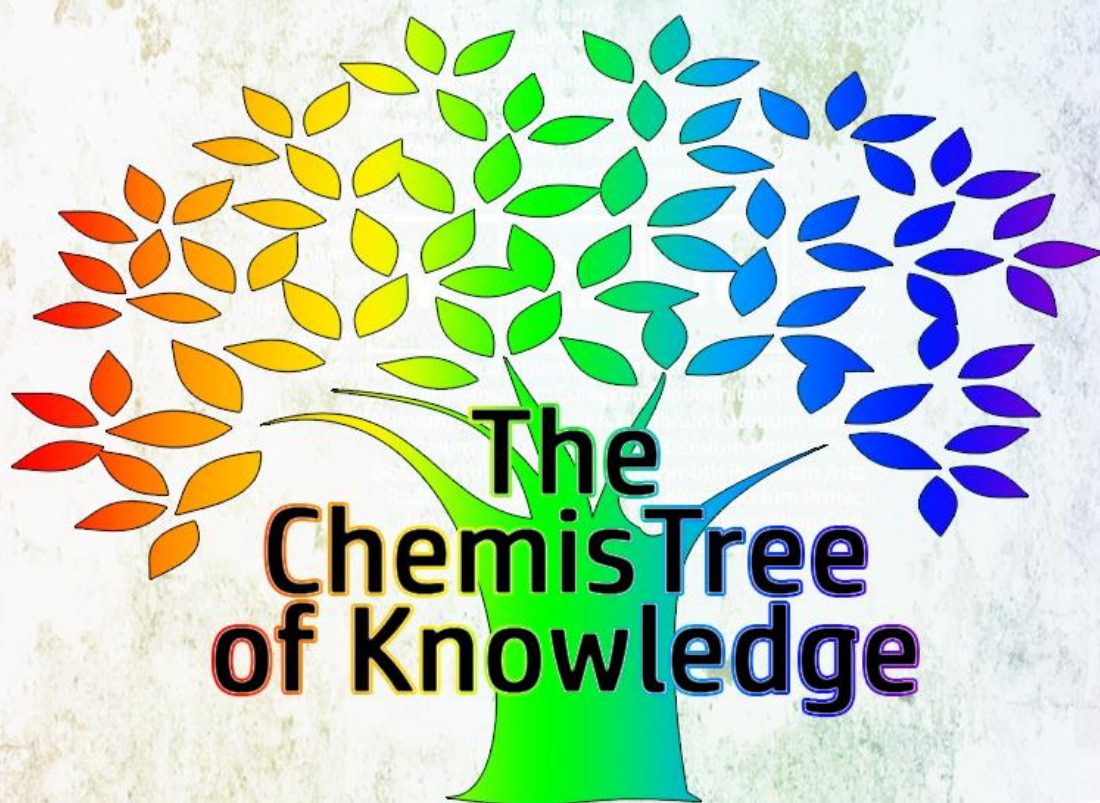




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Mr Barnes Teaches Chemistry



PAPER 1 REVISION GUIDE



AfL	Revision Topic:	P:	Q:	A:
8	States of Matter & Mixtures	04	18	25
9	Separation Techniques	06	19	26
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13	Extracting Metals and Reactivity	14	23	30
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THE PERIODIC TABLE

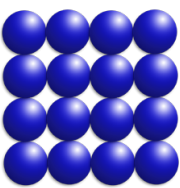
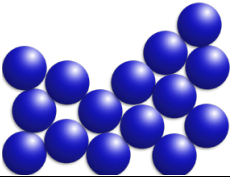
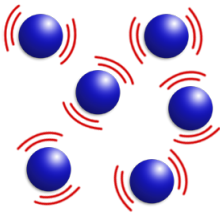
1	2	PERIODIC TABLE										3	4	5	6	7	8
1 H hydrogen 1																	4 He helium 2
7 Li lithium 3	9 Be beryllium 4											11 B boron 5	12 C carbon 6	14 N nitrogen 7	16 O oxygen 8	19 F fluorine 9	20 Ne neon 10
23 Na sodium 11	24 Mg magnesium 12											27 Al aluminium 13	28 Si silicon 14	31 P phosphorus 15	32 S sulfur 16	35.5 Cl chlorine 17	40 Ar argon 18
39 K potassium 19	40 Ca calcium 20	45 Sc scandium 21	48 Ti titanium 22	51 V vanadium 23	52 Cr chromium 24	55 Mn manganese 25	56 Fe iron 26	59 Co cobalt 27	59 Ni nickel 28	63.5 Cu copper 29	65 Zn zinc 30	70 Ga gallium 31	73 Ge germanium 32	75 As arsenic 33	79 Se selenium 34	80 Br bromine 35	84 Kr krypton 36
85 Rb rubidium 37	88 Sr strontium 38	89 Y yttrium 39	91 Zr zirconium 40	93 Nb niobium 41	96 Mo molybdenum 42	[98] Tc technetium 43	101 Ru ruthenium 44	103 Rh rhodium 45	106 Pd palladium 46	108 Ag silver 47	112 Cd cadmium 48	115 In indium 49	119 Sn tin 50	122 Sb antimony 51	128 Te tellurium 52	127 I iodine 53	131 Xe xenon 54
133 Cs caesium 55	137 Ba barium 56	139 La* lanthanum 57	178 Hf hafnium 72	181 Ta tantalum 73	184 W tungsten 74	186 Re rhenium 75	190 Os osmium 76	192 Ir iridium 77	195 Pt platinum 78	197 Au gold 79	201 Hg mercury 80	204 Tl thallium 81	207 Pb lead 82	209 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86
[223] Fr francium 87	[226] Ra radium 88	[227] Ac* actinium 89	[261] Rf rutherfordium 104	[262] Db dubnium 105	[266] Sg seaborgium 106	[264] Bh bohrium 107	[277] Hs hassium 108	[268] Mt meitnerium 109	[271] Ds darmstadtium 110	[272] Rg roentgenium 111							

Atomic Mass (Large #)

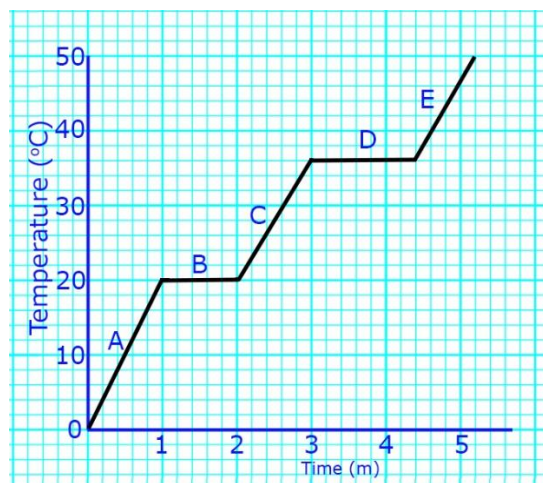
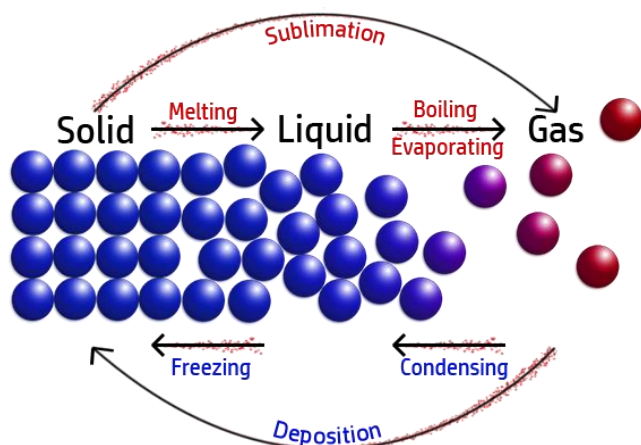
Atomic (proton) Number (Small #)

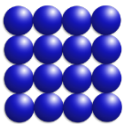
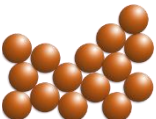
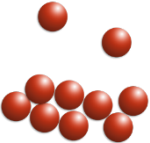
AFL 8: States of Matter and Mixtures:

A. The Particle Model:

	Solid:	Liquid:	Gas:
Particle Model:			
Arrangement of particles:	✳ Regular Pattern ✳ Touching	✳ Random pattern ✳ Touching	✳ Random pattern ✳ Far apart
Movement of particles:	✳ Vibrate slowly about a fixed position	✳ Move/flow around each other	✳ Move fast in all directions

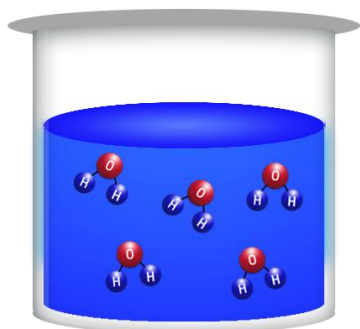
B. State Changes and State Change Graphs:



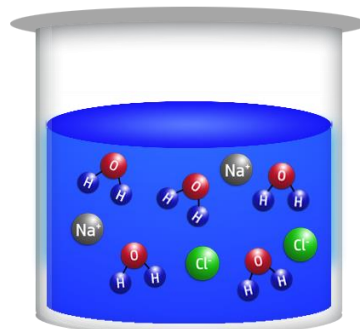
Part of Graph:	Particle Model:	Explanation:
A		✳ Particles are vibrating about a fixed position. ✳ As the temperature increases, the particles will vibrate more.
B		✳ The temperature remains constant. ✳ This is where MELTING is occurring. ✳ The energy is now being used to break some of the weak forces between the molecules (intermolecular forces)
C		✳ At this point, the substance is completely melted, in a random pattern. ✳ As the temperature increases, more particles are starting to vibrate more.
D		✳ The temperature remains constant. ✳ This is where EVAPORATING is occurring. ✳ The energy is now being used to completely break the weak forces between the molecules (intermolecular forces)
E		✳ The particles are now not touching and moving fast in all directions. ✳ As the temperature increases, the particles move faster and faster.

C. Pure or Mixture?

Pure Substance:
Water



Impure/Mixture:
Salt water



A **pure substance** is made up of only one thing.

