

Chem Notes, Year 11, cont

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#Chemistry

Chemistry, Year 11 Notes:

Version 1.3.2

Relevant Topics:

Unit 1:

- Topic 1, Properties and Structure of Atoms
 - Part 1, Periodic Table and Trends
 - Part 2, Atomic Structure
 - Part 3, Introduction to Bonding
 - Part 4, Isotopes
 - Part 5, Analytical Techniques
- Topic 2, Properties and Structure of Materials
 - Part 1, Compound and Mixtures
 - Part 2, Bonding and Properties
- Topic 3, Chemical Reactions – Reactants, Products and Energy Change
 - Part 1, Chemical Reactions
 - Part 2, Exothermic and Endothermic Reactions
 - Part 3, Measurement and Uncertainty and Error
 - Part 4, Fuels
 - Part 5, Mole Concept and Law of Conservation of Mass

Unit 2:

- Topic 1, Intermolecular Forces and Gases
 - Part 1, Intermolecular Forces
 - Part 2, Chromatography Techniques
 - Part 3, Gases
- Topic 2, Aqueous Solutions and Acidity
 - Part 1, Aqueous Solutions and Acidity
 - Part 2, Identifying Ions in Solution
 - Part 3, Solubility
 - Part 4, pH

- Part 5, Reaction of Acids
- Topic 3, Rates of Chemical Reactions
 - Part 1, Rates of Reactions

Unit 1, Topic 1:

Unit 1, Topic 1, Part 1:

Periodic Table and Trends:

- Definitions and trends:
 - Atomic Radii:
 - Valency:
 - Ionic Radii:
 - 1st Ionisation Energy:
 - Electronegativity:
- Successive ionisation energies:
- Metallic versus non-metallic:
 - The “metallic character” of an element describes the tendency to lose electrons (same thing as electropositivity).
- ???

Unit 1, Topic 1, Part 2:

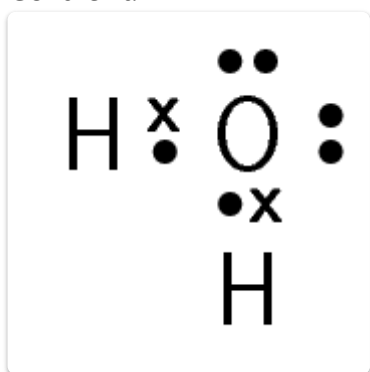
Atomic Structure:

- An atom can be modelled as a nucleus surrounded by electrons in discrete shells held by electrostatic force.
- Notation: *ZAM*
 - A – Mass
 - Z – Atomic number
- Be able to determine the number of protons, neutron and electrons, and the identity of ions and isotopes.
- Orbitals:
 - D
- Be able to write electron configurations (up to Z=36) and use orbital diagrams to represent the character and relative energy of the orbitals (also up spin, down spin stuff).
 - Cr and Cu are exceptions:
 - Cr: $3d^5 4s^1$
 - Cu: $3d^9 4s^2$

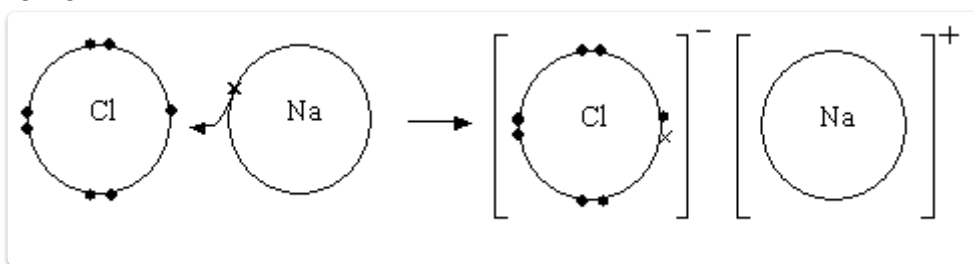
Unit 1, Topic 1, Part 3:

Introduction to Bonding:

- The properties of atoms are derived from their electron arrangements and by the stability of its valence electron shell.
- Bonding is the electrostatic attraction due to the sharing or transfer of electrons from participating atoms.
- Valence electrons are the electrons of the outermost shell. Valency is the measure of an element's capacity to give, accept or share electrons to reach a stable state.
- Ions are formed from losing or gaining electrons (as dictated by its valency). Transition metals can form more than 1 ion.
 - They are charged due to the imbalance of protons (which provide positive charge) and electrons (which provide negative charge)
 - Notated as: Na^+
- Lewis (dot-cross) diagrams:
 - Covalent:



