

Solids, liquids & gases

Solids - Particles arranged regularly and packed closely together. Vibrate in fixed positions. Strong forces between particles.

Liquids - Particles are mostly touching with some gaps. Particles move about at random. Medium forces between particles.

Gases - Particles move at random and quickly. Particles are far apart. Weak forces between particles.

What is melting?

- A solid is heated and turns to liquid.
- Particles have more kinetic energy.
- Substance expands as particles vibrate more.
- Particles can break from fixed positions to move more freely.

What is boiling?

- A liquid when heated turns to gas.
- Particles have more kinetic energy.
- Particles break away from position to escape from liquid.

What is evaporation?

- Particles in liquid at different temperatures.
- Some liquid turns to gas.
- Some particles have enough kinetic energy to turn to gas.

What is freezing?

- The opposite of melting.
- Liquid turns solid.
- Particles lose kinetic energy and move more slowly.
- Particles get closer together and forces of attraction take over.

What is condensation?

- The opposite of evaporation.
- Gas turns to liquid.
- Particles lose energy and come closer together.
- Forces of attraction take over and particles stay together.

What is sublimation?

The transition of a substance directly from solid phase to the gas phase without passing through liquid phase

Describe the molecular structure of a gas

- Molecules are almost totally free of attraction to each other.
- Move freely at high speed colliding with each other and the edge of the container they're in.
- A lack of attraction and movement is why a gas changes shape and volume.

How does a change in temperature effect the movement of molecules in a gas?

- Increased temperature = molecules move faster with more energy.
- Decreased temperature = molecules move slower with less energy.

If a gas is heated in a container with a fixed volume, why will the pressure rise?

- Temperature rises → particles move with more kinetic energy.
- More collisions with container walls.
- Increased collisions against the container walls = increased pressure
- Pressure is directly proportional to the temperature (in kelvin)

What is kinetic theory?

- Used to explain the properties of solids, liquids and gasses
- All matter is made up of tiny particles called atoms
- Particles are constantly moving
- Particles attract each other, with weaker attraction when further apart

What happens when you compress a gas?

- Its pressure increases
- Particles in smaller space hit/collide with walls more often

How will changing volume affect the pressure of a gas at a constant temperature?

- Governed by Boyle's Law
- Pressure is inversely/opposite proportional to volume (if temperature is constant)
- If volume halves, pressure doubles
- If volume doubles, pressure halves

Diffusion

Diffusion - Net movement of particles from an area of high concentration to an area of low concentration.

- Diffusion doesn't require any energy, so it's an passive process
- The smaller the Mr the faster the diffusion occurs

Atoms, elements and compounds

Atoms - Smallest particle of a substance that can exist

Element - contains only one type of atom, can't be split by any chemical means

Compound - Two or more elements chemically combined(Unable to separate)

Mixture - Contains 2 or more elements NOT chemically combined(able to separate)

Atomic structure and the periodic table

Molecule - Two or more atoms bonded together

- Nucleus contains protons and neutrons
- Shells contain electrons

- Mass of protons = 1
- Mass of neutrons = 1
- Mass of electrons = 1/2000

Charges:

- Neutron = 0 (neutral)
- Proton = +1 (positive)
- Electron = -1 (negative)

Atomic number - Number of protons in an atom

Mass number - total number of proton + neutron number

Nucleon number - total number of protons and neutrons in the nucleus

Group number - number of electrons in the outer shell

Period number - number of shells of electrons

- Elements in the same group have the same chemical properties because they have the same number of outer shell electrons

Group 0 - Noble gasses

- Noble gasses are unreactive because they have full outer shell of electrons

Isotopes

2 chlorine isotopes:

- 37 - Cl
- 35 - Cl

Isotope - Atoms of an element with the same number of protons/electron/atomic number but different number of neutrons/atomic mass

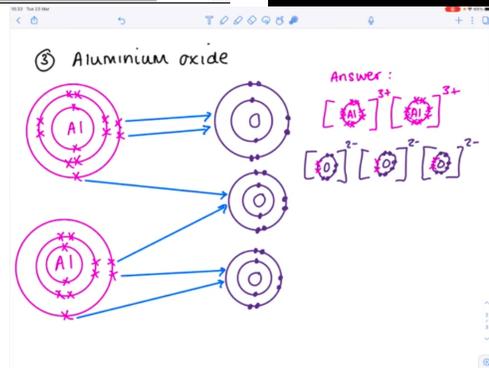
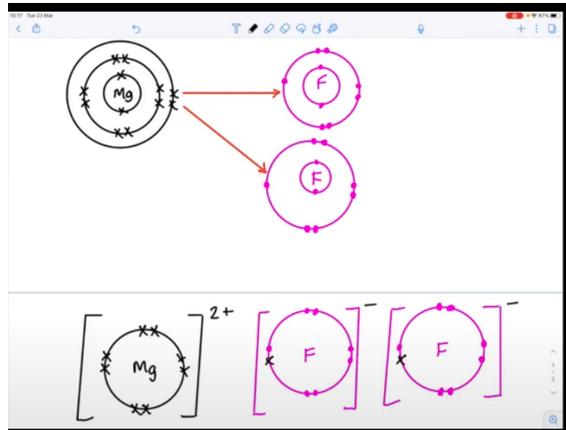
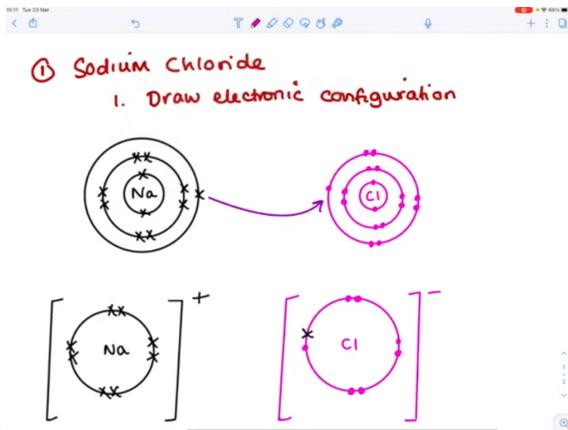
Give two uses of radioactive isotopes

- 1) Carbon dating using carbon- 14
- 1) Carbon- 14 taken in during life
- 1) Radioactive decay shows age of artefacts
-
- 2) Treating cancer (radiotherapy) with cobalt- 60
- 2) Radiation from the isotope kills cancer cells faster than healthy ones

Ionic Bonding

Ion - Charged particle formed by atoms either losing or gaining electrons

- Losing electrons gives positive charge
- Gaining electrons gives negative charge
- Ionic bonding contains a metal and a nonmetal
- Metals always forms positive ions
- Non-metals always forms negative ions

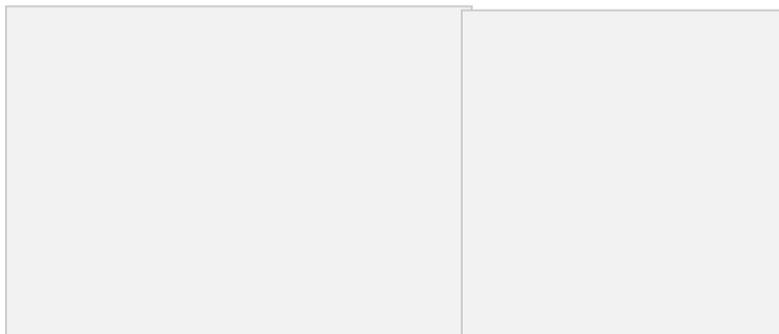


Covalent bonding

- Covalent bonding contains 2 nonmetals

Covalent bond - Shared pair of electrons

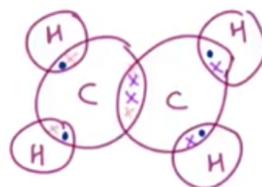
- Electrostatic attraction between positive nucleus and shared pair of negative electrons



carbon dioxide CO_2



Ethene C_2H_4



Giant & simple chemical structures

Ionic bond - Electrostatic attraction between oppositely charged ions (metal + nonmetal)

- Giant ionic structures have high **melting** and **boiling** point
- Strong electrostatic forces attraction between oppositely charged ions
- Requires a lot of energy to break
- Giant ionic structure do not conduct electricity when in solid, because **ions are not free to move**
- Giant ionic structures do conduct electricity when molten because **ions are free to move**

Brittle -

- break apart when hit
- Layers of ions slide so that ions with the same charge end up next to each other

- Like charges repel
- Structure breaks apart

Allotrope - Different forms of the same element

Why does diamond have such a high melting point?

- Giant tetrahedral structure
- Each carbon atom is bonded with 4 others
- Many strong covalent bonds
- Requires a lot of energy to break

Why does graphite have such a high melting point?

- Each carbon atom is bonded to 3 others
- Many strong covalent bonds
- Requires a lot of energy to break

- **Because graphite is bonded with 3 carbon atoms and diamond is bonded with 4 carbon atoms, that's why graphite has a slightly lower melting point than diamond**

Why is graphite used as a lubricant?

