1.1 Atoms, symbols, masses & charges

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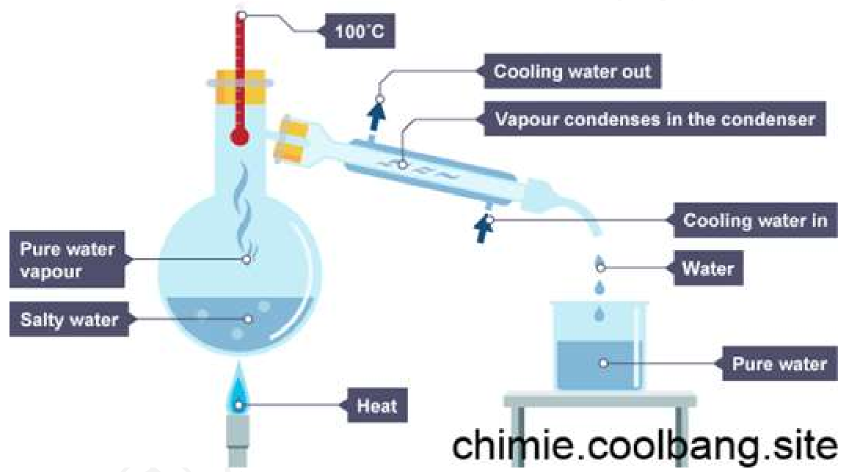
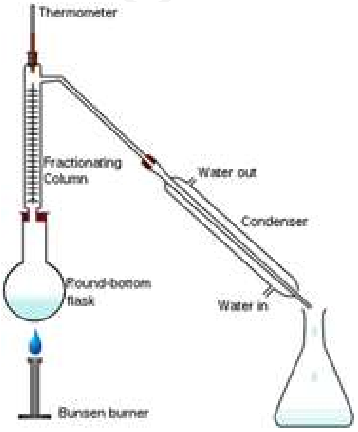
1.1.1 Atoms and elements

Atom	The smallest amount of an element that can exist
Molecule	Two or more atoms chemically bonded together
Element	Contains only one type of atoms
Mixture	Two or more substances combined together but not chemically bonded together.
Compound	Contains two or more different elements chemically bonded together.

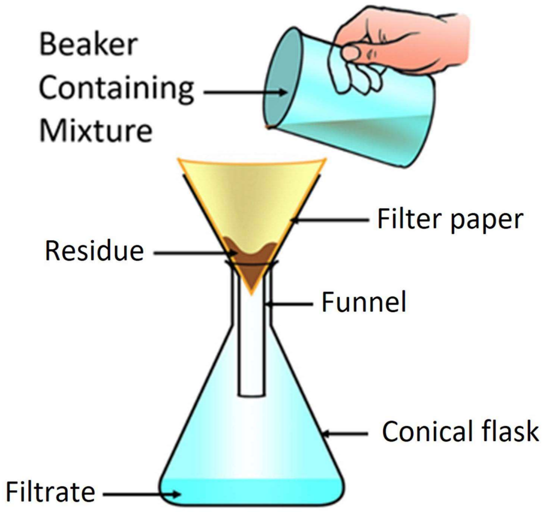
Atom radius = 0.1 nm (10^{-10} m)
Nucleus radium = 10^{-14} m

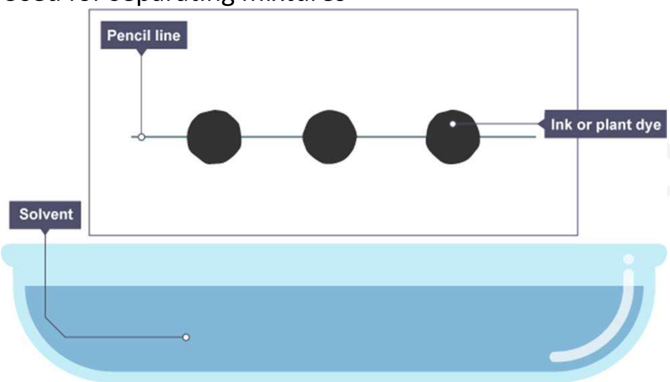
1.1.2 Separating mixtures

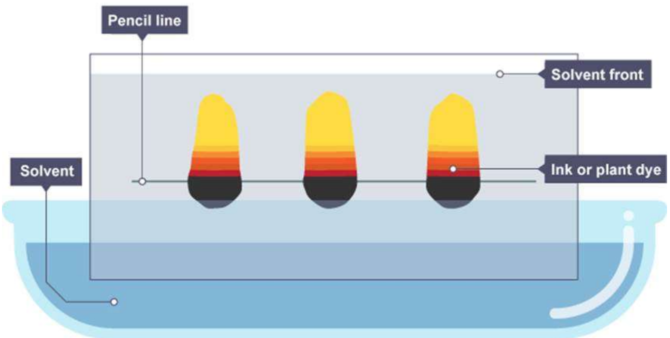
- Mixture can be separated by physical processes
- A physical process is one that doesn't change one substance into another
- A chemical change creates new substances from the substances that were originally present.
- Distillation
  - Distillation works because the different liquids have different boiling points.
  - Simple distillation
    - 

    - Best for separating a liquid and a dissolved solid
  - Fractional distillation
    - 

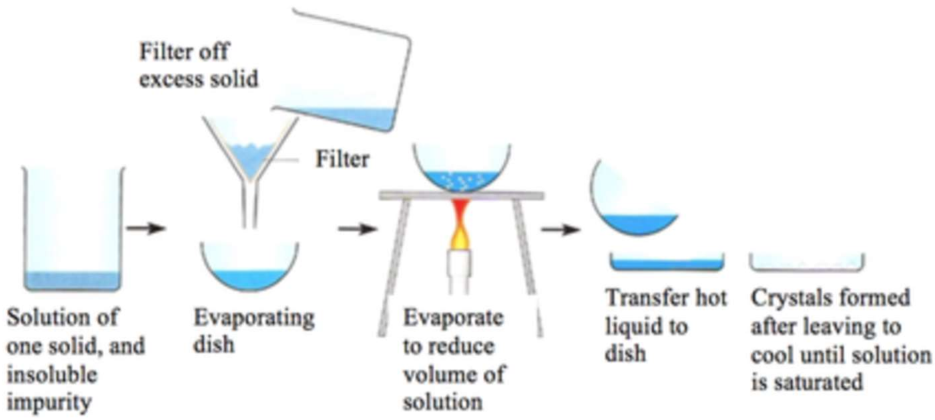
    - For mixture with two or more liquids
- Simple distillation process
  - Heat the solution
  - Water evaporates
  - Vapour cools in condenser
  - Vapour condenses to become water

