Chapter 1: Experimental Chemistry

How are physical quantities measured?

Time
-SI unit: s (seconds)
-60 seconds = 1 minute
-Measured with stopwatch

Temperature
-SI unit: K (kelvin)
-Temperature in kelvin = Temperature in celsius + 273
-Measured with thermometer

Length
-SI unit: m (metre)
-1 kilometre = 1000 metres
-Measured with measuring tape, ruler

Mass
-SI unit: kg (kilogram)
-1 kilogram = 1000 grams
-Measured with beam, electronic balance

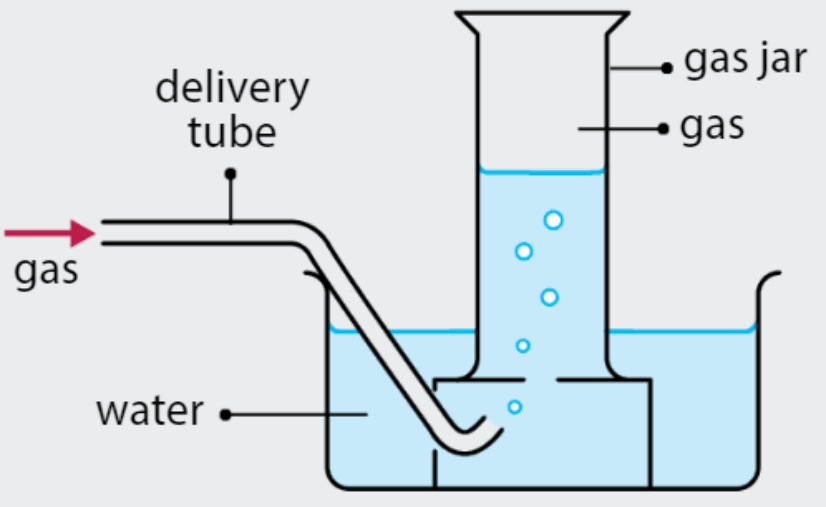
Volume
-SI unit: m^3 (cubic metre)
-1 cubic metre = 100 millilitres
-Measured with pipette (accurate fixed volume up to around 25 cm^3), volumetric flask (big volume up to around 250 cm^3), measuring cylinder (1 d.p. around 20-50 cm^3), burette (2 d.p. 50 cm^3)
-Volume of gases are measured with a gas syringe

Solubility of gases

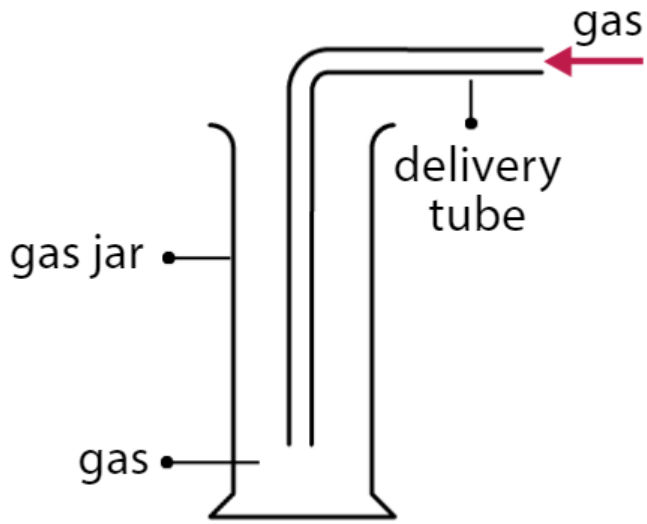
-Gases can dissolve in water
-Insoluble gases: H_2, N_2, CH_4
-Slightly soluble gases: O_2 and CO_2
-Very soluble gases: SO_2, Cl_2, NH_3

How are gases collected?

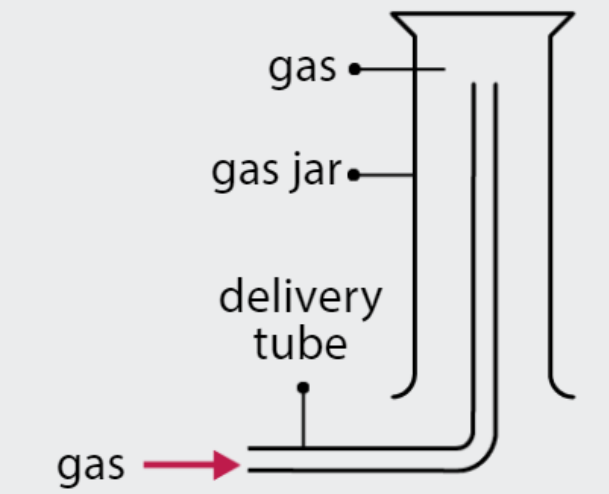
Water displacement
-Gas is bubbled through water and collected in a gas jar
-The gas can only be insoluble or slightly soluble in water
-Density does not affect this method, but note that density of water is 1 g/cm^3
-e.g. hydrogen, oxygen, carbon dioxide



Downward delivery
-Gas is collected in a gas jar and sinks down if denser than air
-Solubility in water does not affect this method
-Must be denser than air
-The Mr of air is 29
-e.g. chlorine, hydrogen chloride, sulfur dioxide



Upward delivery
-Gas is collected in a gas jar and floats up if less dense than air
-Solubility in water does not affect this method
-Must be less dense than air
-The Mr of air is 29
-e.g. ammonia



Calculating the density of gases

-Relative atomic mass (Ar): number of protons and neutrons in an atom
-The relative atomic mass of air is around 29
-Relative molecular mass (Mr): number of protons and neutrons of all atoms in a molecule
-Density = mass / volume

Methods for drying gases

Concentrated sulfuric acid
-Most gases can be dried
-Gases which react with sulfuric acid cannot be used
-The moist gas must be bubbled through the sulfuric acid

Quicklime (calcium oxide)
-Most basic and neutral gases can be used, such as ammonia
-Gases which react to calcium oxide cannot be used
-The gas must be passed through the calcium oxide
-Calcium oxide readily absorbs moisture from the air, so it must be freshly heated before use

Fused calcium chloride (fused means dried)
-Hydrogen, nitrogen, carbon dioxide is regularly used
-Gases which react with calcium chloride cannot be used
-The gas must be passed through the calcium chloride
-Calcium chloride readily absorbs moisture from the air, so it must be freshly heated before use

How are substances in mixtures separated?

What is a mixture?
-A mixture consists of two or more substances that are not chemically combined
-The physical properties of a substance determines the separation techniques used

Seperating solid-solid mixtures

Using magnetic attraction
-Magnetic substances can be separated from non magnetic substances
-Magnetic substances: iron, cobalt, nickel, some alloys
-Waste in a landfill is also sorted this way

Sieving
-A sieve is used to separate solids with different particle sizes
-Large particles can't pass through the smaller pores while small particles can

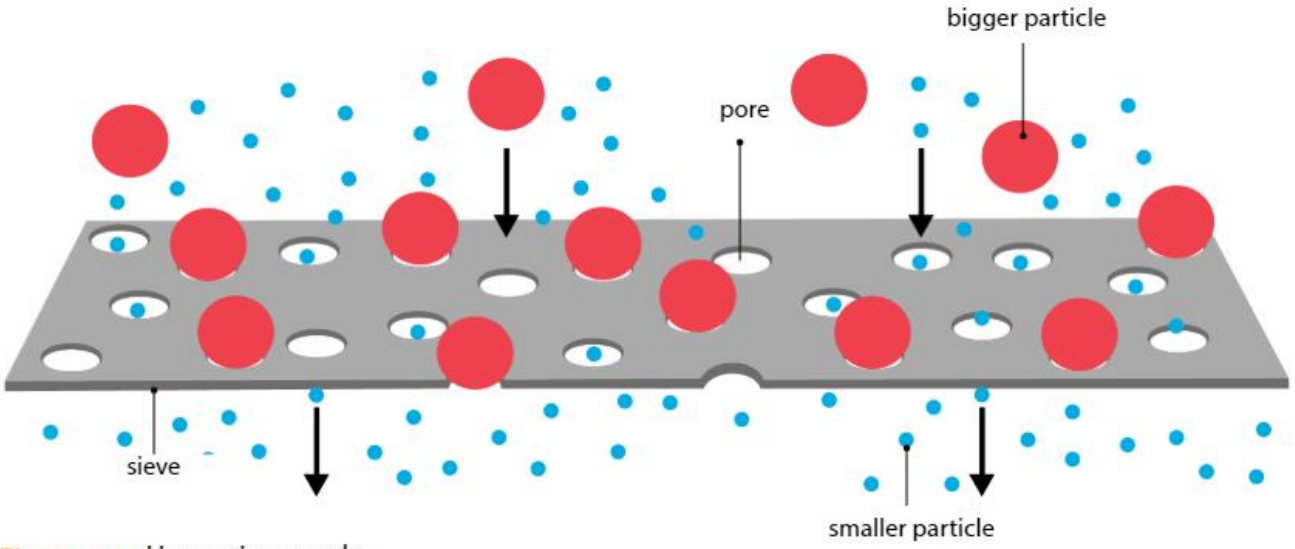


Figure 1.15 How a sieve works

Using suitable solvents
-Solids of different solubility in the solvent can be separated
-Solvent: liquid that dissolves solids
-Solute: solid that dissolves in the solvent
-A solute dissolves in a solvent

Sublimation
-Sublimation is used to separate solids where one substance changes from the solid to the gaseous state directly
-Substances that sublime: naphtalene, ammonium chloride, iodine

Separating solid-liquid mixtures

FIltration
-Seperates an insoluble solid from a liquid
-An insoluble solid is one that is unable to dissolve into the liquid it is placed in
-Filtrate: liquid that passes through the filter
-Residue: solid that remains on the filter

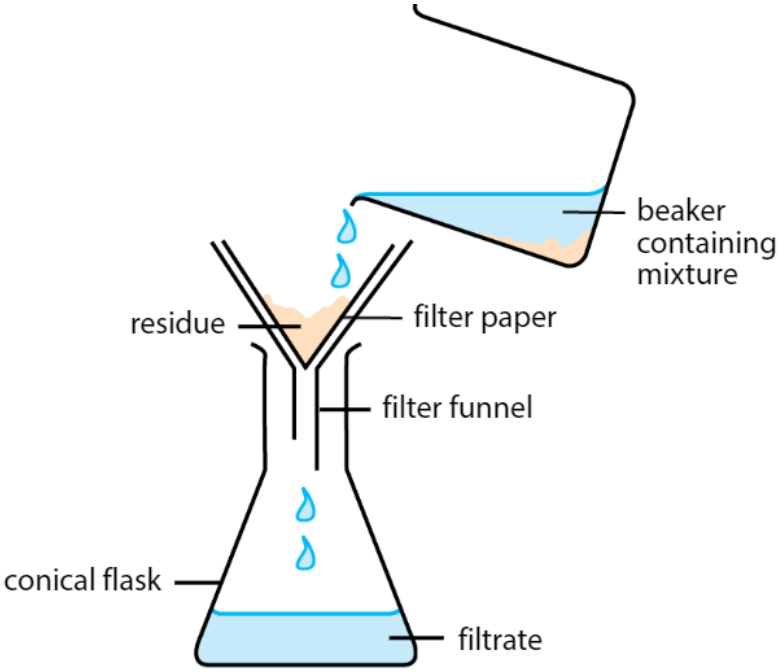


Figure 1.19 Filtration in the laboratory

Evaporation to dryness
-Seperates a dissolved solid from its solvent
-The mixture is heated until all the solvent has vaporised

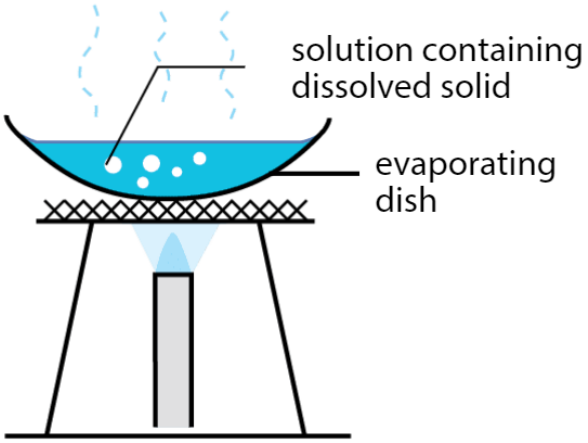


Figure 1.22 Evaporating a solution to dryness

Crystallisation
-Seperates a pure solid from its saturated solution
-A saturated solution is one in which no more solute can be dissolved
-A solution is heated until it is saturated

-It is then cooled gradually until crystals start to form
-The crystals are then filtered from the solvent, washed with distilled water and then dried on filter paper

Simple distillation
-Seperates a pure solvent (liquid) from its solution

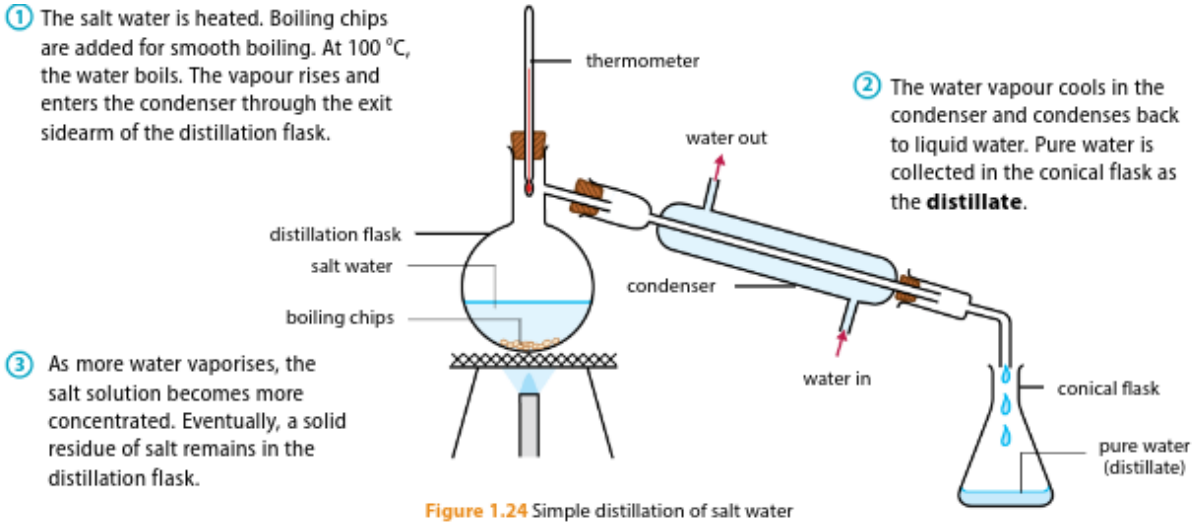


Figure 1.24 Simple distillation of salt water

-Simple distillation can only be used when there is a significant difference in boiling point

Miscible and immiscible liquids

Miscible liquids
-Miscible liquids are liquids which form a uniform (homogeneous) solution when mixed together

Immiscible liquids
-Immiscible liquids are liquids which form a non-uniform (heterogenous) mixture when mixed together

Seperating liquid-liquid mixtures

Seperating funnel
-A separating funnel is used to separate immiscible liquids
-When a heterogenous mixture is left to settle into different layers in a separating funnel, each layer can be removed by opening the tap at the bottom of the flask and collecting it in separate flasks or beakers
-The layers are separated by density, with the densest liquid at the bottom and the lest dense liquid at the top

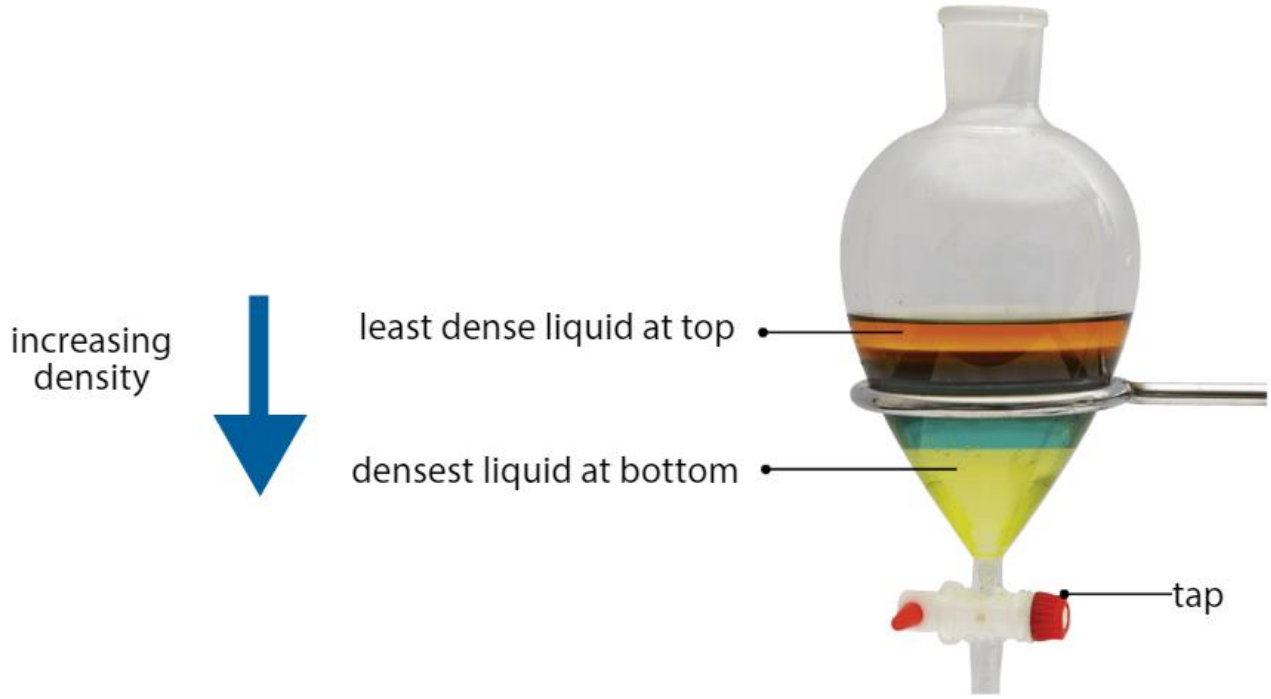


Figure 1.27 Separating funnel with four immiscible liquids

Chromatography

-Chromatography can be used to separate miscible liquids

-A drop of a homogeneous solution is placed at the edge of a piece of paper, and indicated by a line drawn with a pencil as the start

-It is then dipped into a solvent, keeping the solvent just above the start level

-As the solvent is absorbed by the paper, the solution is separated into pure components

-They are separated by solubility in the solvent, with the most soluble at the top and the lest soluble at the bottom

-The furthest point the solvent reache is called the solvent front

-The Rf value of a substance is the ratio of the distance moved by a substance and the distance moved by the solvent until it reaches the solvent front

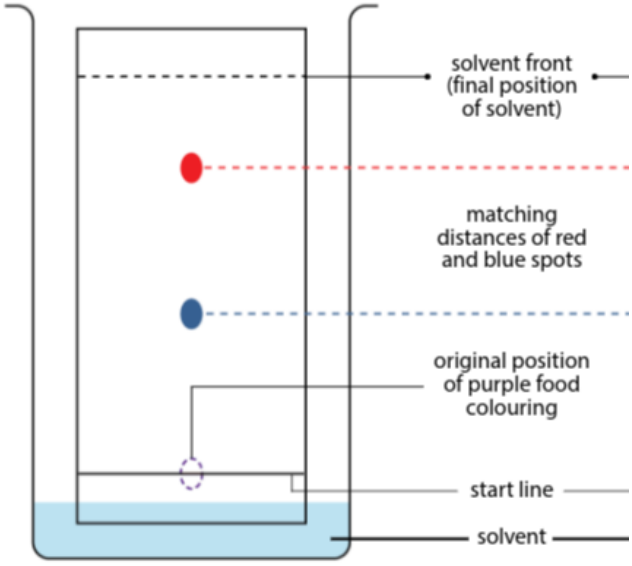


Figure 1.28 Chromatogram of a purple food colouring

Fractional distillation

-When two liquids are miscible, the separating funnel cannot be used to separate them, and the difference in boiling point of the two liquids may be too small for simple distillation to work

-Thus, a fractionating column is added to the distillation set up to separate the two liquids

-The fractionating column contains small glass beads or other small solid objects to provide a larger surface area for vapours to condense onto

Figure 1.35 shows the fractional distillation of an ethanol–water mixture.

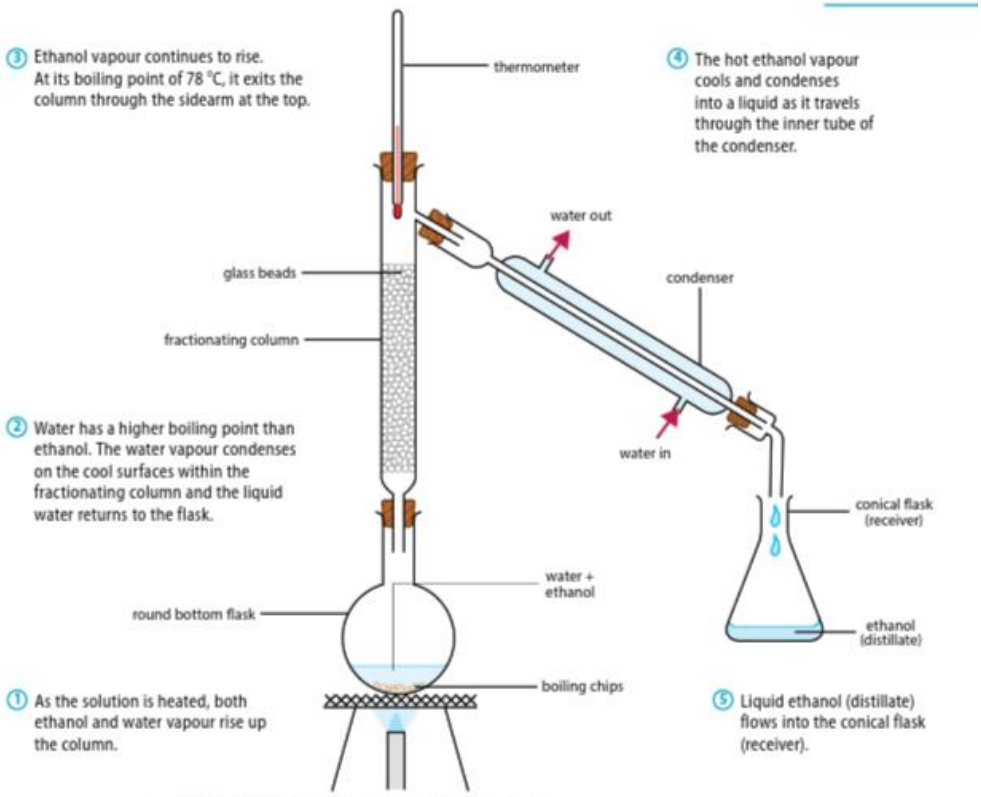


Figure 1.35 Fractional distillation of ethanol and water

How can the purity of substances be determined?

-Mixtures melt or boil at a range of temperatures, thus the greater amount of impurities in a substances, the larger the change in boiling or melting points
-Impurities will lower the melting point and raise the boiling point
-Other than chromatography, melting or boiling point data can be used to determine the purity of substances

Chapter 2: Kinetic Particle Theory

Kinetic particle theory

-The kinetic particle theory states that all matter is made up of tiny particles, which are in constant random motion
-These particles are called atoms, molecules or ions

Solids, liquids and gases

-Based on the kinetic particle theory, adding or removing energy affects the forces experienced by the particles, thus determining their physical state

Solids

-Particles of a solid are very closely packed in an orderly manner, thus they only vibrate or rotate about their fixed positions and have low kinetic energy
-These particles have experience strong attractive forces, so a lot of energy is need to break up their orderly arrangement
-They have a definite volume and shape
-They undergo melting when being converted to a liquid, and undergo sublimation when converted to a gas

Liquids

-Particles of a liquid are more spaced apart compared to solids, so their attractive force is less strong
-Their arrangement is disorderly, unlike solids, and can slide past one another freely
-They have a definite volume, but no definite shape, thus taking the shape of their container
-They undergo freezing when being converted to a solid, and undergo boiling or evaporation when converted to a gas

Gases

-Particles of a gas are very spaced apart, having very weak attractive forces
-Their arrangement is disorderly, and the particles are spaced very far apart
-The particles move around quickly and randomly, and in any direction
-They have no definite volume or shape, thus taking the shape of their container and being able to be compressed
-They undergo condensation when being converted to a liquid

Changes of state

-When a substance is heated, thermal energy is transferred to the substance
-Some of the energy is then converted to kinetic energy, increasing the temperature of the substance, and vice versa when it is cooled
-At certain temperatures, heating or cooling a substances results in a change in state rather than a change in temperature
-These temperatures are known as transition temperatures

-If the temperature and time of a substances is plotted on a graph, where temperature is the y axis and time is the x axis, it is called a heating curve
-When the substance is undergoing a change of state, the temperature of the substance remains constant until all of the substance has changed states as energy is absorbed to overcome the intermolecular forces between the particles or because the particles are attracted to each other

Expansion and contraction

-When a substance is heated below its melting point, the particles cannot spread out freely due to the strong attractive force
-Thus, they vibrate more quickly about their positions, but with a slightly wider spacing than before, leading to expansion
-When a substance is cooled the opposite happens and the particles come closer than before, leading to contraction

Movement of particles

-In liquids, the dissolved particles of a substance will move in the liquid through diffusion
-In gases, the solid particles will collide with the gas particles and move around through diffusion
-Diffusion is the net movement of particles from a region of higher concentration to a region of lower concentration (bio definition with concentration gradient is accepted in chem exam, but chem definition is not accepted in bio exam)

Factors affecting diffusion

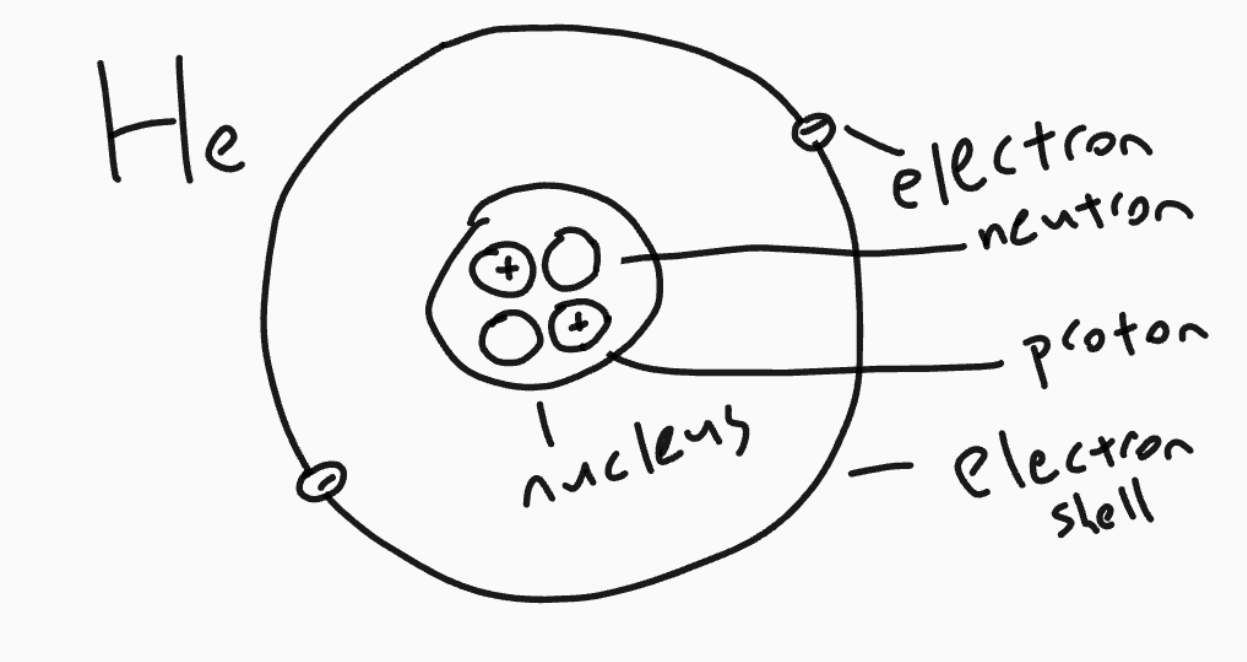
Temperature
-If a particle has a higher temperature, it has more kinetic energy, thus moving around more quickly
-Thus, the higher the temperature, the faster the rate of diffusion

Molecular or atomic mass
-Particles with a larger mass require more kinetic energy to move at a given speed
-Thus, the higher the molecular or atomic mass, the slower the rate of diffusion

Chapter 3: Atomic Structure

What is an atom made up of?

-An atom is the smallest particle that can have the chemical characteristics of an element
-However, there are even smaller particles inside the atom called sub atomic particles
-Different elements have different properties as they have different amounts of sub atomic particles



Sub atomic particles

-Atoms are electrically neutral as the charges of the proton and electron cancel out
-Number of protons = number of electrons
-Protons and nucleus are found in the nucleus, and are thus called nucleons
-An atom has affixed amount of protons, however it can have a different number of neutrons to form different isotopes
-If an atom gains an electron, it becomes a negative ion (anion)
-If an atom loses an electron, it becomes a positive ion (cation)

Sub atomic particle	Relative mass	Relative charge	Location in an atom
proton	1	+1	nucleus
neutron	1	0	nucleus
electron	1/1840 / negligible	-1	electron shell

How many sub atomic particles does an atom have?

Proton (atomic) number (Z)
-Number of protons

Nucleon (mass) number (A)
-Number of protons + number of neutrons

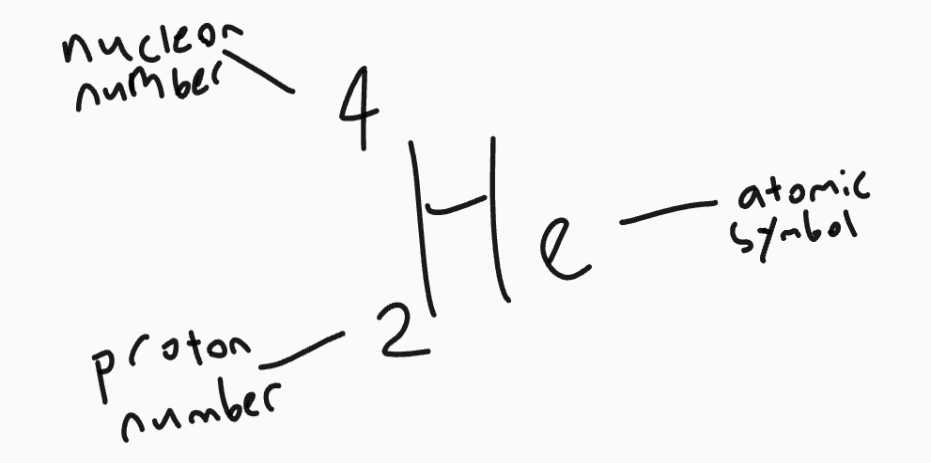
Neutron number
-Nucleon number - proton number

Ar
-Ar is the relative atomic mass of an atom
-It is the nucleon number

Mr
-Mr is the relative molecular mass of an atom
-It is the total nucleon number of all atoms in a molecule (2 or more atoms chemically bonded together)

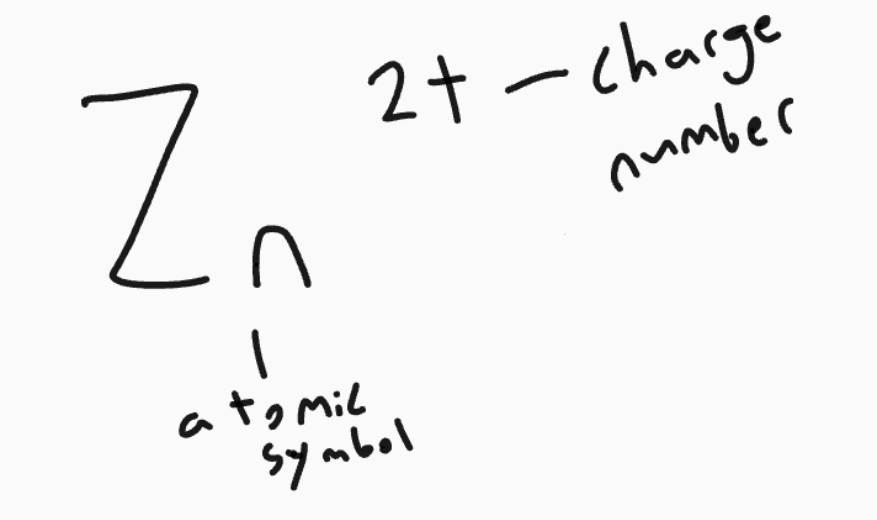
Nuclide notation

-Nuclide notation shows the proton number and nucleon number



Ions

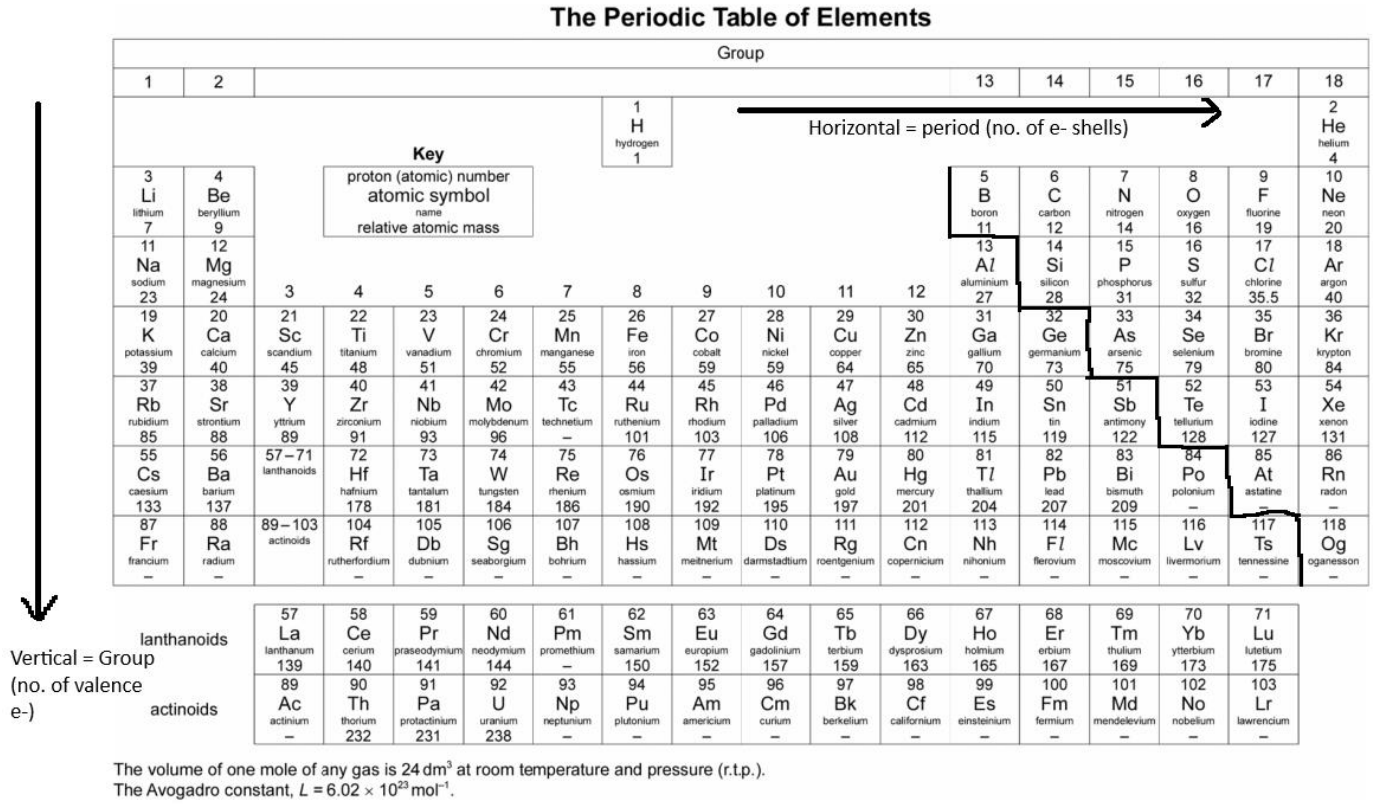
-An ion is the particle formed by the replacement of a hydrogen atom in an acid
-We use charge numbers to indicate the number of electrons that are gained or lost
-A positive ion is called a cation, while a negative ion is called an anion
-Metals usually lose electrons while non metals usually gain electrons
-Thus metals usually form cations while non metals usually form anions



Isotopes

-Isotopes are atoms of the same elements that have the same proton number but different nucleon numbers
-Since the proton number is the same, the isotopes belong to the same element
-Most elements occur naturally as isotopes
-The nucleon number used in the periodic table is found by finding the relative abundance of all isotopes of an element on earth
-e.g. out of all chlorine on earth, 75% is chlorine-35 and 25% is chlorine-37, thus [75(35) + 25(37)] / 100 = 35.5 (nucleon number on periodic table)
-Formula for relative abundance: [p1(m1) + p2(m2)] / 100, where p is percentage on earth and m is nucleon number
-They have the same chemical properties, but different physical properties as they have the same amount of protons but different amounts of neutrons

Reading the periodic table



How are sub atomic particles distributed in an atom?

