PURE CHEMISTRY

UPPER SECONDARY CHEMISTRY 2021

PURE SCIENCE NOTES

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Chapter 1: Kinetic Particle Theory and the 3 States of Matter

Kinetic Particle Theory		
All matter is made of small particles that are continually moving and hence possess kinetic energy.		
Anything that has mass and takes up space is considered matter.		
Gases	**Liquids**	**Solids**
No fixed shape or volume: particles can move freely at high speeds	No fixed shape: particles can slide and roll over each other in translation	Fixed shape and volume: particles can vibrate in fixed positions
Wide spaces in between particles: can be compressed	Very little space in between particles: cannot be compressed	
Particles far apart / widely spaced in a disorderly arrangement: low densities	Particles closely packed in a disorderly arrangement: high densities	Particles closely packed in an orderly arrangement: high densities
Particles are held together by weak forces of attraction	Particles are held together by strong forces of attraction	
Changes in State		
Melting and Boiling	[Describe the original state of the particles.] When the substance is heated, the particles will move faster as they gain more kinetic energy. When particles gain sufficient kinetic energy to overcome the forces of attraction, they will separate and move apart into [liquid / gaseous state]. [Describe the new state of the particles.]	
Condensing and Freezing	[Describe the original state of the particles.] When the substance is heated, the particles will move slower as they lose kinetic energy. When particles lose enough kinetic energy, they are no longer able to overcome the forces of attraction, so there are strong forces of attraction between them. They will get closer in [solid / liquid state]. [Describe the new state of the particles.]	
Volatile liquids are ones with boiling points close to room temperature. They evaporate easily.		

Diffusion		
Diffusion is defined as the random movement of particles of a substance.		
These particles have energy and unrestricted movement. They can move in any direction, though this occurs mainly in liquid and gaseous states. The net movement of the particles is from a region of higher to lower concentration. Eventually, particles are evenly distributed.		
Factors affecting the rate of diffusion		
Molecular mass	Temperature	States of Matter
Heavier molecules travel slower than lighter molecules.	The higher the temperature, the higher the amount of kinetic energy, the faster the molecule.	Gas particles possess the most kinetic energy and are able to move freely and randomly. Liquid particles possess kinetic energy and are able to slide and roll over each other in translation. Solid particles possess the least kinetic energy and are unable to move freely.

Chapters 2-3: Experimental Techniques, Separation and Purification

Separation Methods	
NAME AND TYPE	**KEY STEPS AND USES**
Filtration: solid-liquid	Filter the mixture using a filter funnel and filter paper. Used for filtration in water purification
Crystallisation / Evaporation to Dryness: solute from solution	Heat the solution until saturated. Check for saturation by dipping a cool glass rod into the solution. For crystallisation, let the solution cool and crystallise. Filter to get crystals, wash with distilled water and dry between filter paper. For evaporation to dryness, heat the solution until dryness. Used to separate salt from seawater.
Simple Distillation: two miscible liquids with different boiling points that are far apart	Heat the mixture and add boiling stones for smooth boiling. Condense the vapour in a condenser with cold water entering from the bottom. This ensures the water jacket is filled. The thermometer reads the temperature of the hot vapour passing into the condenser. Used for desalination and production of fermented liquor.

<table>
<tr><td>Fractional Distillation:
two miscible liquids with different boiling points that are close together</td><td>Same steps as simple distillation.
The additional fractionating column provides surface area for higher boiling point gases to cool and condense into liquid, while the gases that are boiling / of lower boiling point do not.
The vapour at top is only the gas with lowest bp and is pure.
Used for refining of crude oil.
Used in the production of liquified oxygen and liquified nitrogen from liquified air.</td></tr>
<tr><td>Separating Funnel:
mixtures of immiscible liquids</td><td>Remove the bottom layer(s) of liquid.
Used to separate oil / chloroform and water.</td></tr>
<tr><td>Sublimation:
solid mixture of a substance that sublimes and others that are stable to heat</td><td>Heat with an inverted funnel above the evaporating dish.
The vapour of the substance that sublimes will resolidify on the inner walls of the filter funnel.
Used mostly for purification of iodine, ammonium chloride or naphthalene.</td></tr>
<tr><td>Paper Chromatography:
determination of the identity and purity of complex mixtures</td><td>Draw a starting line in pencil so there is no ink to separate.
Let the solvent run as far as possible for better separation.
The spots marked on the starting line should be above the solvent level, to prevent them from dissolving.
Cover the boiling tube / beaker to prevent evaporation of the volatile organic solvent.
Apply a locating agent to the paper.
The identity of substance can be found using rf value.
Used to separate pigments in flowers.
Used to test for chemicals in food.
Used to test athletes’ urine samples.</td></tr>
<tr><td colspan="2">rf value = $\frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$
<i>(affected by the type of solvent and temperature)</i></td></tr>
</table>

<table>
<tr><th colspan="3">Testing for Purity</th></tr>
<tr><td>Impurities cause melting / boiling over a range</td><td>Impurities typically decrease melting point and increase boiling points.</td><td>Using paper chromatography.</td></tr>
</table>

Chapter 4: Elements, Compounds and Mixtures

Elements	Compounds	Mixtures
Can be found as monoatomic, diatomic or polyatomic structures of a single element.	Consists of two or more elements chemically combined together in a fixed ratio.	Consists of two or more substances physically combined together in any ratio.
Either metal or non-metal.	Either ionic or covalent. Substance formed does not share properties with constituent elements.	Components retain individual properties.
Cannot be broken down into simpler substances by chemical means.	Can be separated using chemical methods.	Can be separated using physical methods.