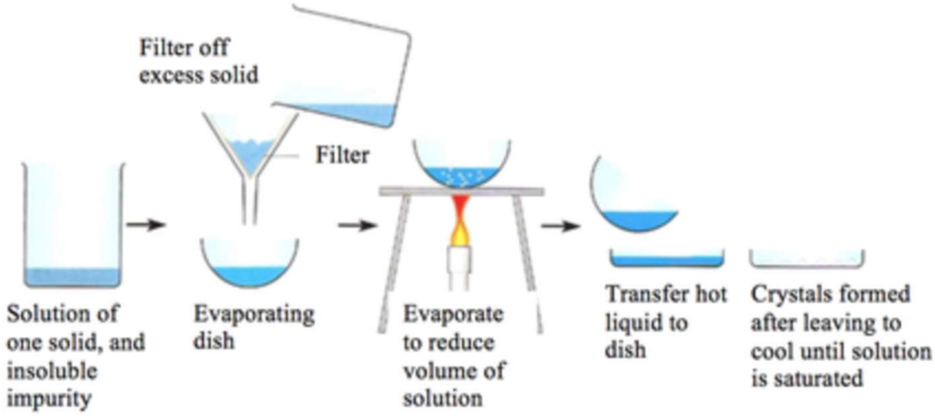
    - Salt stays in the original container
    - Water collected in the beaker
    - Works because salt and water has different boiling points
- Filtration
    - Used for separating a undissolved solid and a liquid
    - 

- Chromatography
    - Used for separating mixtures
    - 



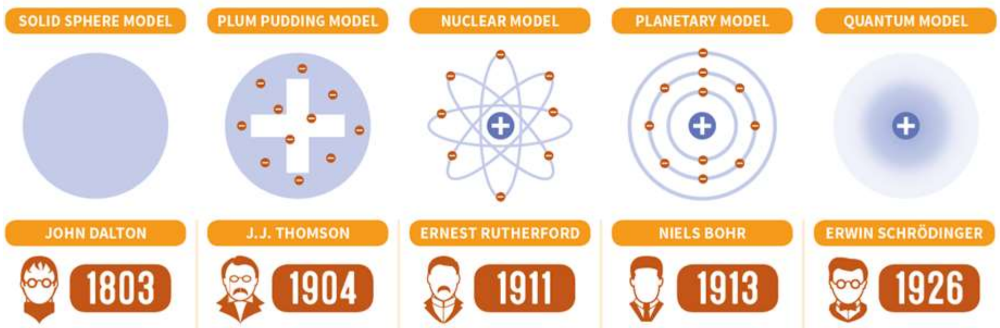
- Crystallisation
    - Used for a dissolved solid in a liquid



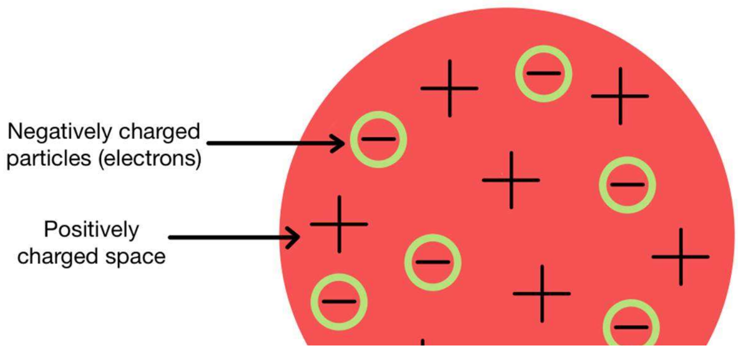


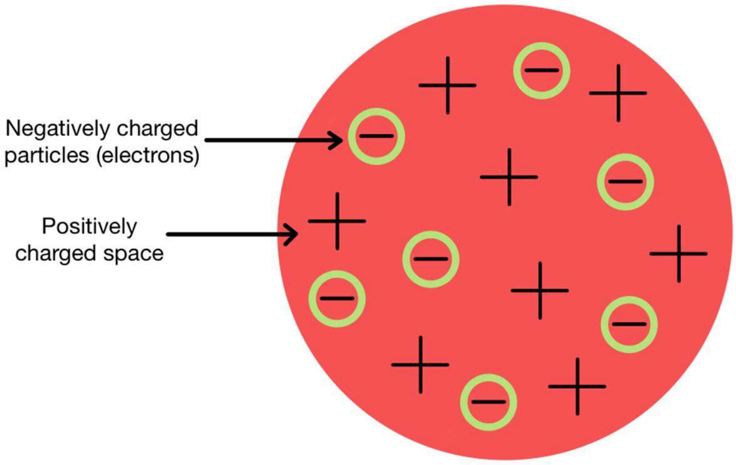
- Crystallisation process
  - Filter off excess liquid
  - Heat to crystallisation point to reduce the volume of the solution
  - Leave the solution to crystallise in a evaporating dish
  - Crystals formed after leaving to cool until solution is saturated

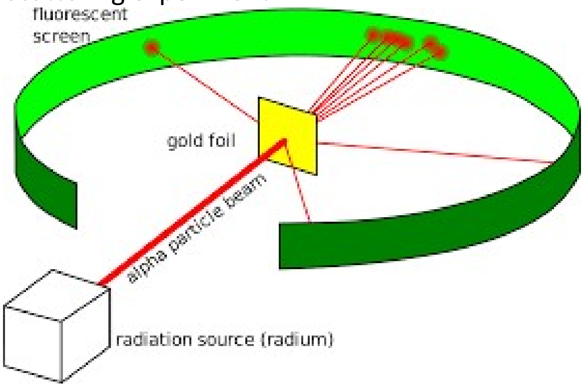
1.1.3 Changing ideas about atoms

- Atomic model dates
  - 

- Discovery of particles
  - Protons
    - Ernest Rutherford 1909
  - Electrons
    - J. J. Thomson 1897
  - Neutrons
    - Proposed by Rutherford in 1920
    - Sir James Chadwick discovered neutrons in the year 1932
- Solid sphere model
  - Substances were made of indestructible atoms that were like tiny, hard spheres, characterized by mass and rearranges and combines themselves with other atoms in chemical reactions
- Cathode ray tube
  - Beams of particles attracted to a positive charge - they must be negative
  - Discover electrons
- Plum pudding model
  - Electrons embedded in a positively charged sphere





- Alpha scattering experiment
  - 

  - Shoot small positively charged alpha particles at the gold foil
  - Thomson's model: all particles will pass through (electrons too small to block)
  - Reality: some bounced right back
- Nuclear model
  - Electrons orbit around the nucleus
  - Nucleus is a dense mass of positively charged particles
- Planetary model
  - Electrons orbit the nucleus in fixed shells or orbitals located at set distances from the nucleus
  - Each orbital has a different energy associated with it, with the higher energy orbitals being located further away from the nucleus
  - Electrons can move from one shell to another
- Quantum model
  - Impossible to determine the exact orbit or speed of electrons, only predict where it will probably be

Electron arrangements

- Electrons in an atom always occupy the lowest energy shell first
- Isotope: Atoms of the same element with different numbers of neutrons
- Boiling point increase going down the table, since the atoms get bigger and heavier and a higher atomic mass
- Left side of equation: reactant
- Right side of equation: product

1.2 Periodic table

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1.2.1 Atomic arrangement in the table

- Periodic table
    - Elements periodically show similar properties
    - Arranged in blocks
- Groups
    - Similar chemical properties
    - Same number of valence electrons in the outer shell
- Argon and Potassium
    - Argon has a larger mass number although it comes before potassium