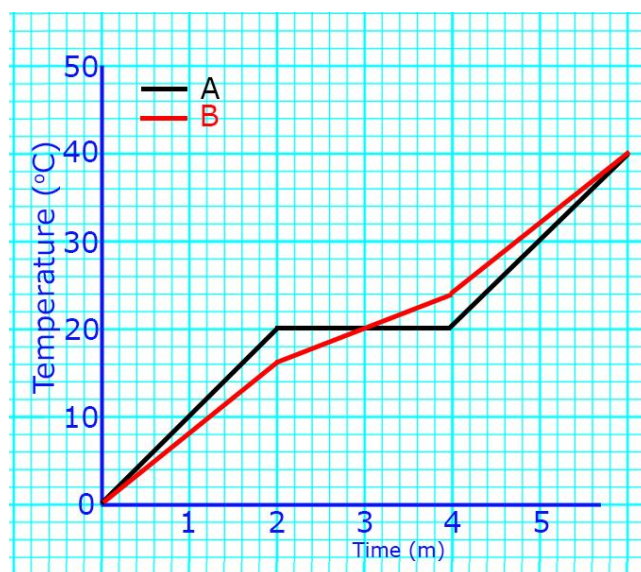
- ✳ There are pure **elements** and pure **compounds**.
- ✳ A **pure element** is only made up of one type of atom.
- ✳ A **pure compound** is only made up of one type of compound – e.g. pure water, H_2O
- ✳ A compound is two different types of elements/atoms **bonded** together – water contains two hydrogen atoms **bonded** to one oxygen atom.

A **mixture** (an impure substance) is made up of more than one thing.

- ✳ A mixture can be a combination of any different element or compound.
- ✳ Salt water (see above) contains two **different compounds** – **water**, H_2O , and **salt**, NaCl .
- ✳ As long as the elements/compounds are **not bonded** together, they can be separated and are classed as mixtures.

D: Pure or Mixture? Melting point graphs:

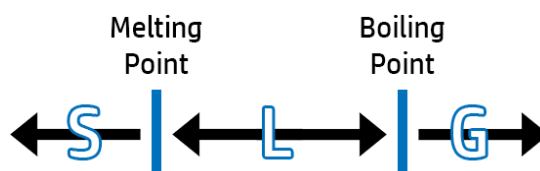
- ✳ When you heat any solid, it will eventually melt.
- ✳ If you have a pure substance, such as distilled water, it will melt at one specific temperature (See plot A, right →).
- ✳ The graph will have a flat section – this is where the intermolecular forces are being broken.
- ✳ For water, this temperature is 0°C .
- ✳ If you have an impure substance, there will be different substances, so their **physical properties** will be different.
- ✳ This means that they will melt over a **range of temperatures**. (See plot B, right →)



E: Solid, Liquid or Gas?

- ✳ If you are asked to predict whether a substance is a solid, liquid or gas. Draw a diagram to show you the melting and boiling points. Label the melting point and boiling point on the graph. If the temperature given in the question is:

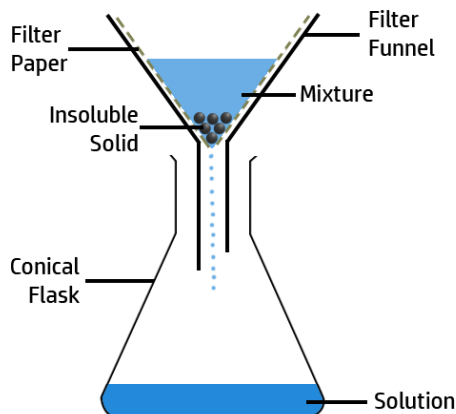
- 🧪 Less than the melting point = **SOLID**
- 🧪 Between melting and boiling point = **LIQUID**
- 🧪 More than the the boiling point = **GAS**



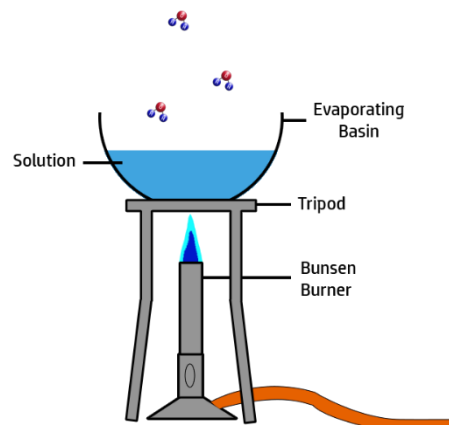
AfL 9: Separation Techniques:

A. Filtration and crystallisation:

Filtration:



Crystallisation:

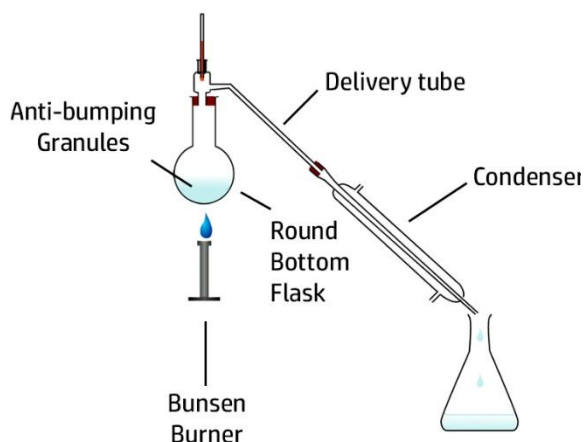


- * Filtration is a technique used to separate a liquid from an **insoluble solid** – a solid that does not dissolve.
- * The filter paper has tiny holes in it that allow water and dissolved solids to pass through.
- * The holes are too small, however, for large particles to pass through – which means solids stay in the filter paper as **residue**.

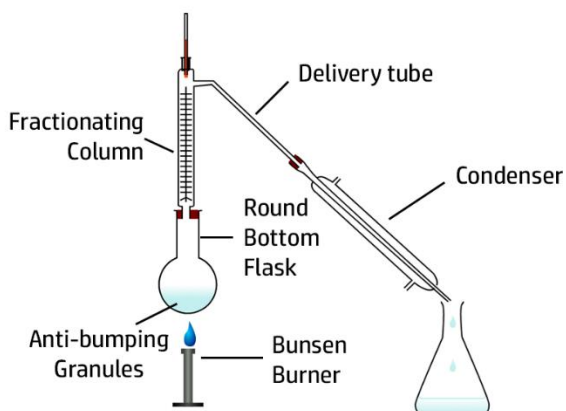
- * Crystallisation is a technique used to separate a liquid from a **soluble solid** – a solid that does dissolve.
- * The Bunsen Burner is used to **evaporate** the water, leaving behind the **crystals** of your salt in the evaporating basin.
- * After half of the water is evaporated, **leave it to cool**. The crystals can then be dried between filter paper.

B. Simple Distillation and Fractional Distillation:

Simple Distillation:



Fractional Distillation:



- * Simple distillation is a technique used to separate two different **liquids** based on their **boiling points**.
- * Example: Separating water (boils at 100°C) and ethanol (boils at 78°C).
 - 🧪 The solution is put into a round bottom flask and heated.
 - 🧪 Once it reaches 78°C, the ethanol will evaporate.
 - 🧪 The gas will rise and move into the **condenser**.
 - 🧪 The condenser is surrounded by cold water, which cools the gas down – turning it back into a **liquid**.
 - 🧪 If you stay **below** 100°C, the water will not evaporate.

- * Fractional distillation is very similar to simple distillation, but is used to separate a mixture that contains lots of different liquids.
- * All that is added is a **fractionating column**.
- * As the mixture is heated, a **temperature gradient** will exist – where it is hotter at the bottom and cooler at the top.
- * The vapour will rise up the column until it falls below its boiling point and will condense back down.
- * Gases with a **lower boiling point** will be able to go **higher** up the column – and will be collected first and will go into the condenser.
- * As the fractionating column continues to warm up, more and more of the fractions will move into the condenser, allowing them all to be collected separately.

C. Chromatography:

Chromatography is used to find out if a substance is pure or not, as well as identifying different substances in a mixture.

To prepare a chromatogram, you need to:

- ✱ Put a line in pencil (which is **insoluble**) on the bottom of the chromatogram.
- ✱ Add your sample(s) to the crosses.
- ✱ Once done, add the chromatogram to water **making sure the ink does not touch the water**. (If it does, the ink will dissolve and not move up the paper!)
- ✱ The stationary phase – the chromatography paper – is the part that the solvent moves up.
- ✱ The mobile phase – the water – is the solvent that dissolves the inks and moves them up the paper.

In the chromatogram on the right, you can see that:

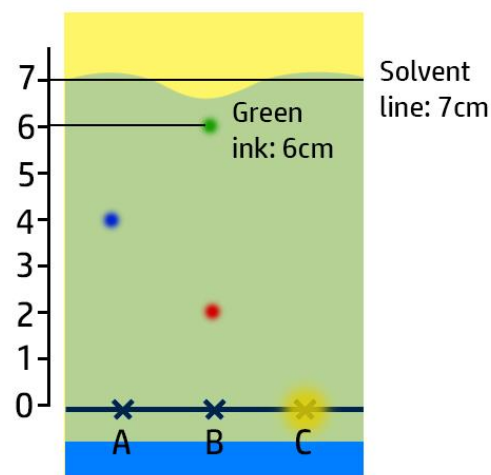
- ✱ Substance A contains one colour – blue.
- ✱ Substance B has two colours – red and green.
- ✱ Substance C is **insoluble** in water.

You can also work out how soluble the colours are:

- ✱ The most soluble colour in this chromatogram is green. It has moved the **furthest**, so is the **most soluble**.
- ✱ Red is the **least soluble** (of the ones that have dissolved!)
- ✱ Yellow is **insoluble**.

The Retention Factor, R_f , is a measure of how soluble an ink is.

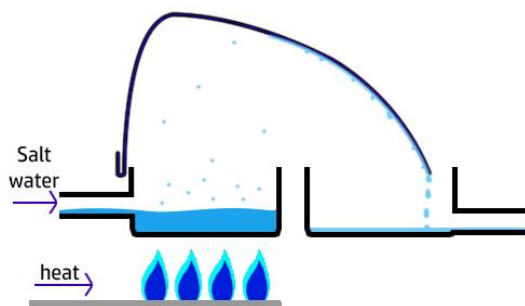
- ✱ To calculate the retention factor, divide the **distance the ink moves** by the **distance the solvent moves**.
- ✱ For the green ink, it has moved 6cm, whilst the water has moved 7cm.
- ✱ Therefore $6\text{cm} / 7\text{cm} = 0.86$.
- ✱ The retention factor will always be between 0 and 1.