-Electrons are constantly moving around the nucleus in zones called electron shells
-Electrons in the outermost shell (valence shell) possess the most energy
-Inner electron shells are filled first
-The first electron shell can hold 2 electrons, and the second and third can usually hold 8 electrons each
-When the outermost shell is fully filled, the atom is stable and unreactive (inert)
-The electrons in the valence shell are known as valence electrons

Electronic configuration

-Electronic configuration is the way you write the amount of electrons in an atom
-e.g. the electronic configuration of oxygen is 2.6 as there are 2 electrons in the first shell and 6 in the second shell

Chapter 4: Chemical Bonding

Why do atoms combine?

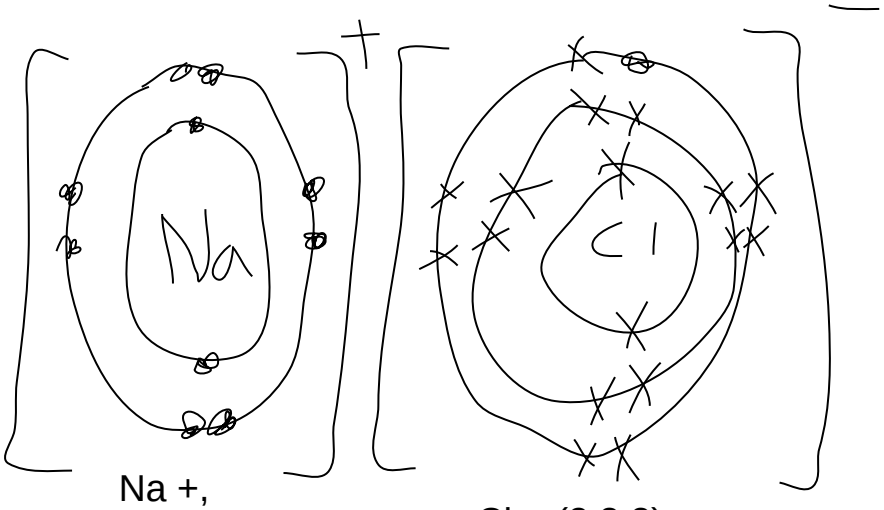
-Some atoms which do not have full valence electron shells combine to achieve the stable electronic configuration of a noble gas
-This occurs through the sharing, gain or loss of electrons

Noble gases

-Noble gases occupy the last group of the periodic table
-They are considered chemically unreactive as they have a full valence electron shell
-They are monoatomic

Ionic bonding

-When a metallic atom reacts with a non metallic atom, the metallic atoms give their electrons to the non metallic atom. This forms two ions which both have the electronic configuration of a noble gas.
-When an atom loses an electron, it becomes a positive cation. Likewise, if an atom gains an electron, it becomes a negative anion.
-If an atom loses two electrons, it becomes a “2+” cation and vice versa.
-If the metallic atoms has not enough electrons to give, then add more of that atom.
-Chlorine has an electronic configuration of 2.8.7. The nearest noble gas has an electronic configuration of 2.8.8. Thus, it needs one electron to attain that electronic configuration. When reacting with sodium, which has an electronic configuration of 2.8.1, it loses one electron to chlorine, attaining the electronic configuration of a noble gas, or 2.8. Chlorine then becomes an anion and sodium becomes a cation.



Na +, (2.8)

Cl -, (2.8.8)

Common cations

Relative charge	Name of cation	Formula
+1	hydrogen	H+
+1	sodium	Na+
+1	potassium	K+
+1	silver	Ag+
+1	all group 1 alkali metals and sometimes hydrogen	nil
+1	ammonium	NH4+ (there are 4 hydrogen atoms, not a charge of +4)
+2	magnesium	Mg2+1
+2	calcium	Ca2+
+2	barium	Ba2+
+2	iron (II)	Fe2+
+2	copper (II)	Cu2+
+2	zinc	Zn2+
+2	lead (II)	Pb2+
+2	all group 2 alkaline earth metals	nil
+3	iron (III)	Fe3+
+3	all group 13 elements	nil
+3	aluminium	Al3+
+4	some group 14 elements	nil

Common anions

Relative charge	Name of anion	Formula
-1	fluoride	F-
-1	chloride	Cl-

-1	bromide	Br-
-1	iodide	I-
-1	hydroxide	OH-
-1	nitrate	NO3- (there are 3 oxygen atoms, not a charge of -3)
-1	manganate (VIII)	MnO4- (there are 4 oxygen atoms, not a charge of -4)
-1	all group 17 halogens and sometimes hydrogen	nil
-2	oxide	O2-
-2	carbonate	CO32- (there are 3 oxygen atoms and a charge of -2, not a charge of -32)
-2	sulfate	SO42- (there are 4 oxygen atoms and a charge of -2, not a charge of -42)
-2	all group 16 elements	nil
-3	phosphate	PO43- (there are 4 oxygen atoms and a charge of -3, not a charge of -43)
-3	all group 15 elements	nil
-4	some group 14 elements	nil

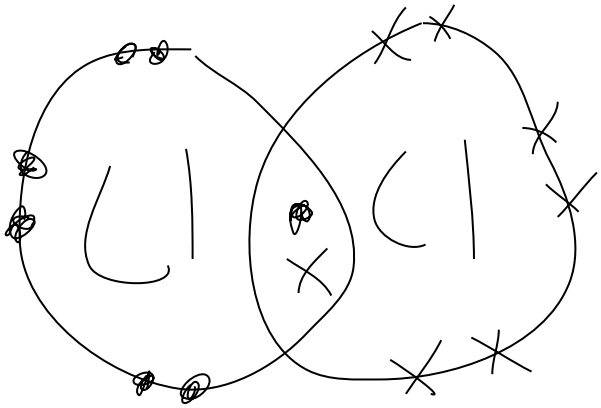
(note that in exam, the number for amount of atoms is written as a subscript while charge is a superscript)

Ionic compounds

-Ionic compounds are two ions held together by ionic bonds, or the strong electrostatic forces between positive and negative ions
-Most ionic compounds are formed by at least one metal cation and one non metal anion
-They have no net charge as the charges cancel out
-Like all compounds, they are formed in a fixed ratio

Covalent bonding

-Non metallic atoms, when taking part in chemical reactions,will try to attain the electronic configuration of noble gases by sharing electrons.
-The number of electrons shared is determined by the number of valence electrons, and the number shared is the same.
-Chlorine has an electronic configuration of 2.8.7. The nearest noble gas has an electronic configuration of 2.8.8. Thus, if two chlorine atoms were to take part in covalent bonding, they would each share on electron so that both atoms attain the electronic configuration of 2.8.8.



Covalent compounds

-Covalent compounds are two or more atoms or molecules held together by either weak intermolecular forces of attraction or strong covalent bonds
-Most covalent compounds do not contain metal atoms
-They have no net charge as the charges cancel out
-Like all compounds, they are formed in a fixed ratio

Chapter 5: Structure and Properties of Materials

Elements

-An element is a pure substance which cannot be broken down into simpler substances by chemical methods
-They can appear as atoms when not bonded to other atoms, molecules when bonded to another of the same atom, or a lattice in a giant metallic structure
-They have a fixed melting and boiling point

Compounds

-A compound is a pure substance which consists of two or more elements chemically combined in a fixed ratio
-The elements can bond through ionic or covalent bonding
-Not all molecules are compounds, as hydrogen gas (H_2) is an element which exists as a molecule
, while carbon dioxide (CO_2) is a compound
-They can be separated, although with difficulty, through processes such as heat decomposition or electrolysis
-They are formed from a chemical reaction and have different properties as compared to their constituent elements
-They have a fixed melting and boiling point

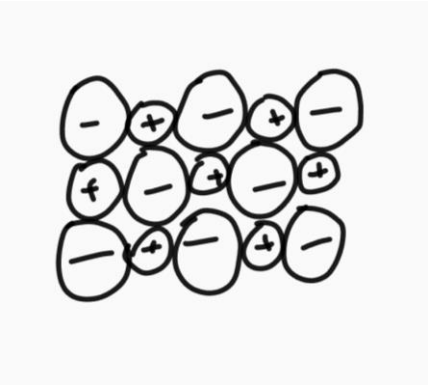
Mixtures

-A mixture is an impure substances consisting of elements, molecules or compounds physically combined together without a fixed ratio
-For example, air is a mixture as it contains various gases, such as hydrogen, oxygen, nitrogen and carbon dioxide
-The gases in air are not combined in a fixed ratio, as air at higher elevations may have different ratio of gases
-Generally, they can be easily separated through heating or filtration
-However, alloys require a large amount of energy to separate
-They are usually formed from physical mixing and have similar properties to their constituent elements, molecules or compounds
-They have a range of melting and boiling points

Giant Ionic lattice structures (GILS)

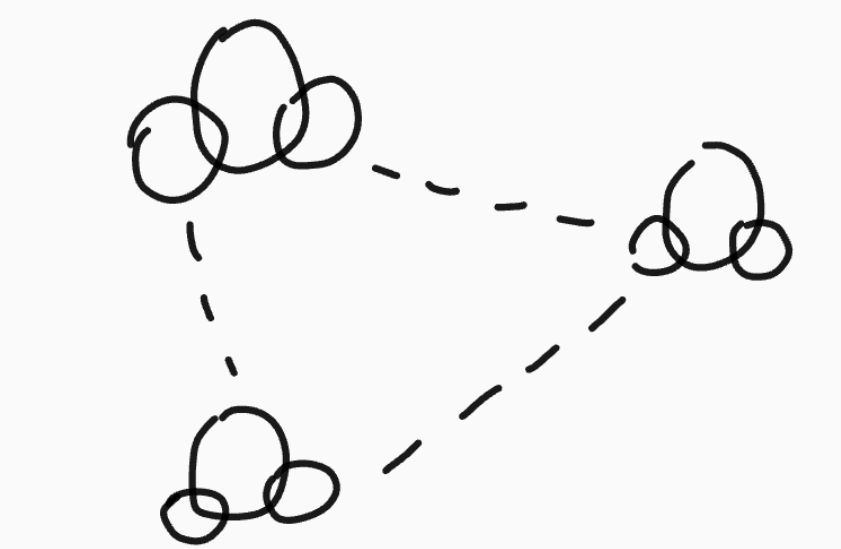
-In the solid state, ionic compounds exist as a 3D structure known as a giant ionic lattice structure
-Each ion is placed next to an ion with the opposite charge
-Thus, each ion is surrounded by 6 ions of an opposite charge
-They have a regular shape as the ions are arranged in an orderly and regular manner
-They are very hard but brittle as the strong electrostatic forces of attraction hold them together

-However, once these bonds are broken, the ions move away from each other and atoms of the same charge move towards each other, increasing the repulsive force between the ions and eventually shattering the structure when the repulsive forces are stronger than the attractive forces
-They have a high melting and boiling point as a lot of heat energy is needed to overcome the strong electrostatic forces of attraction holding them together
-When a giant ionic lattice structure is made up of doubly charged ions (+2, -2), they have a higher melting and boiling point as compared to singly charged ions
-They are soluble in water as the water molecules are attracted to the ions, weakening the electrostatic forces between the ions
-The ions are thus pulled from the lattice structure, dissolving to form an aqueous solution
-They are not soluble in organic solvents
-They can conduct electricity in molten and aqueous states but not solid states, as the rigid lattice structure is broken down and the charged ions are allowed to move around and carry charges, while in solid state, the ions are not able to move around



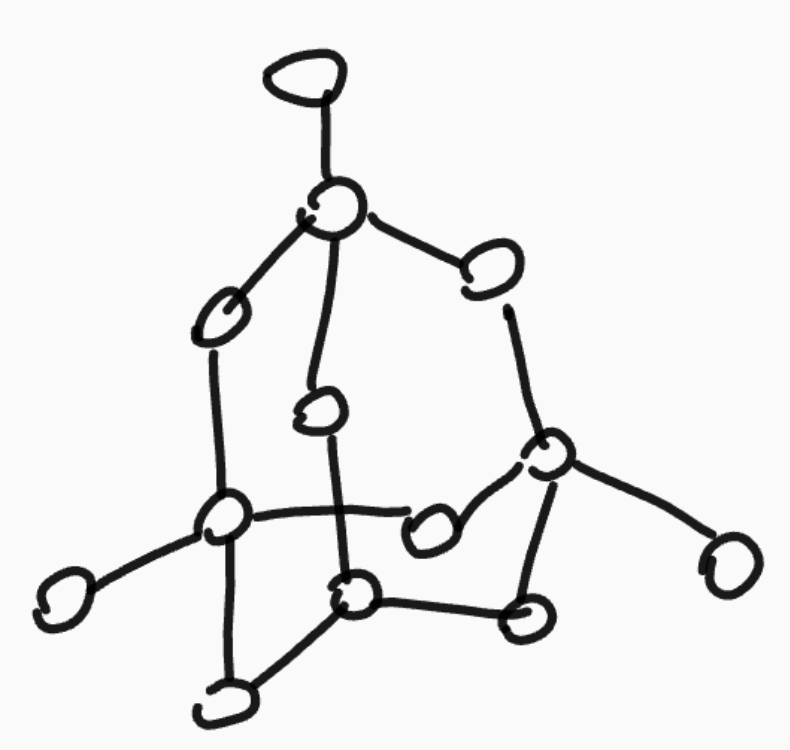
Simple molecular substances (SMS)

-SImple molecular substances, or simple covalent substances, are formed by covalent molecules
-There are weak intermolecular forces between the molecules in the structure of the simple molecular substance
-Thus, they have low melting and boiling points as little heat energy is needed to overcome these forces and are quite volatile (melting and boiling point close to room temperature)
-The larger the size of the molecules, which is proportional to the number of electron shells, the stronger the intermolecular forces of attraction and the higher the melting and boiling points
-They are insulators of electricity as each valence electrons of an atom are used in covalent bonding, so there are no mobile charge carriers
-However, acids which are covalent compounds, such as hydrogen chloride and hydrogen fluoride, can dissolve in water and dissociate into their respective ions, will be able to carry charges and conduct electricity
-However, these compounds are not ionic



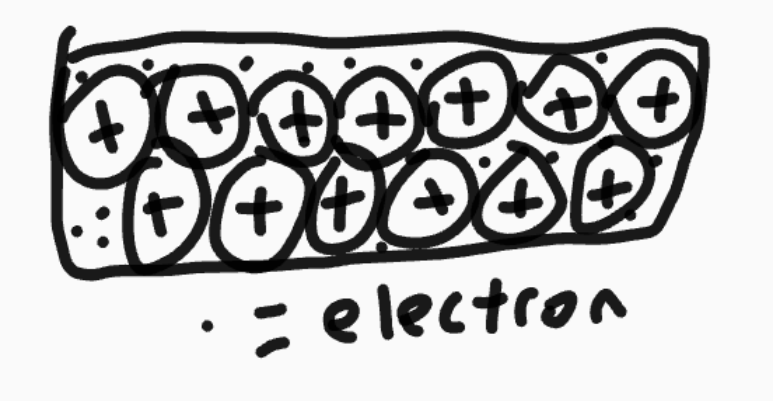
Giant molecular substances (GMS)

-Giant molecular substances, or giant covalent substances, are formed by covalent molecules or individual atoms
-They are joined by strong covalent bonds, unlike simple molecular substances
-Common giant molecular substances include diamond, graphite, silicon and silicon dioxide
-They have a high melting and boiling point as a large amount of energy is needed to break the strong covalent bonds
-They are insoluble in water and organic solvents as the forces of attraction between the solvent molecules and the atoms of the giant molecular substance are not strong enough to break the strong covalent bonds
-They are insulators of electricity as each valence electrons of an atom are used in covalent bonding, so there are no mobile charge carriers
-They are very hard due to the strong covalent bonds between the atoms
-However, graphite can conduct electricity as only 3 of the 4 valence electrons are used for bonding, and the extra valence electron can move and carry a charge
-Graphite is also quite soft as graphite is made up of different layers of carbon atoms, joined by weak intermolecular forces of attraction, which can be overcome easily when a force is applied



Giant metallic substances

-Giant metallic substances are described as a lattice of metal cations, surrounded by a sea of delocalised valence electrons
-These electrons can move and carry a charge, so metals have high electrical and thermal conductivity
-Metals also have a high melting and boiling point, are malleable and ductile, are sonorous (makes a ringing sound when striked), are shiny, are dense, form alkaline oxides and react by losing electrons to form cations
-They have high melting and boiling points as a lot of energy is needed to overcome the strong electrostatic forces of attraction between the metal cations and surrounding sea of delocalised valence electrons
-They are also ductile and malleable as the atoms of pure metals are of the same size and packed in orderly layers, allowing the layers of atoms to slide over each other easily when force is applied



Alloys

-Alloys are a mixture of a metal with another element
-Brass (copper and zinc) and steel (iron and carbon) are both alloys
-They are usually stronger and harder, have increased corrosive resistance and have a lowered melting point (when an impurity is added to a substance, the melting point decreases while the boiling point increases) compared to pure metals, thus they are more commonly used compared to pure metals
-Alloys are strong and hard due to the different sizes of atoms disrupting the orderly layers of atoms and making it difficult for the layers to slide over each other when a force is applied

Chapter 6: Chemical Formulae and Equations

Anions

-Anions are negatively charged non metal ions
-They gain electrons when undergoing ionic bonding with a cation
-When writing the name of the anion in a ionic compound, it has the name of the base element but replaced with -ide at the end
-For example, an ion of fluorine is known as a fluoride ion

Cations

-Cations are positively charged ions, usually made from metals, but not always
-They lose electrons when undergoing ionic bonding with a anion
-Most transition metals (group 3-12) and post transition metals (group 13-15 metals) form a cation with more than one possible charge, which is depicted in brackets with the charge in roman numerals
-For example, a Cu2+ cation is written as copper (II)
-However, some transition and post transition metals only form cations with one charge, such as silver and zinc
-These are written by their names only

Polyatomic ions

-Polyatomic ions are formed from molecules which have gained or lost electrons
-Negatively charged polyatomic ions are also written with an -ide at the end
-When writing more than one polyatomic ion bonded togther, it is written with brackets
-For example, calcium hydroxide is Ca(OH)2
-A compound that contains a polyatomic ion containing oxygen ends with an -ate
-Almost all polyatomic ions are negatively charged

Group number relative to charge of ion

Group number	Charge of ion
1	+
2	+2
13	+3
14	+4 / -4
15	-3
16	-2
17	-

Common formulae

Name	Formula
hydrochloric acid	HCl
sulfuric acid	H2SO4
nitric acid	HNO3
sodium hydroxide	NaOH
potassium hydroxide	KOH
calcium hydroxide	Ca(OH)2
hydroxide ion	OH-
water	H2O
ammonia (not to be confused with ammonium)	NH3
methane	CH4
sulfate ion	SO42- (there are 4 oxygen atoms and a charge of 2-, not a charge of 42-)
carbonate ion	CO32- (there are 3 oxygen atoms and a charge of 2-, not a charge of 32-)
nitrate ion	NO3- (there are 3 oxygen atoms, not a charge of 3-)
phosphate ion	PO43- (there are 4 oxygen atoms and a charge of 3-, not a charge of 43-)
ammonium ion	NH4+ (there are 4 hydrogen atoms, not a charge of 4+)
zinc ion	Zn2+
silver ion	Ag+
copper (II) ion	Cu2+
lead (II) ion	Pb2+
iron (II) ion	Fe2+
iron (III) ion	Fe3+

Writing ionic compounds

-Write down the charges of the cation and the ion
-Add cation or anions in order for the charges to cancel out
-Always write the cation before the anion
-For example, calcium has a charge of 2+ and bromine has a charge of -
-So, you need 2 bromine atoms to cancel out the 2+ charge
-They then form CaBr2, or calcium bromide
-This also works for polyatomic ions
-For example, zinc has a charge of 2+ and a nitrate ion has a charge of -
-So, you need 2 nitrate ions to cancel out the 2+ charge
-They then form Zn(NO3)2, or zinc nitrate

Writing covalent compounds

-Writing covalent compounds is quite similar to ionic compounds, except they need a prefix

Prefix	Number of atoms
mono-	1
di-	2
tri-	3
tetra-	4

-For example, for BCl3, there is 1 boron atom and 3 chlorine atoms
-Thus, it is written as boron trichloride (there is no mono prefix for boron as it is the first element)

Chemical equations

-A chemical equation is an equation which shows a chemical reaction and the reactants and products
-The reactants are always written on the left while the products are on the right
-The equation is written with an arrow instead of the equal sign
-There are state symbols written next to the chemicals showing the state they are in
-Common state symbols are s (solid), l (liquid), g (gas) and aq (aqueous)
-Aqueous means a solid dissolved in a water to form a solution
-For example, when reacting hydrochloric acid and sodium hydroxide solution, it is written as

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H2O\ (l)$

Balancing chemical equations

-When writing chemical equations, the atoms on each side of the equation must be balanced, meaning that there must be the same amount of each atom
-This can be done by writing the unbalanced chemical equation down, then writing the amount of each atom on each side

-If the number of atoms do not match, add molecules of the unbalanced atoms to the left side until they are balanced
-Only add molecules and not singular atoms

-For example, the equation Na (s) + Cl2 (g) → NaCl (s) is not balanced

-There is 1 sodium atom and 2 chlorine atoms on the left side, while there is 1 sodium and chlorine atom on the right side
-Thus, 1 sodium atom needs to be added to the left side

-The balanced equation is then written as 2Na (s) + Cl2 (g) → 2NaCl (s)

-In general, you should balance the metals first, then the non metals, then hydrogen and oxygen

Ionic equations

-When writing ionic equations, the aqueous substances are split up into their respective ions, and the solid, liquid and gaseous substances are left the same
-If an ion appears on both sides of the equation, they are cancelled out

-For example, when sulfuric acid reacts with barium nitrate, they have the balanced equation:

H2SO4 (aq) + Ba(NO3)2 (aq) → BaSO4 (s) + 2HNO3 (aq)

-Since H2SO4, Ba(NO3)2 and 2HNO3 are in aqueous state, we split them up into 2H+, SO42-, Ba2+, 2NO3-, 2H+ and 2NO3- respectively
-2H+ and 2NO3- appears on both sides of the equation, so we can cancel them out

-Then we are left with the ionic equation: SO42- (aq) + Ba2+ (aq) → BaSO4 (s)

Chapter 7: Mole Concept and Stoichiometry

Relative atomic mass

-The relative atomic mass (Ar) is the average mass of one atom of an element relative to 1/12 the mass of an atom of carbon-12

Relative molecular mass

-The relative molecular mass (Mr) of a molecular substance is the average mass of one molecule of that substance relative to 1/12 of the mass of an atom of carbon-12

Percentage mass

-Percentage mass is the percentage of a certain element in a compound
-Formula for percentage mass: [(number of atoms of the element x Ar of the element) / Mr of compound] x 100%

What is a mole?

-A mole is the amount of substance which contains 6.02 x 10^23 particles
-SI unit: mol (mole)
-A particle can be atoms, ions, molecules or even sub atomic particles

Avogadro number

-The avogadro number or avogadro constant is defined as the number of atoms in a 12 g sample of carbon-12
-The avogadro number is equal to 6.02 x 10^23

Moles and mass

-The mass of an atom or molecule is equal to its relative atomic mass or relative molecular mass in grams
-Formula for number of moles: mass / Ar or Mr
-Formula for mass: number of moles x Ar or Mr

Molar gas volume

-One mole of any gas at room temperature and pressure (r.t.p) occupies a volume of 24000 cm^3
-Molar gas volume at r.t.p is 24000 cm^3
-Thus, equal volumes of all gases at r.t.p contain the same number of particles
-Formula for volume of any gas in cm^3: number of moles x molar gas volume in cm^3

Empirical formula

-The empirical formula of any compound has the general formula AxBy

-To find the empirical formula, we find the number of moles of each atom in the formula, then find the ratios of each element in the formula
-For example, 9.6 g of carbon, 25.6 g of oxygen and 0.8 g of hydrogen are in a compound
-No. of mol of C = 9.6 / 12 = 0.8 mol
-No. of mol of O = 25.6 / 16 = 1.6 mol
-No. of mol of H = 0.8 / 1 = 0.8 mol
-Ratio of C:O:H = 0.8:1.6:0.8 = 1:2:1
-Thus the compound is CHO2
-When writing the formula, carbon is written first, followed by hydrogen and the rest of the elements in alphabetic order of chemical symbols
-If carbon is not present, all elements including hydrogen are written in alphabetic order of chemical symbols
-Do note that for gases, do not use the molecular mass of each element if they are diatomic, always use the atomic mass instead

Molecular formula

-The molecular formula of any compound has the general formula n(AxBy), where n is any number
-Formula for n: Mr of molecule / Mr from empirical formula
-For example, the empirical formula of propane is CH2, and the relative molecular mass of propane is 42
-Mr of empirical formula = 12 + 2 = 14
-n = 42 / 14 = 3
-Thus, the molecular formula of CH2 = (CH2)3 = C3H6

Moles and equations

-In chemical equations, the number of molecules of a reactant is equal to molar ratio of the reactants
-For example, in the reaction $2H_2 + O_2 \rightarrow 2H_2O$, there are 2 molecules of hydrogen gas and one molecule of oxygen gas, so the molar ratio of hydrogen:oxygen:water is 2:1:2
-Using this, we can find the mass of a product given a fixed amount of reactant
-For example, if there are 10 g of hydrogen gas, we can find the mass of water formed
-Number of moles of H_2 = 10 / 2 = 5
-Mole ratio of $H_2:H_2O$ = 2:2
-Number of moles of H_2O = 5
-Mass of H_2O = 5 x 18 = 90 g

Percentage purity

-Most chemicals used in a chemical reaction are almost never 100% pure, as they will have some impurities in them

-Formula for percentage purity of a substance: (mass of pure substance used in reaction / mass of impure substance) x 100%

-For example, if 12 g of impure carbon is burned in the presence of oxygen to form 30 g of pure carbon dioxide in the reaction C + O2 → CO2, we can find the percentage purity of the carbon

-Number of moles of CO2 formed = 30 / 44 = 0.682 mol (3s.f.)
-Actual number of moles of C reacted = 0.682 mol (molar ratio of C:CO2 is 1:1)
-Actual mass of C reacted = 0.682 x 12 = 8.18 g (3s.f.)
-Percentage purity of C = (8.18 / 12) x 100% = 68.2% (3s.f.)

Percentage yield

-In a chemical reaction, the actual amount of product will always be less than the theoretical amount of product
-Formula for percentage yield of a substance: (actual yield / theoretical yield) x 100%

-For example, 1.92 g of magnesium metal was heated in the presence of oxygen to form 3 g of magnesium oxide in the reaction 2Mg + O2 → 2MgO

-Number of moles of Mg = 1.92 / 24 = 0.08 mol
-Number of moles of MgO = 0.08 mol (molar ratio of Mg:MgO is 2:2)
-Theoretical yield of MgO = 0.08 x 40 = 3.2 g
-Percentage yield = (3 / 3.2) x 100% = 93.8% (3s.f.)

Moles and solution

-When a soluble substance is dissolved in a liquid the substance is known as the solute, the liquid is known as the sovent and the resultant liquid is called a solution
-In general, we usually use water as the solvent in this chapter, and thus the solution is aqueous
-The concentration of a solution is recorded in mol/dm^3 or g/dm^3
-1 dm^3 = 1000 cm^3 = 1000 ml
-Formula for concentration in mol/dm^3 = no. of moles of substance / volume of solution in dm^3
-Formula for concentration in g/dm^3 = concentration in mol/dm^3 x Mr of solute
-If an acid is monobasic, 1 mol/dm^3 of acid will contain 1 mol/dm^3 of H+ ions, and 1 mol/dm^3 of a dibasic acid will contain 2 mol/dm^3 of H+ ions and so on

Limiting reactant

-When doing chemical reactions, chemists will limit the amount of a certain reactant to ensure it stops at a certain point
-The reactant with fewer moles is the limiting reactant, and the reactant with more moles is the excess reactant

-For example in the reaction Zn + S → ZnS, where 25 g of zinc is used and 30 g of sulfur is used, we can find the limiting reactant

-Number of moles of ZnS is all Zn is used = 25 / 65 = 0.385 mol (3s.f.)

-Number of moles of ZnS if all S is used = 30 / 32 = 0.938 mol (3s.f.)
-Since zinc forms fewer moles of zinc sulfide, zinc is the limiting reactant and sulfur is the excess reactant

Moles and titration

-When conducting a titration, we can find an unknown value in the titration using this formula: [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of acid = [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of alkali

-For example, if 30 cm^3 of sodium hydroxide is used to neutralise 25 cm^3 of 0.1 mol/dm^3 of sulfuric acid in the formula $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$, we can determine the concentration of the sodium hydroxide using the formula above

-By plugging the values into the formula, we get [(0.1 x 25) / 1] = [(M x 30) / 2]
-When solving for M, which is concentration, we get 0.167 mol/dm^3 (3s.f.)

Summary of formulas

-No. of moles (solid, liquid) = mass in grams / Ar or Mr
-Mass in grams (solid, liquid) = no. of moles x Ar or Mr
-No. of particles = no. of moles x 6.02 x 10^23
-Volume of gas in dm^3 = no. of moles x 24
-Volume of gas in dm^3 = (mass in grams / Mr) x 24
-No. of moles (gas) = volume in dm^3 / 24
-Mass in grams (gas) = (volume in dm^3 / 24) x Mr
-Empirical formula = AxBy
-Molecular formula = n(AxBy)
-n = Mr of molecule / Mr from empirical formula
-Percentage purity = (mass of pure substance used in reaction / mass of impure substance) x 100%
-Percentage yield = (actual yield / theoretical yield) x 100%
-Concentration in mol/dm^3 = no. of moles of substance / volume of solution in dm^3
-Concentration in g/dm^3 = concentration in mol/dm^3 x Mr of solute
-[(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of acid = [(concentration in mol/dm^3 x volume in cm^3) / no. of moles in chemical equation] of alkali

Chapter 8: Acids and Bases

What is an acid?

-An acid is a substance which dissociates/ionises in water to form H+ or hydrogen ions
-They have a sour taste and are generally corrosive
-they are electrolytes, which means they can conduct electricity in aqueous solutions
-They turn moist litmus paper red and turn green universal indicator orange if weak or red if strong
-They have a pH below 7

Basicity of acids

-The basicity of an acid is the maximum number of H+ ions produced per mole of acid
-A monobasic acid will produce 1 H+ ion, a dibasic acid will produce 2, a tribasic acid will produce 3 and so on

Mineral vs organic acids

-A mineral acid is an acid which does not occur naturally and is usually man made
-Examples include sulfuric, hydrochloric and nitric acid
-These acids are strong, corrosive and for laboratory and industrial use
-Organic acids occur naturally in fruits and plants
-Examples include citric, ethanoic and tartaric acid
-They are weak, less corrosive and can be used for flavouring food and cooking

Strength vs concentration

-The strength of an acid refers to the extent to which an acid molecule dissociates to forms H+ ions
-A strong acid will dissociate completely to form a high concentration of H+ ions, while a weak acid will dissociate partially to form a low concentration of H+ ions
-The strength of an acid cannot be changed
-However, the concentration of a solution refers to the amount of solute dissolved into 1 dm^3 of water
-A solution can be either dilute or concentrated
-Concentration can be changed by adding more or less solute

Test for acids

Test	Result
2-3 drops of universal indicator	universal indicator turns from green to orange or red depending on the strength of the acid
2-3 drops of methyl orange	methyl orange turns from orange to red

add 2 g of calcium carbonate to the acid	effervescence observed, and carbon dioxide gas is produced
add a piece of magnesium metal to the acid	effervescence observed, and hydrogen gas is produced
pH meter	pH recorded is below 7
a few drops of acid is added to damp blue and red litmus paper	the blue litmus paper turns red while the red litmus paper remains red

Reactions with acids

-Acid + base → salt + water

-Acid + alkali → salt + water

-Acid + carbonate → salt + carbon dioxide + water

-Acid + metal → salt + hydrogen gas

-So for a reaction with Na, NaOH and Na2CO3 with HCl, the equations will be HCl (aq) + NaOH (aq) → NaCl (aq) + H2O (l), 2HCl (aq) + Na2CO3 (aq) → 2NaCl (aq) + CO2 (g) + H2O (l) and 2HCl (aq) + 2Na (s) → 2NaCl (aq) + H2 (g)

What are alkalis and bases?

