

AQA AS level Unit 2 Chemistry notes by Luna Hu

Energetics

Enthalpy: measures the total heat content of a system under constant pressure

Standard enthalpy change: change in enthalpy at constant pressure and temperature. (298 K and 100kPa)

Standard state: State of substance under standard conditions.

Bond dissociation energy: Energy required to break a covalent bond with all species at gaseous state.

Mean bond enthalpy: Average value of the bond dissociation enthalpy for a type of covalent bond taken from a range of different compounds.

Standard enthalpy of reaction: Enthalpy change when 1 mole of a substance is transformed by a chemical reaction with all reactants and products at standard states, under standard conditions.

Standard enthalpy of formation: Enthalpy required to form 1 mole of a substance from its constituent elements with all reactants and products in standard states under standard conditions

Standard enthalpy of combustion: Enthalpy change when 1 mole of a substance is completely combusted in oxygen with all reactants and products in standard states under standard conditions.

The enthalpies of all elements in their standard states are taken as 0. (By definition)

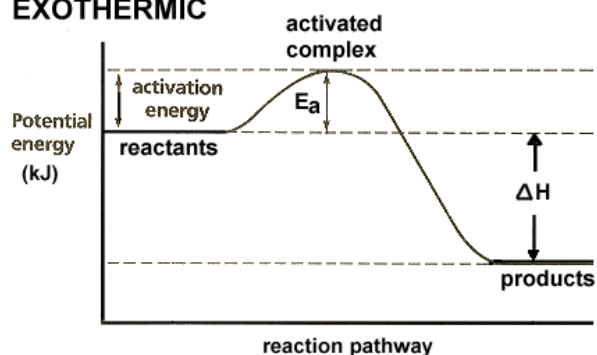
Hess' law: The enthalpy change of a reaction is the same regardless of the route taken from reactants to products.

* Always include state symbols

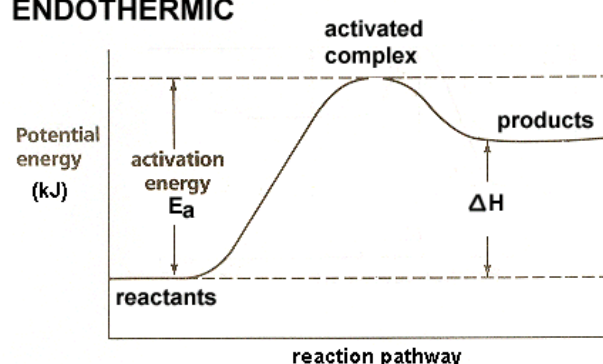
Endothermic: When the amount of energy required to break the bonds is greater than the amount of energy generated in the making of bonds. E.g. Photosynthesis, decomposition

Exothermic: When the amount of energy required to break the bonds is less than the amount of energy generated in the making of bonds. E.g. Combustion, neutralization

EXOTHERMIC



ENDOTHERMIC



Exothermic products are more stable as it has lower energy
Remember to draw single-headed arrows!! NOT DOUBLE HEADED

It is important to measure the enthalpy because:

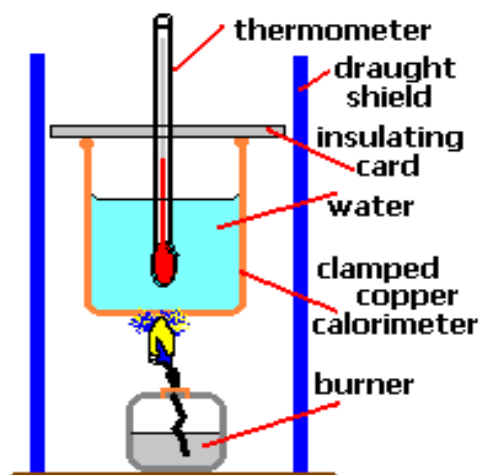
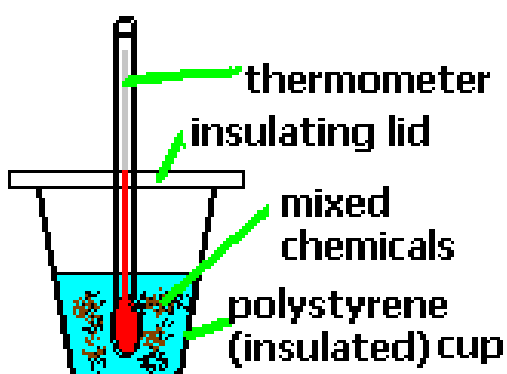
- We can measure the energy of fuels
- Calculate energy requirements
- Predict whether or not a reaction will take place

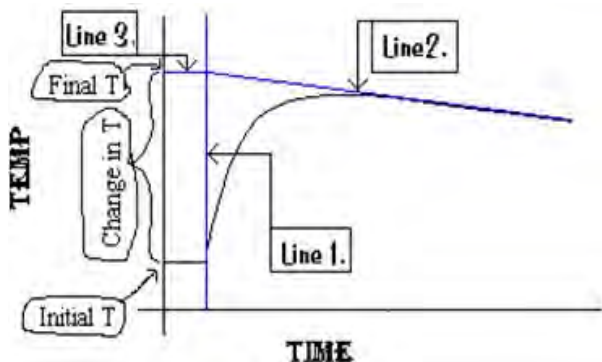
Measuring the enthalpy change of reaction:

- Mass of substance BEING heated
- Temperature change of substance being heated
- Specific heat capacity of the substance (The specific heat is the amount of heat required to change a unit mass of a substance by one degree in temperature.)

Use polystyrene cups (except for combustion) because they are good insulators so less heat is lost to surroundings.
Combustion is less accurate because more heat is lost to surroundings.

A SIMPLE CALORIMETER

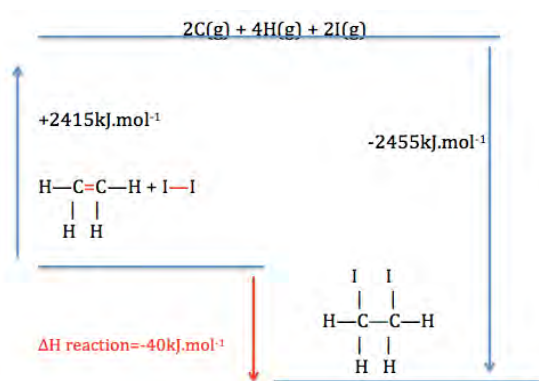




Intersection of 2 blue lines shows the extrapolated but more accurate ΔT , because some heat is still lost to the surroundings immediately when the reactants are added together.