- Carbon atoms are arranged in layers with weak intermolecular forces between them
- Requires little energy to break

Intermolecular - the attractive and repulsive forces that arise between the molecules of a substance

Why doesn't diamond conduct electricity?

- Diamond has no free electrons

Why does graphite conduct electricity?

- Each carbon atom is bonded to 3 others, meaning that there is a delocalised (the 4th electron is free to move) electron

Describe the structure of silicon (IV) oxide (silicon dioxide)

- Each silicon atom bonds covalently to 4 oxygen atoms
- Each oxygen atom bonds to 2 silicon atoms
- Giant structure

Why do simple molecular substances have such low melting points?

- weak intermolecular forces do not require a lot of energy to break

Why do simple molecular substances have increasing boiling point with increasing Mr?

- Boiling breaks the intermolecular forces of attraction between molecules
- Substances with greater Mr have greater intermolecular forces of attraction which need breaking
- Therefore more heat energy is needed to overcome these forces

Metallic bonding

Metallic bond - Electrostatic attraction between positive metal ions and the sea of delocalised electrons

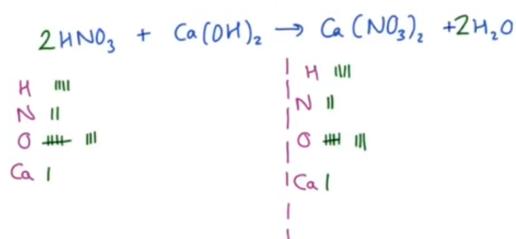
- Delocalised electrons can pass on thermal energy easily
- The sea of delocalised electrons can help carry the current

Malleable - Can be hammered into shape

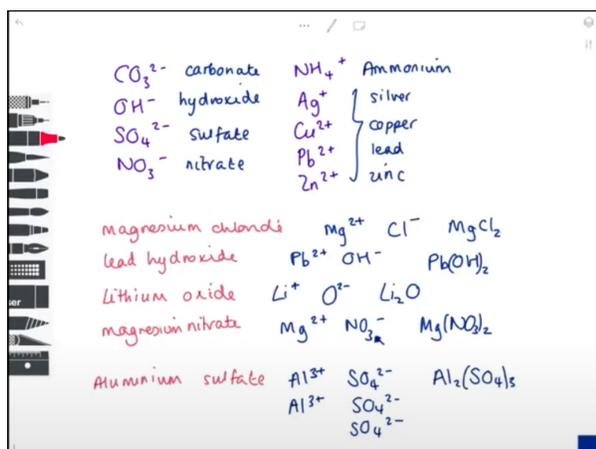
Ductile - Can be drawn into a wire

- Layers of ions can slide over each other, which why they are malleable and ductile

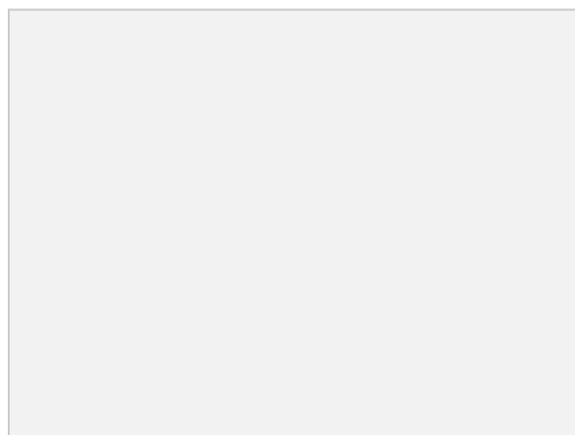
Balancing equations (stoichiometry)



Writing the formulae of common compounds



CO_3^{2-} carbonate	NH_4^+ Ammonium
OH^- hydroxide	Ag^+ silver
SO_4^{2-} sulfate	Cu^{2+} copper
NO_3^- nitrate	Pb^{2+} lead
	Zn^{2+} zinc
magnesium chloride	$\text{Mg}^{2+} \text{Cl}^- \text{MgCl}_2$
lead hydroxide	$\text{Pb}^{2+} \text{OH}^- \text{Pb}(\text{OH})_2$
Lithium oxide	$\text{Li}^+ \text{O}^{2-} \text{Li}_2\text{O}$
magnesium nitrate	$\text{Mg}^{2+} \text{NO}_3^- \text{Mg}(\text{NO}_3)_2$
Aluminium sulfate	$\text{Al}^{3+} \text{SO}_4^{2-} \text{Al}_2(\text{SO}_4)_3$
	$\text{Al}^{3+} \text{SO}_4^{2-}$
	SO_4^{2-}



Relative atomic mass:

- The ratio of the average mass of an element when compared with 1 atom of carbon 12

Mole calculations

A triangle diagram with 'mass' at the top and 'Mr x moles' at the bottom.

$$\text{mass} = \text{Mr} \times \text{moles}$$
$$\text{moles} = \frac{\text{mass}}{\text{Mr}}$$

1. Find the Mr of Ca(OH)_2

$$= 40 + (16 \times 2) + (1 \times 2)$$
$$= 74$$

2. Find the number of moles in 5.4g of CaCO_3

$$\text{moles} = \frac{\text{mass}}{\text{Mr}}$$
$$= \frac{5.4}{(40 + 12 + (16 \times 3))}$$
$$= \frac{5.4}{100} = 0.054$$

Empirical formulae - Simplest ratio of atoms of each element present in a compound

Molecular formulae - Actual number of atoms of each element present in a compound

Empirical formulae: simplest ratio of atoms of each element present in a compound.

Molecular formulae: actual number of atoms of each element present in a compound.

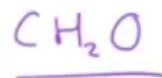
CH_2

C_3H_6

1. A compound contains 40% carbon, 6.73% hydrogen + 53.27% oxygen by mass, determine the empirical formulae.

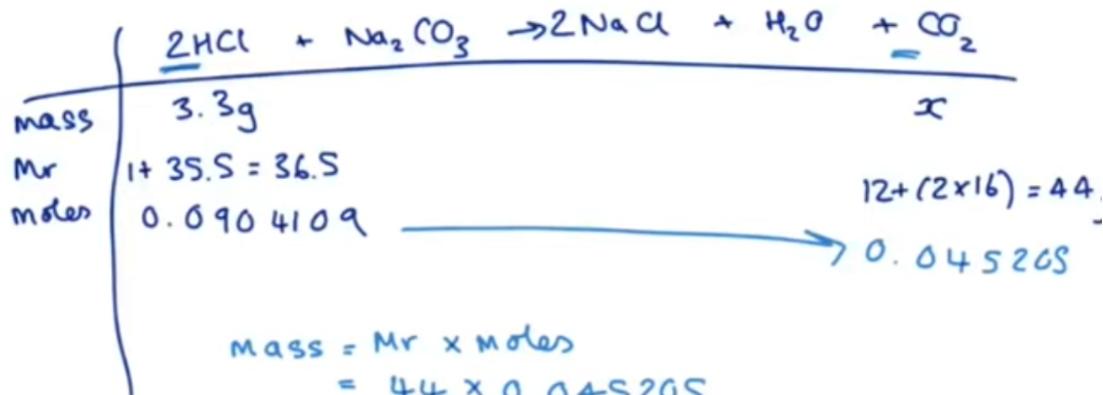
	C	H	O
Mass	40	6.73	53.27
Mr	12.01	1.01	16.00
n	<u>3.33</u>	<u>6.66</u>	<u>3.33</u>
∴ by smallest number	3.33	3.33	3.33

1 : 2 : 1



$$n = \frac{\text{mass}}{\text{Mr}}$$

3.3g HCl reacted with Na_2CO_3 . Calculate the mass + volume of CO_2 collected.



$$\begin{aligned} \text{mass} &= \text{Mr} \times \text{moles} \\ &= 44 \times 0.045205 \\ &= 1.98904 \\ &= 2.00 \text{ g} \end{aligned}$$

1 mol of any gas occupies 24 dm^3
 $0.045205 \times 24 = \underline{1.08 \text{ dm}^3}$



dm^3 to $\text{cm}^3 = \times 1000$ (times by 1000)

acid is in excess

In a reaction, 11.2g of Copper sulfate was obtained when theoretically 12.5g should have been obtained. calculate the percentage yield.

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

$$\begin{aligned} \% \text{ yield} &= \frac{11.2}{12.5} \times 100 \\ &= 89.6\% \end{aligned}$$

Titration (concentration) calculations

25 cm³ of 2.0 mol dm⁻³ HCl reacted with 30 cm³ of NaOH. Calculate the concentration of NaOH.

	$\text{NaOH} + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{O}$	
moles	0.05	← 0.05
conc.	x	2
volume	30/1000	25/1000



$$\begin{aligned}
 x &= \frac{\text{moles}}{\text{vol}} \\
 &= \frac{0.05}{(30/1000)} \\
 &= 1.67 \\
 &= 1.67 \text{ mol dm}^{-3} \text{ (3 s.f.)}
 \end{aligned}$$

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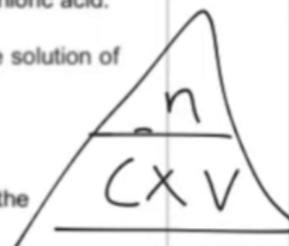
4 (d) A student does a titration to find the concentration of a solution of hydrochloric acid.

