CC9: Mass Calculations knowledge organiser (H)

Lesson 1 Relative formula mass / formulas

The relative formula mass RFM (M_r) of a compound is the mass of all of the atoms in the formula. You can calculate this by multiplying the number of atoms of each element by the Relative atomic mass RAM (A_r) and adding the totals.

1	H_2SO_4	2(1) + 32 + 4(16)	= 98
2	Al_2O_3	2(27) + 3(16)	= 102
3	$Mg(OH)_2$	24 + 2(16) + 2(1)	= 58

Relative atomic mass (RAM) = Mass of an atom
Relative formula mass (RFM) = Mass of a formula (add up the masses of atoms in the formula)

Types of formula
The formula of a compound tells you type and number of elements it contains.
Molecular formula = The normal formula showing the ratio of elements in a compound
Empirical formula = The **simplest** whole number ratio of elements in a compound

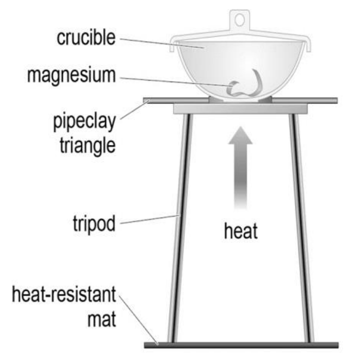
The molecular formula is always a multiple of the empirical formula
e.g. Hydrogen peroxide molecular formula = H_2O_2
empirical formula = HO

Lesson 2 Determining the empirical formula of magnesium oxide

Magnesium burns in oxygen according to the following equation

$$2Mg + O_2 \longrightarrow 2MgO$$

During the reaction the magnesium gains mass as it combines with oxygen particles. The increase in mass is the mass of oxygen. The mass of magnesium at the start is the mass of magnesium.



Mass of Mg at start = 5 g
Mass of MgO at end = 8.3 g

Mass of O added = 3.3 g

From the mass of reactants you can find the empirical formula

Elements	Mg	O
Mass / RAM	5/24	3.3/16
=	0.21	0.21
Simplest whole numbers	1	1

Formula of magnesium oxide = MgO

Lesson 3 Calculation of empirical formula

In questions you will be given the masses of elements in a compound (if you are give % just use this instead of mass)

e.g. 20g of calcium react with 80g of bromine. What is the empirical formula of calcium bromide?

Elements	Ca	Br
Mass / RAM	20 / 40	80 / 80
=	0.5	1
Simplest whole numbers	1	2

Empirical formula of calcium bromide = $CaBr_2$

For empirical formula find the simplest whole number ratio of grams / RAMs !!

If you know the empirical formula and the RFM then you can find the molecular formula

e.g. A compound with empirical formula CH_2 has an RFM of 42. What is the molecular formula?
Mass of empirical formula = 14 42/14 = 3
The molecular formula = C_3H_6

Find the mass of the empirical formula. Divide the RFM by this to get the multiplication factor. Multiply the empirical formula by this to get the molecular formula

Lesson 4 Concentration in g/dm³

The concentration of a solution is the amount of solute (in g) dissolved in 1 dm^3 of solvent

To convert cm^3 to dm^3 DIVIDE by 1000

Where 1 dm^3 = 1000 cm^3 = 1 L

$$\text{Concentration} = \frac{\text{amount of substance (g)}}{\text{volume } (dm^3)}$$

In questions you often have to convert the volume from cm3 to dm3 and then calculate the concentration

e.g. what is the concentration in g/dm^3 of a solution of 5 g of salt in 100cm^3 of water

Concentration = 5 / 0.1 = 50 g/dm^3

Law of conservation of mass

In a chemical reaction particles are not created or destroyed. The mass of all of the particles at the start of the reaction is equal to the mass of all of particles at the end of a reaction

Lesson 5 Moles (HIGHER)

A mole is an amount of substance.
It is the amount of substance that contains the Avogadro number (6.02 x 10^{24}) of particles
One mole of a substance weighs its RFM in grams if is and element and the RFM in grams if it is a compound.

$$\text{moles} = \frac{\text{mass}}{\text{RAM or RFM}}$$

Particles = moles x Avogadro number

You need to be able to convert from mass in grams to moles and from moles to mass in grams.
Once you have found the number of moles, you can find the number of particles by multiplying by the Avogadro number.

Lesson 6 Reacting mass calculations using moles (HIGHER)

Reacting mass calculations involved calculating a mass of something (either a product or reactant) given the mass of other things. The steps involved are :-

- Write a balanced equation for the reaction
- Use the information given to you to calculate the moles of something using moles = $\frac{\text{mass}}{\text{RFM}}$
- Use the balanced equation to find the moles of something else
- Find the mass using the equation mass = moles x RFM

Lesson 7 Stoichiometry and limiting reactant (HIGHER)

The balancing numbers in an equation show the ratio of moles of substances.
These numbers are referred to as the **Stoichiometry** of the equation.
If you know the mass ratios in an equation, you can calculate the Stoichiometry by dividing by the RAM /RFM of each substance.

Stoichiometry = Ratio of moles of each substance in an equation

$$2Mg + O_2 \longrightarrow 2MgO$$

e.g. 2 moles of magnesium react with 1 mole of oxygen to give 2 moles of magnesium oxide
Often in chemical reactions we do not use exact mole ratios but use one of the reactants **in excess**. This reactant is not fully used up in the reaction and the amount of reactant is determined by the other reactant – **the limiting reactant.**

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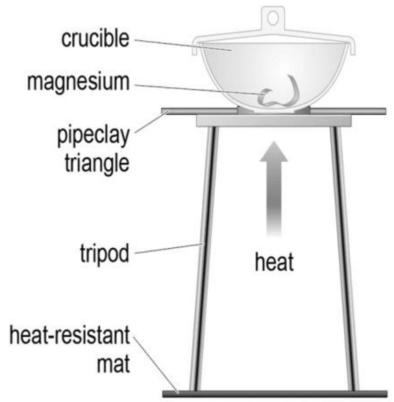

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This assumes the reaction is done in a sealed system

In reality many reactions will lose or gain weight because they on non-sealed system. For example if a gas is given off the mass will decrease.

Lesson 6 Reacting mass calculations

You can find the mass ration in chemical equations by multiplying the RFM of the substance by the balancing number.
You can use mass ratios to calculate how much of a substance reacts with how is produced.

