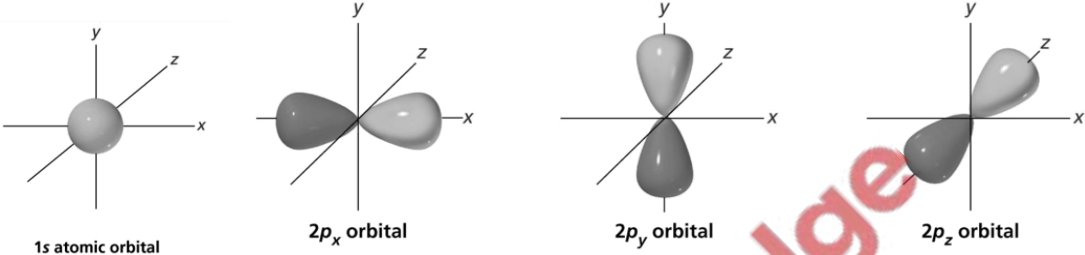


**LIST OF DEFINITIONS FOR  
A'LEVEL CHEMISTRY 9701**

List of Definitions for 9701 A'Level Chemistry

(Black is for AS, Red is for A2)

Atoms, Molecules and Stoichiometry	1.	The <b>relative atomic mass</b> ( $A_r$ ) of an element is defined as the average mass of <b>one atom</b> compared to 1/12 the mass of a $^{12}\text{C}$ atom. OR The <b>relative atomic mass</b> ( $A_r$ ) of an element is defined as the mass of <b>one mole of atoms</b> compared to 1/12 the mass of one mole of $^{12}\text{C}$ atoms.
	2.	The <b>relative isotopic mass</b> of an isotope (of a particular element) is defined as the mass of <b>one isotope</b> compared to 1/12 the mass of a $^{12}\text{C}$ atom.
	3.	The <b>relative molecular mass</b> of a molecule is defined as the average mass of <b>one molecule</b> compared to 1/12 the mass of a $^{12}\text{C}$ atom.
	4.	The <b>relative formula mass</b> of <u>an ionic compound</u> is defined as the average mass of <b>one formula unit</b> compared to 1/12 the mass of a $^{12}\text{C}$ atom. OR The <b>relative formula mass</b> of <u>an ionic compound</u> is defined as the mass of <b>one mole of formula units</b> compared to 1/12 the mass of one mole of $^{12}\text{C}$ atoms.
	5.	A <b>mole</b> of substance is defined as the amount of substance that contains as many entities (atoms, molecules, ions, electrons or any other particles) as the number of atoms in 12g of the carbon-12.  It is equal to $6.022 \times 10^{23}$ , which is called the <b>Avogadro constant or Avogadro number</b> .  $n = \frac{\text{mass}}{M_r}$ ; $n = \text{concentration} \times \text{volume}$ ; $n = \frac{V \text{ (in dm}^3\text{)}}{24}$ (for gases only)
Redox Reactionbs	6.	<b>Oxidation</b> is a process where a chemical species <b>loses</b> electrons; ( <b>Oxidation Is Loss</b> ) <b>Reduction</b> is a process where a chemical species <b>gains</b> electrons. ( <b>Reduction Is Gain</b> )  A <b>redox</b> reaction refers to a reaction where <u>oxidation and reduction occurs simultaneously</u> .
	7.	An <b>oxidising agent</b> is a species that accepts / gains electrons (is reduced) n a reaction. A <b>reducing agent</b> is a species that donates / loses electrons (is oxidised) in a reaction.
	8.	A <b>disproportionation</b> reaction is a redox reaction in which <u>one species is simultaneously oxidised and reduced</u> .  e.g. $2\text{Cu}^+ \rightarrow \text{Cu} + \text{Cu}^{2+}$