The student titrates 25.00 cm³ of hydrochloric acid with sodium hydroxide solution of concentration 0.200 moles per dm³. The equation for the reaction is:



The student added 28.60 cm³ of sodium hydroxide solution to neutralise the hydrochloric acid.

Calculate the concentration of the hydrochloric acid.



	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$		[3 marks]
	5.72×10^{-3}	5.72×10^{-3}	
	x	0.2	$n = C \times V$
	25/1000	28.6/1000	$C = \frac{n}{V}$
	Concentration = 0.229		moles per dm ³

The Avogadro Constant

number of atoms or molecules $\xleftrightarrow[\div \text{Avogadro constant}]{\times \text{Avogadro constant}}$ moles $\xleftrightarrow[\div \text{Mr}]{\times \text{Mr}}$ mass(g)

Avogadro constant = 6.02×10^{23}

1. Calculate the amount in mol of H_2O in a sample of 2.30×10^{23} molecules.

$$n(\text{H}_2\text{O}) = \frac{2.30 \times 10^{23}}{6.02 \times 10^{23}} = 0.382 \text{ mol.}$$

2. Calculate the number of oxygen atom contained in 3 moles of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$

$$\begin{aligned} \text{No. of molecule} &= \text{Moles} \times \text{Avo. constant} \\ &= 3 \times 6.02 \times 10^{23} \\ &= (1.806 \times 10^{24}) \times 6 \\ &= 1.08 \times 10^{25} \end{aligned}$$

Electrolysis

Electrolysis - Breakdown of a molten or aqueous ionic compound by electricity

- Ionic substance must be molten or in solution so that the ions are free to move
- Two electrodes dip into the substance, the electrodes are made out of an inert substance (an unreactive).
- Electrodes are made out of an inert substance e.g. **graphite** or **platinum**

Cation - Positive ion

Anion - Negative ion

- Opposite attracts, so **anions** which are **negative** will be **attracted** to **anode** which is the **positive electrode** and **cation** which are **positive** will be **attracted** to **cathode** which is the **negative electrode**

Anode - Positive electrode

Cathode - Negative electrode

P.A.N.I.C :

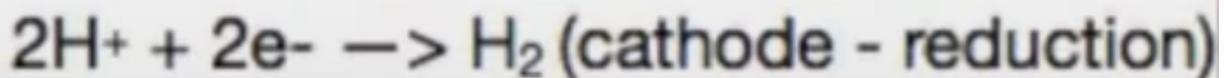
- **P** = Positive
- **A** = Anode
- **N** = Negative
- **I** = Is
- **C** = Cathode

For positive ions:

- The least reactive element discharges

Discharge - The removal or transference of an electric charge, as by the conversion of chemical energy to electrical energy

- In the case of H⁺, Na⁺, OH⁻ and Cl⁻
- H⁺ discharges at the negative electrode because it is least reactive and is positively charged



- In the case of H⁺, Na⁺, OH⁻ and Cl⁻
- Cl⁻ discharges at the positive electrode because it is a halogen and is negatively charged.

Diatomic elements:

- Nitrogen
- Chlorine
- Hydrogen
- Bromine
- Iodine
- Fluorine
- Oxygen

mnemonic :

- Horses - hydrogen
- Need - nitrogen
- Oats - oxygen
- For - fluorine
- Clear - chlorine
- Brown - bromine
- Is - iodine

What are the uses of chlorine?

- Disinfectant
- Kills bacteria in swimming pools
- Make bleach

What are the uses of hydrogen?

- Fuel
- Harden vegetable oil to make margarine

What are the uses of sodium hydroxide?

- Bleach
- Paper making

Compound	Product at anode (+)	Product at cathode (-)	Observations
lead (II) bromide Pb^{2+} Br^{-}	Bromine	lead	lead - cathode - drips - Bromine gas - anode - bubbles off
HCl H^{+} Cl^{-}	Chlorine	Hydrogen	Hydrogen & chlorine bubbled off
aqueous sodium chloride H^{+} OH^{-} Na^{+} Cl^{-}	Chlorine	Hydrogen	Hydrogen & chlorine bubbled off

Electroplating :

- Use metal that you are coating your object with as the anode (+)
- Object being coated (plates) cathode (-)
- Electrolyte is a solution of a soluble compound of the metal
- $Cu^{2+} + 2e^{-} \rightarrow Cu(s)$
- Cathode becomes plates with copper.

What are the uses of electroplating?

- Lots of jewelry - silver plates - made of copper
- Coat steel bumpers with chromium - very hard and very shiny
- Coat steel cans with tin - tin is corrosion resistant

How does a fuel cell work?

- supplied by an external source of fuel (e.g. hydrogen) and oxygen or air
- fuel oxidised to produce a potential difference

Equations in a hydrogen fuel cell:

- $2\text{H}_2 + 4\text{OH}^- \rightarrow 4\text{H}_2\text{O} + 4\text{e}^-$
- $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$

What are the advantages and disadvantages of hydrogen fuel cells?

- Advantages: do not need to be electrically recharged, no atmospheric pollutants, can be a range of sizes
- Disadvantages: hydrogen is highly flammable and hard to store as a gas

Energetics

Exothermic - Heat energy is released

- In exothermic reactions, more energy is needed to make the bonds in the products than is needed to break the bonds in the reactants

Endothermic - Heat energy is taken in

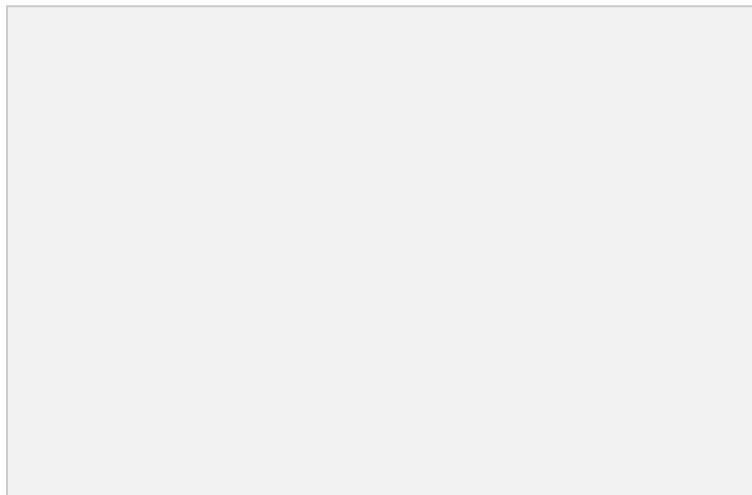
- In endothermic reaction, more energy is needed to break the bonds in the reactants than is needed to make the bonds in the products

Activation energy - the minimum amount of energy required for a reaction to occur

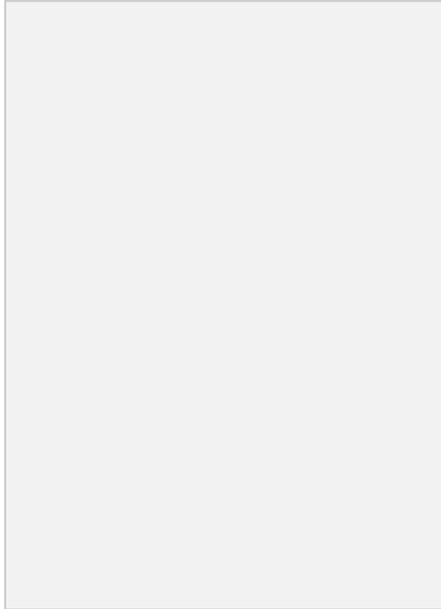
Catalysts - Catalysts is used to speed up the rate of reaction without being used up

How catalysts work?

- Catalysts provide an alternative reaction pathway with lower activation energy
- They are not used up by the reaction



- In an endothermic reaction graph the products are at a higher level than the reactants
- If ΔH is negative it's exothermic
- If ΔH is positive it's endothermic



Physical & Chemical Changes

What is the difference between a chemical and a physical change?

- Physical changes don't make new chemical substances
- Chemical changes take in or give out energy
- Chemical changes are harder to reverse
- Examples of physical changes: mixing, dissolving
- Examples of chemical changes: burning gas

Rates of Reaction

What are the effects of increasing temperature?

- Particles have more kinetic energy
- Collisions occur more frequently and they are harder
- Increased successful collisions
- The rate of reaction will increase

What are the effects of increasing concentration?

- Increased concentration means that there are more particles in the same volume
- Increased frequency of collisions
- The rate of reaction will increase

What are the effects of increasing surface area?