Calculate the reacting mass ratios in the following reaction:

	N_2 +	$3H_2$ →	$2NH_3$
RFM	28	2	17
Balancing number	1	3	2
Mass ratio	28 g +	6 g →	34 g

Mass of reactants 34 g **Mass of products 34 g**

In the reaction 28 g of nitrogen will react with 6 g of hydrogen to produce 34g of ammonia

CC9: Mass Calculations knowledge organiser (C)

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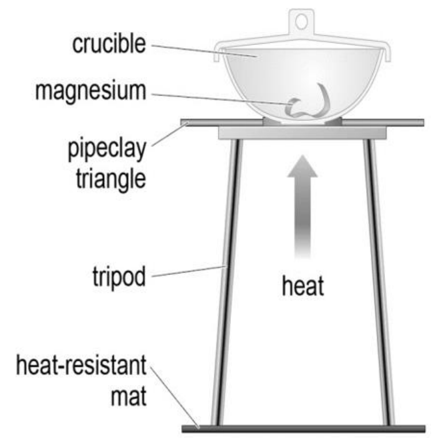

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In reality many reactions will lose or gain weight because they on non-sealed system. For example if a gas is given off the mass will decrease.

CC10: Electrolysis knowledge organiser (H)

Lesson 1 Introduction to electrolysis

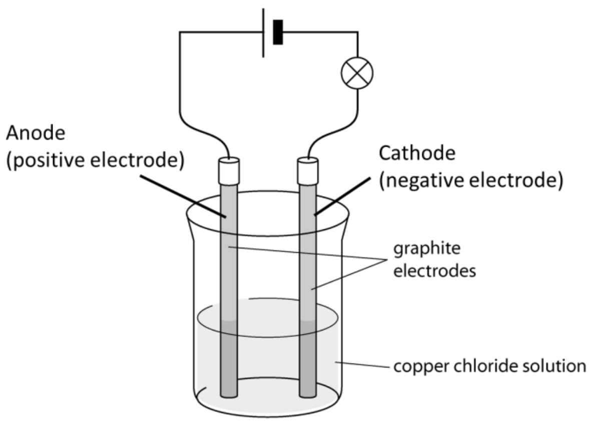
Electrolysis is the process of passing electricity through an electrolyte. This is usually used to split the electrolyte up into the elements it is made from.

Electrolyte : An ionic compound either dissolved in solution or molten
Electrode : The material used to supply electricity to the electrolyte (usually carbon or platinum)

Anode – The positive electrode
Cathode – The negative electrode
Anion – The negative ion
Cation – The positive ion

P Positive
A Anode
N Negative
I Is
C Cathode

Remember the ions are the opposite of the _odes.



During electrolysis the positive ions travel to the negative electrode and the negative ions travel to the positive electrode.
At the electrodes chemical reactions happen which turn the ions back into atoms

Lesson 2 Electrolysis of simple solutions and melts

For certain solutions and all molten electrolytes the electrolyte is turned back into the elements it is made from.

Electrolyte		Product at Anode
Copper chloride	Copper	Chlorine
Lead bromide	Lead	Bromine

Using lead bromide as an example :

Lead bromide (PbI_2) is an ionic compound and contains Pb^{2+} metal ions and I^- non metal ions

The lead ions (Pb^{2+}) are attracted to the negative electrode where they are turned back into lead metal (Pb)

The Iodide ions (I-) are attracted to the positive electrode where they are turned back into Iodine gas (I_2)

Equations at electrodes (HIGHER)
At the Cathode
Pb^{2+} + 2e- → Pb This is Reduction
At the Anode
$2I^-$ - 2e- → I_2 This is Oxidation

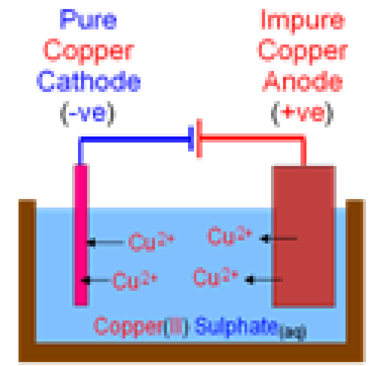
O **Remember**
O Oxidation
I Is
L Loss
R Reduction
I Is
G Gain
of electrons!

Remember the cathode supply's electrons the anode removes electrons

Lesson 3 The electrolysis of copper sulphate using copper electrodes

This is a **special case** for electrolysis where the electrodes are reactive (Usually **inert** electrodes are used)

Because the electrodes are reactive and made from the same metal as the electrolyte, the anode decreases in weight and the cathode increases in weight as copper is transferred from anode to chathode

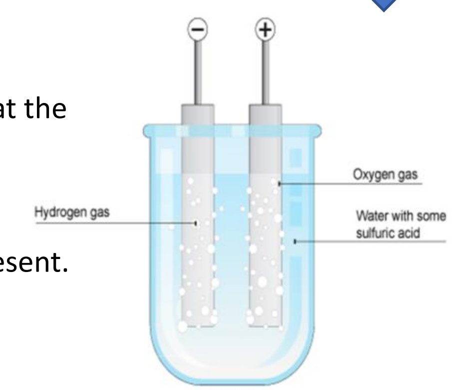


This method can be used to purify copper. If the anode is impure copper, the copper will transfer to the cathode as pure copper and in the process will lose its impurities.

Equations at electrodes (HIGHER)
At the Cathode
Cu^{2+} + 2e- → Cu This is Reduction
At the Anode
Cu - 2e- → Cu^{2+} This is Oxidation

Lesson 4 Products of electrolysis

Water contains H^+ and OH^- ions. If you do electrolysis on water you get Hydrogen gas at the anode (from H^+ ions) and oxygen gas at the cathode (from OH^- ions)



For solution electrolytes there are 4 ions present.

e.g. electrolysis of brine (salt water)

The ions present are Na^+ and Cl^- from NaCl and H^+ and OH^- from water
Both negative ions are attracted to the anode and both positive ions are attracted to the cathode, but only 1 can win!

Rules for competition at electrodes
At the cathode – the least reactive metal will win
At the anode – the least complex ion will win
Halide > Hydroxide > all other negative ions

In this case the products will be chlorine gas at the positive electrode and Hydrogen at the negative

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

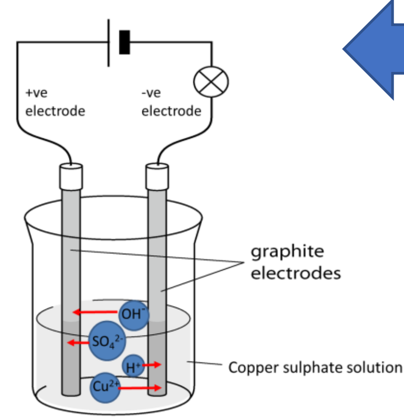
Lesson 5 The electrolysis of copper sulphate using inert electrodes

If you repeat the electrolysis of copper sulphate, but use graphite (inert) electrodes you get copper at the cathode and bubbles of oxygen gas at the anode

Copper sulphate solution contains 4 ions:-

H^+ and the OH^- from water

Cu^{2+} and SO_4^{2-} from copper sulphate



Ions present : Cu^{2+} , H^+ , OH^- , SO_4^{2-}
At the cathode (Copper) Cu^{2+} + 2e- → Cu Red.
At the anode (Oxygen) $4OH^-$ – $4e^-$ → O_2 + $2H_2O$ Ox.