1.2.2 Development of the table

1. Scientists tried to arrange the known elements by their atomic weights
    - The early periodic tables were incomplete
    - Undiscovered elements weren't included - sub-atomic particles weren't discovered
    - Some placed in the wrong order because atomic mass is used rather than atomic number, such as Argon and Potassium
2. Mendeleev
    - Left gaps for undiscovered elements, predicted their properties
    - Swapped the positions of some elements to arrange them into groups of similar properties - explained by sub-atomic particle
    - Elements found filled in the gap and matched the properties predicted
    - Didn't leave gap for noble gases
    - Theoretical

1.2.3 Metals, non-metals & ions

	PROPERTIES OF METALS	**PROPERTIES OF NON-METALS**
Appearance	Shiny & silver coloured	Dull
Hardness	Malleable, ductile, strong	Brittle
Melting point	High melting point	Low melting point
Conductivity	Good conductors of heat & electricity	Poor conductors
Density	Higher density	Lower density
Sonorous?	Yes	No

- Elements lose, gain or share electrons to have 8 electrons in their outer shell.
- Elements that react to form positive ions (cations) are metals.
- Elements that form negative ions (anions) are non-metals

1.2.4 Noble gases

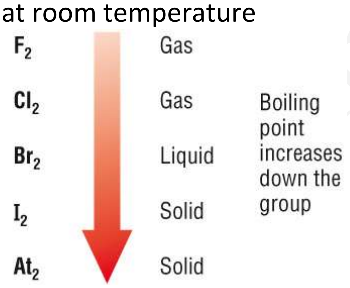
- Group 8
- Full outer shell (8 electrons in its outer shell except for helium which has 2 electrons in its outer shell)
- Very stable and unreactive because their atoms already have stable arrangements of electrons
- Melting and boiling point increase going down the table, since the atoms get bigger and heavier and a higher atomic mass

1.2.5 Alkali metals

- Group 1
- Properties
    - Low density (Lithium, sodium and potassium float on water)
    - Low melting point

    - Very reactive
    - Soft
    - Silver coloured
    - Good conductors
- Reaction with water
    - alkali $metal_{(s)}$ + $water_{(l)}$ → metal $hydroxide_{(aq)}$ + $hydrogen_{(g)}$
    - The metal hydroxide dissolves in the water
    - Metal hydroxide Produces an alkaline solution
    - Lithium
        - Relatively slow
        - Does not melt
        - Fizzing
    - Sodium
        - Melt due to large amount of heat
        - Dash across the water surface
        - Hydrogen released caught fire
    - Potassium
        - More violently
        - Lilac fire because it is hot enough
        - Melts into a shiny ball that dashes across the surface
- Reactivity
    - Going down the table - more shells - outer shell gets further from the nucleus - weaker attraction - easier to lose outer shell electrons - more reactive

<u>1.2.6 Halogens</u>

- Group 7
- Non-metals
- Exists as diatomic molecules
- Going down the table - more shells - outer shell gets further from the nucleus - weaker attraction - harder to gain electrons - less reactive
- Melting and boiling point increase going down the table, since the atoms get bigger and heavier and a higher atomic mass
- State at room temperature
    - 

- Displacement
    - A more reactive halogen can displace a less reactive halogen from an aqueous solution of its salt
    - Take the electron from the less reactive halogen ion

1.3 Transition metals

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1.3.1 Comparison to group 1

- Higher melting points
- Higher densities
- Stronger and harder
- Much less reactive, less vigorously with oxygen and water

★ **1.3.2 Typical properties**

- Form ions with different charges
- Useful as catalysts
- Form coloured compounds

Al	White
Ca	White
Mg	White
Cu^{2+}	Pale greenish blue
Fe^{2+}	Green
Fe^{3+}	Dark-brown

- White = colourless in solutions

2.1 Chemical bonds, structures & properties of matter

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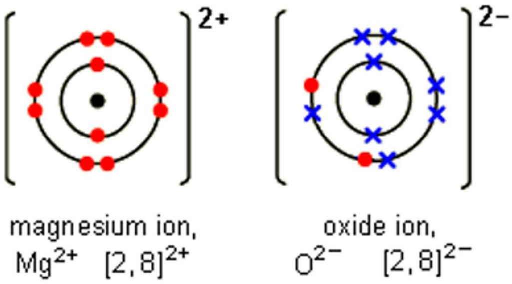
2.1.1 Types of bonding

- Metallic: metal atoms
- Ionic: metals and non-metals, between oppositely charged ions
- Covalent: non-metal atoms, sharing pairs of electrons
- Electrostatic forces of attraction
  - A force of attraction between particles with opposite charges

2.1.2 Ionic bonds

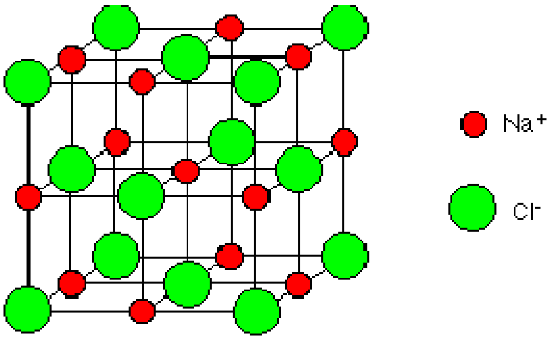
- Definition
  - The strong electrostatic attraction between oppositely charged ions
- Electrons are transferred from metal atoms to non-metals atoms
  - Metals atoms lose electrons and become cations (write equation)
  - Electrons transferred to non-metal atoms
  - Non-metal atoms gain electrons and become anions (write equation)
  - Atoms achieve a full outer shell - stable
  - Ions are attracted to each other and bond together by strong ionic bond
- Non-metals change their name to '-ide ions'
- Dot and cross diagram
  - 

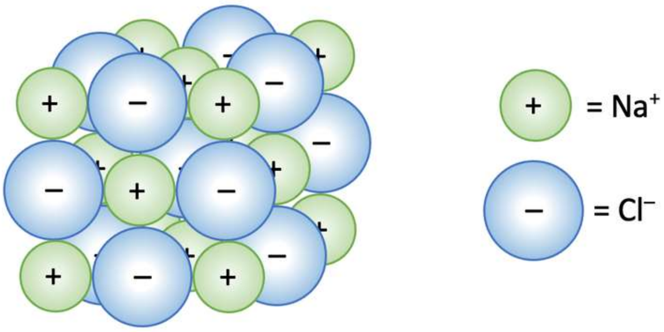
  - Different symbols for electrons in different elements
- Complex ions to remember

Silver	Ag^+	Hydroxide	OH^-
Zinc	Zn^{2+}	Nitrate	NO_3^-
Lead	Pb^{2+}	Sulfate	SO_4^{2-}
Ammonium	NH_4^+	Carbonate	CO_3^{2-}