Atomic Structure	9.	<b>Atomic number</b> of an element refers to the number of protons it contains. <b>Mass number</b> (nucleon number) refers to the sum of the protons and neutrons it contains.
	10.	<b>Isotopes</b> refer to atoms of the same element with the same number of protons but different number of neutrons.
	11.	<p>An <b>atomic orbital</b> is defined as a region of three-dimensional space around the nucleus, whereby there is a 95% chance of locating a particular electron. Each orbital has a characteristic energy level and shape. For 'A' level syllabus, you need to know the shapes of <b>s</b> and <b>p</b> orbitals.</p>  <p style="text-align: center;"> <span style="margin-right: 100px;">1s atomic orbital</span> <span style="margin-right: 100px;">2p<sub>x</sub> orbital</span> <span style="margin-right: 100px;">2p<sub>y</sub> orbital</span> <span>2p<sub>z</sub> orbital</span> </p>
Chemical Bonding	12.	<p><b>Valence-shell electron pair repulsion (VSEPR) theory</b> is a model used to predict the shape of individual molecules based upon the extent of electron-pair electrostatic repulsion.</p> <p>Lone pair – lone pair repulsion &gt; Lone pair – bond pair repulsion &gt; Bond pair – bond pair repulsion</p>
	13.	<b>Metallic bond</b> is the electrostatic attraction between positively charged cations and the 'sea' of delocalised electrons.
	14.	<p>Electrovalent (<b>ionic</b>) <b>bond</b> is the electrostatic attraction between oppositely charged ions which have been formed by the transfer of one or more electrons to achieve the stable electronic configuration of a noble gas.</p> <p><b>Coordination number of an ion in an ionic compound</b> refers to the number of neighboring <b>oppositely</b> charged ions.</p>
	15.	<b>Covalent bond</b> is the electrostatic force of attraction of the nuclei of the 2 atoms for the shared pair(s) of electrons between them.
	16.	<b>Dative / Co-ordinate Covalent bond</b> is a covalent bond in which a pair of electrons is shared between 2 atoms but <b>ONLY ONE</b> of them provides <b>both</b> electrons that make up the bond.
	17.	<b>Electronegativity</b> refers to the ability/tendency of an atom to attract electrons in a bond towards itself. Electronegativity increases across the period and decreases down a group.
	18.	<b>Permanent dipole-permanent dipole interactions</b> are a type of intermolecular forces between polar molecules (molecules with a net dipole moment) which have a simple covalent structure.
19.	<b>Temporary dipole-induced dipole interactions</b> are a type of intermolecular forces between non-polar molecules (molecules with NO NET dipole moment) which have a simple covalent structure.	

