1. General

Experimental Design:

Gas volumes measured using a graduated syringe
Liquid volumes measured using a (i) burette - 2d.p.
(ii) pipette - 1d.p.
(iii) measuring cylinder - various precisions
Use an digital mass balance to measure mass

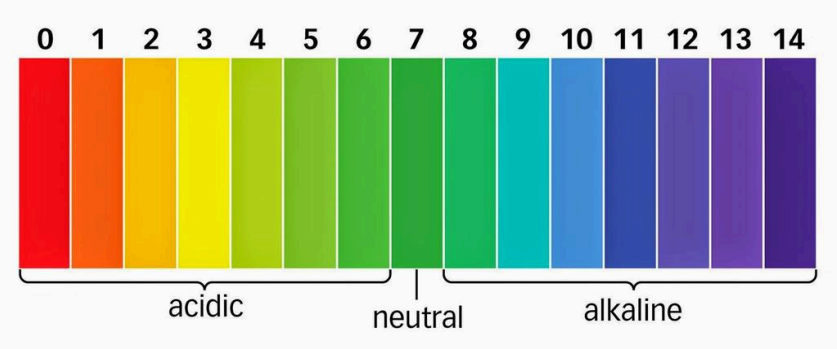
Name of gas	**Solubility in water**	**Density relative to air**	**Collection method**
Hydrogen	Not soluble	Less dense	Displacement of water
Carbon dioxide	Slightly soluble	Denser	Displacement of water
Oxygen	Slightly soluble	Slightly denser	Displacement of water
Chlorine	Soluble	Denser	Downward delivery
Ammonia	Very soluble	Less dense	Upward delivery
Sulfur dioxide	Very soluble	Denser	Downward delivery

Drying agents

- Concentrated sulfuric acid (not for alkaline gases - will react)
- Calcium oxide (not suitable for acidic gases - will react)
- Anhydrous calcium chloride

Anhydrous copper (II) sulfate is used as an indicator for water vapour
It turns from white to blue when hydrated - $CuSO_4 \rightarrow CuSO_4 \cdot 5H_2O$

Universal Indicator



Gas Tests

Gas	**Test**	**Result if present**
Hydrogen	Insert lighted splint into test tube	Lighted splint extinguishes
Oxygen	Insert glowing splint into test tube	Splint relights
Carbon dioxide	Bubble gas through limewater (aq calcium hydroxide)	White ppt (calcium carbonate) forms
Ammonia (NH_3)	Insert moist red litmus paper	Moist red litmus paper turns blue, pungent smell
Chlorine	Insert moist blue litmus paper	Most blue litmus paper turns red then bleaches
Sulfur dioxide	Bubble through aq acidified potassium manganate (VII)	Aq acidified potassium manganate (VII) turns from purple to colourless

Purification Techniques:

If it's solid + solid dissolved in water (solution),

1. Filter mixture and collect residue/filtrate
2. A. Wash residue with deionised water and pat dry
   B. Heat filtrate in an evaporating dish until all the water evaporates, collect remaining solid

   If solid dissolved in water is thermally unstable, do crystallisation instead:
1. Heat solution in evaporating dish, stop heating when there is about ⅓ of solution left
2. Let solution cool to allow crystals to form
3. Filter the mixture to obtain crystals, then pat dry

To collect the solvent from a solution (where solute and solvent have different boiling points) - use simple distillation

To separate a mixture of two miscible liquids, which have different boiling points - use fractional distillation
As vapours go up the fractionating column, they condense when they reach a part of the column that is below their boiling point
Liquid with the lower boiling point is collected as distillate
Fractionating beads increase the surface area for the condensation of vapour, making separation of vapours more efficient

To separate two immiscible liquids - separating funnel
To separate a solid from a solid that sublimes easily under heat - sublimation

2. Bonding and Structure

$^{A}_{Z}X$

A is the nucleon number, Z is the atomic number

Isotopes are atoms with the same number of protons but different number of neutrons

Relative molecular mass represents the average mass of one atom of an element taking into account the different isotopes and their relative abundances (Ar)

Ar = (Mr of isotope) x (abundance in %) + (Mr of other isotope) x (abundance in %)...

Chemical Bonding

All chemical bonds are electrostatic forces of attraction

	Simple molecular structure	**Giant covalent structure**	**Giant ionic structure**	**Giant metallic structure**
Particles in solid	Molecules	Molecules	Cations and Anions	Positive metal ions and sea of delocalised electrons
Bonds between particles	Weak intermolecular forces of attraction	Strong covalent bonds	Strong ionic bonds	Strong metallic bonds
Physical state in room conditions	Aqueous	Solid	Solid	Solid except mercury
Bpt/Mpt	Low	High	High	High
Example	Water, Oxygen, Carbon Dioxide	Diamond	Sodium Chloride	Iron

Physical Properties: mpt, bpt, conductivity and solubility

Ionic Bonds

- Usually formed when a metal reacts with a non-metal
- Soluble in water - the charge of water molecules are attracted to the charges of ions in giant ionic structure and pulls the ions away from their regular arrangement
- Can conduct electricity in solid/aq state due to presence of free-moving ions

Covalent Bonds

- Forces pulling together molecules of non-metals

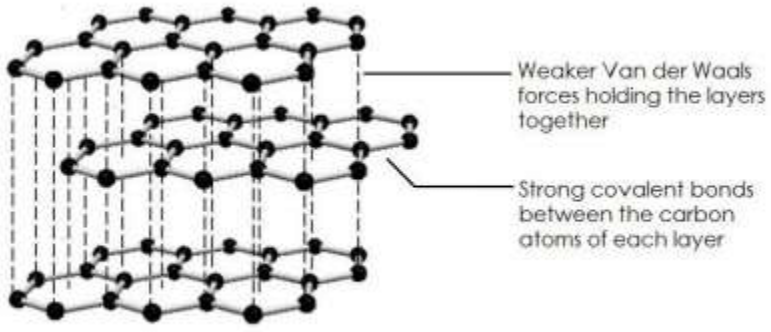
- Giant Covalent: Insoluble in almost all solvents, poor conductors of electricity except graphite (ONLY MENTION GIANT COVALENT FOR GRAPHITE OR DIAMOND)
- Simple Molecular: Insoluble unless polar, do not conduct electricity (e.g ammonia, CO2, water)
- The greater the difference in electronegativities between the 2 bonded atoms in a covalent bond, the greater the polarity of the bond
- The more polar the molecules in a substance, the more soluble it is in water (e.g NH3)

