
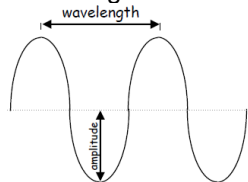









































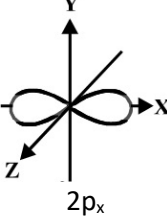
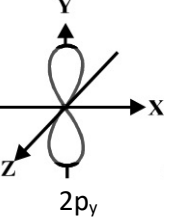
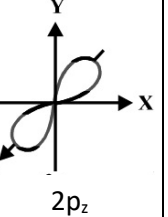
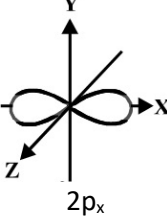
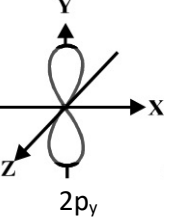
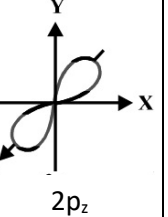
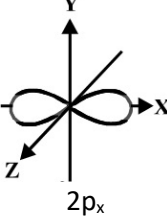
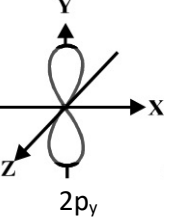
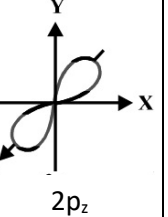
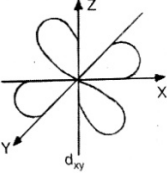
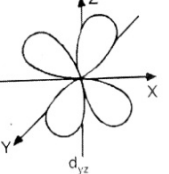
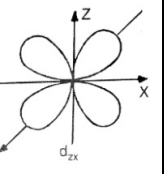
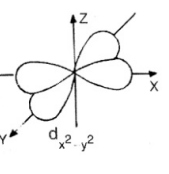
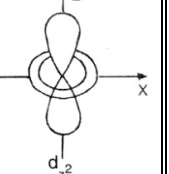
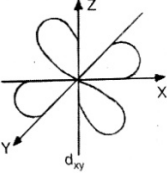
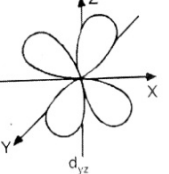
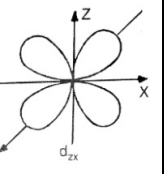
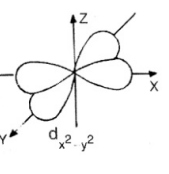
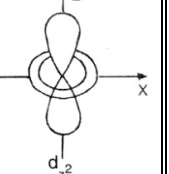
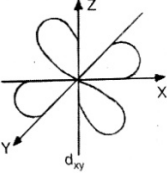
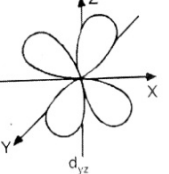
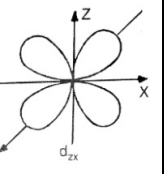
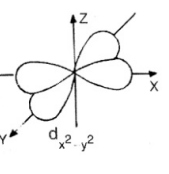
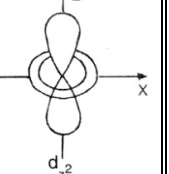
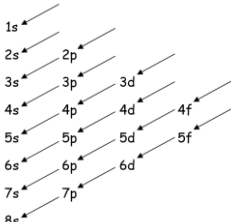
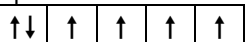
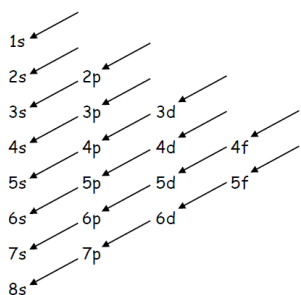
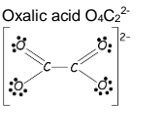
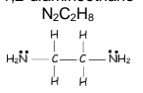
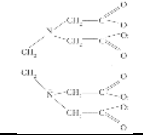
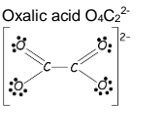
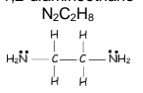
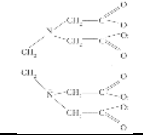
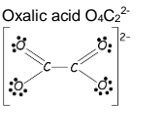
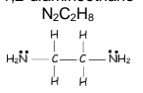
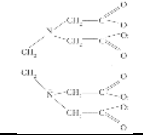
		AH Chemistry: Inorganic Chemistry		Section 1a: Electromagnetic Radiation & Atomic Spectra				Traffic Light																																
								red	amber	green																														
1 2 4	<p>Electromagnetic radiation are waves that have both wavelength and frequency</p>  <p>Wavelength λ : (visible light wavelengths are usually given in nanometres) the distance between wave crest to wave crest</p> <p>Frequency f : (frequency is measured in Hertz) the number of waves per second</p> <p>Speed c : All electromagnetic waves travel at $3 \times 10^8 \text{ m s}^{-1}$</p> <p>The equation $c = f \lambda$ shows the relationship between wavelength and frequency.</p> $c = f \times \lambda \qquad f = \frac{c}{\lambda} \qquad \lambda = \frac{c}{f}$																																							
3	<p>The radiation types of electromagnetic spectrum can be put in order of wavelength.</p> <table border="1"><thead><tr><th>EM Radiation</th><th>Gamma rays</th><th>X rays</th><th>UV radiation</th><th>Visible light</th><th>Infra-Red radiation</th><th>Microwaves</th><th>Radio & TV waves</th></tr></thead><tbody><tr><td>Wavelength</td><td>low</td><td colspan="4"></td><td></td><td>high</td></tr><tr><td>Frequency</td><td>high</td><td colspan="4"></td><td></td><td>low</td></tr><tr><td>Energy</td><td>high</td><td colspan="4"></td><td></td><td>low</td></tr></tbody></table>					EM Radiation	Gamma rays	X rays	UV radiation	Visible light	Infra-Red radiation	Microwaves	Radio & TV waves	Wavelength	low						high	Frequency	high						low	Energy	high						low			
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5	<p>Electromagnetic radiation has a dual nature. It can be described as</p> <ul style="list-style-type: none">a wave with wavelength and frequencya particle																																							
6 7 9	<p>Electromagnetic radiation can be absorbed or emitted by matter</p> <ul style="list-style-type: none">the radiation is behaving as a stream of particles called photonsphotons have quantised energy proportional to the frequency of the radiationhigher the frequency the higher the energy (lower the wavelength the higher the energy)photons in high frequency radiation can transfer greater amounts of energy than photons in low frequency radiation.																																							
8 13 14	<p>When a photon is absorbed, energy is gained by electrons being promoted to higher energy levels.</p> <p>When a photon is emitted, energy is lost by an excited electron moving from higher energy level to a lower energy level</p>																																							
10 11 12	<p>The energy associated by a single photon is: $E = h \times f$ or $E = \frac{hc}{\lambda}$</p> <p>As energy is often given in the unit kJ mol^{-1}</p> <p>The energy associated by a one mole of photon is: $E = L \times h \times f$ or $E = \frac{Lhc}{\lambda}$</p> <table border="1"><thead><tr><th>Symbol</th><th>Quantity</th><th>Units</th></tr></thead><tbody><tr><td>E</td><td>Energy</td><td>kJ mol^{-1}</td></tr><tr><td>L</td><td>Avogadro's Constant</td><td>$6.02 \times 10^{23} \text{ mol}^{-1}$</td></tr><tr><td>h</td><td>Plank's Constant</td><td>$6.63 \times 10^{-34} \text{ J s}$</td></tr><tr><td>f</td><td>Frequency</td><td>Hz or s^{-1}</td></tr><tr><td>λ</td><td>Wavelength</td><td>m</td></tr></tbody></table>					Symbol	Quantity	Units	E	Energy	kJ mol^{-1}	L	Avogadro's Constant	$6.02 \times 10^{23} \text{ mol}^{-1}$	h	Plank's Constant	$6.63 \times 10^{-34} \text{ J s}$	f	Frequency	Hz or s^{-1}	λ	Wavelength	m																	
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15 16	<p>Light energy emitted by an atom produces a spectrum that is made up of a series of lines at discrete (quantised) energy levels.</p> <ul style="list-style-type: none">this provides direct evidence for the existence of these energy levels.each element in sample produces characteristic absorption & emission spectra.These spectra can be used to identify and quantify the element.																																							
17 18	<p>In absorption spectroscopy, electromagnetic radiation is directed at an atomised sample.</p> <ul style="list-style-type: none">radiation is absorbed as electrons are promoted to higher energy levels.an absorption spectrum is produced by measuring how the intensity of absorbed light varies with wavelength.																																							
19 20	<p>In emission spectroscopy, high temperature is used to excite the electrons within atoms.</p> <ul style="list-style-type: none">As the electrons drop to lower energy levels, photons are emitted.emission spectrum of a sample is produced by measuring the intensity of light emitted at different wavelengths.																																							