D. Making Water Potable (Drinkable):

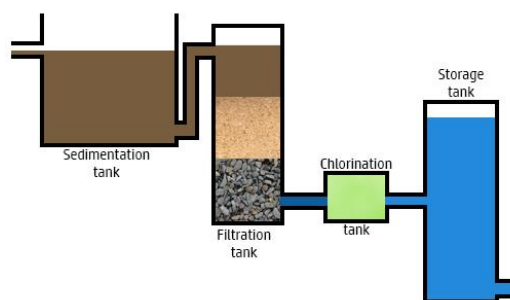
To get drinking water, we need to make sure that it is **potable** – which means safe to drink.

1. Purifying Sea Water - Desalination



- ✱ Salt water is added to a large container and heated using crude oil.
- ✱ The water will evaporate and then be cooled – condensing it back down into water.
- ✱ In the diagram above, the container on the left will be left with the salt impurities, whilst the one on the right will contain pure drinking water.
- ✱ This uses a lot of energy, so is not used to purify water on a large scale.

2. Lakes, rivers and aquifers



- ✱ Step 1: **Screening**. A sieve is used to remove large impurities, such as twigs and leaves.
- ✱ Step 2: **Sedimentation**. The water is left so that the small particles settle out at the bottom.
- ✱ Step 3: **Filtration**. Beds of sand and gravel are used to remove smaller insoluble particles that remain.
- ✱ Step 4: **Chlorination**: Chlorine is added to **sterilise** the water – killing any microorganisms present.

E. Water for Chemical Analysis:

It is important to use pure distilled/**deionised** water (which don't contain any dissolved salts) during chemical analysis because:

- ✱ It could give off incorrect results, or hide correct results
- ✱ It could result in cloudy precipitates being formed
- ✱ It could lead to an incorrect conclusion from chemical analysis.

AFL 10: Acids, Bases and Indicators:

A. Ions and the effect of indicators on acids and bases:

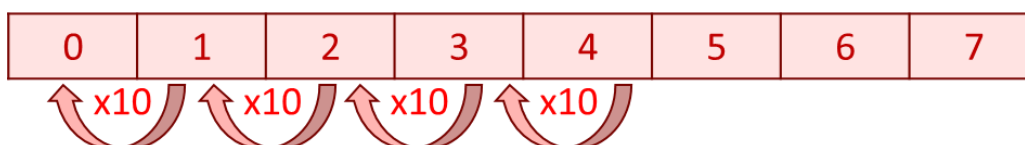
Substance:	pH range:	Ions:	Colour in phenolphthalein:	Colour in Methyl orange:	Colour in Litmus:
Acid	1-6	H ⁺	Colourless	Red	Red
Alkali/Base	8-12	OH ⁻	Pink	Yellow	Blue
Neutral	7	None	Colourless	Orange	Purple

✳ An **alkali** is any **base** that is soluble. All **alkalis** are **bases**, but **not** all **bases** are **alkalis**.

B. Ions and pH:

- ✳ The concentration of an acid is linked to the pH of the substance. Nice and simply, the more H⁺ ions there are, the lower the pH.
- ✳ Every time the pH decreases by one, the concentration of H⁺ ions **increases** by 10 times.
- ✳ For example:

🧪 Calculate how much more concentrated hydrochloric acid is, with a pH of 0, than ethanoic acid, with a pH of 4.



🧪 As you can see, the pH has gone down by 4, so the it is 10 x 10 x 10 x 10 times more concentrated = 10,000 times more concentrated.

C. Dilute, Concentrated, Strong and Weak Acids:

- ✳ The concentration of an acid shows you the amount of hydrogen ions there are in a litre (1 dm³) of water.
- ✳ The strength of an acid tells you how easily they ionise – split up into H⁺ ions.

<u>Dilute</u> The more dilute an acid is, the less H⁺ ions there are per dm³	<u>Concentrated</u> The more concentrated an acid is, the more H⁺ ions there are per dm³
<u>Strong Acids</u> HCl – Hydrochloric acid H₂SO₄ – Sulphuric acid HNO₃ – Nitric acid Strong acids dissociate / ionise completely – producing more H⁺ ions. They have lower pH's (0-2)	<u>Weak Acids</u> CH₃COOH - Ethanoic acid C₆H₈O₇ - Citric acid H₂CO₃ - Carbonic acid Weak acids do not fully ionise – only a small amount dissociate – fewer H⁺ ions. pH = 2-6

D. What happens to the ions during neutralisation?

When any neutralisation reaction occurs, the same basic reaction occurs. Hydrogen ions react with hydroxide ions to form water:



E. Testing for Gases:

Hydrogen: Hydrogen is a flammable gas. If you take a **lit splint** and add it to a test tube containing hydrogen, you will hear a **squeaky pop**.

Carbon Dioxide: Carbon dioxide turns limewater cloudy. **Bubble** the gas through **limewater** and if it goes **cloudy/milky**, carbon dioxide is present.

F. Naming Salts:

When an acid reacts with a base, a salt is formed. You need to know how to name the salts.

Step 1: Naming the Salt

- If you have hydro**chloric** acid, HCl, you get a **chloride** salt
- If you have **nitric** acid, HNO₃, you get a **nitrate** salt
- If you have **sulphuric** acid, H₂SO₄, you get a **sulphate** salt.

For example, if you react **lithium** hydroxide with **nitric** acid, you take the name of the metal – lithium – and add the salt ending, which in this case is a nitrate:



Step 2: The by-products:

You also get by-products in the reaction, depending on what you are reacting the acid with:

- A **metal** on its own will produce **hydrogen** gas, H₂
- A metal **oxide/hydroxide** will produce **water**, H₂O, and
- A **carbonate** will produce **water and carbon dioxide**, H₂O + CO₂

Step 3: Putting it all together:

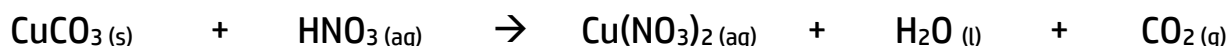
Question: Copper carbonate, CuCO₃ reacts with sulphuric acid. Write the word equation. (2)

- Work out the salt. From my reactants, the metal is **copper**, and the acid gives me a **sulphate** – therefore my salt is **copper sulphate**
- Work out the by-products. I have a **carbonate**; therefore, I have **water** and **carbon dioxide**.



G. What is seen when acids and bases react together?

Look at the chemical equation below. What would you be able to **see** during this reaction?



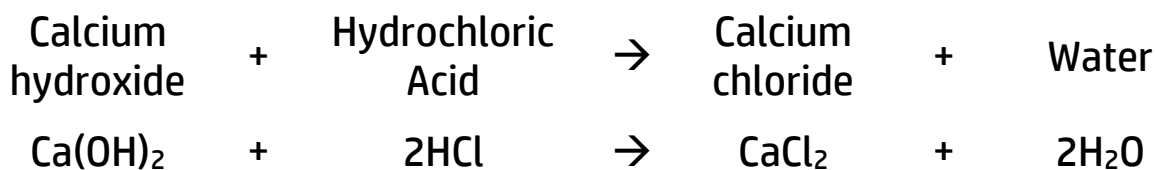
The state symbols show that a solid has disappeared and that a gas has formed. Therefore, we will see:

- The solid has disappeared/dissolved (1)
- There will be bubbling/fizzing/effervescence proving there is a gas (1)

Don't say a gas is formed – you can't **see** that – but you can see bubbles, which **prove** there is a gas.

AfL 11: Preparing Salts:

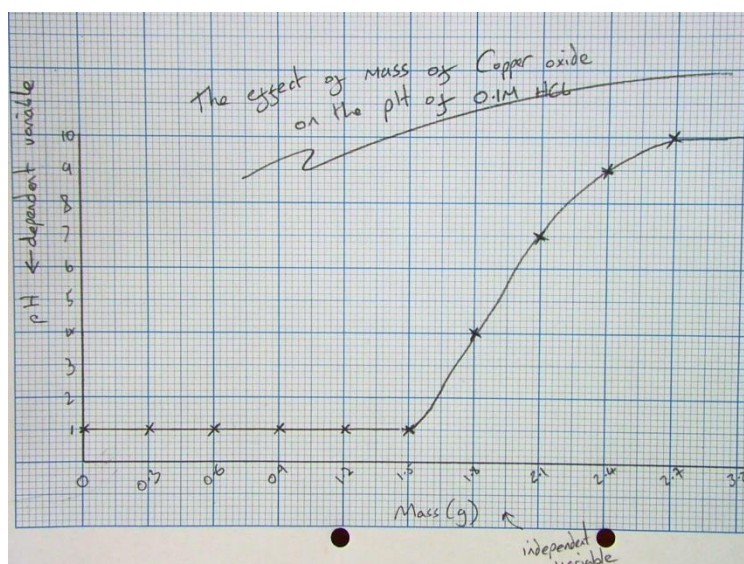
A. Core Practical: Investigating the change in pH when calcium hydroxide is added to an acid:



This is one of the big practicals that you need to know for the exam. You could be asked how to carry it out; to analyse the results; to explain the results or to evaluate the risks.

To investigate the pH of a substance, you will need to be able to do the following:

1. Measure out 50cm³ of hydrochloric acid into a beaker.
2. Record the pH of the solution by putting some Universal Indicator paper onto a white tile and place a drop of the acid onto it.
3. Leave it for 30 seconds to make sure the colour change is complete and then record the pH.
4. Measure out 0.3g of calcium hydroxide and add it to the beaker.
5. Record the pH of the substance and repeat until 2.4g of the solid is added.
6. Plot a graph as seen on the right.



As you can see from the experiment:

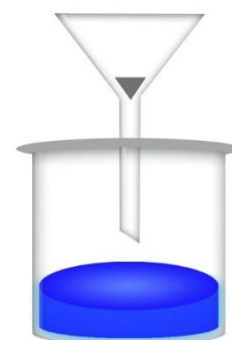
- * The pH stays at 1 until 1.5g of Calcium Hydroxide is added.
- * After 1.5g, the pH starts to increase rapidly and continues until about 2.7g of the solid is added, at which point it stays at pH 10.
- * The reaction is **neutral** at 2.1g.

Risks and management:	Improvements:	Variables:
* Calcium oxide = corrosive/ Calcium hydroxide = irritant – weak goggles and if you get it on your hands, wash it off.	* Use a pH meter instead of Universal Indicator paper. * It is more accurate (1 dp.) than UI paper.	* Independent: Mass of Ca(OH)₂ * Dependent: pH of substance * Control: Volume of acid, concentration of acid, type of tablet, etc.

B. Core Practical: How to prepare soluble salts from an acid and an insoluble reactant:

Example: How can you produce soluble copper sulphate from sulphuric acid and insoluble copper oxide?