Metallic Bonds

- 1, 2 or 3 valence electrons
- Insoluble in water, metal bonds do not dissolve in water
- Metallic bonds are between positive metal ions and sea of delocalised electrons
- Malleable and ductile: layers of atoms can easily slide against each other
- Good conductors of electricity due to movement of delocalised electrons

Graphite vs Diamond

Diamond: Strong covalent bonds holding carbon atoms together in a giant covalent structure
Graphite: Strong covalent bonds holding carbon atoms together in layers, WEAK Van der Waals forces (intermolecular forces of attraction) in between layers.
Graphite has a giant covalent structure where each carbon atom is covalently bonded to 3 other carbons in a hexagonal ring and has a 4th electron that is delocalised within each layer - conduct electricity.



3. Formulae and Mole Concept

Amount = Mole Conc. x Vol
Mass (g) = Mass Conc. x Vol
Mass (g) = Amount x Molar Mass (Mr)

Number of Particles = Amount x (6×10^{23})
Gas @ rtp: Vol = N x 24 dm^3
For gases - Mole ratio = Vol ratio

Percentage Composition by Mass = $\frac{mass\ of\ element}{total\ mass\ of\ compound} \times 100\%$

Percentage yield = $\frac{actual\ yield\ (usually\ given\ in\ question)}{theoretical\ yield\ (usually\ calculated)} \times 100\%$

Percentage purity = $\frac{mass\ of\ pure\ substance\ in\ sample}{mass\ of\ sample} \times 100\%$

Empirical Formula - Simplest ratio of atoms
Molecular formula - Actual number of atoms

Molecular formula = Empirical formula x N

	H	O	…
Assume 100g of _____, mass of element present/g			
Amount/mol (divide by Mr)			
Divide by smallest no. of moles			
Simplest ratio			

The reactant completely used up is the limiting reactant - it determines the amount of products formed
The reactant with more than the required quantity is the excess reactant

4. Acids, Bases and Salts

Reactions
Acid dissociates in water to form H+ ions
Base dissociates in water to form OH- ions

Acid + Base → Salt + Water
Acid + Carbonate → Salt + Water + Carbon Dioxide
Acid + Metal → Salt + Hydrogen
Base + Ammonium Salt (NH4+) → Salt + Water + Ammonia gas (with warming)
Alkali + (some) salt solutions → Salt + Insoluble Hydroxide

Strong acids fully dissociates in water, weak acids partially dissociates
$CH3COOH \rightleftharpoons CH3COO^- + H^+$

HCL - monobasic, 1 mol of HCL dissociates to form 1 mol of H+
H2SO4 - dibasic, 1 mol of H2SO4 dissociates to form 2 mol of H+

pH = -lg[H+] in mol dm^{-3}

Equivalence point of titration graph - mid-point of the vertical part, when the amounts of acid and base exactly react with each other

Types of Oxides

Metallic oxide	Basic oxide	Na2O, CaO, CuO	Reacts with acids
	Amphoteric oxide	ZAP (ZAP hydroxides too)	Reacts with BOTH acids and bases
Non-metallic oxide	Acidic oxide	CO2, SO2, NO2, SiO2	Reacts with bases
	Neutral oxide	H2O, CO, NO	Do not react with acids or bases

Solubility Rules

1. All Group 1, ammonium or nitrate salts are soluble
2. All sulfates are soluble except Barium, Calcium & Lead
3. All chlorides and iodides are soluble except Silver and Lead
4. All carbonates are insoluble except 1
5. All oxides and hydroxides are insoluble except 1

Salt Preparation

Precipitation - Two solutions make insoluble salt (aq + aq → s + aq)

1. Mix 2 aqueous, one with the cation and the other with the anion that make up the salt
2. Filter the mixture
3. Wash the residue with distilled/deionised water and pat dry

Titration - Two solutions make a soluble salt (solution) (aq + aq → aq + H2O (H2O made from a hydroxide)) (USUALLY TO MAKE A GROUP 1/AMMONIUM SALT)

1. Titrate the corresponding acid and base/carbonate with a suitable indicator. Note the end-point titre
2. Repeat titration without indicator
3. Heat salt solution obtained until about half its volume to saturate it
4. Leave it to cool
5. Filter the salt crystals that appear
6. Pat dry the crystals

Excess base - One solution and one insoluble base (solid in excess) make a soluble salt (solution) (s + aq → aq + l/g)

1. Add excess base/carbonate into acid, stirring and heating until no more can dissolve
2. Filter off excess base
3. Heat salt solution obtained until half its volume to saturate it
4. Leave it to cool
5. Filter the salt crystals that appear
6. Pat dry the crystals

5. Redox

Oxidation	**Reduction**
Loss of electron	Gain of electron
Increase of oxidation state	Decrease of oxidation state
Gain of oxygen	Loss of oxygen
Lost of hydrogen	Gain of hydrogen

Oxidation half-eq: $2I^- \rightarrow I_2 + 2e^-$
Reduction half-eq: $Cl_2 + 2e^- \rightarrow 2Cl^-$

Uncombined elements have an oxidation state of 0
Hydrogen is +1
Oxygen is -2
Group 17 is -1
Sum of oxidation states of all the atoms in a polyatomic ion (such as SO_4^{2-}) is equal to its charge (-2)
(O oxidation state of -2, S oxidation state of +6)

Explaining Redox

1. _____ loses electrons to form _____, so _____ is oxidised.
   _____gains electrons to form _____, so _____ is reduced

2. The oxidation state of [element] increased from _____ in A to _____in B, so A is oxidised.
   The oxidation state of [element] decreased from _____ in C to _____in D, so D is reduced.