Characteristics of good fuel:

- High calorific content
- Energy dense
- Easy to store, transport, and cheap



Thermochemical cycles can also be used to represent the enthalpy change of a reaction. 0 is represented by the constituent elements. These diagrams clearly show and compare the differences in enthalpies of reactants and products.

Kinetics

In order for a reaction to occur: particles must collide with $E > E_a$ (activation energy) and at the correct orientation

Rate of reaction: The change in concentration of a substance over a specified amount of time

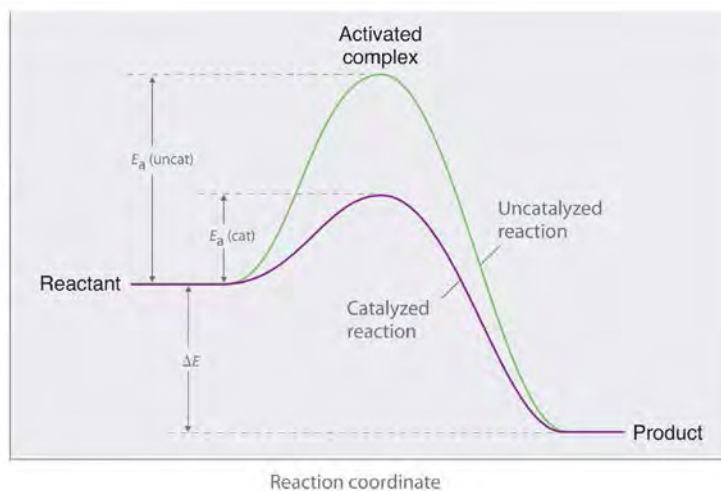
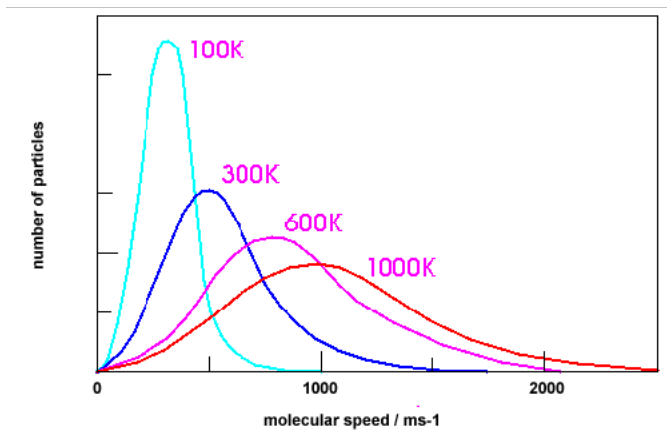
To increase rate of reaction:

- Increase the temperature: Increases the speed of the molecules as well as their energy, meaning the frequency of collision is increased and the proportion of successful collisions per unit time is increased. (+10k \approx double rate of reaction)
- Increase concentration (solution): Increases the number of particles per unit volume, thus the frequency of collisions per unit time increases, though the proportion of successful collisions remains the same.
- Increase pressure (gas): Increases the number of gas particles per unit volume, thus the frequency of collisions per unit time increases, though the proportion of successful collisions remains the same. (for the direction of the reaction of higher # moles of gas to smaller.)
- Increase surface area (solid): Increases the number of particles exposed/able to be reacted, thus the frequency of collision per unit time increases, but the proportion of successful collisions remains unchanged.
- Use a catalyst: lowers the activation energy thus the number of particles with $E > E_a$ increases, thus the number of successful collisions increases.

Represented on a graph:

- + Temperature: steeper, levels off the same.
- + Concentration: steeper, eventually levels off higher (extent depends on concentration)
- + Pressure: Steeper, levels off the same.
- + Surface area: steeper, levels off the same
- + Catalyst: steeper, levels off the same

- Why graph levels off & why gradient decreases: over time, number of reactants/volume decreases, collisions/time decreases, rate of reaction decreases. (eventually the limiting reactant is used up)



- No particle have 0 energy
- The peak is the most probable energy
- Slightly right of the peak is the mean energy
- There is no upper limit to the curve
- Area under the curve remains the same as the number of particles remains constant
- For a reaction to happen, the $E > E_a$.
- A small increase in temperature will result in a much larger increase in the number of particles with $E > E_a$.
- Number of collisions in a given time increases as the temperature increases, but this is not as important to the rate of reaction as the increase in the number of effective collisions.

Activation energy is the minimum energy a particle/molecule requires for a reaction to occur.
 - Molecules pass on energy as they collide. (Some lose, some gain)

Catalysts: Increases rate of reaction without getting used up.

- Provides an alternative route of lower activation energy.

Catalysts don't affect the enthalpy change of a reaction; they only lower the activation energy of forward reaction.

(Nor do they change position of equilibrium)

Types of catalysts:

Heterogeneous- Catalyst is in a different phase to reactants

Homogeneous- Catalyst is in the same phase as reactants

Different phases are separated by distinct boundaries: e.g. water and oil, solid and solution.

How solid catalysts actually work:

Adsorption: reactants form weak bonds with catalyst – to hold them in place to be reacted

Desorption: products break away from metal atoms so more can be reacted on catalyst surface.

Equilibrium – can be reached from either direction but can only be reached in a closed system – evinced when the macroscopic properties of the system do not change with time.

Dynamic equilibrium: When the rate of the forward and backward reaction are equal, and the concentration of reactants and products remain constant in a closed system.

Shifting the position of the equilibrium by changing the conditions can alter the proportion of the reactants and products.

La Châtelier's principle: If a system at equilibrium is disturbed, the equilibrium will shift in the direction that counters the disruption. (But it can't tell how far the equilibrium shifts)

To change conditions of an equilibrium reaction:

- Changing concentrations: Equilibrium will shift to the direction that reduces the concentration of that reactant/product.
- Changing overall pressure: (only works in gas phase) Equilibrium will shift to the direction countering the increase/decrease in pressure
- Changing temperature: Equilibrium will shift to the direction to counter increase/decrease in temperature.
- Catalyst: doesn't shift equilibrium as it increases the rate of forward and backward reaction equally.