1. **Heat the acid in a water bath** to speed up the reaction. (In a fume cupboard to reduce acidic fumes)
2. **Add the copper oxide to the acid.** A reaction will occur, producing copper sulphate – your soluble salt
3. **Continue adding copper oxide until no more reacts.** The acid is now neutral, and you have your salt, water and excess copper oxide.
4. **Use a filter paper to remove the excess copper oxide,** leaving you with the copper sulphate and water
5. **Heat the solution gently.** To evaporate off some of the water.
6. **Leave the solution to cool and crystallise the salt** – then leave to dry.



C: How to prepare soluble salts from an acid and a soluble reactant:

Example: How can you produce soluble sodium chloride from hydrochloric acid and soluble sodium hydroxide?

1. **Measure out a certain amount of alkali (sodium hydroxide) into a conical flask.**
2. **Add an indicator – such as phenolphthalein.** Phenolphthalein will be pink in an alkali/base.
3. **Add the hydrochloric acid slowly until the end point.** For phenolphthalein, the end point will be when it turns colourless.
4. **Repeat the experiment without the indicator,** this gives you just sodium chloride and water.
5. **Heat the solution gently.** To evaporate off some of the water.
6. **Leave the solution to cool and crystallise the salt** – then leave to dry.



D. How to predict insoluble salts:

When it comes to preparing insoluble salts from soluble salts, first you need to know the solubility rules:

Salts:	Soluble:	Insoluble:
Sodium, potassium, ammonium	All soluble	-
Nitrates	All soluble	-
Chlorides	Mostly soluble	Silver chloride and lead chloride
Sulphates	Mostly soluble	Lead, barium and calcium sulphate
Carbonates/hydroxides	Sodium, potassium and ammonium	Mostly insoluble

E. How to prepare insoluble salts:

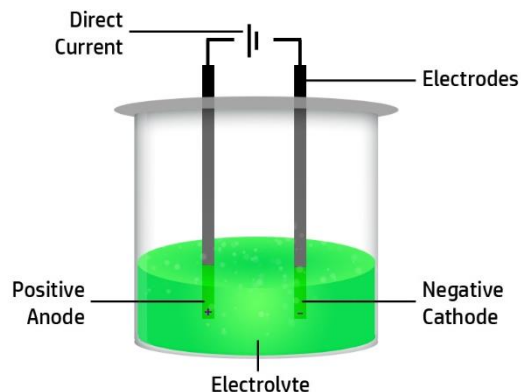
If you wanted to prepare a pure dry precipitate of lead chloride from the above reaction, there are 4 steps you need to follow:

1. **Dissolve** the solid copper chloride and lead nitrate
2. **Mix** them together to produce your copper nitrate and lead chloride
3. **Filter** the solution to give you lead chloride in the filter paper
4. **Wash** the filter paper to remove the soluble copper nitrate
5. **Dry** the filter paper in a desiccator or an oven to leave you with your pure, dry precipitate.

AfL 12: Electrolysis:

A. What is Electrolysis?

- ✳ Electrolysis is the *breaking down* of an **electrolyte** using **electrical energy**.
- ✳ An **electrolyte** is any liquid that contains **ions**. The liquid can be *molten* (melted) or *aqueous* (dissolved).
- ✳ The electrical energy needed is **direct current**.
- ✳ Normally it involved **inert** (unreactive) electrodes which aren't part of the reactions.



B. How do you remember the name and charge of the electrodes?

If you're in the exam and are trying to remember the name of the electrodes, don't forget to PANIC!

P⁺ositive

Anode

N⁻egative

Is

Cathode

The anode is always positive, and the cathode is always negative. To remember that, use the acronym on the left.

Another way to remember it is that:

- ✳ The cations will go to the cathode.
- ✳ Cations are 'paw'sitive.
- ✳ Opposites attract, so the cathode must be negative.

C. What forms at the electrodes?

1. Molten electrolytes: Example – Lead Bromide, PbBr₂ (l)

- ✳ When you have a molten electrolyte, you only have two ions in the liquid – the metal and the non-metal.
- ✳ At the cathode, lead ions (Pb²⁺) will gain electrons and turn back into lead atoms.
- ✳ Half Equation: $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$
- ✳ At the anode, bromide ions will lose electrons and turn back into bromine molecules, Br₂.
- ✳ Half Equation: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$

2. Aqueous electrolytes:

- ✳ When you electrolyse an aqueous solution, you have H⁺ and OH⁻ ions **as well** as the metal and non-metal ions.
- ✳ For example: copper chloride, CuCl₂, will have Cu²⁺ and H⁺ cations as well as Cl⁻ and OH⁻ anions.
- ✳ You need to be able to work out what will form. There are two rules to know:
 - 🧪 Cathode: The **least** reactive ion will form.
 - 🧪 Anode: If there is a **halide** (group 7), that will form – **if not, the hydroxide will go there, and form water and bubbles of oxygen** will.

Aqueous Solution:	Form at cathode:				Form at anode:			
Copper chloride, CuCl ₂ (aq)	Cu ²⁺	✓	H ⁺	✗	Cl ⁻	✓	OH ⁻	✗
Sodium chloride, NaCl (aq)	Na ⁺	✗	H ⁺	✓	Cl ⁻	✓	OH ⁻	✗
Sodium sulphate, Na ₂ SO ₄ (aq)	Na ⁺	✗	H ⁺	✓	SO ₄ ²⁻	✗	OH ⁻	✓
Sulfuric acid, H ₂ SO ₄ (aq)	NA		H ⁺	✓	SO ₄ ²⁻	✗	OH ⁻	✓

D. What is happening at the electrodes? (H)

When the ions move to the electrodes, they turn back into their original elements.

In the example of molten sodium chloride:

- ✱ The sodium ions turn back into sodium metal.
- ✱ The chloride ions turn back into chlorine gas.

Oxidation

Is

Loss

Reduction

Is

Gain

- ✱ They do this by either **losing** or **gaining** electrons.

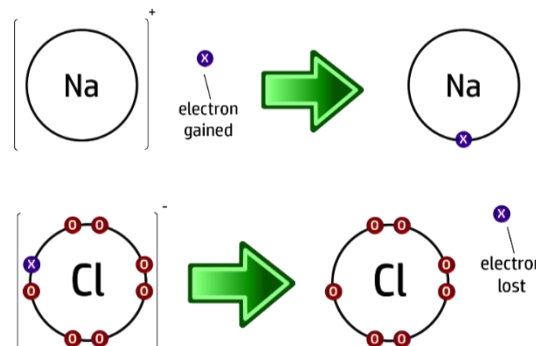
- ✱ Use the oil rig acronym (left) to work out which is oxidation, and which is reduction.

- 🧪 The sodium ions **gain electrons** to turn back into sodium metal.

- 🧪 Sodium is **reduced**.

- 🧪 The chloride ions **lose electrons** turn back into chlorine gas, Cl₂.

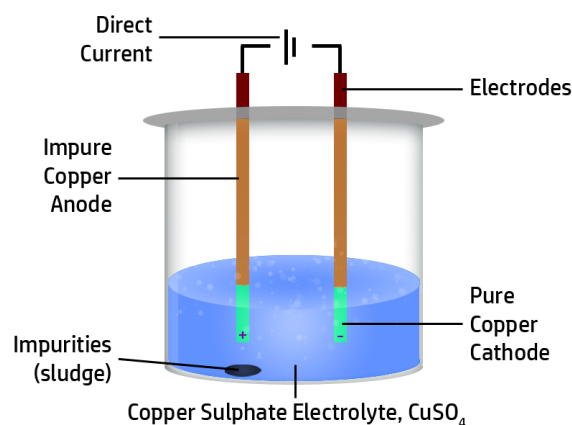
- 🧪 Chlorine is **oxidised**.



E. Core Practical: Electrolysis of Copper Sulphate Solution

1. Electrolysis of CuSO₄ with copper electrodes:

- ✱ During the electrolysis, the copper atoms in the impure copper anode lose two electrons and turn into Cu²⁺ ions (**oxidation**).
- ✱ (H) Anode half equation: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$
- ✱ They then move into the solution and are attracted to the negative cathode.
- ✱ Once there, the Cu²⁺ ions gain electrons (**reduction**) to turn back into copper metal.
- ✱ (H) Cathode half equation: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
- ✱ This causes the mass of the cathode to increase – with pure copper – and the mass of the anode to decrease.
- ✱ The impurities will fall off of the impure electrode and fall to the bottom – this is called **sludge**.



Method:

You will need to be able to explain how to investigate the change of current on the mass of the electrodes in the above set up.

1. Select two clean pieces of copper foil. Label one the anode and one the cathode.
2. Measure the masses of each electrode and record them
3. Turn on the power and use a **variable resistor** to change the current to 0.2A.
4. Leave the power on for 20 minutes.
5. Turn the power off, wash the electrodes and dip into propanone – leave to dry.
6. Measure the mass of the electrodes and calculate the change in mass for the anode and cathode.
7. Repeat for 0.3A, 0.4A and 0.5A.

2. What happens if we use Inert Electrodes?

If you carry out electrolysis of copper sulphate using **inert** (unreactive) **electrodes** instead of copper electrodes, you see different things:

- ✱ The cathode will be coated in copper:
 - 🧪 Half equation: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
- ✱ Oxygen will form at the anode, which will be *seen* as **bubbles**:
 - 🧪 Half equation: $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$

AfL 13: Extracting Metals and Reactivity:

A. The Reactivity Series

- ✳ When you react a metal with water and different acids, you can use the observations to work out how reactive they are.
- ✳ The more bubbles/fizzing there is, the more reactive the metal.
- ✳ This can be put into a reactivity series:

Potassium	These are the most reactive metals:
Sodium	✳ They react violently with acids
Calcium	✳ They react quickly with cold water to form hydrogen and a metal hydroxide
Magnesium	These are less reactive metals:
Aluminium	✳ They react very slowly with cold water
Zinc	✳ They react with steam to form hydrogen and a metal oxide
Iron	✳ They react less quickly with acids
Copper	These are unreactive metals:
Silver	✳ They do not react with cold water
Gold	✳ They do not react with steam
	✳ They do not react with acids

- ✳ As you go up the reactivity series, the tendency of the metal atom to form cations increases.
- ✳ This means that potassium will form cations, K^+ a lot easier than iron.