3. _____ gains oxygen (or loses hydrogen) to become _____, so _____ is oxidised.
   _____ loses oxygen (or gains hydrogen) to become_____ , so _____ is reduced

Agents
Acidified aqueous potassium (VII) manganate $KMnO_4$ to test for REDUCING agents:
Purple, MnO_4^- (aq) → Colourless, Mn^{2+} (aq)

Potassium iodide solution KI to test for OXIDISING agents
Colourless, I^- (aq) → Brown, I_2 (aq)

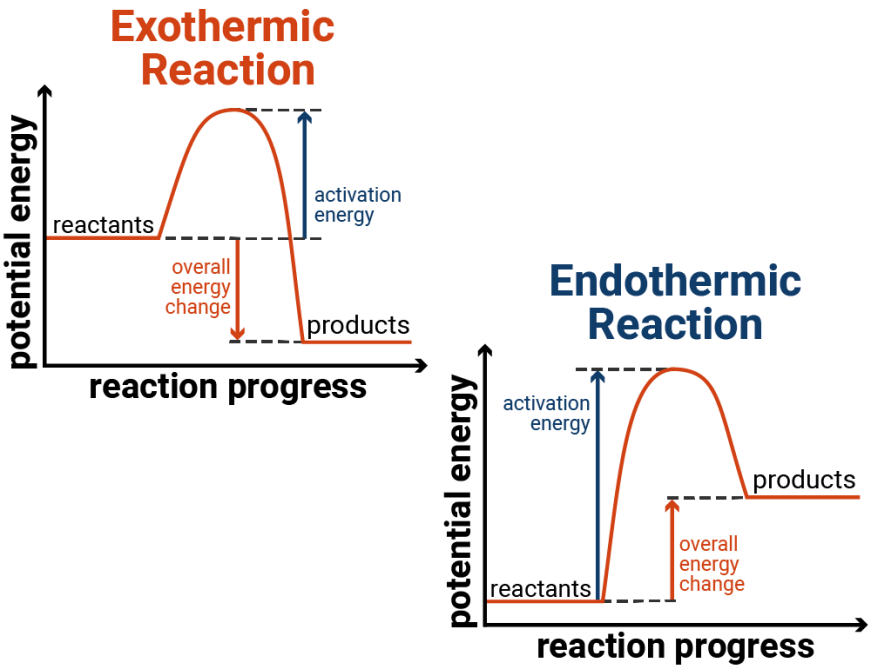
6. Energetics, Kinetics and Equilibrium

Energetics

Exothermic Process	**Endothermic Process**
Heat is released into surroundings	Heat is absorbed from surroundings
Enthalpy change value is negative	Enthalpy change value is positive
Bond making	Bond breaking
Energy content of products < energy content of reactants	Energy content of products > energy content of reactants
E.g combustion, neutralisation, respiration, freezing	E.g photosynthesis, thermal decomposition, melting, electrolysis

Calculation:
ΔH = + (energy absorbed for bond breaking) - (energy released by bond making)

*state what bonds are broken/made



Reaction Kinetics

Activation energy is the minimum amount of energy that reacting particles must possess in order to undergo the reaction. The smaller the Ea, the faster the reaction.

Increased temperature: This increases the number of reactant particles having the activation energy, causing particles to have greater kinetic energy and collide harder and more frequently

Increased concentration: This causes more particles per unit volume and thus more frequent effective collisions between reactant particles
Increased concentration would result in an increased amount of products, since the initial amount of reactant is also increased with its concentration increased

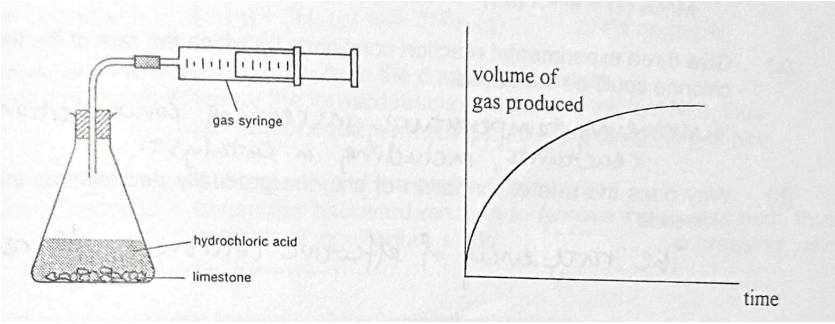
Increased pressure (for gases): This decreases the average distance between reactant particles thus more frequent effective collisions

Increased surface area (decreasing size of solid): This causes more frequent effective collisions between reactant particles

Catalyst: Provides an alternative pathway of lower activation energy, so more reactant particles will have activation energy and there will be more frequent effective collisions between them

A catalyst is not used up by the reaction; it may initially take part in the reaction but will finally be regenerated before the reaction ends

If one of the products is a gas then the usual method of determining the speed of reaction is shown below:



Equilibrium

When the speed of a forward reaction is equal to the speed of a backwards reaction, a state of equilibrium is attained and the equilibrium mixture always contains both the reactants and the products. The reaction does not stop when equilibrium is attained; the amount of reactants and products are just constant

Le Chatelier's Principle states that if a system at equilibrium is subjected to a disturbance, the equilibrium's response is to counteract the disturbance to minimise its effect.