- Increased surface area means an increased frequency of collisions
- The rate of reaction will increase

- You can measure the rates of reaction are given by for example: the change in volume over time, the change in concentration over time
- Marble chips when reacted with hydrochloric acid, they will produce carbon dioxide, so we could measure how quickly the carbon dioxide is produced, either using an top pan balance (it needs a high resolution, because Co2 doesn't weight very much, so we need at least like 0.00 on our weighing scale in order to measure that difference, so when it escapes at the top of a conical flask, we will see the mass decreasing and we can measure that over time) or we can use gas syringes and that will show us the volume of Co2 that is released (this method can't be used if we are measuring hydrogen gas, because it is too light. We won't be able to see a change in the reading on the measuring balance)

Reversible Reactions

Dynamic equilibria:

- Rate if forward and reverse reactions are the same
- Forward and reverse reactions occur at the same time
- No change in the concentrations of reactants or products
- These only true in a closed system

Closed system - Is one where nothing has been added or taken away

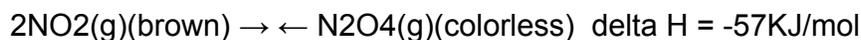
- Position of the equilibria can be changed by altering wither the temperature or the pressure

- Exothermic reaction, the whole reaction gets hotter
- Endothermic reaction, the whole reaction get colder
- Increasing the temperature favors the endothermic reaction, position of equilibria if shift to favor that endothermic reaction
- Decreasing the temperature favors the exothermic reaction, position of equilibria if shift to favor that exothermic reaction
- Increasing pressure favors the side with fewer moles of gas
- Decreasing the pressure favors the side with more moles of gas

Why are the conditions used in the Haber Process described as 'compromised'?



- forward reaction is exothermic and therefore favoured by low temperatures
 - but rates of reaction are too slow at low temperature so 450C temperature is used
 - the forward reaction results in fewer moles of gas so is favoured by high pressures
 - but high pressures are dangerous and expensive so 200 atmospheres is used
-



- Increasing the temperature will favor the endothermic reaction, which means the reverse reaction will be favored, the position of equilibrium shifts to the left. We therefore make more NO₂ and therefore the color changes and it becomes brown
- Increasing the pressure the favor the forward reaction so the position of equilibrium shifts to the right and therefore more N₂O₄ is produced, so the color will change to colorless

What happens when you heat hydrated copper (II) sulphate?

- Blue powder (hydrated copper(II) sulphate) turns white (anhydrous copper (II) sulphate)

What happens when you add water to anhydrous copper (II) sulphate?

- White powder (anhydrous copper (II) sulphate) gets hot and turns blue (hydrated copper(II) sulphate)

Give the word and symbol equation for the conversion of sulfur dioxide to sulfur trioxide?

- Sulfur dioxide + oxygen $\rightarrow \leftarrow$ sulfur trioxide
- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$

Give the condition used for the contact process?

- Vanadium(V) oxide catalyst, V₂O₅
- 450°C
- 2 atm

Give the sources of sulfur dioxide for the contact process?

- Burning sulfur
- Roasting sulfide ores

Give the sources of oxygen for the contact process?

- Air

Redox reaction

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

Redox - Reduction and oxidation occur at the same time

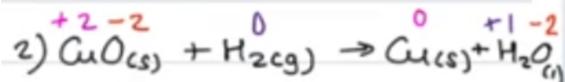
- Reducing agent causes other substance to be reduced
- They themselves are oxidized
- Oxidizing agents causes other substance to be oxidized
- They themselves are reduced

Oxidation numbers

- Oxidation states are a useful tool for the allowing us to identify which species has been oxidized and which has been reduced

Oxidation state/number rules:

- Elements which are not combined with other elements have an oxidation state of zero. E.g O₂, P₄,
- The oxidation number of an uncombined ion is the same as its charge. E.g Na⁺ = +1, Ca²⁺ = +2
- The sum of all the oxidation numbers in a molecule is zero. E.g H₂O, CO₂, HCl. In a complicated ion the sum of oxidation numbers is equal to the charge of the ion. E.g SO₄⁻², NO₃⁻
- Hydrogen = +1 (exception: hydrogen combined as a metal hydride. Here it has an oxidation number of -1)
- Fluorine = -1
- Oxygen = -2 (exception 1: in peroxide = -1 e.g H₂O₂
Exception 2: combined with fluorine = +1)
- Chlorine = -1 (exception: when combined with fluorine = positive
When combined with oxygen = positive)
- For groups 1,2,3, the oxidation number is the same as the group number e.g Na⁺ = +1



Copper has been reduced
+ acts as an oxidising
agent

Hydrogen has been
oxidised + acts as a
reducing

Acids and bases

Acids:

- Donates H⁺ (hydrogen ions) in solution
- pH ≤ 6
- Litmus indicator/paper turns red
- Universal indicator turns
 - Red = strong acid
 - Orange = medium acid
 - Yellow = weak acid
- Methyl orange, turns red in acid

- **Metal + acid → salt + hydrogen**
- **Metal Oxide/hydroxide + acid → salt + water**
- **Metal carbonates + acid → salt + water + carbon dioxide**

- **Metal + oxygen → Base/ Metal oxide**
- **Base/ Metal oxide + water → alkali**
- **Non - Metal + oxygen → acidic oxide / Non - metal oxide**
- **acidic oxide / Non - metal oxide + water → acid**

Bases:

- H⁺ acceptors
- pH ≥ 8
- Universal indicator turns blue/purple
- Litmus paper turns blue
- Methyl orange turns yellow

- Base + Acid \rightarrow salt + water

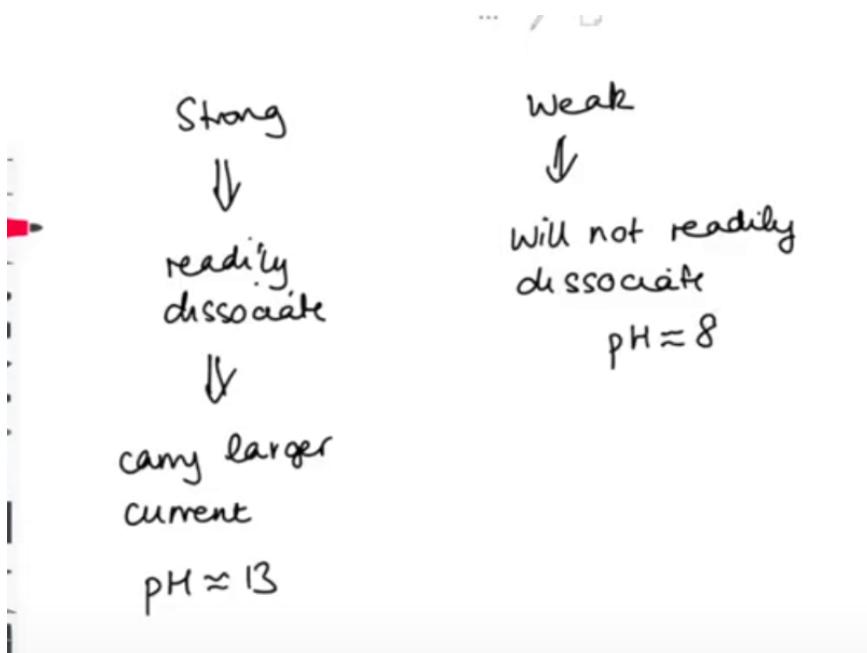
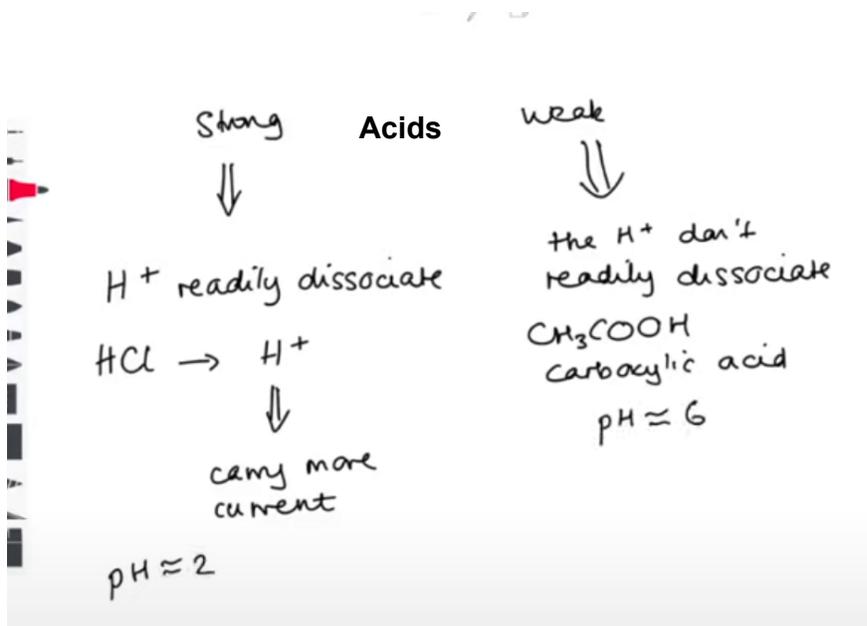
What colour does thymolphthalein indicator turn in acidic solutions?