21	<p>In atomic spectroscopy, the concentration of an element within a sample is related to the intensity of light emitted or absorbed.</p>																																							


	<h1>AH Chemistry: Inorganic Chemistry</h1> <h2>Section 1b: Atomic Orbitals and Electronic Configurations</h2>		<div>Traffic Light</div> <div>redambergreen</div>																																										
22 23 24	<p>Discrete lines observed in atomic spectra can be explained if electrons, like photons, also display the properties of both particles and waves.</p> <ul style="list-style-type: none">electrons behave as standing (stationary) waves in an atom and these are waves that vibrate in time but do not move in space.different sizes and shapes of standing wave possible around the nucleus, known as orbitals.		☺	☹	☹																																								
25	Orbitals can hold a maximum of two electrons.		☺	☹	☹																																								
26	There are four different shapes of orbitals, identified as s, p, d and f		☺	☹	☹																																								
	<table><tr><td>s orbital</td><td></td><td></td><td></td><td colspan="2">s orbitals are circular and increase in size as value of n increases.</td></tr></table>	s orbital							s orbitals are circular and increase in size as value of n increases.																																				
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f orbital	Not required to know f orbitals shapes for AH Chemistry.																																												
27	Electrons within atoms have fixed amounts of energy called quanta.		☺	☹	☹																																								
28a	<p>The principal quantum number (n) is the shell number of an energy level.</p> <ul style="list-style-type: none">the higher the value of n the larger the size of the s-orbital. <table><tr><td>Electron Shell</td><td>1st Shell</td><td>2nd Shell</td><td>3rd Shell</td><td>4th Shell</td><td>5th Shell</td></tr><tr><td>Principal Quantum number</td><td>n=1</td><td>n=2</td><td>n=3</td><td>n=4</td><td>n=5</td></tr></table>	Electron Shell	1 st Shell	2 nd Shell	3 rd Shell	4 th Shell	5 th Shell	Principal Quantum number	n=1	n=2	n=3	n=4	n=5		☺	☹	☹																												
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28b	<p>The angular momentum quantum numbers (l) describes the type of subshell within an electron shell.</p> <ul style="list-style-type: none">The values of l for each shell go from 0 up to n-1 <table><tr><td>Subshell Type</td><td>s</td><td>p</td><td>d</td><td>f</td></tr><tr><td>Angular Momentum Number</td><td>l=0</td><td>l=1</td><td>l=2</td><td>l=3</td></tr></table>	Subshell Type	s	p	d	f	Angular Momentum Number	l=0	l=1	l=2	l=3		☺	☹	☹																														
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Angular Momentum Number	l=0	l=1	l=2	l=3																																									
28c	<p>The magnetic quantum numbers (m_l) describes the orientation of the orbitals within a subshell.</p> <ul style="list-style-type: none">values of each orbital go from -l through 0 up to +l <table><tr><td>Subshell</td><td colspan="7">Values of Magnetic Quantum Number (m_l)</td></tr><tr><td>s (l=0)</td><td></td><td></td><td></td><td>0</td><td></td><td></td><td></td></tr><tr><td>p (l=1)</td><td></td><td></td><td>-1</td><td>0</td><td>+1</td><td></td><td></td></tr><tr><td>d (l=2)</td><td></td><td>-2</td><td>-1</td><td>0</td><td>+1</td><td>+2</td><td></td></tr><tr><td>f (l=3)</td><td>-3</td><td>-2</td><td>-1</td><td>0</td><td>+1</td><td>+2</td><td>+3</td></tr></table>	Subshell	Values of Magnetic Quantum Number (m _l)							s (l=0)				0				p (l=1)			-1	0	+1			d (l=2)		-2	-1	0	+1	+2		f (l=3)	-3	-2	-1	0	+1	+2	+3		☺	☹	☹
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f (l=3)	-3	-2	-1	0	+1	+2	+3																																						
28d	The spin magnetic quantum number (m _s) determines the spin direction of an electron and has values +½ or -½.		☺	☹	☹																																								
29a	<p>The aufbau principle states that electron orbitals fill up in order of increasing energy:</p> <p>1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p, 8s</p> 		☺	☹	☹																																								
29b	<p>Hund's rule states that electrons fill up orbitals singly first to maximise the number of parallel spins but filling each orbital with a second electron.</p> <p>e.g. iron atoms have 26 electrons and has an electronic configuration of</p> <p>1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁶ 4s²</p> <p>1st five d-electrons fill up singly 6th d electrons doubles up.</p> 		☺	☹	☹																																								

29c	The Pauli exclusion principle states that <ul style="list-style-type: none">two electrons in the same atom cannot have the same four quantum numbersno orbital can hold more than two electrons.		😊	😐	😞																																																																												
30	In isolated atoms, orbitals within each subshell are degenerate (equal in energy)		😊	😐	😞																																																																												
31	I can represent the relative energies corresponding to each orbital diagrammatically for the first four shells of a multi-electron atom using orbital box notation.		😊	😐	😞																																																																												
32	Electron configuration of the first 36 elements using orbital box or spectroscopic notation can be written. <table><tr><th>Element</th><th>Spectroscopic Notation</th><th>Electronic Configuration</th></tr><tr><td></td><td></td><th>Orbital Box Notation</th></tr><tr><td>Scandium</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹ 4s²</td><td><table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑</td><td></td><td></td><td></td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table></td></tr><tr><td>Vanadium</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d³ 4s²</td><td><table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑</td><td>↑</td><td>↑</td><td></td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table></td></tr><tr><td>Cobalt</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁷ 4s²</td><td><table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓</td><td>↑</td><td>↑</td><td>↑</td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table></td></tr></table>	