Because
Copper is less reactive than hydrogen so is formed
OH^- is a more simple ion than SO_4^{2-} so oxygen is formed

Lesson 6 Example of electrolysis

Electrolysis of water (acidified with sulphuric acid)
Ions present : H^+ , OH^- , SO_4^{2-}
At the cathode (Hydrogen) 2H+ + 2e- → H_2
At the anode (Oxygen) $4OH^-$ – $4e^-$ → O_2 + $2H_2O$

Electrolysis of sodium sulphate solution
Ions present : Na^+, H^+ , OH^- , SO_4^{2-}
AT the cathode (Hydrogen) $2H^+$ + 2e- → H_2
At the anode (Oxygen) $4OH^-$ – $4e^-$ → O_2 + $2H_2O$

Electrolysis of Brine (Sodium Chloride solution)
Ions present : Na^+, H^+ , OH^- , Cl^-
At the cathode (Hydrogen) $2H^+$ + 2e- → H_2
At the anode (Chlorine) $2Cl^-$ - 2e- → Cl_2
The solution remaining will be alkali sodium hydroxide

Extraction of aluminium (electrolysis of molten aluminium oxide)
Ions present : Al^{3+} , O^{2-}
At the cathode (Aluminium) Al^{3+} + 3e- → Al
At the anode (Oxygen) $2O^{2-}$ – $4e^-$ → O_2

CC10: Electrolysis knowledge organiser (S)

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Electrode : The material used to supply electricity to the electrolyte (usually carbon or platinum)

Anode – The positive electrode
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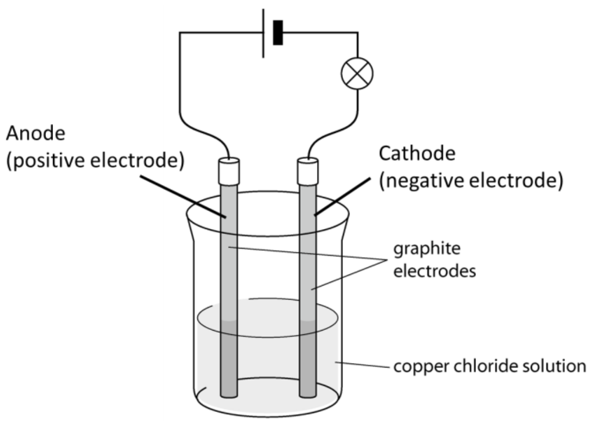

During electrolysis the positive ions travel to the negative electrode and the positive ions travel to the negative electrode.
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Lesson 2 Electrolysis of simple solutions and melts

For most electrolytes the electrolyte is turned back into the elements it is made from.

Electrolyte	Product at Cathode	Product at Anode
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Using lead bromide as an example :

The lead ions (Pb^{2+}) are attracted to the negative electrode where they are turned back into lead metal (Pb)

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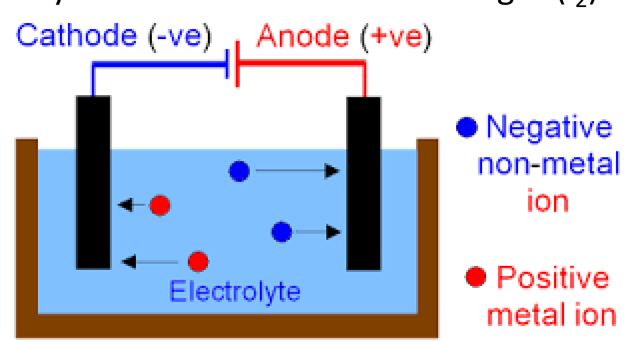

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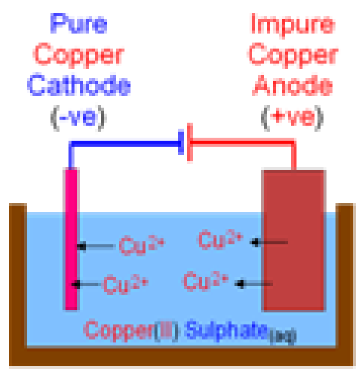

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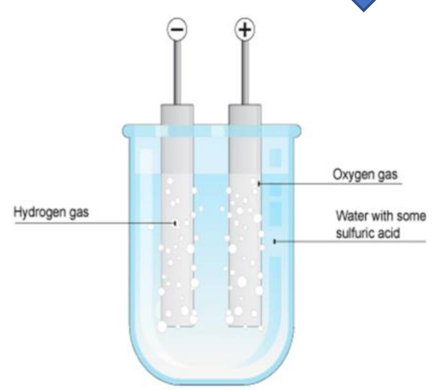

Some solutions (e.g. copper sulphate solution) contain 4 ions and there are some rules for working out which ions win at the electrode

Rules for competition at electrodes
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At the anode – the least complex ion will win
Halide > Hydroxide > all other negative ions

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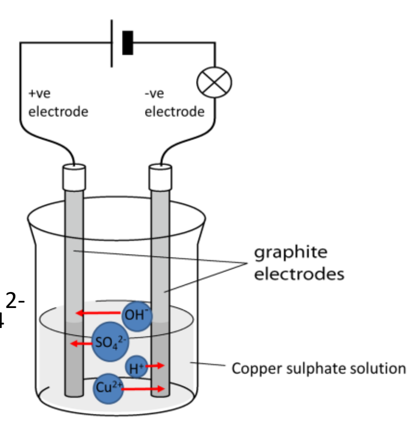

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Ions present : Na^+, H^+ , OH^- , SO_4^{2-}
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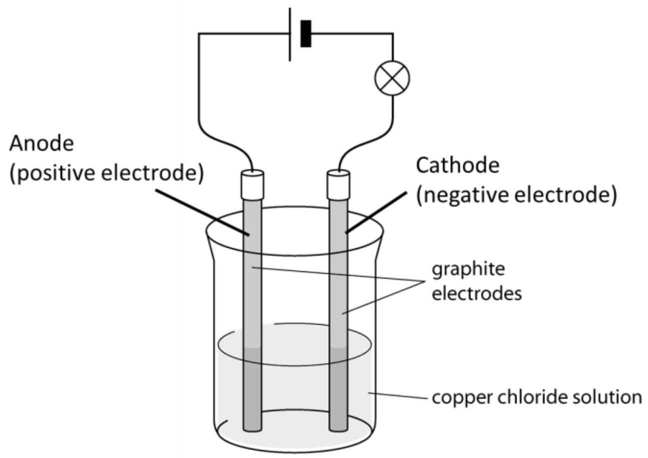