The Haber process, carried out at 200 atm, 450°C, using iron as a catalyst, is a reversible reaction:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$ ΔH is negative

a) Disturbance - Remove some NH_3 from the equilibrium mixture
   System's response - favour the forward reaction to replace the lost NH_3, position of equilibrium shifts right, increasing the yield of NH_3

b) Disturbance - Temperature is increased
   System's response - favour the backward endothermic reaction to absorb the excess heat from surroundings, position of equilibrium shifts left, decreasing the temperature of the equilibrium mixture

c) Disturbance - Increase pressure of equilibrium mixture
   System's response - favour the forward reaction to decrease the total number of molecules to relieve the pressure, position of equilibrium shifts right, increasing yield of NH_3

d) A catalyst does not affect the position of equilibrium or yield as it speeds up both the forward and the backward reaction and helps the reaction reach equilibrium earlier

AvT graph:
If only one of the reactants suddenly spikes, it is the addition/removal of that reactant (concentration)
For gases, if all go through a spike, it is an increase/decrease in pressure or a increase/decrease in volume
No spike, reactants increase and decrease, it is a change in temperature

RoRvT graph:
Forward reaction decreases at a decreasing rate while backward reaction increases at an decreasing rate to meet at a straight line in the middle - equilibrium

7. Periodic Table

Across a period,
Atomic radius decreases, as proton number increases, but the number of shielding electrons remains constant. Valence electrons are pulled closer to the nucleus, making the atom smaller.

Group 1

- Low densities; densities increases down the group
- Melting point decreases down the group
- Soft and silvery
- Reactivity increases down the group as atoms are getting larger and the valence electron is getting further from the nucleus and is more easily lost

Group 17

- 7 valence electrons, made up of diatomic molecules
- Elements become darker down the group
- Reactivity decreases down the group
- The more reactive halogen will displace the less reactive halogen from a solution of its ions
- Melting point increases down the group as molecules are getting larger and have stronger intermolecular forces of attraction which require more energy to break

Group 18

- 8 valence electrons
- Chemically unreactive, provides an inert atmosphere

Transition Metals

- High densities and melting points
- Colourful solid compounds and aqueous solutions
- Variable oxidation states in their compounds (e.g. Iron (II) chloride and Iron (III) chloride)
- Good catalysts (e.g Iron in Haber process)

8. QA

cation	effect of aq sodium hydroxide	effect of aq ammonia
Copper (II) Cu2+	Blue ppt, insoluble in excess	Blue ppt, dissolves in excess to form dark blue solution
Iron (II) Fe2+	Green ppt, insoluble in excess	Green ppt, insoluble in excess
Iron (III) Fe3+	Red-brown ppt, insoluble in excess	Red-brown ppt, insoluble in excess
Ammonium NH4+	Ammonia gas evolves upon heating	-
Calcium Ca2+	White ppt, insoluble in excess	No ppt
Zinc Zn2+	White ppt, soluble in excess to form colourless solution	White ppt, soluble in excess to form colourless solution
Lead Pb2+	White ppt, soluble in excess (think ZAP)	White ppt, insoluble in excess
Aluminium Al3+	White ppt, soluble in excess (think ZAP)	White ppt, insoluble in excess

In order to differentiate from Pb2+ and Al3+, add KI or KCl - will produce a yellow and white ppt respectively with Pb2+ but not with Al3+

anion	test	observations
CO3^2-	Add dilute HCL	Effervescence (CO2)
Cl-	Acidify with HNO3 and add silver nitrate (AgNO3)	White ppt
I-	Acidify with HNO3 and add silver nitrate (AgNO3)	Yellow ppt
NO3^2-	Add aqueous sodium hydroxide (NaOH), aluminium (catalyst) and heat gently	Gas (ammonia) evolved turns moist red litmus paper blue
SO4^2-	Acidify with HNO3 and add barium nitrate (Ba(NO3)2)	White ppt

Why acidify with HNO3 - to remove any interfering ions that can form a precipitate with the added reagent

9. Metals and Electrolysis

Metals

metal	water	steam	dilute acid	extraction by carbon	extraction by hydrogen	thermal stability of carbonate
K	✓	✓	✓			Carbonate not decomposed by heat
Na	✓	✓	✓			
Ca	✓	✓	✓			Carbonate increasingly easier to be decomposed by heat to form CO_2 and the oxide $CuCO_3 (s) \rightarrow CuO (s) + CO_2 (g)$
Mg	✓	✓	✓			
Al		✓	✓			
Zn		✓	✓	✓		
Fe		✓	✓	✓	✓	
Sn		✓	✓	✓	✓	
Pb			✓	✓	✓	
Cu				✓	✓	
Ag				✓	✓	Carbonate decomposed by heat to give metal, CO_2 and O_2 $Ag_2CO_3 (s) \rightarrow Ag_2O (s) + CO_2 (g)$ $2\ Ag_2O (s) \rightarrow 4Ag (s) + O_2 (g)$
Au				✓	✓	

Carbon is between aluminium and zinc in the reactivity series
Hydrogen is between lead and copper in the reactivity series

Thermal stability is opposite of reactivity series

Reaction with water: the **hydroxide** is formed (e.g $2Na + 2H_2O \rightarrow 2$**NaOH** $+ H_2$)
Reaction with steam: the **oxide** is formed (e.g $Mg + H_2O (g) \rightarrow MgO + H_2$)

Extraction by carbon: also can use CO (e.g $Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$ or $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$)
Extraction by hydrogen: $PbO + H_2 \rightarrow Pb + H_2O$ (hydrogen is highly flammable, not use on large scale)

Most metals are extracted from their oxides (ores) by reduction with carbon or carbon monoxide. Those that cannot be extracted through this method are usually extracted by electrolysis.