B. Ores:

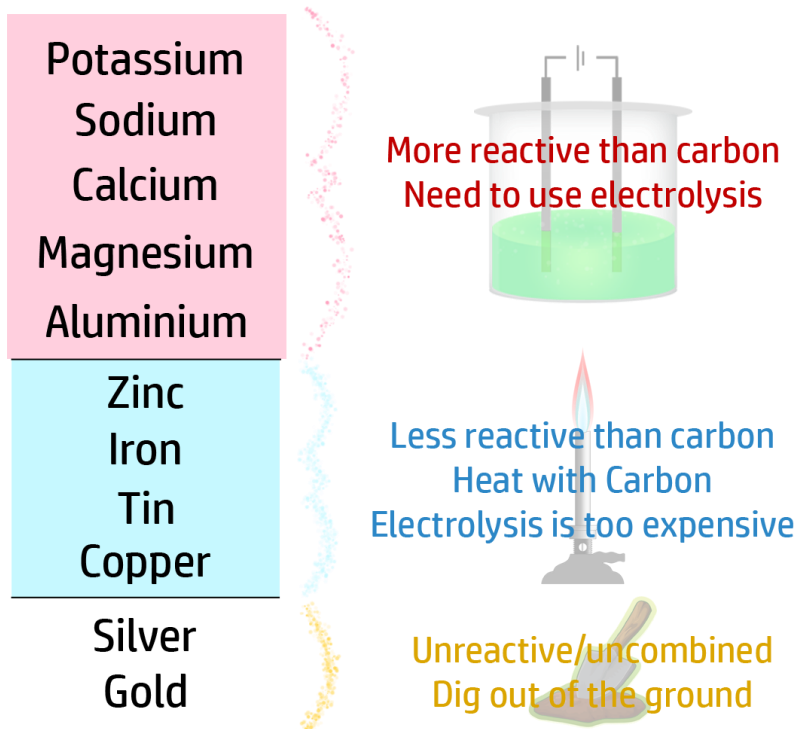
An ore is a rock that contains enough metal to make it financially worthwhile to extract. If it isn't going to give a profit to extract it – it is not an ore!

- ✳ Most metals are found in ores, such as bauxite – aluminium oxide, Al_2O_3
- ✳ Some metals are found uncombined – these are the metals that don't react.

C. Extracting Metals:

There are three ways of removing metals from the ground:

- ✳ **Dig them out** – this can only be done to the **unreactive metals** silver, gold and platinum
- ✳ **Heat them with carbon** – this can only be done with metals less reactive than carbon – the transition metals.
- ✳ **Electrolysis** – this is for elements more reactive than carbon as it is expensive.



D. Biological Methods of Extracting Metals:

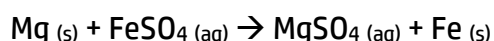
There are two other methods of extracting metals which are used to extract rocks with small amounts of metals in – called **low grade ores**: **Bioleaching and Phytoextraction**

Process:	Method:	Advantages:	Disadvantages:
Bioleaching	✱ Bioleaching uses bacteria to separate metals from low grade ores. ✱ The bacteria produce a solution of copper ions called leachate. ✱ The copper is then extracted by displacement using scrap iron.	✱ Does not require higher temperatures ✱ No harmful gases ✱ Less damage to landscape than mining ✱ Conserves supplies of higher-grade ores	✱ Very slow ✱ Toxic substances and sulphuric acid can be produced – damaging the environment
Phytoextraction	✱ Phytoextraction involves growing plants that absorb metal compounds. ✱ The plants are burnt, forming ash, which the metal can be extracted from.	✱ Can extract metals from contaminated soil ✱ No harmful gases ✱ Less damage to landscape than mining ✱ Conserves supplies of higher-grade ores	✱ Very slow ✱ More expensive than mining some ores ✱ Growing plants depends on the weather/climate

E. Displacement Reactions:

- ✱ When you add a metal to a solution containing a metal, you can predict whether a displacement reaction will occur or not.
- ✱ The more reactive metal wants to be part of the compound – leaving the least reactive as the stand-alone metal.
- ✱ For example:

Magnesium + Iron Sulphate → Magnesium sulphate + Iron



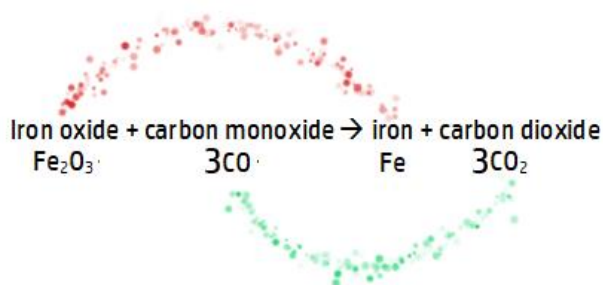
- ✱ Here, magnesium is **more reactive** than iron so **swaps** places with it.

F. Oxidation and Reduction

There is another definition for oxidation and reduction that you need to know:

- ✱ **Oxidation is the addition of oxygen**
- ✱ **Reduction is the removal of oxygen**

For example:



- ✱ The iron oxide has had oxygen removed from it to form iron – Iron oxide is reduced
- ✱ The carbon monoxide has had oxygen added to it to form carbon dioxide – carbon monoxide has been oxidised.

When metals are oxidised, they **corrode**. When iron **rusts**, it isn't just reacting with oxygen, but water as well.

AfL 14: Reversible Reactions and Recycling:

A. Reasons to recycle metals:

- ✳ Conserves Earth's natural resources
- ✳ Less mining of ores is needed. This is good because:
 - 🧪 It doesn't damage the landscape
 - 🧪 It doesn't create noise/dust pollution
- ✳ It can take less energy to recycle than to extract from the ore
- ✳ Metals don't end up wasted in landfill sites.
- ✳ However – it can cost time and money to recycle 😞

B. Life Cycle Assessments

A lifestyle assessment looks at the **total environmental cost** of a product by looking at each stage of the life of a product:

Stage:	Description:
Choice of material	✳ When metals are extracted from their ores, it needs a lot of energy and can produce a lot of pollution. ✳ Raw materials can come from crude oil, which is non-renewable. ✳ Crude oil also gives out greenhouse gases when combustion occurs.
Manufacture	✳ Manufacturing needs a lot of energy ✳ It <i>can</i> cause a lot of pollution ✳ Waste products need disposing of safely – some can be recycled ✳ The water used in lots of manufacturing needs to be safe/unpolluted when put back into the environment
Product use	✳ The products can themselves be harmful – such as: 🧪 Toxic fumes from paint, 🧪 Toxic gases from combustion and 🧪 Fertilisers draining into lakes/rivers causing eutrophication
Disposal	✳ Lots of products are disposed of in landfills, which takes up space and can pollute the land/water ✳ Products can also be burnt – which can give off toxic gases or greenhouse gases

Each of these factors need to be considered when manufacturing a product. If there are multiple ways of producing a product, the one that has the least environmental cost will be chosen:

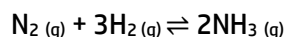
Example: A decision needs to be made about whether to produce a cabinet from two different sources:

Source:	Material	Waste solid produced	Water used (m ³)	Expected lifespan (years)
A	Iron from an iron ore	15,000 kg	8.2	20
B	Recycled iron	5,400 kg	6.5	40
C	Recycled steel	9,000kg	4.5	14

- ✳ You would not choose A because the iron comes from an ore – which takes more energy to extract than recycling. It also produces the most waste and uses more water.
- ✳ You would choose B because it is recycled, which uses less energy, and produces the least waste. It also will last the longest and therefore has the least environmental cost.

C. What is a reversible reaction?

- ✧ A reversible reaction is any reaction that can go in both the forward direction and backwards direction.
- ✧ It has a different symbol: ' $\rightleftharpoons$ ' instead of ' $\rightarrow$ '
- ✧ This can be at the same time, or separately.



In this reaction, you can see that nitrogen and hydrogen react to form ammonia, and ammonia decomposes to form nitrogen and ammonia. To make it easier to write, we use the $\rightleftharpoons$ symbol.

D. Equilibrium

When the forward reaction and backward reaction are occurring at the same rate, and the concentrations **do not change** – it is at **equilibrium**

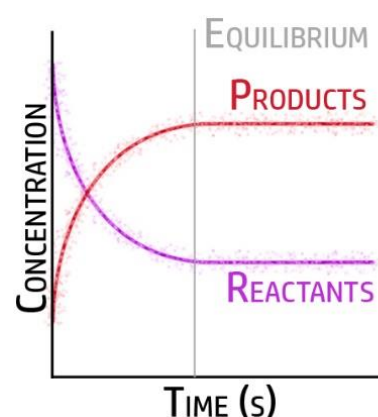
When the forward and backwards reactions are occurring at the **same time and same rate**, it is at **dynamic equilibrium**.

This means that:

1. The concentration does not change.
2. The forward reaction is occurring.
3. The backwards reaction is occurring.
4. The forward reaction and backward reaction are occurring at the same rate/speed

This doesn't mean that there is 50% of products and 50% of reactants though...

- ✧ There may be a higher percentage of products, 80% for example, but it will stay at 80% whilst at equilibrium.
- ✧ Similarly, if there is 60% of reactants and it is at dynamic equilibrium, it will stay at 60%.

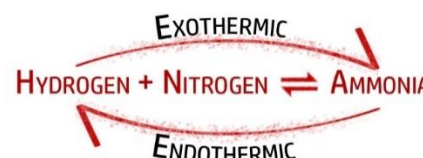


Remember, for something to be at dynamic equilibrium, it must be in a **closed system** – this means that **no** products and reactants should be allowed to escape!

E. The factors affecting equilibrium in the Haber Process:

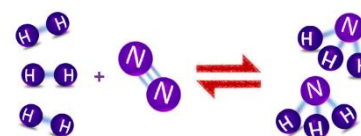
1. Temperature: 450°C

- ✧ Every reversible reaction has an endothermic and an exothermic direction.
- ✧ Increasing the temperature favours the **endothermic direction**.
- ✧ In the Haber Process, the **backwards reaction** is endothermic – which means increasing the temperature will shift the equilibrium to the left, decreasing the yield (amount) of ammonia.
- ✧ If the temperature is too low, the ammonia will form too slowly – so 450°C is a **compromise** between **speed** of reaction and **yield** of ammonia produced.



2. Pressure/Concentration: 200atm

- ✧ When you increase pressure/concentration, the particles will be closer together, meaning more collisions and more reactions.
- ✧ By increasing the pressure/concentration favours the side of the reaction with the **least molecules**.
- ✧ In the Haber process, there are four molecules of reactants – 3xH₂ and 1xN₂, whilst on the right there are only 2 NH₃ molecules. This means that if you increase the pressure there will be **more ammonia**.
- ✧ 200atm is a **compromise** between **yield** of ammonia and **cost** of the equipment.