Applications:

Haber process: $\text{N}_2 (\text{g}) + 3\text{H}_2 (\text{g}) \rightleftharpoons 2\text{NH}_3 (\text{g}) \Delta H^* = -92\text{KJ mol}^{-1}$

Conditions: 400°C, 20000kPa, iron catalyst

(Important liquid fuel/chemical feedstock)

Ethanol: (hydration of ethene) $\text{H}_2\text{C}=\text{CH}_2 (\text{g}) + \text{H}_2\text{O} (\text{g}) \rightleftharpoons \text{CH}_3\text{CH}_2\text{OH} (\text{g}) \Delta H^* = -46\text{KJ mol}^{-1}$

Conditions: 300°C, 6500 kPa

- High pressure (too high) will polymerize the ethane
- Too much steam will dilute catalyst

Methanol: $\text{CO} (\text{g}) + 2\text{H}_2 (\text{g}) \rightleftharpoons \text{CH}_3\text{OH} (\text{g}) \Delta H^* = -91 \text{KJmol}^{-1}$

Conditions: 500K, 10000kPa

Temperature: Costs money to reach and maintain high temperature

Pressure: costs money to reach and maintain high pressure. Equipment! Explosion!

^ Cost may outweigh benefit.

A compromise temperature: high enough to ensure rapid reaction, low enough for decent yield. (Because high temperature results in increased rate of reaction but decreases yield; vice versa)

Redox: reduction-oxidation

Oxidation: gain of oxygen, loss of electron(s)

Reduction: loss of oxygen, gain of electron(s), reduction in charge, gain in hydrogen

Oxidizing agent: accepts electrons (and is reduced)

Reducing agent: donates electrons (and is oxidized)

(ALWAYS REMEMBER TO IDENTIFY THE SPECIES NOT JUST THE SINGLE ION!)

Oxidation state:

Shows distribution of electrons in a compound/molecule.

Charge of element/ion/atom in a compound.

Ionic: number of electrons lost or gained.

Element	Oxidation state
Hydrogen	+1 (except metal hydrides)
G1	+1
G2	+2
Aluminium	+3
Oxygen	-2 (except in peroxide and with F)
Chlorine	-1 (except with F or O)

*Every element has oxidation state of 0 when uncombined

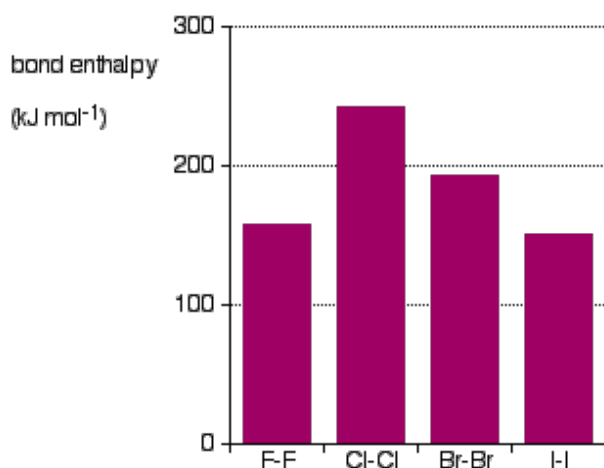
Half equation: an ionic equation used to describe the gain or loss of electrons during a redox process. 2 half equations combine to form balanced equation for redox reaction.

Simplest half equation: no spectator ions

Group 7

Halogen	Atomic number	Electronegativity	Atomic radius	Melting point	State
Fluorine	9	Highest			Pale yellow gas
Chlorine	17				Green gas
Bromine	35				Reddish liquid
Iodine	53		Largest	Highest	Grey solid (purple gas)

Bond enthalpies



Electronegativity decreases down group

- Increased shielding
- Increased distance
- Attraction between bonding electrons and nuclei is decreased.

Atomic radius:

- Increased number of main level shells
- Weaker attraction between e in shells
- More diffuse

Melting point:

- VDW forces
- * electron #, area, shape, shells

Bond enthalpy:

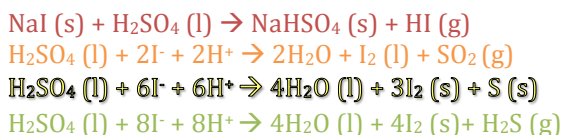
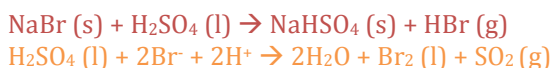
Fluorine is unexpectedly weak because it is very small, thus the non-bonding electrons are too close so they repel each other.

Oxidizing ability increases up the group. (Better at attracting electrons)

Evidenced in displacement reactions. (G7 in G1+G7 compound displaced by more reactive G7)

Reducing ability of halide ions increase down the group. Further away + more shielding so easier to lose electron.

All **solid** sodium halides can react with **CONC** sulfuric acid. Product reflects reducing powers of halide ions.



Neutralization: not redox.

Observe: white fumes (HCl) blue litmus turns red

Observe: Choking gas SO₂

Respective liquids/vaporized.

Observe: Yellow solid sulfur

Observe: rotten egg smell

Identification of group 7 halides:

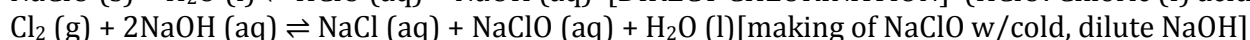
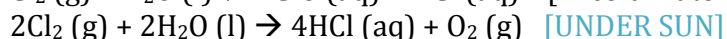
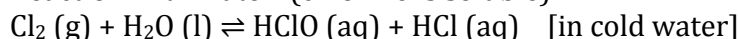
1. Add dilute nitric acid to get rid of CO_3^{2-} and OH^- impurities
2. Add silver nitrate solution into solutions
3. Precipitate forms

Halide	AgF	AgCl	AgBr	AgI
Color	colourless	white	cream	Pale yellow
Further test		Dissolves in dilute ammonia	Dissolves in conc ammonia	Doesn't redissolve.

Uses of chlorine: kill bacteria and sterilize water.

Disproportionation: oxidation state increases and decreases simultaneously.