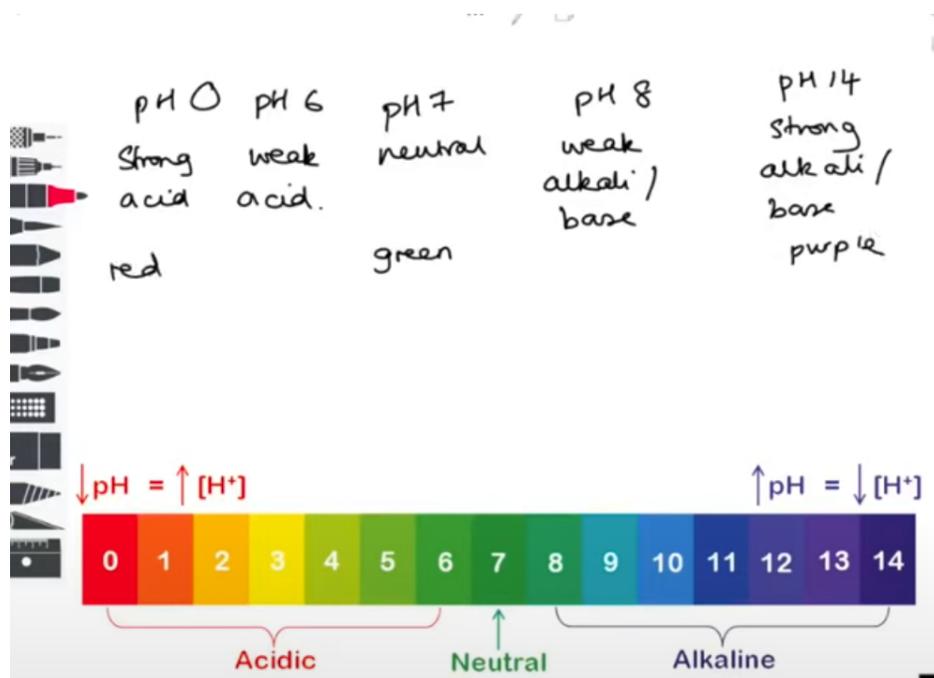
- Colourless

What colour does thymolphthalein indicator turn in alkali solutions?

- Blue

- Difference between alkali and base:
- Alkali is a soluble base
- All alkalis are bases, but not all bases are alkalis





Oxides

What is an acidic oxide?

- In general non metal react with oxygen to make acidic oxides
- E.g carbon burns in oxygen to make carbon dioxide
- Carbon dioxide dissolves in water to make carbonic acids

What is a basic oxide?

- Most metals react with oxygen to make basic oxides
- Basic oxides can neutralize acid
- E.g copper (II) oxide

What is an amphoteric oxide?

- Can be both acidic and basic
- Will react with both acids and bases
- E.g aluminium oxide

What is a neutral oxide?

- Is neither acidic nor basic
- Will not react with either acid or base
- E.g carbon monoxide

Salts

- A salt is formed when the hydrogen of an acid is replaced by a metal or ammonium
- $\text{HCl} + \text{KOH} \rightarrow \text{KCl} + \text{H}_2\text{O}$
- $\text{H}_2\text{SO}_4 + \text{CaCO}_3 \rightarrow \text{CaSO}_4 + \text{H}_2\text{O} + \text{CO}_2$
- Only metals above hydrogen in the reactivity series will react with acids

- All nitrates are soluble
- All potassium, ammonium, and sodium compounds are soluble
- All sulphides are soluble, except:
 - * Lead(II) sulphate
 - * Barium sulphate
 - * Calcium sulphate
- All chloride are soluble, except:
 - * Lead(II) chloride
 - * Silver chloride

- All carbonates are insoluble, except:
 - * Sodium carbonate
 - * Potassium carbonate
 - * Ammonium carbonate
- All hydroxides are insoluble, except:
 - * Sodium hydroxide
 - * Potassium hydroxide
 - * Ammonium hydroxide

To make soluble salts:

- Use appropriate metal, metal oxides, metal hydroxide, metal carbonate
- Acid

- Crystallisation is used to make soluble salts that don't contain K, Na, or NH_3

Use the crystallisation method

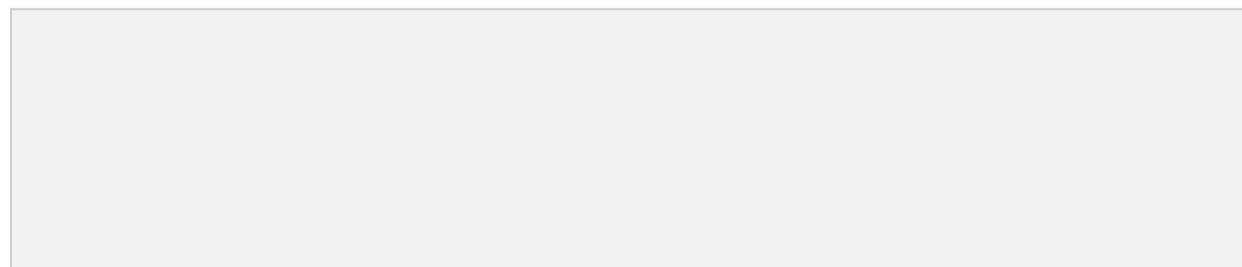
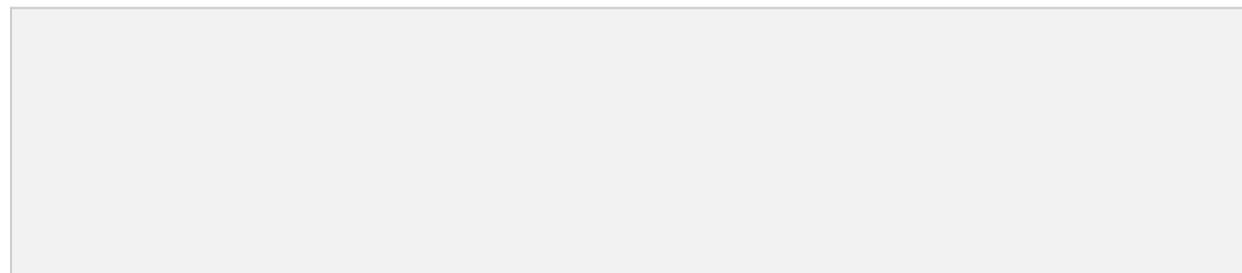
- REACT
- FILTER
- EVAPORATE: heat to evaporate some water
- COOL: collect crystals that form
- DRY: allow the crystals to dry in a warm place or on filter paper

- This method will not work for K, Na, and NH_3 , because they are extremely soluble, if they were added to acid which contains water, they would react with both the acid and the water and they would continuously dissolve away and there will be nothing

to filter and therefore nothing to evaporate. So that's why crystallisation doesn't work in this case

- In this situation you have to use a titration method. We use titration, because we need to know the exact volume of acid and alkali we need to add in order to make the salt.
- REACT: an acid (from a burette) with an alkali (in a conical flask)
- INDICATOR: requires an indicator to show when the alkali has been neutralised (all alkali has been reacted)
- REPEAT: once amounts required have been worked out, add required volumes of acid to alkali without indicator
- EVAPORATE: heat to evaporate some water, this concentrates the solution
- COOL: collect crystals that form
- DRY: allow the crystals to dry in a warm place or on filter paper

- To make an insoluble salt, react two soluble salts



Define a hydrated substance?

- A substance that chemically combines with water

Define anhydrous substance?

- A substance containing no water

Define water of crystallisation?

- The water molecules present in hydrate crystals
- E.g $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$
- E.g $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$

The periodic table

What is the periodic table?

- Method of classifying elements
- Can be used to predict properties of elements

Describe the change from metallic to non-metallic elements across a period?

- Metals lie to the left
- Non metals lie on the right
- Transition elements are in the middle

What is valency?

- The number electrons that are in the outer shell

How does group number relate to the element's character?

- All elements in the same group have same valency
- Elements with same group share similar characteristics
- Number of outer shell electrons dictates whether metallic or non-metallic

Group 1 elements

Group 1 metals - Alkali metals

- Group 1 elements all have the same chemical properties because they all have 1 electron in their outer shell
- As we descend group 1 the elements become more reactive
- As we descend, atoms are larger (have more shells of electrons)
- The outer shell electron is further from the nucleus and is more easily lost
- Group 1 metals are stored in oil
- They are soft and can be cut with a knife, oxidise quickly once cut
- They have low melting and boiling points, they also have low density
- Melting point decrease down the group, Density increases down the group
- React with oxygen to form oxides
- React with cold water to form hydroxides
- React with halogens to form salts
- Rate of reaction with water increases as we go down the group

Observations when added to water:

- Fizz (release H₂)
- Floats, move around, dissolve
- Resulting solution turns universal indicator blue
- Lithium doesn't produce a flame
- Sodium produces an orange flame (when added to water)
- Potassium produces a lilac flame (when added to water)

- Lithium + water → lithium hydroxide + hydrogen

Halogens (group 7)

- Fluorine(yellow colour gas) and chlorine(green colour gas) are gases at room temperature
- Bromine is a red/brown liquid
- Iodine is a grey solid, iodine goes through the process of sublimation. Iodine goes from grey solid to a purple gas
- Halogens react with hydrogen to form hydrogen halides
- Hydrogen + bromine \rightarrow hydrogen bromide (very acidic and poisonous)
- They have low boiling, and melting point
- Poor conductors of electricity and heat

- The reactive halogen will displace the less reactive halogen
- Reactivity decreases as we go down the group

- Iodine is larger (has more shells of electrons)
- The outer shell electrons are further from the nucleus
- Harder to gain extra electron

Displacement summary equations (more reactive halogens displace less reactive elements from their compounds)

E.g.



