

WJEC (Wales) Chemistry A-level

Unit 3: Physical and Inorganic Chemistry Definitions and Concepts

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Definitions and Concepts for WJEC (Wales) Chemistry A-level

Unit 3: Physical and Inorganic Chemistry

3.1 - Redox and Standard Electrode Potential

Anode: Positive electrode where oxidation takes place.

Cathode: Negative electrode where reduction takes place.

Cell potential: A measure of the potential difference between two half cells, calculated by combining two standard electrode potentials. The calculated cell potential can be used to predict the feasibility of a reaction, although this doesn't consider concentration or kinetics.

E.M.F: Electromotive force, measured in volts. The difference between the potential differences of the cathode and anode in an electrochemical cell.

Electrochemical cell: Combination of two half cells, consists of two electron conductors (electrodes) separated by an ionic conductor (electrolyte). Cells are used to measure electrode potentials by reference to the standard hydrogen electrode.

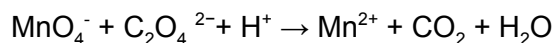
Electrochemical series: A series of elements arranged in order of their standard electrode potentials.

Electrode: A conductor through which electricity enters or leaves the electrolyte in an electrochemical cell.

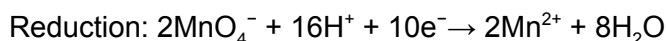
Electrode potential: The potential difference of a cell consisting of a specific electrode as the cathode and the standard hydrogen electrode as the anode. Reduction always takes place at the cathode, and oxidation at the anode.

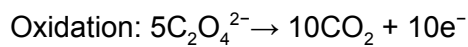
Feasible reaction: For a reaction to be feasible at a given temperature it must occur spontaneously. This means no extra energy is required for the reaction to occur.

Half equation: A full redox equation can be split into two half-equations, one involving oxidation and the other involving reduction. This is useful for balancing complex redox reactions, such as:

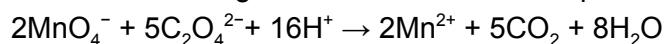


Can be split into:





And combined to give the balanced redox equation:



Hydrogen fuel cell: An electrochemical cell that converts the chemical energy of the fuel (hydrogen) and the oxidising agent (oxygen) into electricity through redox reactions. They do not need to be recharged but require a constant supply of fuel to maintain the potential difference.

Oxidation state: The charge of an ion or a theoretical charge of an atom in a covalently bonded compound assuming the bond becomes ionic.

Redox reaction: A reaction in which reduction and oxidation occur simultaneously.

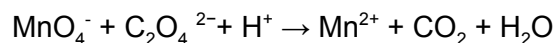
Standard conditions: These conditions are: solutions of 1.0 mol dm^{-3} concentration, a temperature of 298K and 100 kPa pressure.

Standard electrode (redox) potential (E^\ominus): The EMF of a half cell compared with a standard hydrogen half cell. This is measured under standard conditions (1 mol dm^{-3} concentrations, 298K temperature and 1 atm pressure).

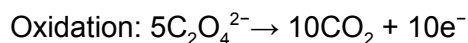
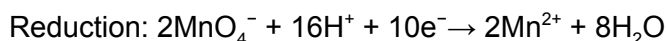
Standard Hydrogen Electrode (SHE): The standard hydrogen electrode is the measuring standard for half-cell potentials. It has a cell potential of 0.00V, measured under standard conditions.

3.2 - Redox reactions

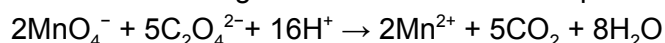
Half equation: A full redox equation can be split into two half-equations, one involving oxidation and the other involving reduction. This is useful for balancing complex redox reactions, such as:



Can be split into:



And combined to give the balanced redox equation:





Oxidation: Process involving the loss of electrons. Results in an increase in oxidation number.

Redox reaction: A reaction in which both reduction and oxidation are occurring simultaneously.

Reduction: Process involving the gain of electrons. Results in a decrease in oxidation number.

Titration: The addition of a solution with a known concentration to a solution with a known volume and an unknown concentration until the reaction reaches neutralisation. This is often indicated by the colour change of an indicator.

3.3 - Chemistry of the p-block

Amphoteric: A substance that is able to react as both an acid and a base.