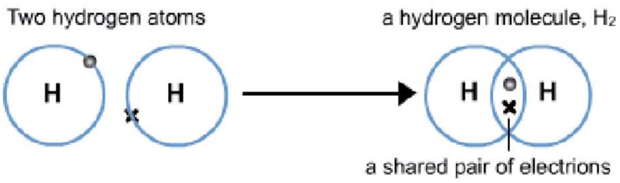
2.1.3 Ionic structures

- An ionic compound is a giant structure of ions known as a lattice
- Ionic compounds are held together by strong ionic bond, which is the strong electrostatic forces of attraction between oppositely charged ions
- These forces act in all directions in the lattice
- Sodium chloride structure

- 

- 

2.1.4a Covalent bonding

- Definition
  - The attraction between the shared pair of electrons and the nuclei of the bonded atoms
- Non-metals share their outermost electrons when they bond together
- Each shared pair of electrons is a covalent bond
- Most covalent substances consist of relatively small molecules
  - They always have the same number and type of atoms in every molecule
- Dot and cross diagram
  - 

  - Usually only show the outermost electron shells for the elements

2.1.4b Polymers

- Large covalent molecules
- Many small monomers joining together
- Do not contain a fixed number of atoms
- Double covalent bond between carbon atoms is broken down into single bond to bond with other carbon atoms
- Allotrope
  - different structures of the same element
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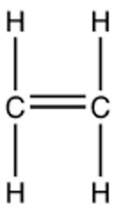
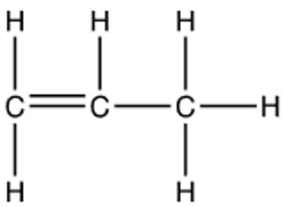
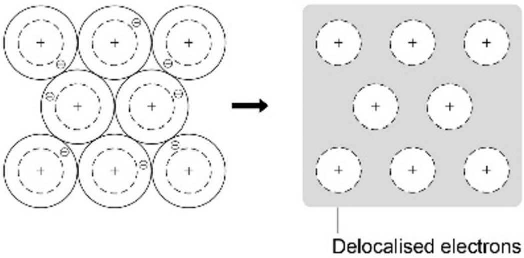
Ethene	
Propene	

2.1.4c Giant covalent structures

- A very small number of non-metals can form lattices with variable number of atoms rather than small molecules
- E.g. silicon (Si), silicon dioxide (SiO_2), silicon carbide (SiC), diamond, graphite

2.1.5 Metallic bonding

- Definition
  - The attraction between the positive metal cations and the shared delocalised electrons between them
- Electrons in the outer shell of metal atoms are delocalised
  - Delocalised electrons: Electrons that are not associated with a particular atom
  - Gives strong bond - metals have high melting points
- Creates a giant structure (lattice with a variable number of atoms)
  - Atoms are arranged in a regular pattern
- Conduct electricity
  - Delocalised electrons carry charge / current through the structure
- Diagrams



2.2 How properties depend on bonding

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2.2.1 Kinetic theory

- Solid
  - Particles vibrate on the spot
  - Close together
  - Regular pattern
  - Strong attractive forces
- Liquid
  - Particles slide past each other
  - Close together
  - Random arrangement
  - Weaker attractive forces
- Gas
  - Particles move around quickly in all directions and spread out to fill all the space available
  - Far apart
  - Random arrangement
  - Force between particles are negligible
- Heating a substance gives the particles more energy
- Particles with more energy can more easily overcome the attractive forces holding them together
- Higher melting / boiling point - stronger intermolecular force between particles, so more hear energy is needed to overcome the intermolecular force

2.2.3 Ionic compounds

- Particles in ionic compounds are ions arranged in a giant ionic lattice
- High melting and boiling point
  - Ions are held together by strong ionic bonds
  - Large amounts of heat energy are needed to break the many strong ionic bonds
- They can conduct electricity when melted or dissolved
  - Ions are free to move so charge can flow through the structure
  - Solid ionic compounds cannot conduct electricity because the ions are in fixed positions so charge cannot flow

2.2.4 Simple covalent substances

- These substances consists of small molecules have relatively low melting and boiling points
  - Usually gases or liquids under room temperature
  - Weak intermolecular forces
  - Strong covalent bonds inside the molecules
  - Intermolecular forces are overcome but not strong covalent bonds when melting
  - Small amount of heat energy needed to break the weak intermolecular forces
- Intermolecular forces increases with the size of the molecules
  - Larger molecules have higher melting and boiling points
- The molecules does not have an overall charge so they don't conduct electricity

2.2.5 Polymers

- Relatively strong intermolecular forces because molecules are very large
- Solid at room temperature

2.2.6 Giant covalent structures

- Solids with high melting points
- Atoms are linked to other atoms by strong covalent bonds so lots of heat energy is needed to overcome the covalent bond

2.2.7 Metals

- Giant structures of atoms with strong metallic bond
- Most metals have high melting and boiling points
- Alloys
    - Definition
        - A mixture of two or more elements , where at least one element is a metal
    - The different sizes of atoms in an alloy distort the layers in the structure, making it more difficult for them to slide over each other, so they are harder than pure metals
    - In pure metals atoms are arranged in layers
        - Layers can easily slide over each other
        - It allows metal to be bent and shaped
        - They are too soft for many uses
    - Most metals in everyday use are alloys

2.2.8 Metal as conductors

- Good conductors of heat because the delocalised electrons in the metal carry electric charge through the metal
- Good conductors of heat because energy is transferred by the delocalised electrons

2.3 Giant structures and properties

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2.3.1 Diamond

- Structure and bonding
  - Each carbon atoms forms 4 covalent bonds with other carbon atoms
  - Forms a giant covalent structure with a variable number of atoms
- Properties
  - Hard + high melting point
    - Atoms held together by strong covalent bond
  - Does not conduct electricity
    - No free moving electrons
    - Charge cannot flow through

2.3.2 Graphite

- Structure and bonding
  - Each carbon atoms forms 3 covalent bonds with other carbon atoms in a giant covalent structure
  - Forming layers of continuous hexagonal rings of carbon atoms
  - No covalent bond between layers, only weak forces
    - Layers are free to slide over each other
    - Makes graphite soft and slippery
  - Strong covalent bonds between carbon atoms within layers
    - A high melting point - lots of heat energy needed to overcome the strong covalent bond
- Good conductor of heat and electricity
  - One electron from each carbon atom is delocalised since each carbon atom only forms 3 bonds
  - Delocalised electron between layers
  - They allows graphite to conduct thermal energy and heat energy

2.3.3.1 Graphene

- Definition
  - Graphene is a single layer of graphite
- Structure
  - A sheet of carbon atoms covalently bonded forming a continuous hexagonal layer
  - One atom thick
- Properties
  - Extremely strong but also light because of the strong covalent bond between atoms
  - Conducts electricity and heat because there are delocalised electrons
  - Useful in fabricating composite materials and in electronics

2.3.3.2 Fullerene

- Definition
  - Molecules of carbon atoms with hollow shapes
- Structure
  - Usually hexagonal rings of carbon but can also have rings of five or seven
- Buckminsterfullerene
  - The first fullerene molecule to be discovered
  - 60 carbon atoms are joined together - C_{60}
  - 12 hexagons and 20 pentagons
- Uses
  - Pharmaceutical delivery
  - Lubricants
  - Catalysts
  - Carbon nanotubes

