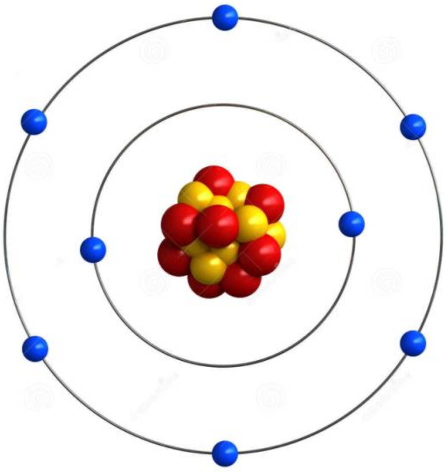
A-level Chemistry

3.1 Physical Chemistry

(Y12)



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Lesson Date(s)	Topic	Learning Outcomes From Specification	Key Words & Equations	Notes (Revision tips, exam technique tips, priorities for revision etc.)
	3.1.1.1 Fundamental Particles	Appreciate that knowledge and understanding of atomic structure has evolved over time. Protons, neutrons and electrons: relative charge and relative mass. An atom consists of a nucleus containing protons and neutrons surrounded by electrons.		

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	3.1.1.2 Mass Number & Isotopes	Mass number (A) and atomic (proton) number (Z). Students should be able to: determine the number of fundamental particles in atoms and ions using mass number, atomic number and charge explain the existence of isotopes. The principles of a simple time of flight (TOF) mass spectrometer, limited to ionisation, acceleration to give all ions constant kinetic energy, ion drift, ion detection, data analysis. The mass spectrometer gives accurate information about relative isotopic mass and also about the relative abundance of isotopes. Mass spectrometry can be used to identify elements. Mass spectrometry can be used to determine relative molecular mass. Students should be able to: interpret simple mass spectra of elements calculate relative atomic mass from isotopic abundance, limited to mononuclear ions.		

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	3.1.1.3 Electron Configuration	Electron configurations of atoms and ions up to Z = 36 in terms of shells and sub-shells (orbitals) s, p and d. Ionisation energies. Students should be able to: 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.		
	3.1.2.1 Relative Atomic Mass & Relative Molecular Mass	Relative atomic mass and relative molecular mass in termsof ^{12}C. The term relative formula mass will be used for ionic compounds. Students should be able to: define relative atomic mass (A_r) define relative molecular mass (M_r).		

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	3.1.2.2 The Mole & Avagadro Constant	The Avogadro constant as the number of particles in a mole. The mole as applied to electrons, atoms, molecules, ions, formulas and equations. The concentration of a substance in solution, measured in mol dm^{-3}. Students should be able to carry out calculations: using the Avogadro constant using mass of substance, Mr, and amount in moles using concentration, volume and amount of substance in a solution. Students will not be expected to recall the value of the Avogadro constant.		
	3.1.2.3 The Ideal Gas Equation	The ideal gas equation $pV = nRT$ with the variables in SI units. Students should be able to use the equation in calculations. Students will not be expected to recall the value of the gas constant, R.		

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	3.1.2.4 Empirical & Molecular Formula	Empirical formula is the simplest whole number ratio of atoms of each element in a compound. Molecular formula is the actual number of atoms of each element in a compound. The relationship between empirical formula and molecular formula. Students should be able to: calculate empirical formula from data giving composition by mass or percentage by mass calculate molecular formula from the empirical formula and relative molecular mass.		

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	3.1.2.5 Balanced Equations & Associated Calculations	Equations (full and ionic). Percentage atom economy is: molecular mass of desired product x100 / sum of molecular masses of all reactants Economic, ethical and environmental advantages for society and for industry of developing chemical processes with a high atom economy. Students should be able to: write balanced equations for reactions studied balance equations for unfamiliar reactions when reactants and products are specified. Students should be able to use balanced equations to calculate: masses volumes of gases percentage yields percentage atom economies concentrations and volumes for reactions in solutions. **Required practical 1:** **Make up a volumetric solution and carry out a simple**		

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	3.1.3.1 Ionic Bonding	Ionic bonding involves electrostatic attraction between oppositely charged ions in a lattice. The formulas of compound ions eg sulfate, hydroxide, nitrate, carbonate and ammonium. Students should be able to: predict the charge on a simple ion using the position of the element in the Periodic Table construct formulas for ionic compounds.		
	3.1.3.2 Nature of Covalent & Dative Covalent Bonds	A single covalent bond contains a shared pair of electrons. Multiple bonds contain multiple pairs of electrons. A co-ordinate (dative covalent) bond contains a shared pair of electrons with both electrons supplied by one atom. Students should be able to represent: covalent bond using a line a co-ordinate bond using an arrow.		
	3.1.3.3 Metallic Bonding	Metallic bonding involves attraction between delocalised electrons and positive ions arranged in a lattice.		

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	3.1.3.4 Bonding & Physical Properties	The four types of crystal structure: ionic metallic macromolecular (giant covalent) molecular. The structures of the following crystals as examples of these four types of crystal structure: diamond graphite ice iodine magnesium sodium chloride. Students should be able to: relate the melting point and conductivity of materials to the type of structure and the bonding present explain the energy changes associated with changes of state draw diagrams to represent these structures involving specified numbers of particles.		