Element	Spectroscopic Notation	Electronic Configuration			Orbital Box Notation	Scandium	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ¹ 4s ²	<table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑</td><td></td><td></td><td></td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table>	↑↓	↑↓	↑↓↑↓↑↓	↑↓	↑↓↑↓↑↓	↑				↑↓	1s	2s	2p	3s	3p	3d				4s	Vanadium	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ³ 4s ²	<table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑</td><td>↑</td><td>↑</td><td></td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table>	↑↓	↑↓	↑↓↑↓↑↓	↑↓	↑↓↑↓↑↓	↑	↑	↑		↑↓	1s	2s	2p	3s	3p	3d				4s	Cobalt	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁷ 4s ²	<table><tr><td>↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓↑↓↑↓</td><td>↑↓</td><td>↑↓</td><td>↑</td><td>↑</td><td>↑</td><td>↑↓</td></tr><tr><td>1s</td><td>2s</td><td>2p</td><td>3s</td><td>3p</td><td colspan="4">3d</td><td>4s</td></tr></table>	↑↓	↑↓	↑↓↑↓↑↓	↑↓	↑↓↑↓↑↓	↑↓	↑↓	↑	↑	↑	↑↓	1s	2s	2p	3s	3p	3d				4s		😊	😐	😞
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37	VSEPR (valence shell electron pair repulsion) theory is used to predict the shapes of molecules and polyatomic ions.																																																																																
38	The number of electron pairs surrounding a central atom is calculated by: $\text{Electron pairs} = \frac{\text{Number of electrons around central atom}}{2} + \text{number of bonds} - \text{charge}$		😊	😐	😞																																																																												
39	Electron pairs are negatively charged and repel each other <ul style="list-style-type: none">electron pairs are arranged to minimise repulsion and maximise separation.both lone pairs and bonding pairs take part in deciding the shape of the electron pairs around the central atom.																																																																																
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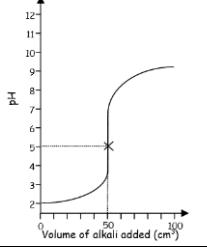
41 42	<p>The shapes of molecules and polyatomic ions are determined by the shapes adopted by the atoms present based on the arrangement of electron pairs.</p> <table border="1"> <tr> <td>2 electron pairs 2 bonding + 0 lone pairs</td><td>3 electron pairs 3 bonding + 0 lone pairs</td><td>4 electron pairs 4 bonding + 0 lone pairs</td><td>4 electron pairs 3 bonding + 1 lone pairs</td><td>4 electron pairs 2 bonding + 2 lone pairs</td></tr> <tr> <td></td><td></td><td></td><td></td><td></td></tr> <tr> <td>Linear</td><td>Trigonal Planar</td><td>Tetrahedral</td><td>Trigonal Pyramidal</td><td>Angular</td></tr> </table>	2 electron pairs 2 bonding + 0 lone pairs	3 electron pairs 3 bonding + 0 lone pairs	4 electron pairs 4 bonding + 0 lone pairs	4 electron pairs 3 bonding + 1 lone pairs	4 electron pairs 2 bonding + 2 lone pairs						Linear	Trigonal Planar	Tetrahedral	Trigonal Pyramidal	Angular		☺	☹	☹			
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43	<table border="1"> <tr> <td>5 electron pairs 5 bonding + 0 lone pairs</td><td>5 electron pairs 3 bonding + 2 lone pairs</td><td>6 electron pairs 6 bonding + 0 lone pairs</td><td>6 electron pairs 4 bonding + 2 lone pairs</td></tr> <tr> <td></td><td></td><td></td><td></td></tr> <tr> <td>Trigonal Bipyramidal</td><td>T-Shaped</td><td>Octahedral</td><td>Square Planar</td></tr> </table> <p>Electron pair repulsions decrease in strength in the order: non-bonding pair/non-bonding pair > non-bonding pair/bonding pair > bonding pair/bonding pair</p> <table border="1"> <tr> <td></td><td></td><td></td></tr> <tr> <td>4 Bonding Pairs + 0 lone pairs Angle between bonds = 109.5°</td><td>3 Bonding Pairs + 1 lone pairs Angle between bonds = 107°</td><td>2 Bonding Pairs + 2 lone pairs Angle between bonds = 104.5°</td></tr> </table> <ul style="list-style-type: none"> Bond angle in NH₃ is narrower than in CH₄ as lone pair pushes bonding pairs closer together due to the increased repulsion of a lone pair compared to a bonding pair. Bond angle in H₂O is even narrower due top two lone pairs pushing them closer together 	5 electron pairs 5 bonding + 0 lone pairs	5 electron pairs 3 bonding + 2 lone pairs	6 electron pairs 6 bonding + 0 lone pairs	6 electron pairs 4 bonding + 2 lone pairs					Trigonal Bipyramidal	T-Shaped	Octahedral	Square Planar				4 Bonding Pairs + 0 lone pairs Angle between bonds = 109.5°	3 Bonding Pairs + 1 lone pairs Angle between bonds = 107°	2 Bonding Pairs + 2 lone pairs Angle between bonds = 104.5°		☺	☹	☹
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44	Metals with an incomplete d subshell in at least one of their ions are called d-block transition metals						😊	😐	😞																					
45	<div>The aufbau principle states that orbitals of the lowest energy fill up first with electrons</div> <div></div> <div>Exceptions to aufbau rule include:</div> <table><thead><tr><th>Element</th><th>Electronic Configuration according to aufbau principle</th><th>Actual Electronic configuration</th><th>Reason</th></tr></thead><tbody><tr><td>Chromium</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁴ 4s²</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁵ 4s¹</td><td>Half-filled 3d⁵ preferred to full 4s²</td></tr><tr><td>Copper</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁹ 4s²</td><td>1s² 2s² 2p⁶ 3s² 3p⁶ 3d¹⁰ 4s¹</td><td>Full 3d¹⁰ preferred to full 4s²</td></tr></tbody></table>					Element	Electronic Configuration according to aufbau principle	Actual Electronic configuration	Reason	Chromium	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁴ 4s ²	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁵ 4s ¹	Half-filled 3d ⁵ preferred to full 4s ²	Copper	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁹ 4s ²	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ¹⁰ 4s ¹	Full 3d ¹⁰ preferred to full 4s ²		😊	😐	😞									
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46	Electrons are lost from the outer electrons shell first, regardless of the order they fill up in according to the aufbau principle. <ul style="list-style-type: none">In transition metals, 4s electrons are removed before 3d electrons when metal ions are formed.						😊	😐	😞																					