2.3.3.3 Carbon nanotubes

- Definition
  - Cylindrical fullerenes with very high length to diameter ratio
- Properties
  - High tensile strength
  - Excellent conductors of heat and electricity
- Uses
  - Nanotechnology
  - Electronics
  - Reinforce materials e.g. tennis rackets

2.4 Nanoscience

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Nanoscience

- Nanoscience is all about substances which are made of nano-sized particles
- A nano-sized particle is between 1nm and 100nm in size

2.4.1 PM

- PM stands for particulate matter (also called particle pollution)
    - A mixture of solid particles and liquid droplets found in the air
- Some particles are large or dark enough to be seen with the naked eye
    - e.g. Dust, dirt, soot, or smoke
- Others are so small that they can only be detected by an electronic microscope
- Pollution particle size:
    - $PM_{2.5}$: very fine particles 100nm to 2500nm
    - PM_{10}: coarse particles 2500nm - 10000nm, often referred as dust
- Surface area to volume ratio
    - Write in the form of 1:n or n:1
    - As particles get smaller their surface area to volume ratio increases
        - If the side of a cube decreases by a factor of 10 the SA to Vol ratio gets 10 time bigger
    - May have different properties from the same material in bulk due to high SA to Vol ratio
    - High SA to Vol ratio means that much smaller quantities are needed for then to be effective compared to normal sized particles

2.4.2 Uses

- Sun cream
    - Traditional: zinc oxide
        - White and opaque
    - Newer sun cream or sun blocks contain Titanium dioxide nanoparticles
        - See-through and not opaque
        - Easier to spread, cover more skin and gets a better protection from UV rays
- Liposome
    - Small fatty nanoparticles that can be absorbed by the skin
    - Cosmetic companies use them to get the ingredients into the skin
    - Pharmaceutical companies are interested in it since it has a possibility to be used to deliver tiny, controlled doses of medicine directly to cells
- Catalysts
    - Catalysts only work on their surface
    - Nanoparticles have a large SA to Vol ratio so needs very small amounts
    - Examples
        - Titanium dioxide acts as a catalyst that breaks down organic matter in self-cleaning glass as a nanofilm
        - Fuel cells
- Silver nanoparticles
    - Kill germs
    - Used in wound dressings to promote healing
    - Also in deodorants and clothing
- Problems
    - Nanoparticles in sun cream can be absorbed by the skin
        - Titanium dioxide in it is a catalyst breaks down organic matter
        - May damage cells and speed up chemical reactions, causing illness or even death
        - Unsure about the damage yet
    - Nanoparticles in clothes will be released in the water when washed
        - They are released into the environment
        - Unsure about the damages to the environment

3.1 Conservation of mass

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3.1.1 Conservation of mass

- The law of conservation of mass
  - No atoms are lost or made during a chemical reaction
  - So the mass of the products equal the mass of the reactants
- Balanced equations
  - During chemical reactions the chemical bonds that hold the atoms together will be re-arranged
  - However, atoms are neither created or destroyed
  - So there must be the same number and type of atoms in both the reactants and products
  - We must always write balanced symbol equations

3.1.2 Relative formula mass

- Definition
  - RFM of a compound is the sum of the relative atomic masses of the atoms in the numbers shown in the formula
- In balanced equations
  - Sum of RFM of the reactants in the quantities shown = Sum of RFM of the products in the quantities shown

3.1.3 Change in mass

- Change in mass in measurement
  - A reactant or product is a gas and its mass is not measured
  - Gas escapes into the atmosphere

3.2 Amount of substances

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3.2.1 Moles

- Chemical amounts are measured in moles
- Symbol is mol
- The mass of 1 mol of a substance in grams is numerically equal to its M_r
- 1 mol of a substance always contain the same number of particles
    - Can be atoms, molecules or ions
- Avogadro constant
    - The number of atoms, molecules or ions in one mole of a given substance
    - Value is 6.02×10^{23} per mole
- Mass = Moles × M_r
- Moles = $\frac{\text{Mass}}{M_r}$

3.2.4 Reagents mass

- Limiting reagent
    - Limit the amount of products
    - Totally used up in the reaction
- Excess reagent (XS)
    - One reagent has some left over in the reaction
    - Ensure that all of the other reactant is used

3.3 - 3.5 Calculations

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3.3.1 Percentage yield

- Even though no atoms are lost or gained in a reaction it is not always possible to obtain the calculated amount of a product
  - The reaction may not go to completion because it is reversible
  - Some of the product may be lost when it is separated from the reaction mixture
  - Some of the reactants may react in ways different to the expected reaction
- Yield
  - The amount of a product obtained
- Percentage yield
  - When compared with the maximum theoretical amount as a percentage
  - $\% \text{ Yield} = \dfrac{\text{Mass of product actually made}}{\text{Maximum theoretical mass of product}} \times 100\%$

3.3.2 Atom economy

- Definition
  - A measure of the amount of starting materials that end up as useful products
- Importance
  - Sustainable development
  - Economic reasons to use reactions with high atom economy
- Formula
  - $\text{Atom economy} = \dfrac{\text{Relative formula mass of desired product from equation}}{\text{Sum of relative formula masses of all reactants from equation}} \times 100\%$

3.4 Concentration

- Formula
  - $\text{Concentration (mol/dm}^3\text{)} = \dfrac{\text{moles (mol)}}{\text{volume (dm}^3\text{)}}$

3.5 Volume of gases

- The volume of one mole of any gas at room temperature and pressure is 24 dm^3.

4.1 Reactivity of metals

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4.1.1 Oxidation and reduction

- Oxidation
    - The gain of oxygen / the loss of electrons
- Reduction
    - The loss of oxygen / the gain of electrons
- Oxidation and reduction always happen together

4.1.2 Reactivity of metals

- Reactivity
    - When metal react with other substances the metal atoms form positive ions
    - The reactivity of metal is its tendency to form positive ions
- Arrange in order of reactivity
    - Reaction of metals with air, water, dilute acids and other substances
    - With air
        - (Group 1) Lithium, sodium and potassium react rapidly with air so they are very reactive
        - Potassium fastest > sodium > lithium
        - Others react slowly with air
    - With water
        - Calcium is more reactive than the others
        - Followed by magnesium
        - The remaining metal react too slowly
    - With dilute acid
        - Calcium and magnesium are the most reactive
        - Zinc > iron because there is more bubbling
        - Copper does not react
- Reactivity series
    - Potassium
    - Sodium
    - Lithium
    - Calcium
    - Magnesium
    - Aluminium
    - *Carbon*
    - Zinc
    - Iron
    - Tin
    - Lead
    - *Hydrogen*
    - Copper
    - Silver
    - Gold
    - Platinum
- Displacement
    - The more reactive metal loses electrons more easily
    - Displace the less reactive metal from a compound

4.1.3 Extraction of metals

- Unreactive metals
    - Found in the Earth as the metal itself (native form)
    - e.g. gold / platinum / silver
- Most metals

