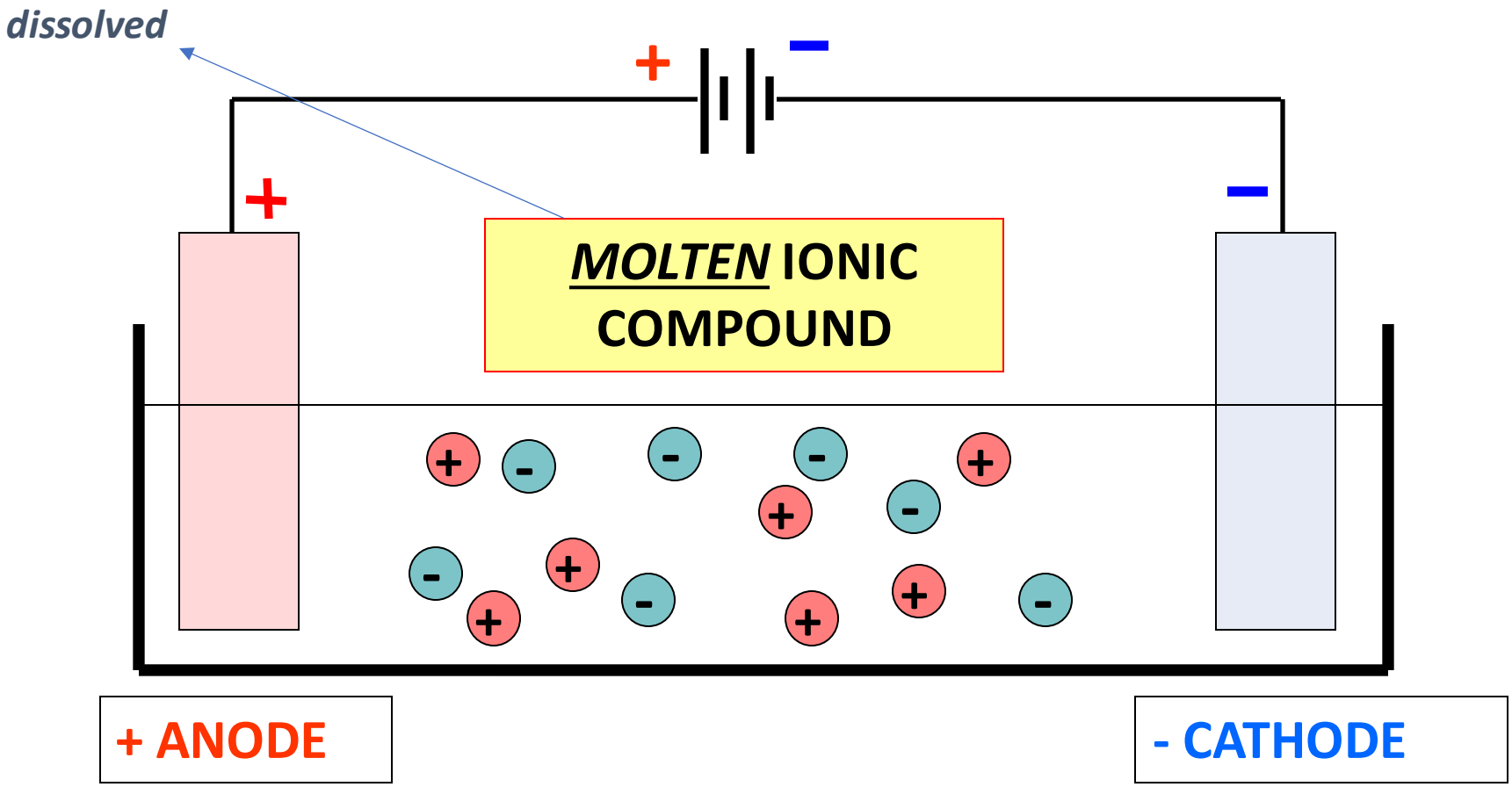
CATHODE

Electrode = an inert rod that conducts electricity (often made of graphite or platinum)



For electrolysis to occur, the **ionic compound must be molten (melted) or aqueous (in solution).**

This is so that the **ions** (*NOT ELECTRONS*) are **able to move around so that they can conduct electricity**.