47	An element is in a particular oxidation state when it has a specific oxidation number. <ul style="list-style-type: none">oxidation number in a free or uncombined element is zerofor single atoms ions, the oxidation number is the same as the charge on the ionhydrogen usually has a oxidation number of +1 (except in hydrides)oxygen usually has an oxidation number of -2 (except in peroxides)The algebraic sum of all the oxidation numbers in a molecule must be zero e.g. In SO₃, three O atoms give 3x oxidation state of -2 and combine to equal -6 therefore the sulphur must have the oxidation state of +6.The algebraic sum of all the oxidation numbers in a polyatomic ion must be equal to the charge on the ion e.g. In SO₄²⁻, four O atoms give 4x oxidation state of -2 and combine to equal -8 therefore the sulphur must have the oxidation state of +6 to allow the overall charge to equal -2.						😊	😐	😞																					
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49	Transition metals can have different oxidation states in its compounds. <ul style="list-style-type: none">compounds of the same transition metal in different oxidation states may have different colours e.g.					<table><thead><tr><th>Ion</th><th>VO₃⁻</th><th>VO²⁺</th><th>V³⁺</th><th>V²⁺</th></tr></thead><tbody><tr><td>Oxidation State of Vanadium</td><td>+5</td><td>+4</td><td>+3</td><td>+2</td></tr><tr><td>Colour</td><td>Yellow</td><td>Blue</td><td>Green</td><td>Violet</td></tr></tbody></table>	Ion	VO ₃ ⁻	VO ²⁺	V ³⁺	V ²⁺	Oxidation State of Vanadium	+5	+4	+3	+2	Colour	Yellow	Blue	Green	Violet	😊	😐	😞						
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52	Reduction occurs when the oxidation number of a species decreases																													
53	Compounds containing metals in high oxidation states are often oxidising agents. <ul style="list-style-type: none">Oxidising agents are reduced themselves which reduces the oxidation number Compounds with metals in low oxidation states are often reducing agents. <ul style="list-style-type: none">Reducing agents are oxidised themselves which increases the oxidation number						😊	😐	😞																					
54	A ligand is defined as a molecule or ion electron donor which bonds to the metal ion by the donation of one or more electron pairs to unfilled metal ion orbitals.					<table><thead><tr><th>Type of Ligand</th><th colspan="2">Monodentate</th><th>Bidentate</th><th>Hexadentate</th></tr></thead><tbody><tr><td rowspan="5">Example</td><td colspan="2"><table><thead><tr><th>Neutral Ligands</th><th>Charged Ligands</th></tr></thead><tbody><tr><td>Water OH₂</td><td>Chloride Cl⁻</td></tr><tr><td>Ammonia NH₃</td><td>Cyanide CN⁻</td></tr><tr><td>Carbon Monoxide CO</td><td>Nitrite NO₂⁻</td></tr><tr><td></td><td>Hydroxide OH⁻</td></tr></tbody></table></td><td></td><td>1,2-diaminoethane N2C2H8 </td><td>Hexadentate E.D.T.A. </td></tr></tbody></table>	Type of Ligand	Monodentate		Bidentate	Hexadentate	Example	<table><thead><tr><th>Neutral Ligands</th><th>Charged Ligands</th></tr></thead><tbody><tr><td>Water OH₂</td><td>Chloride Cl⁻</td></tr><tr><td>Ammonia NH₃</td><td>Cyanide CN⁻</td></tr><tr><td>Carbon Monoxide CO</td><td>Nitrite NO₂⁻</td></tr><tr><td></td><td>Hydroxide OH⁻</td></tr></tbody></table>		Neutral Ligands	Charged Ligands	Water OH ₂	Chloride Cl ⁻	Ammonia NH ₃	Cyanide CN ⁻	Carbon Monoxide CO	Nitrite NO ₂ ⁻		Hydroxide OH ⁻		1,2-diaminoethane N2C2H8 	Hexadentate E.D.T.A. 	😊	😐	😞
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57	The total number of bonds of the ligand(s) to the central transition metal ion is called the co-ordination number <ul style="list-style-type: none">EDTA has a co-ordination number of 6hexaaquacopper(II) [Cu(OH₂)₆]²⁺ has a co-ordination number of 6 as the central Cu²⁺ ion is surrounded by 6 water moleculestetrachloridocuprate(II) [CuCl₄]²⁻ has a co-ordination number of 4 as the central Cu²⁺ ion is surrounded by 4 negative chloride ions						😊	😐	😞																					

56 58	Naming of Complexes from Formula <ul style="list-style-type: none">Ligands listed alphabetically followed by the name of the central metal ionNaming of ligands follow the following rules <table><tr><th>Neutral Ligand</th><th>Naming</th><th>Charged Ligands</th><th>Naming</th></tr><tr><td>Water</td><td>aqua</td><td>-ide ending ligand e.g. chloride</td><td>chlorido</td></tr><tr><td>Ammonia</td><td>ammine</td><td>-ate ending ligand e.g. oxalate</td><td>oxalato</td></tr><tr><td>Carbon monoxide</td><td>carbonyl</td><td>-ite ending ligand e.g. nitrite</td><td>nitrito</td></tr></table> <ul style="list-style-type: none">Mono, di, tri, tetra, penta, prefixes are used for multiple ligands of the same typeIf complex ion is overall a negative ion, the suffix <i>-ate</i> is added to the metal<ul style="list-style-type: none">nickel becomes nickelate(II)iron becomes ferrate(III) [not ironate]copper becomes cuprate(II) [not copperate]If complex ion is overall a positive ion, the metal does not have the suffix <i>-ate</i>The oxidation state of the metal is written after the metal (roman numerals in brackets) e.g. $[\text{Co}(\text{NH}_3)_6]^{2+}$ is hexaamminecobalt(II) $[\text{Fe}(\text{O}_4\text{C}_2)_3]^{3-}$ is trioxalatoferrate(III) Writing Formula from Names of Complexes. <ul style="list-style-type: none">formula of complex ions are written in square bracketsmetal symbol comes firstligands are listed alphabetically irrespective of being charged or neutralatom in ligand which donates pair of electrons written first e.g. OH_2 or $\text{O}_4\text{C}_2^{2-}$overall charge on complex ion written after square brackets e.g. tetrachloridocuprate(II) is written as $[\text{CuCl}_4]^{2-}$ hexaaquacopper(II) is written as $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$	Neutral Ligand	Naming	Charged Ligands	Naming	Water	aqua	-ide ending ligand e.g. chloride	chlorido	Ammonia	ammine	-ate ending ligand e.g. oxalate	oxalato	Carbon monoxide	carbonyl	-ite ending ligand e.g. nitrite	nitrito		☺	☹	☹
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59 60 61	In a complex of a transition metal, the d orbitals are no longer degenerate (equal in energy) <ul style="list-style-type: none">splitting of d orbitals to higher and lower energies occurs when the electrons present in approaching ligands cause the electrons in the orbitals lying along the axes to be repelled.weak field and strong field ligands affect energy differences between subsets of d orbitals.		☺	☹	☹																