    - Found as compounds that require chemical reactions to extract the metal (often oxide)
- Metals less reactive than carbon
    - Extracted from their oxides by reduction with carbon
    - Heating the metal with carbon
    - metal oxide + carbon → metal + carbon dioxide
    - The carbon removes the oxygen from the metal oxide, metal is reduced (displace metal from metal oxide)
- Other metals above carbon in the reactivity series
    - Need to extracted by electrolysis (cannot be reduced)

<u>4.1.4 REDOX</u>

- The metal has gained electrons when they are reduced
- OIL RIG
    - Oxidation is when a substance loses electrons
        - $X \rightarrow X^+ + e^-$
    - Reduction is when a substance gains electrons
        - $X^+ + e^- \rightarrow X$

4.2 Reactions of acids

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4.2.1 With metals

- Acid + metal → metal salt + hydrogen
- Redox reaction
  - Metal lose electrons e.g. $M_{(s)} \rightarrow M^{2+}_{(aq)} + 2e^-$
  - Hydrogen gain electrons e.g. $2H^+ + 2e^- \rightarrow H_2$

4.2.2 Neutralisation

- Neutralisation definition
  - The reaction between an acid and a base to form a salt and water
- Base
  - Definition
    - A substance that reacts with an acid to neutralise it and produce a salt
  - Metal oxide, metal hydroxide, metal carbonate
  - Acid + base → salt + water
- Alkali
  - Definition
    - Substance producing more hydroxide ions than hydrogen ions when dissolved in water
  - Soluble base
  - Acid + alkali → salt + water
  - e.g. $NaOH_{(aq)}$, $KOH_{(aq)}$, $NH_{3(aq)}$, $Na_2CO_{3(aq)}$
- Carbonates
  - Acid + metal carbonate → salt + water + carbon dioxide
  - Na_2CO_3 and K_2CO_3 are soluble
- Neutralisation must have water as a product

4.2.3 Making a salt

- Insoluble salt
  - Reacting 2 salt solutions together
  - Non-metal + sodium (all Na compounds are soluble)
  - Metal + nitrate (all NO_3 compounds are soluble)
- Soluble salts
  - Acid with the non-metal ion (aq) react with solid insoluble substances (e.g. metal oxide, hydroxide, carbonate)
  - (Measure acid volume with measuring cylinder)
  - Put acid in a beaker and warm gently to speed up reaction
  - Keep on adding solid to the acid + stir the mixture until no more solid reacts so solid is in excess (remaining in beaker)
  - Filter off excess solid using filter paper and funnel
  - Solution of salt produced
  - Crystallised to produce solid salts
    - Heat to point of crystallisation
    - Transfer to evaporating dish and leave for evaporation
- Soluble salts with two soluble reactants
  - Titration with indicator → exactly neutralises
  - Repeat without the indicator
- Solubility

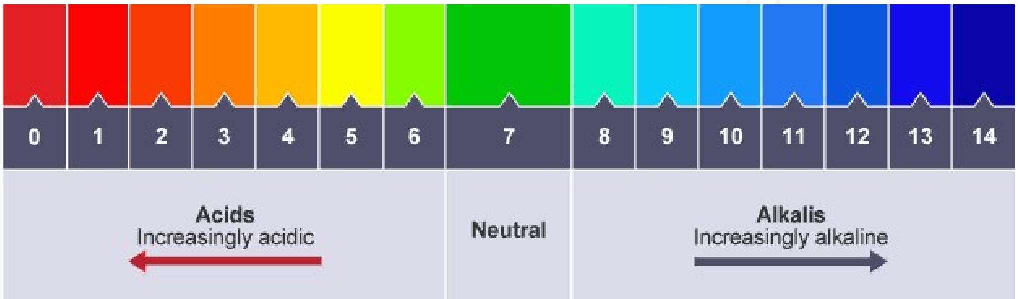
Soluble	**Insoluble**
• All common sodium, potassium and ammonium salts	• Silver chloride, lead chloride

• All nitrates • Most common chlorides • Most common sulphates • Sodium carbonate, potassium carbonate, ammonium carbonate • Sodium hydroxide, potassium hydroxide, ammonium hydroxide	• Lead sulfate, barium sulfate, calcium sulfate • Most common carbonates • Most common hydroxides

4.2.4 Acids and pH

- Definitions

Word	Definition
Acid	Substance producing more hydrogen ions than hydroxide ions when dissolved in water
Neutral	When a substance is neither acidic nor alkaline, and has a pH of 7

- pH scale
  - 0-14
  - A measure of the acidity or alkalinity of a solution
  - Can be measured using a universal indicator or a pH probe
  - Colour of universal indicator
    - 

- pH of solutions
  - Aqueous solution of acid have pH < 7
  - Aqueous solution of alkali have pH > 7
  - Solution with pH 7 = neutral
  - HCl with concentration 1 mol/dm^3 = pH 0
- Ions released
  - Acids release H^+ ions in aqueous solutions
  - Cause the universal indicator to turn red / orange
  - Alkali release OH^- ions in aqueous solutions
  - Cause the universal indicator to turn blue / purple
  - Neutralisation equation
    - $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$
- pH change and ions
  - To change the pH by one unit the concentration change by a factor of 10
  - e.g. Acid of pH 3 have 10 time many $H^+_{(aq)}$ ions as acid of pH 4
- Neutral
  - $[H^+_{(aq)}] = [OH^-_{(aq)}]$

4.2.5 Titrations

- Use
  - Find the exact volume of an acid and an alkali that react together
  - An indicator is used to find the end point
- Process
  - Measure alkali volume with pipette + safety filler and add to conical flask
  - Add a few drops of indicator
    - Phenolphthalein: pink when alkaline, colourless when neutral or acidic

    - Methyl orange: red (acidic) → orange / yellow (alkaline) → orange
  - Fill burette with acid
  - Record initial volume of acid
  - Add acid to alkali and swirl
  - Do a rough first to get approximate volume
  - Slow down when close to end point and go drop by drop
  - When indicator colour change record volume and get volume change
  - Repeat titration until you got a concordant result ($\leqslant$ 0.1 cm^3 apart)

4.2.6 Strong and weak acids

- Releasing ions
  - Strong acids
    - Fully ionised in aqueous solution
    - $HCl \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
    - e.g. HCl, H_2SO_4, HNO_3
  - Weak acids
    - Partially ionised ($\leqslant$ 1%) in aqueous solution
    - e.g. $CH_3COOH \rightleftharpoons H^+_{(aq)} + CH_3COO^-$
    - $\rightleftharpoons$ = reversible reaction, arrow point back longer
    - e.g. ethanoic, citric, carbonic acid
  - For a given concentration of aqueous solutions, the stronger an acid, the lower the pH
- Dilute
  - A relatively small amount of dissolved solute in a given volume
- Concentrated
  - A relatively large amount of dissolved solute in a given volume
- Weak acid
  - An acid that is partially dissociated (<1%) into its ions in an aqueous solution or water
- Strong acid
  - An acid that fully dissociates into its ions in water