When the battery is switched on…
Cations (**metal** ions) are attracted to the negative cathode.
Anions (**non-metal** ions) are attracted to the positive anode.

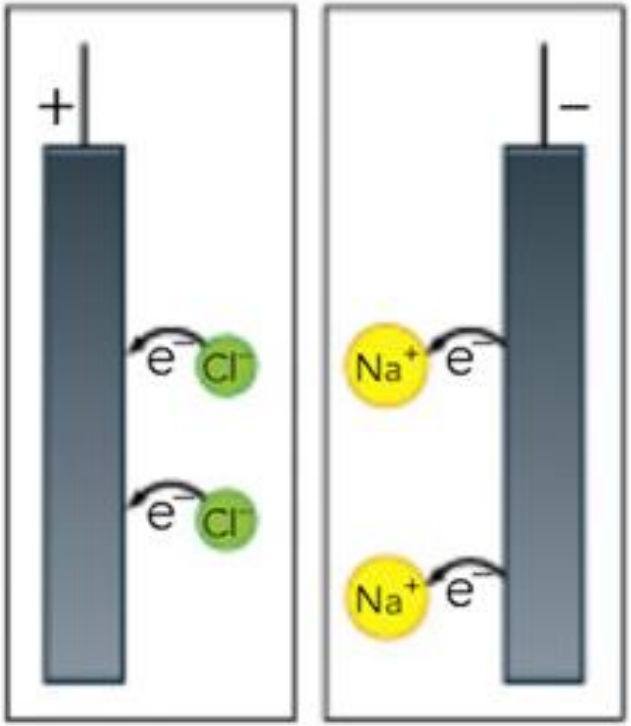
HIGHER ONLY!

At the electrodes, ions transfer electrons so **atoms or molecules are formed at the electrodes.**

e.g. Cl- forms pure chlorine gas (Cl_2)

Remember, all halogens are diatomic! i.e. F_2, Cl_2, I_2, Br_2

e.g. Na+ forms pure sodium (Na)



At the **positive** anode, the anions **lose** electrons

= OXIDATION takes place at the Anode

At the **negative** cathode, the cations **gain** electrons

= REDUCTION takes place at the Cathode

HIGHER ONLY!

OXIDATION
IS
LOSS OF ELECTRONS
REDUCTION
IS
GAIN OF ELECTRONS

Copper and chlorine are produced when molten copper chloride is electrolysed.

Complete the half equation for the reaction at each electrode.

Half equation at negative electrode

Cu^{2+} _____ → _____

Half equation at positive electrode

$2\ Cl^{-}$ → _____

RULES: Cations (metals) will ALWAYS be attracted to the cathode.
Anions (non-metals) will ALWAYS be attracted to the anode.
When the ions, get to the electrodes, they become atoms (lose their charge).

Salt	***Cation?*** (positive ion)	***Anion?*** (negative ion)	***Product formed at the anode?***	***Product formed at the cathode?***
Potassium Bromide				
Aluminium Oxide				
Barium Iodide				
Sodium Chloride				
Calcium Fluoride				

RULES: Cations will ALWAYS be attracted to the cathode.
Anions will ALWAYS be attracted to the anode.
When the ions, get to the electrodes, they become atoms (lose their charge).