Chapter 5: Atomic Structure

Subatomic Particles	Relative Charge and Mass
Proton (p)	+1, 1
Neutron (n)	0, 1
Electron (e)	-1, $^{1}/_{1840}$
Isotopes	
SIMILARITIES	**DIFFERENCES**
Number of protons and electrons	Number of neutrons
Chemical properties — neutrons don't affect reactions	Physical properties — isotopes have different masses
Representation of Atomic Structure	
${}^{A}_{Z}X$	$\overset{Z}{\underset{B}{X}}$
A: nucleon number, exact neutrons and protons B: relative atomic mass: average neutrons and protons amongst all isotopes Z: proton / atomic number, protons X: chemical symbol of element	

Chapters 6-7: Chemical Bonding: Ionic, Covalent and Metallic

<table>
<tr><th>Property</th><th>Ionic compounds</th><th>Simple covalent compounds</th></tr>
<tr><td>Elements in compound</td><td>Metal and non-metal</td><td>Non-metal and non-metal</td></tr>
<tr><td>Type of particle</td><td>Only ions</td><td>Simple molecules</td></tr>
<tr><td>Melting and boiling points</td><td>Higher melting and boiling points, ions held together in a giant crystal lattice by strong ionic bonds.</td><td>Lower melting and boiling points, molecules held in simple molecular structure by weak intermolecular forces.</td></tr>
<tr><td>Solubility</td><td>Most are soluble in water, but insoluble in organic solvents.</td><td>Most are insoluble in water, but soluble in organic solvents.</td></tr>
<tr><td>Electrical conductivity</td><td>Conducts in aqueous and molten states, ions are mobile and able to hold charges.</td><td>Does not conduct as no mobile particles. However, there are exceptions that react with water and form ions in the aqueous state, such as HCl.</td></tr>
<tr><td>Structure</td><td>Giant crystal lattice</td><td>Simple molecular</td></tr>
<tr><td>Bonds present</td><td>Strong ionic bonds</td><td>Weak intermolecular forces</td></tr>
<tr><th colspan="3">Metallic bonding</th></tr>
<tr><td colspan="3">Metals are a lattice of positive ions in fixed positions surrounded by a sea of delocalised electrons.</td></tr>
<tr><td colspan="3">**Metallic bond** is defined as the strong electrostatic forces of attraction between the positive metal ions and the sea of delocalised electrons.</td></tr>
<tr><th colspan="3">Properties of metals</th></tr>
<tr><td>Their atoms function as neutral
→ the sea of delocalised electrons neutralises the lattice of positive ions.</td><td>They have high boiling and melting points
→ metals are giant lattice structures, so the number of electrostatic forces to be broken is extremely large.</td><td>They are malleable and ductile
→ the structure of metals consists of layers of metal ions that can slide over each other when a force is applied, allowing the metals to be rolled into sheets or be drawn into wires.</td></tr>
<tr><td>They are good conductors of electricity in solid and molten states
→ they contain a sea of delocalised electrons (mobile charged particles) that can move in one direction and form an electric current.</td><td colspan="2">They are good conductors of heat
→ their particles are very closely packed, so the vibrations are passed on very quickly, and they contain large numbers of free electrons that drift slowly through the structure.
→ when the metal is heated, the free electrons closest to the heat source are heated, making them move faster as they travel through the metal, colliding with both atoms and other electrons. this makes these other particles vibrate faster.</td></tr>
</table>

<table>
<tr><th colspan="4">Macromolecules</th></tr>
<tr><th>GRAPHITE</th><th>DIAMOND</th><th>SILICON</th><th>SILICON DIOXIDE</th></tr>
<tr><td>Flat hexagonal rings of carbon atoms are arranged in layers. Each carbon atom is covalently bonded to three others.</td><td colspan="2">Tetrahedral arrangement of carbon/silicon atoms. Each carbon/silicon atom is covalently bonded to four others.</td><td>Each silicon atom is covalently bonded to four oxygen atoms in a tetrahedral arrangement. Each oxygen atom is bonded to two silicon.</td></tr>
<tr><td>Only ¾ of the valence electrons are used for bonding. There are weak Van der Waals forces of attraction between layers.</td><td colspan="3">All valence electrons used in covalent bonding.</td></tr>
<tr><td>Soft and slippery. Weak Van der Waals forces of attraction allow layers to slide off easily with force.</td><td>Hard. Strong covalent bonds extend through the whole structure making it rigid.</td><td colspan="2"></td></tr>
<tr><td>Conducts electricity. ¼ of valence electrons are free and mobile.</td><td colspan="3">Does not conduct electricity. All valence electrons are used in bonding. There are no mobile electrons to hold charges.</td></tr>
<tr><td colspan="4">High boiling and melting points. Need to overcome strong covalent bonds within the layer / throughout the entire structure.</td></tr>
<tr><td>Used as pencil lead, a lubricant (It does not burn / decompose on heating), electrodes in batteries.</td><td>Used for jewelry, attached to saws and drills.</td><td></td><td>Its powdered form is used as an abrasive in toothpaste.</td></tr>
</table>

Chapters 8-10: Formulae, Stoichiometry and the Mole Concept

Stoichiometry Formulae		
Number of moles = $\frac{\text{mass of substance in grams}}{\text{molar mass}}$	Percentage yield = $\frac{\text{actual yield}}{\text{theoretical yield}}$	Concentration in moles/dm^3 = $\frac{\text{number of moles in solute}}{\text{volume in } dm^3}$
Number of moles of gas = $\frac{\text{gas volume at r.t.p.}}{24\ dm^3}$	Percentage purity = $\frac{\text{mass of pure substance}}{\text{total mass of sample}}$	Concentration in grams/dm^3 = $\frac{\text{mass of solute in grams}}{\text{volume in } dm^3}$
Relative atomic mass = $\frac{\text{average mass of one atom of the element}}{\frac{1}{12}\text{ the mass of an atom of carbon-12}}$	Relative molecular mass = $\frac{\text{average mass of one molecule}}{\frac{1}{12}\text{ the mass of an atom of carbon-12}}$	Avogadro's constant = 6×10^{23}

