OCR (A) Chemistry A-Level

Module 3 - Periodic Table and Energy

Definitions and Concepts

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Module 3 - Periodic Table and Energy

3.1.1 - Periodicity

Atomic (Proton) Number: the number of protons in the nucleus of an atom.

Bohr Model: describes an atom as a small dense nucleus with electrons orbiting around the nucleus. This model explains different periodic properties of atoms.

Cations: positively charged ions.

d-block: the part of the periodic table in which the elements have their highest energy electron in a d-orbital.

Electron Configuration: the arrangement of electrons into orbitals and energy levels around the nucleus of an atom / ion.

First Ionisation Energy: the removal of one mole of electrons from one mole of gaseous atoms. Factors which affect the first ionisation energy are: the strength of attraction between the electron and the nucleus, the nuclear charge and the atomic radius. There is a small decrease in first ionisation energy due to s- and p-subshell energies (between Be and B) and p-orbital repulsion (between N and O).

Giant Covalent Lattice: a network of atoms bonded by strong covalent bonds (e.g. carbon (diamond, graphite and graphene) and silicon). Giant covalent lattices typically insoluble with a high melting and boiling point due to the presence of strong covalent bonds. They are also poor electrical conductors as they don't contain mobile charged particles.

Giant Metallic Lattice Structure: the structure of all metals, made up of cations and delocalised electrons. Giant metallic structures are typically insoluble with a high melting and boiling points due to strong electrostatic forces of attraction between cations and electrons. Metals are good electrical conductors due to the presence of delocalised electrons (mobile charges).

Group: a column in the periodic table.

Melting Point: the temperature at which a solid melts and becomes a liquid. This increases from giant metallic to giant covalent structures then decreases to simple molecular structures.

Metallic Bonding: strong electrostatic attraction between cations and delocalised electrons.

p-block: the part of the periodic table in which the elements have their highest energy electron in a p-orbital.

Period: a row in the periodic table.

Periodicity: a repeating trend in physical and chemical properties across the periods of the periodic table.

s-block: the part of the periodic table in which the elements have their highest energy electron in an s-orbital.

Successive Ionisation Energies: the energy required to remove each electron one-by-one from one mole of gaseous atoms / ions.

3.1.2 - Group 2

Base: a substance that can accept H^+ ions from another substance. Group 2 compounds can be used as bases: $Ca(OH)_2$ is used to neutralise acidic soils in agriculture and $Mg(OH)_2$ and $CaCO_3$ are used as antacids to treat indigestion.

Electron Configuration: the arrangement of electrons into orbitals and energy levels around the nucleus of an atom/ ion. Group 2 elements have an s^2 outer shell electron configuration.

First Ionisation Energy: the removal of one mole of electrons from one mole of gaseous atoms. Factors which affect the first ionisation energy are: the strength of attraction between the electron and the nucleus, the nuclear charge and the atomic radius.

Group 2 Oxide: a compound with the general formula MO, where M is a group 2 element. When group 2 oxides react with water, they form an alkaline solution, with alkalinity increasing down the group.

Oxidation: the loss of electrons / increase in oxidation number.

Redox: a reaction in which oxidation of one element and reduction of another occurs. During a redox reaction involving group 2 elements, 2 electrons are lost to form 2+ ions. Group 2 elements undergo redox reactions with water, oxygen and dilute acids.

Reduction: the gain of electrons/ decrease in oxidation number.

Second Ionisation Energy: the removal of one mole of electrons from one mole of gaseous 1+ ions to form one mole of 2+ ions.

3.1.3 - The Halogens

Boiling Point: the temperature at which a liquid boils and becomes a gas. Boiling point increases down group 7 due to the increasing strength of London Forces between the halogen molecules.

Diatomic Molecules: molecules that are made up of 2 atoms. Halogens are diatomic.

Displacement Reaction: a reaction in which one atom is replaced by another. Halogens can undergo displacement reactions as their reactivity decreases down the group. The more reactive halogen will displace the less reactive halogen from a solution of its salt.

Disproportionation: the oxidation and reduction of the same element. Examples include the water treatment (reacting chlorine with water) and bleach formation (reacting chlorine with cold, dilute aqueous sodium hydroxide).

Electron Configuration: the arrangement of electrons into orbitals and energy levels around the nucleus of an atom / ion. The halogens have a s^2p^5 outer shell electron configuration.

Induced Dipole-Dipole Interactions: forces of attraction between molecules caused when the random movement of electrons creates a temporary dipole in one molecule which then induces a dipole in a neighbouring molecule.

London Forces: another name for induced dipole-dipole interactions.

Oxidation : the loss of electrons / increase in oxidation number.

Precipitation Reaction: a reaction in which two aqueous solutions are combined to form an insoluble salt (a precipitate). Halide anions undergo precipitation reactions with aqueous silver ions.

Redox reaction: a reaction in which oxidation of one element and reduction of another occurs. During a redox reaction involving group 7 elements, 1 electron is gained to form 1- ions.

Reduction: the gain of electrons/ decrease in oxidation number.

Water Treatment: the addition of chlorine to water to kill bacteria. The risks associated with the use of chlorine to treat water are the hazards of toxic chlorine gas and the possible risks from the formation of chlorinated hydrocarbons.