Reaction with water: (chlorine is soluble)



NaClO (Sodium Chlorate (I)): Making bleach and water disinfection

Group 2 - alkaline earth metals

Element	Atomic number	Metallic radius	MP/BP	Ionization energy
Magnesium	12		v	Highest
Calcium	20		Highest	
Strontium	38		^	
Barium	56	Largest	^	

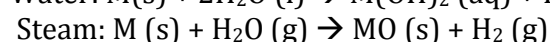
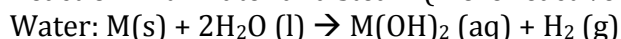
Size increases: extra filled main level of electrons

MP/BP decreases: sea of delocalized electrons further away from nuclei, weaker attraction.

[Exception: magnesium: lattice structure is different.]

Ionization energy decreases: more shielding and distance from nuclei thus easier to remove

Reaction with water and steam (more reactive down the group)



(M = metal)

Anion	Trend
Hydroxide	More soluble down
Carbonate	Less soluble down
Sulfate	Less soluble down

Uses of group 2 compounds:

*Calcium hydroxide used to treat acidic soil.

*Magnesium sulfate is used as a laxative.

*Barium sulphate is used as x-ray lining (your body doesn't absorb it since its very insoluble)

*Barium chloride is used to test for sulfates because it forms white precipitate with it.

Extraction of metals:

Metals combine with sulfur and oxygen

Roast sulfide in air → oxide

(SO₂ produced. Collected → sulfuric acid)

^ with water and oxygen

Gangue: clay and rock impurities

Sometimes temperature required for metal to react with Carbon is too high (uneconomic)

Ideal reducing agent: Cheap, readily available, don't cause pollution, no harmful byproduct.

Iron (haematite) [continuous]

$C + O_2 \rightarrow CO_2$ [Exothermic 2000k]

$CO_2 + C \rightarrow 2CO$ [reducing agent]

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

Byproduct: SiO₂

Recycling: reduces scrap in landfill

No CO₂ produced (energy used to melt produced CO₂)

Easily separable: magnetic

Manganese: reduction with coke (impure carbon) or carbon monoxide

Copper (malachite): reduction with coke (old method) or carbon monoxide

$CuCO_3 \rightarrow CuO + CO_2$ [Thermal decomposition]

$2CuO + C \rightarrow 2Cu + CO_2$

New method:

- Spray Cu low-grade ore with dilute acid in the presence of thiobacillus ferrooxidans.
- Cu extracted from solution with scrap iron
- + Scrap iron is cheap
- + No CO₂ produced & less energy use than reduction with carbon

Aluminium (bauxite) [continuous]

Al₂O₃ dissolve in molten cryolite: Na₃AlF₆ [melting point 1240 k] 1000 k less!

Pass through current of 300000A (VERY HIGH COST)

Cathode: $Al^{3+} + 3e^- \rightarrow Al$

Anode: $O^{2-} \rightarrow \frac{1}{2}O_2 + 2e^-$

^ @ high temp, O₂ reacts with carbon anode (CO₂)

Recycling: uses 5% of energy required to make them

+less CO₂

Increased transportation costs (CO₂ emitted in the process)

Titanium (Rutile) [batch]

Strong, low density, resistant to corrosion

$TiO_2 + 2Cl_2 + C \rightarrow TiCl_4 + CO_2$ [TiCl₄ distilled off. Temp: 1173 k]

$TiCl_4 + 4Na \rightarrow Ti + 4NaCl$ [inert argon atmosphere: stop metals reacting w/O&N; 1300k]

• Can't use carbon as reducing agent, as @ high temp, metal react w/Carbon → carbide (brittle)

EXPENSIVE:

- Using a more reactive metal as reducing agent
- Batch process (heat back up)
- 3 step process

Tungsten → filament lamps

• Can't use carbon as reducing agent as @ high temp, metal react w/Carbon → carbide (brittle)

$WO_3 + 3H_2 \rightarrow 3H_2O + W$

(H₂ highly flammable)

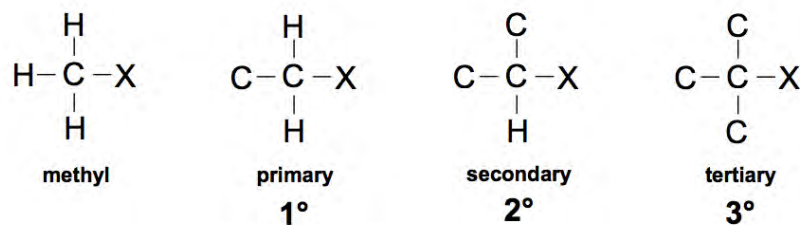
Haloalkanes have polar bonds with the carbon it is bonded to, as they are all more electronegative than carbon.

- Polarity decreases down the group as difference in electronegativity decreases.
- The polarity is not sufficient to make them soluble in water.
- Increased branching decreases melting/boiling point: (decreased surface area therefore less vdw)

The reactivity of the haloalkane depends on :

Bond enthalpy and polarity, but the effects of bond enthalpy is greater than that of polarity.

The stronger/higher the bond enthalpy, the less reactive/harder it is for the halogen to leave.



Tertiary Haloalkanes are more reactive than primary Haloalkanes

Nucleophilic substitution: to introduce new functional groups into organic compounds.

The carbon is electron deficient, it can be attacked by reagents that are electron rich or have electron rich areas.

1. The nucleophile provides the electrons for a new bond (using its lone pair)
2. Carbon can only have 8 outermost electrons thus the Halide ion is displaced (lowest bond enthalpy with C)

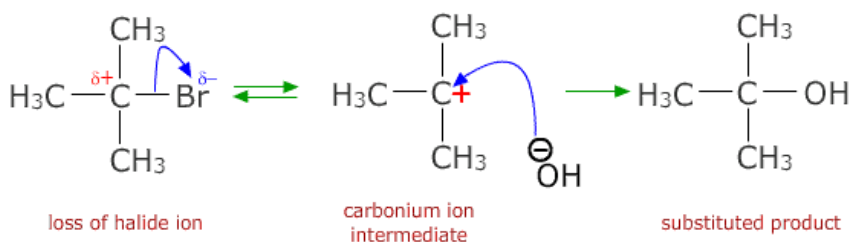
Nucleophile: electron pair donor

- Has a lone, unshared pair of electron on electronegative atom. (to form a covalent bond)
- OH⁻, CN⁻, and NH₃ (neutral)
- Bond between C - Halogen breaks (heterolytic fission) resulting in production of ions.