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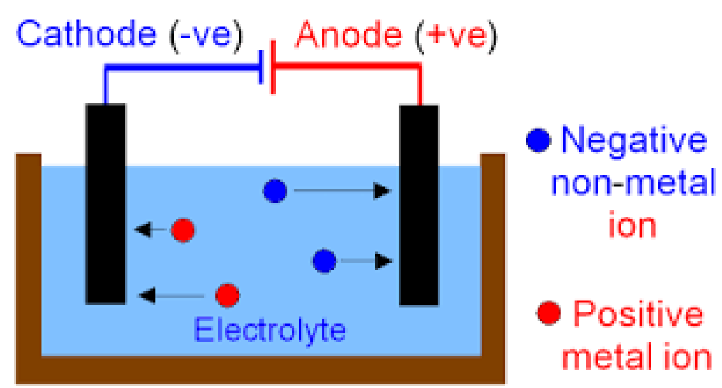

Remember metals form positive ions Non- metals form negative ions

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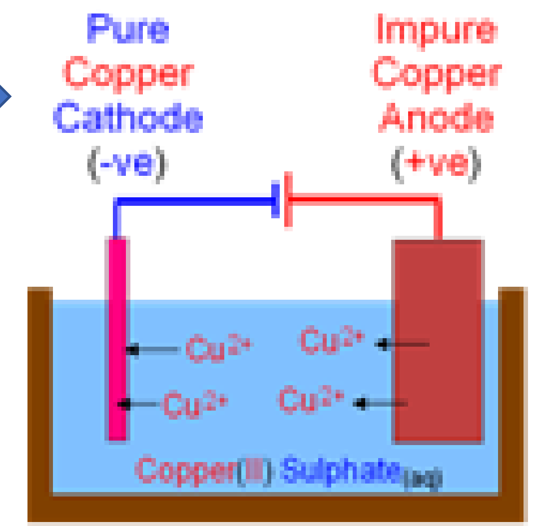

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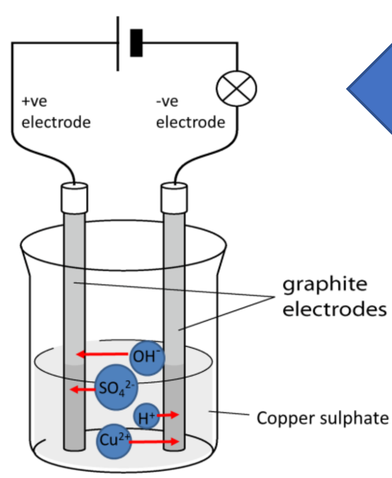

Ions present : Cu^{2+}, H^+ , OH^- , SO_4^{2-}
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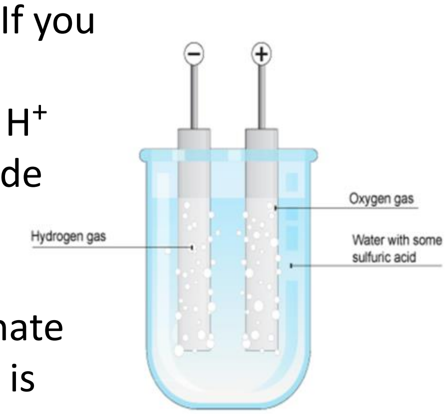
At the anode (Oxygen from water)

Lesson 5 Products of electrolysis

Water contains H^+ and OH^- ions. If you do electrolysis on water you get Hydrogen gas at the anode (from H^+ ions) and oxygen gas at the cathode (from OH^- ions)



Some solutions (e.g. copper sulphate solution) contain 4 ions and what is formed at the electrode is not always what you expect.

There are rules for working out what is formed at the electrode

Lesson 6 Example of electrolysis

Electrolysis of sodium sulphate solution
Ions present : Na^+, H^+ , OH^- , SO_4^{2-}
AT the cathode (Hydrogen from water)
At the anode (Oxygen from water)

Electrolysis of Brine (Sodium Chloride solution)
Ions present : Na^+, H^+ , OH^- , Cl^-
At the cathode (Hydrogen from water)
At the anode (Chlorine from sodium chloride)
The solution remaining will be alkali sodium hydroxide

CC11: Extraction of metals knowledge organiser (H)

Lesson 1 Reactivity of metals

Metals react by losing electrons to form positive ions.
Different metals have different reactivity and the metals which lose electrons most easily are the most reactive.

Metals react with water and acid:

	Metal	Reaction with water	Reaction with dilute acid	Tendency of metal atoms to form cations
Reactive metals	potassium	react with cold water to form hydrogen and a metal hydroxide	react violently	increasing ability of metal atoms to form positive ions
	sodium			
	calcium		react to form hydrogen and a salt solution	
Medium Reactive metals	magnesium	react very slowly, if at all, with cold water but react with steam to form hydrogen and a metal oxide		
	aluminium			
	zinc			
	iron			
Un-reactive metals	copper	do not react with cold water or steam	do not react	
	silver			
	gold			

Reaction of metals with acid (MASH)

metal + acid → salt + hydrogen

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

Reactions of metals with water

metal + water → metal hydroxide + hydrogen

$2Na + 2H_2O \rightarrow 2NaOH + H_2$

Lesson 2 Displacement reactions

In a displacement reaction a more reactive metal steals the non-metal from a less reactive metal compound.

Magnesium + copper sulphate → magnesium sulphate + copper

$Mg (s) + CuSO_4 (aq) \rightarrow MgSO_4 (aq) + Cu (s)$

In the reaction magnesium is more reactive than copper so steals the sulphate displacing the copper

potassium	most reactive	K
sodium		Na
calcium		Ca
magnesium		Mg
aluminium		Al
carbon		C
zinc		Zn
iron		Fe
tin		Sn
lead		Pb
hydrogen		H
copper		Cu
silver		Ag
gold		Au
platinum	least reactive	Pt

Redox Reactions (HIGHER)

Displacement reactions are REDOX reactions
In a redox reaction something is REDuced by gaining electrons and something is OXidised by losing electrons.

You can write displacement equations as ionic equations by ignoring spectator ions.

You can then write half equations showing what is oxidised and what is reduced

Symbol equaton : $Mg (s) + CuSO_4 (aq) \rightarrow MgSO_4 (aq) + Cu (s)$

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

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

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

In the displacement reaction magnesium is oxidised and copper is reduced.

Lesson 3 Extracting metals from ores

Most metals are reactive and are not found naturally as metals, but as metal compounds (usually metal oxides) in metal ores. The method used to extract the metal from its ore depends on the reactivity of the metal.

potassium	most reactive	K	Metals more reactive than carbon (e.g. Aluminium) must be extracted by electrolysis
sodium		Na	
calcium		Ca	
magnesium		Mg	
aluminium		Al	
carbon		C	Metals less reactive than carbon (e.g. Iron) can be extracted by heating with carbon
zinc		Zn	
iron		Fe	
tin		Sn	
lead		Pb	
hydrogen		H	
copper		Cu	
silver		Ag	Unreactive metals (e.g. gold) are found naturally as the metal
gold		Au	
platinum	least reactive	Pt	

Extracting metals with carbon

When metal ores (metal oxides) are heated with carbon the carbon can act like a metal and 'steal' the oxygen to make carbon dioxide. e.g.