Salt	***Cation?*** (positive ion)	***Anion?*** (negative ion)	***Product formed at anode?***	***Product formed at cathode?***
Potassium Bromide	K^+	Br-	Bromine (Br_2)	Potassium (K)
Sodium Chloride	Na^{+1}	Cl^{-1}	Chlorine ($Cl_{2)}$	Sodium (Na)
Aluminium Oxide	Al^{3+}	O^{2-}	Oxygen (O_2)	Aluminium (Al)
Barium Iodide	Ba^{2+}	I^-	Iodine (I_2)	Barium (Ba)
Calcium Fluoride	Ca^{+2}	F^{-1}	Fluorine ($F_{2)}$	Calcium (Ca)

Model exam question:

Metal and non metal bonded together

(a) Zinc chloride is an ionic substance.
Complete the sentence.

When zinc chloride is molten, it will conduct Electricity .

(1)

Liquid, ions are free to move.

(b) Zinc ions move towards the negative electrode where they gain electrons to produce zinc.

(i) Name the product formed at the positive electrode.

Chlorine

(1)

Only one other element left in molten zinc chloride!

(ii) Explain why zinc ions move towards the negative electrode.

Why do metal ions move towards the negative electrode?

Zinc forms positively charged ions.

Opposites attract so it moves towards the negative electrode.

What charge do metal ions have?

(2)

Model exam question:

Solid compound, ions can't move!

Lead bromide was placed in the tube and the circuit was switched on. The light bulb did not light up.

The tube was heated and soon the bulb lit up. The observations are shown in the table.

Positive electrode	Negative electrode
red-brown gas	silver liquid

What products are formed from lead bromide?

(a) What is meant by *electrolysis*?

Using electricity to break down compounds into elements

(2)

(b) Why did the lead bromide conduct electricity when the tube was heated?

The solid compound is melted, so ions could move

(1)

(c) Name the substances formed at the:

positive electrode; Bromine (Br_2)

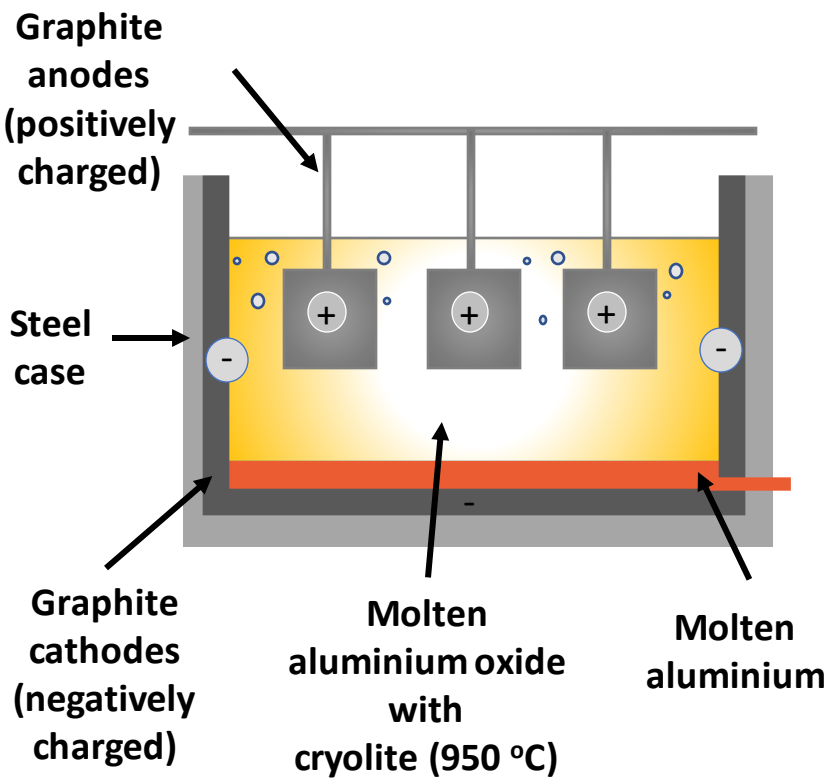
negative electrode. Lead (Pb)

(2)