Substitution: Chemical changes which replace atoms/groups of atoms by another atom/groups of atoms.

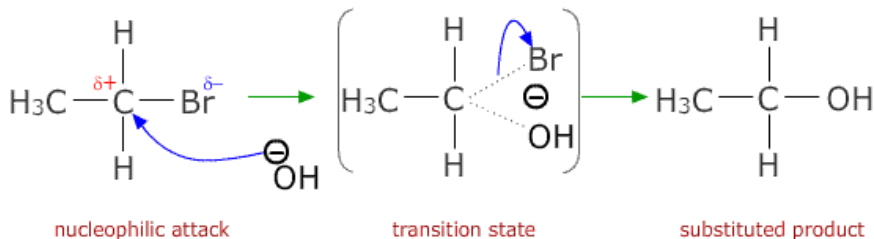
Hydrolysis: compound split in a reaction by water.

Unimolecular nucleophilic substitution - S_N1

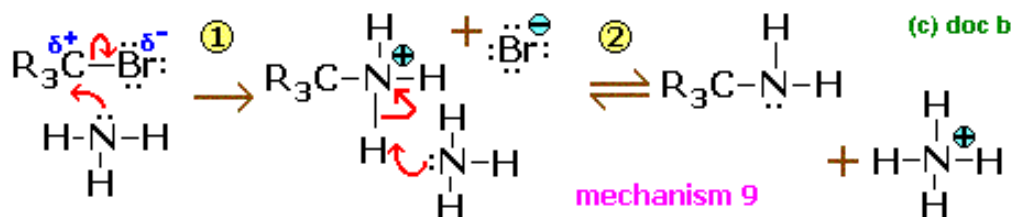


- S_N1 favored by tertiary Haloalkanes

Bimolecular nucleophilic substitution - S_N2



- S_N2 favored by primary Haloalkanes



Ammonia as nucleophile

Organic compound formed is a primary amine.

Methylamine,
Ethylamine,
Propylamine etc.

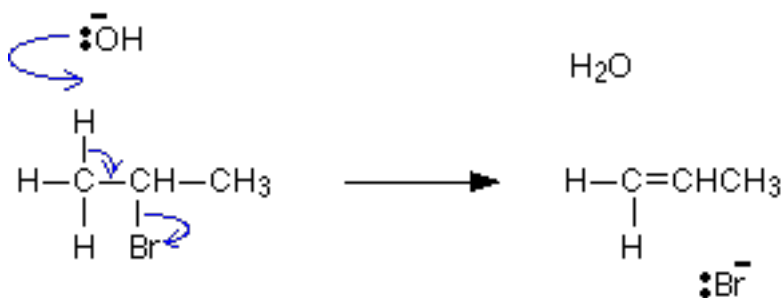
Note: with CN⁻, nitriles are formed.

Ethanenitrile, Propanenitrile

(the C in CN must be taken into account)

Elimination: Hydrogen halide is eliminated from the molecule, leaving a double bond. (Alkene) A.K.A. hydrolysis

Under different conditions (No water present, only G1 hydroxide dissolved in ethanol, and heated) the OH⁻ will act as a base, removing the H⁺ ion from the haloalkane.



OH⁻ removes H⁺ (dative covalent bond)
 C-H electron moves down - x2 bond =
 C-Br has lowest bond enthalpy hence it leaves.

Formation of Haloalkanes – Free radical substitution reaction (chain reaction)



INITIATION

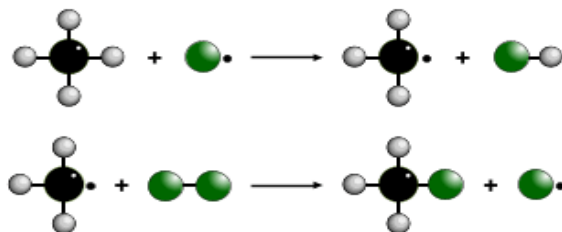
Free radicals formed



Initiation – The formation of free radicals (from non-free radicals)
 Initiated by UV light.
 eg. $\text{Cl}_2 \rightarrow \text{Cl}\cdot + \text{Cl}\cdot$

PROPAGATION

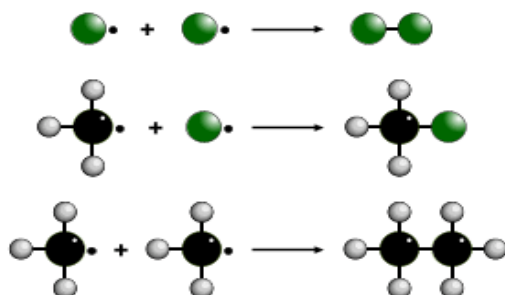
Free radicals used and re-generated



Propagation – ‘Recycling’ free radicals (free radical to free radical and back to initial free radical) 2 stages
 Eg. $\text{Cl}\cdot + \text{CH}_4 \rightarrow \text{HCl} + \cdot\text{CH}_3$
 $\cdot\text{CH}_3 + \text{Cl}_2 \rightarrow \text{CH}_3\text{Cl} + \cdot\text{Cl}$

TERMINATION

Free radicals removed



Termination – Free radicals removed (free radical with free radical → no radical)
 Eg. $\text{Cl}\cdot + \text{Cl}\cdot \rightarrow \text{Cl}_2$
 $\cdot\text{CH}_3 + \cdot\text{CH}_3 \rightarrow \text{C}_2\text{H}_6$
 $\text{Cl}\cdot + \text{CH}_3 \rightarrow \text{CH}_3\text{Cl}$

CFCs

- Very unreactive in normal conditions
- Goes high up in the atmosphere before decomposed by UV light.

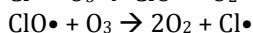
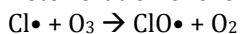
Uses:

- Short chain – Aerosol, refrigerants, propellants
- Long chain – **solvent**, dry cleaning

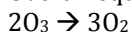
Ozone: shields earth from excess UV radiation which causes blindness and skin cancer in humans

Ozone at ground level causes lung irritation, degradation of plastics and paints.