Displacement reaction: more reactive metal replace the ion or compound of a less reactive metal

Alloying:

- Metals are soft and malleable due to the regular layers of atoms that are able to slide over each other easily.
- Alloying mixes other metals that have atoms of different sizes, so they are unable to form regular layers and cannot slide over each other
- Alloys are therefore harder and stronger
- Brittle too: When a force is focussed on a disruption in these regular layers, alloys can easily break

- Iron is alloyed with copper to make steel
- Mild steel (~0.25% carbon) used to make car bodies, bridges, railway lines
- Hard steel (~1% carbon) is harder and more brittle, used to make tools
- Stainless steel (other additives besides carbon) is used to make cutlery and surgical instruments

Corrosion of Metal:
The corrosion (or rusting) of metals (specifically iron) requires oxygen and water

Rusting can be prevented by:

1. Surface protection: Coating metal with a barrier to block air and water from reaching the metal underneath (paint or oil or plastic or a metal such as tin or chromium)

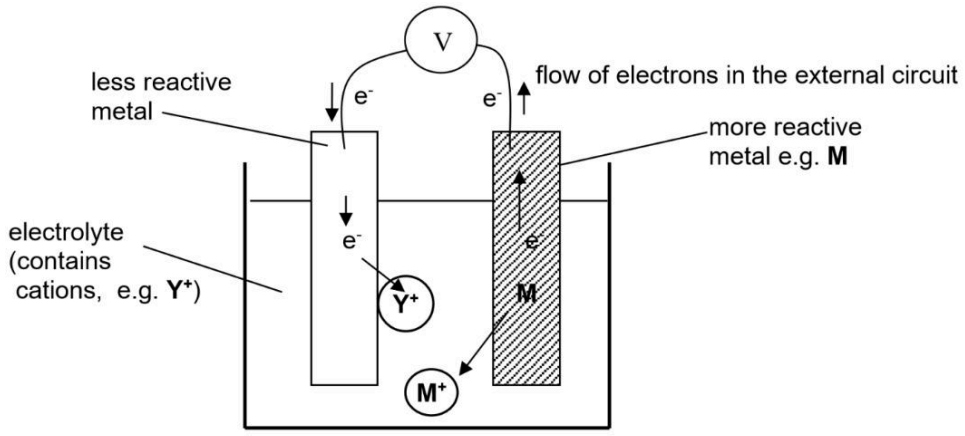
2. Sacrificial protection: Coat/have metal in contact with a more reactive metal. The more reactive metal reacts with oxygen and water in place of the less reactive metal, thus protecting it. Even if the layer of more reactive metal is scratched to expose the less reactive metal underneath, the less reactive metal is still protected (until all of the more reactive metal is corroded off)

3. Alloying: In stainless steel, the iron is mixed with carbon and copper and nickel. The chromium and nickel prevents the iron from reacting with water and oxygen

Aluminium is covered by a layer of non-corroding, non-porous aluminium oxide that protects and prevents the metal from reacting, so it does not corrode easily

Electric Cell vs Electrolytic Cell

Electric Cell:



The two electrodes are 2 different metals
An electrolyte contains free, mobile ions
Electrons flow from the more reactive metal to the less reactive metal, and are then 'captured' by the cations in the electrolyte
The greater the difference in reactivity of the metals, the greater the voltage produced

The more reactive metal oxidises and loses electrons and moves from M to the other electrode. The M+ ions produced due to the oxidation of M will enter the electrolyte, so M becomes smaller as the reaction progresses
M → M+ + e-

The less reactive metal receives the electrons and passes them to the electrolyte. The cation in the electrolyte, Y+, accepts the electrons to undergo reduction. A substance is produced at the less reactive metal, dependent on the identity of the cation Y+ present in the electrolyte
Y+ + e- → Y

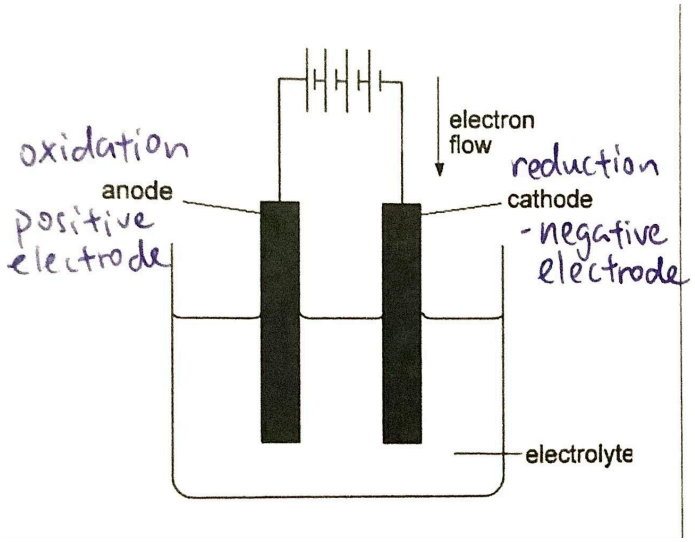
Overall algebraic eq: M + Y+ → M+ + Y

Clarifications:
In electric cells, **CRAO**

Anode - negative electrode: it is the **more reactive metal** and electrons are generated continuously through **oxidation**, so this is the source of negative charge and called the **negative electrode**.

Cathode - **positive** electrode - **reduction** occurs here - **less reactive** metal

Electrolytic Setup:



When an electric current is passed into an electrolyte, the free ions lose their random movement. Electrolysis is an ENDOTHERMIC reaction

CRN APO:

The positive cations will be attracted to the **negative electrode (cathode)**; **reduction** occurs at the cathode where cations gain electrons and are discharged, forming products

The electrode is connected to the negative terminal of the power source and receive electrons from it, so it is negative.

The negative anions will be attracted to the **positive electrode (anode)**; **oxidation** occurs at the anode where anions lose electrons and are discharged, forming products

If an electrolyte is molten, it contains only 1 type of cations and anions that will be reduced and oxidised respectively. If an electrolyte is aqueous, it would also contain H+ and OH- ions, and the ions discharged will be dependent on the factors affecting electrolysis.