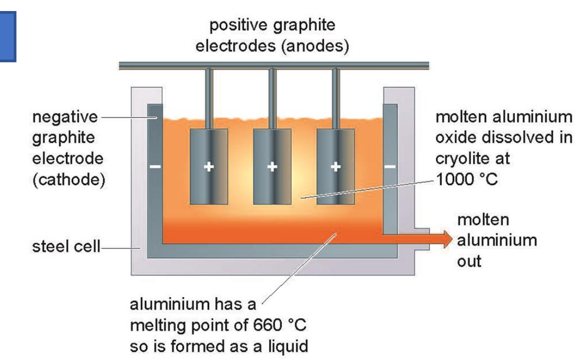
Iron oxide + carbon → Iron + carbon dioxide

$2FeO + C \rightarrow 2Fe + CO_2$

This can only be used for metals that are less reactive than carbon.

Lesson 4 Extracting metals using electrolysis

Metals that are more reactive than carbon must be extracted from their ores by electrolysis.
e.g. Aluminium
Aluminium is found as its ore Bauxite Al_2O_3
This is heated to high temperature so that it melts and is then electrolysed.



Ions present : Al^{3+}, O^{2-}
At the cathode (Aluminium)
$Al^{3+} + 3e- \rightarrow Al$
At the anode (Oxygen)
$2O^{2-} - 4e^- \rightarrow O_2$

Extracting copper

Copper can be obtained from copper rich ores by heating (smelting). But these ores are running out so biological methods are used to increase the concentration of copper
Bio leaching – uses bacterial to concentrate the copper
Phytomining – uses plants to concentrate the copper

Lesson 5 Redox

Reduction and oxidation reaction happen together and are given the term Redox.

Redox can be thought of in terms of oxygen	**(HIGHER)** in terms of electrons
Reduction is losing oxygen	Reduction is gaining electrons
Oxidation is gaining oxygen	Oxidation is loss of electrons

Corrosion
In corrosion reactions metals gain oxygen to give the metal oxide
Metals are therefore oxidised

$2Fe + O_2 \rightarrow 2FeO$

More reactive metals gain oxygen more easily and corrode easily

Metal extraction from ores
In metal extraction metal oxides lose oxygen and carbon gains oxygen. The metal is reduced and carbon is oxidised

$2FeO + C \rightarrow 2Fe + CO_2$

Redox and electrons (HIGHER)

Oxidation and reduction half equations can be used to show the transfer of electrons in redox reactions. Half equations can be used to show Redox in displacement reactions and for the reactions occurring at the electrodes during electrolysis

Lesson 6 Recycling and life cycle assessment

Metals should be recycled at the end of their use. This uses less metal ores which can be in short supply and can require a lot of energy and cause environmental problems to extract.

When considering how environmentally sustainable a product is scientists may do a Life Cycle assessment. This looks at all of the stages in a products life and evaluates the impact of each on the environment.

Stages involved:
Obtaining raw materials
Manufacture
Use
Disposal

Each of these stages should be evaluated in terms of the materials used, the energy used, the waste produced etc.

Stage of life cycle	Cotton tee shirt	Polyester tee shirt
1 Obtaining and processing raw material	Energy / resources to grow, harvest and process cotton.	Energy / resources required to drill crude oil make polyester
2 Manufacturing the product	Energy / resourses to spin cotton thread, weave it into cloth etc	Energy / resources to spin polyester thread weave it into cloth etc
3 Using the product	How long will product last? How is it washed? Temp?	How long will product last? How is it washed? Temp?
4 Disposal	Biodegradable and can be put in landfill or recycled	Non-biodegradable Can be recycled.

CC11: Extraction of metals knowledge organiser (S)

Lesson 1 Reactivity of metals

Metals have different reactivity and are placed in a reactivity series of metals. The most reactive metals are at the top
Metals react with water and acid:

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zinc		Zn	
iron		Fe	
tin		Sn	
lead		Pb	
hydrogen		H	
copper		Cu	Unreactive metals (e.g. gold) are found naturally as the metal
silver		Ag	
gold		Au	
platinum	least reactive	Pt	

Extracting metals with carbon

When metal ores (metal oxides) are heated with carbon the carbon can act like a metal and 'steal' the oxygen to make carbon dioxide. e.g.

Iron oxide + carbon ⟶ Iron + carbon dioxide

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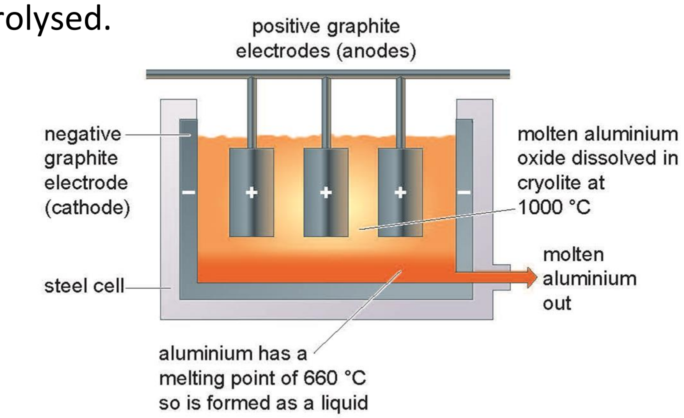

Ions present : Al^{3+} , O^{2-}
At the cathode (Aluminium)
At the anode (Oxygen)

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Oxidation is gaining oxygen

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Metals react with water and acid:

Reaction of metals with acid (MASH)

metal + acid $\longrightarrow$ salt + hydrogen

e.g.

magnesium + acid $\longrightarrow$ magnesium chloride + hydrogen

Reactions of metals with water

metal + water $\longrightarrow$ metal hydroxide + hydrogen

e.g.

Sodium + water $\longrightarrow$ sodium hydroxide + hydrogen

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In a displacement reaction a more reactive metal steals the non-metal from a less reactive metal compound.