Electrolysis of molten aluminium oxide

Play video:
https://www.youtube.com/watch?v=mvDHeYI-a00

Aluminium oxide → aluminium + oxygen

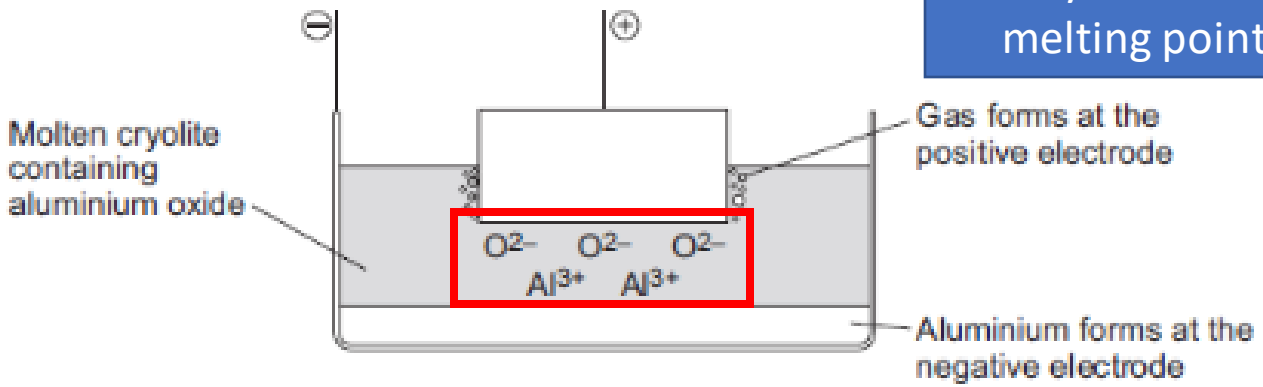


- Aluminium goes to the negative electrode and sinks to bottom. Oxygen forms at positive electrodes.
- Aluminium oxide is mixed with **cryolite** to ***lower the melting point of the mixture*** and save energy.
- **Oxygen reacts with the positive electrode** making carbon dioxide. The ***electrode needs replacing*** due to this reaction: $C + O_2 \rightarrow CO_2$

(c) Aluminium is produced by electrolysis of a molten mixture of aluminium oxide and cryolite.
This is shown in **Figure 3**.

Cryolite reduces the melting point of Al_2O_3

Figure 3



(i) Name a gas produced at the positive electrode.

Must be an anion (-)

Oxygen

(1)

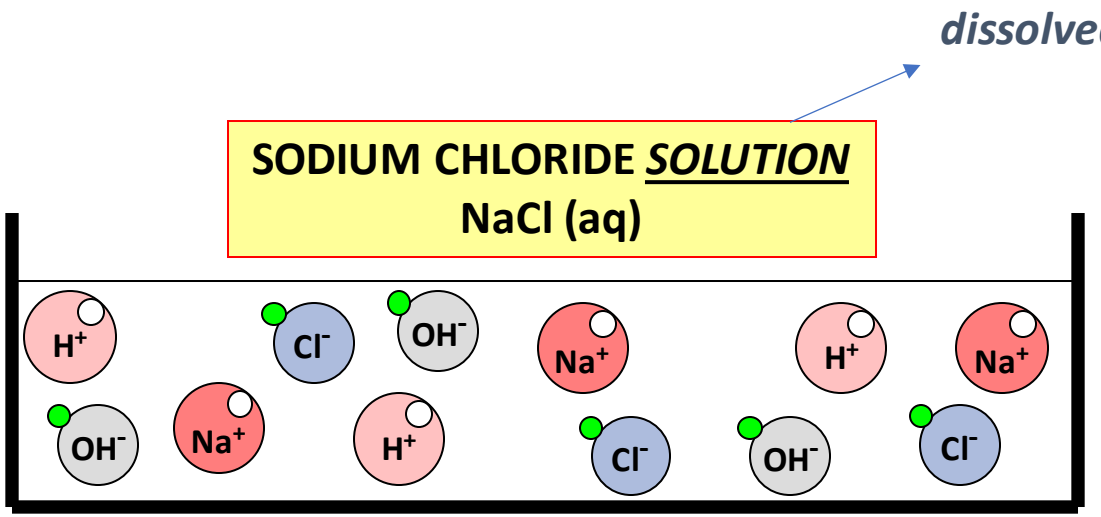
(ii) Aluminium ions move to the negative electrode.