Disproportionation reaction: A reaction in which a species is both oxidised **and** reduced, seen as both an increase and a decrease in oxidation number for that species.

Oxidation: Process involving the loss of electrons. Results in an increase in oxidation number.

Oxidation state: The charge of an ion or a theoretical charge of an atom in a covalently bonded compound assuming the bond becomes ionic.

Oxidising ability: Oxidising ability is the ability of a substance to act as an oxidising agent and accept electrons.

Oxidising agent: Electron acceptors. A substance that can oxidise another species by being reduced.

p-block element: Elements in groups 3-8/0 of the periodic table. p-block non-metals generally undergo reduction reactions.

Reducing agent: Electron donors. A substance that can reduce another species by being oxidised.

Reduction: Process involving the gain of electrons. Results in a decrease in oxidation number.

Thermal stability trend: As you go down the period more heat is required for the thermal decomposition of Group 2 nitrates and carbonates because the ions increase in size and therefore have greater thermal stability.

Valence shell: The outermost shell that is occupied by electrons is the valence shell.



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.

3.4 - Chemistry of the d-block Transition Metals

Adsorption: The process by which reactants form bonds to the surface of a solid catalyst. This weakens the covalent bonds in the reactants and brings them closer together, allowing the reaction to occur more easily.

Catalyst: A substance that speeds up the rate of a reaction without being used up.

Catalytic activity of transition metals: Due to transition metals having variable oxidation states they can catalyse other reactions by acting as an oxidising/reducing agent.

Colours in transition metal complexes: Colour arises when some of the wavelengths of visible light are absorbed and the remaining wavelengths of light are transmitted or reflected. The d-orbital electrons move from the ground state to an excited state when light is absorbed.

Complex ion: An ion which has a central metal atom (typically a transition element) surrounded by ligands. The ligands are bound to the transition metal centre by dative coordinate bonds.

Coordination number: The number of dative covalent bonds formed with the central metal ion.

Dative covalent (coordinate) bond: A type of covalent bond in which one bonding atom provides both electrons in the bonding pair.

d-block: The block of elements in the middle of the periodic table. Most d-block elements are transition metals. Elements in the d block have their outer electron in the d orbital.

Heterogeneous catalysts: Catalysts that are in a different phase or state to the species in the reaction.

Homogeneous catalysts: Catalysts that are in the same phase or state as the species in the reaction.

Hydrogenation: A reaction between H_2 and another substance, this often reduces or saturates a compound. These reactions usually require a catalyst like nickel.

Ligand: An atom, ion or molecule that forms a coordinate bond with a central transition metal ion by donating a pair of electrons.



Ligand exchange: A reaction in which one ligand in a transition metal complex is replaced by another. Typically, these reactions are associated with a colour change.

Octahedral: The shape of a transition metal complex with a coordination number of 6 and a bond angle of 90° . E.g. $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$.

Oxidation state: The charge of an ion or a theoretical charge of an atom in a covalently bonded compound assuming the bond becomes ionic.

Splitting of d-orbitals: When a ligand forms a bond with a central metal ion it affects the d orbitals differently depending on how they are arranged in space. Consequently the d-orbitals split into two sets of degenerate (same energy) orbitals; in octahedral complexes there are two orbitals that are higher in energy and three that are lower, in tetrahedral complexes there are three orbitals higher and two lower.

Tetrahedral: One of the possible shapes of a transition metal complex with a coordination number of 4 and a bond angle of 109.5° . E.g. CuCl_4^{2-} .

Transition metal elements: d-block elements that can form one or more stable ions with an incomplete d-subshell. Transition elements have more than one oxidation state, form coloured ions and can often act as catalysts.

Variable oxidation states: Due to transition metals having incomplete d subshells they have the ability to have variable stable oxidation states. This is due to the 4s and 3d subshells being similar in energies.

3.5 - Chemical Kinetics

Activation energy: The minimum energy required for a reaction to occur.

Arrhenius equation: $k = Ae^{-E_a/RT}$ where k is the rate constant, A is the pre-exponential factor, E_a is the activation energy, R is the gas constant and T is the temperature.

First order reactant: A substance in the reactant mixture where the rate of reaction is directly proportional to the concentration of this substance.

Order of reaction: Tells you how the reactant's concentration will affect the rate of reaction. In the rate equation, it is the power to which the concentration of the reagent is raised.
e.g. rate = $k [\text{A}]^2[\text{B}]$.



Rate constant: Relates the rate of a chemical reaction at a given temperature to concentration of substances in the reaction mixture.

Rate determining step: The slowest step in a multi-step reaction. The overall rate is decided by this step. Species present in the rate determining step will also be in the rate equation.

Rate equation: Describes the relationship between the rate of chemical reaction and the concentrations/pressures of reagents.

Rate = $k[A]^m[B]^n$, where k is the rate constant and where m and n are the orders of reaction with respect to reactants A and B.

Rate of reaction: The measure of the amount of product formed or reactant used over time. The units of rate of reaction may be given as g/s, cm^3/s or mol/s.

Second order reactant: A substance in the reactant mixture where the rate of reaction is directly proportional to the square of the concentration of this reactant.

Total order of a reaction: Sum of all the individual orders of all the reactants in a chemical reaction.

Zero order reactant: A substance in the reactant mixture where the rate of reaction is independent of the concentration of this reacting substance.

3.6 - Enthalpy Changes for Solids and Solutions

Atomisation enthalpy ($\Delta_{\text{at}}H$): The energy required for the formation of one mole of gaseous atoms from an element in its standard state under standard conditions.

Born-Haber cycle: Calculates the lattice enthalpy by applying Hess's law and comparing the standard enthalpy change of formation of the ionic compound to values such as ionisation energy and electron affinity.

Endothermic reaction: An endothermic reaction takes in energy from the surroundings so the temperature of the surroundings decreases. In an endothermic reaction, the energy needed to break bonds in the reactants is greater than the energy released from forming new bonds in the products.

Enthalpy of hydration: The enthalpy change when one mole of gaseous ions is dissolved in water to form one mole of aqueous ions under standard conditions.





Enthalpy of solution: The enthalpy change when one mole of solute is dissolved in water to infinite dilution so that the ions no longer interact under standard conditions.

Exothermic reaction: An exothermic reaction is one that releases energy to the surroundings so the temperature of the surroundings increases. In an exothermic reaction, the energy needed to break existing bonds is less than the energy released from forming new bonds.

Ionic compound: A compound made up of anions and cations which are held together by ionic bonds, which arise due to the electrostatic attraction between oppositely charged ions. These structures are neutral overall.

Lattice dissociation enthalpy: The enthalpy change when one mole of a solid ionic compound is completely dissociated into its gaseous constituent ions under standard conditions. This is an endothermic process.

Lattice formation enthalpy: The enthalpy change when one mole of a solid ionic compound is formed from its gaseous constituent ions under standard conditions. This is an exothermic process.

Solubility: The ability of a given substance to dissolve in a solvent.

3.7 - Entropy and Feasibility of Reactions

Endothermic reaction: A reaction that takes in energy from the surroundings. The energy of the products is higher than the reactants.

Entropy: A measure of the disorder of a system. The units of entropy are $\text{JK}^{-1}\text{mol}^{-1}$. On a molecular level, gases are more disordered than liquids, which are more disordered than solids. A reaction that produces a greater number of molecules than the number of reactants molecules will have a positive entropy change, as more random arrangements of these molecules will exist, in other words, the system will become more disordered.

Entropy change: This can be calculated by finding the difference between the standard entropies of the products and the reactants:

$$\Delta S_{\text{total}} = \sum \Delta S_{\text{products}} - \sum \Delta S_{\text{reactants}}$$

If the entropy change for a reaction is positive, the products are more disordered than the reactants. If the entropy change for a reaction is negative, the products are less disordered than the reactants.



Exothermic: A reaction that releases energy to the surroundings. The energy of the reactants is higher than the products.

Feasible reaction: For a reaction to be feasible at a given temperature it must occur spontaneously. This means no extra energy is required for the reaction to occur.

Gibbs free-energy change: A measure of the feasibility of a chemical reaction. This equation relates Gibbs free energy to enthalpy change, entropy change and temperature. A negative Gibbs free energy change means the reaction is feasible and a positive Gibbs free energy change means it is not feasible.

$$\Delta G = \Delta H - T\Delta S_{\text{system}}$$

3.8 - Equilibrium Constants

Endothermic reaction: A reaction that takes in energy from the surroundings. The energy of the products is higher than the reactants.