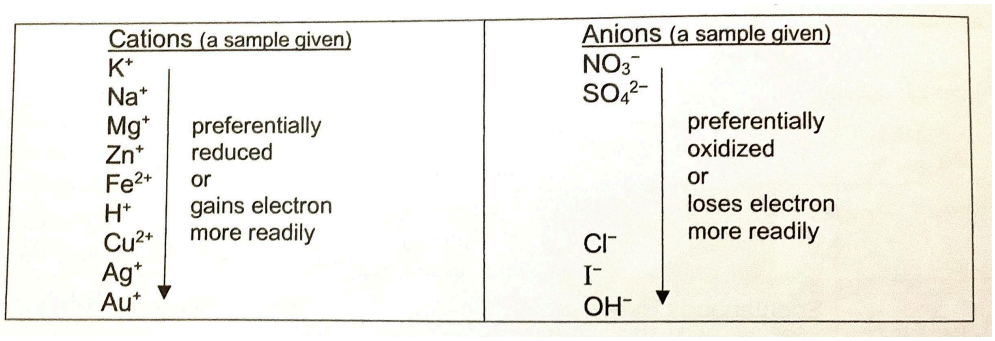
<u>IMPORTANT EQUATIONS</u>

Oxidation of OH-: $4OH^- \rightarrow 2H_2O + O_2 + 4e^-$
Reduction of H+: $H^+ + 2e^- \rightarrow H_2$

For 4 moles of electrons that pass through the circuit, 1 mole of oxygen and 2 moles of hydrogen gas are formed - therefore volume ratio of oxygen to hydrogen produced is theoretically **2:1**, however it is not exactly so as some oxygen dissolves back into the electrolyte while volume of hydrogen gas remains constant (O_2 is slightly soluble in water while H_2 is insoluble)

Factors affecting electrolysis:

1. Electro-chemical series: which ion is "**preferentially oxidised/reduced** and discharged" - phrase like this



2. Concentration effect - **affects between halides and hydroxide ions - for cations it does affect but only those ions in near vicinity, not between ions where the distance between them in the electrochemical series is too great** (e.g between H+ and Na+)

   Dilute - OH- > Cl-/I-
   Concentrated - Cl-/I- > OH-

3. Nature of electrode
   Inert electrodes are graphite and platinum; they do not interfere with the electrolysis

   Active electrodes are other metals, it **takes part in the reaction** and is **preferentially oxidised** over OH- and other anions, so the active anode eventually becomes smaller and dissolves away (also check if cathode is preferentially reduced over the cations)

Industrial applications of electrolysis:

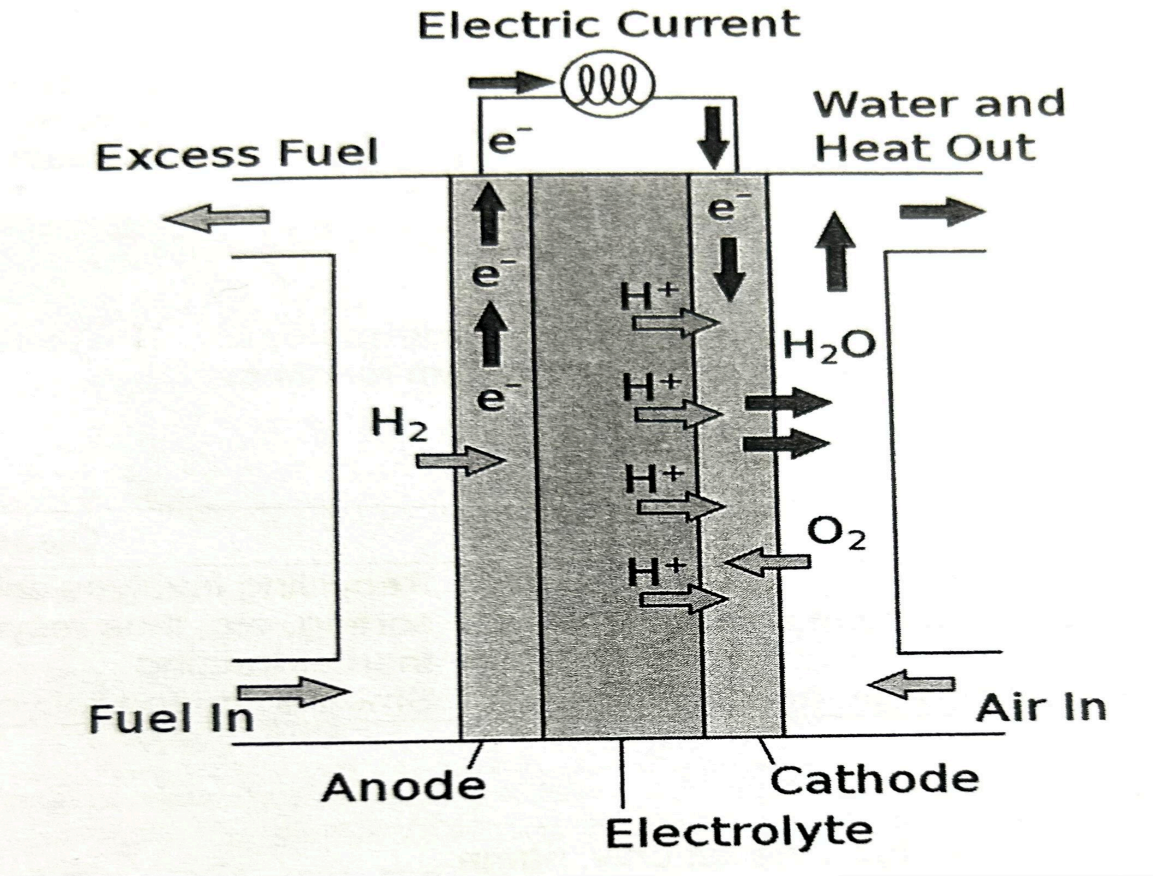
1. Electroplating - e.g in copper plating, the **object to be plated is the cathode** and the anode is the metal (copper), and the electrolyte is copper (II) sulfate solution. If object to be plated is non-conductive, object is first covered with a layer of graphite.