Ionic equations of halogen displacement reactions:

E.g.



Transitions Metals

What are transition elements?

- Metals with high densities, high melting points
- Form coloured compounds
- Often act as catalyst
- Have variable oxidation states

Noble Gases

What are noble gases?

- In the group VIII (also known as group 0)
- Unreactive
- Monoatomic
- Gases
- Have full outer shell of electrons
- No need to gain or lose electrons

What are noble gases used for?

- Providing inert atmosphere
- Argon in lamps
- Helium in balloons

Properties Of Metal

- They have high melting and boiling points
- They are good conductors of heat and electricity
- They are shiny
- They are sonorous (which means when we hit it they make a noise)
- They are malleable and ductile
- Metal lose electrons and become positive ions
- Metal form basic oxides
- They partake in ionic bonding

Malleable - Can be hammered into shape

Ductile - Can be drawn into wire

Properties Of Non-Metal

- They are dull and not shiny
- They tend to have low boiling and melting points
- They are brittle
- They form acidic oxides
- They gain electrons in bonding and become negative ions
- They partake in covalent and ionic bonding

Brittle - shatter when hit

Uses Of Metal

- Al is used to manufacture planes due to its low density
- Al is also a good conductor of heat

- Cu is used for electrical wiring
- Good conductor of heat and electricity

Alloys And Their Properties

Alloy - Mixture of two or more metals

- Steel is an alloy of iron, it contains both iron and carbon
- Alloys are made of ions of different sizes, layers can't slide so easily

- Low carbon steel (contains less than 0.25% carbon). It is used to make car bodies, bridges, and ship building
- Low carbon steel rusts easily and is very heavy (high density)

- Stainless steel contains nickel, chromium, and iron
- It has high resistance to corrosion
- It is used to make cutlery, saucepans, and gardening tools

What is zinc used for?

- Galvanising - coating iron to protect from corrosion
- Making brass - as alloy with copper. It is hard, strong, and shiny
- Brass is used for door lock, keys, and musical instruments

The reactivity Series

- Potassium (most reactive)
- Sodium
- Lithium
- Calcium
- Magnesium
- Aluminium
- Carbon
- Zinc
- Iron
- Hydrogen
- Copper
- Silver
- Gold

- Silver and Gold are found native in the Earth's crust
- Presence of oxide layer means that aluminium is less reactive

We have an unknown metal and we don't know how reactive it is, what can we do to try and determine it's position in the reactivity series?

- React unknown metal with **cold water**. Only very reactive metals such as those found in group one will react with cold water(will produce metal hydroxide). If no reaction than
- React with steam(will produce metal oxide). If no reaction than
- React with an acid(will produce salt + hydrogen)

- Only metals more reactive than hydrogen will react with acids

How does a reaction with carbon show reactivity?

- Carbon is more reactive than some metals
- Carbon reduces their oxides to the metal

How does a reaction with another metal oxide show reactivity?

- If more reactive, the metal reacts with oxygen to form oxide
- A metal will reduce the oxide of a less reactive metal

How does a reaction with ions of other metals in solution show reactivity?

- A metal displaces a less reactive metal
- From solutions of its compounds

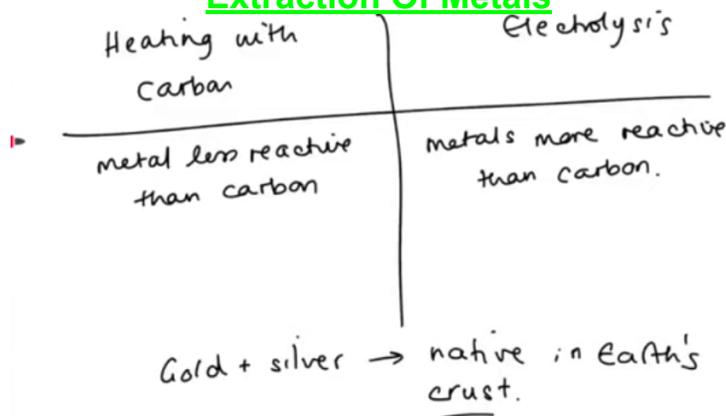
Corrosion Of Metals

- Rusting needs both oxygen and water (salt increases the rusting process)
- Barrier methods e.g. paint, oil, grease

Galvanising - Iron is coated in zinc which is more reactive

- Zn forms Zn^{2+} and donates electrons to Fe, preventing formation of Fe^{3+}

Extraction Of Metals

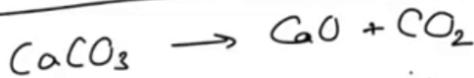
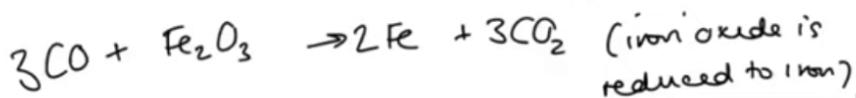
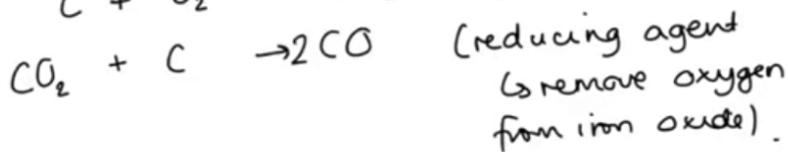
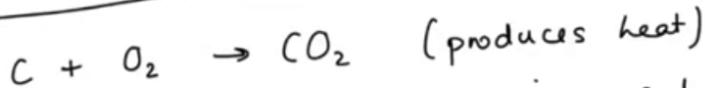


Blast Furnace → purifying iron
iron oxide - haematite

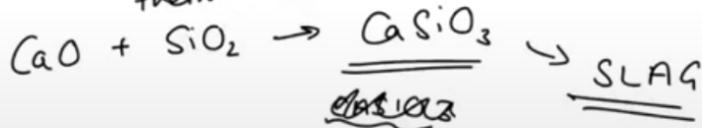
- heating with carbon
- carbon is more reactive than iron.

Raw materials

- coke (carbon)
- iron (III) oxide
- limestone - removes acidic impurities



thermal decomposition



- Aluminium is more reactive than carbon and therefore cannot be reduced in this way
- Electrolysis is used to extract aluminium
- Al extraction is expensive due to high use of electricity
- The molten solution contains dissolved aluminium ore and cryolite.
- Cryolite is used to reduce the melting point of aluminium.
- C anodes react with oxygen forming carbon dioxide which burns them away

- $2O_2 - 4e^- \rightarrow O_2$ (reaction in anode)
- $Al^{3+} + 3e^- \rightarrow Al$ (reaction in cathode)

Cryolite - A mineral consisting of a fluoride of sodium and aluminium. It is colourless to white, also brownish, reddish and rarely black

Water

Physical test for water - Boils at 100°C

- Every pure substance has one distinct boiling point

Chemical test for water - Turns white anhydrous copper (II) sulphate blue

- Coagulant makes particles stick together

Nitrogen And Fertilisers

What does fertilisers contain:

- Nitrogen
- Phosphorus
- potassium
- This type of fertilisers are called **NPK fertiliser**

Uses of nitrogen:

- It is used by plants to produce amino acids and proteins

Uses of phosphorus:

- Which is needed in the healthy growing of roots
- It helps crops ripen

Uses of potassium:

- It is used to produce protein
- This helps them resist diseases

How is ammonia displaced from its salt?

- Heat ant ammonia compound
- With a strong base
- Base displaces ammonia from compound

Air Quality And Climate

Air contains:

- 78% nitrogen (N₂)
- 21% oxygen (O₂)
- 0.04% carbon dioxide (CO₂)
- <1% water vapour, Noble gases

Pollutants:

- Carbon dioxide(CO₂), is formed by the complete combustion of fuel. It is a greenhouse gas, it causes global warming. Which causes climate change and the extinction of species
- Carbon monoxide(CO), is formed by the incomplete combustion of fuel. It is a toxic gas, it combines with red blood cells in the blood and therefore lowers the oxygen carrying capacity
- Methane(CH₄), comes from the decomposition of vegetation and the waste gases from the digestion of animals (rice paddy fields, cattle farming). It is a greenhouse gas
- Oxides of nitrogen, it is formed from the high temperature in car engines, the high temperature causes the O₂ + N₂ to react. It causes acid rain, photochemical smog and respiratory problems
- Sulphur dioxide, sulphur impurities in crude oil. It causes acid rain

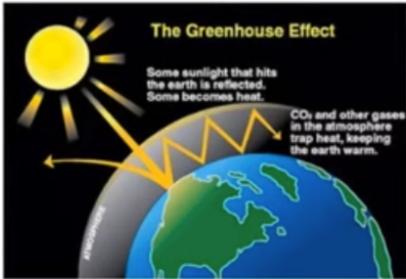
Particulates - this is caused by unclean fuel and these leads to increased respiratory problem and also increased cancers

ENHANCED GREENHOUSE EFFECT

① Sun emits thermal energy in the form of infra-red radiation which enters the Earth's atmosphere

② some thermal energy is absorbed by the Earth; some is reflected back out into space

by carbon dioxide, methane



The Greenhouse Effect

Some sunlight that hits the earth is reflected. Some becomes heat.