Ionic equations	
Displacement	$M\ (s) + N^+\ (aq) \rightarrow N\ (s) + M^+\ (aq)$
Metal-acid	$M\ (s) + 2H^+\ (aq) \rightarrow M^+\ (aq) + H_2\ (g)$
Precipitation	$M\ (aq) + F\ (aq) \rightarrow MF\ (s)$
Neutralisation	$OH^-\ (aq) + H^+\ (aq) \rightarrow H_2O\ (l)$
F: non-metal, M, N: metals, H^+: hydrogen, OH^-: hydroxide, H_2O: water	

Chapters 11-12: Acids, Bases and Salts [including QA]

Types of Oxides			
Basic	**Amphoteric**	**Acidic**	**Neutral**
Metal oxides	ZAP [Zn, Al, Pb] oxides	Non-metal oxides with multiple oxygen atoms	Non-metal oxides with a single oxygen atom
React with water to form acids	Insoluble in water		
React with alkalis to form salt and water	React with both alkalis and acids to form salt and water.	React with acids to form salt and water.	Does not react with either alkalis or acids.

Properties of Acids and Alkalis			
ACIDS	React with bases to form salt and water.	React with carbonates to form salt, water and carbon dioxide.	React with metals to form salt and hydrogen.
ALKALIS	React with acids to form salt and water.	React with ammonium salts to form salt, ammonia gas and water when heated gently.	

Methods of salt preparation		
TITRATION	**EXCESS INSOLUBLE SOLID**	**PRECIPITATION**
Acid + alkali / soluble carbonate → soluble salt	Acid + metal / base / carbonate → soluble salt	Solution containing cation + solution containing anion → insoluble salt + other product
Conduct titration.	Add excess of solid to acid. Filter the mixture.	Add excess of one of the solutions and stir until no more precipitate forms. Filter the mixture.

Basicity, Strength and Concentration		
BASICITY	**STRENGTH**	**CONCENTRATION**
Basicity refers to the number of moles of H$^+$ or OH$^-$ ions that are being dissociated per mole of an acid / alkali.	Strength refers to how easily an acid / alkali dissociates when dissolved in water.	Concentration refers to how much of an acid (the solute) is dissolved in the solution.
Monobasic: 1 mole of H$^+$ or OH$^-$ ions are dissociated per mole of acid / alkali. Dibasic: 2 moles of H$^+$ or OH$^-$ ions are dissociated per mole of acid / alkali. Tribasic: 3 moles of H$^+$ or OH$^-$ ions are dissociated per mole of acid / alkali.	⇌ : Weak acids / alkalis undergo partial dissociation. → : Strong acids / alkalis undergo full dissociation.	The concentration of a solution is given by the amount of a solute dissolved in a unit volume of the solution.

Solubility rules	
Soluble	All SPA salts and nitrates All halides except lead (ii) and silver All sulfates except lead (ii), calcium and barium
Insoluble	All carbonates except SPA carbonates All hydroxides and oxides except SPA hydroxides / oxides

Sub-chapter: Qualitative Analysis

Test for anions		
ANION	**TEST**	**TEST RESULT**
Carbonate (CO_3^{2-})	Add dilute acid	Effervescence of carbon dioxide
Chloride (Cl^-)	Acidify with dilute nitric acid Add aqueous silver nitrate	White precipitate
Iodide (I^-)		Yellow precipitate
Nitrate (NO_3^-)	Add aqueous sodium hydroxide then aluminium foil / powder Warm carefully	Ammonia produced
Sulfate (SO_4^{2-})	Acidify with dilute nitric acid Add aqueous barium nitrate	White precipitate

Test for cations		
CATION	**EFFECT OF AQ SODIUM HYDROXIDE**	**EFFECT OF AQ AMMONIA**
Aluminium (Al^{3+})	White precipitate. Soluble in excess to form a colourless solution.	White precipitate. Insoluble in excess.
Ammonium (NH_3^-)	Ammonia produced on warming.	
Calcium (Ca^{2+})	White precipitate. Insoluble in excess.	No precipitate.
Copper [ii] (Cu^{2+})	Light blue precipitate. Insoluble in excess.	Light blue precipitate. Soluble in excess to form a dark blue solution.
Iron [ii] (Fe^{2+})	Green precipitate. Insoluble in excess.	Green precipitate. Insoluble in excess.
Iron [iii] (Fe^{3+})	Red-brown precipitate. Insoluble in excess.	Red-brown precipitate. Insoluble in excess.
Lead [ii] (Pb^+)	White precipitate. Soluble in excess to form a colourless solution.	White precipitate. Insoluble in excess
Zinc (Zn^{2+})	White precipitate. Soluble in excess to form a colourless solution.	White precipitate. Soluble in excess to form a colourless solution.
ALUMINIUM VS LEAD [ii]		
Add a soluble chloride to the solution with the cation. Lead [ii] chloride is insoluble → white precipitate observed. Aluminum chloride is soluble → no precipitate observed.		

Chapter 13: Redox Reactions

What is a redox reaction?

It is a chemical reaction which involves the oxidation of a substance and reduction of another substance in the same reaction.

Rules for Assigning Oxidation Numbers / States

Oxidation states are hypothetical numbers assigned to an individual element by a set of rules.

1. The oxidation state of any element in the free / uncombined state is 0.
2. The oxidation state of any monoatomic ions is equal to the charge on the ion.
3. Combined oxygen: oxidation number is -2. (When oxygen is in a compound, its oxidation number is -2)
   Exception: The oxidation number of oxygen in peroxides is -1.
4. Combined hydrogen: oxidation number is +1. (When hydrogen is in a compound, its oxidation number is +1)
   Exception: The oxidation number of hydrogen in metal hydrides is -1.
5. In any neutral compound, the sum of all oxidation states must add up to 0.
6. In any polyatomic ion, the sum of all oxidation states must add up to the atom's net charge.