   During the plating process, the copper anode dissolved, forming Cu2+ and replenishing the Cu2+ used up for plating. Since **for every 1 mol of Cu2+ reduced to Cu at the cathode another 1 mol of Cu is oxidases to Cu2+ at the anode**, the blue colour of the solution remains the same as **concentration of Cu2+ is constant**.

2. Purification of copper - To get 100% pure copper, **impure copper is made the anode** and a small piece of pure copper the cathode, and the electrolyte is a copper (II) sulfate solution. The anode becomes smaller and the cathode larger as electrolysis progresses.

Hydrogen Fuel Cell

Hydrogen (as a fuel) burns cleanly - only producing water, and it also produces more energy per gram when burnt.



1. Hydrogen enters through an inlet. A catalyst at the anode speeds up the oxidation of hydrogen, producing electrons: $H_2 \rightarrow 2H^+ + 2e^-$
2. The electrons run in the external wire to go do work, and returns back to the cell, entering the current
3. At the cathode, the electrons react with oxygen in the air, as well as the H^+ ions that have diffused from the anode to the cathode: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
4. **OVERALL REACTION: $2H_2 + O_2 \rightarrow 2H_2O$ + electricity**
5. Here, the anode is the negative terminal, where electrons are running out of the cell

10. Organic Chemistry

Fractional Distillation of Crude Oil:

1. Crude oil is heated in a furnace so that it evaporates
2. The vapours are passed up the fractionating column, which is cool at the top and hot at the bottom, and they begin to cool down and condense
3. Smaller hydrocarbons with lower boiling points are collected at the top column
4. Larger hydrocarbons with higher boiling points are collected at the bottom column

Cracking of hydrocarbons: (condition: high amounts of heat, catalyst AlO3 and SiO2(Silicon (V) oxide))

- Breaking larger hydrocarbons into smaller molecules of alkanes, alkenes and/or hydrogen
- The mixture of products is separated through fractional distillation
- Cracking is used to:
    - Change heavy fractions (lubricating oil) in low demand to lighter fractions (petrol) in high demand
    - Make alkenes (especially ethene)
    - To make hydrogen gas

Homologous series: A family of organic compounds with the same general formula and similar chemical properties

Members in the same homologous series:

1. Same functional group, general formula, similar chemical properties
   Functional group: An atom or a group of atoms that give a molecule its characteristic properties

2. Has a gradual change in their physical properties as the **number of carbon increases**: **Mp/Bp increases, density increases, viscosity increases, flammability decreases**
   Hydrocarbons are simple molecules held by intermolecular forces, as size of molecules increase, strength of intermolecular forces increase, mpt/bpt increase

3. Differs from the next by a -CH2- unit

	Alkanes	Alkenes	Alcohols	Carboxylic Acids
General Formula	CnH(2n+2)	CnH2n	CnH(2n+1)OH	CnH(2n+1)COOH (where n = 0 for the first member)
Functional Group	C-C	C=C (alkene functional group)	-O-H (alcohol functional group)	-COOH (carboxylic acid functional group)

Naming convention: 1 Carbon: Meth-
2 Carbon: Eth-
3 Carbon: Prop-

4 Carbon: But-

Isomers are compounds with the same molecular formula but different structural formula

Reactions

Alkanes:

- Saturated hydrocarbons
- Mostly gases and liquids at rtp (simple molecular structure), first 4 members are gases

1. Complete Combustion: Alkane + O_2 → CO_2 + H_2O
   Incomplete Combustion: Alkane + O_2 → CO + H_2O

   Algebraic Eq for Complete Combustion: $C_xH_y + (x + y/4)O_2 \rightarrow xCO_2 + (y/2)H_2O$

2. Substitution with Chlorine: CH_4 (g) + Cl_2 (g) → CH_3Cl (g) + HCl (g) (if one hydrogen atom is replaced) methane + chlorine → chloromethane + hydrochloric acid
   Condition: UV Light

   Reaction happens 1 by 1, H_2 will not be produced from the rxn. If two hydrogen atoms are replaced, the product is dichloromethane. If three are replaced, the product is trichloromethane. If four are replaced, the product is tetrachloromethane.

Alkenes:

- Unsaturated hydrocarbons - presence of C=C double bond
- Alkenes more reactive than inert alkanes due to carbon-carbon double bond
- Mostly gases and liquids at rtp (simple molecular structure), first 4 members are gases

1. Combustion - similar to Alkanes
   a. Produces more soot as greater percentage of carbon by mass

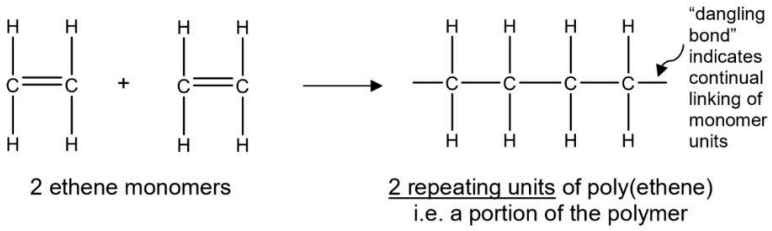
2. Addition reactions:
   a. Addition of hydrogen (hydrogenation): alkene + H_2 → alkane
      Condition: nickel catalyst, heat
      To convert vegetable oil (alkene) into margarine (alkane)

   b. Addition of bromine (bromination): alkene + Br_2 → 1,2-dibromo(alkane)
      TEST for alkene (unsaturation)! Brown aqueous bromine will turn colourless in the presence of alkenes

   c. Addition of steam (hydration): alkene + H_2O (g) → alcohol
      Condition: H_3PO_4 phosphoric acid catalyst, high temp (300°C), high pressure (60 atm)
      Water is in the form of steam