-Bases are chemical substances which include all metal oxides and hydroxides
-Alkalis are substances which dissociate/ionise in water to from OH- or hydroxide ions
-Alkalis are basically soluble bases
-Alkalis are soapy and corrosive when concentrated
-They turn red litmus paper blue
-They have a bitter taste
-They have a pH above 7
-They turn green universal indicator purple if strong and blue if weak
-The stuff on strength and concentration can be applied to alkalis too

Test for alkalis

Test	Result
2-3 drops of universal indicator	universal indicator turns from green to purple or blue depending on strength
pH meter	pH recorded is above 7
a few drops of alkali is added to damp blue	the red litmus paper turns blue while the blue

and red litmus paper	litmus paper remains blue

Reactions with alkalis

-Alkali + acid → salt + water

-Alkali + ammonium salt → salt + ammonia gas + water

-Neutralisation reaction ionic equation: $H^+ (aq) + OH^- (aq) \rightarrow H_2O (l)$

pH

-pH measures the concentration of H+ ions in a solution
-Formula for pH: -log[H+], where log is log base 10, and H+ is the concentration of hydrogen ions in mol/dm^3

Types of oxides

Type of oxide	Contains	Reacts with	Examples
acidic (dissolves in water to form acid)	oxides of non metals	base	SO_2, CO_2, NO_2
basic (some dissolve in water to form alkali)	oxides of metals (mostly group 1 and 2)	acid	Na_2O, MgO, CaO
amphoteric	oxides of metals (near the line separating metals and non metals on the periodic table)	both acids and bases	PbO, ZnO, Al_2O_3
neutral (insoluble in water)	oxides of non metals	nothing	CO, NO, H_2O

-Amphoteric oxides will act as an acid or an alkali, so if you put it in a solution of both acids and alkalis, they will react with both

Chapter 9: Salts

What are salts?

-A salt is a chemical compound formed by the replacement of one or more hydrogen ions of acids by a metallic ion
-They usually contain at least one metallic cation and one non metallic anion

Acid vs normal salts

-Acid salts contain hydrogen ions
-Examples include sodium bicarbonate ($NaHCO_3$) and sodium bisulfate ($NaHSO_4$)
-Normal salts do not contain hydrogen ions
-Examples include sodium carbonate (Na_2CO_3) and sodium sulfate (Na_2SO_4)
Anhydrous vs hydrated salts

Anhydrous salts	Hydrated salts
do not contain water	contain water
usually a powder	usually a crystal
$CuSO_4$ anhydrous copper (II) sulfate, which is a white powder	$CuSO_4 . 5H_2O$ copper (II) sulfate pentahydrate, which is a blue crystal

-The water in hydrated salts are chemically combined with the salt, and are different from salts dissolved in water

Solubility of salts

Type of salt	Soluble salts	Insoluble salts
nitrates	all nitrates	nil
ammonium	all salts	nil
group 1 (sodium, potassium etc)	all salts	nil
chlorides, bromides, iodides (group 17 halogens)	all except lead and silver	lead and silver
sulfates	all except lead, barium and calcium	lead, barium and calcium (sparingly soluble)
carbonates	group 1, ammonium	all except group 1 and ammonium
hydroxides	group 1, ammonium (calcium,	all except group 1 and

	barium and strontium are sparingly soluble)	ammonium

Preperation of salts

-Salts can be prepared through three different methods, precipitation, titration and reaction of acid with insoluble substances

Precipitation

-Precipitation is used to prepare insoluble salts
-A fixed volume of two soluble salt solutions, where one contains the cation of the salt and one contains the anion of the salt is added and allowed to react
-The mixture is then filtered to obtain the insoluble salt
-The salt is then washed with distilled water and dried between filter paper

Titration

-Titration is used to prepare group 1 and ammonium salts
-Using a pipette, 25 cm^3 of a known concentration of aqueous alkali is added into a conical flask
-A few drops of indicator is then added to the conical flask
-Dilute acid is added from a burette to the alkali solution, swirling the flask continuously
-Once the indicator changes colour, stop adding acid as the alkali has been neutralised
-Record the volume of acid used
-Repeat the previous steps with the recorded value of acid but without adding indicator
-Heat to evaporate the solution until a third of it is left
-Cool the hot saturated solution and allow crystals to form
-Filter to obtain the crystals, and wash with cold distilled water and dry between filter paper

Reaction of acids with insoluble substances

-When preparing soluble non group 1 or ammonium salts, excess solid metal carbonate, oxide or hydroxide is added to a fixed volume of acid until the solid can no longer dissolve or no more effervescence occurs if carbonate was used
-Filter the mixture to remove the excess solid
-Heat to evaporate the solution until a third of it is left
-Cool the hot saturated solution and allow crystals to form
-FIlter to obtain the crystals, and wash with cold distilled water and dry between filter paper

Chapter 10: Ammonia

The haber process

-The haber process is the way ammonia is made industrially

-Nitrogen and hydrogen are the raw materials in the process, in the reaction $N_2 (g) + 3H_2 (g) \rightleftharpoons 2NH_3 (g)$

-The '⇌ ' symbol indicates that it is a reversible reaction, which is a reaction that can go both backwards and forwards at the same time

-Nitrogen is obtained from the fractional distillation of liquid air, and hydrogen is obtained from the cracking or breaking down of crude oil fractions

Conditions for the haber process

-Since the reaction to produce NH_3 is reversible, both the reactions $N_2 (g) + 3H_2 (g) \rightarrow 3NH_2 (g)$ and $2NH_3 (g) \rightarrow N_2 (g) + 3H_2 (g)$ occur at the same time

-The first reaction is known as the forward reaction, and the second is known as the backward reaction

-At r.t.p, the forward reaction does not occur, hence a high pressure and relatively high pressure is needed to start the reaction, along with an iron catalyst

-Since the reaction is reversible, some of the NH_3 produced can break down and decompose back into N_2 and H_2, hence pressure and temperature have to be carefully controlled to obtain the most NH_3 for the least cost

-A higher pressure leads to a higher yield and higher rate of reaction, but it is more costly to maintain a high pressure

-A lower temperature leads to a higher yield as it lowers the rate of decomposition of NH_3, but also a lower rate of reaction

-Hence, a compromise is made, and the optimal conditions for the haber process are a pressure of 250 atm, a temperature of 450°C, and in the presence of a finely divided Fe catalyst

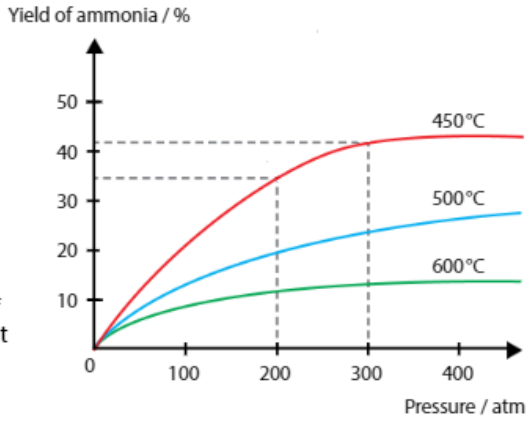


Figure 10.3 Graph showing the yield of ammonia at different temperatures and pressures

Haber process plant

-The haber process plant is the plant where NH3 is manufactured

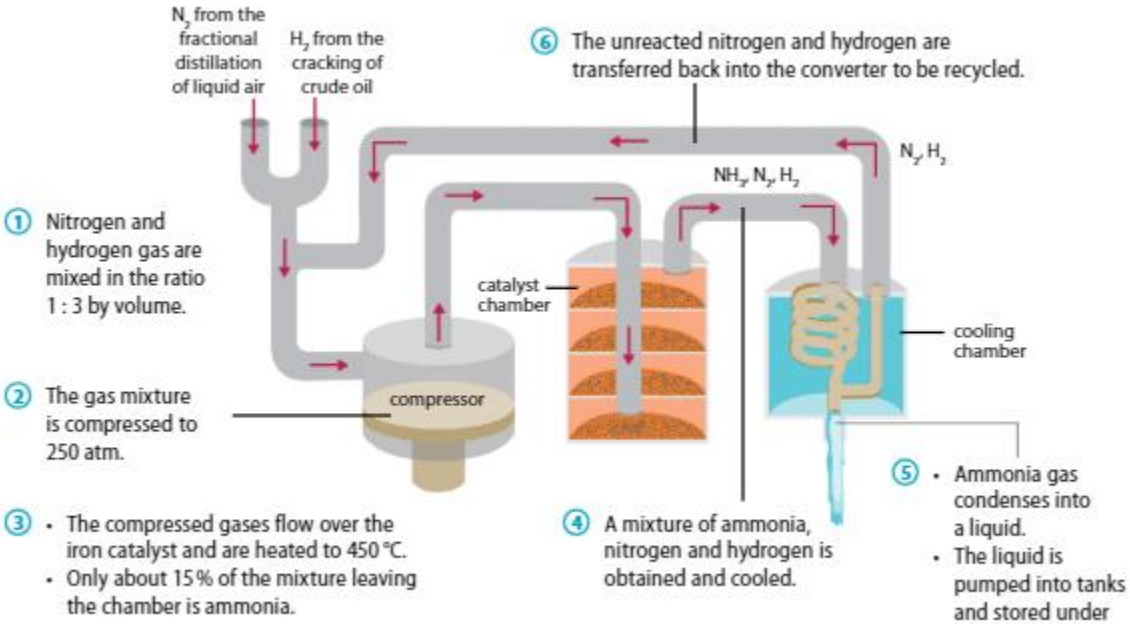


Figure 10.5 The manufacture of ammonia in a Haber process plant

Chapter 11: Qualitative Analysis

What is qualitative analysis?

-Qualitative analysis is a process of identification of cations and anions in an unknown salt
-This can be done through various tests, such as tests for gases, cations, anions, solubility, heating,oxidisng agent and reducing agent

Colour of substances

Colour	Possible identity
white solid	ammonium salts and most group 1-3 substances
black solid	copper (II) oxide, iron (II) oxide
grey solid	most solid metals, such as iron, magnesium and zinc
blue solid	most copper (II) salts, such as copper (II) sulfate and copper (II) nitrate
green solid	iron (II) salts, some copper (II) salts such as copper (II) carbonate
brown solid	iron (III) salts, copper metal

How to record a produced gas

-If a gas is liberated in a chemical reaction, the following must be recorded
-Observations (effervescence, colour, odour)
-How the gas was tested and result of the test
-Identity of gas from test

Test for gases

Gas	Colour and odour	Test	Observation	Nature of gas
H2	colourless and odourless	place a lighted splint at the mouth of the test tube	lighted splint extinguished with a pop sound	neutral
O2	colourless and odourless	insert a glowing splint into the test tube	gas relights glowing splint	neutral
CO2	colourless and odourless	bubble gas into limewater	gas forms white precipitate in limewater; if more carbon	acidic

			dioxide is bubbled into limewater, white precipitate dissolves	
Cl2	green and pungent	place moist litmus papers at the mouth of the test tube	gas turns moist blue litmus paper red and bleached/white	acidic and bleaching
SO2	colourless and pungent	place filter paper dipped in purple potassium manganate (VII) at the mouth of the test tube	gas turns purple potassium manganate (VII) colourless	acidic and reducing agent
NH3	colourless and pungent	place moist litmus papers at the mouth of the test tube	gas turns moist red litmus paper blue	alkaline

How can cations be identified?

-A cation in an aqueous salt can be identified using sodium hydroxide and aqueous ammonia
-These react with the salt to from the hydroxide of the cation, which is usually a precipitate
-The following must be recorded when testing for cations
-The colour of precipitate formed; if any
-Whether the precipitate is soluble/insolublein excess reagent (about twice the original volume of the sample)

Test for cations

Cation	NaOH	Excess NaOH	NH3 (aq)	Excess NH3 (aq)	Respective hydroxide/precipitate
group 1	no precipitate	nil	no precipitate	nil	group 1 hydroxide
NH4+	no precipitate, ammonia gas produced on warming	nil	no precipitate	nil	NH4OH
Cu2+	blue precipitate	insoluble in excess	blue precipitate	soluble in excess to form a dark blue solution	Cu(OH)2

Fe2+	dirty green precipitate, turns brown when left standing	insoluble in excess	dirty green precipitate, turns brown when left standing	insoluble in excess	Fe(OH)2
Fe3+	reddish brown precipitate, forms when Fe(OH)2 is left standing	insoluble in excess	reddish brown precipitate, forms when Fe(OH)2 is left standing	insoluble in excess	Fe(OH)3
Ca2+	white precipitate	insoluble in excess	no precipitate	nil	Ca(OH)2
Zn2+	white precipitate	soluble in excess to from a colourless solution	white precipitate	soluble in excess to from a colourless solution	Zn(OH)2
Al3+	white precipitate	soluble in excess to from a colourless solution	white precipitate	insoluble in excess	Al(OH)3
Pb2+	white precipitate	soluble in excess to from a colourless solution	white precipitate	insoluble in excess	Pb(OH)2

-To distinguish aluminum salts from lead (II) salts, potassium iodide is added to both samples
-A yellow precipitate is formed with the lead (II) salt but no visible reaction will be seen with the aluminum salt
-As a side note for practical, Al(OH)3 is formed as gelatinous precipitate, which can help differentiate it further
-However this should not be used as an explanation for why the cation is Al3+
-In general, the ionic equation of the cation test is: An+ (aq) + OH- (aq) = A(OH)n (s)

Test for anions

Anion	Test	Observation	Chemical/ionic equation
CO32-	add dilute acid and test the gas evolved	effervescence is observed, gas evolved forms white precipitate in limewater, gas is CO2	1. CO32- (aq) + 2H+ (aq) → CO2 (g) + H2O (l) 2. CO2 (g) + Ca(OH)2 (aq) → CaCO3 (s) + H2O (l)

Cl-	add dilute HNO3, followed by AgNO3 or add dilute HNO3, followed by Pb(NO3)2	a white precipitate of AgCl forms or a white precipitate of PbCl2 forms	Ag+ (aq) + Cl- (aq) → AgCl (s) or Pb2+ (aq) + 2Cl- (aq) → PbCl2 (s)
I-	add dilute HNO3, followed by AgNO3 or add dilute HNO3, followed by Pb(NO3)2	a yellow precipitate of AgI forms or a yellow precipitate of PbI2 forms	Ag+ (aq) + I- (aq) → AgI (s) or Pb2+ (aq) + 2I- (aq) → PbI2 (s)
NO3-	add alkali follwed by a piece of aluminum; warm and test for gas evolved	gas evolved turns moist red litmus paper blue, gas is NH3	1. NO3- (aq) → NH4+ (aq) 2. NH4+ (aq) + OH- (aq) → NH3 (g) + H2O (l)
SO42-	add dilute HNO3, followed by Ba(NO3)2 or add dilute HCl, followed by BaCl2	a white precipitate of BaSO4 forms	Ba2+ (aq) + SO42- (aq) → BaSO4 (s)

-Acid is added to the test for chlorides, iodides and sulfates to test for carbonates, as carbonates can give false results

Chapter 12: Oxidation and Reduction

What is oxidation?

-A substance is oxidised when it gains oxygen, loses electrons, loses hydrogen, or increases in oxidation state
-Oxidation is the opposite of reduction

Examples of oxidation

-In the reaction $2Ca + O2 \rightarrow 2CaO$, Ca is oxidised as it gains oxygen to form CaO

-In the same reaction, Ca has a charge of 0 in Ca and a charge of 2+ in CaO, hence it loses electrons and is oxidised

-In the reaction $H2S + Cl2 \rightarrow 2HCl + S$, H2S is oxidised as it loses hydrogen to form S

What is reduction?

-A substance is reduced when it loses oxygen, gains electrons, gains hydrogen, or decreases in oxidation state
-Reduction is the opposite of oxidation

Examples of reduction

-In the reaction $Zn + CuO \rightarrow ZnO + Cu$, CuO is reduced as it loses oxygen to form Cu

-In the same reaction, Cu has a charge of 2+ in CuO and a charge of 0 in Cu, hence it gains electrons and is reduced

-In the reaction $H2 + Cl2 \rightarrow 2HCl$, Cl2 is reduced as it gains hydrogen to form HCl

Redox reactions

-Reduction and oxidation occur simultaneously
-A redox reaction is a chemical reaction which involves the oxidation of a substance and the reduction of another substance

-For example, in the reaction $Mg + H2O \rightarrow MgO + H2$, Mg is oxidised as it gains oxygen and H2O is reduced as it loses oxygen, so since oxidation and reduction both occur, the reaction is a redox reaction

Oxidation states

-The oxidation state of an atom is the charge of that atom if it would have existed as an ion in a compound
-This is true even if the atom is covalently bonded
-So, the oxidation state of Cl in NaCl and HCl is the same

-When a substance loses or gains electrons, its oxidation state increases or decreases respectively

Rules of oxidation states

-The oxidation state of a free element is 0, so the oxidation state of Cu, Fe, Cl_2 and H_2 are all 0
-The oxidation state of a simple ion is the same as the charge of the ion, so the oxidation state of Na^+ is +1 and the oxidation state of Cl^- is -1
-Group 1 elements have an oxidation state of +1 in their compounds
-Group 2 elements have an oxidation state of +2 in their compounds
-Hydrogen usually has an oxidation state of +1 in their compounds, however in metal hydrides its -1, such as in NaH
-Oxygen usually has an oxidation state of -2 in their compounds, however in peroxides its -1, such as in H_2O_2
-The oxidation states of all atoms or ions in a compound add up to zero
-In polyatomic ions, the sum of all the oxidation states of atoms in the ion add up to the charge of the ion, so the sum of oxidation states in SO_4^{2-} is -2

Examples of oxidation states

-In the reaction $KBr + Cl_2 = Br_2 + 2KCl$, the oxidation state of Br is -1 in KBr and 0 in Br_2
-Since the oxidation state of Br increases from KBr to Br_2, Br is oxidised
-In the same reaction, the oxidation state of Cl is 0 in Cl_2 and -1 in KCl
-Since the oxidation state of Cl decreases from Cl_2 to KCl, Cl is reduced

Oxidising agents

-An oxidising agent causes another substance to be oxidised by losing oxygen, gaining hydrogen or gaining electrons
-It is reduced at the end of the reaction
-For example, in the reaction $Zn + CuO \rightarrow ZnO + Cu$, CuO is the oxidising agent as Zn is oxidised
-Common oxidising agents include oxygen, chlorine and potassium manganate (VII)
-In exams, always prove oxidising agents by explaining how another substance is oxidised

Reducing agents

-A reducing agent causes another substance to be reduced by gaining oxygen, losing hydrogen or losing electrons
-It is oxidised at the end of the reaction
-For example, in the reaction $Zn + CuO \rightarrow ZnO + Cu$, Zn is the reducing agent as CuO is reduced
-Common reducing agents include hydrogen, reactive metals, carbon and potassium iodide
-In exams, always prove reducing agents by explaining how another substance is reduced

Tests for reducing and oxidising agents

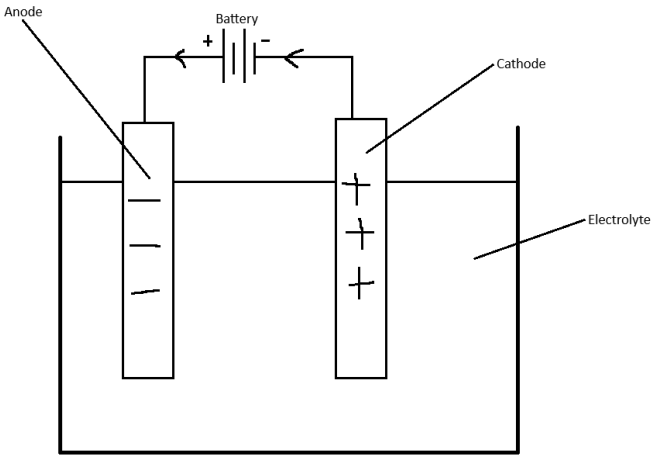
-If an oxidising agent reacts with an unknown substance, that unknown substance is a reducing agent and vice versa
-Colourless aqueous potassium iodide (KI) is added to an unknown solution
-If the solution turns from colourless to yellow-brown, oxidising agent is present as the iodide ions are oxidised and turn into brown elemental iodine
-Purple aqueous potassium manganate (VII) ($KMnO_4$) is added to an unknown substance
-If the solution turns from purple to colourless, reducing agent is present as the manganate (VII) ions are reduced into manganese (II) ions
-Both these tests can be applied to gases, as a strip of filter paper with a spot of reactant is used
-If the spot turns yellow-brown or colourless depending on the test, the gas is a oxidising or reducing agent

Chapter 13: Electrochemistry

Electrolysis

-Electrolysis is the process of using electricity to break down or decompose a compound
-The compound is usually an ionic compound in the molten or aqueous state, or an aqueous acid or alkali
-Electrolysis takes place in an electrolyic cell

Parts of an electrolytic cell

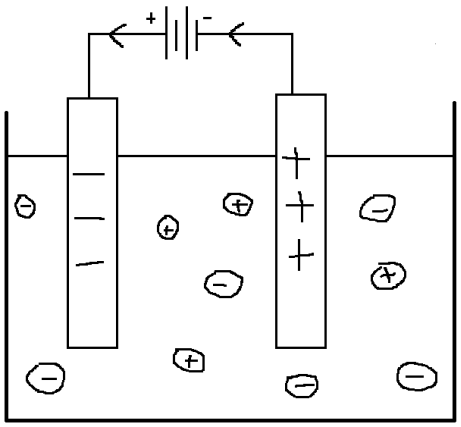


Battery
-The battery acts as an electron pump, drawing electrons away from the anode, making the anode positively charged
-The electrons then enter the positive terminal of the battery and exit at the negative terminal
-The electrons are then supplied to the cathode, making it negatively charged

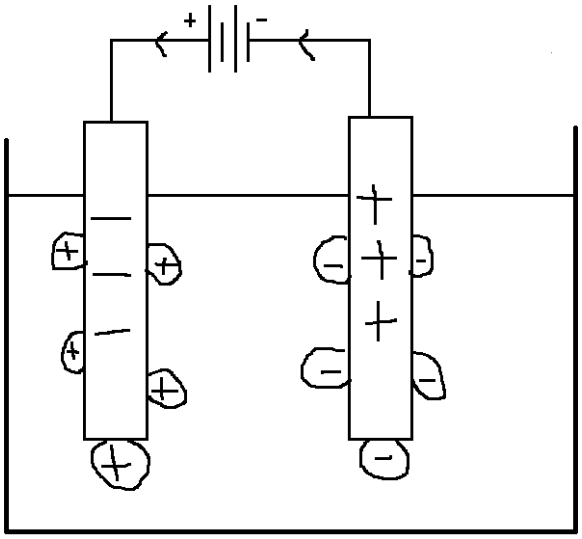
Electrodes
-There are two electrodes, the anode and the cathode
-The anode attracts anions, hence it is positively charged
-Oxidation occurs at the anode
-The cathode attracts cations, hence it is negatively charged
-Reduction occurs at the cathode
-The electrodes are usually made of carbon, graphite or metal plates, and remain chemically unchanged after a reaction

Electrolyte
-An electrolyte conducts electricity, as it contains free moving ions which allow electricity to flow through
-It is a molten ionic compound or aqueous solution
-It is decomposed to form cations and anions

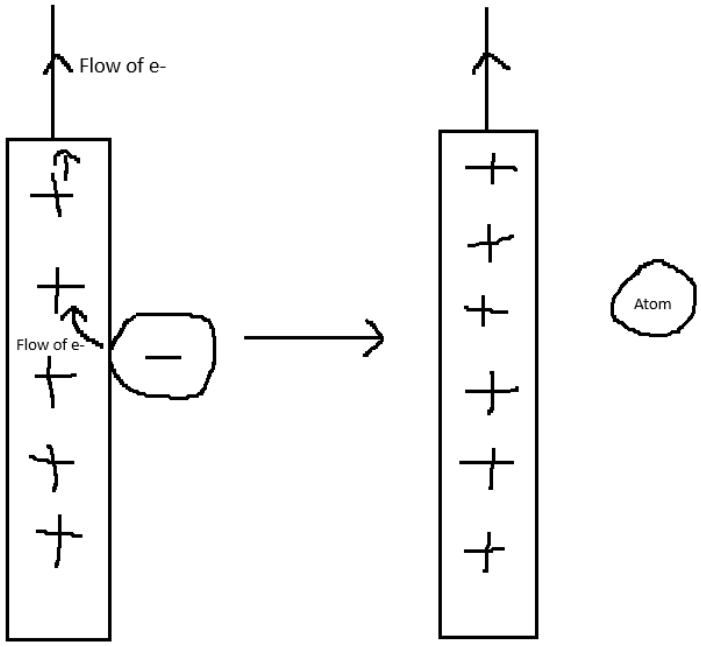
How electrolysis works



-During electrolysis, electrons flow from the positive terminal to the negative terminal of the battery
-For a split second, the ions in the electrolyte looks like the image above



-However, very quickly, the ions will be attracted to the oppositely charged cathode, and the ions will start to discharge
-When cations or anions gain or lose electrons to form atoms, we say they are discharged
-Oxidation will occur at the anode, and reduction will occur at the cathode
-You can visualise it like in the image below, where electrons are received or given up from each ion since they are ‘connected’ to the electrode



Electrolysis of molten compounds

-Molten compounds, such as salts, can have their individual elements separated through electrolysis sinc the ions are separated when molten
-For example, NaCl is a binary compound
-A binary compound is one made up of two elements, usually a metal cation and non metal anion
-Molten NaCl contains Na+ and Cl- ions
-At the anode, the negatively charged Cl- ions are attracted to the anode
-The Cl- ions then lose electrons to form Cl2 gas, and are oxidised
-The Cl- ions are said to be discharged
-At the cathode, the positively charged Na+ ions are attracted to the cathode
-The Na+ ions then gain the electrons lost from the Cl- ions to form Na metal, and are reduced
-The Na+ ions are said to be discharged
-For electrolysis of molten compounds, inert electrodes (graphite or platinum) are used
-This is because the products formed can be quite reactive, and reactive electrodes could react with them

Writing equations for molten compounds

-A half equation shows the chemical reactions happening at each electrode
-They usually have ions on the left side, and products on the right, as well as electrons (e-) in the equation as well
-State symbols are also required

-For example, for the reaction above, the half equation at the anode is $2Cl^-$ (l) → Cl_2 (g) + $2e^-$, and the half equation at the cathode is Na^+ (l) + e^- → Na (l)

-The reactants and products are liquids or gases, as the reaction is taking place at a very high temperature

-Overall equations represent the full reaction in the electrolytic cell

-To make an overall equation, you first balance the electrons on in both half equations, so for the example above, the half equation at the anode remains the same, while the half equation at the cathode becomes $2Na^+$ (l) + $2e^-$ → 2Na (l)

-Note that when balancing electrons, the ions and atoms must also be multiplied by the same number the electrons are multiplied by

-The left hand side of the half equations are then put together, and the same is done for the right hand side, giving you $2Na^+$ (l) + $2e^-$ + $2Cl^-$ (l) → 2Na (l) + Cl_2 (g) + $2e^-$

-The electrons are then cancelled out, giving you $2Na^+$ (l) + $2Cl^-$ (l) → 2Na (l) + Cl_2 (g)

-Finally, the ions are then combined, giving you 2NaCl (l) → 2Na (l) + Cl_2 (g)

Electrolysis of aqueous solutions

-The presence of water will affect the products formed at the electrodes, as water partially ionises to form H^+ and OH^-

-However at each ion, only one ion is selectively discharged, so you use the electrochemical series explained below

-For example, in dilute NaCl solution, there are Na^+, Cl^-, H^+ and OH^- ions

-At the anode, the negatively charged Cl^- and OH^- ions are attracted to the anode

-Since the solution is dilute and Cl^- is higher than OH^- in the electrochemical series, OH^- is selectively discharged over Cl^-

-The OH^- ions then lose electrons to form O_2 gas and water, and are oxidised

-The OH^- ions are said to be discharged, and Cl^- ions remain in solution

-At the cathode, the positively charged Na^+ and H^+ ions are attracted to the cathode

-Since Na^+ is higher than H^+ in the electrochemical series, H^+ is selectively discharged over Na^+

-The H^+ ions then gain the electrons lost from the OH^- ions to form H_2 gas, and are reduced

-The H^+ ions are said to be discharged, and Na^+ ions remain in solution

-The electrolyte will become more concentrated over time, as more water is lost through H_2 and O_2 gas than created from the electrolysis of OH^-

-For electrolysis of aqueous solutions, inert electrodes are also used

Electrochemical series

-For the selective discharge of ions for aqueous solutions, we look at the position of ions on the electrochemical series, as well as the concentration of electrolyte

-For cations:
K+
Na+
Ca2+
Mg2+
Zn2+
Fe2+
Pb2+
H+
Cu2+
Ag+
-Note that the electrochemical series for cations is related to the reactivity series for metals, where reactivity increases as you go up the electrochemical series
-The ions lower in position will be selectively discharged over it, so H+ will always be selectively discharged over Na+, and Cu2+ will always be selectively discharged over H+
-Do note that for transition metal cations e.g. Fe2+ or Mn2+, always use the ion which loses the least electrons, so use Fe2+ over Fe3+
-For example, in dilute CuSO4 solution, there are Cu2+, SO42-, H+ and OH- ions
-At the anode, the negatively charged SO42- and OH- ions are attracted to the anode
-Since SO42- is higher than OH- in the electrochemical series, OH- is selectively discharged over SO42-
-The OH- ions then lose electrons to form O2 gas and water, and are oxidised
-The OH- ions are said to be discharged, and SO42- ions remain in solution
-At the cathode, the positively charged Cu2+ and H+ ions are attracted to the cathode
-Since H+ is higher than Cu2+ in the electrochemical series, Cu2+ is selectively discharged over H+
-The Cu2+ ions then gain the electrons lost from the OH- ions to form Cu metal, and are reduced
-The Cu2+ ions are said to be discharged, and H+ ions remain in solution
-The electrolyte becomes more acidic as OH- ions are removed while H+ ions remain in solution
-For anions:
SO42-
NO3-
Cl-
Br-
I-
OH-
-The ions lower in position will be selectively discharged over it, so OH- will always be selectively discharged over NO3-
-However, for the halide ions (Cl-, Br-, I-), in concentrated solutions, Cl- will be in higher concentration than OH- ions, and Cl- ions are discharged in preference to OH- ions
-This is because the two anions are close together in the electrochemical series
-For example, in concentrated NaCl solution, there are Na+, Cl-, H+ and OH- ions
-At the anode, the negatively charged Cl- and OH- ions are attracted to the anode

-Since the solution is concentrated and Cl- is quite close to OH- in the electrochemical series, Cl- is selectively discharged over OH-
-The Cl- ions then lose electrons to form Cl2 gas, and are oxidised
-The Cl2- ions are said to be discharged, and OH- ions remain in solution
-At the cathode, the positively charged Na+ and H+ ions are attracted to the cathode
-Since Na+ is higher than H+ in the electrochemical series, H+ is selectively discharged over Na+
-The H+ ions then gain the electrons lost from the OH- ions to form H2 gas, and are reduced
-The H+ ions are said to be discharged, and Na+ ions remain in solution
-The electrolyte becomes more alkaline as H+ ions are removed while OH- ions remain in solution
-In exams, when they just say aqueous solution, you can assume it is dilute

Writing equations for aqueous solutions and final electrolyte

-For the electrolysis of H+ and OH- ions, the half equations are $2H^+ (aq) + 2e^- \rightarrow H_2 (g)$ and $4OH^- (aq) \rightarrow 2H_2O (l) + O_2 (g) + 4e^-$ respectively

-The method of making overall equations is still the same, so for dilute NaCl, the overall equation is $2H_2O (l) \rightarrow 2H_2 (g) + O_2 (g)$ (you cancel out the extra 2 H_2O on both sides)

-However, for reactions such as concentrated NaCl and dilute $CuSO_4$, there is a bit of difference
-If the reactant ions cannot form the compound originally used in the reaction i.e. NaCl, $CuSO_4$ or H_2O etc, you just write them as ions

-So the overall equations for concentrated NaCl and dilute $CuSO_4$ are $2H^+ (aq) + 2Cl^- (aq) \rightarrow H_2 (g) + Cl_2 (g)$ and $2Cu^{2+} (aq) + 4OH^- (aq) \rightarrow 2Cu (s) + 2H_2O (l) + O_2 (g)$ respectively

-Final electrolyte is found by combining the leftover ions after electrolysis
-For dilute NaCl, since H+ and OH- are discharged, Na+ and Cl- remain, so the final electrolyte is NaCl (aq)
-For concentrated NaCl, since H+ and Cl- are discharged, Na+ and OH- remain, so the final electrolyte is NaOH (aq)
-For dilute $CuSO_4$, since Cu2+ and OH- are discharged, H+ and SO42- remain, so the final electrolyte is H_2SO_4 (aq)
-Do note that since electrolysis for aqueous solutions is done at r.t.p, the state symbols of products follow that, and the reactants are always (aq)

Electrolysis with reactive anode

-Reactive anodes are anodes made of metals other than Pt or graphite, such as copper and silver
-These anodes will ionise to form ions, and anions such as OH- will not discharge and instead remain in the electrolyte
-For example, in dilute $CuSO_4$ solution, there are Cu2+, SO42-, H+ and OH- ions

-However, if Cu electrodes are used, OH^- and SO_4^{2-} will not discharge at the anode

-Instead, the anode will ionise or oxidise in the half equation $Cu\ (s) \rightarrow Cu^{2+}\ (aq) + 2e^-$

-The anode will also decrease in mass and size

-At the cathode, the nothing changes, so the half equation is still $Cu^{2+}\ (aq) + 2e^- \rightarrow Cu\ (s)$

-Do note that if metals such above H^+ such as Zn are used as both anodes and cathodes, the half equation at the anode will be $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-$, but the half equation at the cathode will be $2H^+\ (aq) + 2e^- \rightarrow H_2\ (g)$

Electrolytic purification

-Impure metals such as copper can be purified using electrolysis

-For example, to purify Cu, an aqueous solution of a Cu salt is used as an electrolyte, so $CuSO_4$ can be used, and pure Cu as the cathode and impure Cu as the anode

-At the anode, the Cu will ionise to form Cu^{2+} and $2e^-$, while the impurities are left behind, and usually settle to the bottom of the electrolytic cell

-At the cathode, pure Cu is then deposited onto the cathode, hence purifying the Cu

Electroplating

-Electroplating is the process of depositing a layer of metal onto another substance using electrolysis

-Electrolysis forms a thin layer of metal on the object, which can be used for decorative or protective purposes

-To electroplate a metal onto another object, an aqueous solution of the metal's salt is used as the electrolyte, the anode is the metal that will be plated onto the object, and the cathode is the object