3. Catalyst: Iron

- ✧ Catalysts are chemicals that speed up the rate of a reaction, but don't get used up.
- ✧ This means that they can be used continuously.
- ✧ **Iron** is the catalyst used in the Haber Process.
- ✧ It does not change the position of the equilibrium, but it does speed up the production – so used producing ammonia is quicker.

Apply Task 1: States of Matter & Mixtures

Question:		Marks:															
1:	i. Draw the particle model for a solid, liquid and gas. (3) ii. Describe the movement of solids, liquid and gas particles (2)	5															
2:	i. Draw the melting point curve for a solid and label where melting and evaporating are occurring. (3) ii. Explain, in terms of particles, why the melting point curve is flat when melting is occurring (2)	5															
3:	i. Describe the difference between a pure element and a mixture. (2) ii. Draw the melting point curve for a mixture and explain why it is different to a pure solid. (3)	5															
4:	Which of the following substances will be a solid at 50°C, and a gas at 200°C <table border="1"> <thead> <tr> <th>Substance:</th><th>Melting point (°C)</th><th>Boiling point (°C)</th></tr> </thead> <tbody> <tr> <td>A</td><td>30</td><td>250</td></tr> <tr> <td>B</td><td>60</td><td>180</td></tr> <tr> <td>C</td><td>250</td><td>400</td></tr> <tr> <td>D</td><td>20</td><td>100</td></tr> </tbody> </table>	Substance:	Melting point (°C)	Boiling point (°C)	A	30	250	B	60	180	C	250	400	D	20	100	1
Substance:	Melting point (°C)	Boiling point (°C)															
A	30	250															
B	60	180															
C	250	400															
D	20	100															

Red:	Amber:	Green:	Blue:
1-4	5-8	9-12	13-16

Apply Task 2: Separation Techniques

Question:		Marks:
1:	Rock salt contains rock – an insoluble solid – and salt – a soluble solid. i. Explain how to separate the insoluble rock from a solution of rock salt. (3) ii. Explain how to separate the soluble salt from a solution of salt water. (3)	6
2:	A colouring in some sweets was analysed using paper chromatography. One of the dyes in the colouring moved 2 cm up the paper while the solvent moved 8 cm. i. What is the R_f value of this dye? (1) ☐ A 0.25 ☐ B 2 ☐ C 4 ☐ D 6 ii. Describe how to set up a paper chromatogram (4)	5
3:	There are two types of distillation – fractional and simple distillation. i. Describe the difference between simple and fractional distillation (2) ii. Explain how a mixture of 3 liquids could be separated using fractional distillation. The boiling points for the liquids are: X-35°C, Y-70°C and Z-150°C (3)	5
4:	Water needs to be made potable before it is safe to drink. i. Describe how desalination is used to purify water from salt water (3) ii. Describe the steps used to turn lake water into drinking water (3)	6

Red:	Amber:	Green:	Blue:
1-6	7-13	14-17	18-22

Apply Task 3: Acids, Bases & Indicators

Question:		Marks:
1:	An acid can be reacted with a base in a titration experiment. i. The general equation for the reaction of an acid with a base is (1) ☐ A acid + base → alkali + water ☐ B acid + base → salt + carbon dioxide ☐ C acid + base → salt + water + hydrogen ☐ D acid + base → salt + water ii. Write the ionic equation to show what happens to the ions during neutralisation (1) iii. Describe the test for hydrogen (2) iv. Describe the test for carbon dioxide (2)	6
2:	A titration is used to determine the exact volumes of sulfuric acid and sodium hydroxide solution that neutralise each other. i. State the name of an indicator that could be used, and the colour change seen in this titration. (3) • Indicator: _____ • Colour in sodium hydroxide solution: _____ • Colour when neutral: _____ ii. As the pH of the solution of sulfuric acid and sodium hydroxide changes from pH 3 to pH 5, the concentration of H⁺ ions changes. Calculate how much less concentrated the solution is at pH 5. (2)	5
3:	Acids can be strong, dilute, concentrated or weak. i. Describe the difference between a strong and a weak acid (2) ii. Describe the difference between a dilute and concentrated acid (2) iii. Explain how a weak, dilute acid can be altered to increase the speed of a reaction. (2)	6
4:	When sodium carbonate, Na₂CO₃, is added to nitric acid, HNO₃, a neutralisation reaction occurs and sodium nitrate, Na₂(NO₃)₂, is produced. i. Write the word equation for the reaction. (2) ii. Describe what you would see when solid sodium carbonate is added to dilute nitric acid. (1) iii. Write the balanced equation for the reaction. (3)	6

Red:	Amber:	Green:	Blue:
1-6	7-12	13-18	19-23

Apply Task 4: Preparing Salts

Question:		Marks:
1:	When calcium hydroxide is added to hydrochloric acid, calcium chloride and water are formed. i. Describe an experiment to investigate how changing the mass of copper hydroxide added to hydrochloric acid affects the pH of the solution. (4) ii. Describe a hazard from the experiment and suggest how to minimise the risk (2)	6
2:	Copper sulfate is a soluble salt. Copper sulfate can be prepared by reacting copper oxide with dilute sulfuric acid. Copper oxide is an insoluble solid. Describe how you would prepare some pure, dry crystals of copper sulfate by reacting excess copper hydroxide with dilute sulfuric acid.	6
3:	Sodium nitrate is a soluble salt. Sodium nitrate can be prepared by reacting soluble sodium hydroxide with nitric acid. Describe how you would prepare some pure, dry crystals of sodium nitrate by reacting sodium hydroxide with nitric acid.	6
4:	When sodium hydroxide and copper nitrate react together, a precipitate forms. i. Write the word equation for the reaction. (1) ii. Name the precipitate formed. (1) iii. Explain how to produce a pure, dry precipitate from soluble copper nitrate and sodium hydroxide. (4)	6

Red:	Amber:	Green:	Blue:
1-6	7-12	13-18	19-24

Apply Task 5: Electrolysis

Question:	Marks:
Electrolysis of hydrochloric acid produces chlorine and hydrogen. 1: i. Explain what is meant by electrolysis. (2) ii. Explain what an electrolyte is. (2) iii. Describe the test to show that a gas is chlorine.(2)	6
When copper sulfate solution is electrolysed using inert electrodes, oxygen is formed at the positively charged anode and copper is formed at the negative cathode. 2: i. Explain what happens to the copper ions at the cathode (2) ii. Name the cations and anions present in copper sulfate solution? (2) iii. What forms at the anode and cathode for molten sodium chloride, NaCl? (2)	6
Sodium chloride solution is electrolysed using a direct electric current. 3: i. Which of these ions will be attracted to the cathode during the electrolysis of sodium chloride solution? (1) ☐ A H⁺ ions only ☐ B H⁺ and Na⁺ ions ☐ C Cl⁻ ions only ☐ D Cl⁻ and OH⁻ ions ii. Chlorine is one of the products of the electrolysis. The half-equation for the production of chlorine is $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ Explain whether the chloride ion has been oxidised or reduced. (2) iii. The electrolysis of sodium chloride solution does not produce metallic sodium. State what change you would make to the electrolyte to obtain metallic sodium. (1)	4
Copper sulfate solution was electrolysed using copper electrodes. Before the electrolysis, the masses of the electrodes were determined. After the electrolysis, the anode had decreased by 0.85g, whilst the cathode had increased by 0.79g. 4: i. Explain, in terms of electrons, what happens to the copper ions during this investigation. (4) ii. Explain why the change in mass wasn't the same for both the anode and cathode (2)	6

Red:	Amber:	Green:	Blue:
1-5	6-12	13-17	18-22

Apply Task 6: Extracting Metals and Reactivity

Question:		Marks:
1:	i. Describe how sodium reacts with hydrochloric acid and with cold water. (2) ii. Describe how magnesium reacts with hydrochloric acid and with cold water. (2) iii. Describe how copper reacts with hydrochloric acid and with cold water. (2)	6
2:	When magnesium is added to copper sulfate, CuSO_4, a displacement reaction occurs and magnesium sulfate, MgSO_4, is produced. i. Write the word equation for the displacement reaction (1) ii. Write the balanced equation for the displacement reaction (2) (H) iii. Write the ionic equation for the reaction. (2) (H) iv. Use the answers from part i-iii to explain why this is a REDOX reaction (2)	7
3:	Metals are obtained from the Earth's crust by different methods. Some metals are found uncombined, but others have to be extracted from their ores by electrolysis or by heating the ore with carbon. Explain, using aluminium, gold and iron as examples, how the method used to obtain the metal is related to its position in the reactivity series and to the cost of the extraction process.	6
4:	Two other methods of extracting ores are bioleaching and phytoextraction. Evaluate the two methods suggesting advantages and disadvantages for each.	6
5:	Look at the word equation for the extraction of lead from its ore using carbon monoxide: Lead oxide + carbon monoxide $\rightarrow$ lead + carbon dioxide Explain where oxidation and reduction are occurring during this reaction (4)	4

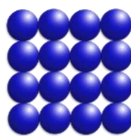
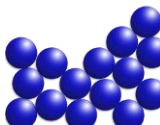
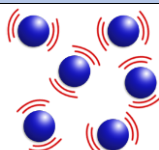
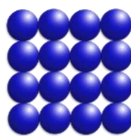
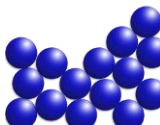
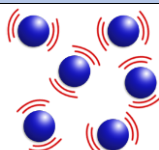
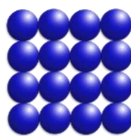
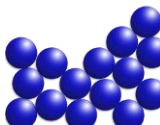
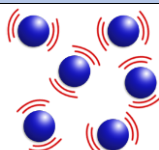
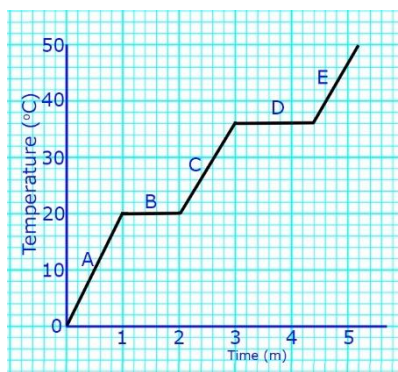
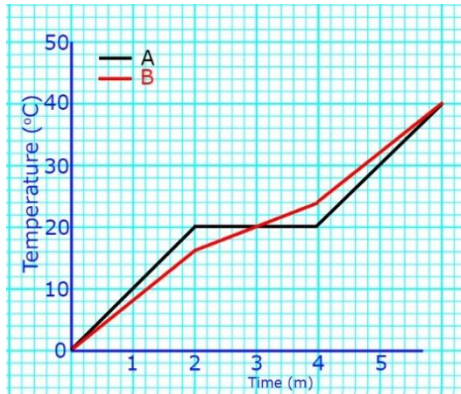
Red:	Amber:	Green:	Blue:
1-7	8-14	15-21	22-29