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	3.1.3.5 Shapes of Simple Molecules & Ions	Bonding pairs and lone (non-bonding) pairs of electrons as charge clouds that repel each other. Pairs of electrons in the outer shell of atoms arrange themselves as far apart as possible to minimise repulsion. Lone pair–lone pair repulsion is greater than lone pair–bond pair repulsion, which is greater than bond pair–bond pair repulsion. The effect of electron pair repulsion on bond angles. Students should be able to explain the shapes of, and bond angles in, simple molecules and ions with up to six electron pairs (including lone pairs of electrons) surrounding the central atom.		
	3.1.3.6 Bond Polarity	Electronegativity as the power of an atom to attract the pair of electrons in a covalent bond. The electron distribution in a covalent bond between elements with different electronegativities will be unsymmetrical. This produces a polar covalent bond, and may cause a molecule to have a permanent dipole. Students should be able to: use partial charges to show that a bond is polar explain why some molecules with polar bonds do not have a permanent dipole.		

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	3.1.3.7 Forces Between Molecules	Forces between molecules: permanent dipole–dipole forces induced dipole–dipole (van der Waals, dispersion, London) forces hydrogen bonding. The melting and boiling points of molecular substances are influenced by the strength of these intermolecular forces. The importance of hydrogen bonding in the low density of ice and the anomalous boiling points of compounds. Students should be able to: explain the existence of these forces between familiar and unfamiliar molecules explain how melting and boiling points are influenced by these intermolecular forces.		

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	3.1.4.1 Enthalpy Change	Reactions can be endothermic or exothermic. Enthalpy change (ΔH) is the heat energy change measured under conditions of constant pressure. Standard enthalpy changes refer to standard conditions ie 100 kPa and a stated temperature (eg $\Delta H^{298\,\ominus}$). Students should be able to: define standard enthalpy of combustion ($\Delta cH^{\ominus}$) define standard enthalpy of formation ($\Delta fH^{\ominus}$).		
	3.1.4.2 Calorimetry	The heat change, q, in a reaction is given by the equation $q = mc\Delta T$ where m is the mass of the substance that has a temperature change ΔT and a specific heat capacity c. Students should be able to: use this equation to calculate the molar enthalpy change for a reaction use this equation in related calculations. Students will not be expected to recall the value of the specific heat capacity, c, of a substance. **Required practical 2 Measurement of an enthalpy change.**		

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	3.1.4.3 Applications of Hess's Law	Hess's law. Students should be able to: use Hess's law to perform calculations, including calculation of enthalpy changes for reactions from enthalpies of combustion or from enthalpies of formation.		
	3.1.4.4 Bond Enthalpies	Mean bond enthalpy. Students should be able to: define the term mean bond enthalpy use mean bond enthalpies to calculate an approximate value of ΔH for reactions in the gaseous phase explain why values from mean bond enthalpy calculations differ from those determined using Hess's law.		

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	3.1.5.1 Collision Theory	Reactions can only occur when collisions take place between particles having sufficient energy. This energy is called the activation energy. Students should be able to: define the term activation energy explain why most collisions do not lead to a reaction		
	3.1.5.2 Maxwell – Boltzmann Distribution	Maxwell–Boltzmann distribution of molecular energies in gases. Students should be able to: draw and interpret distribution curves for different temperatures.		

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	3.1.5.3 Effect of temperature on reaction rate	Meaning of the term rate of reaction. The qualitative effect of temperature changes on the rate of reaction. Students should be able to: use the Maxwell–Boltzmann distribution to explain why a small temperature increase can lead to a large increase in rate. **Required practical 3** **Investigation of how the rate of a reaction changes with temperature.**		
	3.1.5.4 Effect of Concentration & Pressure	The qualitative effect of changes in concentration on collision frequency. The qualitative effect of a change in the pressure of a gas on collision frequency. Students should be able to: explain how a change in concentration or a change in pressure influences the rate of a reaction.		

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	3.1.5.5 Catalysts	A catalyst is a substance that increases the rate of a chemical reaction without being changed in chemical composition or amount. Catalysts work by providing an alternative reaction route of lower activation energy. Students should be able to: use a Maxwell–Boltzmann distribution to help explain how a catalyst increases the rate of a reaction involving a gas.		

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	3.1.6.1 Chemical equilibria and Le Chatelier's principle	Many chemical reactions are reversible. In a reversible reaction at equilibrium: forward and reverse reactions proceed at equal rates the concentrations of reactants and products remain constant. Le Chatelier's principle. Le Chatelier's principle can be used to predict the effects of changes in temperature, pressure and concentration on the position of equilibrium in homogeneous reactions. A catalyst does not affect the position of equilibrium. Students should be able to: use Le Chatelier's principle to predict qualitatively the effect of changes in temperature, pressure and concentration on the position of equilibrium explain why, for a reversible reaction used in an industrial process, a compromise temperature and pressure may be used.		

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	3.1.6.2 Equilibrium Constant Kc for Homogenous Systems	The equilibrium constant K_c is deduced from the equation for a reversible reaction. The concentration, in mol dm^{-3}, of a species X involved in the expression for K_c is represented by [X] The value of the equilibrium constant is not affected either by changes in concentration or addition of a catalyst. Students should be able to: construct an expression for K_c for a homogeneous system in equilibrium 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		

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	3.1.7 Oxidation, Reduction & Redox Equations	Oxidation is the process of electron loss and oxidising agents are electron acceptors. Reduction is the process of electron gain and reducing agents are electron donors. The rules for assigning oxidation states. Students should be able to: 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.		