Explain why.

Opposites attract

So positively charged aluminium ions move towards the negatively charged electrode.

(2)

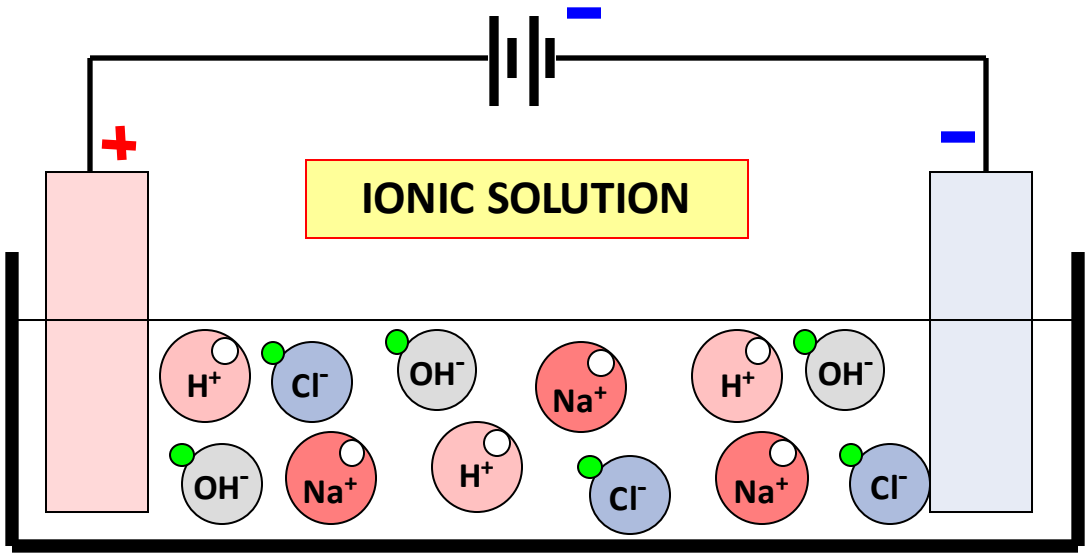


In an ionic **solution** (eg sodium chloride solution), there will be FOUR types of ion present:

- TWO from the ionic compound (**Na^+** + **Cl^-**)

- TWO from the water (**H^+** + **OH^-**)

$$H_2O \rightarrow H^+ + OH^-$$



Cations (positive ions) are attracted to the negative cathode.

Anions (negative ions) are attracted to the positive anode.