4.3 Electrolysis

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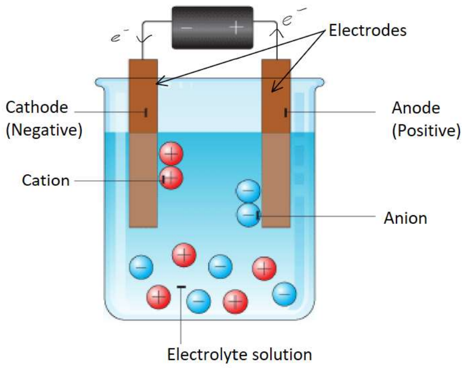
4.3.1 Electrolytes

- Electrolytes
  - Created when an ionic compound is melted or dissolved in water so the ions are free to move
  - The liquid or solution can conduct electricity and can be electrolysed
- Forms of electrolytes
  - Molten ionic substances
  - Solutions of ionic substances
  - Water
- Electrolysis
  - Using an electric current to separate a substance into its components
  - Steps
    - Passing an electric current through electrolytes cause the ions to move to the electrodes
    - Negatively charged ions are attracted the anode (positive electrode)
    - Positively charged ions are attracted the cathode (negative electrode)
    - Because electrodes attract ions of opposite charges
    - Ions are discharged at the electrodes and become neutral atoms
    - Negatively charged ions lose electrons (oxidised)
    - Electrons that are released at anode travel through the wire to the cathode
    - Positively charged ions gain electrons (reduced)
    - The atoms may combine to form molecules at both electrodes
  - 

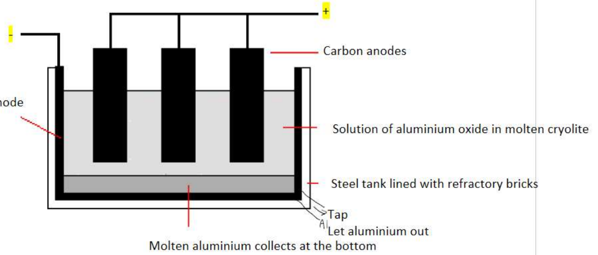
4.3.2 Electrolysis of compounds

- Inert electrode
  - Ones that do not take part in the electrolysis process
  - Include graphite and platinum
  - Used to electrolyse molten compound and break it down into its elements
    - Molten - free moving ions carry charge through the structure of the substance
  - Metal ions discharged at cathode, non-metal ions discharged at anode

4.3.3 Extraction of metals

- Extract metals by electrolysis
  - Used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon
  - Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current
- Ore
  - Compound from which a metal can be extracted for profit
- Extraction of aluminium
  - Extracted from bauxite (its ore, Al_2O_3)
  - Al_2O_3 has a very high melting point (2045℃) so melting is not practical

- Too much energy needed
- Dangerous
- Expensive
- Difficult to find a suitable container

- Al_2O_3 is dissolved in molten cryolite (Sodium aluminium fluoride, Na_3AlF_6)
    - Lower melting point, below 950℃
    - Less energy needed to melt
    - 950℃ is the temperature electrolysis is carried out
- Al formed at cathode made of graphite (lined with cell)
- Molten metal is denser than the solution so it sinks to the bottom and is periodically tapped off
- At anode oxygen is produced and reacts with carbon in graphite electrode to form CO_2
- Anode gets used up and needs to be frequently replaced
- 

4.3.4 Electrolysis of solutions

- Process
    - Both positive ions will be attracted to cathode and both negative ions will be attracted to anode
    - Only the less reactive ion will be discharged (gain / lose electrons)
- Reactivity
    - $F_2 > O_2 > Cl_2 > Br_2 > I_2$
    - Assume O_2 is from OH^- and others such as SO_4 and NO_3 don't react
    - Hydrogen - compare to reactivity series

4.3.5 Summary

- Positively charged ions gain electrons at the cathode so the reactions are reductions
- Negatively charged ions lose electrons at the anode so the reactions are oxidations
- Half equations
    - Lose: e.g. $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$
    - Gain: e.g. $2H^+ + 2e^- \rightarrow H_2$

5.1 Exothermic and endothermic reactions

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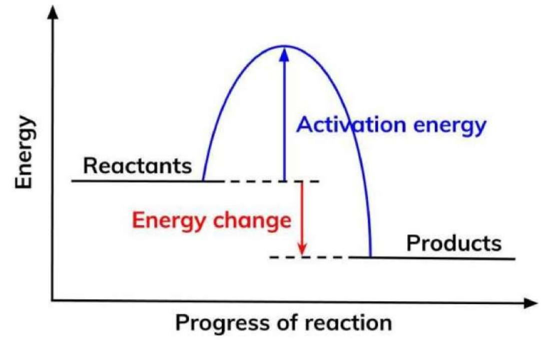
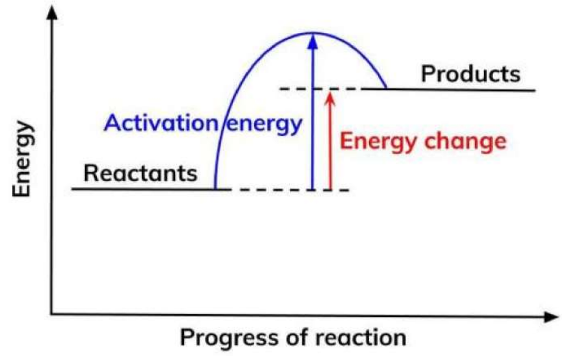
5.1.1 Conservation of energy

- Conservation of energy
  - Energy cannot be created or destroyed
  - Only converted from one form to another
  - The amount of energy in the universe is constant
- Energy is conserved during a chemical reaction
  - The amount of energy in the universe at the end of a reaction is the same as the amount of energy before the reaction
  - If a reaction transfers energy to the surroundings the product molecules must have less energy than the reactants by the amount transferred
- Exothermic
  - Energy is released to the surroundings as heat
  - Cause the temperature of both the surroundings and the mixture to increase
  - If need heating only heat at start
  - Include:
    - Combustion reactions (fuels are burned)
    - Oxidation reactions
    - Neutralisation
  - Everyday examples
    - Self-heating cans and hand warmers
- Endothermic
  - Energy is absorbed from the surroundings
  - Temperature of the surroundings fall, temperature of the mixture increase
  - The reaction may need to be heated to make it happen
    - If need heating then need to continually heating it
  - Include:
    - Thermal decompositions
    - Sherbet (reaction of citric acid and sodium hydrogencarbonate)
    - Electrolysis
  - Everyday examples
    - Sports injury packs
    - Photosynthesis
- Measuring temperature change
  - Use polystyrene cup + lid on top to reduce heat lost to the surroundings
  - Put the polystyrene cup in a beaker for stability

5.1.2 Collision theory

- Conditions for a chemical reaction to occur between two particles
  - The reacting particles must collide with enough energy and at the right angle (correct orientation)
- Activation energy
  - The minimum amount of energy needed for a reaction to happen
- Reaction profile
  - Show the relative energies of reactants and products, activation energy and the overall energy change of a reaction
  - Exothermic