-The cathode has to be conductive to allow electrolysis to take place

-For example, to electroplate a Fe mug with Ag, $AgNO_3$ is used as the electrolyte, Ag is the anode, and the Fe mug is the cathode

-At the anode, Ag will ionise in the half equation $Ag\ (s) \rightarrow Ag^+\ (aq) + e^-$, and the anode will decrease in mass and size

-At the cathode, Ag^+ ions from the electrolyte and anode will discharge as Ag metal onto the Fe mug, in the half equation $Ag^+\ (aq) + e^- \rightarrow Ag\ (s)$

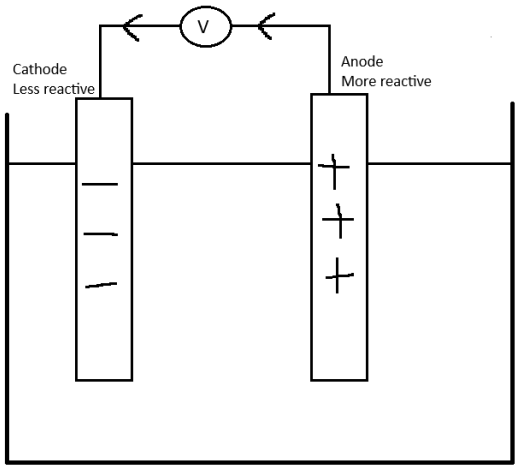
Simple cell

-A simple cell is a device that converts chemical energy into electrical energy, also known as an electric cell

-A simple cell is made by placing two different metals in contact with an electrolyte, and the two metals act as electrodes

-The electrons always flow from the more reactive metal to the less reactive metal

-Hence, the more reactive metal becomes positively charged and is the anode or negative electrode, and vice versa



-Note that there is no bettery and instead there is a voltmeter, as electricity is generated by the difference
in reactivity of the metals
For example, if Zn and Cu electrodes are used and dilute H_2SO_4 is used as an electrolyte, since Zn is more reactive, it becomes the anode and Cu becomes the cathode
-Zn then loses electrons and discharges to form Zn^{2+} ions, hence decreasing in mass and size
-The electrons then flow along the wire into the Cu electrode, making it negatively charged
-This flowing of electrons is what produces electricity
-The larger the difference in reactivity of the two metals used based on the reactivity series, the greater the greater the voltage produced
-For example, Mg and Cu electrodes will produce 2.7 V of electricity, while Fe and Cu electrodes only produce 0.8 V of electricity
-If the electrodes are the same metal, no current will flow and no electricity is produced

Writing equations for simple cell

-For simple cell, the half equation of the anode follows the same rule as when a reactive anode is used, so the metal at the anode will always ionise
-For the cathode, you follow the electrochemical series with reference to the electrolyte to determine the half equation, so the cathode metal does not matter when writing the half equation

-For the example above, the anode half equation will be $Zn\ (s) \rightarrow Zn^{2+}\ (aq) + 2e^-$

-For the cathode, since the electrolyte is H_2SO_4, the only cations are H^+

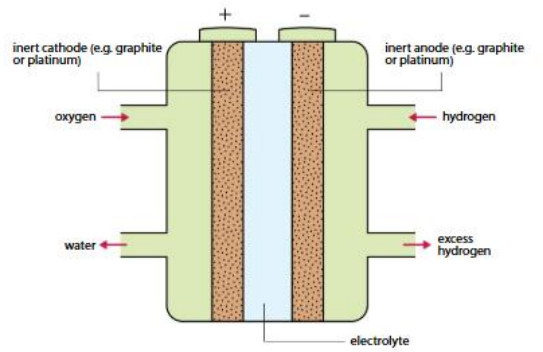
-Hence the reaction at the cathode is $2H^+\ (aq) + 2e^- \rightarrow H_2\ (g)$

-For overall equation, you just balance the electrons, add up products and reactants, and then remove the electrons

-So the overall equation for the example will be $Zn\ (s) + 2H^+\ (aq) \rightarrow Zn^{2+}\ (aq) + H_2\ (g)$

Hydrogen fuel cell

-A fuel cell is a chemical cell in which the reactants are continuously supplied to produce electricity directly
-In our syllabus, we only learn about the hydrogen fuel cell, where H2 and O2 gas is used as fuel



-In a hydrogen fuel cell, the electrolyte is always an alkali
-Hence, OH- ions can react with the H2 gas supplied at the anode in the half equation $2H_2 (g) + 4OH^- (aq) \rightarrow 4H_2O (l) + 4e^-$, where excess H2 is removed at the anode
-The products (water and electrons) then move through the porous inert anode, through the electrolyte and to the cathode, where they then react with O2 gas supplied at the cathode in the half reaction $O_2 (g) + 2H_2O (l) + 4e^- \rightarrow 4OH^- (aq)$, where excess water is removed at the cathode
-This means that as long as O2 and H2 gas is being constantly supplied, the reaction can continuously produce electricity, as OH- ions are being produced and used up at a constant rate
-At the cathode, O2 gas is reduced to form OH- ions, and at the anode, H2 gas is oxidised to form water and electrons
-The overall reaction is $O_2 (g) + 2H_2 (g) \rightarrow 2H_2O (l)$

Advantages of fuel cell

-H2 is a renewable fuel as it can be obtained via the electrolysis of water or the cracking of hydrocarbons
-Hydrogen fuel cells produce no harmful by products, as only water is produced
-It is more efficient than fuel burning electricity as a larger percentage of chemical energy stored in fuel ends up as useful electricity in a fuel cell

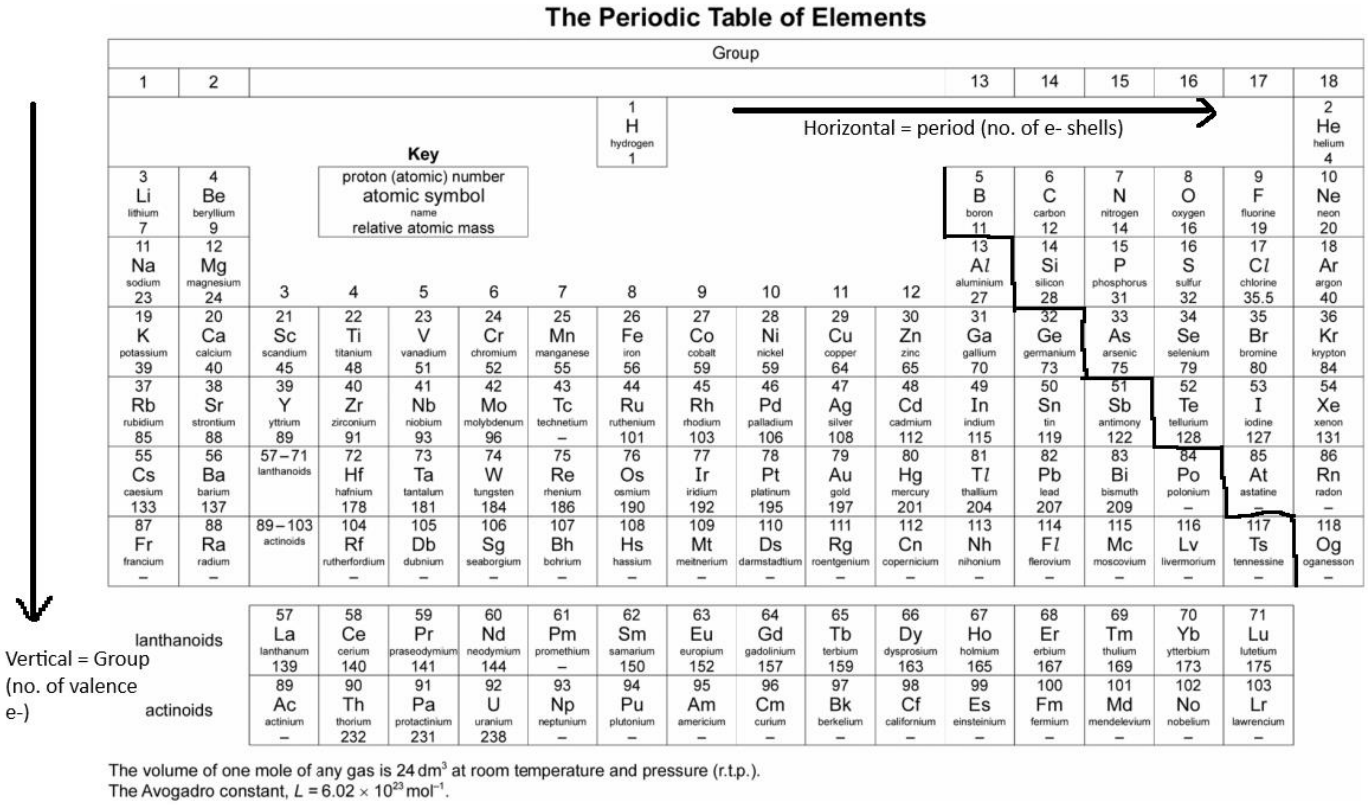
Disadvantages of fuel cell

-It is difficult to store and transport H2 safely as it is a highly flammable gas at r.t.p, so it is often transported in high pressure cylinders which can be dangerous to handle

-Extraction of hydrogen is expensive, as electrolysis of water and cracking of hydrocarbons are both expensive
-The production of hydrogen still uses electricity, which produces air pollutants such as CO, NO_2 and SO_2, which means a hydrogen fuel cell is not pollution free

Chapter 14: The Periodic Table

The periodic table



-On the periodic table, elements are arranged in order of increasing proton number
-Groups, numbered 1 to 18, are the vertical columns, and periods, number 1 to 7, are the horizontal rows
-At r.t.p, most elements are solids, but Hg and Br are liquids and H2, F2, O2, N2, Cl2 and the noble gases are gases

Periodic trends

Physical properties
-All elements in the same period have the same number of electron shells, so period 2 has 2 electron shells etc
-The number of valence electrons corresponds with the group number, so group 1 has 1 valence electron etc
-Metallic properties tend to decreases across each period while non metallic properties tend to increase, for example in period 2, the first three elements are metallic, while those from group 14 to 17 are non metallic

Chemical properties
-There is a gradual trend in the type of bonds formed by the elements in compounds from ionic to covalent as we go across the period

-They form oxides which change from strongly basic to amphoteric to strongly acidic across the period

Ionisation energy

-Ionisation energy is the amount of energy required to remove the most loosely bound electron (valence electron) from an atom
-A higher ionisation energy of an atom makes it harder to remove the valence electron from the atom, resulting in a more stable atom and a less reactive atom

Trends across a period

	Na	Mg	Al	Si	P	S	Cl
	1	2	13	14	15	16	17
nce e-	1	2	3	4	5	6	7
on	1+	2+	3+	-	3-	2-	1-
perty	Metal			Metalloid	Non-metal		
	Giant metallic structure			Giant molecular structure	Simple molecular structure		
xide	Basic		Amphoteric	Acidic			
f ompounds	Giant ionic lattice structure			Giant molecular structure	Simple molecular structure		
us / pm	190	145	118	111	98	87	79
nergy / kJmol^-1	496	738	578	787	1012	1000	1251
xidising agent	Reducing agent			→			Oxi

-Note that B, Si, Ge, As, Sb, Te, Po and Lv are considered metalloids, and have properties of both metals and non metals
-Atomic radius generally decreases as there are the same number of electron shells across the period but the number of electrons increases, hence increasing the electrostatic forces of attraction between the valence electrons and the positive nucleus, and these valence electrons are held more tightly to the nucleus, decreasing atomic radius
-Ionisation energy generally increases as atomic radius decreases, so valence electrons are closer to the positive nucleus, increasing the electrostatic forces of attraction between the

nucleus and valence electrons, so the valence electrons are held more tightly and strongly to the nucleus and are harder to remove
-Note that for these two (atomic radius and ionisation energy), the trend is general and may not be true from element to element, and we do not need to memorise exact values

Trends down a group

-The elements in a group have the same number of valence electrons, so these elements have similar chemical properties
-Since they have the same number of valence electrons, they also form ions with similar charges and hence formulas of compounds formed by these elements are similar
-There is a gradual change in physical properties as we move down the group
-Atomic radius generally increases as the number of electron shells increases down a group
-Ionisation energy generally decreases as atomic radius increases, so valence electrons are further away from the positive nucleus, decreasing the electrostatic forces of attraction between the nucleus and valence electrons, so the valence electrons are held less tightly and less strongly to the nucleus and are easier to remove
-Note that for these two (atomic radius and ionisation energy), the trend is general and may not be true from element to element, and we do not need to memorise exact values

Group 1

-The group 1 elements are also known as the alkali metals

Physical properties
-They are metals with giant metallic structures, and are good conductors of electricity
-Their ionic compounds are white solids that are soluble in water to form colourless solutions
-They are soft solids that can be cut with a knife, and have low melting and boiling points
-They have low densities, and Li, Na and K all float on water

Electronic structure
-They have 1 valence electron, and hence form ions with a 1+ charge
-They form ionic compounds with similar chemical formulas, so NaCl, KCl, LiCl etc

Chemical properties
-They are very reactive with water and form an alkali and hydrogen gas as products
-They form thermally stable compounds due to their high reactivity and hence do not decompose on heating
-They react with acids to form salts and H_2 gas as well as with steam to form metal oxides and H_2 gas
-They react with O_2 gas to form metal oxides as well as with chlorine to form metal chlorides
-They are powerful reducing agents and have low ionisation energy

Group trends
-Atomic size, density and reducing property increases down the group

-Reducing property is just how well the element acts as a reducing agent
-Reactivity increases down the group because the size of the atom increases, decreasing the electrostatic forces of attraction between the positive nucleus and valence electron, making it easier to lose the valence electron, increasing reactivity
-Melting point and boiling point decrease down the group as with more electron shells, there are weaker electrostatic forces of attraction between the nucleus and valence electrons, resulting in a weaker metallic bond which require less heat energy to break, hence a decreasing melting and boiling point

Group 17

-The group 17 elements are also known as the halogens

Physical properties
-They are non metals which form diatomic molecules with a simple molecular structure
-They have low melting points and boiling points due to weaker intermolecular forces of attraction between the molecules
-They are coloured and soluble in water and organic solvents
-At r.t.p,
-Fluorine is a yellow gas
-Chlorine is a green gas
-Bromine is a brown liquid
-Iodine is a black solid (iodine sublimates to form a purple gas, and aqueous iodine is brown)
-Astatine is a black solid
-Fluorine compounds are addded to tap water and toothpaste to prevent tooth decay
-Chlorine is added to tap water and swimming pools to kill harmful bacteria
-Iodine is used as an antiseptic for cuts and wounds

Electronic structure
-They have 7 valence electrons and hence form ions with a 1- charge

Chemical properties
-They are very reactive non metals which react to form stable compounds
-They react with alkali metals to form ionic compounds, and react with non metals to form halides
-They react with hydrogen to form halides which when dissolved in water, form acids, and are powerful oxidising agents
-A more reactive halogen can displace a less reactive halogen from its halide or salt solution
-For example, chlorine will displace iodine from sodium iodide because chlorine is more reactive than iodine in the reaction Cl_2 (g) + 2NaI (aq) → 2NaCl (aq) + I_2 (aq)
-Observation: the solution turns from colourless to brown
-Ionic equation: Cl_2 (g) + 2I^- (aq) → 2Cl^- (aq) + I_2 (aq)

Group trends
-Atomic size and density increases down the group
-Oxidising property decreases down the group
-Oxidising property is just how well the element acts as a oxidising agent
-Reactivity decreases down the group because as the size of the atom increases, it becomes more difficult for the nucleus to attract one more electron
-Melting point and boiling point increases down the group because as the size of the halogen molecules increases, the intermolecular forces of attraction between the molecules increases, and more energy is required to overcome these forces and cause boiling/melting
-Elements get darker down the group

Group 18

-The group 18 elements are also known as the noble gases

Physical properties
-They are colourless monoatomic gases with low melting and boiling points
-They are insoluble in water
-Helium is used to fill up weather balloons and airships
-Neon is used in advertising strip lights
-Argon is used to weld stainless steel and fill electric bulbs to protect filaments from oxidising in air

Electronic structure
-They have what is known as a stable octet (duplet for helium) structure, hence they do not form ions
-Every element reacts to form a stable octet structure, or a stable duplet structure for hydrogen, lithium and beryllium

Chemical properties
-Since they have a fully filled valence electron shell and hence have a stable electronic configuration, theory are very chemically unreactive and are inert
-They do not gain, lose or share electrons and do not form compounds

Group trends
-Atomic size and density increases down the group
-Melting and boiling point also increases down the group due to increases intermolecular forces of attraction between molecules and hence, more energy is required to overcome these forces and cause boiling/melting

Transition elements

-The transition elements or transition metals, are the elements from group 3 to 11, except scandium

-They are metals with strong metallic bonding and a giant metallic structure resulting in high melting and boiling points
-They have high densities
-They have variable oxidation states in their compounds, so for example manganese has an oxidation state of +2 in $MnCl_2$ and an oxidation state of +4 in MnO_2
-They form coloured compounds, so for example $Fe(OH)_2$ is green while $Fe(OH)_3$ is orange
-Transition metals and their compounds are good catalysts, for example iron in the haber process
-Tungsten is used as filaments for light bulbs
-Iron is used as a catalyst in the haber process for manufacturing ammonia
-Platinum is a catalyst when making nitric acid
-Manganese (IV) oxide is a catalyst for the decomposition of hydrogen peroxide to make oxygen
-Nickel is a catalyst to manufacture margarine from vegetable oils
-Note that the roman numeral next to the transition element in a compound denotes the oxidation state, so +3 in iron (III) chloride and +2 copper (II) sulfate

Chapter 15: The Reactivity Series

Reactivity series

-The reactivity series of metals allows for the prediction of the reactivity of different metals by their reactions with water, steam and HCl, as well as their tendency to form its positive ion
-The reactivity series is as follows:
Potassium
Sodium
Calcium
Magnesium
Carbon
Zinc
Iron
Lead
Hydrogen
Copper
Silver
-Reactivity decreases down the reactivity series, so potassium is the most reactive, and silver is the least reactive

	K	Na	Ca	Mg	Zn	Fe	Pb	Cu
to	Decreasing reactivity →							
	Violent		Readily	Slowly	Does not react			
	Explosive			Violent	Readily	Slowly	Does not r	
	Explosive		Violent	Fast	Moderately Fast	Slowly	Does not r	
ive f H2	Does not react				Reduced with increasing ease →			
tive	Does not react	Decompose to respective oxide + CO_2 gas with increasing ease →						

Reaction of metals with water, steam and acid

-In general, when reacting with these reagents, the metal will displace hydrogen from the water, steam or acid

-Metal + water → metal hydroxide + hydrogen gas

-Metal + steam → metal oxide + hydrogen gas

-Metal + acid → salt + hydrogen gas

-So for a reaction with Na, the equations will be $2Na (s) + 2H_2O (l) \rightarrow 2NaOH (aq) + H_2 (g)$, $2Na (s) + H_2O (g) \rightarrow Na_2O (s) + H_2 (g)$ and $2Na (s) + 2HCl (aq) \rightarrow 2NaCl (aq) + H_2$

Displacement of metals

-A more reactive metal can displace a less reactive metal from its solution

-For example, in the reaction $Zn (s) + CuSO_4 (aq) \rightarrow ZnSO_4 (aq) + Cu (s)$, Zn is able to displace the Cu in $CuSO_4$ to form $ZnSO_4$ as Zn is more reactive than Cu

Displacement of metals from metal oxides

-A more reactive metal can displace a less reactive metal from its metal oxide

-For example, in the reaction $2Al (s) + Fe_2O_3 (s) \rightarrow Al_2O_3 (s) + 2Fe (l)$, Al is able to displace the Fe in Fe_2O_3 to form Al_2O_3 as Al is more reactive than Fe

-This is also known as the thermite reaction, which is a very exothermic reaction, hence why the Fe produced is molten

-Strong heating is also required for the reaction to occur

Reduction of metal oxides with carbon or hydrogen

-Carbon and hydrogen can both act as reducing agents to extract metals

-For example, in the reactions $2ZnO (s) + C (s) \rightarrow 2Zn (s) + CO_2 (g)$ and $ZnO (s) + H_2 (g) \rightarrow Zn (s) + H_2O (g)$

-These reactions both yield a pure metal, and require strong heating

-However, carbon is a more powerful reducing agent than hydrogen as a lower temperature is needed to reduce ZnO to Zn with carbon than with hydrogen

-Carbon is also cheaper than hydrogen

Heating of metal carbonates

-The more reactive a metal, the more thermally stable its compound, and the opposite is true for less reactive metals

-When a metal carbonate is heated (given that it reacts), it will form a metal oxide and CO_2 gas, for example in the reaction $CuCO_3 (s) \rightarrow CuO (s) + CO_2 (g)$

-Hence, thermal stability of the carbonate is greatest for very reactive metals like group 1 metals such as K and Na, hence K_2CO_3 and Na_2CO_3 will not decompose, and no CO_2 gas will be produced since they are very stable to heat
-For less reactive metals, the less reactive the metal, the more CO_2 gas is collected in the same amount of time
-For Ag, it's carbonate is so unstable because of Ag being so unreactive that when heated, it will decompose entirely to form Ag metal, O_2 gas and CO_2 gas in the reaction $2Ag_2CO_3\ (s) \rightarrow 4Ag\ (s) + O_2\ (g) + 2CO_2\ (g)$

Valence electrons

-The atomic structure of an element can affect its reactivity
-If there are more valence electron shells, the element will be more reactive and react more vigorously, as the valence electrons are further away from the atomic nucleus, hence there are weaker electrostatic forces of attraction between the valence electrons and the atomic nucleus and the element can lose the electron more easily
-For example, K is more reactive than Na and reacts more vigorously with water as Na has only 3 valance shells, while K has 4 valence shells

Extraction of metals

-In the earth's crust, metals are not found as pure elements, but rather mostly as compounds (mainly oxides) which are called ores
-We can extract these metals through chemical extraction methods, mainly electrolysis of molten oxides and carbon reduction
-Electrolysis of molten oxides is the process of using electricity to break down or decompose a compound, and is used to extract ores of reactive metals (K, Na, Ca, Mg)
-This is because these ores are very stable to heat
-Carbon reduction is the process of heating the compound strongly with carbon or carbon monoxide, where carbon or carbon monoxide acts as a reducing agent, and is used for moderately reactive metals (Zn, Fe, Pb, Cu, Ag)
-For unreactive metals (Au), they do not require extraction using chemical methods as they appear as pure elements due top their unreactivity

Rusting

-Rusting is a reaction where iron reacts with water and oxygen to form rust
-Rusting is also known as the oxidation or corrosion of iron
-Rust is a brown solid, with the chemical name of hydrated iron (III) oxide and formula $Fe_2O_3 . xH_2O$
-Rusting is a term for iron, and only iron
-The general equation for rusting is $4Fe\ (s) + 3O_2\ (g) + 2xH_2O\ (l) \rightarrow 2Fe_2O_3 . xH_2O\ (s)$
-Corrosion is the reaction between any metal with oxygen and water

Rust prevention

Protective layer

-A coating, such as paint, grease or zinc, forms a barrier protection or protective layer around the iron and prevents the exposure of it to oxygen and water, preventing rusting

-Common protective layers include paint, grease, plastic, chrome or tin plating, or zinc through galvanising

-The protective layer of paint or plastic may be able to be scratched off easily, so it is not as effective as other methods

-Oil or grease has to be renewed as well

-The metal used to plate the iron should be preferably more reactive to it, so that even if the coating scratches, the coating will still corrode in preference to the iron

Sacrificial metal

-A metal which is more reactive than iron, such as Mg or Zn, can cover the surface of the iron, be placed near the iron (connected to the iron via insulated cable) or be placed in contact with the iron

-Since the Mg or Zn is more reactive than iron, they lose electrons more readily, hence the Mg or Zn will corrode in preference to iron, preventing the iron from rusting

-Note that there does not have to be a layer of metal, because as long as the metal is in contact with the iron, the iron will not rust

-For example, ships often have zinc attached to steel hulls to prevent the iron in the steel from rusting

Alloys

-Stainless steel is the best known rust resistant alloy of iron, made up of iron, nickel and chromium

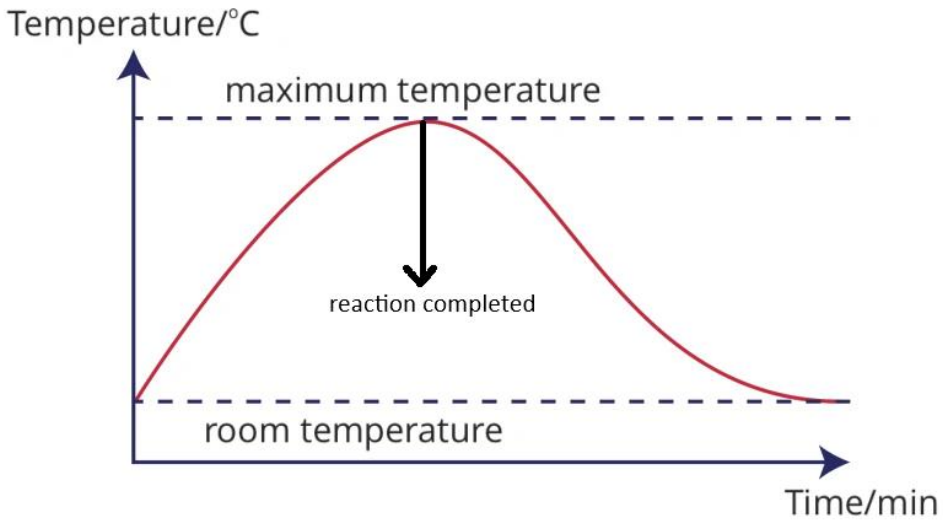
-On exposure to air and water, a very hard coating of chromium (III) oxide, or Cr_2O_3, is formed on the surface of the steel, protecting it from rusting

-However, using stainless steel is the most expensive form of rust prevention

Chapter 16: Chemical Energetics

Exothermic reactions

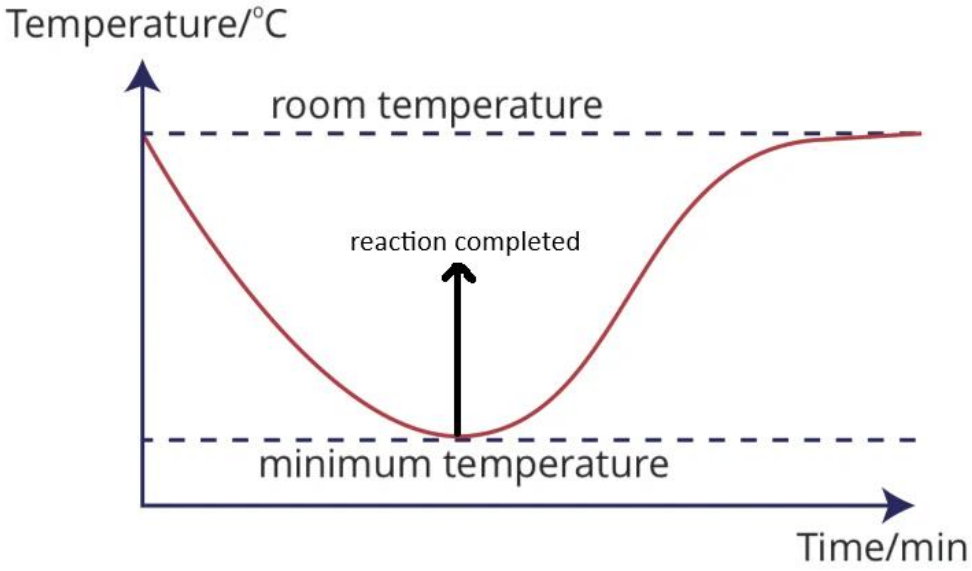
-Exothermic reactions are defined as reactions which give out heat to its surroundings
-This results in the increase in temperature of the surroundings



-Examples of exothermic reactions include combustion, respiration, neutralisation, oxidation and the addition of sodium to water
-Examples of exothermic processes include condensation and freezing

Endothermic reactions

-Endothermic reactions are defined as reactions which absorb heat from its surroundings
-This results in a decrease in temperature of the surroundings



-Examples of endothermic reactions include thermal decomposition of metal carbonates, photosynthesis and electrolysis
-Examples of endothermic processes include boiling, melting, sublimation and dissolving some ionic compounds in water such as NH4Cl

Reaction phases

-In a chemical reaction, there are usually two reaction phases, bond breaking and bond forming
-This is because in order to convert reactants into products, bonds between the reactants must first be broken, before bonds in the products are formed
-Bond breaking requires energy, hence it absorbs energy from the surroundings and is an endothermic process
-Bond forming releases energy to the surroundings, hence it is an exothermic process
-The energy released or absorbed is in the form of thermal energy or heat

Enthalpy change

-We call the amount of net energy given out or absorbed in a reaction the enthalpy change of a reaction
-Symbol for enthalpy change: ΔH
-Formula for enthalpy change: ΔH = total energy absorbed for bond breaking in reactants - total energy given out for bond making in products
-Unit for enthalpy change: kJ (kilojoules)
-ΔH is a negative value for exothermic reactions, as the energy released during bond formation exceeds the energy absorbed during bonds breaking
-ΔH is a positive value for endothermic reactions, as more energy is absorbed to break bonds than the energy released when bonds are made
-To explain why a reaction is exo/endothermic using ideas of bond breaking and forming, you say that the total energy absorbed to break the bonds in (reactants) is less/more than the energy given out to form the bonds in (products)
-Note that it’s less for exothermic reactions, and more for endothermic reactions

Bond energy

-Bond energy is the amount of energy absorbed to break one mole of a chemical bond, and is also the amount of energy released when one mole of that bond is formed
-Unit for bond energy: kJ/mol (kilojoules per mole)
-Some common bond energies are listed below, though they do not need to be memorised

t bond	Bond energy / kJ mol^-1	Covalent bond	Bond energy / kJ mol^-1
	436	H-Br	366
	158	H-N	388
	244	O=O	496

	224	C=C	612
	215	C=O	743
	568	N≡N	945
	432	C≡C	838

-Note that '-' means a single bond, '=' means a double bond and '≡' means a triple bond

-This information can be used to determine the nature of a reaction by finding the enthalpy change of the reaction

-For example, when hydrogen gas and fluorine gas react, they form hydrogen fluoride in the reaction H2 (g) + F2 (g) → 2HF (g)

-1 mole of H2 has 1 H-H bond,1 mole of F2 has 1 F-F bond, and 2 moles of HF has 2 H-F bonds
-Hence, energy released to form bonds in 2 moles of HF = 2 mol x 568 kJ/mol = 1136 kJ
-Energy absorbed to break bonds in 1 mole of H2 = 1 mol x 436 kJ/mol = 436 kJ
-Energy absorbed to break bonds in 1 mole of F2 = 1 mol x 158 kJ/mol = 158 kJ
-Thus, ΔH = (436 + 158) - (1136) = -542 kJ
-This can be simplified to ΔH = (436 + 158) - (568 x 2) = -542 kJ
-Since ΔH is negative, the reaction is exothermic

Activation energy

-The activation energy, Ea, is the minimum amount of energy that colliding reactant particles must posses to react with each other
-When reactant particles with energy greater than or equal to the activation energy collide, they are able to break their bonds and form new ones

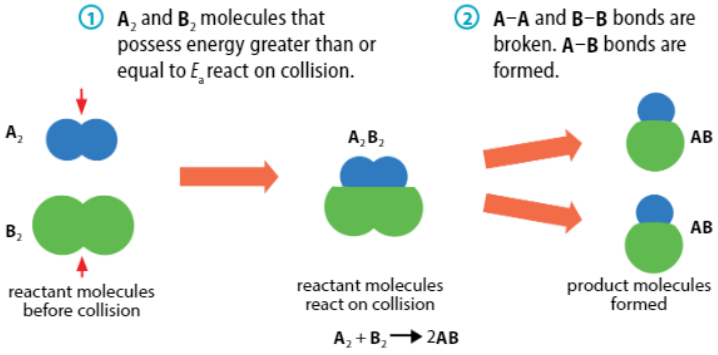


Figure 16.6 Reactant particles with sufficient energy will react upon collision.

-However if they posses energy less than the activation energy, they are unable to break their bonds and no reaction occurs

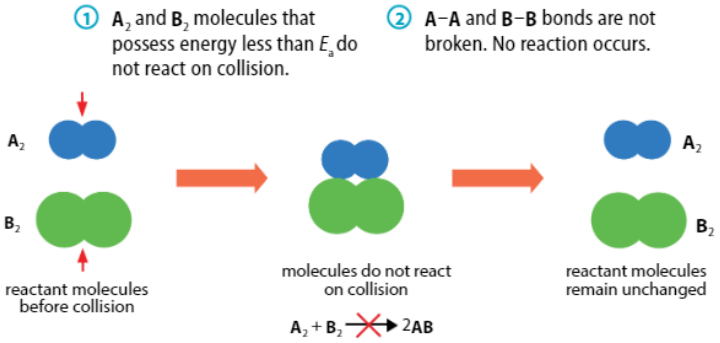
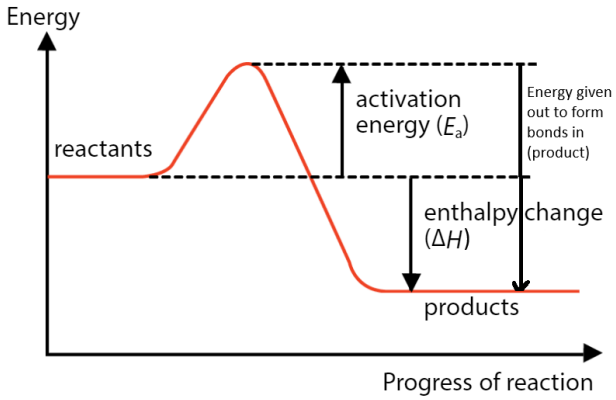


Figure 16.7 Reactant particles with insufficient energy will not react upon collision.