62	Ligands can be placed in a spectrochemical series based on their ability to split d orbitals. $\text{I}^- < \text{Br}^- < \text{Cl}^- < \text{F}^- < \text{H}_2\text{O} < \text{NH}_3 < \text{CN}^-$		☺	☹	☹																
63 64 65	The colour of transition metal complexes can be explained in terms of d-d transitions. <ul style="list-style-type: none">$\text{dx}^2\text{-y}^2$ and dz^2 orbitals are raised to a higher energy level due to electrostatic repulsion from the ligands in the complextransition metals can absorb light because photons (at a particular wavelength) excite electrons in the lower d-orbitals (ground state) up to a higher energy d-orbital (excited state) <div><div><div>$\uparrow\downarrow$</div>$\text{d}_{x^2-y^2}$</div><div>$\uparrow\downarrow$</div>d_{z^2}</div> <div>$\uparrow\downarrow$</div> d_{xy} <div>$\uparrow\downarrow$</div> d_{xz} <div>$\uparrow\downarrow$</div> d_{yz} <div>absorption of energy</div> <div><div>$\uparrow\downarrow$</div>$\text{d}_{x^2-y^2}$</div> <div>$\uparrow\downarrow$</div> d_{z^2} <div>$\uparrow\downarrow$</div> d_{xy} <div>$\uparrow\downarrow$</div> d_{xz} <div>$\uparrow\downarrow$</div> d_{yz} <ul style="list-style-type: none">light of one colour is absorbed, then the complementary colour will be observed.electrons transition to higher energy levels when energy corresponding to the ultraviolet or visible regions of the electromagnetic spectrum is absorbed.		☺	☹	☹																
66 67 68	Transition metals can act as catalysts as they can form a variable number of bonds due to the availability of unoccupied and half-filled d-orbitals <ul style="list-style-type: none">allows the easier formation of intermediate complexesprovides reaction pathways of lower energy to proceedvariability of oxidation state of transition metals is important factor.transition metal reverts to original oxidation state once the reaction is complete		☺	☹	☹																
69a	Homogeneous catalysts are in the <u>same</u> state as the reactants.		☺	☹	☹																
69b 70	Heterogeneous catalysts are in the <u>different</u> state as the reactants. <ul style="list-style-type: none">Heterogeneous catalysts work by the adsorption of reactant molecules <table><tr><td></td><td></td><td></td><td></td></tr><tr><td>Reactant molecule collides with catalyst</td><td>Reactant molecule adsorbs to catalyst</td><td>Activated Complex forms</td><td>Product molecule(s) desorbs from catalyst</td></tr></table>					Reactant molecule collides with catalyst	Reactant molecule adsorbs to catalyst	Activated Complex forms	Product molecule(s) desorbs from catalyst		☺	☹	☹								
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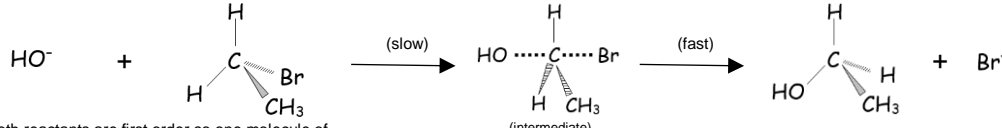
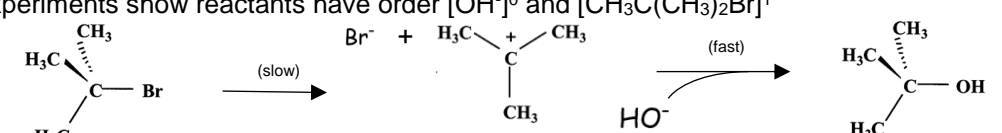
	AH Chemistry: Physical Chemistry	JAB chem	Traffic Light								
			red	amber	green						
71 72	A chemical reaction is in equilibrium when the composition of the reactants and products remains constant indefinitely. <ul style="list-style-type: none">equilibrium constant (<i>K</i>) characterises the equilibrium composition of the reaction mixture		☺	☹	☹						
73	For the general reaction:	$aA + bB \rightleftharpoons cC + dD$	☺	☹	☹						
	The equilibrium expression is:	$K = \frac{[A]^a [B]^b}{[C]^c [D]^d}$									
	where:	<div><div>[A], [B], [C] & [D] are the equilibrium concentrations of A, B, C and D</div><div>a, b, c & d are the stoichiometric coefficients in the balanced equation</div></div>									
74 76	<p>The value of equilibrium constant can be calculated:</p> <ul style="list-style-type: none">the equilibrium constant has no units. <p>e.g. 0.8mol of nitrogen and 1.5mol of hydrogen react by the Haber Process in a one litre container to give ammonia with an equilibrium concentration of 0.4mol.</p> <p>Equation: $N_2 + 3H_2 \rightleftharpoons 2NH_3$</p> <p>Mole ratio</p> <table><tr><td>1 mol</td><td>3mol</td><td>2mol</td></tr><tr><td>0.2mol</td><td>0.6mol</td><td>0.4mol</td></tr></table> <p>(Reactants Left) (0.6mol leftover) (0.9mol leftover)</p> <p>As container has 1 litre volume, the number of moles is equal to concentration in mol l⁻¹</p> $K = \frac{[NH_3]^2}{[N_2] [H_2]^3} = \frac{[0.4]^2}{[0.6] \times [0.9]^3} = 0.366$		1 mol	3mol	2mol	0.2mol	0.6mol	0.4mol	☺	☹	☹
1 mol	3mol	2mol									
0.2mol	0.6mol	0.4mol									
75	<p>The value of the equilibrium constant <i>K</i> indicates the position of equilibrium.</p> <ul style="list-style-type: none">A very high value of <i>K</i> (well above 1) indicated equilibrium far to the RIGHTA very low value of <i>K</i> (well below 1) indicated equilibrium far to the LEFT		☺	☹	☹						
77	Concentrations of pure solids and pure liquids at equilibrium are taken as constant and given a value of 1 in the equilibrium expression.		☺	☹	☹						
78	The numerical value of the equilibrium constant depends on the reaction temperature and is independent of concentration and/or pressure.		☺	☹	☹						
79	<div><div>For endothermic reactions</div><div><ul style="list-style-type: none">a rise in temperature causes an increase in <i>K</i>yield of the product is increased</div></div>	<div><div>For exothermic reactions</div><div><ul style="list-style-type: none">a rise in temperature causes a decrease in <i>K</i>yield of the product is decreased.</div></div>	☺	☹	☹						
80	The presence of a catalyst does not affect the value of the equilibrium constant.		☺	☹	☹						
81	<p>In water and aqueous solutions, water molecules form an equilibrium with hydronium and hydroxide ions. This ionisation of water can be represented by:</p> $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$ <p>water molecule water molecule hydronium ion hydroxide ion</p>		☺	☹	☹						
82	Hydronium ion has the formula H ₃ O ⁺ _(aq) and is a hydrated proton and is often represented by the shorthand H ⁺ _(aq)										
83	Water is described as amphoteric as it can act as an acid or a base.		☺	☹	☹						