The products formed at each electrode depends on reactivity…

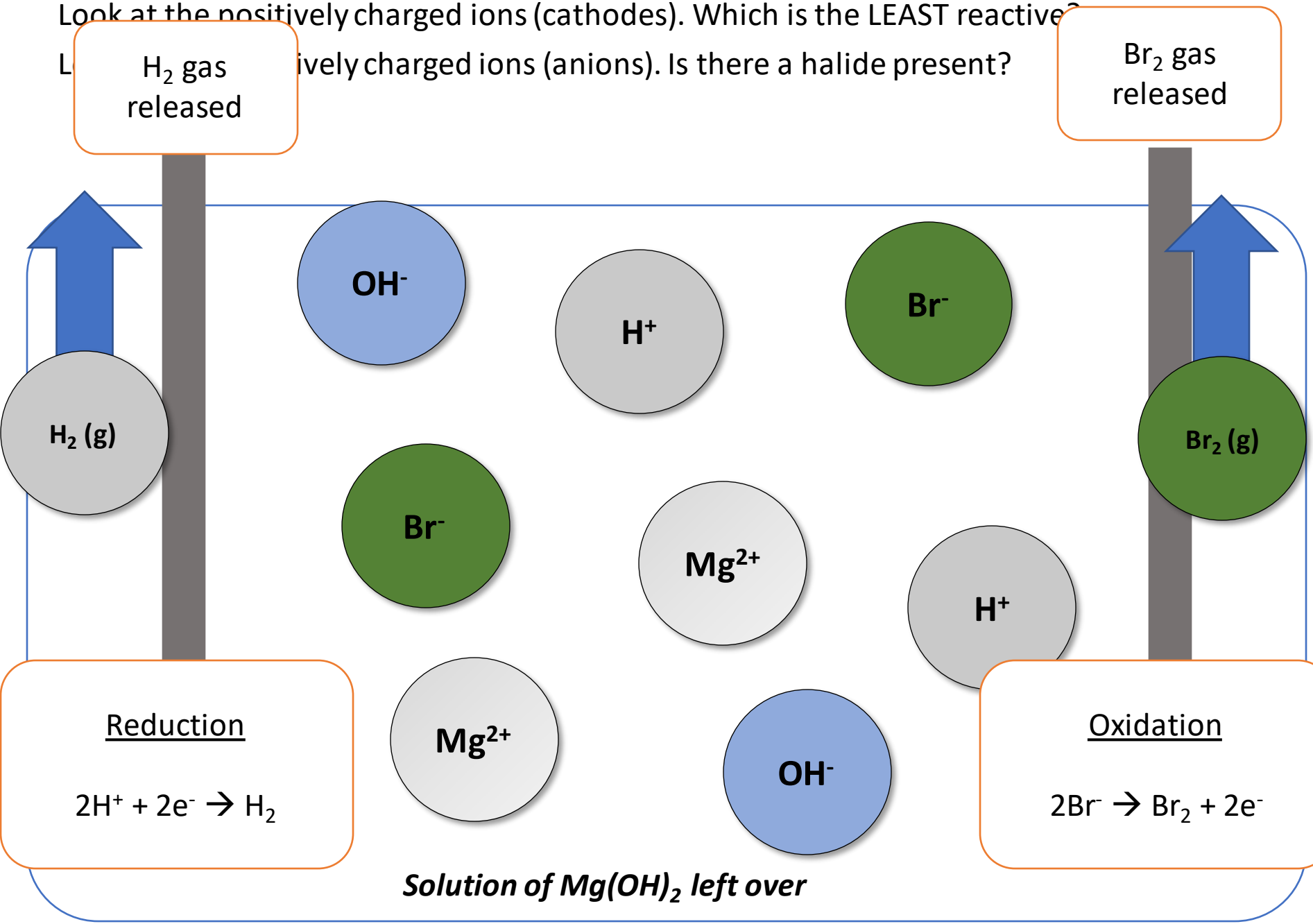
*At the **negative electrode** (anode), **hydrogen** is produced if it is **less reactive** than the metal. If the metal is **gold, copper or silver**, then the metal is produced.*

*At the **positive electrode** (cathode), if **halogens** are present then the **halogen is produced**. If halogens are **not** present, **oxygen** is produced.*

Electrolysis of Magnesium Bromide solution

Look at the positively charged ions (cathodes). Which is the LEAST reactive?

L ively charged ions (anions). Is there a halide present?