- 

- Endothermic
    - 

5.1.3 Bond energy

- Energy changes during a reaction
    - Energy must be supplied to break bonds in the reactants
    - Energy is released when bonds in the products are formed
    - Overall energy change = sum of energy needed to break bonds minus sum of energy released when forming bonds
- Exothermic energy change
    - Energy released from forming new bonds > energy needed to break existing bonds
- Endothermic energy change
    - Energy needed to break existing bonds > energy released from forming new bonds

5.2 Chemical cells and fuel cells

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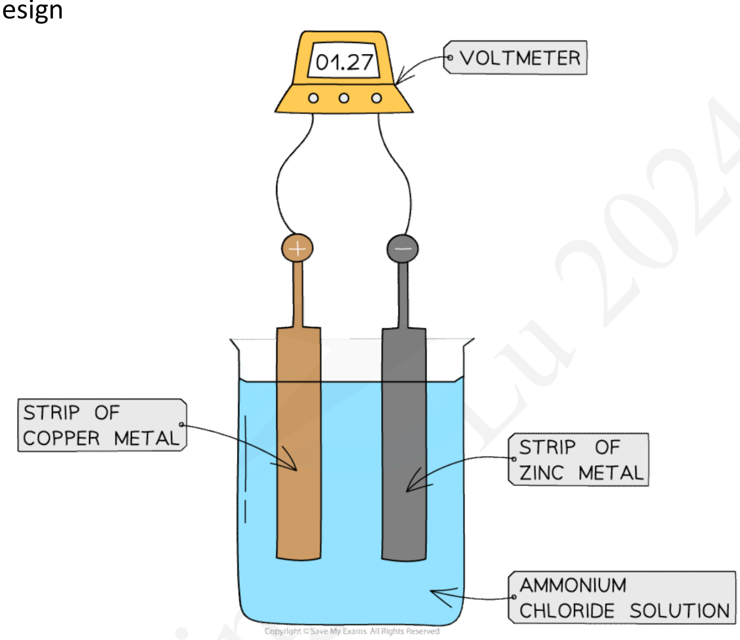
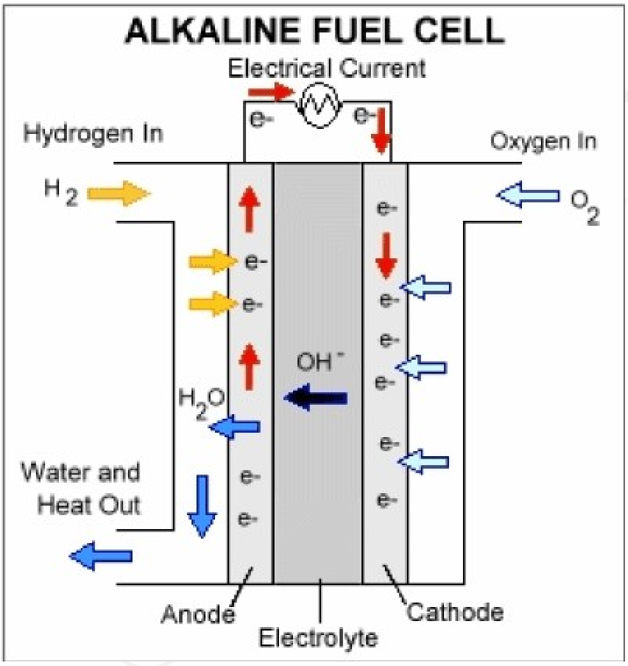
5.2.1 Cells and batteries

- Cell
  - A simple cell can be made by connecting two different metals in contact with an electrolyte so a complete circuit is formed
  - Cells contain chemicals which react to produce electricity
  - The voltage produced by a cell is dependent upon a number of factors
    - Type of electrode
    - Type of electrolyte
    - Temperature
    - Concentration of electrolyte
  - Simplest design
    - 

- Batteries
  - Consist of two or more cells connected together in series to provide a greater voltage than a single cell
- How electrode and electrolyte affect voltage
  - e.g. zinc = negative + copper = positive
  - The electrons give the more reactive electrode a negative charge as it releases electrons
  - Sets up a charge difference between the electrodes
  - The electrons then flow around the circuit to the positive electrode
  - The difference in the ability of the electrodes to release electrons / difference in reactivity causes a voltage to be produced
  - The greater the difference in the metals reactivity then the greater the voltage produced
  - The electrolyte used also affects the voltage as different ions react with the electrodes in different ways
- Non-rechargeable vs. rechargeable cells / batteries
  - Non-rechargeable
    - Chemical reactions stop when one of the reactants has been used up = no voltage produced
    - e.g. alkaline batteries
  - Rechargeable
    - Can be recharged because the chemical reactions are reversed
    - Reversed when an external electrical current is supplied

5.2.2 Fuel cells

- Fuel cell
  - Produce voltage continuously as long as they are supplied with
    - An external source of fuel (e.g. hydrogen)
    - Oxygen in air
  - Fuel is oxidised electrochemically within the fuel cell to produce a potential difference
- Hydrogen-oxygen fuel cell
  - Potential alternative to rechargeable cells and batteries
  - Hydrogen and oxygen are used to produce a voltage
  - Hydrogen is oxidised to produce water
  - Water is the only product
  - These have continuous fuel supply so do not go flat or need recharging
- Hydrogen fuel cell equations
  - Overall: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$
  - Alkaline conditions
    - Negative electrode / anode: $2H_{2(g)} + 4OH^-_{(aq)} \rightarrow 4H_2O_{(l)} + 4e^-$
    - Positive electrode / cathode: $O_{2(g)} + 2H_2O_{(l)} + 4e^- \rightarrow 4OH^-_{(aq)}$
  - Acidic conditions
    - Negative electrode / anode: $2H_{2(g)} \rightarrow 4H^+_{(aq)} + 4e^-$
    - Positive electrode / cathode: $O_{2(g)} + 4H^+_{(aq)} + 4e^- \rightarrow 2H_2O_{(l)}$
  - Diagram
    - 

- Comparing energy cells

	Non-rechargeable cells	Rechargeable cells	Fuel cells
Example uses	• Alkaline cells	• Cells in remote controls and clocks • Lithium-ion cells e.g. in electronic devices	• Hydrogen fuel cell (e.g. cars, spacecraft)
Advantages	• Cheap to manufacture	• Can be recharged many times before being recycled, reducing the use of resources • Cheaper than fuel cells • Cheaper than non-rechargeable cells in the long run	• Easy to maintain as there are no moving parts • Water is the only waste product → no pollution • Continuous supply of electricity (as long as there is a continuous supply of fuel)
Disadvantages	• Disposal of lots of used cells which contain toxic chemicals	• More expensive than non-rechargeable cells to manufacture	• Very expensive to manufacture • Need a constant supply of

	• End up in land fill sites = pollution • Expensive to recycle	• Takes time to recharge • Can only be recharged for a certain times before running out	hydrogen fuel = flammable gas • Fossil fuels may be used to make hydrogen → pollution