Deterioration of ozone layer:



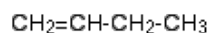
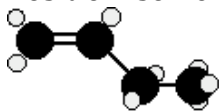
Overall equation:



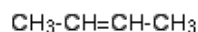
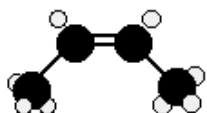
Alkenes

- Unsaturated hydrocarbons (1 or more c-c double bond) – decolorizes bromine water
- Planar molecule
- Restricted rotation (overlapping p orbital → π orbital)
- More reactive than alkanes as it's very susceptible to electrophilic attacks (the double bond)

Position isomers: (difference in position of double bond)

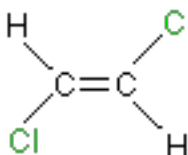


but-1-ene

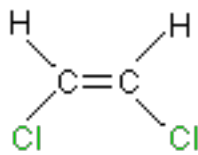


but-2-ene

Stereoisomerism: Same structural formula, bonds arranged differently in space (due to lack of free rotation about double bond)



trans-1,2-dichloroethene



cis-1,2-dichloroethene

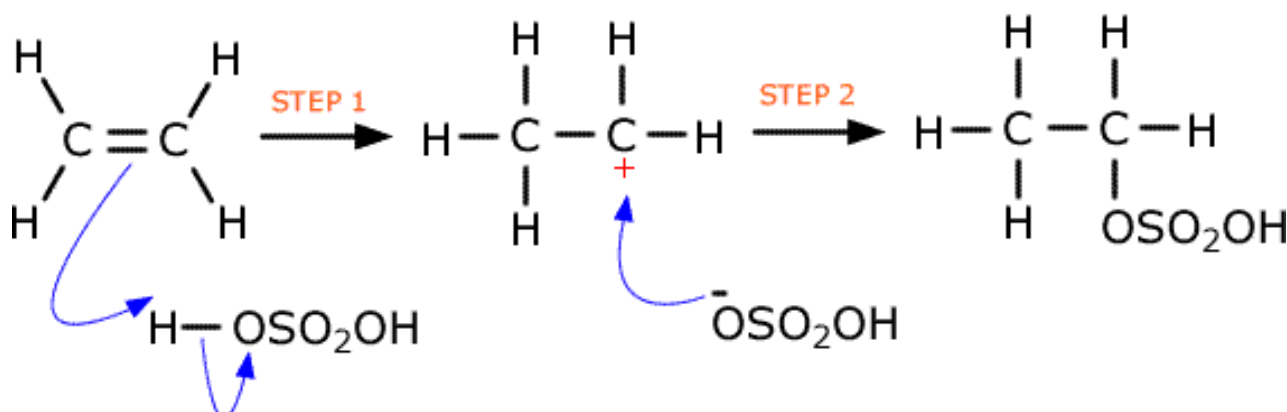
If the higher priority groups are on the same side: z isomer
If the higher priority groups are on the opposite side: e isomer

Determining higher priority:
Highest atomic number.

Electrophilic addition

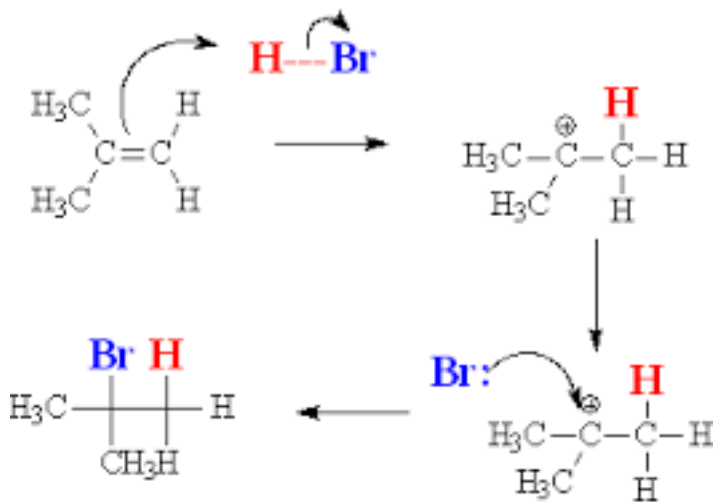
- the double bond is a high electron density area
- Electrophile is attracted to the double bond
 - Positively charged (area or ion)
 - Accepts pair of electrons
 - Carbocation formed
 - now negative ion forms a bond

Electrophilic addition of sulfuric acid to an alkene



1. δ^+ Hydrogen in H_2SO_4 is attracted to electron dense area (double bond)
2. Accepts an electron, forming a carbocation
3. Now negative ion attracted to + carbocation

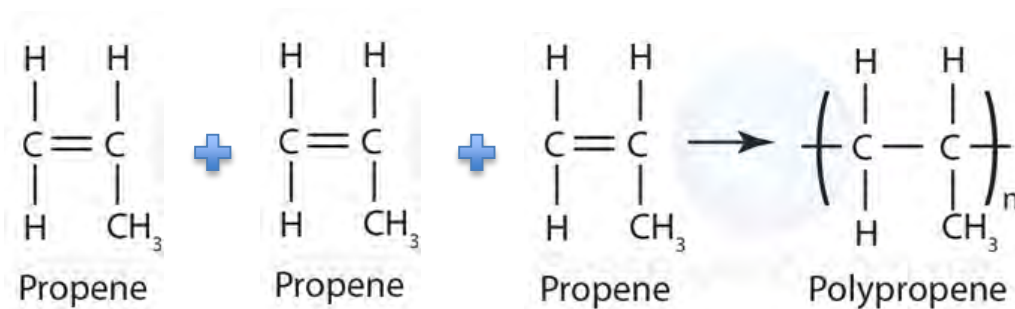
Note: the electrophile can be a neutral molecule eg. $\text{Br}-\text{Br}$. Negative charge area in double bond induces a temporary charge on both Br



Note: Tertiary carbocation is much more stable than primary as alkyl groups have a tendency to release electrons (positive inductive effect) which stabilizes the positive charge of the carbocation. Thus the more the C⁺ is attached to alkyl groups, the more likely it will be the main intermediate carbocation.

Polymerization of alkenes

Polymers: long chain molecules formed when many small molecules (monomers) join up in long chains



(Addition polymerization of propene) – not additional!