84 85	<p>The dissociation constant <i>K_w</i> for the ionisation of water is known as the ionic product:</p> $K_w = [H_3O^+][OH^-]$ <ul style="list-style-type: none">value of <i>K_w</i> varies with temperature<i>K_w</i> is approximately 1 x 10⁻¹⁴ at 25°C.		☺	☹	☹						
86	<p>The relationship between pH and the hydronium H₃O⁺ ion concentration is given by:</p> $pH = -\log_{10}[H_3O^+] \quad [H_3O^+] = 10^{-pH}$		☺	☹	☹						
87	In water and aqueous solutions with a pH value of 7, the concentrations of H ₃ O ⁺ _(aq) and OH ⁻ _(aq) are both equal to 10 ⁻⁷ mol l ⁻¹ at 25°C.		☺	☹	☹						
88	<p>If the concentration of H₃O⁺_(aq) or OH⁻_(aq) is known, the concentration of the other ion can be calculated using the ionic product <i>K_w</i> (or by using pH + pOH = 14.)</p> <p>e.g. Calculate the [OH⁻] if [H₃O⁺] = 0.025mol l⁻¹</p> $[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1 \times 10^{-14}}{0.025} = 4 \times 10^{-13} \text{ mol l}^{-1}$		☺	☹	☹						

89 90 91	<p>The Brønsted-Lowry definition of acids and bases are:</p> <table><tr><td>Acid</td><td>Loses a proton (H⁺) to form the conjugate base</td></tr><tr><td>Base</td><td>Gains a proton (H⁺) to form the conjugate acid</td></tr><tr><td>Conjugate Acid</td><td>Formed when the base gains a proton (H⁺)</td></tr><tr><td>Conjugate Base</td><td>Formed when the acid loses a proton (H⁺)</td></tr></table> <p>For example:</p> $\text{CH}_3\text{COOH}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{CH}_3\text{COO}^{-}_{(\text{aq})} + \text{H}_3\text{O}^{+}_{(\text{aq})}$ <p style="text-align: center;">acid base conjugate base conjugate acid</p>	Acid	Loses a proton (H ⁺) to form the conjugate base	Base	Gains a proton (H ⁺) to form the conjugate acid	Conjugate Acid	Formed when the base gains a proton (H ⁺)	Conjugate Base	Formed when the acid loses a proton (H ⁺)		☺	☹	☹																
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92	<p>Strong acids fully dissociate into their ions e.g. $\text{HCl}(\text{g}) + \text{H}_2\text{O}(\text{l}) \longrightarrow \text{H}_3\text{O}^{+}(\text{aq}) + \text{Cl}^{-}(\text{aq})$</p> <p>Weak acids partially dissociate into their ions e.g. $\text{CH}_3\text{COOH}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COO}^{-}(\text{aq}) + \text{H}_3\text{O}^{+}(\text{aq})$</p>		☺	☹	☹																								
93 94	<p>Examples of strong and weak acids & alkalis include:</p> <table><tr><td>Strong Acid</td><td>Weak Acid</td><td>Strong Base</td><td>Weak Base</td></tr><tr><td>hydrochloric acid</td><td>ethanoic acid</td><td>Sodium hydroxide</td><td>Ammonia solution</td></tr><tr><td>sulphuric Acid</td><td>carbonic acid</td><td>Potassium hydroxide</td><td></td></tr><tr><td>nitric Acid</td><td>sulphurous acid</td><td>Lithium hydroxide</td><td></td></tr></table>	Strong Acid	Weak Acid	Strong Base	Weak Base	hydrochloric acid	ethanoic acid	Sodium hydroxide	Ammonia solution	sulphuric Acid	carbonic acid	Potassium hydroxide		nitric Acid	sulphurous acid	Lithium hydroxide			☺	☹	☹								
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95	<p>Carboxylic acids (and other weak acids) partially dissociate into ions, staying mainly as molecules.</p> <table><tr><td>Weak Acid</td><td>Equilibrium Equation</td></tr><tr><td>Ethanoic acid</td><td>$\text{CH}_3\text{COOH}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COO}^{-}(\text{aq}) + \text{H}_3\text{O}^{+}(\text{aq})$</td></tr><tr><td>Sulphur Dioxide solution</td><td>$\text{SO}_{2(\text{g})} + 2\text{H}_2\text{O}(\text{l}) \rightleftharpoons 2\text{H}_3\text{O}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$</td></tr><tr><td>Carbon Dioxide solution</td><td>$\text{CO}_{2(\text{g})} + 2\text{H}_2\text{O}(\text{l}) \rightleftharpoons 2\text{H}_3\text{O}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$</td></tr></table>	Weak Acid	Equilibrium Equation	Ethanoic acid	$\text{CH}_3\text{COOH}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{COO}^{-}(\text{aq}) + \text{H}_3\text{O}^{+}(\text{aq})$	Sulphur Dioxide solution	$\text{SO}_{2(\text{g})} + 2\text{H}_2\text{O}(\text{l}) \rightleftharpoons 2\text{H}_3\text{O}^{+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$	Carbon Dioxide solution	$\text{CO}_{2(\text{g})} + 2\text{H}_2\text{O}(\text{l}) \rightleftharpoons 2\text{H}_3\text{O}^{+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$		☺	☹	☹																
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96	<p>Ammonia, and amines, only partially dissociate into ions and mainly stays as molecules</p> <table><tr><td>Weak Base</td><td>Equilibrium Equation</td></tr><tr><td>Ammonia solution</td><td>$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$</td></tr><tr><td>1-aminomethane solution</td><td>$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$</td></tr></table>	Weak Base	Equilibrium Equation	Ammonia solution	$\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$	1-aminomethane solution	$\text{CH}_3\text{NH}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{CH}_3\text{NH}_3^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$		☺	☹	☹																		
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97	<p>Equimolar solutions of weak and strong acids/bases have the following properties:</p> <table><tr><td>Property</td><td>Strong Acid</td><td>Weak Acid</td><td>Property</td><td>Strong Base</td><td>Weak Base</td></tr><tr><td>pH Value</td><td>lower</td><td>Higher (Nearer pH=7)</td><td>pH Value</td><td>higher</td><td>Lower (nearer pH=7)</td></tr><tr><td>Conductivity</td><td>Higher</td><td>Lower</td><td>Conductivity</td><td>Higher</td><td>Lower</td></tr><tr><td>Reaction Rate</td><td>Higher</td><td>Lower</td><td>Reaction Rate</td><td>Higher</td><td>Lower</td></tr></table>	Property	Strong Acid	Weak Acid	Property	Strong Base	Weak Base	pH Value	lower	Higher (Nearer pH=7)	pH Value	higher	Lower (nearer pH=7)	Conductivity	Higher	Lower	Conductivity	Higher	Lower	Reaction Rate	Higher	Lower	Reaction Rate	Higher	Lower		☺	☹	☹
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98	<p>The acid dissociation constant or the equation $\text{HA} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^{+} + \text{A}^{-}$ is:</p> $K_a = \frac{[\text{H}_3\text{O}^{+}][\text{A}^{-}]}{[\text{HA}]}$ <p>∴ pK_a = -log₁₀ K_a</p> <p>NB AS H₂O is both a reactant and the solvent, [H₂O] is given the value = 1 and cancels out of equation.</p>		☺	☹	☹																								