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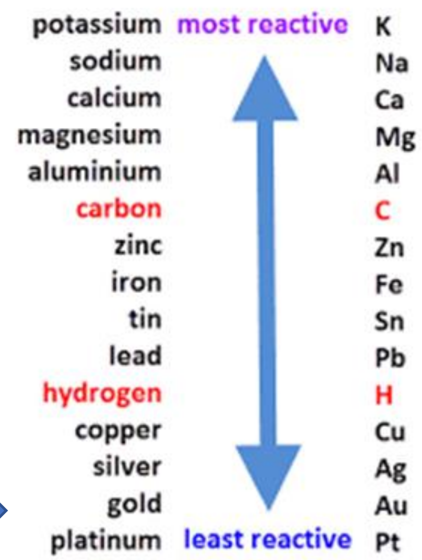

Magnesium is more reactive than copper so steals the sulphate displacing the copper

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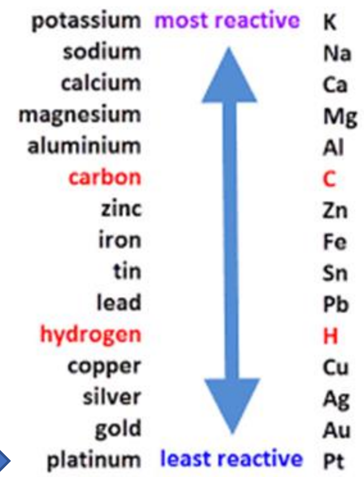
Most metals are extracted from metal ores. The method used to extract the metal from its ore depends on the reactivity of the metal.



Metals more reactive than carbon use electrolysis

Metals less reactive than carbon are extracted by heating with carbon

Unreactive metals (e.g. gold) are found naturally

Extracting metals with carbon

When metal ores (metal oxides) are heated with carbon the carbon can act like a metal and 'steal' the oxygen to make carbon dioxide. e.g.

Iron oxide + carbon $\longrightarrow$ Iron + carbon dioxide

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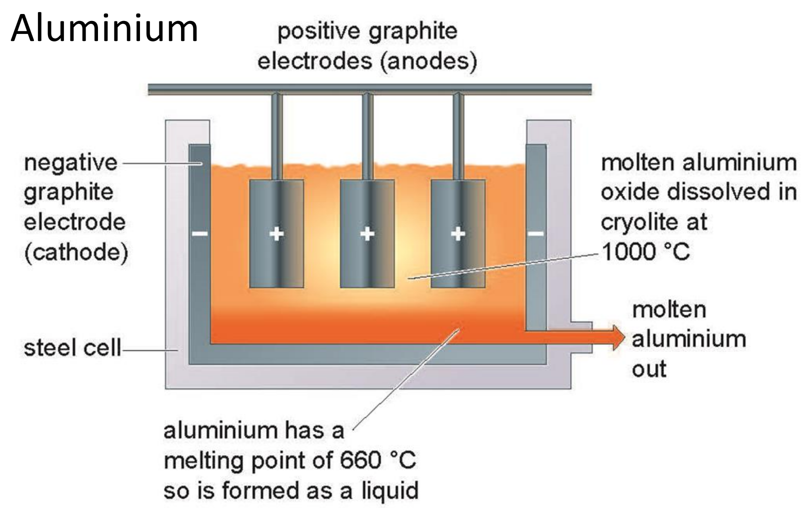
e.g. Aluminium



Electrolyte = molten aluminium oxide

Ions present : Al^{3+}, O^{2-}

At the cathode (Aluminium)

At the anode (Oxygen)

Lesson 5 Redox

Reduction and oxidation reaction happen together and are given the term **Redox.**

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Oxidation is gaining oxygen

Corrosion

In corrosion reactions metals gain oxygen to give the metal oxide. Metals are therefore oxidised

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In metal extraction metal oxides lose oxygen and carbon gains oxygen. The metal is reduced and carbon is oxidised

Lesson 6 Recycling and life cycle assessment

Metals should be recycled at the end of their use.

- uses less metal ores
- uses less energy
- better for the environment

When considering how environmentally sustainable a product is scientists may do a Life Cycle assessment.

Stages involved:

Obtaining raw materials
Manufacture
Use
Disposal

Each of these stages should be evaluated in terms of the materials used, the energy used, the waste produced etc.

CC12: Equilibria knowledge organiser (H)

Lesson 1 Reversible reactions

In some chemical reactions the products can recombine to make the reactants.
These reactions are known as reversible reactions and are given the symbol. $\rightleftharpoons$

If a **reversible reaction** takes place in a **closed system** a **dynamic equilibrium** is established.

In a dynamic equilibrium both the forward and the reverse reactions are happening at the same rate.

Rate of forward reaction = Rate of reverse reaction

Energy in reversible reactions

If a reversible reaction is exothermic (gives our heat energy) in the forward direction, it will be endothermic in the reverse direction (it will take in the same amount of heat energy)

e.g. The formation of ammonia in the Haber Process is a reversible reaction :

$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$ + Heat

The reaction is exothermic in the forward direction (it gives out heat when ammonia is made) but endothermic in the reverse direction (it takes in heat when Nitrogen and Hydrogen are made)

Lesson 2 The position of the equilibrium

A reversible reaction reaches that reaches equilibrium will have an equilibrium position. This position is not always in the middle (50% reactants and 50% products) The position of the equilibrium determines how much products or reactants are made.

Position of equilibrium is on RHS

e.g. $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$ + Heat

To make as much ammonia as possible in the Haber Process the position of the equilibrium should lie to the RHS of the equation

Changing the position of the equilibrium (HIGHER)

The position of the equilibrium (and therefore how much product is made) depends on :

- Temperature
- Pressure
- Concentration

To understand how these effect the position of the equilibrium you need to know the **equilibrium law:**

If the temperature, pressure or concentration is changed the equilibrium position will shift to try to minimise the change

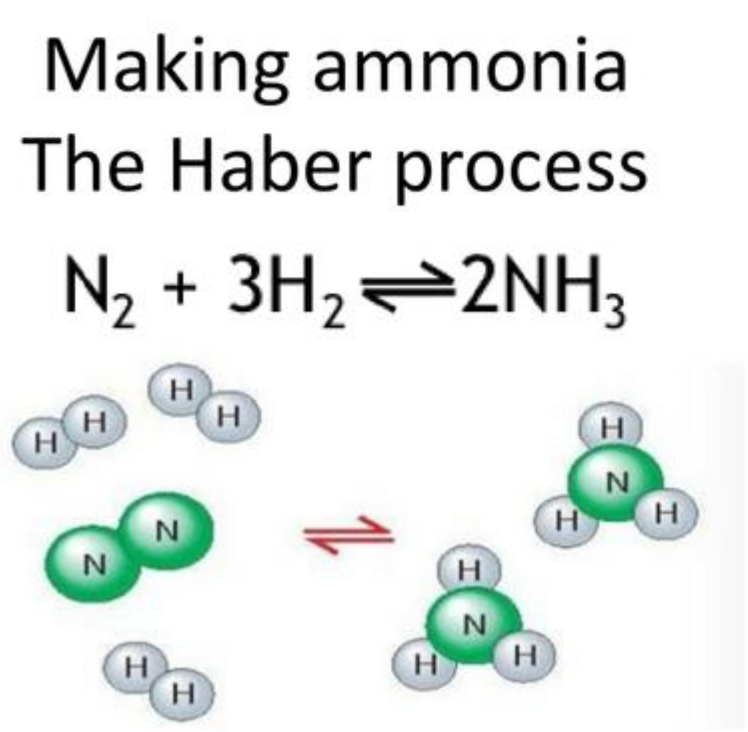
Lesson 3 The effect of temperature, pressure and concentration on the position of equilibrium (HIGHER)

The effects of concentration, pressure and temperature on the position of an equilibrium are:

1 concentration
If the concentration of reactant is increased more products are made to use them up.
If the concentration of products are decreased more products will be made to replace them.