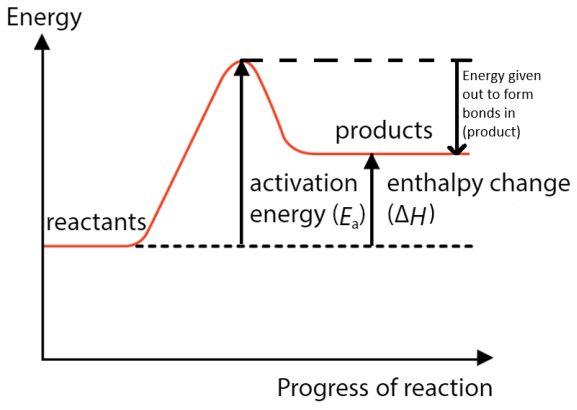
-The lower the activation energy, the faster the rate of reaction
-The activation energy can be lowered with a catalyst, which provides an alternative route with a lower activation energy for a reaction to occur

Energy profile diagrams

-Energy profile diagrams are used to show the enthalpy change, activation energy and total energy change of a reaction
-The energy profile diagram of an exothermic and endothermic reaction is shown below



Exothermic Reaction

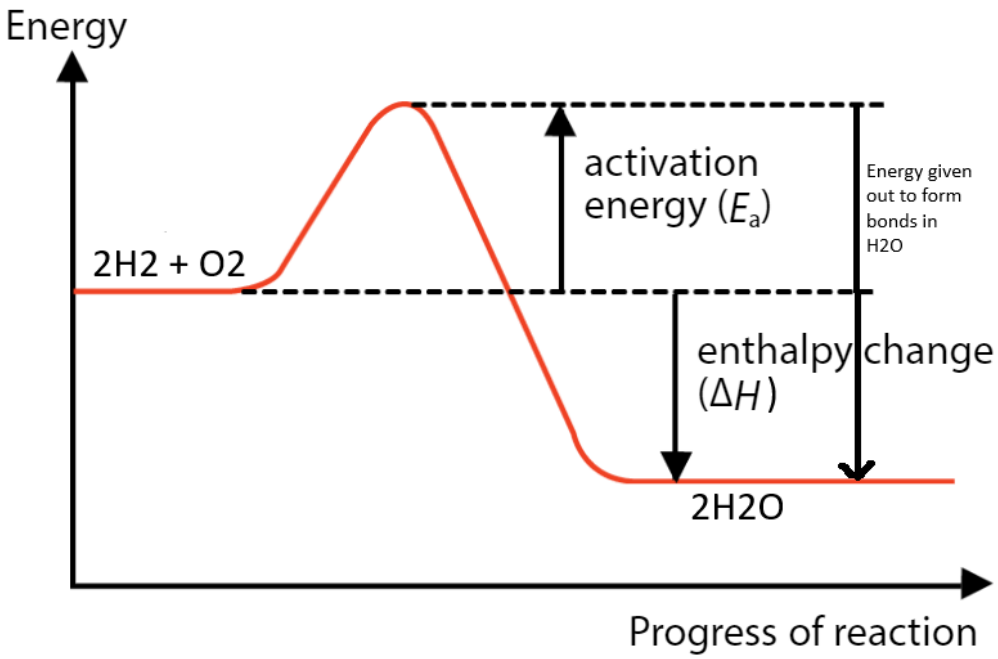


Endothermic Reaction

-For example, the reaction 2H2 (g) + O2 (g) → 2H2O (g) is exothermic

-The reaction is exothermic as the total energy content of the reactants is more than the total energy content of the products

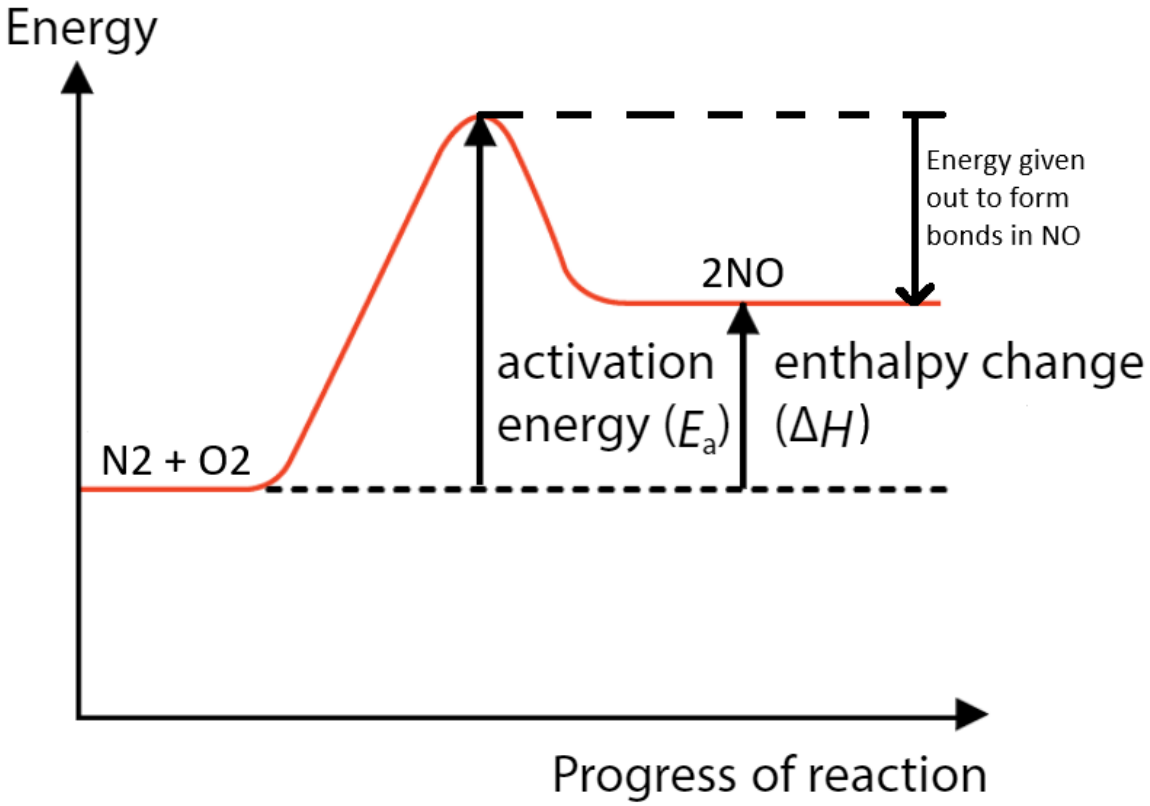
-The energy profile diagram is shown below



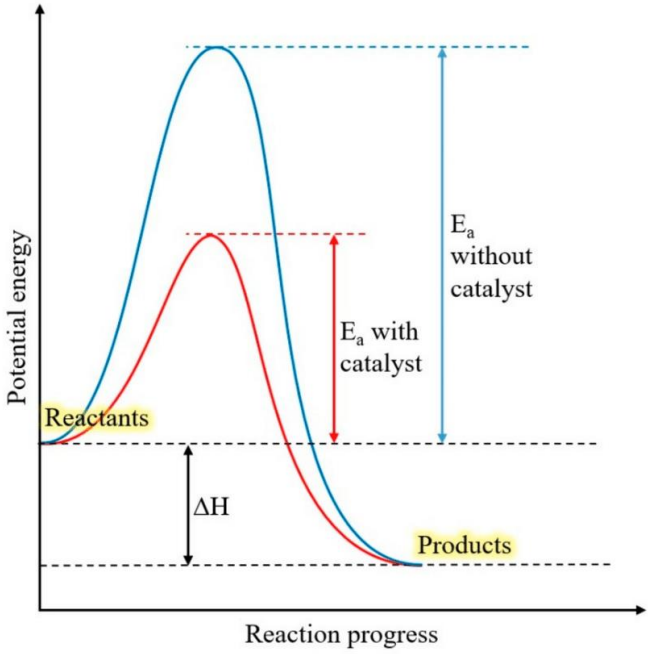
-For example, the reaction N2 (g) + O2 (g) → 2NO (g) is endothermic

-The reaction is endothermic as the total energy content of the reactants is less than the total energy content of the products

-The energy profile diagram is shown below



-If a catalyst is used, the new energy profile diagram would have a lower peak, corresponding to the lower activation energy

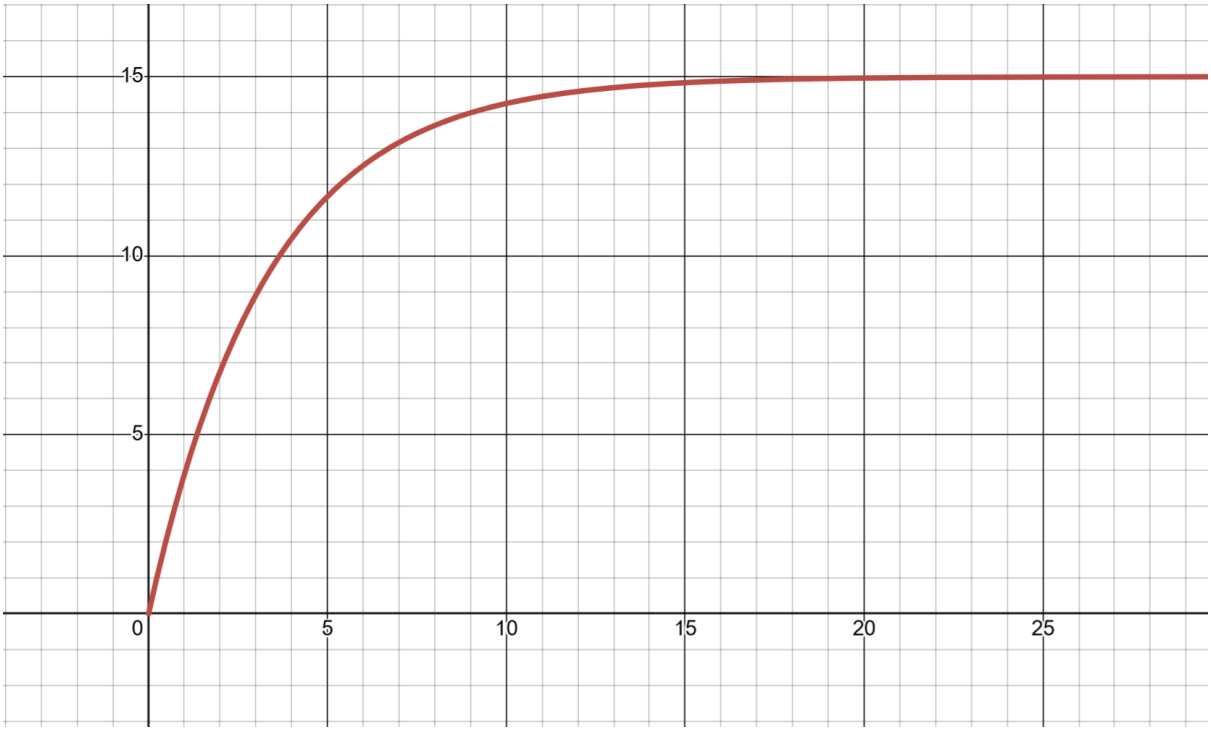


Chapter 17: Rate of Reactions

What is the speed of a reaction?

-The rate of reaction is how fast a reaction take place as the reaction proceeds with time
-Formula for rate of reaction: amount of reactant used / time taken or amount of product formed / time taken

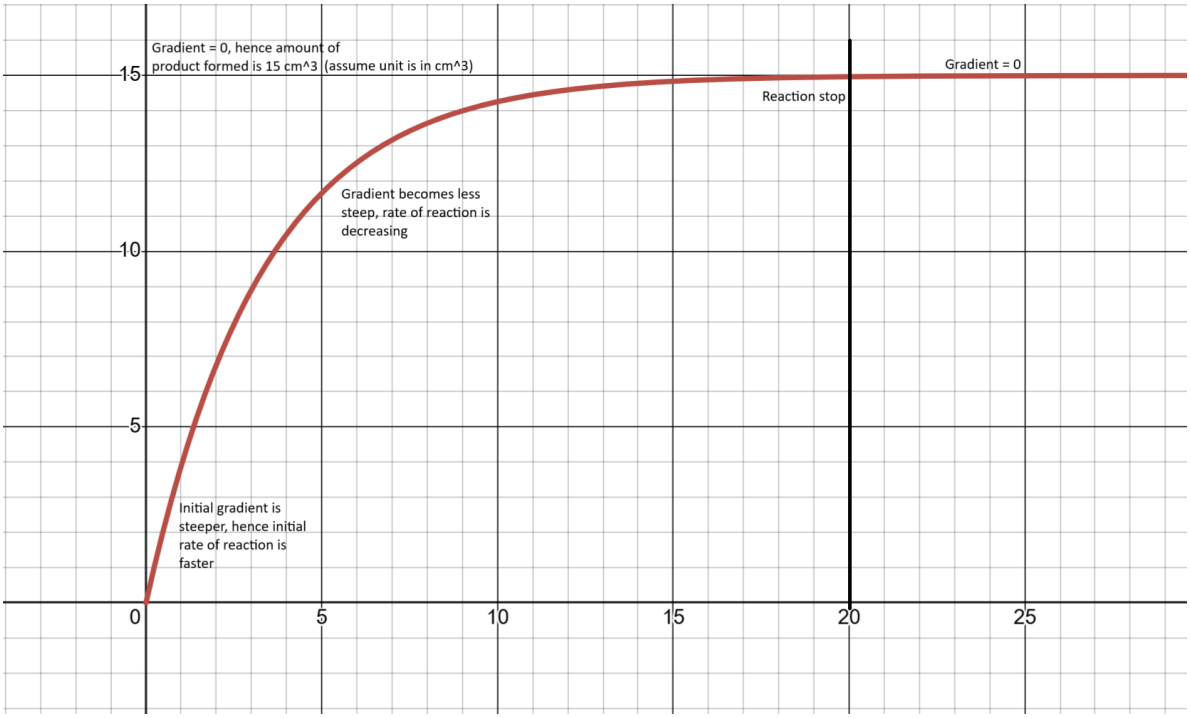
-In a reaction where a gas is produced, such as in Mg + HCl → MgCl2 + H2, the rate of reaction can be shown in a graph, where the y axis is the amount of gas produced and the x axis is the time



-As seen in the graph, the rate of reaction slows down over time
-This is because the concentration of the limiting reactant decreases over time as more product is formed, reducing the rate of reaction

Reading a reaction graph

-Three things can be obtained from a reaction graph, the rate of reaction, reaction stop and product formed
-The rate of reaction is the initial gradient of the graph before the reaction stops, where the steeper the gradient, the higher the rate of reaction
-The reaction stop is the point where the reaction stops, and can be determined by the point where the gradient reaches 0
-The amount of product formed is read off the graph when the gradient is 0



Energy in reactions

-For a chemical reaction to take place, a minimum amount of energy must be possessed by the molecules
-This is called the activation energy, and it is also the total energy absorbed to break the bonds in the reactants
-Energy is also released or absorbed in chemical reactions
-When energy is released, the reaction is known as an exothermic reaction
-When energy is absorbed, the reaction is known as an endothermic reaction

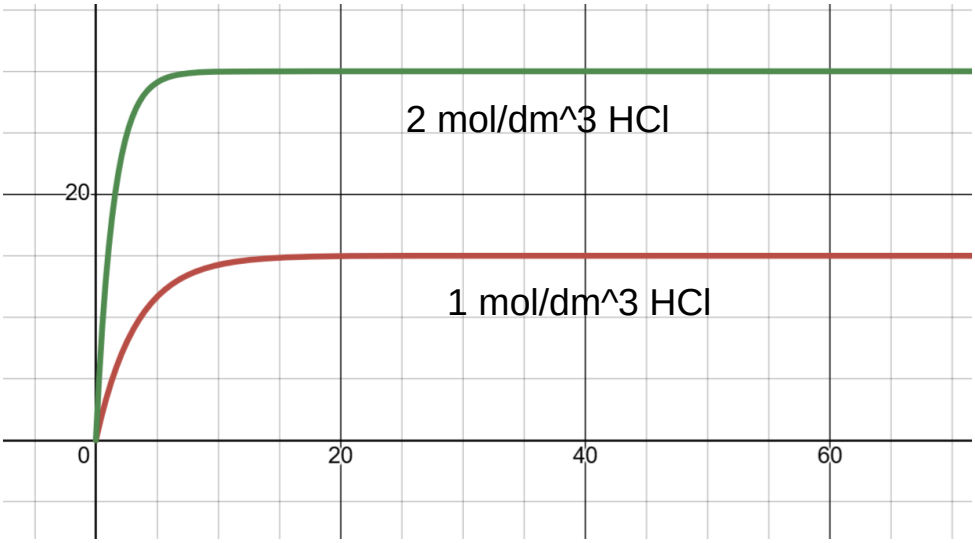
Collision theory of particles

-Particles do not always react when they collide, as sometimes they just bounce apart
-In order to react, the collisions of the particles must have energy equal to or greater than the activation energy for the reaction for the collision to be effective
-So, in order to increase the rate of reactions, we have to either increase the frequency at which particles collide, or increase the energy with which they collide

Concentration

-The concentration of the limiting reactant can affect the rate of reaction
-As the concentration increases, the number of particles per unit volume increases, which increases the frequency of effective collisions between particles
-Hence, the higher the concentration, the higher the rate of reaction
-Furthermore, the amount of product formed is also proportional to the concentration, hence if the concentration is doubled, the amount of product formed is also doubled

-For example, in the reaction of Mg + HCl → MgCl2 + H2, the rate of reaction will be faster if 2 mol/dm^3 HCl is used instead of 1 mol/dm^3 HCl, and the volume of H2 produced will be doubled

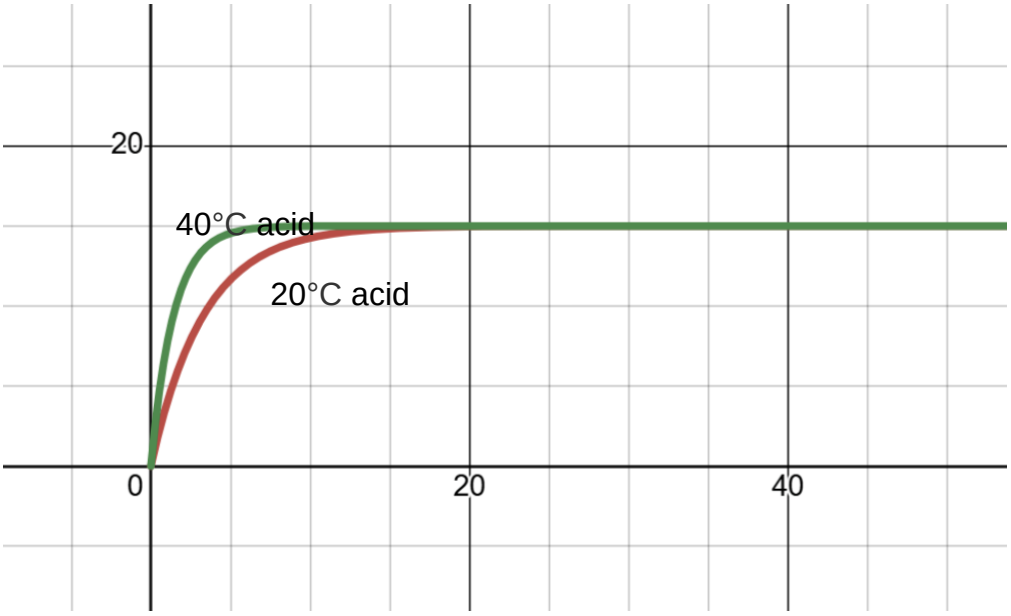


-The y axis is the volume of H2 produced and the x axis is the time

Temperature

-Temperature can affect the rate of reaction
-As the temperature increases, the kinetic energy of the particles increases and the particles move faster
-More particles also have energy higher or equal to the activation energy
-This increases the frequency of effective collisions between particles
-Hence, the higher the temperature, the higher the rate of reaction

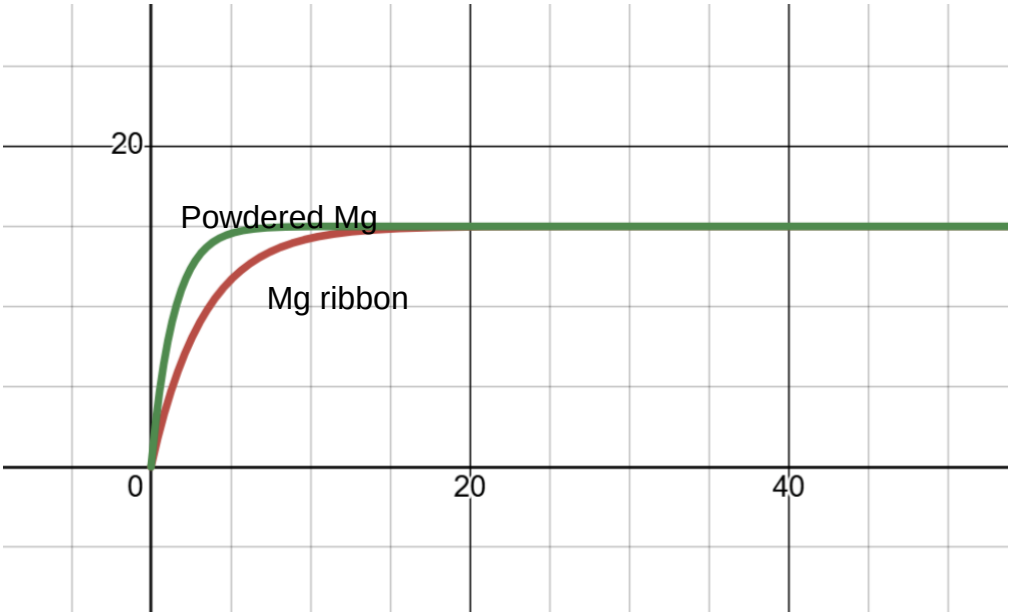
-For example, in the reaction of Mg + HCl → MgCl2 + H2, the rate of reaction will be faster if hot acid is used instead of cold acid



-The y axis is the volume of H2 produced and the x axis is the time

Particle size

-The size of the particles used can affect the rate of reaction
-The smaller the size of particles of a solid, the greater the surface area per unit mass of the solid
-This increases the frequency of effective collisions between particles
-Hence, the smaller the particle size, the higher the rate of reaction
-For example, in the reaction of Mg + HCl → MgCl2 + H2, the rate of reaction will be faster if magnesium powder is used instead of a magnesium ribbon



-The y axis is the volume of H2 produced and the x axis is the time

Catalyst

-The presence of a catalyst can affect the rate of reaction
-A catalyst is a transition metal or compound of a transition metal, and it provides an alternative route with a lower activation energy for a reaction to occur
-They are unchanged after the reaction and hence can be reused after the reaction takes place
-More reactant particles will have energy equal to or greater than the activation energy, and more reactant particles are able to collide with enough energy for the reaction to occur
-This increases the frequency of effective collisions between particles
-Hence, a catalyst increases the rate of reaction
-For example, in the reaction of N2 + H2 → NH3, the presence of a suitable catalyst such as iron metal can speed up the reaction
-Enzymes are biological catalysts

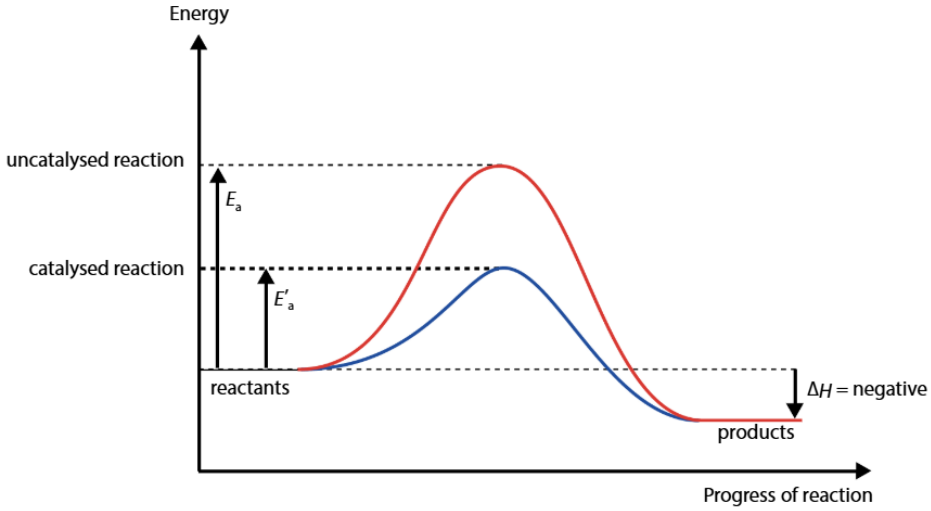
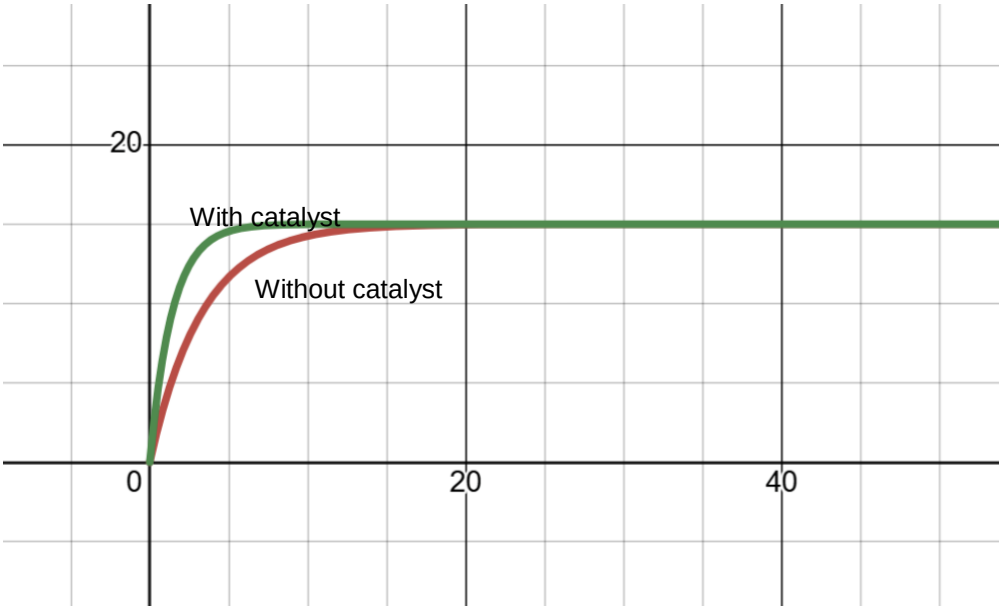


Figure 17.24 A catalyst lowers the activation energy of a reaction.

-ΔH is the change in enthalpy, E_a is the activation energy without a catalyst and E'_a is the activation energy with a catalyst



-The y axis is the volume of NH3 produced and the x axis is the time

Enzymes

-Enzymes are biological catalysts which catalyse biological reactions, such as the digestion of food
-They are made of proteins
-They are temperature sensitive and have a optimum temperature, at which catalytic activity is the highest
-When below the optimum temperature, the enzymes are dormant, and when above the optimum temperature, they are dentured and cannot catalyse reactions again
-They are also pH sensitive, and hav a optimum pH where catalytic activity is the highest
-When placed in a pH varying from their optimum pH, the enzymes will start to denature
-They have a specific catalytic action, and a specific enzyme will only catalyse a specific reaction
-This is because the active site of the enzyme can only fit certain reactant molecules
-Like catalysts, they can catalyse multiple reactions one after another as they are unchanged after the reaction

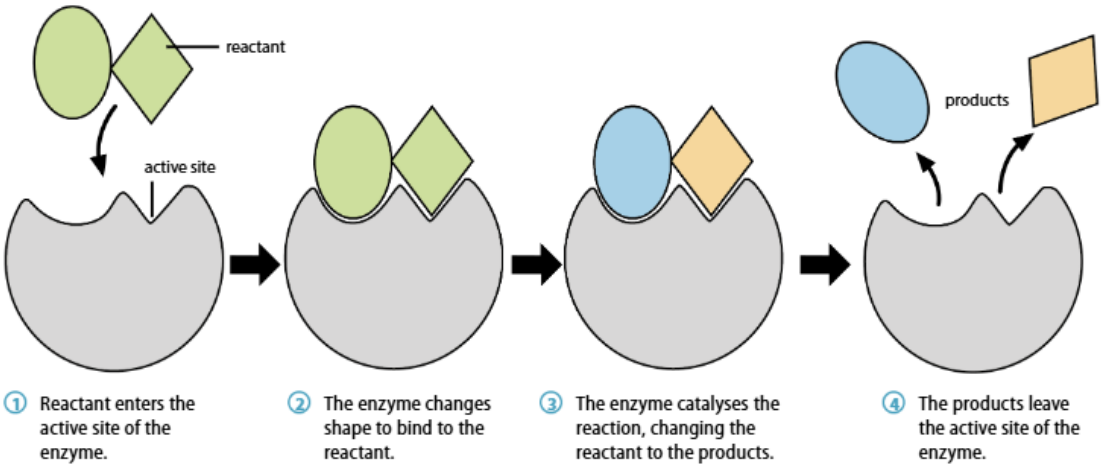
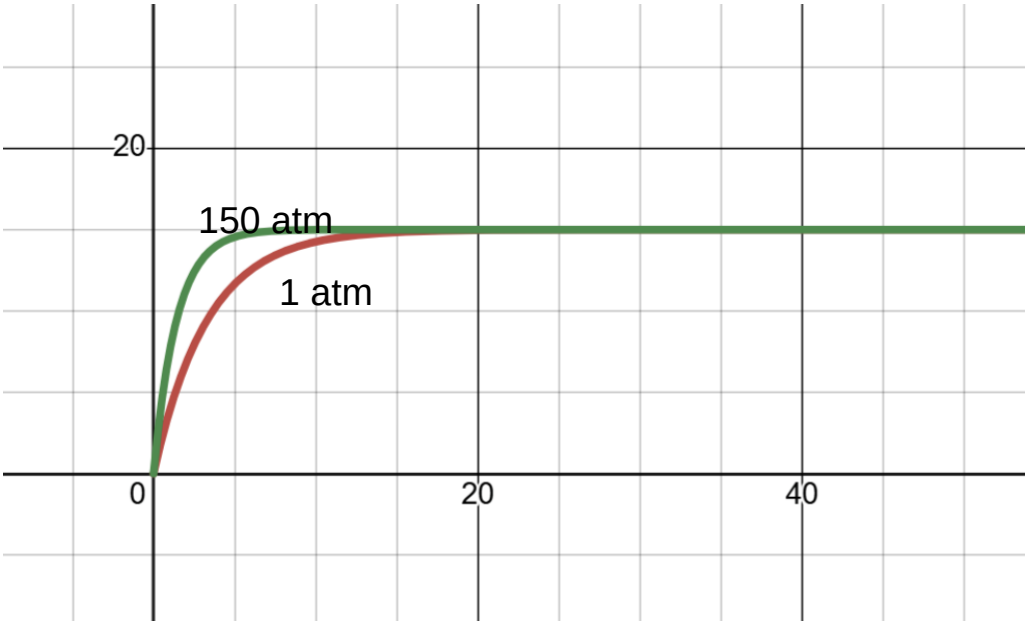


Figure 17.26 The catalytic activity of enzymes

Pressure

-The pressure of a gas can affect the rate of reaction
-The higher the pressure, the greater the number of reactant molecules per unit volume
-This increases the frequency of effective collisions between particles
-Hence, the higher the pressure, the higher the rate of reaction

-For example, in the reaction of N2 + H2 → NH3, if the reaction takes place at 150 atm instead of 1 atm, the rate of reaction will be faster (1 atm is the pressure experienced at sea level)

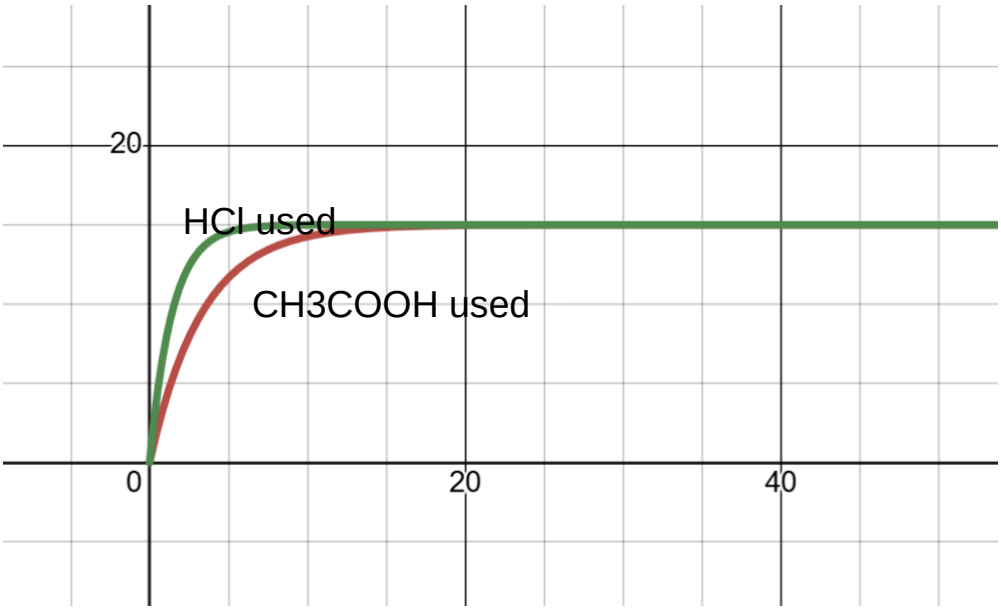


-The y axis is the volume of NH3 produced and the x axis is the time

Strength of acid

-The strength of the acid used can affect the rate of reaction
-A strong acid will ionise completely in water to yield a high concentration of H+ ions, but a weak acid will only partially ionise in water to yield a low concentration of H+ ions
-This means there will be more H+ ions per unit volume to react with in a strong acid
-This increases the frequency of effective collisions between H+ ions and reactant particles
-Hence, the stronger the acid, the higher the rate of reaction
-For example, CH3COOH (ethanoic acid) is a weak acid while HCl is a strong acid

-Therefore, the reaction of Mg + HCl → MgCl2 + H2 will take place faster than the reaction of Mg + CH3COOH → (CH3COO)2Mg + H2

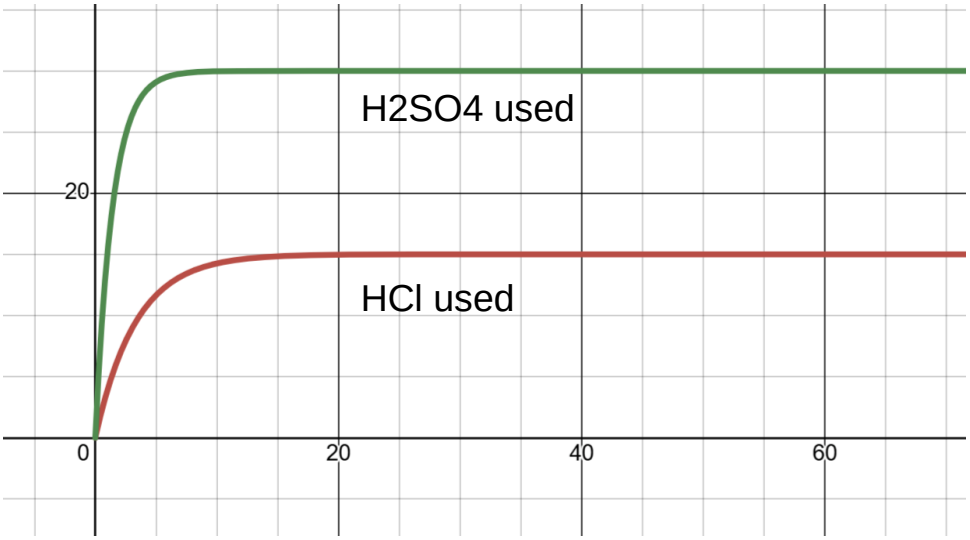


-The y axis is the volume of H2 produced and the x axis is the time

Basicity of acid

-The basicity of an acid can affect the rate of reaction
-A dibasic acid will produce twice the amount of H+ ions per mole of acid as compared to a monobasic acid
-This means there will be more H+ ions per unit volume to react with in a dibasic acid
-This increases the frequency of effective collisions between H+ ions and reactant particles
-Hence, the higher the basicity of an acid, the higher the rate of reaction
-Furthermore, the amount of product formed is also proportional to the basicity of the acid, hence if the basicity of an acid is doubled, the amount of product formed is also doubled
-For example, H2SO4 is a dibasic acid while HCl is monobasic