Apply Task 7: Reversible Reactions and Recycling

Question:		Marks:									
1:	Evaluate the importance of recycling metals, such as lead.	4									
2:	When a new product is designed, you must look at the life cycle assessment for that product. i. Describe the four stages of a life cycle assessment (4) ii. Carrier bags can be made from either paper or plastic. Look at the table below showing information on the use and disposal of both: <table border="1" style="margin-left: 40px;"> <thead> <tr> <th>Material:</th><th>Use:</th><th>Disposal:</th></tr> </thead> <tbody> <tr> <td>Plastic</td><td>Reusable, strong</td><td>Non-Biodegradable</td></tr> <tr> <td>Paper</td><td>Reusable, tears easily</td><td>Biodegradable</td></tr> </tbody> </table> Explain which material is best to use for a carrier bag: (/2)	Material:	Use:	Disposal:	Plastic	Reusable, strong	Non-Biodegradable	Paper	Reusable, tears easily	Biodegradable	6
Material:	Use:	Disposal:									
Plastic	Reusable, strong	Non-Biodegradable									
Paper	Reusable, tears easily	Biodegradable									
3:	When nitrogen and hydrogen react to form ammonia, the reaction can reach a dynamic equilibrium. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ i. The $\rightleftharpoons$ symbol in the word equation shows that the reaction goes forwards and backwards at the same time. Give the name of this type of reaction. (1) ii. Explain what is meant by a dynamic equilibrium. (3)	4									
4:	In industry, the reaction between nitrogen and hydrogen is affected by the conditions used. i. The pressure used is 250 atmospheres. Explain how the use of a higher pressure would affect the equilibrium yield of ammonia. (3) ii. The reaction between nitrogen and hydrogen to form ammonia is exothermic. The temperature used is 450°C. Explain how the use of a lower temperature would affect the equilibrium yield of ammonia (4) iii. Even at 450°C, the reaction is very slow. State what is used in industry to overcome this problem (1)	8									

Red:	Amber:	Green:	Blue:
1-5	6-10	11-15	16-22

Check Task 1: States of Matter & Mixtures

Question:																	
1:	i. Draw the particle model for a solid, liquid and gas. (3) <table><tr><td></td><td>Solid:</td><td>Liquid:</td><td>Gas:</td></tr><tr><td>Particle Model:</td><td></td><td></td><td></td></tr></table> ✱ MP1: Solid particles touching and regular pattern / rows (1) ✱ MP2: Liquid particles touching and random pattern (1) ✱ MP3: Gas particles not touching / far apart and motion lines (1) ii. Describe the movement of solids, liquid and gas particles (2) Any 2 from: ✱ Solid: Vibrating about a fixed position (1) ✱ Liquid: Flowing / moving randomly (1) ✱ Gas: Moving fast in {all/random} directions (1)		Solid:	Liquid:	Gas:	Particle Model:				5							
	Solid:	Liquid:	Gas:														
Particle Model:																	
2:	i. Draw the melting point curve for a solid and label where melting and evaporating are occurring. (3) ✱ Graph like one on right (1) ✱ Melting point marked at B (1) ✱ Evaporation labelled at D (1) ii. Explain, in terms of particles, why the melting point curve is flat when melting is occurring (2) ✱ Energy not used to raise temperature (1) ✱ Energy used to break {intermolecular forces / bonds} (1) 	5															
3:	i. Describe the difference between a pure element and a mixture. (2) ✱ Pure element: Only one {substance/type of atom} (1) ✱ Mixture: More than one type of {substance / atom / compound} not bonded (1) ii. Draw the melting point curve for a mixture and explain why it is different to a pure solid. (3) ✱ Graph drawn like B on the right (1) ✱ “melting line” diagonal, not flat (1) ✱ More than one melting point / melts over different temperatures (1) 	5															
4:	Which of the following substances will be a solid at 50°C, and a gas at 200°C <table><tr><td>Substance:</td><td>Melting point (°C)</td><td>Boiling point (°C)</td></tr><tr><td>A</td><td>30</td><td>250</td></tr><tr><td>B</td><td>60</td><td>180</td></tr><tr><td>C</td><td>250</td><td>400</td></tr><tr><td>D</td><td>20</td><td>100</td></tr></table>	Substance:	Melting point (°C)	Boiling point (°C)	A	30	250	B	60	180	C	250	400	D	20	100	1
Substance:	Melting point (°C)	Boiling point (°C)															
A	30	250															
B	60	180															
C	250	400															
D	20	100															

Check Task 2: Separation Techniques

Question:		
1:	Rock salt contains rock – an insoluble solid – and salt – a soluble solid. i. Explain how to separate the insoluble rock from a solution of rock salt. (3) * Filtration (1) * Add the mixture to filter paper (1) * The solid will stay in the filter paper (1) * The solution will go into the conical flask (1) ii. Explain how to separate the soluble salt from a solution of salt water. (3) * Heat / evaporate / crystallisation (1) * Evaporate half of the liquid (1) * Leave it to cool (1) * Dry the crystals between filter paper (1)	6
2:	A colouring in some sweets was analysed using paper chromatography. One of the dyes in the colouring moved 2 cm up the paper while the solvent moved 8 cm. i. What is the R_f value of this dye? (1) ☒ A 0.25 ii. Describe how to set up a paper chromatogram (4) * Step 1: Draw a line (and cross) in pencil (1) – which is insoluble (1) * Step 2: Add the ink to the cross (1) * Step 3: Add the paper to the chromatography column (1) making sure the ink does not touch the water (1) * Step 4: Remove the paper when the water is near the top (1)	5
3:	There are two types of distillation – fractional and simple distillation. i. Describe the difference between simple and fractional distillation (2) * Simple = separates less liquids (1) * Fractional = separates many liquids (1) using a temperature gradient (1) ii. Explain how a mixture of 3 liquids could be separated using fractional distillation. The boiling points for the liquids are: X-35°C, Y-70°C and Z-150°C (3) * Step 1: Heat the mixture (1) * Step 2: When the top of the fractionating column gets to 35°C, X will move into the condenser (1) * Step 3: X will condense and turn back into a liquid (1) * Step 4: Continue to heat until the column gets to 70°C (1)	5
4:	Water needs to be made potable before it is safe to drink. i. Describe how desalination is used to purify water from salt water (3) * Heat the salt water / evaporate the water (1) * The {gas/steam/water vapour} will hit the lid and condense (1) * The liquid is collected (1) ii. Describe the steps used to turn lake water into drinking water (3) * Step 1: Sieving – large pieces removed (1) * Step 2: Sedimentation – small insoluble pieces settle at the bottom (1) * Step 3: Filtration – {sand/gravel} used to filter out small pieces (1) * Step 4: Chlorination – chlorine kills bacteria (1)	6

Check Task 3: Acids, Bases & Indicators

Question:		
1:	An acid can be reacted with a base in a titration experiment. i. The general equation for the reaction of an acid with a base is (1) ☐ D acid + base → salt + water ii. Write the ionic equation to show what happens to the ions during neutralisation (1) * $H^+ + OH^- \rightarrow H_2O$ iii. Describe the test for hydrogen (2) * Lit splint (1)...squeaky pop (1) iv. Describe the test for carbon dioxide (2) * (Bubble through) limewater (1) which goes cloudy (1)	6
2:	A titration is used to determine the exact volumes of sulfuric acid and sodium hydroxide solution that neutralise each other. i. State the name of an indicator that could be used, and the colour change seen in this titration. (3) - Indicator: <u>Methyl orange / phenolphthalein / litmus (1)</u> - Colour in sodium hydroxide solution: <u>Yellow / pink / blue (1)</u> - Colour when neutral: <u>Orange / colourless / purple (1)</u> ii. As the pH of the solution of sulfuric acid and sodium hydroxide changes from pH 3 to pH 5, the concentration of H^+ ions changes. Calculate how much less concentrated the solution is at pH 5. (2) * pH change: $2 = 10 \times 10$ (1) * = 100 times less concentrated (1) * 100 times gets both marks (1)	5
3:	Acids can be strong, dilute, concentrated or weak. i. Describe the difference between a strong and a weak acid (2) * Strong acids ionise fully into H^+ ions (1) * Weak acids only ionise partially (1) ii. Describe the difference between a dilute and concentrated acid (2) * Concentrated = lots of H^+ ions * Dilute = less H^+ ions iii. Explain how a weak, dilute acid can be altered to increase the speed of a reaction. (2) * Make it more concentrated (1) * Evaporate some of the water (1) * Add more H^+ ions in the same volume (1)	6
4:	When sodium carbonate, Na_2CO_3, is added to nitric acid, HNO_3, a neutralisation reaction occurs and sodium nitrate, $NaNO_3$, is produced. i. Write the word equation for the reaction. (2) * Sodium carbonate + nitric acid → sodium nitrate (1) + carbon dioxide + water (1) ii. Describe what you would see when solid sodium carbonate is added to dilute nitric acid. (1). Any one from: * Bubbles (1) * Solid {dissolving / disappearing} (1) iii. Write the balanced equation for the reaction. (3) * $Na_2CO_3 + 2HNO_3 \rightarrow 2NaNO_3 + CO_2 + H_2O$ * LHS (1); RHS (1); Balancing (1)	6