- Ionic



- It should show all valence electrons and up to 4 electron pairs
- Misc stuff:
 - Be able to figure out ion charge based on its position in the periodic table
 - Be able to name covalent and ionic compounds based on their chemical formula
 - Be able to identify bonding and lone pairs in Lewis (dot-cross) diagrams
 - Bonding pairs are electron pairs used in bonding
 - Lone pairs are electron pairs not used in bonding

Unit 1, Topic 1, Part 4:

Isotopes:

- Versions of an elements that have a different number of neutrons, but an equal number of protons; hence a different relative atomic mass.
- It has identical chemical properties (since the protons are the same) but different physical properties (since the mass of different).
- Relative atomic mass is relative to 1/12 of the mass of Carbon-12 (not to Hydrogen-1).
- IUPAC Notation:
 - ^{12}C and ^{14}C
 - C-12 and C-14

Unit 1, Topic 1, Part 5:

Analytical Techniques:

- Not needed lol

Unit 1, Topic 2:

Unit 1, Topic 2, Part 1:

Compounds and Mixtures:

- Pure substances versus mixtures
 - Pure substances are of one element or compound
 - Mixtures are two or more different substances that aren't chemically joined
- Measurable Properties
 - Pure substances (e.g., water) has distinct measurable physical/chemical properties (e.g., boils at 100°C). Mixtures have properties dependant on the identity and relative amount of the substances that make up the mixture.
 - Properties include:
 - Melting and boiling points
 - Electrical and thermal conductivity
 - Solubility
 - Strength
 - Density
 - Reactivity
- Homogenous versus Heterogenous (both are mixtures, not pure substances)
 - Homogenous – mixtures of uniform composition
 - E.g., Salt dissolved in water
 - Heterogenous – mixtures of non-uniform composition
 - E.g., A piece of lead in water
- Nanomaterials (idk why QCAA put it in this section):
 - Defined as substances that contain particles 1-100nm size

- Be able to interpret data to evaluate the physical properties of pure substances and mixtures
 - If a range is presented, it's probably a mixture

Unit 1, Topic 2, Part 2:

Bonding and Properties:

- Properties of Ionic Compounds
 - High melting point
 - Brittleness
 - Able to conduct when dissolved in liquid
- Ionic compounds have a lattice like structure held together by electrostatic force of the ions.

Bond Types & Properties

- Hydrocarbons
 - Alkane
 - No double bonds in the hydrocarbon
 - Alkene
 - A double bond is present in a hydrocarbon
 - Benzene
 - A ring of 6 hydrocarbons
- Be able to:
 - Compare melting/boiling points of various compounds:
 - Comparing the bonding type first:
 - Network covalent -> Ionic -> Metallic -> Simple molecular
 - When comparing between similar types:
 - Simple molecular:
 - Consider the types of intermolecular force
 - Hydrogen -> Dipole-Dipole -> Intermolecular
 - Network covalent:
 - Consider the
 - ????
 - Ionic:
 - Consider the changes of the ions:
 - More change=higher melting/boiling points
 - Consider the atomic size
 - Smaller=higher melting/boiling points
 - Metallic:
 - Consider the changes of the ions:

- More change=higher melting/boiling points
- Consider the atomic size
 - Smaller=higher melting/boiling points

Unit 1, Topic 3:

Unit 1, Topic 3, Part 1:

Chemical Reactions:

- Chemical reactions are indicated by:
 - The production of a new substance.
 - A change in energy (e.g., heat or light)
- Types of chemical reactions:
 - Single Displacement
 - Double Displacement
 - Combustion
 - Decomposition
 - Combination
- Both chemical reactions and phase change involve energy changes; as indicated by temperature or light.
- Be able to balance equations, use state symbols

Unit 1, Topic 3, Part 2:

Exothermic and Endothermic Reactions:

Unit 1, Topic 3, Part 3:

Measurement and Uncertainty and Error:

- Reliability/Precision and Validity/Accuracy:
 - Reliability:
 - The degree of consistency in measurement
 - Validity:
 - The degree the data accurately reflects the concept being measured
- Random Versus Systematic Error:
 - Random:
 - Random statistical fluctuation in the data, not necessarily reoccurring.
 - Reduces reliability/precision
 - Systematic:
 - Biases in measurement that cause data to differ by a consistent amount
 - Reduces validity/accuracy

Most of the formulas are on the formula sheet anyway. Ones that aren't:

- Absolute uncertainty of a measurement
 - Least count/2

Unit 1, Topic 3, Part 4:

Fuels:

Experimentally calculating enthalpy of different fuels:

- m = mass of water w
- c = specific heat capacity ($4.18\text{J/g}^\circ\text{C}$) w
- assumes no heat loss

Allows the comparison of multiple fuels. Can be used in as the enthalpy in chemical equations.