Oxidation	Reduction
Gain in oxygen, loss of hydrogen	Loss in oxygen, gain in hydrogen
Gain in oxidation state	Loss in oxidation state

OIL RIG: Oxidation **I**s **L**oss of electrons, **R**eduction **I**s **G**ain of electrons

Oxidising agents	Reducing agents
Oxidizes another substance and gains electrons, undergoing reduction	Reduces another substance and gains electrons, undergoing oxidation
$KMnO_4$: purple Mn^{7+} to colourless Mn^{2+} $K_2Cr_2O_7$: orange Cr^{6+} to green Cr^{3+} $FeCl_3$: yellow Fe^{3+} to pale green Fe^{2+}	$FeSO_4$: pale green Fe^{2+} to yellow Fe^{3+} KI : colourless I^- to brown I_2
Halogens: Colour change depends on halogen. Oxygen: Involved in combustion reactions.	Halides: Colour change depends on halogen. Metals: Lose electrons to form ions. The more reactive, the stronger it is as a reducing agent. Hydrogen: Oxidises to form H_2O. Carbon: Oxidises to form CO or CO_2. Carbon monoxide: Oxidises to form CO_2.

Hydrogen peroxide can act as both an oxidising and reducing agent, depending on the reactants it reacts with.

Chapter 14: Metals

Alloys		
A mixture of a main metal with metallic / non metallic elements.		
Why are alloys stronger than pure metals?		
Addition of foreign atoms that are a different size than the majority of the atoms.	Disruption of the regular arrangement of atoms in the pure metal.	Layers of atoms cannot slide over each other as easily, so alloys are harder.

Reactions of metals				
Mnemonic: Please stop calling me a cute zebra, I like hardworking cunning smart girls.				
METAL	**EXTRACTION**	**RXN WITH WATER (HYDROXIDE PROD.)**	**RXN WITH STEAM (OXIDE PROD.)**	**RXN WITH ACID (HYDROGEN PROD.)**
Potassium	Electrolysis	Very violently. Darts on the surface of the water with lilac (K) / yellow (Na) flame. Hissing sound, sparks and heat given off.	Explosively. Hydrogen gas explodes.	Explosively.
Sodium				
Calcium		Moderately. Sinks. Effervescence seen when hydrogen is produced. White solid seen as CaOH only slightly soluble in water.	Violently.	Vigorously.
Magnesium		Very slowly.	Vigorously. White light observed. White ash / solid observed.	
Aluminium		No reaction occurs. Aluminium naturally forms an oxide layer that is impervious.		
CARBON				
Zinc	Reduced by carbon.	No reaction.	Moderately.	Moderately.

Iron	Reduced by carbon or hydrogen.		Slowly.	Slowly.
Lead			No reaction.	Very slight slow reaction which stops quickly. Insoluble lead [ii] chloride prevents further reaction.
HYDROGEN				
Copper	Reduced by carbon or hydrogen.	No reaction.	No reaction.	No reaction.
Silver				
Gold	Found uncombined.			

Reactions of metal compounds

METAL	**CARBONATE RXN WITH HEAT**	**OXIDE RXN WITH HEAT**	**OXIDE RXN WITH C/CO**	**OXIDE RXN WITH H_2**
Potassium	Stable to heat	Stable to heat	Not reduced when heated	Not reduced when heated
Sodium				
Calcium	Decomposes to metal oxide and carbon dioxide			
Magnesium				
Aluminium				
Zinc			Reduced with increasing ease, forming metal and carbon dioxide	
Iron				Reduced with increasing ease forming metal and steam
Lead				
Copper				
Silver	Decomposes to metal, oxygen and carbon dioxide	Decomposes to metal and oxygen		

Extraction of iron
Haematite (Fe_2O_3), **limestone ($CaCO_3$)** and **coke** added to the top of the blast furnace. **Hot air** is blown in through pipes from the bottom as a source of heat.

<table>
<tr><td colspan="3">Refractory lining, typically magnesium oxide, is an ionic compound with a high melting point used to line the insides of the blast furnace.</td></tr>
<tr><th>Production of the key reducing agent</th><th>Key reaction producing iron</th><th>Removal of acidic impurities</th></tr>
<tr><td>1. $C + O_2 \rightarrow CO_2$
Coke reacts with oxygen from hot air to form carbon dioxide.</td><td rowspan="2">3. $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$
a. Carbon monoxide reduces iron [iii] oxide / haematite, to form molten iron.
b. Iron flows to the bottom of the furnace where it is tapped off.
c. Waste gases (CO, CO_2, N_2) escape through the top of the furnace.</td><td>4. $CaCO_3 \rightarrow CaO + CO_2$
Limestone is decomposed by heat to form calcium oxide and carbon dioxide.</td></tr>
<tr><td>2. $CO_2 + C \rightarrow 2CO$
The carbon dioxide formed reacts with more coke to form carbon monoxide.</td><td>5. $SiO_2 + CaO \rightarrow CaSiO_3$
Basic CaO reacts with acidic SiO_2 to form slag that flows to the bottom of the furnace and floats on top of molten iron. The slag is tapped off to be used in making roads.</td></tr>
</table>