2 pressure
If the pressure is increased the equilibrium will shift to the side of the reaction with less molecules of gas to try to reduce the pressure.

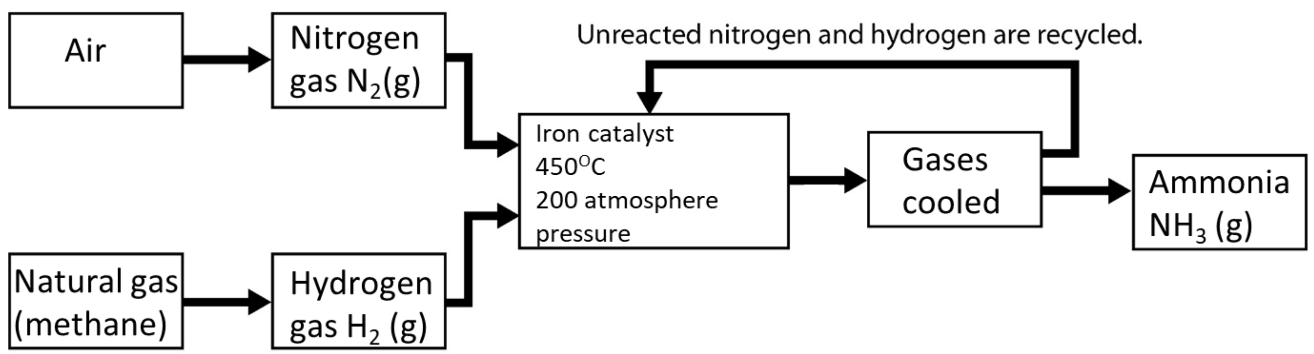
3 temperature
If the temperature is decreased the equilibrium will shift to the side that produces heat to try to heat it up.



Lesson 4 The Haber Process

The Haber process is the industrial process used to make ammonia (NH_3) from Nitrogen and Hydrogen

Nitrogen gas comes from the fractional distillation of air
Hydrogen gas comes from methane



The reaction is reversible and reaches an equilibrium

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g) \quad + \text{Heat}$$

The conditions use for the reaction are Temperature = 450 °C (low) Pressure = 200 atmospheres (high) Catalyst = Iron catalyst

Choosing conditions to maximise the yield of ammonia (HIGHER)

Conditions are chosen to ‘shift’ the position of the equilibrium to the RHS to increase the amount of ammonia formed
1 Concentration - Use high concentration of Nitrogen and Hydrogen and drain the ammonia off as it forms
2 Pressure – The reaction has more gas molecules on the LHS so more ammonia is formed at high pressure
3 Temperature – The forward reaction is exothermic so more ammonia is formed at low temperature
4 Catalyst – does not effect equilibrium, but speeds up rate of the reaction.

CC12: Equilibria knowledge organiser (S)

Lesson 1 Reversible reactions

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In a dynamic equilibrium both the forward and the reverse reactions are happening at the same rate.

Rate of forward reaction = Rate of reverse reaction

The formation of ammonia from nitrogen and hydrogen is an example of a reversible reaction that is allowed to reach equilibrium

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g) \quad + \text{Heat}$$

Lesson 2 The position of the equilibrium

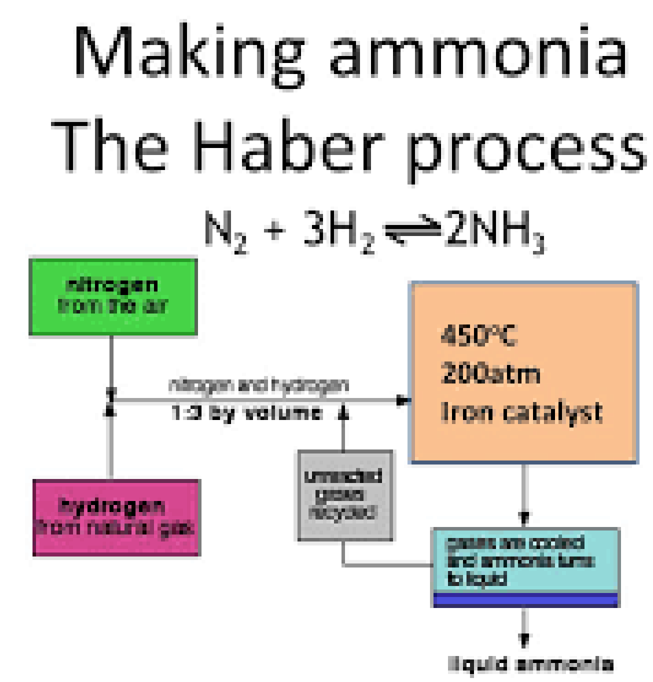
A reversible reaction reaches that reaches equilibrium will have an equilibrium position. This position is not always in the middle (50% reactants and 50% products) The position of the equilibrium determines how much products or reactants are made.

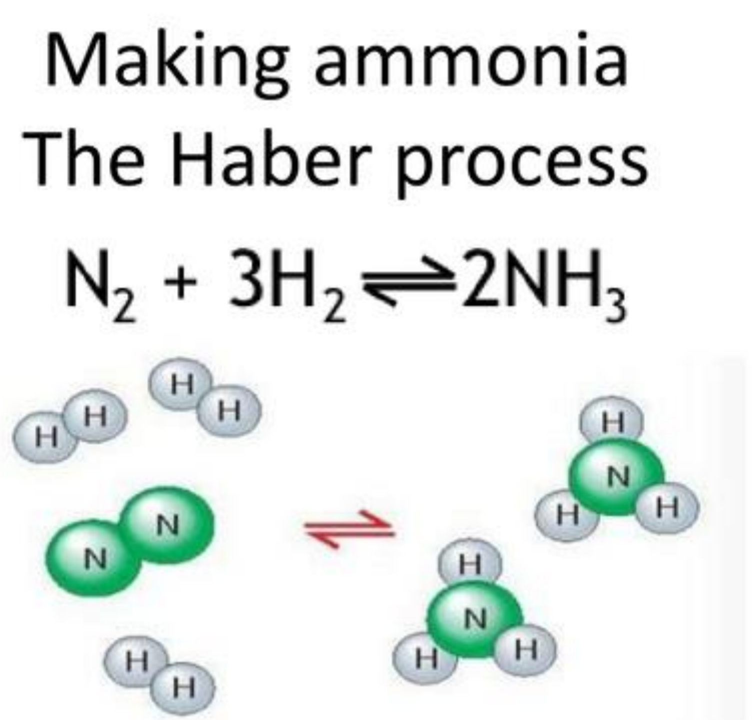
Position of equilibrium is on RHS

e.g. $N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g) \quad + \text{Heat}$

To make as much ammonia as possible in the Haber Process the position of the equilibrium should lie to the RHS of the equation

The position of the equilibrium is effected by the conditions that the reaction is carried out. In the Haber Process the conditions are chosen to shift the position of the equilibrium to the Right Hand Side and make large amounts of product.