Ask about sustainability and energy generation types (which is on the textbook).

Unit 1, Topic 3, Part 5:

Mole Concept and Law of Conservation of Mass:

- Concepts
 - Law of conservation of mass
 - States that mass is maintained throughout the reaction
 - Empirical formula:
 - A formula giving the proportions of the elements in a compound in simplest form
 - Limiting Reactant:
 - The reactant which is fully consumed by the reaction; and determines the amount of product produced.
 - Misc:
 - Be able to use stoichiometric ratios
 - Be able to determine yields, mass and moles using these ratios
 - Be able to differentiate theoretical and actual yield
- Avogadro's number:
 - 6.02×10^{23}
- Percentage Composition:
 - $\%age\ Composition = \frac{\text{Mass}}{\text{Total Mass}} \times 100\%$
- Percentage Yield:
 - $\%age\ Yield = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$
- Determining Empirical Formula:

Atomic Calculations

- Determining Molecular Formula (using Empirical Formula):
 - $\text{Molar mass} = \text{Empirical Molar Mass} \times n$
 - Then multiply every element's number by n

Mandatory Practical:

- Determine the empirical formula of a compound from reactions involving mass change.

Unit 2, Topic 1:

Unit 2, Topic 1, Part 1:

Intermolecular Forces:

- Valence Shell Electron Pair Repulsion (VSEPR) theory
 - States that electrons in the valence shell form pairs (up and down), which repel other electrons.
- Shapes of molecules
 - The shape of a molecule depends on the number of lone pairs and the number of bonded atoms.
 - A lone pair has more repulsion than an atom.

Bonding Shapes

- a short cut allows you to figure out the shape of the molecule based on the group
- Polarity
 - Symmetry
 - If a molecule is symmetrical, it is non-polar
 - Electronegativity
 - If the difference in the electronegativity of two covalent molecules is between 0.4 and 1.7, the bond is polar covalent (forms dipoles)
 - If the molecule is non-symmetrical and has polar covalent bonds, it is a polar molecule
 - Diagrams make it a whole lot easier
- Types of intermolecular forces (in order of strength):
 - Dispersion
 - Intermolecular bonds formed from the random distribution of electrons, causing temporary dipoles (which interact with each other).
 - Dipole-Dipole
 - Intermolecular bonds formed by the interaction of polar molecules
 - Hydrogen bonding

- Stronger dipole-dipole forces between hydrogen and either nitrogen, oxygen or fluorine
- Observable properties and intermolecular forces
 - Vapour pressure
 - The stronger the intermolecular forces, the weaker the vapour pressure is (and the easier it is to evaporate it)
 - Melting/boiling point
 - Stronger intermolecular forces are harder to separate and require more energy. Thus, molecules with stronger intermolecular forces will have a higher melting/boiling point.
 - Solubility
 - Like dissolves like
 - Solubility relies on the ability of a solvent to overcome the intermolecular forces of the solute. Stronger intermolecular forces are harder to separate.
- Mandatory Practical:
 - Construct 3D models of different molecular shapes

Unit 2, Topic 1, Part 2:

Chromatography Techniques:

- Definitions:
 - Chromatography:
 - A technique is used to separate mixtures, identify substances, and determine the composition/purity of a mixture/sample.
 - Mobile Phase:
 - A substance that flows through a chromatographic system, separating the sample over the stationary phase.
 - Stationary Phase:
 - The phase over which the sample and mobile phase moves across, separating the sample into its compounds
- There are 4 main techniques:
 - Paper chromatography:
 - Through capillary action, the solvent will travel up the stationary phase (the paper), bringing the sample along with it. Compounds that are more attracted to the mobile phase (the solvent) will be dragged further, separating from other compounds.
 - The competition between the phases pulling at the components drives the separation between the parts of the sample (since the polarity of the compounds affects how much they are pulled).
 - Thin-layer chromatography (TLC):
 - Paper chromatography but better lol (faster and with more separation).