99	<p>The approximate pH of a weak acid can be calculated using:</p> $\text{pH} = \frac{1}{2}\text{p}K_a - \frac{1}{2}\log_{10}c$ <p>e.g. calculate the pH of 0.25mol l⁻¹ solution of ethanoic acid (pK_a = 4.76)</p> $\begin{aligned} \text{pH} &= \frac{1}{2}\text{p}K_a - \frac{1}{2}\log_{10}c \\ \text{pH} &= \frac{1}{2} \times 4.76 - \frac{1}{2} \times \log_{10}(0.25) \\ \text{pH} &= 2.38 - \frac{1}{2} \times (-0.60) \\ \text{pH} &= 2.38 - (-0.30) \\ \text{pH} &= 2.68 \end{aligned}$		☺	☹	☹																								
100	<table><tr><td>Acid Type</td><td>Base Type</td><td>pH of solution of Soluble Salt formed</td></tr><tr><td>Strong Acid</td><td>Strong Base</td><td>Neutral solution</td></tr><tr><td>Weak Acid</td><td>Strong Base</td><td>Alkaline solution</td></tr><tr><td>Strong Acid</td><td>Weak Base</td><td>Acidic Solution</td></tr></table>	Acid Type	Base Type	pH of solution of Soluble Salt formed	Strong Acid	Strong Base	Neutral solution	Weak Acid	Strong Base	Alkaline solution	Strong Acid	Weak Base	Acidic Solution		☺	☹	☹												
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101	<p>The names of salts are worked out from the individual acids and bases used:</p> <table><tr><td>Acid Used</td><td>Base Used</td><td>Name of Salt</td><td>Acid Used</td><td>Base Used</td><td>Name of Salt</td></tr><tr><td>hydrochloric acid</td><td>sodium hydroxide</td><td>sodium chloride</td><td>ethanoic acid</td><td>magnesium hydroxide</td><td>magnesium ethanoate</td></tr><tr><td>sulphuric acid</td><td>Potassium hydroxide</td><td>potassium sulphate</td><td>sulphurous acid</td><td>calcium hydroxide</td><td>calcium sulphite</td></tr><tr><td>nitric acid</td><td>lithium hydroxide</td><td>lithium nitrate</td><td>carbonic acid</td><td>ammonia solution</td><td>ammonium carbonate</td></tr></table>	Acid Used	Base Used	Name of Salt	Acid Used	Base Used	Name of Salt	hydrochloric acid	sodium hydroxide	sodium chloride	ethanoic acid	magnesium hydroxide	magnesium ethanoate	sulphuric acid	Potassium hydroxide	potassium sulphate	sulphurous acid	calcium hydroxide	calcium sulphite	nitric acid	lithium hydroxide	lithium nitrate	carbonic acid	ammonia solution	ammonium carbonate		☺	☹	☹
Acid Used	Base Used	Name of Salt	Acid Used	Base Used	Name of Salt																								
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nitric acid	lithium hydroxide	lithium nitrate	carbonic acid	ammonia solution	ammonium carbonate																								
102	<p>Salt solutions can have different concentrations of H₃O⁺(aq) and OH⁻(aq):</p> <table><tr><td>Sodium ethanoate solution has pH greater than 7</td><td>Ammonium chloride solution has pH less than 7</td></tr><tr><td>Sodium ethanoate solid fully dissociates into ions on dissolving. Ethanoate ions collide with H₃O⁺ ions to form molecules of ethanoic acid: $\text{CH}_3\text{COO}^{-}(\text{aq}) + \text{H}_3\text{O}^{+}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ <small>ethanoate ion hydronium ion ethanoic acid molecule water</small></td><td>Ammonium chloride solid fully dissociates into ions on dissolving. Ammonium ions collide with hydroxide ions to form molecules of NH₃: $\text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$ <small>ammonium ion hydroxide ion ammonia molecule water</small></td></tr><tr><td>H₃O⁺(aq) ions removed from solution as they join up with CH₃COO⁻(aq) ions. Equilibrium in water shifts to RIGHT to replace missing H₃O⁺(aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small></td><td>OH⁻(aq) ions removed from solution as they join up with NH₄⁺(aq) ions. Equilibrium in water shifts to RIGHT to replace missing OH⁻(aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small></td></tr><tr><td>[OH⁻] > [H₃O⁺] as H₃O⁺(aq) ions are removed and OH⁻(aq) build up</td><td>[H₃O⁺] > [OH⁻] as OH⁻(aq) ions are removed and H₃O⁺(aq) build up</td></tr></table>	Sodium ethanoate solution has pH greater than 7	Ammonium chloride solution has pH less than 7	Sodium ethanoate solid fully dissociates into ions on dissolving. Ethanoate ions collide with H ₃ O ⁺ ions to form molecules of ethanoic acid: $\text{CH}_3\text{COO}^{-}(\text{aq}) + \text{H}_3\text{O}^{+}(\text{aq}) \rightleftharpoons \text{CH}_3\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ <small>ethanoate ion hydronium ion ethanoic acid molecule water</small>	Ammonium chloride solid fully dissociates into ions on dissolving. Ammonium ions collide with hydroxide ions to form molecules of NH ₃ : $\text{NH}_4^{+}(\text{aq}) + \text{OH}^{-}(\text{aq}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l})$ <small>ammonium ion hydroxide ion ammonia molecule water</small>	H ₃ O ⁺ (aq) ions removed from solution as they join up with CH ₃ COO ⁻ (aq) ions. Equilibrium in water shifts to RIGHT to replace missing H ₃ O ⁺ (aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small>	OH ⁻ (aq) ions removed from solution as they join up with NH ₄ ⁺ (aq) ions. Equilibrium in water shifts to RIGHT to replace missing OH ⁻ (aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small>	[OH ⁻] > [H ₃ O ⁺] as H ₃ O ⁺ (aq) ions are removed and OH ⁻ (aq) build up	[H ₃ O ⁺] > [OH ⁻] as OH ⁻ (aq) ions are removed and H ₃ O ⁺ (aq) build up		☺	☹	☹																
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H ₃ O ⁺ (aq) ions removed from solution as they join up with CH ₃ COO ⁻ (aq) ions. Equilibrium in water shifts to RIGHT to replace missing H ₃ O ⁺ (aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small>	OH ⁻ (aq) ions removed from solution as they join up with NH ₄ ⁺ (aq) ions. Equilibrium in water shifts to RIGHT to replace missing OH ⁻ (aq) ions. $\text{H}_2\text{O}(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^{+}(\text{aq}) + \text{OH}^{-}(\text{aq})$ <small>water water hydronium ion hydroxide ion</small>																												
[OH ⁻] > [H ₃ O ⁺] as H ₃ O ⁺ (aq) ions are removed and OH ⁻ (aq) build up	[H ₃ O ⁺] > [OH ⁻] as OH ⁻ (aq) ions are removed and H ₃ O ⁺ (aq) build up																												
103	<p>Buffer solutions have a pH which remains approximately constant when small amounts of acid, base or water are added.</p> <ul style="list-style-type: none">Large amounts of acid or base will overpower the buffer solution.		☺	☹	☹																								