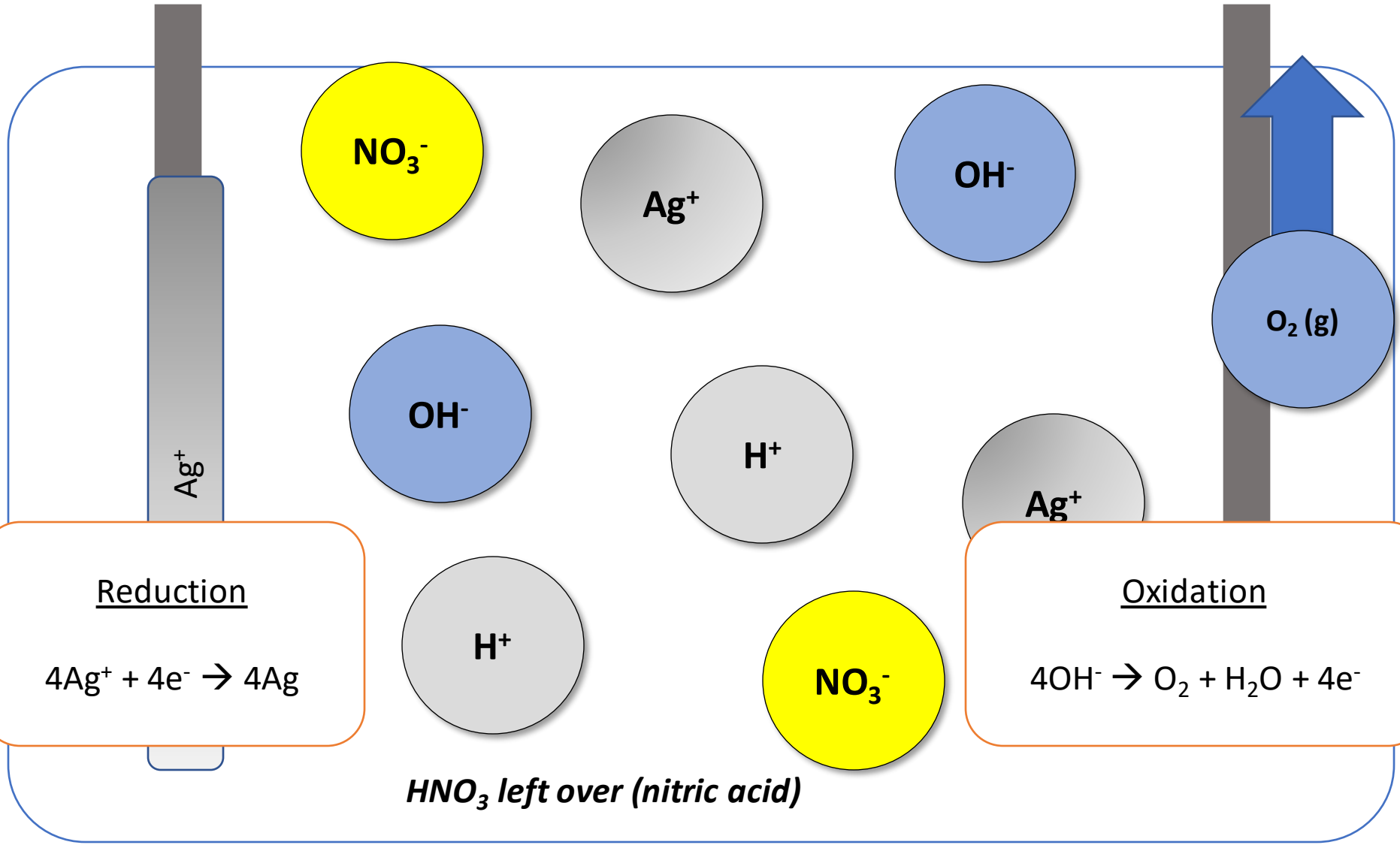
Electrolysis of Silver Nitrate solution

Silver builds up on electrode

L vely charged ions (cathodes). Which is the LEAST reactive

L vely charged ions (anions). Is there a halide present?