3.1.4 - Qualitative Analysis

Anion: a negatively charged ion.

Ammonium Ion: an ion with the formula NH_4^+. The test for ammonium ions is a reaction with warm NaOH, which forms NH_3.

Cation: a positively charged ion.

Carbonate: a salt containing the CO_3^{2-} anion. A reaction between a carbonate and H^+ will form $CO_2(g)$.

Halide: a salt containing a group 7 anion. Cl^-, Br^- and I^- can be tested for using a solution of silver ions as this reaction forms a coloured precipitate. The solubility of the precipitate is then tested using dilute and concentrated ammonia.

Qualitative Analysis: identifies the elements present in a substance, typically using test-tube reactions.

Sulfate: a salt containing the anion SO_4^{2-}. A reaction between SO_4^{2-} and $Ba^{2+}(aq)$ will form a precipitate.

3.2.1 - Enthalpy Changes

Activation Energy: the minimum energy required for a reaction to take place.

Average Bond Enthalpy: the energy required to break one mole of gaseous bonds. Actual bond enthalpies may differ from the average as the average bond enthalpy considers a particular bond in a range of molecules.

Endothermic: a reaction which takes in energy (ΔH is positive). More energy is required to break bonds than is released by making bonds.

Enthalpy (H): a value that represents the heat content of a system.

Enthalpy Change (ΔH): the change in the heat content of a system during a reaction. This can be determined from experimental results using $q = mc\Delta T$ (where q is the heat change of the surroundings, m is the mass of the surroundings, c is the specific heat capacity and ΔT is the change in temperature).

Enthalpy Change of Combustion ($\Delta_c H$): the enthalpy change that takes place when one mole of a substance is completely combusted.

Enthalpy Change of Formation ($\Delta_f H$): the enthalpy change that takes place when one mole of a compound is formed from its elements.

Enthalpy Change of Neutralisation ($\Delta_{neut}H$): the enthalpy change that takes place when one mole of water is formed from a neutralisation reaction.

Enthalpy Change of Reaction ($\Delta_r H$): the enthalpy change that is associated with a particular chemical equation.

Enthalpy Profile Diagram: shows the difference in the enthalpy of reactants and products as well as the activation energy of a reaction.

Exothermic: a reaction which gives out energy (ΔH is negative). More energy is released by bond making than is used in bond breaking.

Hess' Law: the enthalpy change of a reaction is independent of the route it takes.

Standard Conditions: a pressure of 100 kPa and a temperature of 298K.

Standard State: the physical state (s, l, g, aq) of a substance under standard conditions.

3.2.2 - Reaction Rates

Boltzmann Distribution: a graph showing the distribution of the energies of molecules in relation to the activation energy. Increasing temperature will increase the proportion of molecules with energy above the activation energy. A catalyst lowers the activation energy meaning more molecules will have sufficient energy to react.

Catalyst: a substance that speeds up the rate of a reaction without being used up. A catalyst allows the reaction to proceed via a different route with a lower activation energy. Catalysts are important in terms of economics and sustainability as they enable processes to take place at lower temperatures meaning less energy is required (this decreases the combustion of fossil fuels so CO_2 emissions are reduced).

Collision Theory: the theory which states that molecules must collide with sufficient energy at the correct orientation for a reaction to occur.

Concentration: the amount of a substance that is dissolved per unit volume of solution. Increasing the concentration increases the rate of reaction as there are more molecules in the same volume meaning more frequent successful collisions.

Enthalpy Profile Diagram: shows the difference in the enthalpy of reactants and products as well as the activation energy of a reaction.

Heterogeneous Catalyst: a catalyst that is in a different state to the reactants (e.g. a solid catalyst with gaseous reactants).

Homogeneous catalyst: a catalyst which is in the same state as the reactants.

Pressure: the force that a gas exerts on the walls of a container. Increasing the pressure increases the rate of reaction as there are more molecules in the same volume (or the same number of molecules in a smaller volume) meaning more frequent successful collisions.

Rate of Reaction: a measure of how quickly a reactant is used up/ a product is formed. Rate can be determined by measuring concentration, the volume of gas produced or the mass lost over time.

3.2.3 - Chemical Equilibrium

Catalyst: a substance that increases the rate of a reaction without being used up. In an equilibrium, the rates of both the forwards and reverse reactions are increased meaning the position of equilibrium is unchanged. A catalyst will increase the rate at which the equilibrium is established.

Compromise: balancing several different factors in order to get the best possible outcome. In industry, the position of the equilibrium and rate of reaction must be considered when deciding the conditions for the reaction. These factors must also be balanced with safety and economics.

Dynamic Equilibrium: a closed system in which the rates of the forward and reverse reactions are equivalent. The concentrations of reactants and products don't change.

Homogeneous Equilibrium: an equilibrium in which all reactants are in the same state.

Le Chatelier's Principle: when a system in dynamic equilibrium is subject to change, the position of equilibrium will shift to minimise the change. This principle is used to determine the effect of changing pressure, temperature or concentration on the position of equilibrium.

K_c: the equilibrium constant that is equal to the concentration of products raised to their stoichiometric coefficients divided by the concentration of reactants to the power of their stoichiometric coefficients. When K_c is greater than 1, the equilibrium favours the products. When K_c is less than 1, the equilibrium favours the reactants.