- Still relies on capillary action
- Still relies on a difference in polarity (more polar stuff goes further)
- Same process but paper is replaced with a glass slide with a layer of silica gel
 - The silica gel is very polar (more than the solvent). Therefore, compounds that are more polar are not dragged as far on the plate.
 - The solvent must have the ability to be able to dissolve the components of the mixture
- The position of each separated dot is measured using 'retardation factor' (R_f)
 - $R_f = \frac{\text{Distance traveled by solute}}{\text{Distance traveled by solvent}}$
 - Since the solute cannot travel further than the solvent, R_f is never greater than 1.
 - If R_f values are too similar, different solvent mixtures and stationary phases can be tried.
 - It can be used to identify the substance too.
- Gas chromatography (GC):
 - A separation technique used for small organic molecules that can withstand high temperatures (e.g., alcohol).
 - It works by:
 - An inert gas (called the eluant/carrier gas) is used to push the sample through a column in an oven. The gas (e.g., nitrogen) is the mobile phase.
 - The sample is injected into the carrier gas and evaporates (due to the pressure)
 - The column is a capillary tube packed with silica that is coated with liquid (the liquid acts as the stationary phase).
 - In the column, the different compounds in the sample are separated
 - Smaller and/or less polar particles leave the column first
 - Larger and/or more polar particles leave the column last
 - At the end of the column, a detector detects and records them.
 - The time it takes a sample to elute is measured by 'retention time' (R_t)
 - This is used to identify the compound
 - GC can also measure the amount of analyte. The amount is reflected by the height of the peak on a gas chromatography reading (higher=more).
 - The exact amount can be determined by finding the area of each triangle-like reading and multiplying it by a proportionality constant (which is determined by the specific gas chromatograph and the measurement taken).
 - GC graphs:
 - x-axis is the time taken for the analytes to reach the detector
 - y-axis is intensity/response (just means the amount of that analyte)
- High-performance liquid chromatography (HPLC):

- Used on larger organic molecules or substances that can't be heated (and hence, cannot be used in GC)
- Very similar to GC
- It works by:
 - A pump is used to move a liquid (the mobile phase) into a column (stationary phase)
 - The compounds of the sample separate in the column
 - A detector measures the time for each analyte to elute.
- Similar to GC, *Rt* can also be used to determine the amount and identity of analytes

Chromatography

- more techniques exist but these are the ones the syllabus mentions

Unit 2, Topic 1, Part 3:

Gases:

- In the same conditions (number of moles, pressure, temperature), all gases will inhabit the same amount of volume.
 - $n \text{ (mol)} = \frac{\text{Volume (L)}}{\text{Molar Volume (L/mol)}}$

Standard Conditions

- "Room Temperature and Pressure" (RTP) and SLC are interchangeable.
- Kinetic Theory of Gases:
 - The size of the particles is negligible compared to the size of the container
 - The intermolecular forces between molecules are negligible
 - The collisions between molecules are perfectly elastic (no kinetic energy lost)
 - The particles move in a straight line randomly until they collide with other particles or a wall
 - Pressure is caused by collisions of the particles with the wall
 - Temperature is a measure of the average kinetic energy of the molecules.
- A theoretical gas that obeys these laws perfectly is an **ideal gas**.
 - The opposition of an ideal gases is usually called a **real gas**.
 - Particles have volume and intermolecular forces between each other
 - To make it similar to an ideal gas:
 - Lower the pressure
 - Raise the temperature
 - Use non-polar particles
- At absolute 0, there is not kinetic energy (and hence velocity) in the particles
- Gas laws:
 - Boyle's Law:

- $V \propto \frac{1}{P}$
- $P_1V_1 = P_2V_2$
- Charles' Law:
 - $V \propto T$
 - $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
 - Degrees Kelvin must be used in Charles' Law and any laws that derive from it.
- These laws combined make the combined gas law:
 - $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
 - Use degrees kelvin
- All above gas laws rely on at least 1 variable being constant. The ideal gas law is a general equation that can be used instead:
 - $PV = nRT$
 - Pressure [kPa]
 - Volume [L]
 - n [mol]
 - Temperature [*K]
 - R is a constant (8.314 J K⁻¹ mol⁻¹)
 - Real gases will slightly deviate though
- Mainly able to do problems with gas

Unit 2, Topic 2:

Unit 2, Topic 2, Part 1:

Aqueous Solutions and Molarity:

- Key Terms:
 - **Solute** – A component dissolved into a solvent
 - **Solvent** – A component that a solvent is dissolved into
 - **Solution** – A homogenous mixture of two or more substances
 - **Concentration** – The abundance/density of a solute in a solvent
 - **Molarity** – Concentration measured as mols/litre
 - **Mass Concentration** – Concentration measured as grams/litre
 - **Parts per Million (ppm)** – Concentration measured as a ratio of solute mass to solvent mass
 - Different from the others as their volume includes the solvent
- Unique Properties of Water
 - **Hydrogen Bonding** – The oxygen and hydrogen bonds facilitate hydrogen bonding.
 - **As a result:**
 - **Boiling Point** – For a small covalent compound, water has a high boiling point