CO₂ and other gases in the atmosphere trap heat, keeping the earth warm.

Human activities e.g. deforestation, combustion

- ↳ ↑ greenhouse gases
- ↳ more thermal energy is absorbed + trapped.
- ↳ Average temp. rising (enhanced greenhouse effect)
- ↳ less thermal energy lost to space
- ↳ GLOBAL WARMING

How do we reduce the effects of human activity on the earth?

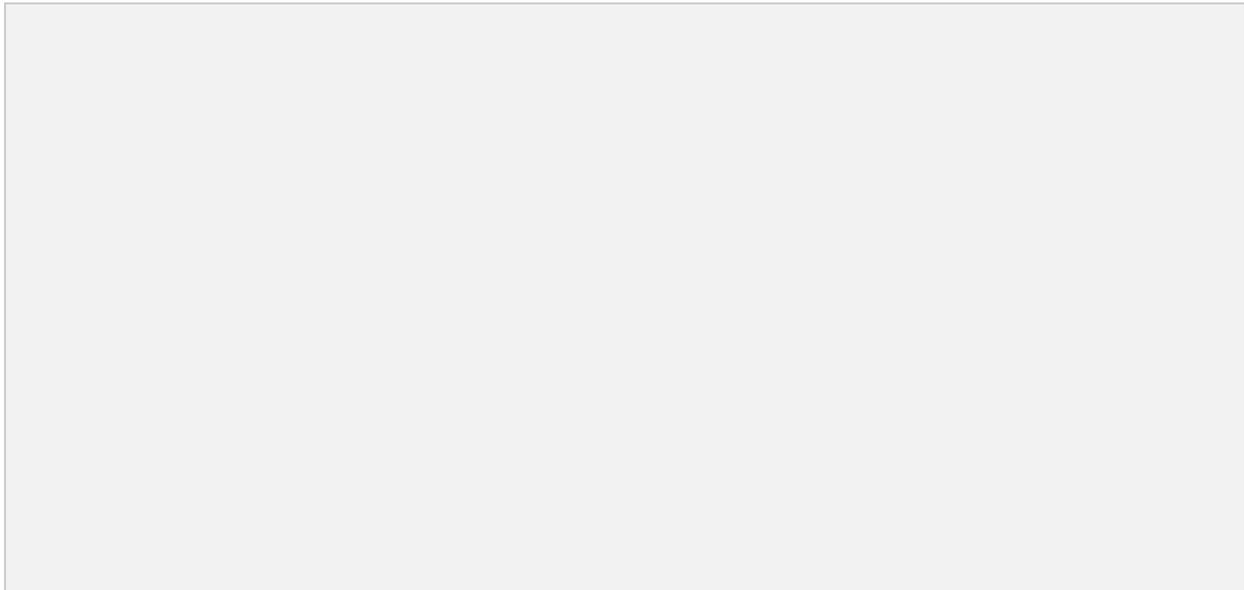
- Plant more trees, CO₂ is used up by trees in photosynthesis. The by product of which is Oxygen
- Decrease livestock farming, fewer cows means there is a reduction of methane gas
- Decrease use of fossil fuels, complete combustion will lead to less CO₂ released. Incomplete combustion will lead to less CO released. Less acid rain

- We can use hydrogen as a fuel. Which produces H₂O as waste product
- Use renewable energy, fewer polluting gases produced. Solar, wind, tidal.

Acid rain reduction:

- Use catalytic converters
- Reduce emission of SO₂
- Use low sulphur fuels
- Carry out flue gas desulfurization with calcium oxide

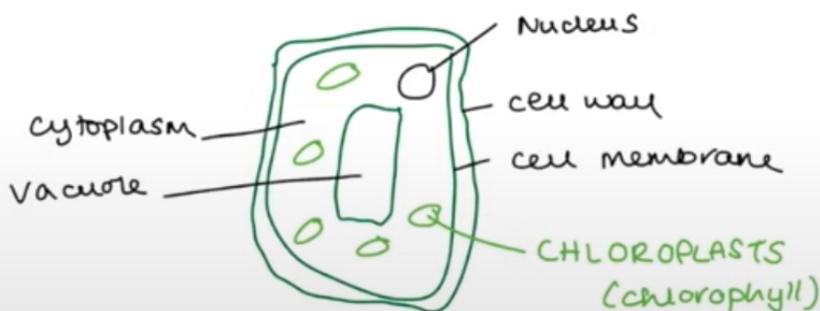
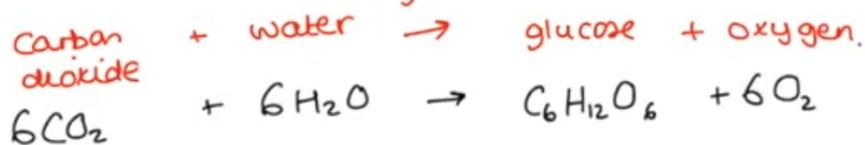
Flue gas - Waste gases that comes from combustion of any fuels



Photosynthesis

↳ plants → produces food (glucose)

light energy



Organic Chemistry

Hydrocarbon - A compound containing hydrogen and carbon only

- Alkanes are generally unreactive, except when burned

Saturated - Contains single C bonds only

Unsaturated - contain C double bond C or c triple bond C

Isomer - Same molecular formula but different structural formula

Homologous series - Compounds with the same functional group and the same general formula

- Same chemical properties
- Trend in physical properties
- Same functional group

Functional group - Atom or group of atoms which determine the chemical properties of a compound

Fuels & Fractional Distillation

Crude oil - Contains a mixture of hydrocarbons

Fuel - A substance which releases energy when burned

Fractional distillation:

- Heat crude oil until it evaporates
- It rises, cools and condenses
- Different fractions have different boiling points and therefore condense at different positions
- Longer chains condense at the bottom where it is hottest

Fraction - Group of compounds with similar boiling points

Viscosity - How runny a substance is

- Honey - very viscous
- Water - not viscous

Flammability - How readily a substance sets alight

Volatility - How a readily a substance turns into a gas

Alkanes, Alkenes & Cracking

- Cracking breaks large hydrocarbon chains into smaller, more useful ones
- Shorter chain hydrocarbon make better fuels as they are more flammable
- Cracking 600-700°C, silica or alumina catalyst

Test for an alkene/ unsaturated hydrocarbon (C double bond C):

- Add bromine water
- Orange to colourless

Alcohols

ALCOHOLS - OH

Number of Carbon atoms	Name	Displayed formula
1C	methanol	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{O}-\text{H} \\ \\ \text{H} \end{array}$
2C	ethanol	$\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array}$
3C	propanol propan-1-ol propan-2-ol	$\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{O}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array}$ $\begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{O}-\text{H} \quad \text{H} \end{array}$

We can oxidise alcohols by:

- Combustion (burning in air)
- Natural reaction with oxygen in the air due to action of microorganism (microbial oxidation)
- Heat with acidified potassium dichromate in the presence of dilute sulphuric acid

- Ethanol is oxidised to ethanoic acid
- Alcohols produce carboxylic acids when they are oxidised

Uses of alcohols:

- Alcoholic drinks
- Fuels
- Perfumes (they are good solvents)

Production of alcohol:

- Fermentation
- Hydration of ethene

Hydration of ethene:

- Ethene comes from crude oil (non - renewable)
- Hydration of ethene uses high temperature (300°C) and pressure (60 atm)
- It is a continuous process
- $\text{C}_2\text{H}_4 + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_5\text{OH}$

- It makes pure alcohol

Fermentation:

- Fermentation uses sugar cane (renewable resource), yeast acts as the catalyst
- Fermentation uses low temperature and pressure
- Fermentation is a batch process
- $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$
- It makes impure alcohol

Carboxylic Acids

<u>CARBOXYLIC ACIDS</u>		
Number of C atoms	Name	Displayed formula
1C	methanoic acid	$ \begin{array}{c} \text{O} \\ \parallel \\ \text{H} - \text{C} - \text{O} - \text{H} \end{array} $
2C	ethanoic acid	$ \begin{array}{c} \text{H} \\ \\ \text{H} - \text{C} - \text{C} = \text{O} \\ \quad \quad \\ \text{H} \quad \quad \text{O} - \text{H} \end{array} $
3C	propanoic acid	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H} - \text{C} - \text{C} - \text{C} = \text{O} \\ \quad \quad \quad \\ \text{H} \quad \text{H} \quad \quad \text{O} - \text{H} \end{array} $

Describe the reaction of ethanoic acid with metals

- Acid + Metal \rightarrow Salt + Hydrogen
- e.g. Ethanoic acid + Magnesium \rightarrow Magnesium ethanoate + Hydrogen
- $Mg + 2CH_3COOH \rightarrow (CH_3COO)_2Mg + H_2$
 - o Solution fizzes as hydrogen gas produced
 - o Hydrogen burns with squeaky pop

Describe the reaction of ethanoic acid with bases

- Acid + Base \rightarrow Salt + Water
- e.g. Ethanoic acid + Sodium Hydroxide \rightarrow Sodium ethanoate + Water
- $NaOH + CH_3COOH \rightarrow CH_3COONa + H_2O$
 - o Neutralisation reaction

Describe the reaction of ethanoic acid with carbonates

- Acid + Carbonate \rightarrow Salt + Water + Carbon dioxide
- e.g. Ethanoic acid + Copper carbonate \rightarrow Copper ethanoate + Water + Carbon dioxide
- $CuCO_3 + 2CH_3COOH \rightarrow (CH_3COO)_2Cu + CO_2 + H_2O$
 - o Solution fizzes as CO_2 gas produced
 - o Solution turns blue (copper ethanoate)

How is ethanoic acid made?