-Therefore, the reaction of Mg + H2SO4 → MgSO4 + H2 will take place faster than the reaction of Mg + HCl → MgCl2 + H2, and the volume of H2 produced will be doubled



-The y axis is the volume of H2 produced and the x axis is the time

Determining rate of reaction

-The rate of reaction can be determined by the change in temperature, formation of a precipitate, change in colour, change in pH value, change in gas volume and change in mass

Change in temperature

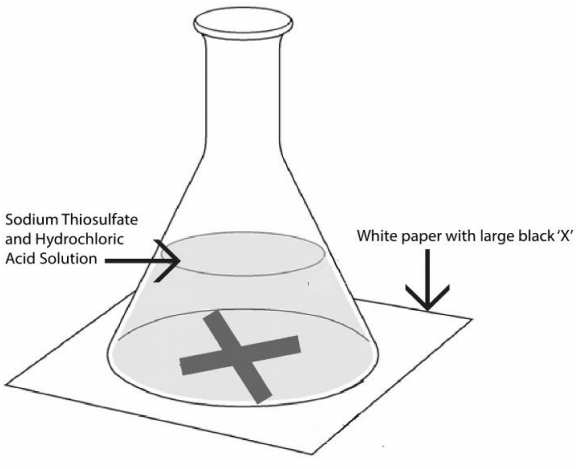
-The reaction $CH_4 + O_2 \rightarrow CO_2 + H_2O$ is exothermic, hence it releases energy, mainly in the form of heat

-Hence we can determine the rate of reaction by finding how the temperature changes with respect to time

Formation of a precipitate

-In the reaction $Na_2S_2O_3 + HCl \rightarrow NaCl + S + SO_2 + H_2O$, S is formed as an insoluble yellow precipitate

-We can then conduct the experiment in a conical flask and place a paper with a cross below it, and determine the time taken for the cross to disappear
-Hence we can determine the rate of reaction by finding how fast the cross disappears



Change in colour

-In the reaction Zn + CuSO4 → ZnSO4 + Cu, the reaction mixture will slowly turn from blue to colourless

-Hence we can determine the rate of reaction by finding the rate of colour change

Change in pH

-In the reaction NaOH + HCl → NaCl + H2O, the reaction mixture will slowly turn from acidic to neutral to alkaline as HCl is added to the NaOH

-Hence we can determine the rate of reaction by finding the rate of pH change

Change in gas volume

-In the reaction Mg + HCl → MgCl2 + H2, H2 gas is produced

-Hence we can determine the rate of reaction by finding the rate of gas produced

Change in mass

-In the reaction Na + H2O → NaOH + H2, the mass of the reaction mixture will slowly decrease as H2 gas is produced and leaves the container the reaction is taking place in

-Hence we can determine the rate of reaction by finding the rate of change in mass

Chapter 18: Fuels and Crude Oil

Fuels

-A fuel is a substance that burns in oxygen to produce heat energy
-There are 3 types of fuels, solid, liquid and gaseous
-Solid fuels include coal and charcoal
-Liquid fuels include petroleum, biogas, ethanol and biodiesel
-Gaseous fuels include natural gas and hydrogen

Fossil fuels

-Fossil fuels are formed from decayed plants and animals from millions of years ago
-It consists of 40% coal, 30% petroleum and 20% natural gas
-Fossil fuels are non renewable sources of energy as there are limited amounts of these sources and they cannot be replaced quickly enough to keep up with our rate of use

Natural gas

-Natural gas is mainly made up of 70-90% methane and other alkanes, and is formed from the decay of animal or plant waste
-It is considered the cleanest fossil fuel in terms of its impact on the environment
-Natural gas is usually transported in the liquid state as liquids take up less space than gases for transportation and storage

Petroleum

-Petroleum is also known as crude oil, and is a dark brown viscous liquid mixture
-It consists of a mixture of hydrocarbons and must be separated using fractional distillation into many smaller useful fractions or fuel mixtures
-During the fractional distillation of crude oil, the oil is heated and then it evaporates
-The fractionating column is cooler at the top than at the bottom, so as the vapours rise up the column they cool down
-Each fraction then condenses at a different temperature, and the fractions with the lowest boiling point condenses and are removed at the top, while those with the highest boiling points condense at the bottom

Decreasing height of fractionating column ↓	Fraction	State at r.t.p	Uses
	petroleum gas	gas	fuel for cooking
	petrol	liquid	fuel for motorcars
	naphtha	liquid	making chemicals such as alcohols and plastics

	kerosene (paraffin)	liquid	fuel for jet aircraft and fuel for heating and cooking
	diesel oil	liquid	fuel for diesel engines of buses, taxis and lorries
	lubricating oil	liquid	lubricants for machines, as well as for waxes and polishes
	bitumen	solid	surfacing roads

-As you go down the fractionating column, the molecular size of the fractions increase
-As molecular size increases, melting and boiling points also increase, as there is an increase in intermolecular forces of attraction between particles, hence more heat energy is required to overcome these forces
-As molecular size increases, density increases
-As molecular size increases, viscosity increases, as there is an increase in intermolecular forces of attraction between molecules
-As molecular size increases, flammability increases, as the percentage of carbon in the mixtures of molecules increases
-90% of crude oil is used as a fuel and a low percentage is used as chemical feedstock to make new products
Hence if the demand for crude oil can be reduced, then more crude oil can be used to manufacture a wide range of products such as plastics and synthetic rubber

Biofuels

-Biofuels are alternative renewable energy sources to crude oil and natural gas
-Biofuels come from plants or animals, and can be replaced relatively quickly after use
-One example of biofuel is bioethanol, which is ethanol obtained from the fermentation of plants such as sugarcane or corn in the presence of yeast
-The main use of ethanol as an alternative fuel is as a replacement to petrol in cars

Advantages	Disadvantages
-Ethanol is renewable while petrol is not -Ethanol burns cleaner than petrol, as it does not produce SO_2 while producing less CO_2 per unit mass than petrol -When producing ethanol, plants are grown which absorb CO_2 from the atmosphere, hence offsetting some of the CO_2 produced from burning ethanol	-Ethanol burns to produce half as much energy per mass of petrol -Engines must be modified to run on ethanol -The demand for biofuel crops means greater demand on rainforest land -Although ethanol is in theory carbon neutral, this is not accounting fr the CO_2 emissions associated with growing, harvesting and transporting the crops, as well as the CO_2 when producing the ethanol

Hydrogen

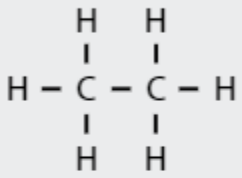
-Hydrogen is another alternative fuel that can be used for cars
-It is manufactured from the reaction of methane from natural gas with steam, from the cracking of crude oil fractions and through the electrolysis of water
-Hydrogen produces more energy per gram as compared to ethanol, but ethanol is easier to handle as it is a liquid at r.t.p while hydrogen is a gas, so it is more difficult to store
-Some methods of producing hydrogen also release CO_2 and other pollutants into the atmosphere

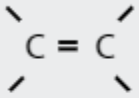
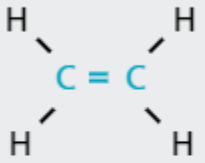
Chapter 19: Hydrocarbons

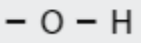
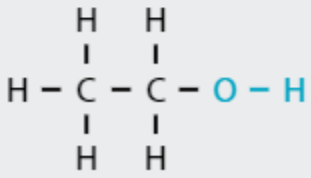
Homologous series

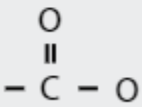
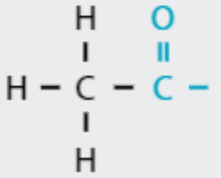
-A homologouse series consists of a family of compounds with the same general formula, and similar chemical properties because they have the same functional groups

Table 19.1 Some homologous series and their functional groups

Alkanes	
Functional group: Alkanes do not have any functional groups. There are only C–C and C–H bonds.	Example: 

Alkenes	
Functional group:  carbon–carbon double bond	Example: 

Alcohols	
Functional group:  hydroxyl group	Example: 

Carboxylic Acids	
Functional group:  carboxyl group	Example: 

-In general organic compounds in the same homologous series have the same functional group, the same general formula and have similar chemical properties
-Each successive member of the homologouse series differ by one unit of CH2
-There is a gradual change in physical properties such as boiling point or viscosity as we go down the series from one member to the next

Naming organic compounds

-Organic compounds are compounds which contain carbon
-The name of an organic compound is divided into two parts, the prefix which denotes the number of carbon atoms, and the suffix which denotes the homologous series

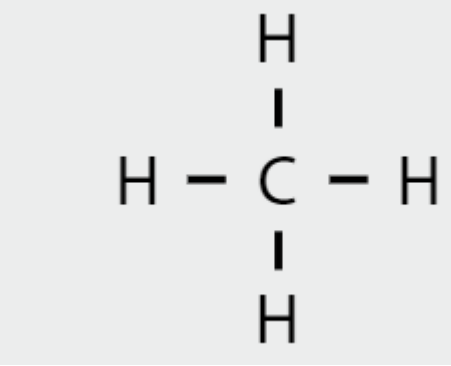
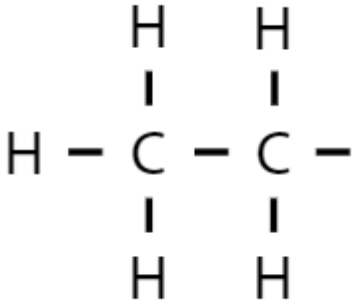
Prefix	Number of carbon atoms	Suffix	Homologous series

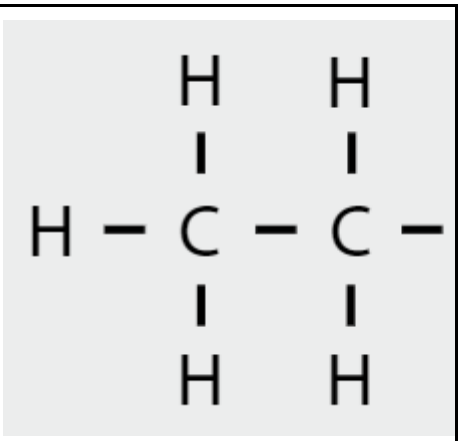
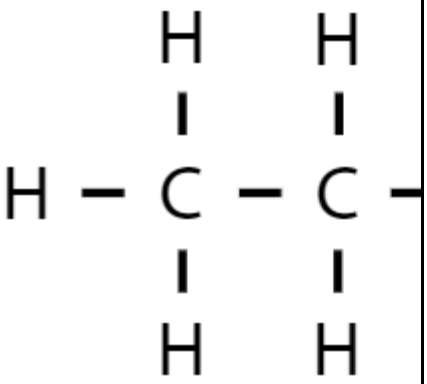
meth-	1	-ane	alkane
eth-	2	-ene	alkene
prop-	3	-ol	alcohol
but-	4	-oic acid	carboxylic acid

-For example, butanol would be an alcohol with 4 carbon atoms

Alkanes

-Alkanes are hydrocarbons that contain only carbon-carbon single bonds (C-C) and carbon-hydrogen bonds (C-H)
-A hydrocarbon is a compound that contains only hydrogen and carbon
-A carbon-carbon single bond is a single covalent bond between two carbon atoms
-The general formula of alkanes is CnH2n+2, where n is the number of carbon atoms in each molecule
-When there is 1 carbon molecule, n=1, so the formula of that alkane is CH4

Name	Molecular formula	Full structural formula
methane	CH4	
ethane	C2H6	

propane	C3H8	
butane	C4H10	

Physical properties of alkanes

-Alkanes are simple molecular substances with weak forces of attraction between molecules, hence having low melting and boiling points
-Melting and boiling points increase down the homologous series, as forces of attraction between alkane molecules increase as size increases
-Methane, butane, ethane and propane are all gases at r.t.p
-Alkanes become more viscous down the homologous series as their molecular sizes increase
-Alkanes are insoluble in water but soluble in most organic solvents such as tetrachloromethane (TCM)(CCl_4)
-Liquid alkanes are also often used as organic solvents themselves

Chemical properties of alkanes

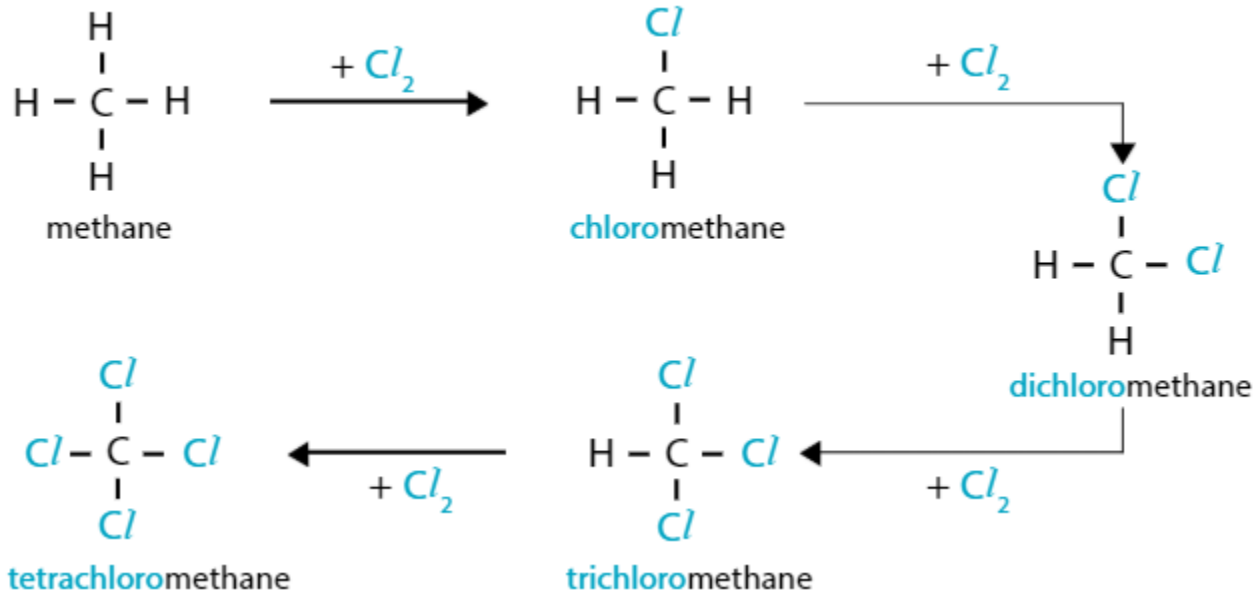
-Alkanes are generally unreactive since C-C and C-H bonds are generally hard to break
-However, alkanes can undergo combustion and substitution reactions
-Alkanes burn in excess oxygen to produce carbon dioxide and water in a combustion reaction
-For example, the combustion of butane is $2C_4H_{10}\ (g) + 13O_2\ (g) \rightarrow 8CO_2\ (g) + 10H_2O\ (g)$
-When there is insufficient oxygen, incomplete combustion occurs, producing soot (elemental carbon) and carbon monoxide

-For example, the incomplete combustion of butane is C4H10 (g) + 3O2 (g) → CO (g) + 3C (s) + 5H2O (g)

-Even if undergoing incomplete combustion, some CO2 can still be produced since some alkane molecules can still undergo complete combustion, though not shown in the equation above
-Alkanes can react with halogens in the presence of UV light in substitution reactions

-For example, methane can react with chlorine in the reaction CH4 (g) + Cl2 (g) → (UV light) CH3Cl (g) +HCl (g)

-In this reaction, a hydrogen atom is substituted by a chlorine atom to form chloromethane
-More hydrogen atoms can be replaced with chlorine atoms to form other compounds, such as dichlormethane, trichloromethane etc

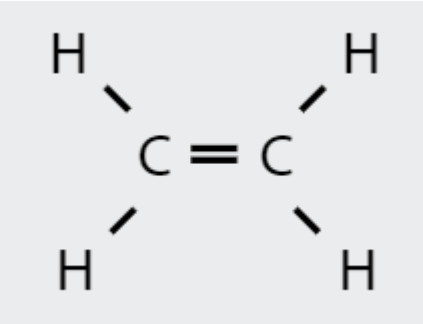
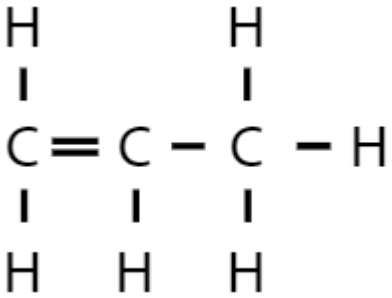
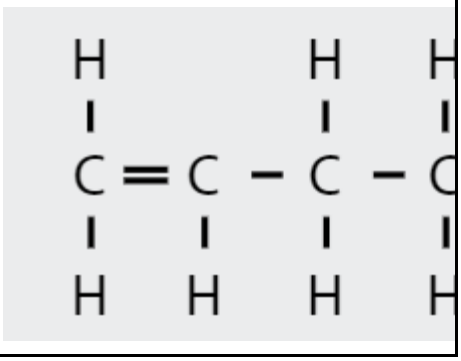


-The prefix corresponds to the number of halogen atoms and the type of halogen, and the suffix corresponds to the original alkane, so tetrabromoethane is an ethane molecule where 4 hydrogen atoms have been substituted with 4 bromine atoms

Alkenes

-Alkenes are hydrocarbons that contain carbon-carbon double bonds (C=C)
-A carbon-carbon double bond is a double covalent bond between two carbon atoms, which mean two electrons are shared from each carbon atom in the covalent bond
-The general formula of alkenes is CnH2n, where n is the number of carbon atoms in each molecule
-When there is 2 carbon molecule, n=2, so the formula of that alkane is C2H4
-The functional group of alkenes is the C=C bond

Name	Molecular formula	Full structural formula

ethene	C2H4	
propene	C3H6	
butene	C4H8	

Physical properties of alkenes

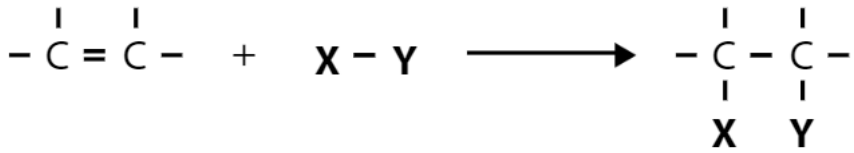
-Alkenes are simple molecular substances with weak forces of attraction between molecules, hence having low melting and boiling points
-Melting and boiling points increase down the homologous series, as forces of attraction between alkene molecules increase as size increases
-Ethene, propene and butene are all gases at r.t.p
-Alkenes become more viscous down the homologous series as their molecular sizes increase
-Alkenes are insoluble in water but soluble in most organic solvents such as TCM or hexane
-However, unlike liquid alkanes, they aren't used as solvents as commonly due to their reactivity interfering with reactions

Chemical properties of alkenes

-Alkenes can undergo combustion and addition reactions
-Like alkanes, alkenes burn in excess oxygen to produce carbon dioxide and water

-For example, the combustion of ethene is $C_2H_4 (g) + O_2 (g) \rightarrow 2CO_2 (g) + 2H_2O (g)$

-Compared to alkanes, the percentage by mass of carbon in alkene molecules is higher, hence alkenes are more likely to undergo incomplete combustion and produce carbon monoxide and elemental carbon
-Hence alkenes usually burn with a sootier flame
-The C=C bonds in alkenes are reactive, hence alkenes will readily undergo addition reactions, where the general formula is as shown:

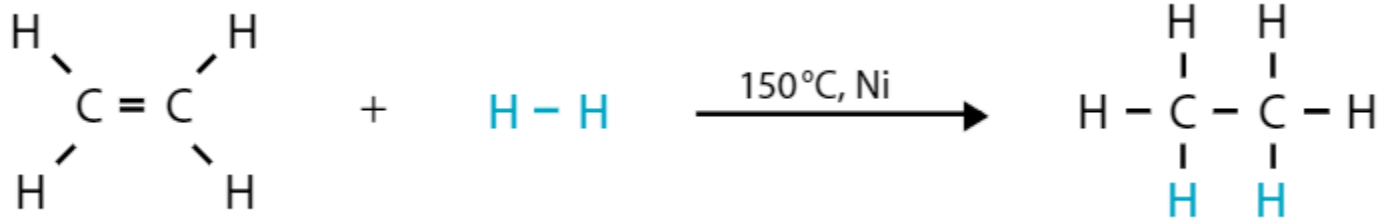


-The C=C bond breaks to form new single bonds, and as a result, each carbon atom that was part of the C=C bond forms 4 single covalent bonds instead
-Not all reactants can be added to the C=C bond, and some examples of reactants include H_2, Br_2, steam and other alkenes

Addition of hydrogen

-The addition reaction of an alkene with H_2 is also known as hydrogenation
-This can only occur at temperatures of around 150°C, and in the presence of a catalyst such as nickel

-The equation of the hydrogenation of ethene is: $C_2H_4 (g) + H_2 (g) \rightarrow C_2H_6 (g)$

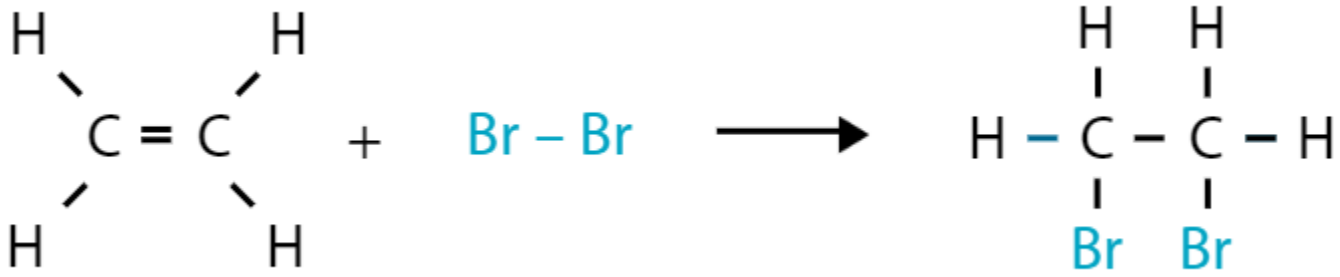


Addition of bromine

-The addition reaction of an alkene with Br_2 is also known as bromination
-Other halogens, such as Cl_2 and I_2 also work

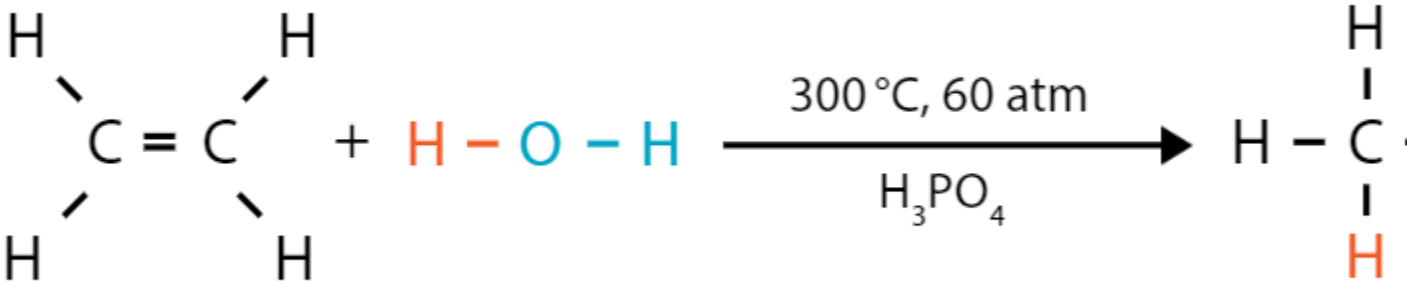
-The equation of the bromination of ethene is: $C_2H_4 (g) + Br_2 (l) \rightarrow CH_2BrCH_2Br (l)$

-The brown colour of the bromine disappears and a colourless liquid (CH_2BrCH_2Br or 1,2-dibromoethane) is formed



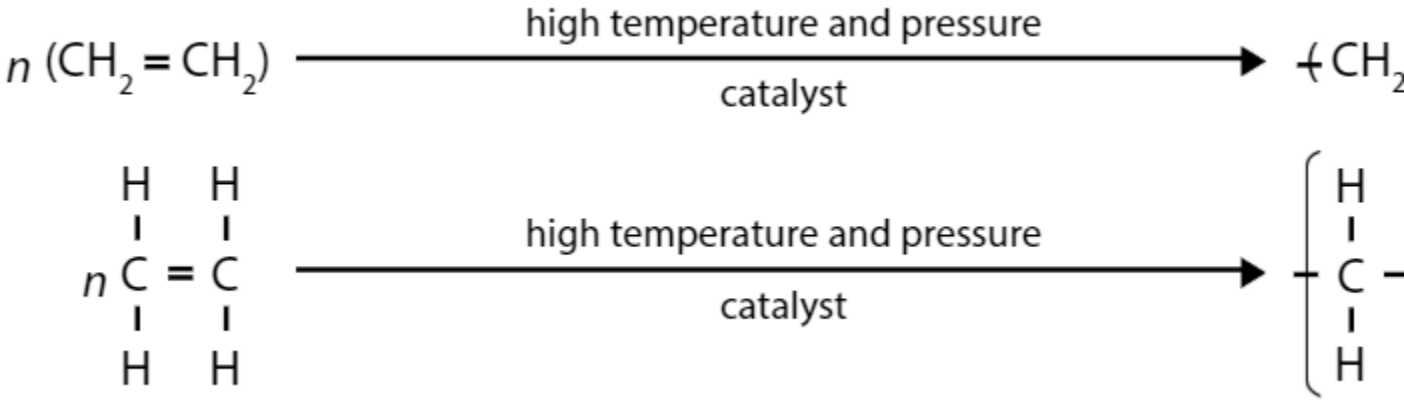
Addition of steam

--The addition reaction of an alkene with steam is also known as hydration
-This can only occur in temperatures of around 300°C, pressures of around 60 atm, and in the presence of a suitable catalyst, such as phosphoric acid (H3PO4)
-The equation for the hydration of ethene is: C2H4 (g) + H2O (g) → C2H5OH (aq)



Addition polymerisation

-At high temperatures and pressures, and in the presence of a catalyst, ethene molecules which in this case are monomers, can react with each other to form a long molecule called polyethene, which is a polymer
-This reaction is used to manufacture plastics



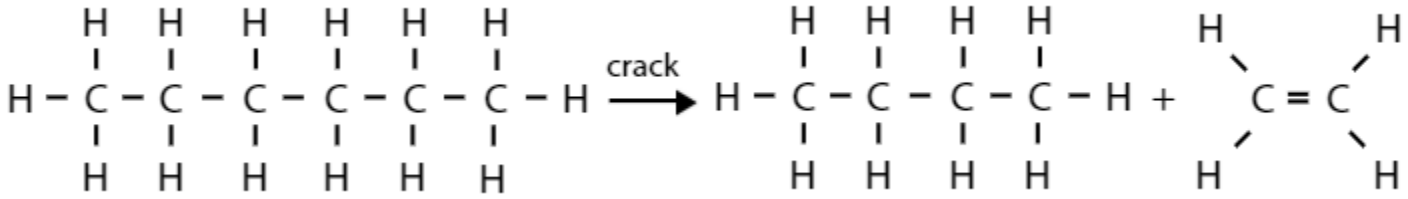
Cracking

-Cracking is a process in which larger hydrocarbon molecules, usually alkanes, are broken down into smaller hydrocarbon molecules

-Hydrogen gas can also be obtained in some cracking processes

There are many possible cracking reactions, for example the cracking equation for hexane is:

C6H14 → C4H10 + C2H4



-However, in the industry, a catalyst is used to speed up the process, known as catalytic cracking

-In industries, catalytic cracking is usually carried out in the presence of a catalysts such as Al2O3 and SiO2, high temperatures of around 500°C to 700°C, and pressures of around 1 atm

-The general equation for catalytic cracking is: long chain alkanes → mixture of short chain alkenes + mixture of short chain alkanes/hydrogen gas

-For example, the catalytic cracking of C2H6 is: C2H6 → C2H4 + H2, and the catalytic cracking of decane is: C10H22 → C2H4 + C8H18

Importance of cracking

-Cracking is important as it converts the less useful components of crude oil, usually the long chain alkanes, into shorter chain alkenes such as ethene and propene which are starting materials for many industrial processes and shorter chain alkanes, such as petrol, which are in high demand as fuel

-Cracking allows the supply from refining crude oil to match industrial demand for shorter chain alkanes and alkenes, while also producing H2 gas as a byproduct which is used in important industrial processes such as the haber process

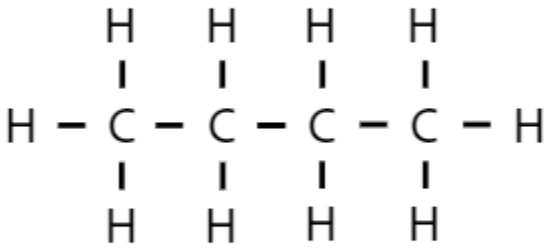
Isomers

-Compounds that have the same molecular formulae but different structural formulae are called isomers
-Isomerism can occur when the length of the main carbon chain differs
-Straight chain iosmers contain the longest main carbon chains while branched chain iomsers contain shorter main carbon chains
-Isomers have different physical properties such as melting and boiling points

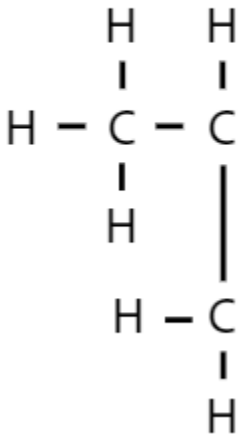
Isomers of alkanes

-For methane, ethane and propane, they do not have isomers
-However, butane has isomers
-Butane and methylpropane are isomers
-Butane is a straight chain alkane as all the carbon atoms are joined up in a row, while methylpropane is a branched chain alkane because it has a branch or side chain, in this case -CH3
-Butane has a higher boiling point than methylpropane, as for straight chain isomers, there is a greater surface area which increases the intermolecular forces of attraction between molecules, increasing the boiling point
-This is true for all straight chain and branched chain isomers

Butane (C_4H_{10})



Methylpropa

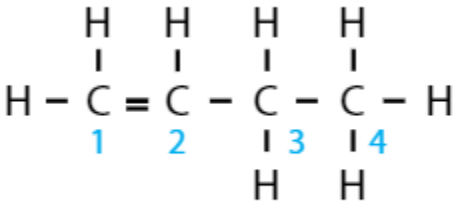


Alkyl groups

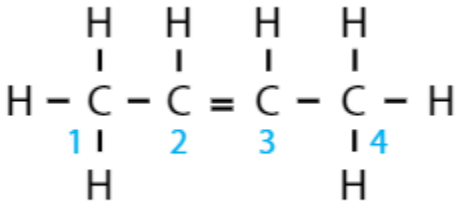
-An alkyl group is formed by removing a hydrogen atom from an alkane molecule, and it has the same general formula CnH2n+1
-For example, in methylpropane, the side chain is a methyl group
-Longer side chains result in different alkyl groups, such as ethyl groups (-C2H5) or propyl groups (-C3H7)

Isomers of alkenes

-Unlike alkanes, due to the C=C bond in alkenes, different placements of this C=C bond can result in different straight chain isomers, while alkanes only have 1 straight chain isomer
-For example, but-1-ene and but-2-ene are 2 straight chain isomers of butene, where the C=C bond is placed to the side or in the centre respectively



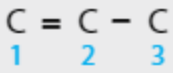
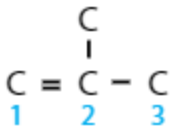
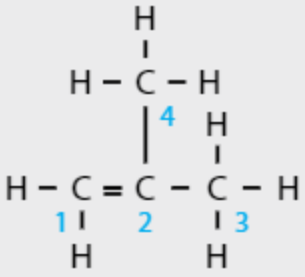
but-1-ene



but-2-ene

-Branched chain isomers of alkenes can be drawn by shifting the placement of the alkyl groups in straight chain alkenes

Table 19.10 Deducing the full structural formula of the branched-chain isomer of butene

Step 1	Draw the main carbon chain. It should only have 3 carbons. There is only one way to place the double bond here as placing the double bond between C2 and C3 will lead to the same structure.	
Step 2	Determine the position of the side chain. To make a branched isomer, the side chain can only be attached to C2. Attaching the side chain to carbons C1 and C3 which are located at the ends of the main carbon chain will give you the structure of a straight-chain alkene.	
Step 3	Add in the hydrogen atoms such that all carbon atoms form four bonds.	Thus, the structural formulae of the branched-chain isomer of butene is: 

Saturated and unsaturated hydrocarbons

-Although both alkanes and alkenes are hydrocarbons, they have different chemical properties
-Alkanes are known as saturated hydrocarbons as they only have single carbon bonds and each carbon atom is bonded to 4 other atoms
-Alkenes are known as unsaturated hydrocarbons as they contain double carbon bonds in addition to single carbon bonds
-You can think of a saturated hydrocarbon as one that has a lot more hydrogen atoms, or is saturated with hydrogen atoms, due to the C-C bond allowing for more hydrogen atoms to bond per carbon atom as compared to with a C=C bond

Test for unsaturation

-Aqueous bromine is commonly used to test for unsaturation, as aqueous bromine reacts with alkenes in a bromination reaction
-Hence, aqueous bromine remains brown when shaken with alkanes but decolourises or turns from brown to colourless when shaken with alkenes

Fats and oils

-A fat molecule contains 3 hydrocarbon chains that can be saturated or unsaturated
-When 2 or more C=C bonds exist in the same chain, we call it a polyunsaturated chain
-Polyunsaturated fats contain hydrocarbon chains with two or more C=C bonds in each chain

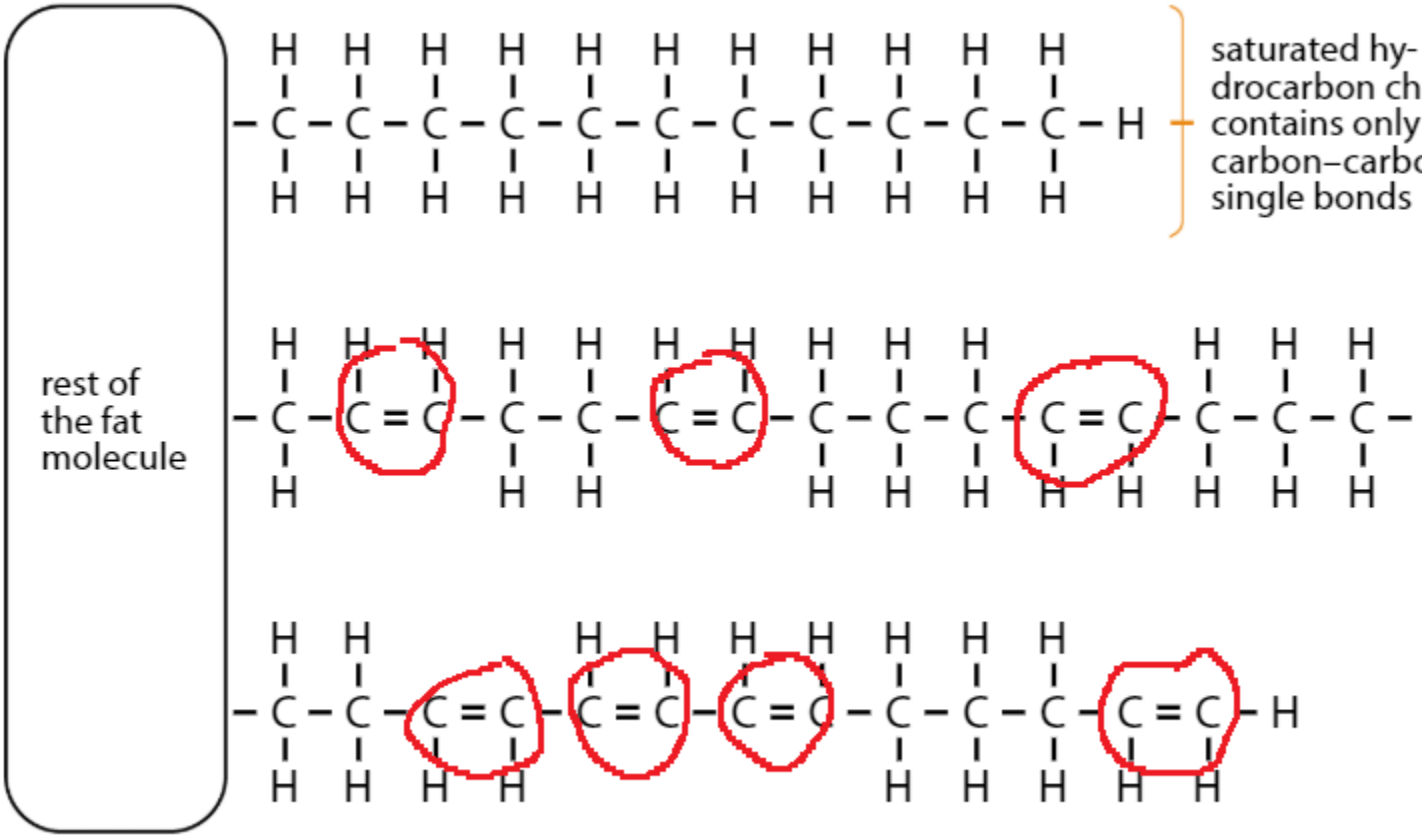
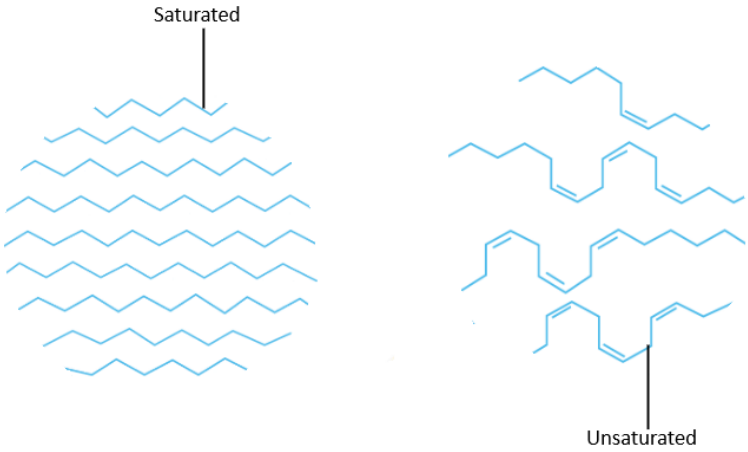


Figure 19.25 A fat molecule contains three hydrocarbon chains.