- **Surface Tension** – Water has a high surface tension
- **Solvent** – Water is a strong polar solvent
- **Density** – Water has the highest density at 4°C, since it forms a lattice structure
- Solvent saturation:
 - **Unsaturated**: Below full solvent concentration
 - **Saturated**: At maximum solvent concentration
 - **Supersaturation**: Above full saturation (by saturating a solution and cooling it)
- Spectator ions are delocalised ions in the chemical reactions that aren't exchanged.
 - If the question specifies for a **total** ionic equation, include them (put the ratios in front too).
 - If the question specifies for a **net** ionic equation, don't include them.
- Make sure you can calculate/predict moles, concentrations and saturation stuff using formulas

Unit 2, Topic 2, Part 2:

Identifying Ions in a Solution:

- Solubility Rules:
 - In water:
 - Soluble (aq):
 - All group I (and ammonium)
 - sodium, potassium and ammonium
 - (no exceptions)
 - All group VII (or just these three at least)
 - chlorides, bromides and iodides
 - (except with lead and silver)
 - All nitrates and acetates
 - NO₃ and CH₃COO
 - (AgCH₃COO is only slightly soluble)
 - All sulphates
 - SO₄
 - (except with lead, barium, strontium, silver and calcium)
 - Insoluble (s):
 - All carbonates, phosphates and sulphites
 - CO₃, PO₄, SO₃
 - (except group I and ammonium),
 - All sulphides
 - S
 - (except group I, group II and+ ammonium)
 - All oxides and hydroxides

- O, OH
- (except with sodium, potassium, ammonium, strontium and barium)
- Also, Calcium hydroxide (limewater) is partially soluble
- From QCAA Formula and Data Book:

Solubility of selected compounds at 298 K

	bromide	carbonate	chloride	hydroxide	iodide	nitrate	oxide	phosphate	sulfate
aluminium	s	—	s	i	s	s	i	i	s
ammonium	s	s	s	s	s	s	—	s	s
barium	s	i	s	s	s	s	s	i	i
calcium	s	i	s	p	s	s	p	i	p
cobalt(II)	s	i	s	i	s	s	i	i	s
copper(II)	s	—	s	i	i	s	i	i	s
iron(II)	s	i	s	i	s	s	i	i	s
iron(III)	s	—	s	i	s	s	i	i	s
lead(II)	p	i	s	i	i	s	i	i	i
lithium	s	s	s	s	s	s	s	—	s
magnesium	s	i	s	i	s	s	i	p	s
manganese(II)	s	i	s	i	s	s	i	p	s
potassium	s	s	s	s	s	s	s	s	s
silver	i	i	i	i	i	s	i	i	p
sodium	s	s	s	s	s	s	s	s	s
zinc	s	i	s	i	s	s	i	i	s

Key:

Abbreviation	explanation
s	soluble in water (solubility greater than 10 g L ⁻¹)
p	partially soluble in water (solubility between 1 and 10 g L ⁻¹)
i	insoluble in water (solubility less than 1 g L ⁻¹)
—	no data

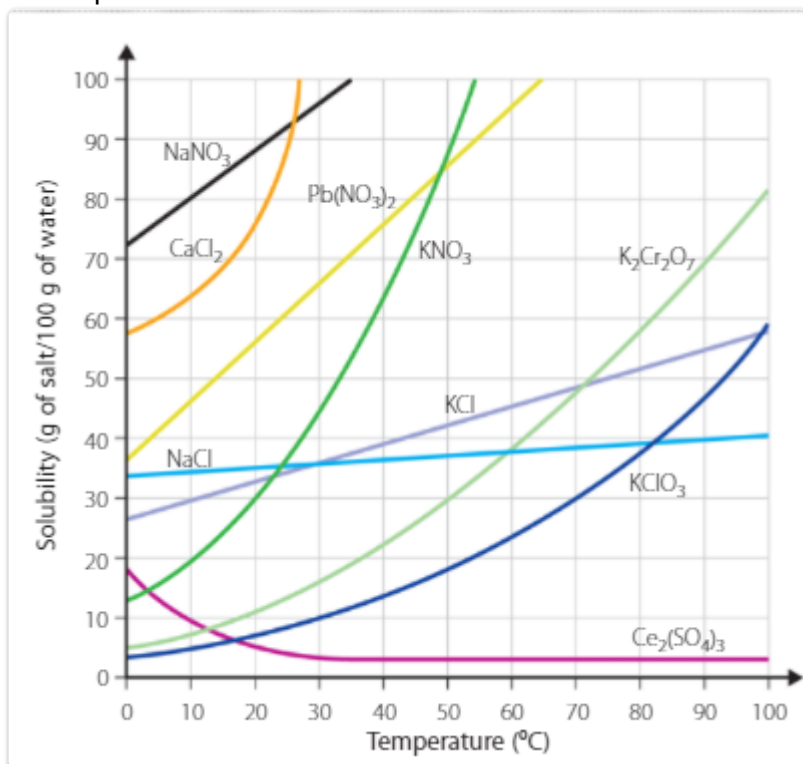
- Determining Specific Ions:
 - Positive Ions (Cations):