- By oxidation of ethanol
 - o Bacterial oxidation - Ethanol reacts with oxygen in the air
 - o Under reflux - Ethanol reacts with acidified potassium manganate(VII)

Polymers

Polymers - A polymer is a large molecule formed from many small molecules known as monomers

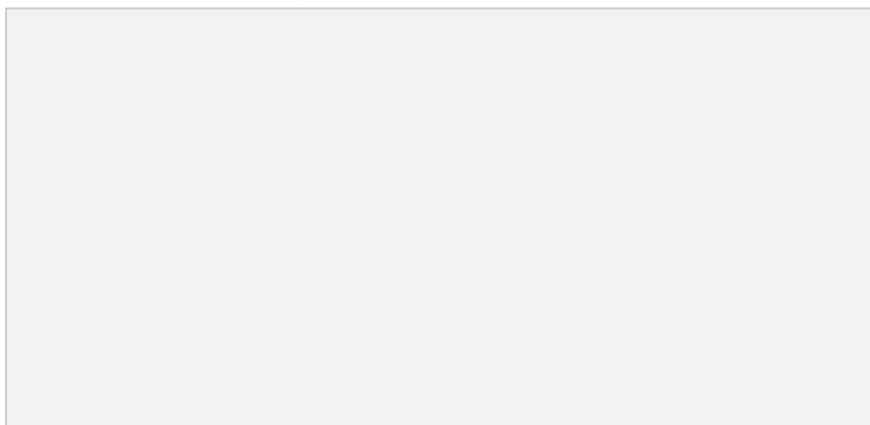
Addition polymers - Formed from many monomers joining together. No byproducts are formed.

Condensation polymers - Small molecules are lost in the reaction. I.e. water. Can be converted into its original monomers

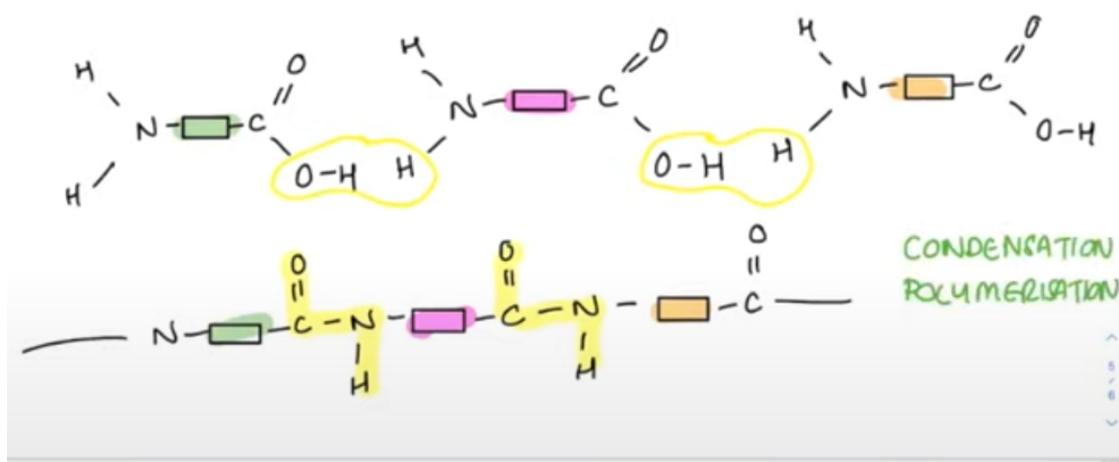
- Plastics are difficult to dispose of because they are inert (unreactive)

methods of disposal:		
	Adv.	Disadv:
Landfill + oceans	- cheap - no greenhouse or toxic gases	- ugly, smelly, noisy - use large areas of land. - Long lasting
Incineration	- requires little space - can produce heat for local homes/offices / to produce electricity	- Expensive to build + maintain the plant - produces toxic + greenhouse gases - Ash produced occupies landfill

Proteins - Natural polyamides. Amino acid monomers



Structure of a protein:



Experimental Design

Solute - A solid which dissolves in a solvent

Solvent - The liquid in which the solute dissolves

Solution - Mixture of a solute and solvent

Saturated solution - A solution in which no more solute can dissolve

Chromatography - Used to separate dyes, inks, food colourings

- Ink which travels the furthest is the most soluble

$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$$

How can we separate colourless substances?

- Use a locating agent
- Apply to chromatogram

Why is purity important?

- Prevent contamination
- Needed for food
- Safety of drugs

Separation And Purification

Filtration - Used to separate an **INSOLUBLE** solute from a solvent

Evaporation - Used to separate a **SOLUBLE** solute from a solvent (also known as the crystallisation method)

Immiscible - Liquids which do not mix

- Separating funnel is used to separate two immiscible liquids

Simple distillation - To separate liquids of different boiling points

Fractional distillation - Used to separate liquids with many different boiling points

Pure substance - Substance which contains only 1 type of material. E.g. 1 element or 1 compound

- Pure substance have a fixed boiling point

Identification Of Ions And Gases

- Hold a lighted splint over the gas and if there's a squeaky pop, hydrogen is present
- Oxygen relights a glowing splint
- Carbon dioxide turns lime water cloudy
- Chlorine bleaches damp litmus paper
- Ammonia turns damp red litmus paper blue

How do we identify sulphur dioxide?

- Colourless, poisonous acidic gas
- Soak filter paper in acidified aqueous potassium manganate (VII)
- Paper goes from purple to colourless

Flame test:

- Clean nichrome wire (dipped in HCl)
- Dip in sample
- Hold in roaring blue flame (yellow sooty flame obscures the colour of the flame)

- Li^+ = red flame
- Na^+ = yellow flame
- K^+ = lilac flame
- Ca^{2+} = orange/red flame
- Cu^{2+} = blue/green flame

What tests are used to identify aqueous cations?

Ion	Add aqueous sodium hydroxide	Add aqueous ammonia
Aluminium	White precipitate forms Precipitate dissolves if more added	White precipitate forms Adding more has no effect
Ammonium	Ammonia gas given off	No reaction
Calcium	White precipitate forms Adding more has no effect	No precipitate, or very slight white one
Chromium (III)	Grey-green precipitate Dissolves if add more- grey-green solution	Grey-green precipitate forms Adding more gives purple solution
Copper (II)	Pale blue precipitate Adding more has no effect	Pale blue precipitate Dissolves if add more- deep blue solution
Iron (II)	Pale green precipitate	Pale green precipitate
Iron (III)	Red-brown precipitate	Red-brown precipitate
Zinc	White precipitate Dissolves if add more	White precipitate Dissolves if add more

Testing for halides:

- Add nitric acid
- Add silver nitrate
- $\text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl}$. AgCl is a **WHITE** ppt
- $\text{Ag}^+ + \text{Br}^- \rightarrow \text{AgBr}$. AgBr is a **CREAM** ppt
- $\text{Ag}^+ + \text{I}^- \rightarrow \text{AgI}$. AgI is a **YELLOW** ppt

How do you detect anions?

Ions	Test	Result
Halides (chloride, bromide, iodide)	Equal volume dilute nitric acid Add aqueous silver nitrate	White = chloride Cream = bromide Yellow = iodide
Sulfate (SO_4^{2-})	Equal volume dilute HCl Add barium nitrate	If sulphate present – white precipitate forms
Sulfite (SO_3^{2-})	Add equal volume dilute HCl Heat gently	Sulfur dioxide given off (Potassium manganate turns from purple to colourless)
Nitrate (NO_3^-)	Add small amount dilute sodium hydroxide Add aluminium foil Heat gently	Ammonia gas given off if nitrate present
Carbonate ions (CO_3^{2-})	Add dilute HCl	Mixture bubbles and gives off CO_2 CO_2 turns limewater milky