<table>
<tr><th colspan="2">Steel</th></tr>
<tr><td>High carbon steel is stronger and brittle.</td><td>Low carbon steel is softer and easily shaped.</td></tr>
<tr><th colspan="2">Prevention of Rusting</th></tr>
<tr><td colspan="2">Rusting is a common term for the corrosion of iron and its alloys that occurs in the presence of oxygen and moisture. This process is accelerated by salt.</td></tr>
<tr><th>SURFACE / BARRIER PROTECTION</th><th>SACRIFICIAL PROTECTION</th></tr>
<tr><td>Coating and painting to isolate iron from the environment with unreactive, inert substances.</td><td>Attaching zinc or magnesium to the iron. Both metals are more reactive and corrode instead.</td></tr>
<tr><td>When they are coated onto iron, the substances will form a protective layer. When scratched off, the iron will not be protected anymore.</td><td>Galvanisation is when a layer of zinc is coated onto the iron. Zinc is used as it is cheap, adheres well to steel and provides cathodic protection to the steel layer.</td></tr>
<tr><th colspan="2">Recycling of metals</th></tr>
<tr><td colspan="2">The main aim of recycling of metals is to preserve finite metal ores.</td></tr>
<tr><th>ADVANTAGES</th><th>DISADVANTAGES</th></tr>
<tr><td>- Recycling creates jobs.
- It conserves energy, reduces the cost of disposing trash in landfills or solid waste incinerators and allows the extraction of valuable metals.
- It reduces pollution and the chances of leaching of metals into soil and river water.</td><td>- Recycling can pollute the environment as metal fumes may be released during smelting and less valuable metals may be thrown away.
- Time and energy is needed in sorting and transporting recycled materials.</td></tr>
</table>

Chapter 15: Electrolysis

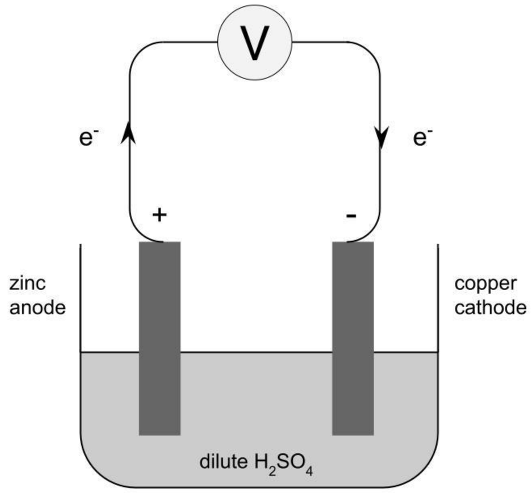
What is electrolysis?
The conduction of electricity by a molten electrolyte or an aqueous electrolyte that occurs when an electric current is passed through it that causes it to decompose.
Memorise: $4OH^- (aq) \rightarrow 2H_2O (l) + O_2 (g) + 4e^-$

ANODE	CATHODE	ELECTROLYTE
Anode is the positive electrode that attracts anions and discharges them to form non-metals.	Cathode is the negative electrode that attracts cations and discharges them to form metals.	Electrolyte is a compound that conducts an electric current in molten / aqueous state and is decomposed by it.
Anions migrate here and lose electrons → oxidation	Cations migrate here and gain electrons → reduction	

Molten vs. Aqueous electrolytes	
MOLTEN	**AQUEOUS**
Only contains the ions of the compound.	Contains the ions of the compound and water.
Only the ions of the compound are discharged.	Rules must be applied to determine which ion is discharged as hydrogen and hydroxide ions are present in the solution.
Applications of electrolysis	
PURIFICATION OF COPPER	**ELECTROPLATING**
Impurities drop down from impure Cu anode.	Object be plated is placed at the cathode.
Copper deposits on the pure Cu cathode.	Metal to be plated with is placed at the anode.
Electrolyte must contain Cu ions.	Electrolyte must be the metal to be plated with.

Factors affecting preferential discharge of ions in aqueous electrolytes		
The lower the ion in the electrochemical series, the more easily it is discharged.	**Anions:** F^- SO_4^{2-} $NO3_3^{2-}$ Cl^- Br^- I^- OH^-	If an electrolyte is concentrated, Cl^-, Br^- and I^- ions are discharged in preference to OH^- ions. F^-, SO_4^{2-} and NO_3^{2-} will not be discharged no matter what.
The pH of a solution can change during electrolysis depending on the concentration of OH^- and H^+ ions.		

<table>
<tr><td>Active (Cu, Ag) electrodes tend to ionise during electrolysis.
Inert (C, Pt) electrodes do not ionise during electrolysis.</td><td rowspan="3">Electrolysis of copper (ii) sulfate:
<table>
<tr><th>Inert Electrode</th><th>Copper Electrode</th></tr>
<tr><td>Effervescence observed, hydrogen discharged. Anode size remains unchanged.</td><td>No effervescence observed. Anode decreases in size, copper in anode ionises, forms copper [ii] ions in solution.</td></tr>
<tr><td>Blue colour of electrolyte fades over time as copper [ii] ions are discharged and removed from solution.</td><td>Blue colour of electrode remains unchanged as copper [ii] ions discharged at cathode are replaced at the same rate as those formed at anode.</td></tr>
</table></td></tr>
<tr><td>if the anode is an active anode, no ions will be discharged and the anode will ionise instead.</td></tr>
<tr><td>Water is a weak electrolyte, so small amounts of sulfuric acid are added in electrolysis of water.</td></tr>
</table>

<table>
<tr><th colspan="2">Simple/Electrochemical Cell</th></tr>
<tr><td colspan="2">A simple cell is made up of two metals of different reactivities which are connected to each other and dipped in an electrolyte. Simple cells are convenient and portable sources of energy.</td></tr>
<tr><td>
</td><td>The more reactive metal (zinc) loses electrons to the less reactive metal (copper), which generates a current.
As it ionises, the more reactive metal (zinc) decreases in size.
This causes the less reactive metal (copper) to be negatively charged and thus be the cathode.
Cations in the electrolyte are attracted to the less reactive metal (copper) and are discharged.
The greater the difference in reactivity between the two metals, the greater the voltage and current generated.</td></tr>
</table>