- Copper (II) with sodium hydroxide (NaOH) turns blue
- Iron (II) with sodium hydroxide (NaOH) turns green
- Iron (III) with sodium hydroxide (NaOH) turns brown
- Negative Ions (anions):
 - Carbonate (CO₃²⁻) ions:
 - Identified with dilute hydrochloric acid (HCl).
 - If carbonate ions are present, fizzling will be present (CO₂)
 - Since both the hydrochloric acid and the carbonate ions are delocalised, they can be expressed as the following:
 - Example: $\text{CO}_3^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
 - Limewater can be used to ensure the gas is CO₂.
 - Sulphate (SO₄²⁻) ions:
 - Identified with barium chloride (BaCl₂). Hydrochloric acid must be added first (HCl) to acidify the solution.
 - If sulphate ions are present, a white precipitate (BaSO₄) is formed.
 - [Apparently] The hydrochloric acid is added to ensure that any carbonate ions that might be present are removed; since they also form a white precipitate, giving a false positive (¯_(ツ)_/¯).
 - Example (sodium): $\text{Na}_2\text{SO}_4(\text{aq}) + \text{BaCl}_2(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{NaCl}(\text{aq})$
 - Halide (Group VII) ions:
 - Identified with silver nitrate (AgNO₃). Nitric acid is first added to acidify the solution.
 - If halides are present, a precipitate will form:
 - The exact halide can be determined by its colour:
 - Chloride: White
 - Bromide: Cream
 - Iodide: Yellow
 - [Apparently] The nitric acid is added to ensure that any carbonate ions that might be present are removed; since they also form a precipitate, giving a false positive (¯_(ツ)_/¯).
 - Example (Chloride): $\text{NaCl}(\text{aq}) + \text{AgNO}_3(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$
- Be able to write ionic equations with formulas and phases; and use them to make predictions.
- Mandatory Practical:
 - Use precipitation reactions to identify cations and anions
 - This practical just systematically tries 6 different cations and 4 different anions and finds trends between them.

Unit 2, Topic 2, Part 3:

Solubility:

- Energy is used to separate the ionic bonds, but energy is released when ions bond with water. For a substance to be soluble, the energy released from bonding must be greater than the energy required to separate the solute ions.
 - Ion-dipole bonds form between water molecules and solute ions.



- Factors that affect solubility by:
 - Changing energy required to separate ions:
 - Strength of the bond
 - Arrangement (e.g., lattices are hard to separate)
 - Sizes of ions
 - Charges of the ions
 - Changing energy released from ion-dipole bonds:
 - Strength of ion-dipole bonds
 - Geometry of the ions (e.g., polyatomic)
 - Sizes of ions
 - Charges of the ions
- While organic molecules are usually non-polar (since they are covalent), if enough polar groups are present (e.g., -OH , C=O , -NH), it increases solubility. If more carbon and hydrogen atoms are present, it becomes less soluble (hydrophobic). Additionally, some organic molecules fully dissociate ions, having a significantly higher solubility (e.g., ethanoic acid)
 - For example, ethanol and glucose are slightly soluble
- Temperature increases the internal energy, increasing the energy of collisions, increasing solubility

- Solubility curves can be used to determine the amount of solvent able to dissolved at a given temperature (for a specific solution)
- Gas solubility decreases with temperature as the additional energy allows the gas to overcome the dipole forces of the water and return to the atmosphere.
- Be able to analyse solubility curves

Unit 2, Topic 2, Part 4:

pH:

- pH can be considered using the Arrhenius model, in which pH is dependant on the concentration of H⁺ ions.
 - Similarly, a base/alkaline can be measured by OH⁻ concentration.
- The pH scale is used to measure acidity
 - It is logarithmic by factors of 10
 - It is neutral at 7
 - More than 7 being alkaline, less than 7 being acidic
- A strong acid/base will dissociate more H⁺ (or OH⁻) ions per molecule.
 - This can be reflected by electrical conductivity.
 - The hydrogen must be bonded through ionic bonding to be dissociated
- Be able to compare acidity using an indicator (or multiple indicators).
- Mandatory Practical:
 - Examine the strengths of strong and weak acids:
 - Using an indicator
 - Using and electronic pH meter
 - Using an electric current