Equilibrium: A chemical state in which the forward and reverse reactions of a process occur at the same rate. This means there is no overall change in the concentrations of the reactants and products.

Equilibrium constant (K): A value that relates the amount of products and reactants at equilibrium in a reversible reaction at a specific temperature. K is unaffected by pressure and the presence of a catalyst but is affected by temperature.

Exothermic: A reaction that releases energy to the surroundings. The energy of the reactants is higher than the products.

K_c: A value that relates the concentrations of products and reactants present at equilibrium in a reversible reaction at a specific temperature. The equilibrium constant that is equal to the concentration of products raised to their stoichiometric coefficients divided by the concentration of reactants raised to the power of their stoichiometric coefficients. Liquids and solids are not included in heterogeneous K_c expressions as their concentrations effectively remain constant.

K_p: A value that relates the amounts of gaseous products and gaseous reactants present at equilibrium in a reversible reaction at a specific temperature. For example:





$$K_p = \frac{p(C)^2}{p(A) \times p(B)^3}$$

where $p(X)$ is the partial pressure of each gas at equilibrium.

Mole fraction: A value used to calculate partial pressure.

$$\text{Mole Fraction} = \frac{\text{Number of moles of a particular gas}}{\text{Total number of moles of all gases in the mixture}}$$

Partial pressure: The pressure that would be exerted by one gas in a mixture if it occupied the container alone.

$P_A = P X_A$ where P_A is the partial pressure of A, P is the total pressure and X_A is the mole fraction of A.

Reversible reaction: A reaction in which the products from the reaction can react together to form the original reactants. The direction of reversible reactions can be changed by changing the conditions.

3.9 - Acid-Base Equilibria

Brønsted-Lowry acid: Proton donors. These species release hydrogen ions in solution.

Brønsted-Lowry base: Proton acceptors.

Buffer: A solution which is able to resist changes in pH when small volumes of acid or base are added.

End point: The point during the titration when the indicator changes colour. A suitable indicator should change colour near the equivalence point (so should have a pH range within the pH change during the equivalence point).

Equivalence point: The point when full neutralisation occurs. In acid-base titrations it is the point where all the acid/base has been neutralised and $[H^+] = [OH^-]$. This is the vertical section of an acid/base titration curve.

Hydrolysis: A reaction in which a molecule is broken down by its reaction with water.

Indicator: Indicators are chemical solutions whose colour depends on the pH of the solution they are in. Methyl orange and phenolphthalein are indicators commonly used in titrations. Methyl orange is red in acid and yellow in alkali. Phenolphthalein is colourless in acid and pink in alkali.



K_a : Acid dissociation constant, a quantitative measure of the strength of an acid in solution. The larger the K_a value the stronger the acid, since it means the acid is largely dissociated into its ions.

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

K_w : Ionic product of water. At 298K, $K_w = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$.

$$K_w = [H^+][OH^-]$$

pH: A value that represents the acidity or alkalinity of a solution. Acidic solutions have a pH of less than 7 while alkali solutions have a pH of greater than 7. Neutral solutions have a pH of 7.

$$\text{pH} = -\log[H^+]$$

$$[H^+] = 10^{-\text{pH}}$$

pH curve: A graph of pH against the volume of acid/base that is added to a solution. The curve usually has a vertical part which can be identified as the equivalence point.

pH meter: An instrument which measures the pH of a solution. A digital pH meter is preferred to a universal indicator as it can give a more precise value.

pH scale: The pH scale, from 0 to 14, is a measure of the hydrogen ion concentration and tells you about the acidity or alkalinity of a solution. It can be measured using a universal indicator or a pH probe.

Strong acid: An acid which dissociates/ionises almost completely in water. This means nearly all the H^+ ions will be released. E.g. HCl.

Strong base: A base which dissociates/ionises almost completely in water. E.g. NaOH.

Titration: An experimental technique used to determine the concentration of an unknown solution by using a second solution with a known concentration.

Weak acid: Acids which only dissociate/ionise very slightly in water so that only a small number of H^+ ions are released. E.g. Ethanoic acid.

Weak base: Bases which only slightly dissociate/ionise in water. E.g. NH_3 .