Properties of unsaturated and saturated fats

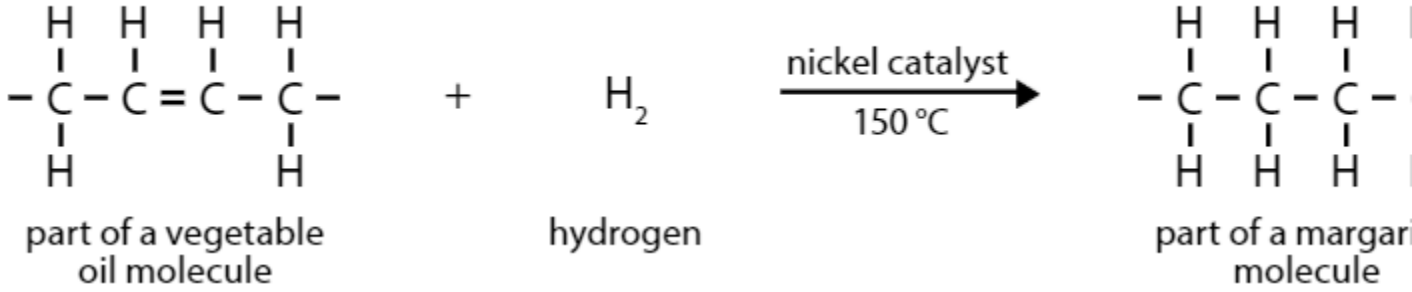
-The saturated and unsaturated hydrocarbon chains in fat molecules result in different properties
-The greater the level of saturation (less C=C bonds), the higher the melting point
-This is because saturated fats have hydrocarbon chains which stack well, while the presence of C=C bonds in unsaturated fats causes a bend in the structure, resulting in these chains not stacking as well

-This means saturated hydrocarbon chains are closer together and have stronger intermolecular forces of attraction, resulting in more energy needed to overcome these forces and a higher melting point
-Saturated fats are usually solids at r.t.p, while unsaturated fats, such as oils, are usually liquids at r.t.p



Manufacture of margarine

-In industries, margarine is a solid at room temperature and can be obtained from the hydrogenation of vegetable oil
-This process is done as margarine is more convenient to use and can usually be kept longer than vegetable oil
-This can only occur at temperatures of around 150°C, and in the presence of a catalyst such as nickel as well as with hydrogen gas

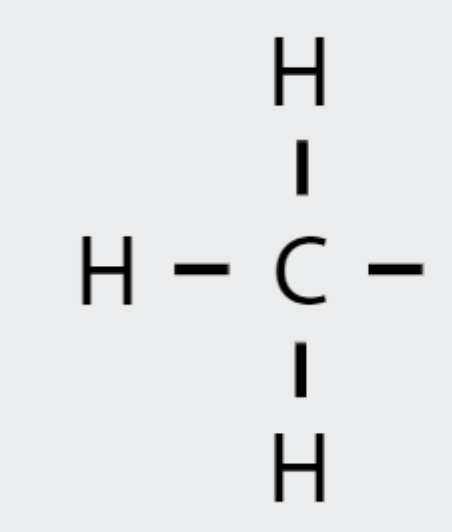
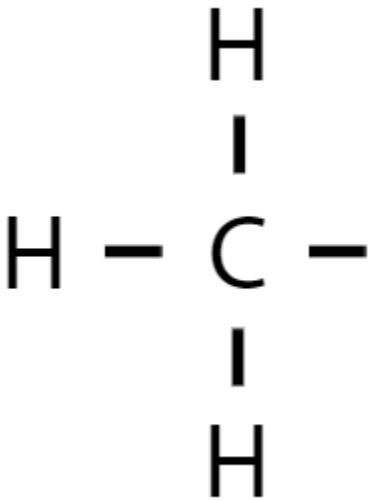


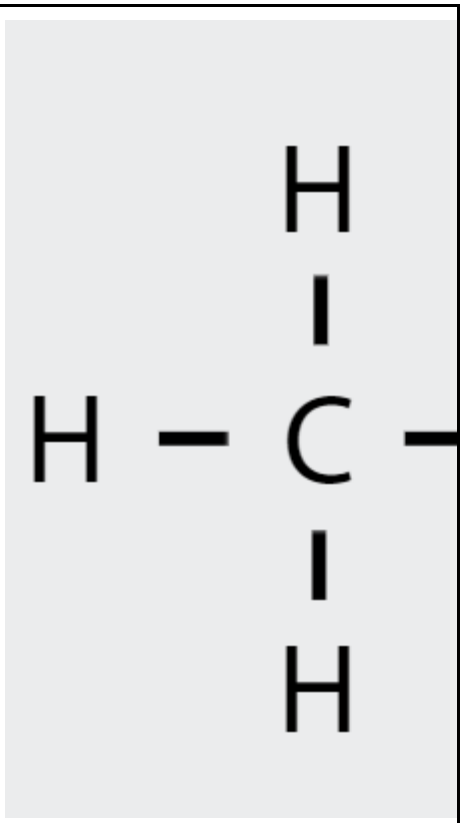
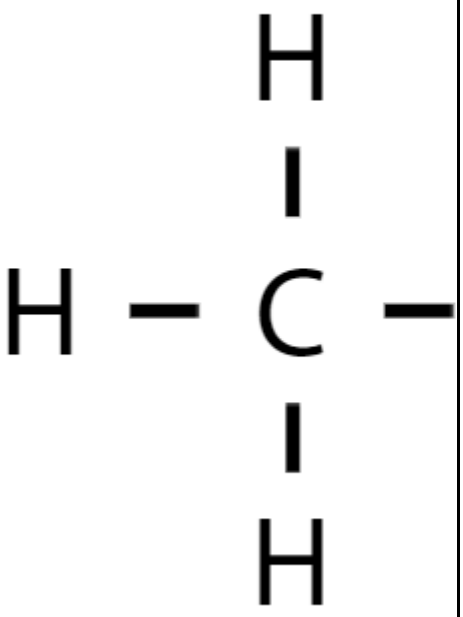
-As more hydrogen molecules are added to the oil molecule, there are fewer and fewer C=C bonds and the melting point of the oil increases, hence why vegetable oil is a liquid while margarine is a solid at r.t.p
-However, this process is a partial hydrogenation, which means some of the margarine produced still contains unsaturated hydrocarbons
-Trans fats can also be formed due to the reaction of the hydrocarbon chain with the catalyst, which are associated with heart diseases

Chapter 20: Alcohols, Carboxylic Acids and Esters

Alcohols

-Alcohols are organic compounds with the hydroxyl (-OH) functional group
-Since alcohols all contain carbon, hydrogen and oxygen atoms, they are not hydrocarbons
-The general formula of alcohols is CnH2n+1OH, where n is the number of carbon atoms in each molecule
-When there is 1 carbon molecule, n=1, so the formula of that alcohol is CH3OH
-The -OH group is always connected to the carbon atom at the end of the carbon chain for a straight chain alcohol
-This carbon atom is known as the terminal carbon atom

Name	Molecular formula	Structural formula	Full structural formula
methanol	CH3OH	CH3OH	
ethanol	C2H5OH	CH3CH2OH	

propanol	C3H7OH	CH3CH2CH2OH	
butanol	C4H9OH	CH3CH2CH2CH2OH	

-Isomers of alcohols can be obtained by shifting the alkyl groups like alkanes and alkenes, but also by shifting the hydroxyl groups
-Note that the structures shown above are the straight chain isomers of the alcohols

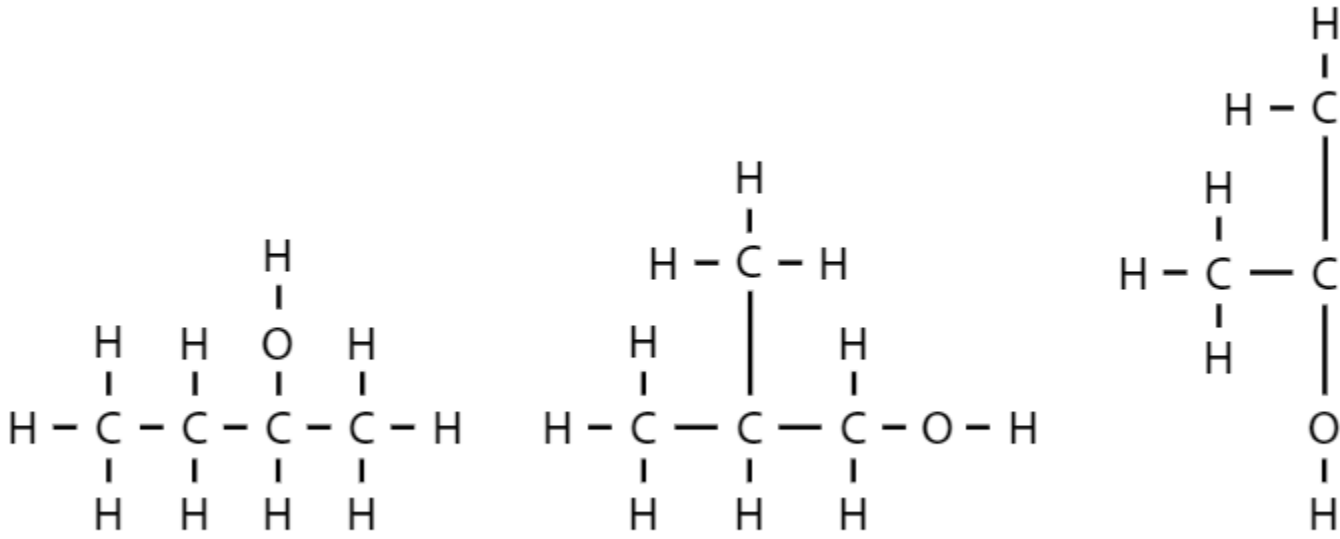


Figure 20.5 Isomers of butanol

Physical properties of alcohols

-Shorter chain alcohols are volatile liquids at r.t.p, and have higher melting and boiling points compared to hydrocarbons with the same number of carbon atoms, for example ethane has a boiling point of -89°C, while ethanol has a boiling point of 78°C
-Solubility decreases down the homologous series as the carbon chain increases in length
-Melting and boiling points increase down the homologous series as the molecular size increases, increasing intermolecular forces of attraction between alcohol molecules

Chemical properties of alcohols

-Alcohols can undergo combustion, oxidation and esterification reactions
-The complete combustion of an alcohol produces carbon dioxide and water

-For example, the complete combustion of methanol is $2CH_3OH\ (l) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 4H_2O\ (g)$

-Most alcohols can undergo oxidation when heated with an oxidising agent such as acidified potassium manganate (VII)
-For alcohols where the -OH group is attached to the terminal carbon atom, a carboxylic acid is obtained upon oxidation

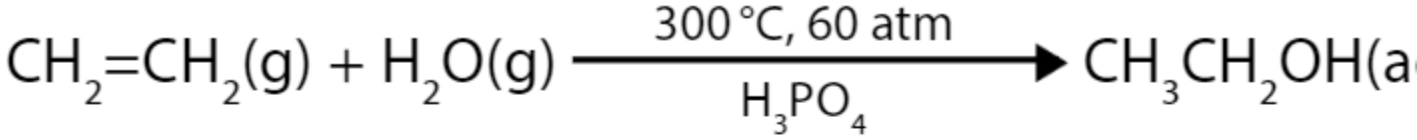
-For example, the oxidation of straight chain ethanol is $CH_3CH_2OH\ (aq) + 2[O] \rightarrow CH_3COOH\ (aq) + H_2O\ (l)$

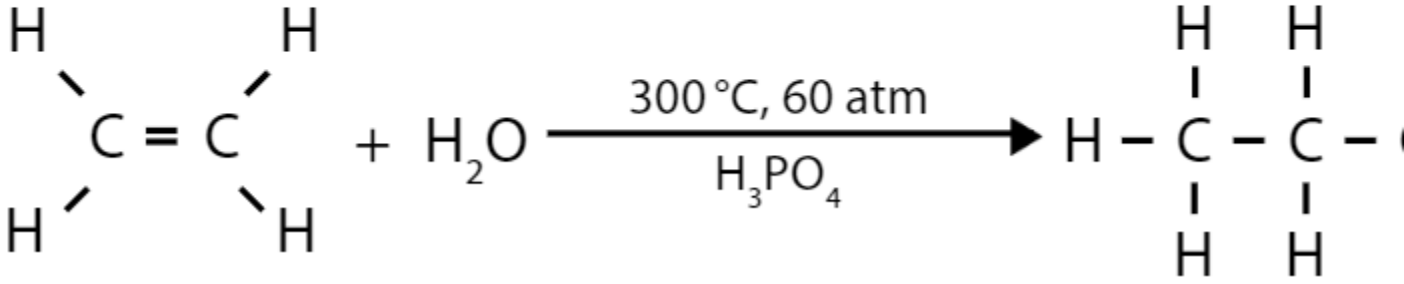
-Note that [O] represents the oxygen from the oxidising agent
-The carboxylic acid produced has the same number of carbon atoms as the initial alcohol, so ethanoic acid in this case

Formation of ethanol

-Ethanol is used as the main ingredient in sanitisers and fuel additives, as well as an organic solvent in industry
-Hence there are many ways to make ethanol

Catalytic addition of steam to ethene
-This is the main method for the industrial production of ethanol
-This can only occur at 300°C and 60 atm, and with an H3PO4 catalyst
-The equation of the catalytic addition of steam to ethene is: C2H4 (g) + H2O (g) → CH3CH2OH (aq)
b





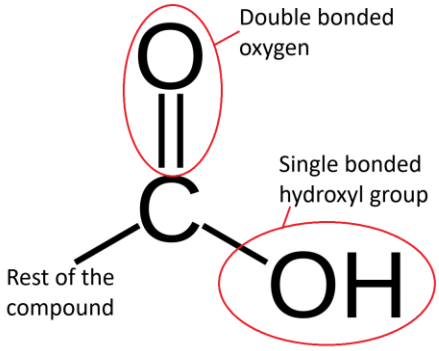
-Ethene required for this reaction is usually obtained from the refining and cracking of crude oil, hence this method is not sustainable as it requires finite resources
-Note that this is a hydration reaction

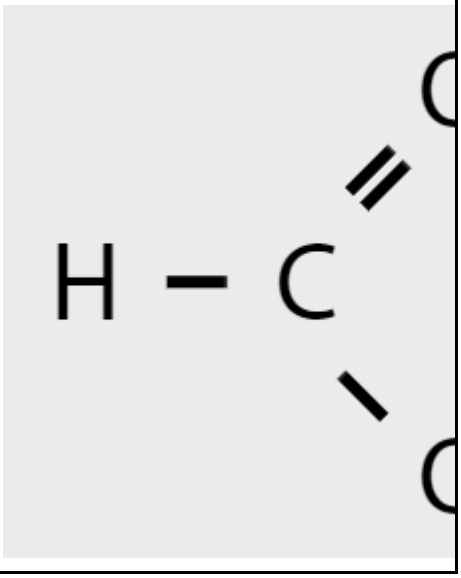
Fermentation of glucose
-Alcoholic drinks such as wine and beer contain ethanol, and to produce this ethanol fruits, vegetables or grains are fermented
-Alcohol fermentation is a process where microorganisms such as yeast act on carbohydrates such as glucose to produce ethanol and carbon dioxide
-In industries, fermentation takes place in fermentation tanks, and when fermentation is complete, the yeast settles to the bottom and alcohol is pumped out of the tanks
-The equation for the fermentation of glucose is: C6H12O6 (aq) → 2C2H5OH (aq) + 2CO2 (g)

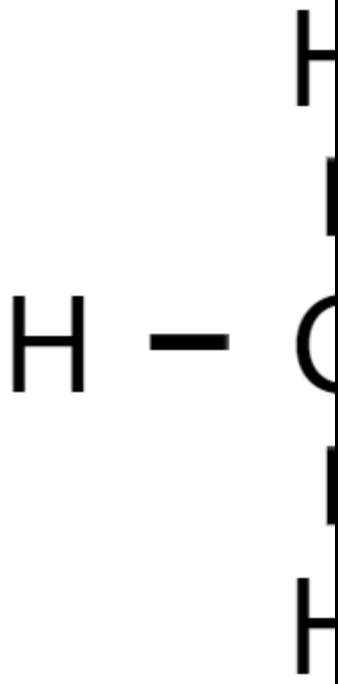
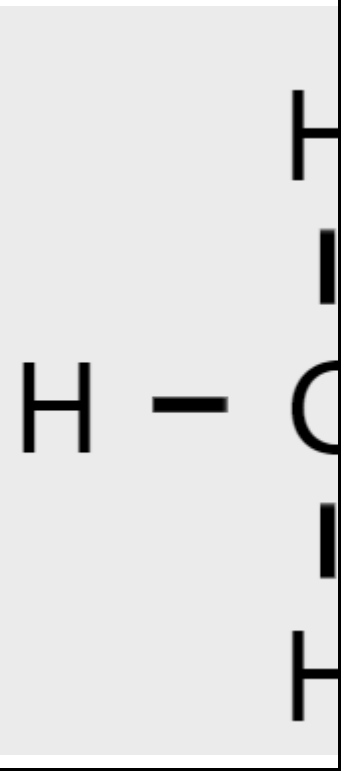
-This is a slower reaction compared to the catalytic addition of steam to ethene, but it is more sustainable as the starting materials come from plants, which are considered renewable resources

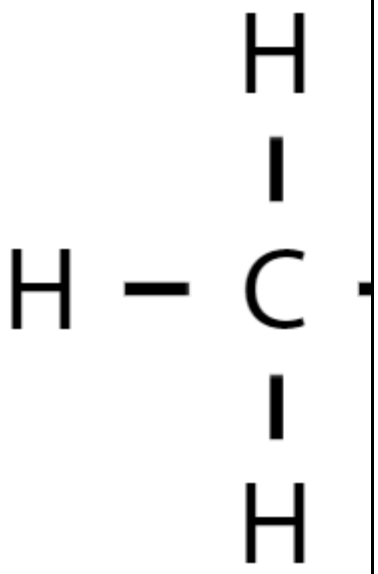
Carboxylic acids

-Carboxylic acids are organic compounds with the carboxyl (-COOH) functional group
-Since carboxylic acids also all contain carbon, hydrogen and oxygen atoms, they are not hydrocarbons, like alcohols
-The general formula of all carboxylic acids is CnH2n+1COOH, where n is the number of carbon atoms in the molecule minus one
-When there are 2 carbon molecules, n=1, so the carboxylic acid is CH3COOH
-As shown below, the -COOH group consists of a oxygen double bonded to a carbon atom and a hydroxyl group single bonded to the same carbon atom



Name	Molecular formula	Structural formula	Full structural formula
methanoic acid	CH2O2	HCOOH	

ethanoic acid	C2H4O2	CH3COOH	
propanoic acid	C3H6O2	CH3CH2COOH	

butanoic acid	C4H8O2	CH3CH2CH2COOH	

Physical properties of carboxylic acids

-Carboxylic acids have higher melting and boiling points compared to alcohols and hydrocarbons with the same number of carbons, for example ethane has a boiling point of -89°C, ethanol has a boiling point of 78°C and ethanoic acid has a boiling point of 118°C
-Boiling point increases down the homologous series as the molecular size increases, increasing intermolecular forces of attraction between carboxylic acid molecules

Chemical properties of carboxylic acids

-Carboxylic acids are weak acids because they partially ionise i water to form a carboxylate ion andhydrogen ion in the equation RCOOH (aq) ⇌ RCOO- (aq) + H+ (aq), where R represents the alkyl group, or the rest of the carboxylic acid (methyl for ethanoic acid, ethyl for propanoic acid etc), and RCOO- is the carboxylate ion (ethanoate for ethanoic acid, propanoate for propanoic acid etc)
-Since carboxylic acids are still acids, they still react with the same reactions as other acids
-Equation for reaction of carboxylic acid and reactive metal: 2CH3COOH (aq) + Mg (s) → (CH3COO)2Mg (aq) + H2 (g)
-Equation for reaction of carboxylic acid with carbonate: 2HCOOH (aq) + Na2CO3 (s) → HCOONa (aq) + CO2 (g) + H2O (l)

-Equation for reaction of carboxylic acid with base: 2CH3CH2COOH (aq) + Ca(OH)2 (s) → (CH3CH2COO)2Ca (aq) + 2H2O (l)

-Note that the carboxylate ion comes first when writing compounds, where the cation essentially replaces the hydrogen atom in the -COOH group

Formation of ethanoic acid

-Ethanoic acid is a colourless liquid at r.t.p and is just pure vinegar as vinegar is dilute ethanoic acid, and it is used in food as a preservative and flavouring
-There are 2 ways to form ethanoic acid, both involving the oxidation of ethanol

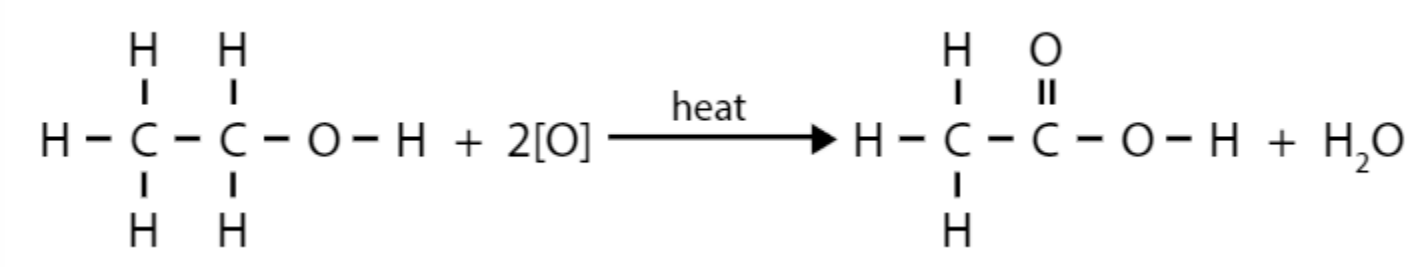
Atmospheric oxygen
-In industry, ethanoic acid is often produced by the oxidation of a solution of ethaol with atmospheric oxygen
-Ethanol reacts with atmospheric oxygen in the presence of a type of bacteria present in air to become ethanoic aid

-The equation for the oxidation of ethanol by atmospheric oxygen is: CH3CH2OH (aq) + O2 (g) → CH3COOH (aq) + H2O (l)

-This is also why wine can turn into vinegar after some time

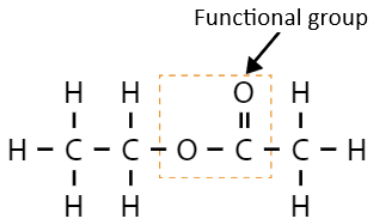
Acidified potassium manganate (VII)
-In a lab, ethanoic acid is prepared by heating a mixture of ethanol and acidified potassium manganate (VII)

-The equation for the oxidation of ethanol by acidified potassium manganate (VII) is: CH3CH2OH (aq) + 2[O] → CH3COOH (aq) + H2O (l)



Esters

-Esters are organic compounds formed by the reaction of alcohols with carboxylic acids
-The functional group of an ester is an L-shaped group with an oxygen atom single bonded to a carbon atom, double bonded to another oxygen atom, as shown below



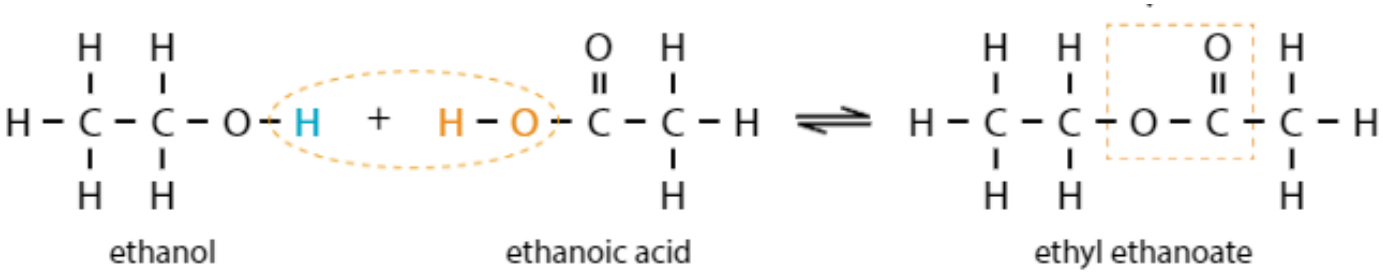
-Fats are also made up of esters, and contain 3 functional groups hence being called triesters

Formation of esters

-When a carboxylic acid is warmed with alcohol in the presence of a few drops of concentrated H2SO4 as a catalyst, an ester and water are formed

-For example, ethanol and ethanoic acid react in the equation: CH3CH2OH (aq) + CH3COOH (aq) ⇌ CH3COOC2H5 (aq) + H2O (l)

-During the formation of the ester, the -OH group from CH3COOH is replaced by the CH3CH2-O- group from the CH3CH2OH
-This reaction is known as an esterification reaxtion or condensation reaction
-During the reaction, the O-H bond in the alcohol is broken, and the C-O bond in the acid is broken, causing the -H and H-O- to form the water molecule
-The alcohol and acid then join together to form the ester, with the parts where the bonds are broken forming the ester linkage, which is also the functional group



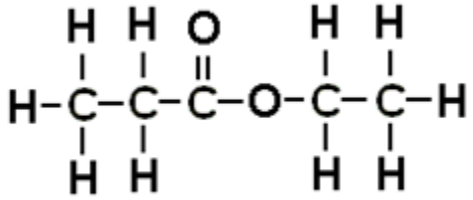
Formulae and naming of esters

-The general formula of an ester with one ester linkage is CnH2n+1COOCmH2m+1, where n is the number of carbon atoms in the carboxylic acid minus one, and m is the number of carbon atoms in the alcohol
-When CH3COOH and C2H5OH react, n=1 and m=2, so the ester formed is CH3COOCH3
-The name of an ester consists of two parts, the first part from the alcohol and the second from the carboxylic acid
-For example, an ester derived from ethanol and propanoic acid will be called ethyl propanoate

Number of carbon atoms in each part	1st part (alcohol)	2nd part (carboxylic acid)

1	methyl	methanoate
2	ethyl	ethanoate
3	propyl	propanoate
4	butyl	butanoate

-The 2nd part is the same as the names of the respective carboxylate ions
-To name an ester given the structure, first identify the ester linkage (find the parts with oxygen)
-Then, identify the alcohol part as the part connected to the single bonded oxygen of the ester linkage
-Then identify the acid part on the other side
-Count the number of carbon atoms for each part and name the ester accordingly, noting that the carbon in the ester linkage belongs to the carboxylic acid

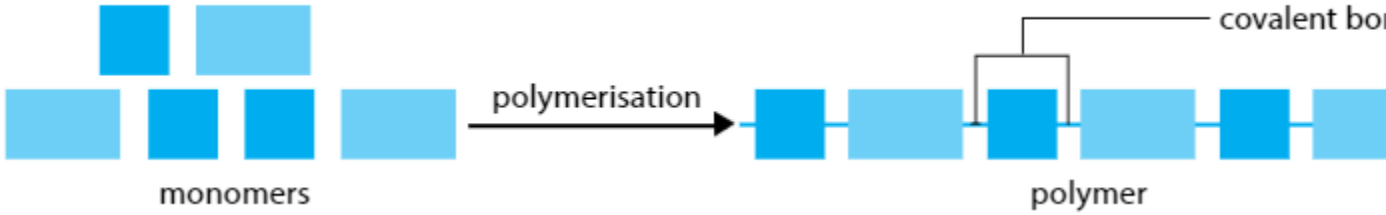


-For example in the structure above, the alcohol part is on the right and the carboxylic acid part is on the left
-The alcohol part has 2 carbon atoms, so ethyl
-The carboxylic acid has 3 carbon atoms, so propanoate
-Hence the ester is ethyl propanoate

Chapter 21: Polymers

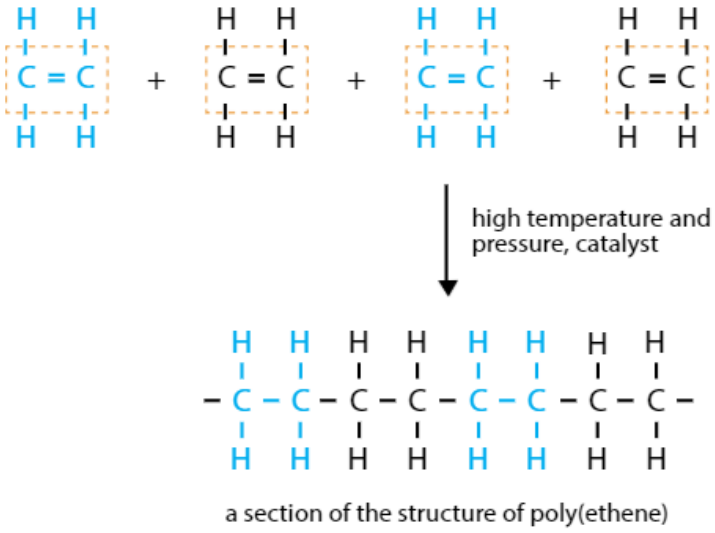
What is a polymer?