Chapter 16: The Periodic Table

<table>
<tr><th colspan="3">Group I. Alkali Metals</th></tr>
<tr><td>Good conductors of heat and electricity</td><td>Soft and malleable</td><td>Low melting point that decreases down the group</td></tr>
<tr><td>Density of elements increases down the group</td><td>Chemical reactivity increases down the group</td><td>Strength of the metallic bonds decrease down the group</td></tr>
<tr><td>Tend to be stored in oil as they are very reactive</td><td>Their compounds are extremely stable to heat</td><td>Extracted by electrolysis</td></tr>
</table>

Group VII. Halogens

Very reactive non metals, decreases down the group	Form molecular crystals, so their melting and boiling points are low and increase down the group

Colours of the halogens get darker down the group

Halogens are slightly soluble in water and very soluble in organic solvents (Ethanol, CCl_4, C_2H_5OH)

Halogen	State (r.t.p.)	Colour	Colour in H_2O
F_2	gas	pale yellow	
Cl_2		greenish yellow	pale yellow
Br_2	liquid	reddish brown	reddish brown
I_2	solid	black (purple vapor)	dark brown
At_2		black	

Group 0. Noble Gases

Stable electronic configuration	Provide an inert atmosphere	Exist as monatomic gases

Transition Elements

High melting and boiling point, high densities	Excellent conductivity of heat and electricity	Hard and strong but have good malleability and ductility

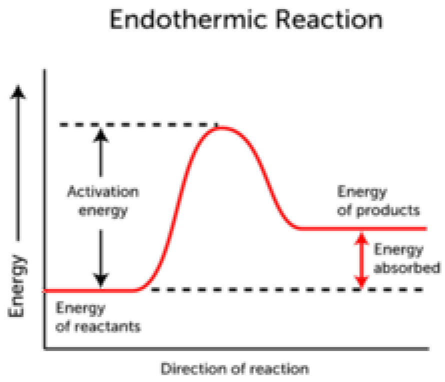
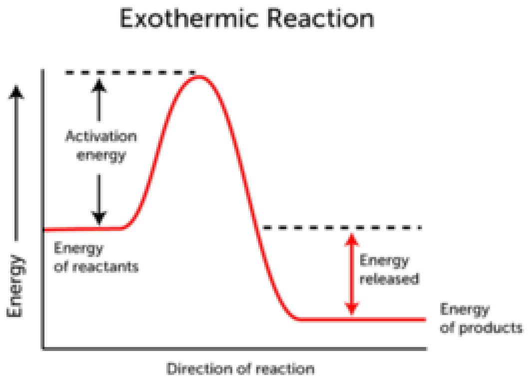
ELEMENT	ION	COLOUR OF COMPOUND
Iron	Fe^{2+}	pale green
	Fe^{3+}	yellow, orange, brown
Copper	Cu^{+}	red, cream
	Cu^{2+}	blue, green
Chromium	Cr^{2+}	blue
	Cr^{3+}	green
Manganese	Mn^{2+}	colourless
	Mn^{7+}	purple

Only transition elements:
- Can act as catalysts
- Have ions of different charges
- Form compounds of different colours

Not all transition elements have these traits, but most do.

Chapter 17: Energy from Chemicals

Energy changes in reactions
Energy must be absorbed when bonds are broken, so bond breaking is endothermic. Energy must be released when bonds are formed, so bond breaking is exothermic.
Enthalpy change (ΔH) is the energy exchanged between the system and the surroundings. It can be calculated using bond energies. Absolute enthalpy (H) of a substance cannot be calculated.
The larger the amount of energy absorbed during bond breaking or released during bond forming, the stronger the bond.

ENDOTHERMIC REACTIONS	**EXOTHERMIC REACTIONS**
There is an increase in the system's overall energy, while the energy of the surroundings decreases by the same magnitude.	There is a decrease in the system's overall energy, while the energy of the surroundings increases by the same magnitude.
Energy released in bond formation is less than energy absorbed in bond breaking.	Energy released in bond formation is more than energy absorbed in bond breaking.
Positive enthalpy change as the system gains energy from the surroundings.	Negative enthalpy change as the system loses energy from the surroundings.
	
Examples: - Thermal decomposition - Photosynthesis - Decomposition of CFCs by sunlight - Dissolving of ammonium chloride and sodium nitrate - Melting and boiling [Physical processes]	Examples: - Combustion of fuels - Neutralisation - Metal-acid reactions - Displacement reactions - Respiration - The Haber process

Types of Fuels

Fuels are substances that are burnt to produce energy. They are exothermic reactions.

Heat (enthalpy) of combustion is the energy given out when a fuel burns completely and is measured in kilojoules per mole.

The energy density of a fuel is the amount of energy given out per gram of fuel burnt.

HYDROCARBON / FOSSIL FUELS

- The most commonly used fuels that are depleting at a fast rate.
- Made from decayed plants and animals that lived millions of years ago.

Disadvantages:
- During incomplete combustion due to a limited supply of air, soot (unburnt hydrocarbons) and carbon monoxide are formed.
- Carbon dioxide is formed in the burning of fossil fuels.
- The burning of coal produces large amounts of ash and sulfur dioxide.
- The depletion of natural resources.

HYDROGEN-OXYGEN FUEL CELLS

$$\mathbf{2H_2\ (g) + O_2\ (g) \rightarrow 2H_2O\ (l)}$$

- Hydrogen is the fuel with the highest energy density.
- Hydrogen (from electrolysis of water or cracking of oils) and oxygen (from fractional distillation of liquid air) is passed through an electrolyte in a fuel cell.
- A good alternative to the combustion of fossil fuels.

Applications:
- Used in NASA's space shuttles to generate electricity.
- Used to generate electricity in certain buses.