	20.	<b>Hydrogen bonds</b> are a special case of permanent dipole-permanent dipole interactions, whereby there is an attractive interaction of a hydrogen atom with an electronegative atom, such as nitrogen, oxygen or fluorine (typically from another molecule). Do not confuse this with a covalent bond between H and N, O or F.																																								
The Gaseous State	21.	<p><b>Basic Assumptions of kinetic theory of gases</b></p> <ul style="list-style-type: none"> <li>• Gases consist of small particles of <b>negligible size/volume</b> as compared to the size of the container.</li> <li>• Gas particles have <b>negligible intermolecular forces of attraction</b> between each other.</li> <li>• Collisions between gas particles are <b>perfectly elastic</b>. I.e. there is no loss of kinetic energy upon collision.</li> </ul>																																								
	22.	<p><b>Ideal Gas equation <math>PV = nRT</math></b> Use <b>Pa</b> for Pressure, <b>m<sup>3</sup></b> for volume and <b>K</b> for temperature</p> <table border="1"> <tr> <td rowspan="3"><b>Pressure</b></td> <td>1 atm</td> <td>=</td> <td><math>1.01 \times 10^5 \text{ Pa}</math></td> </tr> <tr> <td>1 bar</td> <td>=</td> <td><math>1 \times 10^5 \text{ Pa}</math></td> </tr> <tr> <td>760 mmHg</td> <td>=</td> <td><math>1.01 \times 10^5 \text{ Pa}</math></td> </tr> <tr> <td rowspan="2"><b>Volume</b></td> <td>1 dm<sup>3</sup></td> <td>=</td> <td><math>10^{-3} \text{ m}^3</math></td> </tr> <tr> <td>1 cm<sup>3</sup></td> <td>=</td> <td><math>10^{-6} \text{ m}^3</math></td> </tr> <tr> <td><b>Temperature</b></td> <td colspan="3"><math>T(\text{K}) = T(^{\circ}\text{C}) + 273</math></td> </tr> <tr> <td rowspan="3"><b>At s.t.p</b></td> <td>Temperature</td> <td>=</td> <td>273 K (0°C)</td> </tr> <tr> <td>Pressure</td> <td>=</td> <td><math>1.01 \times 10^5 \text{ Pa}</math> (1 atm)</td> </tr> <tr> <td>Molar volume</td> <td>=</td> <td><math>0.0224 \text{ m}^3</math></td> </tr> <tr> <td rowspan="3"><b>At r.t.p (standard conditions)</b></td> <td>Temperature</td> <td>=</td> <td>298 K (25°C)</td> </tr> <tr> <td>Pressure</td> <td>=</td> <td><math>1.01 \times 10^5 \text{ Pa}</math> (1 atm)</td> </tr> <tr> <td>Molar volume</td> <td>=</td> <td><math>0.024 \text{ m}^3</math></td> </tr> </table>	<b>Pressure</b>	1 atm	=	$1.01 \times 10^5 \text{ Pa}$	1 bar	=	$1 \times 10^5 \text{ Pa}$	760 mmHg	=	$1.01 \times 10^5 \text{ Pa}$	<b>Volume</b>	1 dm <sup>3</sup>	=	$10^{-3} \text{ m}^3$	1 cm <sup>3</sup>	=	$10^{-6} \text{ m}^3$	<b>Temperature</b>	$T(\text{K}) = T(^{\circ}\text{C}) + 273$			<b>At s.t.p</b>	Temperature	=	273 K (0°C)	Pressure	=	$1.01 \times 10^5 \text{ Pa}$ (1 atm)	Molar volume	=	$0.0224 \text{ m}^3$	<b>At r.t.p (standard conditions)</b>	Temperature	=	298 K (25°C)	Pressure	=	$1.01 \times 10^5 \text{ Pa}$ (1 atm)	Molar volume	=
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Chemical Energetics	23.	<b>Hess' law</b> states that the change in enthalpy accompanying a reaction is independent of the path taken between the initial and final states.	
	24.	The standard enthalpy change of <b>reaction</b> ( $\Delta H_{\text{rxn}}^{\circ}$ ), is the enthalpy change when molar quantities of reactants (as specified by the chemical equation) react to form products under standard conditions 25°C and 1 atm.	
	25.	The standard enthalpy change of <b>formation</b> of a compound ( $\Delta H_{\text{f}}^{\circ}$ ), is the enthalpy change when <u>1 mole</u> of a pure compound in a specified state is formed from its <u>constituent elements</u> in their standard states, under standard conditions 25°C and 1 atm.	
	26.	The standard enthalpy change of <b>combustion</b> of a compound ( $\Delta H_{\text{c}}^{\circ}$ ), is the enthalpy change when <u>1 mole</u> of that compound is <u>completely burnt</u> in oxygen under standard conditions 25°C and 1 atm.	
	27.	The standard enthalpy change of <b>neutralisation</b> ( $\Delta H_{\text{neu}}^{\circ}$ ), is the enthalpy change when an acid and a base react to form <u>1 mole of water</u> under standard conditions 25°C, and 1 atm.	
	28.	The standard enthalpy change of <b>atomisation</b> of an element ( $\Delta H_{\text{atom}}^{\circ}$ ), is the enthalpy change when <u>1 mole of atoms</u> in the gaseous state is formed from the element in its normal physical state under standard conditions 25°C and 1 atm.	
	29.	The <b>bond dissociation energy</b> of a bond is the energy required to break one mole of chemical bonds between two atoms in a molecule in the gaseous phase.	
	30.	The <b>first ionisation energy</b> of an element ( $\Delta H_{\text{1st I.E.}}^{\circ}$ ), is the energy required to remove 1 mole of electrons from 1 mole of <u>gaseous atoms</u> , to form 1 mole of <u>gaseous singly charged cations</u> . $M_{(\text{g})} \rightarrow M_{(\text{g})}^{+} + e^{-}$	
	31.	The <b>second ionisation energy</b> of an element ( $\Delta H_{\text{2nd I.E.}}^{\circ}$ ), is the energy required to remove 1 mole of electrons from 1 mole of <u>gaseous singly charged cations</u> , to form 1 mol of <u>gaseous doubly charged cations</u> . $M_{(\text{g})}^{+} \rightarrow M_{(\text{g})}^{2+} + e^{-}$	
	32.	<b>The first electron affinity</b> of an element ( $\Delta H_{\text{1st E.A.}}^{\circ}$ ), is the enthalpy change when 1 mol of electrons are added to 1 mol of <u>gaseous atoms</u> , to form 1 mol of <u>gaseous singly charged anions</u> .	
	33.	<b>Lattice energy</b> is the energy evolved when 1 mole of an ionic solid is formed from its constituent gaseous ions under standard conditions 25°C and 1 atm.	
	34.	The standard enthalpy change of <b>hydration</b> of a gaseous ion ( $\Delta H_{\text{hyd}}^{\circ}$ ), is the enthalpy change when 1 mole of hydrated aqueous ions is formed from the gaseous ions under standard conditions 25°C and 1 atm. .	
	35.	The standard enthalpy change of <b>solution</b> of an ionic compound ( $\Delta H_{\text{soln}}^{\circ}$ ), is the enthalpy change when 1 mole of an ionic compound is dissolved in a large excess of water under standard conditions 25°C and 1 atm. $\Delta H_{\text{solution}} = \sum \Delta H_{\text{hydration}} - \text{Lattice energy}$	
	Chemical Energetics	36.	<b>Entropy (S)</b> measures the degree of <u>disorder</u> in a system. The entropy of a system increases when the matter or energy in the system becomes more random in its arrangement. A system that has a high degree of disorder/randomness is said to have a large entropy. Gases have the highest entropy followed by liquids and solids.
		37.	<b>The Gibbs Free Energy change, <math>\Delta G</math></b> , is the limiting maximum useful work that can be