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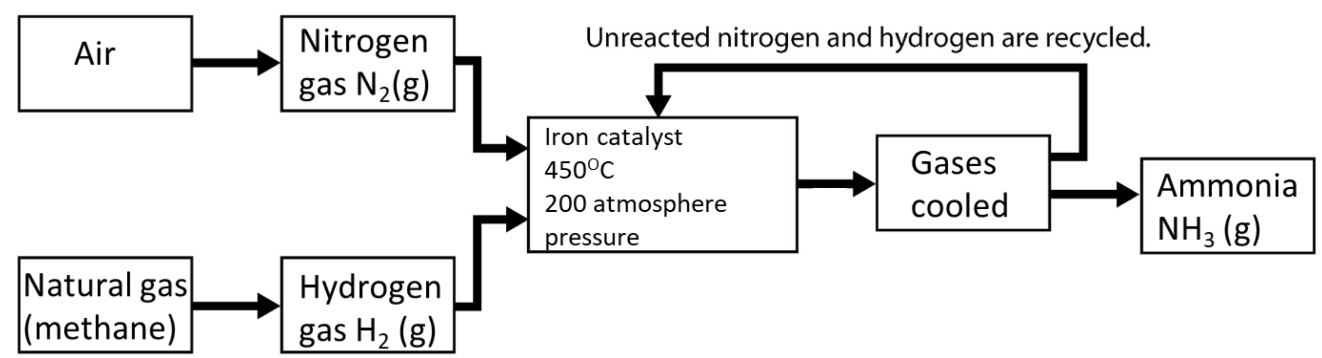

Nitrogen gas comes from the fractional distillation of air
Hydrogen gas comes from methane

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g) \quad + \text{Heat}$$

The reaction is reversible and reaches an equilibrium.

The conditions use for the reaction are

Temperature = 450 °C
Pressure = 200 atmospheres
Catalyst = Iron catalyst

These conditions are chosen to make a large amount of ammonia as cheaply as possible

CC12: Equilibria knowledge organiser (C)

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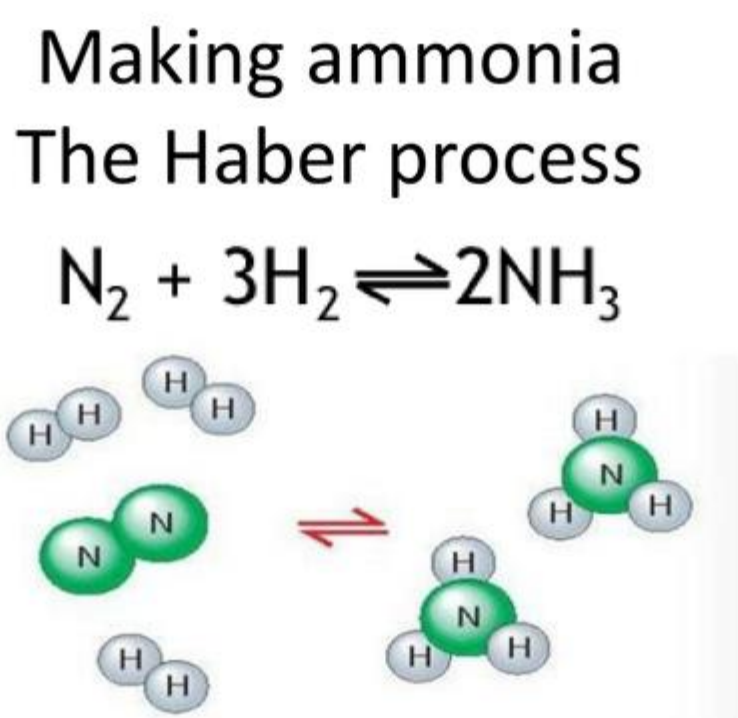

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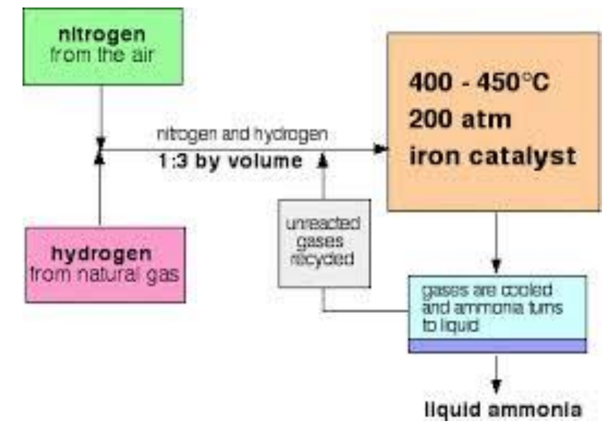

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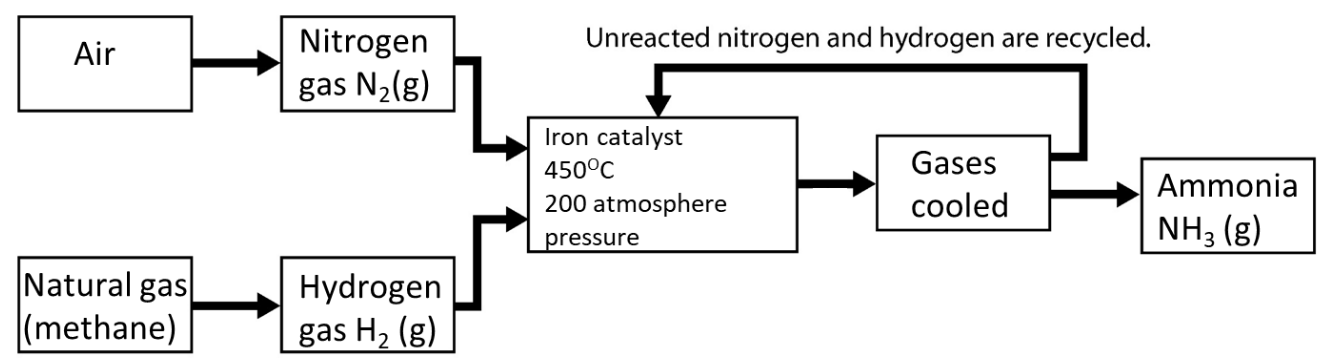

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