Unit 2, Topic 2, Part 5:

Reactions of Acids:

- General formulas:
 - Acid and Base (aka; neutralisation reaction):
 - Acid + Base/MetalOxide -> Salt + Water
 - (There are exceptions, e.g., ammonia as a base doesn't form water)
 - Acid and Metal (aka; a corrosion reaction):
 - Acid + Metal -> Salt + Hydrogen Gas
 - Acid and Carbonate:
 - Acid + Carbonate -> Salt + Carbon Dioxide + Water
 - Acid and Hydrogen Carbonate/Bicarbonate:
 - Acid + Hydrogen Carbonate -> Salt + Carbon Dioxide + Water
- Constructing and representing reactions (notation stuff):

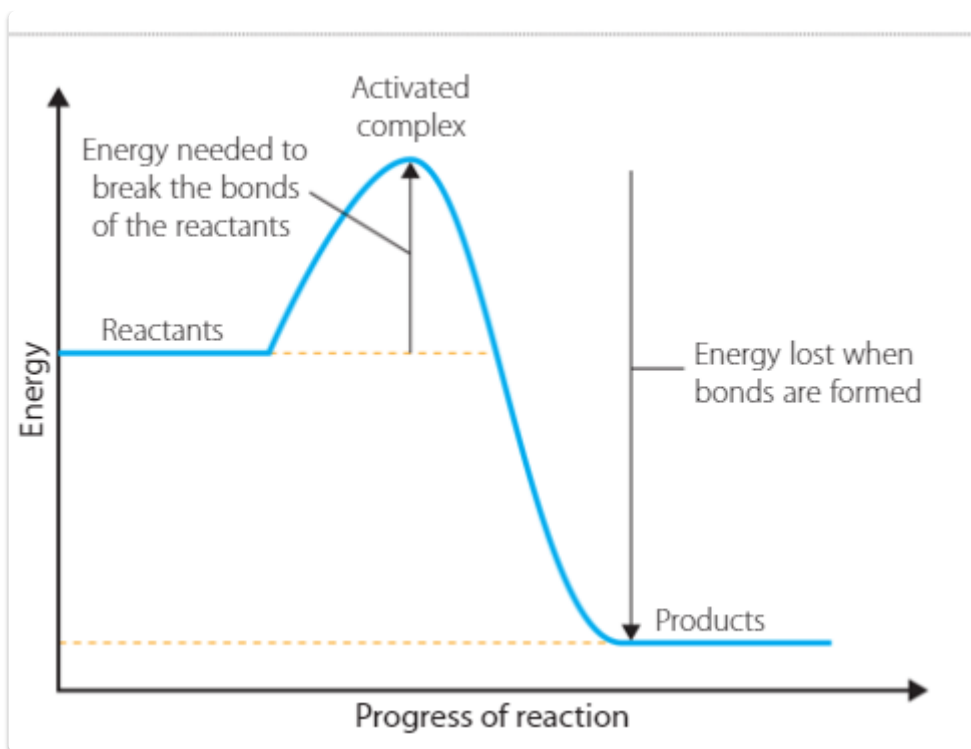
- Neutral species equation:
 - Ignore spectator ions
 - E.g., $\text{HCl (aq)} + \text{NaOH (aq)} \rightarrow \text{NaCl (aq)} + \text{H}_2\text{O (l)}$
- Complete ionic equation:
 - Show spectator ions
 - E.g., $\text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)} + \text{Na}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)} \rightarrow \text{Na}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)} + \text{H}_2\text{O (l)}$
- Net Ionic Equation:
 - Remove spectator ions
 - E.g., $\text{H}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)} \rightarrow \text{H}_2\text{O (l)}$

Unit 2, Topic 3:

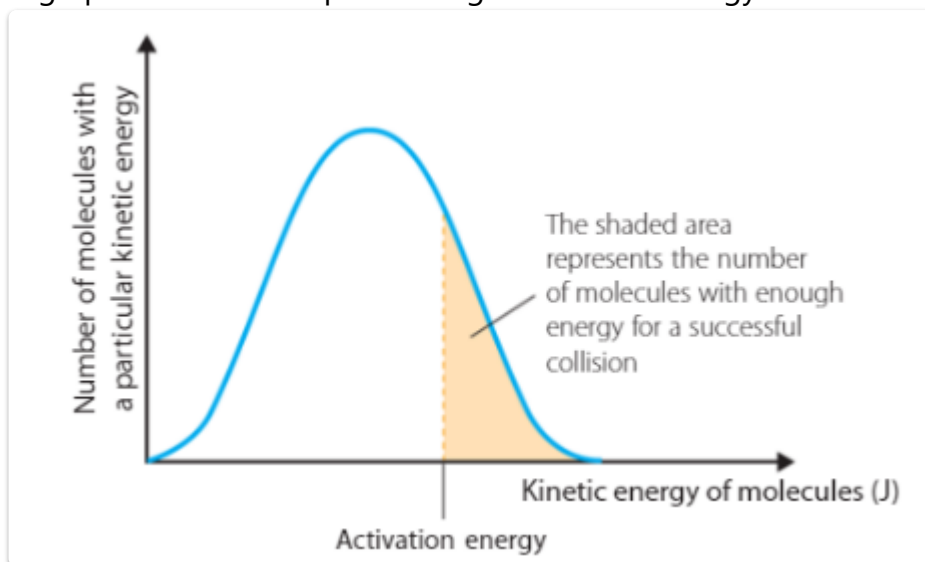
Unit 1, Topic 3, Part 1:

Rates of Reaction:

- Increase rate of reaction by:
 - Increasing temperature
 - Increasing surface area
 - Increasing pressure (in a gaseous system)
 - Increasing concentration
 - Adding a catalyst
- These are justified by collision theory, which states that reactants must collide with a sufficient amount of energy to react.
 - The initial required energy is the activation energy, which is dependent on the strength and number of the existing chemical bonds.
- Energy profile diagrams:
 - Shows the activation energy and energy released from the bonds



- Can be used to determine if a reaction is exothermic or endothermic
- A catalyst provides an "alternative reaction pathway" with a lower activation energy, increasing the proportion of collisions that lead to chemical change.
- Metal nanoparticles can also be used as a catalyst
- Maxwell-Boltzmann curves:
 - A graph of number of particles against kinetic energy.



- Looks like a normal distribution.
- The activation energy will be a point on the x axis
 - Particles with more energy will react
 - A catalyst translates the activation energy to the left.
- Measuring rate of reaction:
 - Rate of formation
 - Mass of product formation/time.

- Volume of product formation/Time
- Rate of depletion
 - Mass of reactant depletion/time
 - Volume of reactant depletion/time
- Be able to analyse experimental data:
 - Be able to calculate change in concentration
 - Be able to graph volume or mass against time
- Mandatory Practical:
 - Investigating rate of reaction:
 - Place a beaker with CaCO_3 and HCl on a scale and graph mass with time. (since one of its reactants, CO_2 is gaseous and will escape)

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Active Recall:

Recreate the portions you've done as questions, painful one if u want

Active Recall:

Unit 1, Part 1, Topic 1, Periodic Table and Trends:

- Regurgitate periodic trends and why it is the case:
 - Properties:
 - Atomic Radii
 - Valencies
 - Ionic Radii
 - 1st ionisation energy
 - Electronegativity
 - Reactivity (where applicable)
 - Group trends (where applicable)
 - Where:
 - Across a period
 - Across a row
- Be able to define each of the above as well
- Explain and compare metallic properties of elements (also define it)
- Successive ionisation energy and electron configuration
- Oxides amphoteric thing
- Interpreting data

Unit 1, Part 1, Topic 2, Atomic Structure:

Unit 1, Part 1, Topic 3, Introduction to Bonding:

- Why do atoms bond?
- Why do ions come about?
- Valency definition
- Types of bonds and why (and in which molecules)
-

Unit 1, Part 1, Topic 4, Isotopes:

- Notation for isotopes? (both ways)
- Chemical and physical properties? (in relation to other isotopes)
- Justify the above and why?
- Electron configuration?
- What is relative atomic mass relative to?

Unit 1, Part 1, Topic 5, Analytical Techniques:

- Explain mass spectrometry, what it does and how
- Explain the flame test, what it does and how
- Explain absorption and emission spectra, what it does and how
- Evidence for discrete energy levels, in absorption and emission spectra
- Explain energy level photon stuff
- Know how to interpret flame test/AAS for metallic ions
- Know how to calculate stuff with relative atomic mass and isotopes (e.g., average based on mass and abundancy)

Unit 1, Part 2, Topic 1, Compounds and Mixtures:

- Types of pure substances?
- Measurable properties?
- Examples of properties?
- Heterogenous versus homogenous?
- Size range for nanoparticles?
- Specific properties of nanoparticles?
- How to interpret data about pure/non-pure substances
- check active recall stuff

Unit 2, Topic 1, Part 2, Chromatography Techniques: