

AQA AS and A-level Chemistry

3.1 Physical Chemistry

3.1.1 Atomic Structure

Determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge

Explain the existence of isotopes

Interpret simple mass spectra of elements

Calculate the relative atomic mass from isotopic abundance, limited to mononuclear ions

Define first ionisation energy

Write equations for first and successive ionisation energies

Explain how first and successive ionisation energies in period 3 (Na-Ar) and in group 2 (Be-Ba) give evidence for electron configuration in sub-shells and in shells

3.1.2 Amount of substance

Define relative atomic mass (A_r) and relative molecular mass (M_r)

Carry out calculations using the Avogadro constant

Carry out calculations using mass of substance, M_r and amount in moles

Carry out calculations using concentration, volume and amount of substance in a solution

Use the ideal gas equation $pV = nRT$ in calculations (variables in SI units)

Calculate empirical formula from data giving composition by mass or percentage mass

Calculate molecular formula from the empirical formula and relative molecular mass

Write balanced equations (full and ionic) for reactions studied

Balance equations for unfamiliar reactions when reactants and products are specified

State the economic, ethical and environmental advantages for society and for industry for developing chemical reactions with high atom economy

Use balanced equations to calculate: Masses, volumes of gases, percentage yield, atom economies; and concentrations & volumes for liquids

REQUIRED PRACTICAL 1: make up a volumetric solution and carry out simple acid-base titration

3.1.3 Bonding

Ionic bonding involves electrostatic attraction between oppositely charged ions in a lattice

Predict the charge on a simple ion using the position of the element in the periodic table

Construct formulas for ionic compounds (e.g. sulfate, hydroxide, nitrate, carbonate and ammonium)

Single covalent bond contains a shared pair of electrons; multiple bonds contain multiple pairs of electrons; co-ordinate (dative covalent) bond shares a pair of electrons, both supplied by 1 atom

Represent a covalent bond using a line; co-ordinate bond using an arrow

Metallic bonding involves attraction between delocalised electrons and positive ions arranged in a lattice

The structures of: diamond, graphite, ice, iodine, magnesium and sodium chloride as examples of one of these 4 crystal structures: ionic, metallic, macromolecular, molecular

Relate the melting point and conductivity of materials to the type of structure and bonding present

Explain the energy changes associated with changes of state

Draw diagrams to represent these structures involving specified numbers of particles

Explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs) surrounding the central atom

Pairs of electrons as clouds that reel each other, arranging themselves as far apart as possible; with lone pair lone pair repulsion being greater than pair bond, pair bond repulsion
Define electronegativity
Use partial charges to show that a bond is polar
Explain why some molecules with polar bonds do not have a permanent dipole
Explain the existence of: permanent dipole-dipole forces; induced dipole-dipole (van der Waals, dispersion, London) forces; hydrogen bonding; between familiar and unfamiliar molecules
Explain how melting and boiling points are influenced by these intermolecular forces
3.1.4 Energetics
Understand reactions can be exothermic or endothermic and that enthalpy change (ΔH) is the heat energy change measured under conditions of constant pressure
Understand the term standard conditions
Define standard enthalpy change of combustion ($\Delta_c H^\theta$) and standard enthalpy change of formation ($\Delta_f H^\theta$)
Use the equation $q=mc\Delta T$ to calculate the molar enthalpy change for a reaction and in related calculations
REQUIRED PRACTICAL 2: Measurement of an enthalpy change
Use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation
Define the term mean bond enthalpy
Use mean bond enthalpies to calculate an approximate value of ΔH for reactions in the gaseous state
Explain why values from mean bond enthalpy calculations differ from those determined using Hess's law
3.1.5 Kinetics
Define the term activation energy
Explain why most collisions do not lead to a reaction
Draw and explain Maxwell-Boltzmann distribution curves for different temperatures
Define the term: rate of reaction
Use the Maxwell-Boltzmann distribution curve to explain why a small increase in temperature can lead to a large increase in rate
REQUIRED PRACTICAL 3: Investigation of how the rate of a reaction changes with temperature
Explain how a change in concentration or a change in pressure influences the rate of a reaction (collision frequency)
Define the term catalyst and explain how they work (activation energy; alternative pathway)
Use a Maxwell-Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas
3.1.6 Chemical equilibria, Le Chatelier's principle and K_c
Explain what is happening in a reversible reaction at equilibrium
Use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium (Catalysts do not affect it)
Explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used
Construct an expression for K_c for a homogeneous system in equilibrium (using $[X]$ for a species X of mol dm^{-3} concentration)
Calculate a value for K_c from the equilibrium concentrations for a homogeneous system at constant temperature
Perform calculations involving K_c
Predict the qualitative effects of changes of temperature on the value of K_c

3.1.7 Oxidation, reduction and redox equations

Understand the terms oxidation and reduction in terms of electrons

Work out the oxidation state of an element in a compound or ion from the formula

Write half equations identifying the oxidation and reduction processes in redox reactions

Combine half equations to give an overall redox equation