3. Addition Polymerisation
   - When there is only one type of product formed through polymerisation
   - Many small alkene molecules can link together in a repeating manner to form a polymer - the C=C double bond breaks, allowing the monomers to join

- Monomer is unsaturated, polymer is saturated



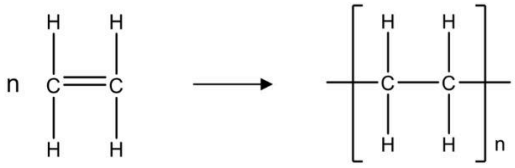
2 ethene monomers

2 repeating units of poly(ethene)
i.e. a portion of the polymer

Uses of Poly(ethene): Plastic (plastic bag, plastic bottles)

Naming convention: Monomer - ethene, Polymer - Poly(ethene)
Monomer - propene, Polymer - Poly(propene)

The **equation for the polymerisation** of ethene to produce poly(ethene) is:



n ethene monomers

poly(ethene)

Alcohols:

- All alcohols are colourless liquids at room conditions, first four members are soluble in water

1. Preparation of Ethanol
    a. Preparation from ethene: $C_2H_4 + H_2O \rightarrow C_2H_5OH$ - a mixture of ethanol and steam are passed over heat and a catalyst (phosphoric acid)

    b. Preparation by **fermentation**: $C_6H_{12}O_6 \rightarrow 2\ C_2H_5OH + 2\ CO_2$
    glucose → 2 ethanol + 2 carbon dioxide
    Condition: 37°C, yeast, absence of oxygen - fermentation is an anaerobic process + to prevent the oxidation of ethanol into ethanoic acid

    Disadvantages of fermentation:
    - Slow
    - Fermentation gives a low yield as a high concentration of ethanol kills the yeast
    - Need to carry out fractional distillation to obtain pure ethanol from reaction mixture

    Uses of ethanol: solvent for organic compounds, alcoholic drinks (ethanol) and fuel

2. Combustion: alcohol + $O_2 \rightarrow CO_2 + H_2O$
3. Dehydration: alcohol → alkene + H_2O
    Condition: dehydrating agent - concentrated sulfuric acid

4. Oxidation: alcohol + 2[O] → carboxylic acid + H_2O
    Condition: heat, oxidising agent - acidified aq potassium (VII) manganate ($KMnO_4$) - the 2[O] represents oxygen that comes from the oxidising agent

KMnO4 can be used to **test for alcohol** - turns from purple to colourless as it is reduced

Oxygen in the air can also oxidise alcohols - beer and wine left exposed in the air becomes sour due to the formation of ethanoic acid

5. Esterification: see carboxylic acids section

Carboxylic Acids:

- Vinegar is a solution of ethanoic acid and water, ethanoic acid is also known as acetic acid
- Acids with two carboxyl groups have names ending with -dioic
- First 4 members are liquids at room conditions. They partially dissociate in water and are weak acids

1. Preparation of Ethanoic Acid:
    a. From ethanol: see oxidation
    b. From methane: methane is used to manufacture methanol and carbon monoxide - and these two reacted together in the presence of a catalyst forms ethanoic acid

    Uses: manufacture esters which are used as solvents, insecticides, drugs

2. Acidic Properties: reacts with base, some metals and carbonates
    Test for carboxylic acids - react with carbonate since effervescence (carbon dioxide) is observed

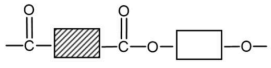
3. Esterification: alcohol + carboxylic acid ⇌ ester + H2O
    Condition: concentrated sulfuric acid, heat

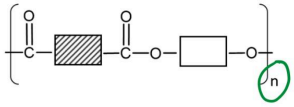
    Uses of esters (colourless liquid with a sweet smell): perfumes, flavouring

    Examples: ethanoic acid + propanol ⇌ propyl ethanoate + water ($CH_3COOH + C_3H_7OH \rightleftharpoons CH_3COOC_3H_7 + H_2O$) - ester bond R - COO - R (including C=O and C-O)
    This is a reversible reaction, forward reaction is esterification and backwards reaction is **hydrolysis** - catalyst used would be dilute HCl

Condensation Polymerisation:

- Many monomers joining to form a polymer and at the same time removing small molecules like water - forming an ester bond between two monomers when a carboxylic acid group (-COOH) reacts with an alcohol group (-OH)

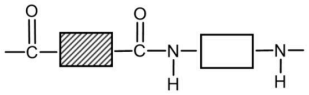
One repeat unit of Terylene is 

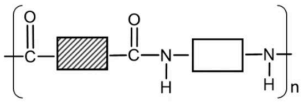
Terylene is represented as 

n is a large number

Terylene is an example,

- Another condensation polymer is nylon, which is a polyamide. The amide bond between its monomers is formed from an amine group (-NH2) and a carboxylic acid group (-COOH), with the loss of H2O

One repeat unit of Nylon is 

Nylon is represented as 

n is a large number

Terylene and Nylon are used in the production of clothes, parachutes, fishing lines and sleeping bags

Plastics are usually non-biodegradable, there are 2 methods of recycling plastics:

1. Physical methods: breaking down plastics e.g poly(ethene) into smaller pieces, melting them into pellets that can be made into ‘new’ products

2. Chemical methods:
    a. Cracking - larger molecules can be broken down into smaller molecules by applying high temperatures in the presence of a catalyst
    b. Depolymerisation - a polyester can be converted to its monomers by reacting with acids like dilute hydrochloric acid - reaction is called hydrolysis. The acid provides H+ ions that act as a catalyst