-A polymer is a very large organic molecule built from many small units called monomers
-There are different types of polymers, mainly natural and synthetic polymers
-Natural polymers include wood, natural rubber, silk, starch and cellulose
-Synthetic polymers include poly(ethene), nylon, terylene, polyvinyl chloride (PVC) and polypropylene
-The process of joining a large number of monomers to form a polymer is called polymerisation
-The monomers are joined together by covalent bonds in a polymer



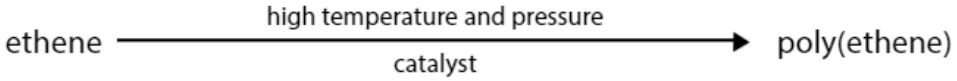
Addition polymerisation

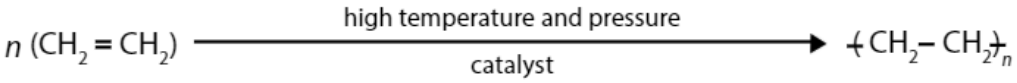
-Addition polymerisation occurs when unsaturated monomers join together without losing any molecules or atoms to form addition polymers
-Since the monomer used for addition polymerisation is usually unsaturated, that means alkenes can undergo addition polymerisation
-Poly(ethene) or polyethylene is the simplest addition polymer, produced by the addition polymerisation of ethene monomers
-At high temperatures and pressure, and in the presence of a catalyst, the C=C bonds in ethene break
-Each monomer then forms bonds with two other monomers, and eventually the polymer poly(ethene) is formed

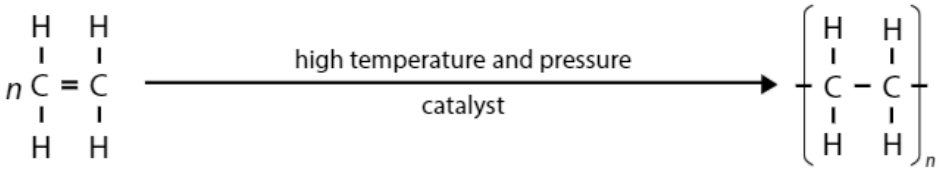


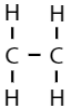
a section of the structure of poly(ethene)

-The equation for the polymerisation of ethene can be written as shown below, where n represents the number of repeat units, and is a large number



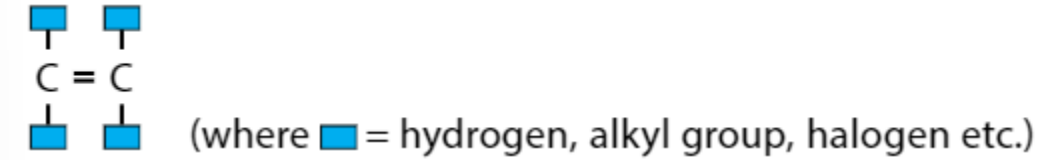




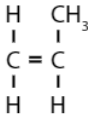
-A polymer is usually made up of similar units joined together called repeat units, and the repeat unit of poly(ethene) is 

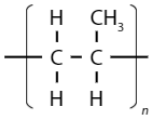
Deducing the structure of an addition polymer and monomer

-It is easier to deduce the structure of an addition polymer if the structural formula is written as shown below



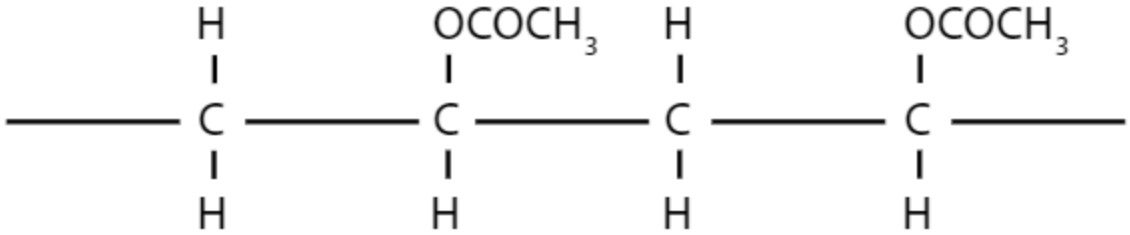
-This is because to deduce the structure you just have to replace the C=C bond with a C-C bond to form the polymer

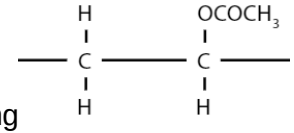
-For example, the formula of propene can be written as 

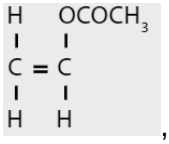
-Hence if you replace the C=C bond with a C-C bond, you get  , which is the formula for poly(propene)

-However, it is more difficult to deduce the structure of a monomer from a polymer

-For example, the structural formula of a section of PVA (polyvinyl acetate) is shown below

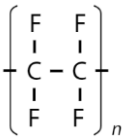


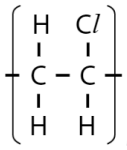
-In order to deduce the structural formula for the monomer of PVA, you first have to find the repeat unit in the polymer, in this case being 

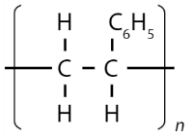
-Then, you convert the C-C bond in the repeat unit into a C=C bond to get the structural formula for the monomer, so in this case the monomer is  , or vinyl acetate

Uses of addition polymers

-Poly(ethene) is a type of plastic, and is used for making items such as plastic bags, toys, buckets and cling wraps, as it can be easily molded into various shapes

-Many frying pans are coated with a layer of polytetrafluoroethylene,  , as it is heat resistant and has non stick properties

-Polyvinyl chloride (PVC),  , is used to make thin gloves, pipes, raincoats and flooring mats

-Polystyrene,  , is hard, light and brittle, and is used to make disposable containers

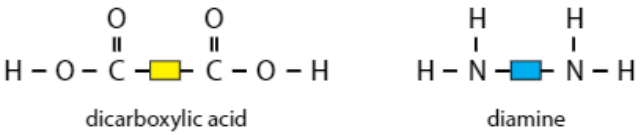
Condensation polymerisation

-Condensation polymerisation occurs when monomers combine to form condensation polymers, with the removal of small molecules such as water

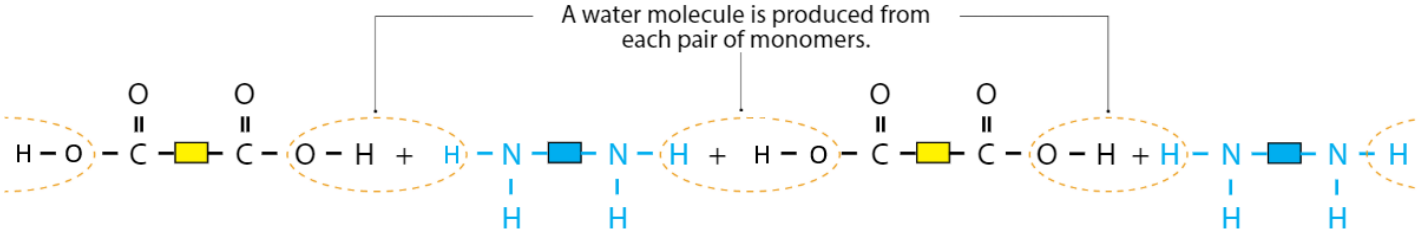
-There are two main groups of condensation polymers, polyamides and polyesters

Nylon

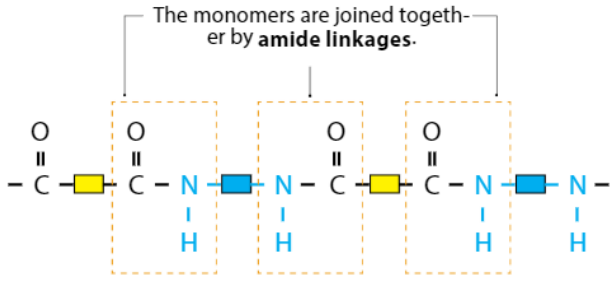
-Nylon is a condensation polymer and can be made from a dicarboxylic acid (molecule with two -COOH groups) and a diamine (molecule with two -NH2 groups)
-The dicarboxylic acid and diamine can be represented as shown below

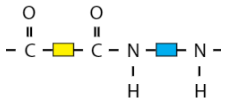


-When the dicarboxylic acid and diamine react, a water molecule is produced from each pair of the monomer



-These water molecules are then removed and the monomers are joined together by amide linkages to form nylon



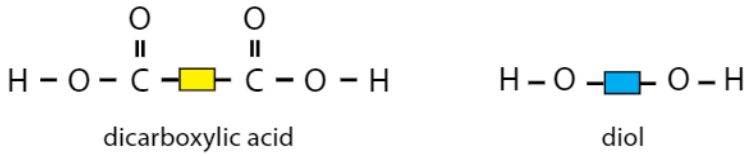
-The repeat unit of nylon is 

-The amide linkage is usually written as 

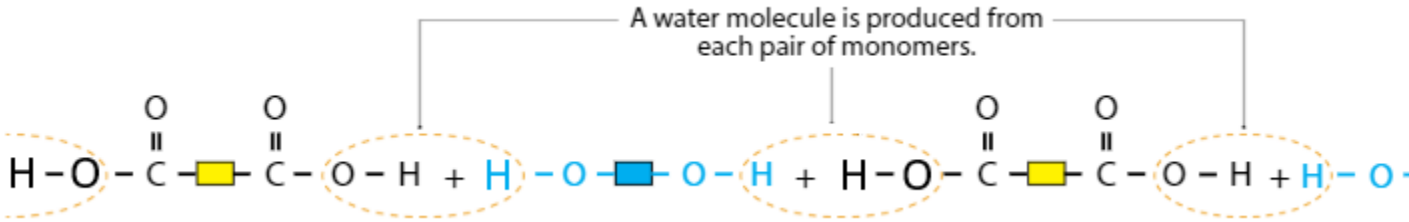
-Since nylon contains many amide linkages, it is a polyamide
-Note that for both polyamides and polyesters, 1 mole of water is produced per repeat unit of polymer, or nH2O

Terylene

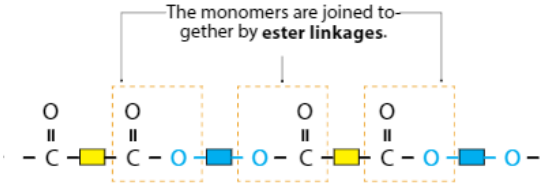
-Terylene is another example of a condensation polymer, and can be made from a dicarboxylic acid and a diol (molecule with two -OH groups)
-The dicarboxylic acid and diol can be represented as shown below

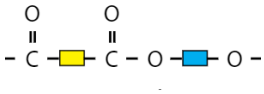


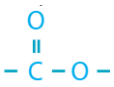
-When the dicarboxylic acid and diol react, a water molecule is produced from each pair of the monomer



-These water molecules are then removed and the monomers are joined together by ester linkages to form terylene



-The repeat unit of terylene is 

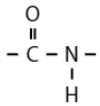
-The ester linkage is usually written as  and is identical to the ester linkage found in regular esters

-Since terylene contains many ester linkages, it is a polyester

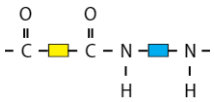
Deducing the structure of a condensation polymer and monomer

-In general, polyamides are formed when dicarboxylic acid and diamine are combined, while polyesters are formed when carboxylic acid and diol are combined

-When polyamides are formed, water is formed from the -OH of the -COOH group and the -H from the -NH2 group and is removed before the monomers are linked together by amide linkages

-Hence in the repeat unit we see the functional group  in the middle, while the dicarboxylic end will have the -OH removed from the -COOH group while the diamine end will have the -H removed from the -NH2 group

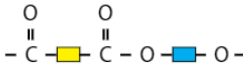
-So if the monomers are HOOC–⬜–COOH and H_2N–⬜–NH_2, the repeat unit would be

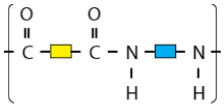


-When polyesters are formed, water is formed from the -OH of the -COOH group and the -H from the -OH group and is removed before the monomers are linked together by ester linkages

-Hence in the repeat unit we see the functional group  in the middle, while the dicarboxylic end will have the -OH removed from the -COOH group while the diol end will have the -H removed from -OH groups

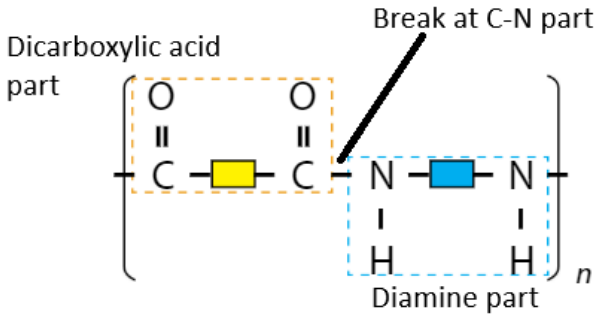
-So if the monomers are HOOC–⬜–COOH and HO–⬜–OH, the repeat unit would be



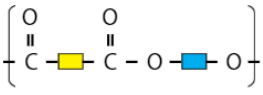
-When given a polyamide  , we know that the monomers are dicarboxylic acid and diamine

-To obtain the structural formula of the dicarboxylic acid, you break the amide linkage at the C-N part, and then add O-H to the C=O to get HOOC–⬜–COOH

-To obtain the structural formula of the diamine, you also break the amide linkage at the C-N part, and then add -H to the N-H to get H_2N–⬜–NH_2

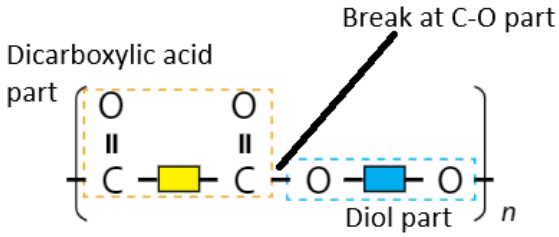


-To distinguish the dicarboxylic acid from the diamine, look out for N and C

-When given a polyester  , we know that the monomers are dicarboxylic acid and diol

-To obtain the structural formula of the dicarboxylic acid, you break the ester linkage at the C-O part, and then add O-H to the C=O to get HOOC–⬜–COOH

-To obtain the structural formula of the diol, you also break the ester linkage at the C-O part, and then add -H to the -O to get HO–⬜–OH



-To distinguish the dicarboxylic acid from the diol, look out for C

Uses of synthetic fibers

-Nylon and terylene are both synthetic fibers
-Synthetic fibers are usually strong and can be drawn into long thin strands without breaking, hence they can be used to make fishing lines
-Clothes made from synthetic fibers are shrink proof and crease proof, and are also easier to wash and dry
-Tents and sleeping bags are made from terylene, as unlike natural fibers, synthetic fibers do not shrink when exposed to water

Effect of plastics on our environment

-Increasingly, plastics are being used in place of natural materials such as wood as they are relatively cheap, easily moulded into various shapes, light, tough, durable and waterproof
-However, most plastics are non-biodegradable and hence cannot be broken down by bacteria or other living organisms in the soil
-This can result in many environmental problems

Land pollution	Water pollution	Air pollution
-Since plastics do not decompose, they take up a lot of space in landfills	-Plastics thrown into the sea endanger marine animals, as for example turtles can mistake plastic bags for food and choke on them - Plastics can clog up rivers and drains, which can become breeding grounds for mosquitoes	-Plastics are mostly flammable, and when plastics are incinerated, they produced toxic gases, as for example PVC produces corrosive HCl gas on combustion

Recycling plastics

-There are two methods to recycle plastics, the physical and chemical method
-However, regardless of the method, plastic waste must undergo pre-treatment, as not all plastics can be recycled and the method used for different types of plastics can be different

-Pre-treatment of plastic waste involves sorting by different methods (e.g through manual sorting or by density), washing to remove contaminants and shredding or grinding to smaller pieces

Physical method
-The physical method of recycling plastics is also known as mechanical recycling, and the chemical composition of the plastic remains the same in the recycled plastic
-After pre-treatment, small pieces of plastics such as poly(ethene) are melted, cooled, pulled into long, thin strands and then cut into pellets, which can then be made into new products

Figure 21.30 Plastic pellets obtained from the physical method of recycling

Chemical method
-The chemical method of recycling plastics converts plastic waste into different raw materials through different chemical reactions, mainly cracking and depolymerisation
-Plastic waste such as poly(ethene) can undergo cracking under high temperature and with a catalyst to obtain short chain alkanes and alkenes, which can be used as fuel and to make other chemicals individually
-Depolymerisation is a process in which polymers are broken down into their monomers
-These monomers can then be converted into other useful chemicals
-Polyesters can undergo depolymerisation through acid hydrolysis, which involves breaking down a polyester using water in the presence of an acid catalysts to form the products of a dicarboxylic acid and a diol

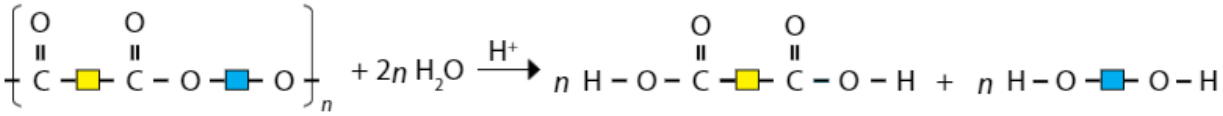


Figure 21.31 Acid hydrolysis of a polyester

Issues with recycling plastics

-It is important to minimise pollution caused by plastic waste and we should first reduce the plastic we use, then reuse plastics before finally recycling plastics, as recycling plastics still has some issues

al issues	Economic issues	Social issues

ing plastics, both methods can lead ntal issues as if wastewater om the recycling process is not erly before being discharged, it can water and cause water pollution	-Recycling plastic waste is more expensive than disposal in landfills and incineration, as there is the cost of transporting the waste to the processing plant, the cost of sorting and cleaning the waste -The processes involved in recycling plastic waste also require manpower, machines and energy, which cost money -It may not be economically viable to recycle plastics, as recycled plastics usually have a lower market value than virgin plastics made from raw materials, hence it would be difficult for recycling businesses to survive if the cost of recycling plastic waste is too high	-In Singapore, many people find it m convenient to just throw all their was hence recyclable plastics are dispos are not recycled -Many people are not aware of the p tom recycle plastics, as people may recyclable items with leftover food, c to be incinerated instead of being re -People also throw non recyclable ite recycling bins, increasing the time a separate recyclables, slowing down -It takes time and effort for communi recycling as a lifestyle, hence it may immediately effective

-To address these issues, the government can implement strict measures and laws to minimise potential environmental issues

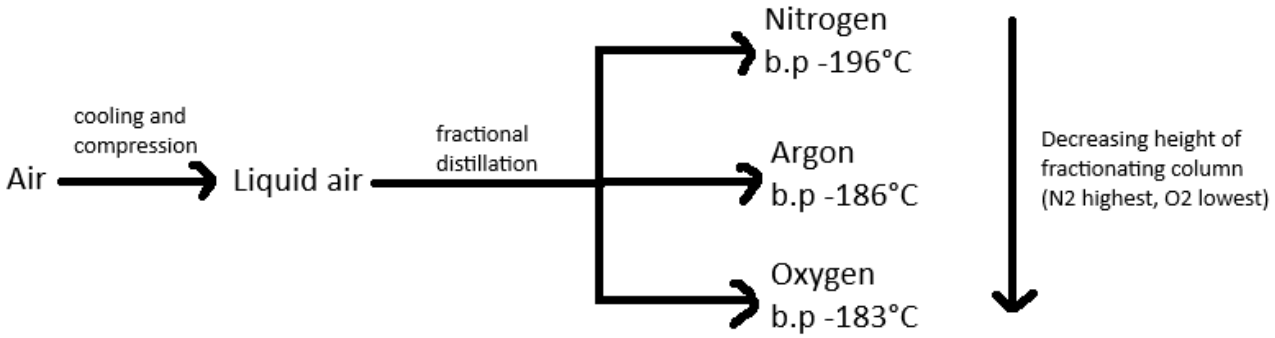
-Recycling programs can be introduced to educate the public on the correct way to recycle plastic waste

-Creative activities can be introduced to encourage recycling

Chapter 22: Maintaining Air Quality

Composition of air

-Dry air is composed of 78% nitrogen, 21% oxygen and the leftover being carbon dioxide and noble gases (mainly argon)
-These gases can be separated from each other through fractional distillation as they each have a unique boiling point
-Atmospheric air is liquified at very low temperatures and then allowed to slowly warm up inside a chamber attached to a fractionating column
-Nitrogen has the lowest boiling point so it is collected first, followed by argon and then oxygen



Air pollutants

-Air pollution is the introduction of unwanted and harmful chemicals into the atmosphere

Cause	Effect	Strategies to reduce
-Incomplete combustion of carbon containing fuels and petrol in car engines	-Toxic and causes respiratory problems -Binds irreversibly to haemoglobin in red blood cells, reducing their ability to carry oxygen -Causes headaches, fatigue and death -Can oxidise to form CO2	-Supply excess O2 to ensu complete combustion -Use catalytic converters in
-In car engines and factories as high temperatures provide energy for N2 and O2 to react and form NOx in the reaction N2 (g) + O2 (g) → 2NO (g) and 2NO (g) + O2 (g) → 2NO2 (g) -Also forms during thunderstorms as heat energy released by lightning causes N2 and O2 in the air to react and form NOx	-Causes breathing difficulties by irritating the lungs -Irritates the eyes -Can lead to inflammation of the lungs -Can react with rainwater to form acid rain	-Use catalytic converters in

-During volcanic eruptions -Combustion of fossil fuels which contain sulfur such as coal and crude oil in the reaction S (s) + O2 (g) → SO2 (g)		-Remove sulfur from fossil f before burning, however thi too expensive and difficult -Remove SO2 from waste g gases through flue gas desulfurisation
-Formed by the photochemical reaction between NO2 and unburnt hydrocarbons in the presence of sunlight -Also formed by the photochemical reaction between oxygen in the presence of sunlight -A photochemical reaction is a reaction catalysed or initiated by sunlight or UV light	-Irritates eyes and lungs, can cause asthma attacks	-Rudece vehicle emissions pollutants
-Formed when hydrocarbons combust with a lack of oxygen in vehicle combustion engines	-Irritates eyes and lungs -Can react with NO2 in a photochemical reaction to produce O3 as well as smog, which is a brown haze	-Supply excess O2 to ensu complete combustion
-Formed by the bacterial decay of plant and animal matter as well as the decay of rubbish in landfills -Also exhaled as waste gases from cattle	-Major greenhouse gas which can lead to global warming	Nil

Acid rain

-Unpolluted rainwater typically has a pH of around 5.0 to 5.5
-However, when rain combines with excess amounts of SO2 and NO2, it can have a pH around 4.0, which is around 10 times as acidic as normal rainwater

-This is because of the reactions of SO2 and NO2 with water, being 2SO3 (g) + 2H2O (l) →

2H2SO4 (aq) (SO2 oxidises in air to form SO3) and 4NO2 (g) + 2H2O (l) + O2 (g) → 4HNO3 (aq)

-In word form, this means that SO2 oxidises in air to form SO3, which then dissolves in rainwater to form sulfuric acid, and NO2 dissolves in rainwater to form nitric acid
-The formation of these strong acids is what lowers the pH
-Acid rain can react with metals and with carbonates in marble and limestone, damaging structures made of metal and marble
-Acid rain can reduce the pH of natural water bodies to below 4.0, killing fish and other aquatic life
-Acid rain also leaches important nutrients from the soil and destroys plants

Liming

-The effects of acid rain on forests and water bodies can be lowered through the application of $CaCO_3$
-The carbonate is added to the soil and water bodies to remove some of the excess acid contributed by acid rain through a process called liming
-However, liming has its limits as acid rain can affect vast areas, and liming is also expensive and a temporary measure

Catalytic converters

-Catalytic converters are located midway in the exhaust system of a vehicle, and contains a coating of catalysts made up of platinum, palladium and rhodium
-This speeds up the conversion of harmful substances in the vehicle exhaust into less harmful substances
-The catalysts are arranged in honeycomb structures that maximise the surface area for gases to interact with
-Gases that leave the engine typically contain CO_2, CO, N_2, NO_2, NO, O_2, C_xH_y and H_2O
-NO, NO_2, C_xH_y and CO are all air pollutants that can be removed by the catalytic converter in the following reactions

-$2CO\ (g) + 2NO\ (g) \rightarrow 2CO_2\ (g) + N_2\ (g)$

-$2CO\ (g) + O_2\ (g) \rightarrow 2CO_2\ (g)$

-$2NO_2\ (g) \rightarrow N_2\ (g) + 2O_2\ (g)$

-$C_xH_y\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$

-Even though the catalytic converter increases the amount of CO_2, it is still a preferred method of air pollution control as CO_2 is non toxic
-However CO_2 is still harmful as it is a greenhouse gas

Figure 22.6 The internal honeycomb structure of a catalytic converter

Flue gas desulfurisation

-Fossil fuels can have a lot of sulfur in them, but it is too expensive and difficult to remove sulfur from the fuel itself
Hence, SO_2 is removed from exhaust gases or flue gases in power plants that burn coal and fuel oil, as well as waste incinerators through flue gas desulfurisation

-The most widely used method of flue gas desulfurisation is wet scrubbing using a $CaCO_3$ slurry, which is formed when limestone is mixed with water

-The slurry droplets are sporayed through the flue gas and the $CaCO_3$ reacts with the SO_2 in the following reaction: $CaCO_3 (s) + SO_2 (g) \rightarrow CaSO_3 (g) + CO_2 (g)$

-The calcium sulfite ($CaSO_3$) is then further oxidised in the reaction $2CaSO_3 (s) + O_2 \rightarrow 2CaSO_4 (s)$

-The calcium sulfate can then be hydrated to form hydrated calcium sulfate or gypsum in the reaction $CaSO_4 (s) + 2H_2O (l) \rightarrow 2CaSO_4 . 2H_2O (s)$

-Gypsum can be used to make drywalls, fertilisers and plaster

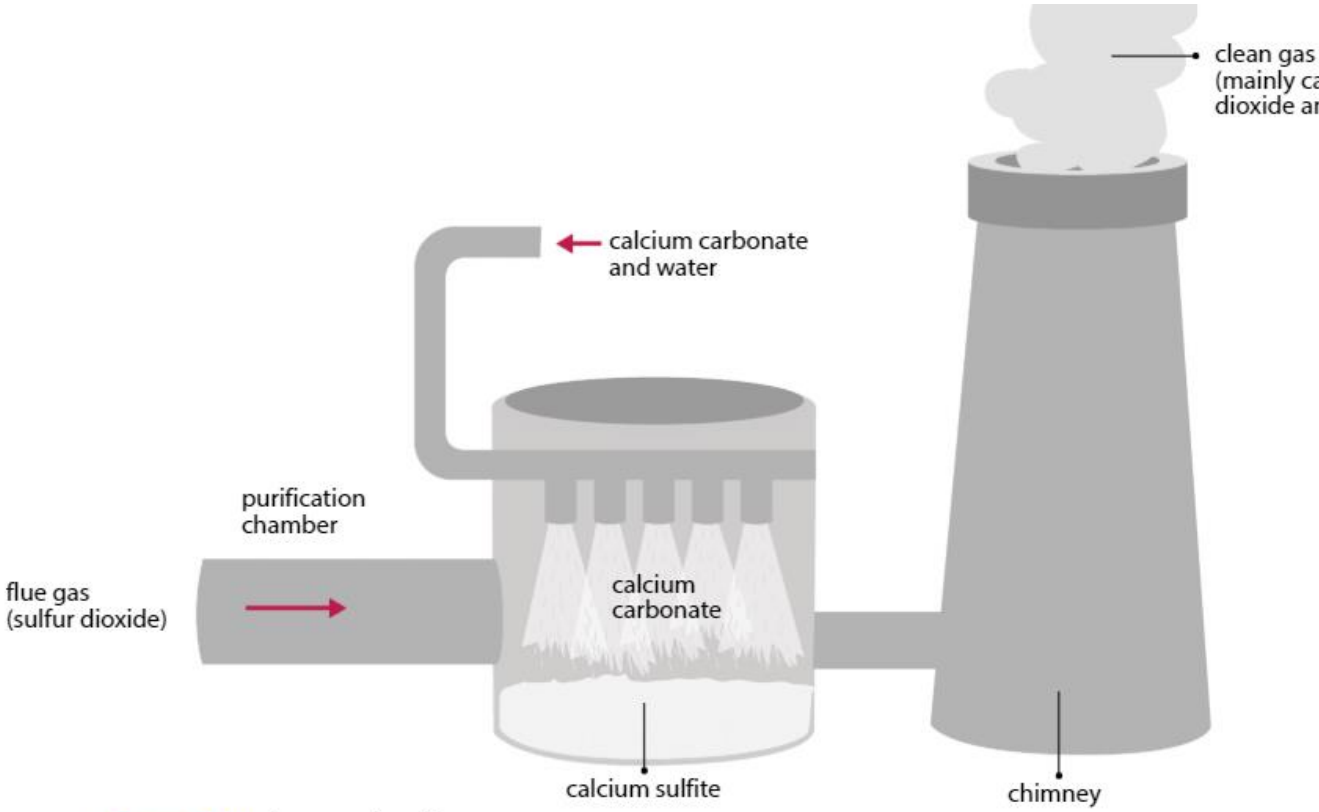


Figure 22.8 Flue gas desulfurisation

Low sulfur fuels

-Flue gas desulfurisation set ups require a lot of space and can be quite costly, so it is not practical to install them in vehicles

-Hence, countries enforce the use of low sulfur fuels in vehicles to reduce SO_2 emissions

-Excess sulfur can be removed from fossil fuels by treating them with H_2, while producing hydrogen sulfide as a byproduct

The ozone layer

-The ozone layer is around 15 to 30 km above sea level
-Although O_3 is considered a pollutant at ground level, its presence in the atmospher is beneficial as it can absorb UV radiation, reducing the amount of UV radiation that reaches the earth's surface
-O_3 molecules form when O_2 molecules interact with UV radiation from the sun
-The reaction that forms O_3 is reversible, hence O_3 can also break up to reform O_2 when they absorb UV radiation in the reaction $3O_2\ (g) \rightleftharpoons 2O_3\ (g)$
-This cycle of O_3 breaking up and then reforming happens naturally in the stratosphere and maintains the ozone layer
-Note that O_3 is an oxidising agent and can be tested for with potassium iodide

Depletion of the ozone layer

-Chlorofluorocarbons (CFCs) can be found in some aerosol propellants, refrigerants and certain kinds of plastics, and contain chlorine, fluorine and carbon
-These compounds are very light, hence they can rise up into the stratosphere when released, before interacting with UV radiation to produce Cl atoms
-These Cl atoms then react with O_3 to form ClO and O_2 in the reaction $Cl + O_3 \rightleftharpoons ClO + O_2$
-ClO can then break up another O_3 molecule in the reaction $ClO + O_3 \rightleftharpoons Cl + 2O_2$, releasing the Cl atom in the process and causing a chain reaction
-This results in O_3 molecules beong destroyed faster than they can be replaced by nature, hence causing holes in the ozone layer to form, depleting the ozone layer
-This is bad as the holes allow more UV radiation to reach the surface, resulting in eye cataracts and skin cancer with prolonger exposure
-Thankfully, many countries have now agreed to ban the use of CFCs, and the ozone layer is slowly healing

The carbon cycle

-The carbon cycle is the mechanism that maintains the level of carbon dioxide in the atmosphere
-For the atmosphere to maintain a constant amount of CO_2, the rate of removal of CO_2 must be equal to the rate of return of CO_2

Removal of CO_2	Return of CO_2
-Oceans and other large bodies of water absorb CO_2 as some marine plants use it to photosynthesise, and it is also conveyed to carbonic acid and $CaCO_3$ that make up the shells of marine organisms which accumulate on the seabed when they die, causing the CO_2 to	-Living things gain energy for life processes by breaking down glucose into CO_2 and water, in the equation $C_6H_{12}O_6\ (aq) + 6O_2\ (g) \rightarrow 6CO_2\ (g) + 6H_2O\ (l)$, which is then released back into

remain in the ocean -Plants take in CO_2 and water to make glucose and O_2 in the presence of sunlight in the equation $6CO_2 (g) + 6H_2O (l) \rightarrow C_6H_{12}O_6 (s) + 6O_2 (g)$	the atmosphere -Most fossil fuels contain mostly CH_4, which is burnt in the presence of O_2 to produce energy, producing CO_2 as a byproduct in the reaction $CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O (g)$ -CO_2 is also released when organic matter decays and when bacteria decomposes dead organic matter

Global warming

-Global warming is the increase in the average temperature of the earth's surface due to increasing amounts of greenhouse gases in the atmosphere
-An increase in greenhouse gas emissions results in an accelerated greenhouse effect
-Greenhouse gases are mainly CO_2 and CH_4

Greenhouse effect

-Solar radiation reaches the eart's atmosphere, where some of it is reflected into space and some passes through the atmosphere and becomes heat
-As the earth's surface is warmed, some of the heat rises and escapes into space, while the remaining heat is trapped in the atmosphere by greenhouse gases, further warming the planet
-This is known as the greenhouse effect
-A certain level of greenhouse effect is essential for life, as without it earth would be too cold to live in
-However, too strong of a greenhouse effect would result in surface temperatures being too hot to sustain life
-Activities like the burning of fossil fuels and large scale cutting down of forests are causing some greenhouse gases, especially CO_2, to build up in the atmosphere
-This means that CO_2 is being added to the atmosphere at a higher rate than it is being removed, resulting in global warming
-Cattle is a significant source of CH_4, which a greenhouse gas 80 times as strong as CO_2

Climate change

-Climate change is the result of global warming
-This is because the climate in an area is influenced by factors such as ocean temperature, wind speed and direction, which can all be disturbed by the temperature changes of global warming

Climate change phenomena	Effects
Change in rainfall patterns	-Shifting wind patterns can lead to changes in rainfall patterns

	-Lush areas can receive too little rain and turn into deserts, while other regions may receive too much rain and flood more -This can result in fertile land becoming barren and us no longer being able to produce enough food
Heat waves	-Heat waves would become more common, resulting in some parts of the world being too hot for humans to live in -Wildfires can also become more common, destroying entire ecosystems
Tropical storms	-Typhoons, hurricanes and cyclones are all forms of tropical storms -They form over warm water and move westwards, bringing strong winds, sheets of rain and an ultra high tide known as a strom surge -Hence, the warmer the oceans, the stronger and more frequent these storms become -Tropical storms are very destructive, endangering human lives, destroying property and damagning farmland
Ocean warming and acidification	-Marine organisms can be very sensitive to water temperatures, with coral reefs, a key marine habitat, being a good example, bleaching and dying due to rising sea temperatures -This greatly reduces marine biodiversity as habitats are lost -Populations of fish species, many of which humans depend on for food, are also severely cut
Ocean acidification	-Warmer oceans absorb CO2 more quickly, forming carbonic acid which decreases the pH of the water -Crustaceans such as crabs, corals, oysters and plankton are negatively impacted as they rely on carbonate minerals to form their shells -However, acidified water removes these carbonates through acid-carbonate reactions, while also dissolving existing shells

	-This makes these organisms more vulnerable to injury and predation, or just killing them directly
Glacial retreat and the melting of polar ice caps	-About 70% of the world's freshwater is locked up as ice in polar ice caps -A rise in global temperatures have caused these glaciers and polar ice caps to melt and shrink -Many communities rely on a steady stream of melt-water from glaciers for drinking and irrigation -Hence, their lives will be greatly impacted if the glaciers deplete faster than they can be replenished -Sea levels would also start to rise due to the melting of these ice caps, making coastal cities more prone to being permanently flooded -Adding a large amount of freshwater to the ocean can also disrupt ocean currents, which redistribute heat around the globe to stabilise global climate patterns