

AQA Chemistry A-level

Physical Chemistry II

Definitions



Definitions and Concepts for AQA Chemistry A-level

Physical Chemistry II

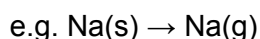
1.8: Thermodynamics

Bond enthalpy: The amount of energy required to break one mole of the stated bond in the gas phase.

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.

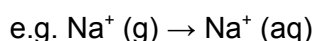
Covalent character: The partial sharing of electrons between atoms that have an ionic bond.

Enthalpy of atomisation: Enthalpy change when one mole of gaseous atoms is formed from the elements in its standard states. It is always endothermic.



Enthalpy of formation: Enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions.

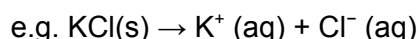
Enthalpy of hydration: Enthalpy change when one mole of a gaseous ion is completely dissolved in water under standard conditions.



Enthalpy of lattice dissociation: Enthalpy change when one mole of a solid ionic compound is converted into its gaseous ions.

Enthalpy of lattice formation: Enthalpy change when one mole of a solid ionic compound is formed from its gaseous ions.

Enthalpy of solution: Enthalpy change when one mole of ionic solid completely dissolves in water under standard conditions to form an infinitely dilute solution.

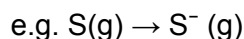


Entropy: A measure of the disorder of a system. The units of entropy are $\text{JK}^{-1}\text{mol}^{-1}$. On a molecular level, gases are more disordered than liquids, which are more disordered than solids. A reaction that produces a greater number of molecules than the number of reactants molecules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



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.

First electron affinity: Enthalpy change when one electron is added to each atom in one mole of gaseous atoms to form one mole of gaseous 1- ions.



First ionisation energy: The energy required to remove one mole of electrons from one mole of gaseous atoms of an element to form one mole of gaseous 1+ ions.

Gibbs free-energy change: A measure of the feasibility of a chemical reaction.

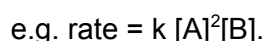
$$\Delta G = \Delta H - T\Delta S_{\text{system}}$$

Ionic character: The amount the electrons are shared between two atoms. Limited electron sharing corresponds with a higher percentage of ionic character.

1.9: Rate Equations

Concentration-time graph: A graph which can be used to deduce the rate of reaction by drawing tangents to the curve.

Order of reaction: Tells you how the reactant's concentration will affect the rate of reaction. In the rate equation, it is the appropriate power to which the concentration of the reagent is raised.



In this example, the reaction is 1st order w.r.t B, and 2nd order w.r.t A. Therefore increasing the concentration of B by a factor of 2 will increase the rate the same factor. However, increasing the concentration of A by 3 will increase the rate by a factor of $3^2 = 9$.

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

Rate-concentration data: Rate-concentration data or graphs can be used to deduce the order (0, 1 or 2) with respect to a particular reactant.

Rate constant: Relates the rate of a chemical reaction at a given temperature to the product of the concentrations of reactants.

Rate determining step: The slowest step in a multi-step reaction. The overall rate is decided by this step since species occurring in the RDS will also occur in the rate equation.



Rate equation: Describes the relationship between the rate of chemical reaction and the concentrations/pressures of reagents. $\text{Rate} = k[A]^m [B]^n$, where k is the rate constant and where m and n are the orders of reaction with respect to reactants A and B.

1.10: Equilibrium Constant K_p for Homogeneous Systems

Catalyst: Increases the rate of reaction by providing an alternative reaction pathway with a lower activation energy. A catalyst does not affect the equilibrium constant since it increases the rate of the forward and backward reaction equally.

Equilibrium: A reversible reaction is at equilibrium when the rate of the forward reaction equals the rate of the backward reaction. The concentrations of the reactants and products remains constant.

Equilibrium constant (K_p): A value that expresses the relationship between the amounts of gaseous products and gaseous reactants present at equilibrium. E.g.



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

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

Homogeneous system: A system where all the reactants and products are in the same phase.

Mole fraction of gas: 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 pressures: The pressure exerted by a particular gas in a mixture in a closed system. Related to mole fraction:

$$\text{Partial pressure} = \text{Mole fraction} \times \text{Total pressure.}$$

Reversible reaction: A reaction in which the products can react together to form the original reactants.

Total pressure: Sum of all partial pressures.

1.11: Electrode Potentials and Electrochemical Cells

Anode: Positive electrode where oxidation takes place.



Cathode: Negative electrode where reduction takes place.

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 a substance or region.

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.

EMF: The difference between the potential differences of the cathode and anode in an electrochemical cell.

Fuel cell: An electrochemical cell that converts the chemical energy of a fuel (often hydrogen) and an oxidising agent (often oxygen) into electricity through redox reactions. They do not need to be recharged.

Non-rechargeable cell: A cell that is designed to be used only once since the reactions involved are irreversible.

Rechargeable cell: cells in which reversible reactions occur, allowing them to be recharged to regain their cell potential.

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

Salt bridge: A porous substance soaked with a solution of an inert, strong electrolyte, e.g. a filter paper soaked in $\text{KNO}_3(\text{aq})$. The salt ions flow through the bridge to complete the cell and balance charges in solutions.

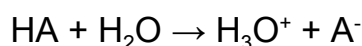
1.12: Acids and Bases

Acidic buffer: A buffer containing a weak acid and its conjugate base, e.g. a solution of acetic acid and sodium acetate.

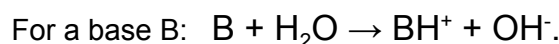
Basic buffer: A buffer containing a weak base and its conjugate acid, e.g. a solution of ammonia and ammonium chloride.



Brønsted-Lowry acid: Proton donors. They release hydrogen ions when mixed with water. The hydrogen ions mix with water molecules to form hydroxonium ions:



Brønsted-Lowry base: Proton acceptors. When in solution they bond to hydrogen ions from water molecules.



Buffer solution: A solution which resists change in pH when small amounts of strong acid/base are added.

Diprotic acid: An acid which can release two H^+ ions upon dissociation, e.g. H_2SO_4 .

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. When titrating an acid with a base, it is the point when all acid has been neutralised, and so $[\text{H}^+] = [\text{OH}^-]$.

Indicators: Indicators are chemical solutions which will change colour when there is a change of pH. Methyl orange and phenolphthalein are indicators commonly used in titrations. Methyl orange is red in an acid and changes to yellow in an alkali. Phenolphthalein is colourless in an acid and changes to pink in an 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.

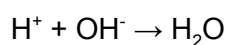
$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

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

$$K_w = [\text{H}^+][\text{OH}^-]$$

Monoprotic acid: An acid which can release only one H^+ upon dissociation, e.g. HCl.

Neutralisation: A reaction between an acid and a base which react together to form water and a salt. The ionic equation for neutralisation:



pH: A value expressing the acidity or alkalinity of a solution. A value of 0-6 implies an acidic solution, 7 implies a neutral solution and 8-14 implies an alkaline solution.

pH curve: A graph that plots pH against volume of acid/base that is added to a solution will produce a pH curve. 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 better than 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 which 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 almost completely in water. This means nearly all the H^+ ions will be released. E.g. HCl.

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

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

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

Weak base: A base which only slightly ionises in water. E.g. NH_3 .