O_2 gas released



Reduction

$4Ag^+ + 4e^- \rightarrow 4Ag$

Oxidation

$4OH^- \rightarrow O_2 + H_2O + 4e^-$

HNO_3 left over (nitric acid)

Task: Complete the table to show what product is formed at each electrode

(REACTIVITY: K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{3+} H^+ Cu^{2+} Ag^+ Au^{3+})

Compound	State	Ions	Negative electrode	Positive Electrode
potassium chloride	molten	K^+ Cl^-	potassium, K	chlorine, Cl_2
aluminium oxide	molten			
copper chloride	solution			
sodium bromide	solution			
silver nitrate	solution			
potassium chloride	solution			
zinc sulphate	solution			

(REACTIVITY: K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{3+} H^+ Cu^{2+} Ag^+ Au^{3+})

Compound	State	Ions	Cathode (-)	Anode (+)
potassium chloride	molten	K^+ Cl^-	potassium	chlorine
aluminium oxide	molten	Al^{3+} O^{2-}	aluminium	oxygen
copper chloride	solution	Cu^{2+} Cl^- H^+ OH^-	copper	chlorine
sodium bromide	solution	Na^+ Br^- H^+ OH^-	hydrogen	bromine
silver nitrate	solution	Ag^+ NO^{3-} H^+ OH^-	silver	oxygen
potassium chloride	solution	K^+ Cl^- H^+ OH^-	hydrogen	chlorine
zinc sulphate	solution	Zn^+ SO_4^{2-} H^+ OH^-	hydrogen	oxygen

(REACTIVITY: K^+ Na^+ Ca^{2+} Mg^{2+} Al^{3+} Zn^{2+} Fe^{3+} H^+ Cu^{2+} Ag^+ Au^{3+})

(ii) Sodium chloride solution contains two types of positive ions, hydrogen ions (H^+) and sodium ions (Na^+).

Why is hydrogen produced at the negative electrode and **not** sodium?

Tick (✓) **one** box.

Hydrogen is a gas. ☐

Hydrogen is less reactive than sodium. X

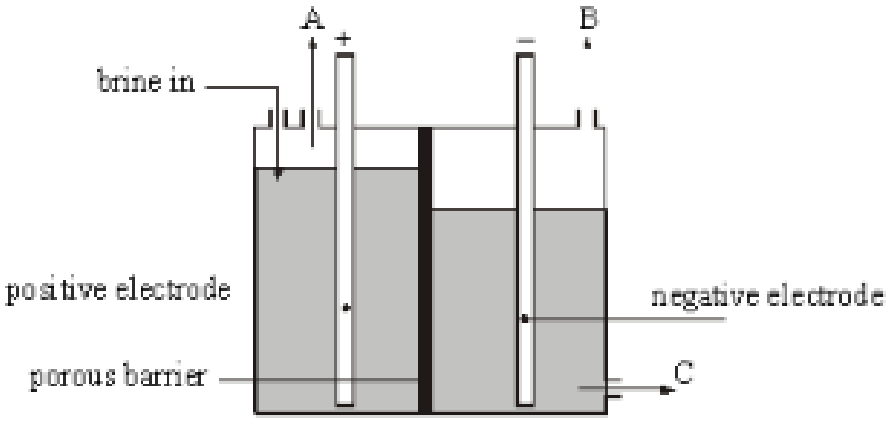
Hydrogen ions move faster than sodium ions. ☐

(1)

Background information

Think about the reactivity series. Why might H_2 be released rather than sodium?

Sodium hydroxide, hydrogen and chlorine can all be made in one industrial process. Electricity is passed through aqueous sodium chloride solution (brine). The diagram below shows a cell that can be used for this process.



(a) Name A, B and C.

Gas A Chlorine

Gas B Hydrogen

Solution C Sodium hydroxide

(b) Balance the equations for the reactions at the electrodes.

(i) 2 Cl^- – 2 e^- → Cl_2

(ii) 2 H^+ + 2 e^- → H_2

(c) Name the compound in this cell which produces the hydrogen ions.

Water

(1)

Step 1: Balance the chlorine ions

Step 2: Figure out how many each electrons each chlorine needs to gain

Step 3: Multiply it by the total number of chlorines we have.

Step 4: Do the same for hydrogen

Hint: Number of electrons in each half equation will always be the same!