Check Task 4: Preparing Salts

Question:		
1:	When calcium hydroxide is added to hydrochloric acid, calcium chloride and water are formed. i. Describe an experiment to investigate how changing the mass of copper hydroxide added to hydrochloric acid affects the pH of the solution. (4) * Measure out hydrochloric acid (1) * Measure the pH (with pH paper) (1) * Add {0.3g/specific mass} of copper hydroxide (1) * Remeasure the pH (1) * Keep adding more solid and measuring the pH (1) ii. Describe a hazard from the experiment and suggest how to minimise the risk (2) * Hazard: Acid is an irritant (1) * Risk: {If on hands wash off / wear goggles to protect eyes / wear gloves} (1)	6
2:	Copper sulfate is a soluble salt. Copper sulfate can be prepared by reacting copper oxide with dilute sulfuric acid. Copper oxide is an insoluble solid. Describe how you would prepare some pure, dry crystals of copper sulfate by reacting excess copper hydroxide with dilute sulfuric acid. * Step 1: Add copper hydroxide (1) until the solution is neutral (1) * Step 2: Filter the mixture (1) * Step 3: Heat the solution / evaporate the water (1) * Step 4: Stop heating when about half of the solution has evaporated (1) * Step 5: Leave the solution to cool (1) * Step 6: Dry the crystals using filter paper (1)	6
3:	Sodium nitrate is a soluble salt. Sodium nitrate can be prepared by reacting soluble sodium hydroxide with nitric acid. Describe how you would prepare some pure, dry crystals of sodium nitrate by reacting sodium hydroxide with nitric acid. * Step 1: Add sodium hydroxide to a conical flask (1) * Step 2: Add phenolphthalein indicator – it will go pink (1) * Step 3: Add hydrochloric acid to a burette (1) * Step 4: Add the acid until the indicator goes colourless (1) * Step 5: Repeat without the indicator (1) so that it is pure * Step 6: Heat the solution / evaporate the water / crystallise (1) * Step 7: Leave the solution to cool (1) * Step 6: Dry the crystals using filter paper (1)	6
4:	When sodium hydroxide and copper nitrate react together, a precipitate forms. i. Write the word equation for the reaction. (1) * Sodium hydroxide + copper nitrate → sodium nitrate + copper hydroxide (1) ii. Name the precipitate formed. (1) * Copper hydroxide (1) iii. Explain how to produce a pure, dry precipitate from soluble copper nitrate and sodium hydroxide. (4) * Dissolve (1) * Mix (1) * Filter (1) * Wash (1) * Dry (1)	6

Check Task 5: Electrolysis

Question:		
1:	Electrolysis of hydrochloric acid produces chlorine and hydrogen. i. Explain what is meant by electrolysis. (2) ✳ Breaking down of {an electrolyte / a substance} (1) ✳ Using (electricity / direct current) (1) ii. Explain what an electrolyte is. (2) ✳ A liquid (1) ✳ containing ions (1) iii. Describe the test to show that a gas is chlorine.(2) ✳ Blue Litmus paper (1)... ✳ ...(turns red then) bleaches (1)	6
2:	When copper sulfate solution is electrolysed using inert electrodes, oxygen is formed at the positively charged anode and copper is formed at the negative cathode. i. Explain what happens to the copper ions at the cathode (2) ✳ Copper ions gain (2) electrons (1) ✳ Copper ions are reduced (1) ✳ Copper ions turn into copper atoms (1) ✳ $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ (2 marks) ii. Name the cations and anions present in copper sulfate solution? (2) ✳ Cations: {Hydrogen / H^+} and {Copper / Cu^{2+}} (1) ✳ Anions: {Hydroxide / OH^-} and {Sulfate / SO_4^{2-}} (1) iii. What forms at the anode and cathode for molten sodium chloride, NaCl? (2) ✳ Anode: Chlorine / Cl_2 (1) ✳ Cathode: Sodium / Na (1)	6
3:	Sodium chloride solution is electrolysed using a direct electric current. i. Which of these ions will be attracted to the cathode during the electrolysis of sodium chloride solution? (1) ☒ H^+ and Na^+ ions ii. Chlorine is one of the products of the electrolysis. The half-equation for the production of chlorine is $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$ Explain whether the chloride ion has been oxidised or reduced. (2) ✳ Oxidised (1) ✳ Lost electrons (1) iii. The electrolysis of sodium chloride solution does not produce metallic sodium. State what change you would make to the electrolyte to obtain metallic sodium. (1) ✳ Melt it (1)	4
4:	Copper sulfate solution was electrolysed using copper electrodes. Before the electrolysis, the masses of the electrodes were determined. After the electrolysis, the anode had decreased by 0.85g, whilst the cathode had increased by 0.79g. i. Explain, in terms of electrons, what happens to the copper ions during this investigation. (4) ✳ Anode: Cu atoms turn into Cu^{2+} (1) by losing 2 electrons (1) ✳ The Cu^{2+} ions move into the solution, decreasing the mass of the anode (1) ✳ Cathode: Cu^{2+} ions turn into Cu atoms (1) by gaining 2 electrons (1) and increasing the mass of the cathode (1) ii. Explain why the change in mass wasn't the same for both the anode and cathode (2) ✳ Sludge / impurities from the anode (1) get left in the solution (1)	6

Check Task 6: Extracting Metals and Reactivity

Question:		
1:	i. Describe how sodium reacts with hydrochloric acid and with cold water. (2) * {Violently/fast} with acids (1) * {React with cold water / Form hydrogen and a hydroxide} (1) ii. Describe how magnesium reacts with hydrochloric acid and with cold water. (2) * React {slower/less vigorously} with acids (1) * {React slowly with cold water / Form hydrogen and an oxide} (1) iii. Describe how copper reacts with hydrochloric acid and with cold water. (2) * Doesn't react with acids (1) * Doesn't react cold water (1)	6
2:	When magnesium is added to copper sulfate, CuSO_4, a displacement reaction occurs and magnesium sulfate, MgSO_4, is produced. i. Write the word equation for the displacement reaction (1) * Magnesium + copper sulfate → copper + magnesium sulfate (1) ii. Write the balanced equation for the displacement reaction (2) * $\text{Mg} + \text{CuSO}_4 (1) \rightarrow \text{Cu} + \text{MgSO}_4 (1)$ (H) iii. Write the ionic equation for the reaction. (2) * $\text{Mg}_{(s)} + \text{Cu}^{2+}_{(aq)} (1) \rightarrow \text{Cu}_{(s)} + \text{Mg}^{2+}_{(aq)} (1)$ (H) iv. Use the answers from part i-iii to explain why this is a REDOX reaction (2) * Magnesium has lost electrons (1) and has been oxidised (1) * Copper ions have gained electrons (1) and been reduced (1) * Oxidation and reduction are occurring (1)	7
3:	Metals are obtained from the Earth's crust by different methods. Some metals are found uncombined, but others have to be extracted from their ores by electrolysis or by heating the ore with carbon. Explain, using aluminium, gold and iron as examples, how the method used to obtain the metal is related to its position in the reactivity series and to the cost of the extraction process. * Aluminium is more reactive than carbon (1) * ...so, you can't heat with carbon (1) * You must use electrolysis to extract aluminium (1) * Iron is less reactive than carbon (1) * ...so, you heat with carbon (1) * You don't use electrolysis because it is too expensive (1) * Gold is unreactive (1) * ...so, you can just dig it out of the ground (1)	6
4:	Two other methods of extracting ores are bioleaching and phytoextraction. Evaluate the two methods suggesting advantages and disadvantages for each. Both: Max of two from: * + Don't produce harmful gases (1) * + Less damage to the landscape than mining (1) * + Conserves supplies of ores (1) * - very slow Bioleaching: Max of two from: * + Doesn't require high temperatures (1) * - Toxic substances and sulphuric acid can be produced (1) Phytoextraction: Max of two from: * + Can extract metals from contaminated soil (1) * - Can be more expensive than mining ores (1) * - Growing plants depends on the weather (1)	6
3:	Look at the word equation for the extraction of lead from its ore using carbon monoxide: Lead oxide + carbon monoxide → lead + carbon dioxide Explain where oxidation and reduction are occurring during this reaction (4) * Lead oxide has {had oxygen removed / lost oxygen} (1) * Lead oxide has been reduced (1) * Carbon monoxide has gained oxygen (1) * Carbon monoxide has been oxidised (1)	4

Check Task 7: Reversible Reactions and Recycling

Question:											
1:	Evaluate the importance of recycling metals, such as lead. Any 4 from: * Conserves Earth's natural resources (1) * Less mining of ores which means it doesn't damage the landscape (1) * Less mining of ores which means it doesn't create dust/noise pollution (1) * Less energy than extracting more (1) * Doesn't end up in landfills (1)	4									
2:	When a new product is designed, you must look at the life cycle assessment for that product. i. Describe the four stages of a life cycle assessment (4) * Stage 1: Choice of Materials (1) * Stage 2: Manufacture of Product (1) * Stage 3: Use of Product (1) * Stage 4: Disposal of Product (1) ii. Carrier bags can be made from either paper or plastic. Look at the table below showing information on the use and disposal of both: <table border="1" style="margin-left: auto; margin-right: auto;"> <thead> <tr> <th>Material:</th><th>Use:</th><th>Disposal:</th></tr> </thead> <tbody> <tr> <td>Plastic</td><td>Reusable, strong</td><td>Non-Biodegradable</td></tr> <tr> <td>Paper</td><td>Reusable, tears easily</td><td>Biodegradable</td></tr> </tbody> </table> Explain which material is best to use for a carrier bag: (/2) Either: * Paper because it {will break down / is biodegradable} (1) * Plastic is non-biodegradable (1) OR * Plastic because it {is strong / won't break} (1) * Paper will tear easily (1)	Material:	Use:	Disposal:	Plastic	Reusable, strong	Non-Biodegradable	Paper	Reusable, tears easily	Biodegradable	6
Material:	Use:	Disposal:									
Plastic	Reusable, strong	Non-Biodegradable									
Paper	Reusable, tears easily	Biodegradable									
3:	When nitrogen and hydrogen react to form ammonia, the reaction can reach a dynamic equilibrium. $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ i. The $\rightleftharpoons$ symbol in the word equation shows that the reaction goes forwards and backwards at the same time. Give the name of this type of reaction. (1) * Reversible (1) ii. Explain what is meant by a dynamic equilibrium. (3) * forward and back reactions take place (at the same time) (1) * rate of the forward and back reactions is the same (1) * no (overall) change in the amount/concentration/mass/ volume of each substance / reactant / product (1) * no observable change (1)	4									
4:	In industry, the reaction between nitrogen and hydrogen is affected by the conditions used. i. The pressure used is 250 atmospheres. Explain how the use of a higher pressure would affect the equilibrium yield of ammonia. (3) * higher pressure favours forward reaction/equilibrium shifts to the right (1) * because decrease in volume / number of molecules/sides with lower volume (1) * yield increases (1) ii. The reaction between nitrogen and hydrogen to form ammonia is exothermic. The temperature used is 450°C. Explain how the use of a lower temperature would affect the equilibrium yield of ammonia (4) * lower temperature favours forward reaction (1) * equilibrium shifts to the right (1) * because (forward) reaction is exothermic (1) * yield increases (1) iii. Even at 450°C, the reaction is very slow. State what is used in industry to overcome this problem (1) * Iron / catalyst (1)	8									