- Repeating unit: smallest group of atoms that produce the polymer. (basically the monomer)
- low density: free radical substitution (branching, don't pack together)

Polymer + name	Uses
Polyethene (CH ₂ =CH ₂)	Plastic bags, bottles
Polypropene (CH ₃ -CH=CH ₂)	Rope, milk crates
Polychloroethene (Cl-CH=CH ₂)	Vinyl records
Polypropenenitrile (CN-CH=CH ₂)	Clothing
Poly 1,1,2,2-tetrafluoroethene (CF ₂ =CF ₂)	cookware

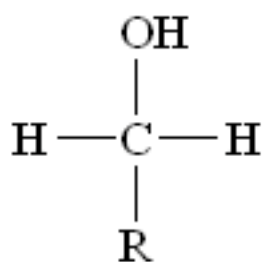
Plastic problems:

- not biodegradable
- Landfill sites
- Problem for animals when ingested

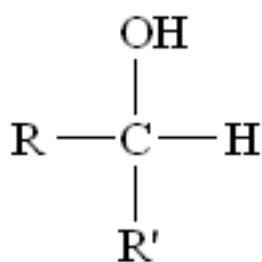
Solutions: reduce plastic usage or recycle

Mechanical recycling	Feedstock recycling
<ul style="list-style-type: none"> - separate different types of plastic - wash once - ground up into small pellets - Melted and remoulded 	<ul style="list-style-type: none"> - Heat plastic - Break polymer into monomer - make new plastics <p>- can only be done limited amount of times: plastic's properties degrades each time heated (chains break)</p>

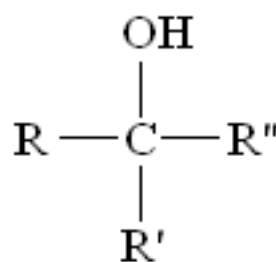
Alcohols $C_nH_{2n+1}OH$



primary (1°)
alcohol



secondary (2°)
alcohol



tertiary (3°)
alcohol

- OH = hydrogen bonding (thus higher mp bp than alkanes with similar RMM)
- Shorter chain alcohols soluble in water

Making alcohol:

- Fermentation

1. Anaerobic respiration: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$ [using YEAST]

2. compromise temperature : 35°C

- anaerobic to prevent oxidation

Biofuel: fuels produced from vegetable matter and organic waste

+ **carbon neutral:** an activity that has no net annual carbon (greenhouse gas) emissions to the atmosphere

+ idea of carbon sink; recycling CO_2

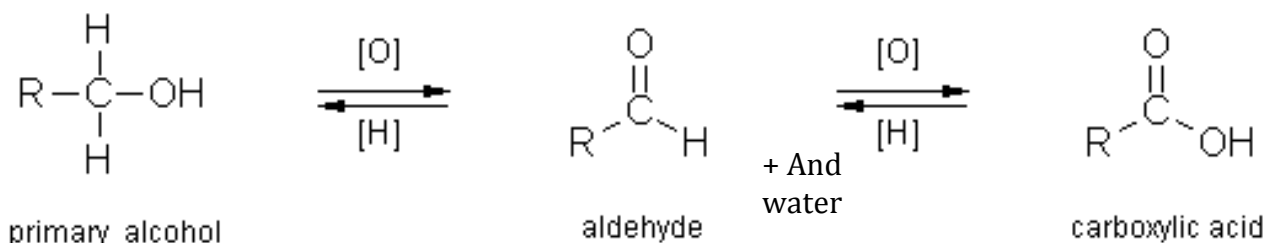
- Hydration of ethene (covered in previous chapter)

Comparing the 2 methods of manufacturing ethanol:

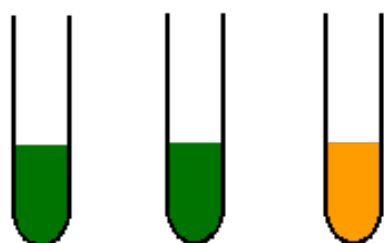
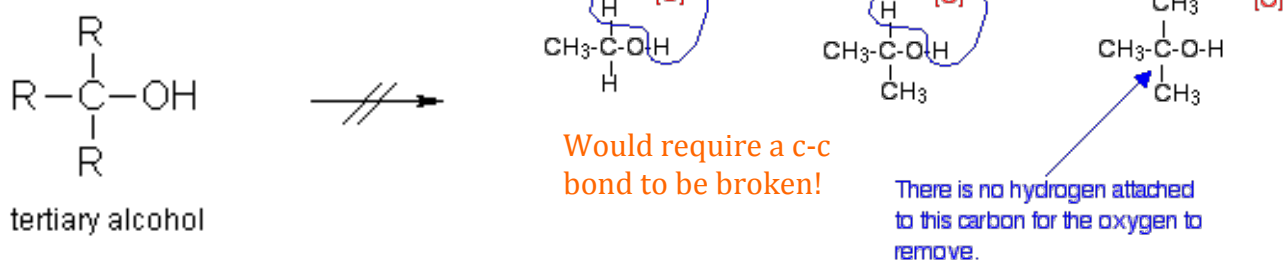
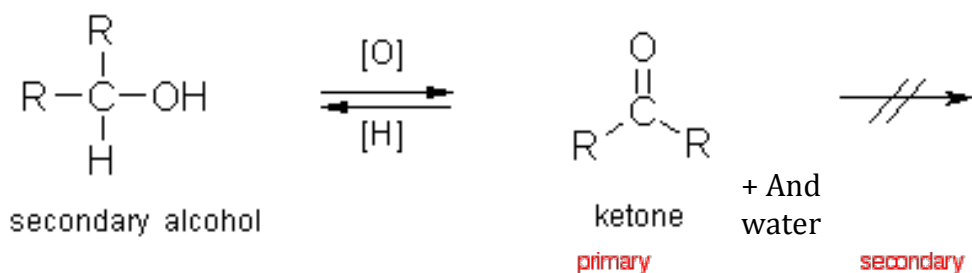
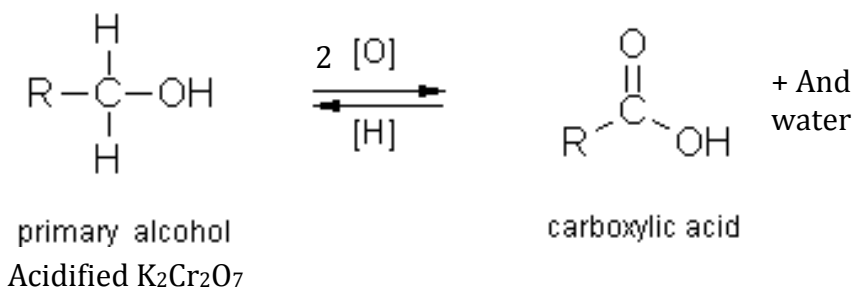
	Fermentation	Hydration
+	Renewable Carbon neutral Sugar canes cheap to grow Low temperature (35)	Continuous High purity, yield, atom economy, fast
-	Batch Low purity Low yield & slow Low atom economy Fermentation stops at 15% Labour intensive	Non-renewable High energy and pressure (energy use)