	<p>obtained from a reaction, at constant pressure. When <math>\Delta G &lt; 0</math>, the reaction is spontaneous.</p> <p>At standard state conditions,</p> $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$
	<p>38. The <b>standard electrode potential</b> of an element is the potential difference between the element and its aqueous ion of <math>1.00 \text{ mol dm}^{-3}</math> relative to that of the standard hydrogen electrode at 1 atm and 298 K.</p>
	<p>39. The <b>standard cell potential</b>, <math>E^\circ_{\text{cell}}</math>, is the potential difference between two standard half cells measured under standard conditions. When <math>E^\circ_{\text{cell}} &gt; 0</math>, the reaction is spontaneous.</p> $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
Chemical Equilibria	<p>40. <b>Dynamic Equilibrium</b> refers to a reversible reaction in which the <i>forward and the backward reactions</i> are taking place at the <u>same rate</u> and concentrations of reactants and product are <u>constant</u>.</p>
	<p>41. <b>Le Chatelier's Principle</b> states that if a system in <u>equilibrium</u> is subjected to a change which <u>disturbs</u> the equilibrium, the system will <u>respond</u> in such a manner as to <u>reduce or counteract</u> the effect of the change.</p> <p>Industrial application: <b>Haber Process: 450 °C – 500 °C, 200 atm – 300 atm</b></p>
Ionic Equilibria	<p>42. Monoprotic or monobasic acids can donate only one proton. E.g. HCl, HNO<sub>3</sub> and CH<sub>3</sub>COOH Diprotic or dibasic acids can donate two protons. E.g. H<sub>2</sub>SO<sub>4</sub>, H<sub>2</sub>S and H<sub>2</sub>CO<sub>3</sub></p> <p>A Bronsted <b>acid</b> is a proton donor.</p> <p>A strong acid is one that <b>dissociates completely</b> in aqueous solution to give H<sub>3</sub>O<sup>+</sup> ions.  <math>\text{HA}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{A}^-_{(\text{aq})}</math></p> <p>Weak acids only <b>dissociate partially</b> in aqueous solution forming ionic equilibrium systems  <math>\text{HA}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}_3\text{O}^+_{(\text{aq})} + \text{A}^-_{(\text{aq})}</math></p> <p><math>K_a</math> provides an accurate measure of the extent to which a <u>weak acid</u> is dissociated.</p> $K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{H}_2\text{O}][\text{HA}]}$ <p><b><math>K_a</math> is only affected by changes in temperature</b></p>