ADVANTAGES	DISADVANTAGES
Produces only water when burnt and thus does not have any pollutant products.	Hydrogen is a gas at room temperature and must be kept in a thick pressurised container which is heavy, causing it to be harder to store and transport compared to liquid fuels.
Highly efficient as chemical energy is converted directly to electrical energy, with little loss of energy.	A gram of hydrogen takes up a lot of space compared to liquid fuels due to its low density.
Fuel cells can produce energy indefinitely as there is a continuous supply of fuel.	Hydrogen is flammable and explosive.
The hydrogen obtained from electrolysis of water is renewable.	Obtaining hydrogen from electrolysis of water requires a lot of energy and uses fossil fuels.

Chapter 18: Rate of Reaction

Chemical reactions

Chemical reactions occur when:
- Reactant particles collide with one another
- Colliding particles possess energy equal to or greater than the activation energy

Measuring speed of reaction can be done through:
- Measuring the volume of gas produced
- Measuring the change in mass

Factors affecting rate of reaction

FACTOR	EXPLANATION
Concentration of reactants	With an increased concentration, there are more reactant particles per unit volume, which increases the frequency of collisions and thus increases the frequency of effective collisions.
Pressure of gaseous reactants	With an increased pressure, reactant particles are closer together, which increases the frequency of collisions and thus increases the frequency of effective collisions.
Temperature of reactants	With an increased temperature, the particles move at higher speeds due to higher kinetic energy, which increases the frequency of collisions. In addition, more particles have energy equal or greater to the activation energy of the reaction. This causes the frequency of effective collisions to increase.
Particle size of solid reactants	With an increased particle size, there is a greater total surface area for reactant particles to collide with, increasing the frequency of collisions and thus that of effective collisions.
Presence of catalysts	The presence of catalysts causes the activation energy of a reaction to be lowered as they provide an alternative pathway for the reaction to proceed. This causes more particles to have energy equal or greater to the activation energy with the catalyst, increasing the frequency of effective collisions.

Catalysts

A catalyst is a substance which increases the speed of a chemical reaction and remains chemically unchanged at the end of the reaction.

- Only a small amount of a catalyst is needed to speed up the reaction.
- Lowers the activation energy of a reaction.
- Impurities can prevent catalysts from working.
- A catalyst is selective in its action — different catalysts speed up different reactions.

Catalysts are used to speed up various industrial processes like the Haber process (iron) and catalytic converters (platinum or rhodium). Enzymes are biological catalysts that catalyse chemical reactions in plants and animals.

Chapter 19: Ammonia

<table>
<tr><th colspan="6">Properties of Ammonia</th></tr>
<tr><th colspan="2">PHYSICAL</th><th colspan="2">CHEMICAL</th><th colspan="2">USES</th></tr>
<tr><td colspan="2">Colorless alkaline gas with a pungent smell</td><td colspan="2">Acts as a weak alkali</td><td colspan="2">To produce fertilisers</td></tr>
<tr><td colspan="2">Lighter than air and easily liquefied / very soluble</td><td colspan="2">Precipitates insoluble hydroxides of metals</td><td colspan="2">As a cooling agent in refrigerators</td></tr>
<tr><th colspan="6">Laboratory Preparation of Ammonia</th></tr>
<tr><th colspan="2">DISPLACEMENT FROM SALT</th><th colspan="2">HEATING OF CERTAIN SALTS</th><th colspan="2">NO_3 + NaOH + Al</th></tr>
<tr><td colspan="2">Ammonia is displaced from its salts by heating with a strong alkali (E.g.: NaOH, KOH, CaOH)</td><td colspan="2">Thermal dissociation: ⇌ Compounds that decompose on heating but reform the original compound on cooling</td><td colspan="2">Addition of sodium hydroxide and Devarda’s alloy to a solution of nitrate ions.</td></tr>
<tr><td colspan="2">Ammonium salt + base → salt + water + ammonia</td><td colspan="2">Thermal decomposition: → Compounds that decompose on heating and do not reform on cooling</td><td colspan="2">$3NaNO_3 + 8Al + 5NaOH + 18H_2O \rightarrow 3NH_3(g) + 8NaAl(OH)_4$</td></tr>
<tr><th colspan="6">Haber Process</th></tr>
<tr><th colspan="3">RAW MATERIALS</th><th colspan="3">OPTIMAL CONDITIONS FOR PRODUCTION</th></tr>
<tr><td colspan="3">Nitrogen gas from fractional distillation of liquefied air</td><td colspan="3">Nitrogen and hydrogen are dried and mixed in the proportion 1:3</td></tr>
<tr><td colspan="3">Hydrogen gas from cracking of crude oil</td><td colspan="3">Gasses passed over finely divided iron catalyst at 450°C at 200 atm</td></tr>
<tr><td colspan="3"></td><td colspan="3">The final mixture contains about 10% ammonia gas which is cooled then removed. Unreacted reactants are recycled for further production.</td></tr>
<tr><th colspan="6">Calcium Hydroxide</th></tr>
<tr><td colspan="3">Calcium hydroxide is added to soil to increase the pH of acidic soils caused by acid rain.</td><td colspan="3">Calcium hydroxide should not be added at the same time as ammonium fertilisers as they will react to form ammonia, calcium salt and water, causing ammonia and hydrogen to be removed from the soil.</td></tr>
</table>

Chapter 20: Air and Atmosphere

Composition of Air
Air is a mixture of several gases.

GAS	COMPOSITION OF DRY AIR BY VOLUME / %
Nitrogen	78
Oxygen	21
Other gases: - Noble gases (Mostly argon) - Carbon dioxide	1 0.97 0.03

Pollutants
Air pollution is the condition in which air contains a high concentration of chemicals that may harm living things or damage non-living things. It is caused by solid particles and poisonous gases in the air.