Oxidation of alcohols:

Excess alcohol, dilute acid, $K_2Cr_2O_7$, heat and distill



Excess [O], conc acid, $K_2Cr_2O_7$, heat and reflux



primary secondary tertiary

Alcohols are

- Neutral, free of water (anhydrous $CuSO_4$ or cobalt chloride)
- Reacts w/solid phosphorous (V) chloride \rightarrow HCL fumes

\leftarrow Add potassium dichromate acidified with sulfuric acid

- warm in water bath

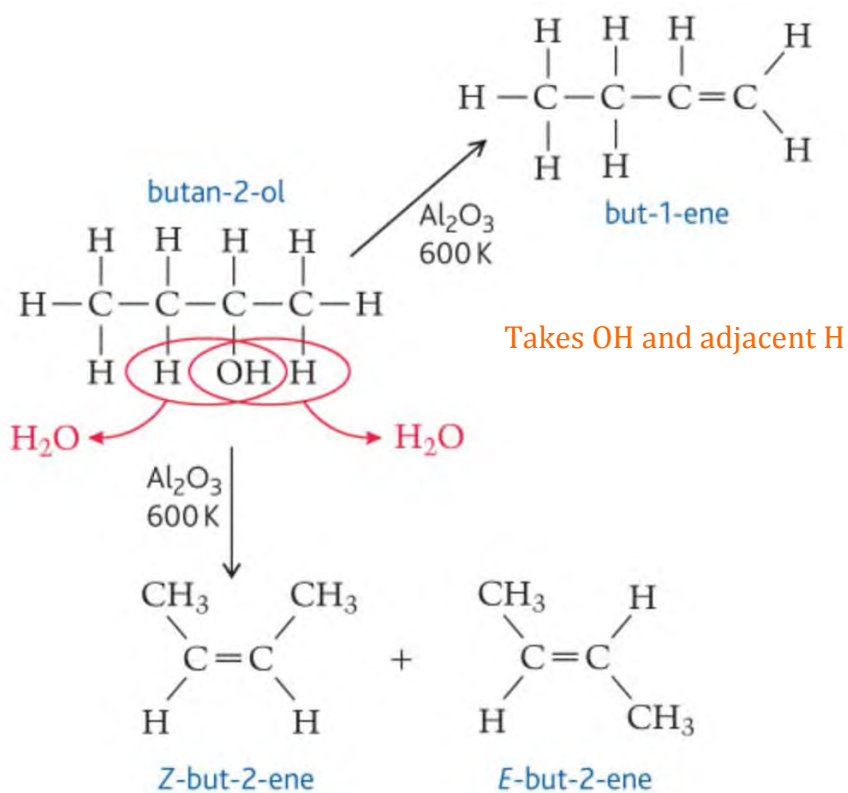
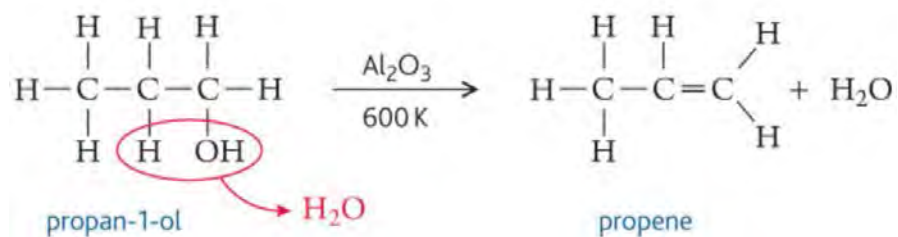
\leftarrow that happens (1° and 2° turns green, 3° stays orange)

Test for carboxylic acid: add metal carbonate \rightarrow CO_2 bubbles

	Tollen's ($AgNO_3 + NH_3$)	Fehling's/benedict's ($Cu(II) +$ warming)
Aldehyde	Mirror	Blue \rightarrow red ppt
ketone	No Δ	No Δ

Elimination reaction: When a small molecule leaves parent molecule

Dehydration of alcohols: pass through hot concentrated sulfuric acid or Al_2O_3



Analytical techniques

Infrared spectroscopy: Identify functional groups in organic molecules

- Most compounds absorb infrared radiation
- Wavelengths absorbed corresponds to natural frequency for vibration of bonds
- IR re-emitted from sample of decreased intensity (transmittance)

Fingerprint region (below 1500cm^{-1})

- complex vibrations of whole molecule
- unique to particular molecule

(identify isomers by superimposing with IR spectrums in database)

Greenhouse gases have large IR absorption ranges.

EXAM TECHNIQUES

General:

- **FORMS** acid rain
- **Making** plastic
- **Making** bleach
- State effects (rather than just naming)
- Don't just write p/d orbital: specify shell number

Why is a process used:

- See product (pollution? Harmful?)
 - See reactant (Harmful? Expensive? Explosive? Recycling?)
 - Time and energy considerations (conditions)
 - Atom economy
-
- Identifying species: the whole thing, not just the ion!
 - Simplest ionic equation: write out whole equation and rid spectator ions.
 - Mechanisms: use curly arrow.

Calculations:

- Look back at equation to check (moles)
- Enthalpy of formation/reaction/combustion: divide by # moles!
 - Do plenty of past papers
 - Read the syllabus
 - Mark papers carefully, use the mark scheme to streamline answers
 - Make note of mistakes and review them
 - Read examiner reports if necessary

Good luck and don't leave revision to the last minute!