Ionic Equilibria	43.	<p>A Bronsted <b>base</b> is a proton acceptor.</p> <p>A strong base is one that dissociates completely in aqueous solution to give OH<sup>-</sup> ions.  <math>B_{(aq)} + H_2O_{(l)} \rightarrow BH^+_{(aq)} + OH^-_{(aq)}</math></p> <p>Weak bases only dissociate partially in aqueous solution forming ionic equilibrium systems.  <math>B_{(aq)} + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}</math>  <math display="block">K_b = \frac{[BH^+][OH^-]}{[B]}</math></p> <p><b>K<sub>b</sub> is only affected by changes in temperature</b></p>
	44.	<p><b>pH</b> is defined as the negative logarithm to base 10 of [H<sub>3</sub>O<sup>+</sup>] ⇒ pH = -log<sub>10</sub>[H<sub>3</sub>O<sup>+</sup>]</p> <p>pOH is thus the negative logarithm to base 10 of [OH<sup>-</sup>] ⇒ pOH = -log<sub>10</sub>[OH<sup>-</sup>]</p> <p>pH + pOH = 14</p>
	45.	<p>Water ionizes itself to a very small extent to give H<sub>3</sub>O<sup>+</sup> and OH<sup>-</sup> ions.</p> $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)} \quad \Delta H = +ve$ <p>The equilibrium constant for the above system is given the symbol K<sub>w</sub> and is known as the ionic product of water.</p> $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ mol}^2\text{dm}^{-6} \text{ (at } 25^\circ\text{C)}$ <p>As the <b>auto-ionisation of water</b> is an endothermic process, when temperature is increased, equilibrium shifts to the right to absorb the heat. [H<sub>3</sub>O<sup>+</sup>] and [OH<sup>-</sup>] increase by the same amount, K<sub>w</sub> increases.</p> <p><b>K<sub>w</sub> is only affected by change in temperature</b></p>
	46.	<p>An <b>acidic/alkaline buffer</b> solution is an aqueous solution consisting of a mixture of a weak acid and its conjugate base or a mixture of a weak base and its conjugate acid. It has the property that it resists changes in pH when a small amount of acid or base is added to it.</p>

Reaction Kinetics	47.	<p><b>Reaction rate</b> is defined as the increase in concentration of one of the products per unit time or the decrease in concentration of one of the reactants per unit time.</p> <p>e.g. <math>A + B \rightarrow 2C</math></p> $\text{Rate} = -\frac{d[A]}{dt} = -\frac{d[B]}{dt} = \frac{1}{2} \times \frac{d[C]}{dt}$
	48.	<p>Types of rate:</p> <p><b>Initial rate</b> is change in concentration of reactants or product at time <math>t = 0</math>.</p> <p><b>Instantaneous rate</b> is rate of reaction at any given time/instant.</p> <p><b>Average rate is total concentration of reactant used or total concentration of product formed over total time.</b></p>
	49.	The minimum energy which colliding molecules must possess for successful collision/reaction is called the <b>activation energy, <math>E_a</math></b> .
	50.	<p>Rate law or rate equation is the mathematical relationship between the rate of a reaction and the concentration of the reactants in a reaction. E.g. <math>A + B \rightarrow \text{Products}</math></p> <p>The rate law is</p> $\text{Rate} = k[A]^m[B]^n$ <p>where <math>k</math> is the rate constant</p> <p><math>m</math> is the order of reaction with respect to reactant A</p> <p><math>n</math> is the order of reaction with respect to reactant B</p> <p><b><math>(m + n)</math> is the overall order of the reaction</b></p>
	51.	<b>The order of reaction</b> with respect to a particular reactant is the power to which the concentration of that reactant is raised in an experimentally determined rate equation / rate law.
	52.	The <b>rate constant, <math>k</math></b> , is the proportionality constant in the experimentally determined rate law.
	53.	The <b>half-life (<math>t_{1/2}</math>)</b> of a reaction is the time taken for the concentration of a reactant to fall to half its initial value. It is constant only for a first order reaction as it is independent of reactant concentrations.
	54.	A <b>catalyst</b> is a substance that increases the rate of a reaction by providing an alternative reaction pathway that has lower activation energy.
Transition	55.	<b>Biological catalysts</b> such as enzymes are very selective in the reactions that they catalyze, and some are absolutely specific, operating for only one substance in only one reaction. For reactions that normally produce a pair of optical isomers (racemic mixture) when carried out in the lab, enzymes are able to selectively produce one optical isomer in the body.
	56.	A <b>transition element</b> is a d-block element which is able to form one or more stable ions with a partially filled d-subshell.