POLLUTANT	SOURCE	EFFECTS
Carbon monoxide CO (Dangerous as it is colourless and odorless)	Incomplete combustion of carbon-containing fuels	Fatigue, headache and death. CO binds irreversibly with the haemoglobin in blood in place of oxygen, leading to a lack of supply of oxygen to cells and tissues.
Oxides of nitrogen NO, NO_2, N_2O etc.	Natural lightning activity Combustion of nitrogen gas at high temperatures - $N_2 + O_2 \rightarrow NO$ - $2NO + O_2 \rightarrow 2NO_2$	Irritation to eyes and lungs, causing breathing difficulties. Reacts with water to form acid rain that corrodes metals, cement, limestone and marble. Acid rain also leaches nutrients in soil and causes soil / water to be acidic, harming aquatic life and plants.
Sulfur dioxide SO_2	Volcanic eruptions Combustion of fossil fuels that contain sulfur - $S + O_2 \rightarrow SO_2$	
Methane CH_4	Decaying matter Digestion of food	Enhanced greenhouse effect.
Ozone O_3	Product of the reaction of unburnt hydrocarbons with NO_2 Pollutant when in the lower atmosphere	Damages crops. Forms photochemical smog that irritates eyes and lungs, causing breathing difficulties.
Unburnt hydrocarbons C_xH_y	Incomplete combustion of fuels	Reacts with NO_2 to form ozone. Carcinogenic with prolonged exposure.

Chlorofluorocarbons $C_xCl_yF_z$	Produced to be used as propellants for aerosols, coolants in refrigerators or to help manufacture packing foam.	Deplete the ozone layer by being broken down by UV radiation to form chlorine atoms, which break down ozone to form chlorine oxide and oxygen gas. Since chlorine oxide is unstable, another oxygen atom will steal the oxygen atom the chlorine is bonded with. The chlorine atom then attacks another ozone molecule.

<table>
<tr><th colspan="2">The Ozone Layer</th></tr>
<tr><th>IMPORTANCE</th><th>PROBLEMS WITH DEPLETION</th></tr>
<tr><td>Ozone in the stratosphere filters out 99% of dangerous ultraviolet radiation from the sun.</td><td>Ozone depletion allows more ultraviolet radiation to penetrate the atmosphere, causing an increase in skin cancer and cataract issues.</td></tr>
<tr><th colspan="2">Preventive methods</th></tr>
<tr><th>CATALYTIC CONVERTER IN VEHICLES</th><th>FLUE GAS DESULFURIZATION</th></tr>
<tr><td>Converts CO, NO and unburnt hydrocarbons into less harmful substances.</td><td rowspan="2">Removes SO_2 from the air</td></tr>
<tr><td>$2CO + O_2 \rightarrow 2CO_2$
$2NO + CO \rightarrow N_2 + 2CO_2$</td></tr>
<tr><td>Unburnt hydrocarbons are converted by combustion to carbon dioxide and water vapour.</td><td rowspan="2">$SO_2 + CaCO_3 \rightarrow CaSO_3 + CO_2$
$SO_2 + CaO_3 \rightarrow CaSO_3$
Pass the sulfur dioxide through wet calcium oxide / calcium carbonate. They will react to form calcium sulfite.</td></tr>
<tr><td>Consists of a stainless steel chamber with a platinum or rhodium catalyst within. This catalyst is coated on ceramic tubes and exists as a fine layer.</td></tr>
<tr><td>Vehicles with catalytic converters cannot use leaded petrol as the lead coats the catalyst, making it ineffective.</td><td>$2CaSO_2 + O_2 \rightarrow 2CaSO_4$
This calcium sulfite is further oxidised to calcium sulfate.</td></tr>
<tr><th colspan="2">The Carbon Cycle</th></tr>
<tr><td colspan="2">- The mechanism that maintains the level of carbon dioxide in the atmosphere.
- Comprises a sequence of events that maintains a constant amount of carbon dioxide.
- The rate of removal of carbon from the air is equal to the rate of return of carbon to the air.
- Three main processes that produce carbon dioxide: respiration, combustion and decay.
- Two main processes that remove carbon dioxide: photosynthesis and ocean uptake.</td></tr>
</table>

PRODUCTION OF CARBON DIOXIDE	**REMOVAL OF CARBON DIOXIDE**
Respiration: $C_6H_{12}O_6 + 6O_2 \rightarrow 6H_2O + 6CO_2 + \text{energy}$ - All living things respire. - During respiration, carbon and hydrogen containing compounds (glucose) in food are converted into carbon dioxide and water.	Photosynthesis: $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ - Photosynthesis occurs in organisms with chloroplasts, especially plants. - During photosynthesis, plants convert carbon dioxide and water into glucose and oxygen in the presence of sunlight and chlorophyll.
Combustion: - Fuels contain carbon and hydrogen. When they are burned, carbon dioxide and water is produced. - Combustion of natural gas contains mainly methane and has the equation $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{heat energy}$	Ocean uptake: - The world's oceans serve as a massive sink that traps carbon dioxide. - Much of the dissolved carbon dioxide is used by plants and marine organisms. - Most of it is eventually converted into calcium carbonate in the form of shells and skeletons of marine organisms. - Over time, the shells and skeletons of dead organisms settle on the seabed to form limestone.
Decay and decomposition: - Decomposers feed on the remains of dead plants and animals. - These decomposers respire and give out carbon dioxide.	

Global Warming

Global warming is the increase in the Earth's average temperature due to the build-up of greenhouse gases in the atmosphere.

Effects of global warming:

- More occurrences of unusual weather conditions (Heat waves, droughts, floods, tsunamis).
- A decrease in crop yield worldwide because vegetated areas may become deserts.
- Melting of large quantities of ice in the North and South Poles, causing ocean levels to rise and low-lying countries and coastal areas to flood.
- The rapid evaporation of water from the Earth's surface. This will result in the carbon dioxide dissolved in the oceans being released into the atmosphere and further contributing to an enhanced greenhouse effect.