	57.	A <b>ligand</b> is an anion or neutral molecule with at least one lone pair of electrons available for dative bonding to the central metal atom or ion.
	58.	A <b>complex ion</b> contains a central metal atom or ion bonded to ligands by coordinate (or dative) bonds.
	59.	<b>Co-ordination number</b> indicates the number of co-ordinate (or dative) bonds around the central metal ion.
Organic Chemistry	60.	<b>Empirical formula</b> is the simplest formula that shows the <u>ratio</u> of each kind of atom in a molecule. e.g. C <sub>2</sub> H <sub>5</sub> is the empirical formula for C <sub>4</sub> H <sub>10</sub>
	61.	<b>Molecular formula</b> shows the <u>actual number</u> of each kind of atoms in a molecule. e.g. C <sub>4</sub> H <sub>10</sub>
	62.	<b>Structural formula</b> shows how the atoms are connected to each other in a molecule. e.g. CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>3</sub>
	63.	<b>Displayed/full formula</b> shows <u>all the bonds</u> and relative placing of all the atoms in a molecule. e.g. $  \begin{array}{cccc}  \text{H} & \text{H} & \text{H} & \text{H} \\    &   &   &   \\  \text{H}-\text{C} & -\text{C} & -\text{C} & -\text{C}-\text{H} \\    &   &   &   \\  \text{H} & \text{H} & \text{H} & \text{H}  \end{array}  $
	64.	<b>Homologous series</b> are compounds have the same general formula and functional group and each homologue differs from its neighbor by a fixed group of atoms (e.g. -CH <sub>2</sub> ). As we go down a homologous series, the chemical properties remain unchanged but there is a gradual change in physical properties. Examples of homologous series are alkanes, alkenes, alcohols.....
	65.	<b>Structural isomerism</b> refers to compounds with the same molecular formula but different structural formula. E.g. CH <sub>3</sub> COOCH <sub>3</sub> and C <sub>2</sub> H <sub>5</sub> COOH
	66.	<p><b>Stereoisomerism</b> refers to compounds that have the same molecular formula but with different spatial arrangements.</p> <ul style="list-style-type: none"> <li>• <b>Geometric isomers</b> have same carbon skeleton with double bonds restricting free rotation. For geometric isomerism to exist, there must be two different groups of atoms bonded to each side of the C=C bond.</li> <li>• <b>Optical isomers</b> are non-superimposable mirror images of each other (enantiomers). Isomers have at least one chiral C atom, i.e. there are four different groups attached and have no plane of symmetry. An equal proportion of enantiomers forms a <b>racemic mixture</b> which is optically inactive.</li> </ul>

67.	The <b>primary structure</b> of a protein shows the exact order (or unique sequence) of the $\alpha$ -amino acids held by peptide/amide linkages along the polypeptide chain. The primary structure determines what the protein is, how it folds and its function.
68.	The <b>secondary structure</b> refers to the detailed configurations of the polypeptide chain. In a protein molecule, the long chain of amino acid units may be coiled into an $\alpha$ -helix or folded into a $\beta$ -pleated sheet. Both structures are stabilized by <u>hydrogen bonds</u> between the N-H group of one amino acid residue and the C=O group of another along the main chain.
69.	<p>The <b>tertiary structure</b> of the protein refers to the overall 3-dimensional shape of the <b>entire</b> protein involving folding or coiling of the chains. It shows how protein molecules are arranged in relation to each other.</p> <p>There are four types of R group interactions which hold the tertiary structure in its shape.</p> <ul style="list-style-type: none"> <li>✓ van der Waals' forces (induced dipole-induced dipole bonding) exist when non-polar R groups (e.g. alkyl or aryl groups) come close together. They are usually found on the inside of globular proteins where, because they are hydrophobic, they do not interfere with solubility.</li> <li>✓ hydrogen bonding between polar groups (e.g.. <math>-\text{CH}_2\text{OH}</math>, <math>-\text{COOH}</math> and <math>-\text{NH}_2</math> groups).</li> <li>✓ ionic bonding eg. <math>-\text{COO}^-</math>, <math>-\text{NH}_3^+</math>, and <math>&gt;\text{NH}_2^+</math>.</li> <li>✓ disulfide linkages eg. <math>-\text{SH}</math> or <math>-\text{CH}_2-\text{S}-\text{S}-\text{CH}_2-</math> groups.</li> </ul>
70.	<p><b>Quaternary structure</b> of proteins refers to the spatial arrangement of its protein subunits. It shows how the individually folded protein subunits are packed together to yield large structures. This only applies to proteins that contain two or more polypeptide chains. The individual polypeptide chains are called the subunits. E.g. haemoglobin contains 4 subunits, each containing a haem group.</p> <p>It is stabilized by the same R-group interactions that stabilise the tertiary structure.</p>
71.	<p><b>Denaturation</b> is the loss of biological activity of a native protein. When proteins are denatured, the secondary and tertiary structures are disrupted i.e. the R group interactions are broken or destroyed. Note that the primary structure remains unaffected.</p> <p>Factors that can lead to denaturation include extremes in pH, temperature, ionic salts, heavy metal compounds, presence of organic solvents etc.</p>

All the best for 'A' Level ☺