104	I can describe what an acidic and a basic buffer consists of.				
	Type	Description	Example		
	Acid Buffer	salt of weak acid dissolved in a weak acid	sodium ethanoate dissolved in ethanoic acid		
	Basic Buffer	salt of weak base dissolved in a weak base	ammonium chloride dissolved in ammonia solution		
105	Acid Buffer		Basic Buffer		
	e.g. sodium ethanoate dissolved in ethanoic acid solution. large concentration from weak acid large concentration from dissolved salt $\text{CH}_3\text{COOH}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{CH}_3\text{COO}^{-}_{(\text{aq})} + \text{H}_3\text{O}^{+}_{(\text{aq})}$ ethanoic acid molecule water ethanoate ion hydronium ion When acid is added to buffer: Equilibrium shifts to LEFT as added $\text{H}_3\text{O}^{+}_{(\text{aq})}$ in added acid join up with ethanoate ions and form ethanoic acid molecules. When alkali is added to buffer: Equilibrium shifts to RIGHT as $\text{H}_3\text{O}^{+}_{(\text{aq})}$ ions are neutralised by the alkali and ethanoic acid molecules dissociate into ions to replace $\text{H}_3\text{O}^{+}_{(\text{aq})}$ ions.		e.g. ammonium chloride dissolved in ammonia solution. large concentration from weak acid large concentration from dissolved salt $\text{NH}_3_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{NH}_4^{+}_{(\text{aq})} + \text{OH}^{-}_{(\text{aq})}$ ammonia molecule water ammonium ion hydroxide ion When acid is added to buffer: Equilibrium shifts to RIGHT as $\text{OH}^{-}_{(\text{aq})}$ ions are neutralised by the acid and ammonia NH_3 molecules dissociate into ions to replace $\text{OH}^{-}_{(\text{aq})}$ ions. When alkali is added to buffer: Equilibrium shifts to LEFT as added $\text{OH}^{-}_{(\text{aq})}$ in added alkali join up with ammonium NH_4^{+} ions and form ammonia molecules.		
106	The approximate pH of a buffer solution is calculated using: e.g. Calculate the pH of a buffer where 3.74g of sodium ethanoate (CH_3COONa) is dissolved 0.20mol l^{-1} ethanoic acid and the final volume of the buffer is 100 cm^3 . $\text{gfm } \text{CH}_3\text{COONa} = (2 \times 12) + (3 \times 1) + (2 \times 16) + (1 \times 23) = 82 \text{ g mol}^{-1}$ $\text{no. of mol} = \frac{\text{mass}}{\text{gfm}} = \frac{3.74}{82} = 0.0456 \text{ mol}$ $\text{concentration} = \frac{\text{no. of mol}}{\text{volume}} = \frac{0.0456 \text{ mol}}{0.1 \text{ litres}} = 0.456 \text{ mol l}^{-1}$ $\text{pH} = \text{pK}_a - \log_{10} \frac{[\text{acid}]}{[\text{salt}]} = 4.76 - \log_{10} \frac{0.2}{0.456} = 4.76 - \log_{10}(0.439)$ $= 4.76 - (-0.358)$ $= 5.12$				
	Indicators are weak acids in which the dissociation can be represented as: $\text{HIn}_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})} \rightleftharpoons \text{H}_3\text{O}^{+}_{(\text{aq})} + \text{In}^{-}_{(\text{aq})}$				
108	The dissociation constant K_{In} for an acid indicator is: $K_{\text{In}} = \frac{[\text{H}_3\text{O}^{+}] [\text{In}^{-}]}{[\text{HIn}]}$				
	The colour of an acid indicator is distinctly different from that of its conjugate base. <ul style="list-style-type: none"> the colour of an indicator is determined by the ratio of $[\text{HIn}]$ to $[\text{In}^{-}]$ the theoretical point at which colour changes is when $[\text{H}_3\text{O}^{+}] = K_{\text{In}}$ the colour change is assumed to be distinguished when $[\text{HIn}]$ and $[\text{In}^{-}]$ differ by a factor of 10 the pH range over which a colour change occurs can be estimated by the expression: $\text{pH} = \text{pK}_a \pm 1$ 				
114	Suitable indicators can be selected from pH data or titration curves. Which of the following indicators should be used in the titration of aqueous potassium hydroxide solution with aqueous ethanoic acid A. Phenolphthalein (pH range 8.3-10.0) B. Bromothymol blue (pH range 6.0-7.6) C. Methyl Red (pH range 4.2-6.3) D. Methyl Orange (pH range 3.1-4.4)		Which indicator would be best in the following titration of ammonium hydroxide and sodium hydroxide?  A. Phenolphthalein (pH range 8.3-10.0) B. Bromothymol blue (pH range 6.0-7.6) C. Methyl Red (pH range 4.2-6.3) D. Phenol Red (pH range 6.8-8.4)		
	Answer: Potassium hydroxide reacting with ethanoic acid will produce a salt solution with a pH in the alkaline region of the pH scale. Phenolphthalein is the only listed indicator where the colour change pH range is entirely in the alkaline region of the pH scale.		Answer: The titration curve clearly shows the neutralisation point to be around pH=5 so an indicator with a pH range of around 4.0-6.0 would be best used. Methyl Red is the only list indicator which is close to this pH range.		

	AH Chemistry: Physical Chemistry	JAB chem	Traffic Light												
			red	amber	green										
115	The standard enthalpy of formation, ΔH° _f is defined as the enthalpy change for the formation of one mole of a substance from its elements in their natural state. e.g. Enthalpy of combustion of ethanol: 2C _(s) + 3H _{2(g)} + ½O _{2(g)} → C ₂ H ₅ OH _(l)			😊	😐	😞									
116	Standard state of a substance is its most stable state at a pressure of 1 atmosphere and a specified temperature (usually 298K).			😊	😐	😞									
117	I can calculate the standard enthalpy change of a reaction using: $\Delta H^\circ = \sum \Delta H^\circ_f (\text{products}) - \sum \Delta H^\circ_f (\text{reactants})$ e.g. calculate ΔH° for the following reaction: 2ZnS _(s) + 3O _{2(g)} → 2ZnO _(s) + 2SO _{2(g)}			😊	😐	😞									
	<table><tr><th>Substance</th><th>ΔH°_f (kJ mol⁻¹)</th></tr><tr><td>2ZnS_(s)</td><td>-206</td></tr><tr><td>O_{2(g)}</td><td>0</td></tr><tr><td>ZnO_(s)</td><td>-350</td></tr><tr><td>2SO_{2(g)}</td><td>-297</td></tr></table>	Substance	ΔH° _f (kJ mol ⁻¹)				2ZnS _(s)	-206	O _{2(g)}	0	ZnO _(s)	-350	2SO _{2(g)}	-297	$\begin{aligned} \Delta H^\circ &= \sum \Delta H^\circ_f (\text{products}) - \sum \Delta H^\circ_f (\text{reactants}) \\ &= (2 \times -350) + (2 \times -297) - (2 \times -206) + (3 \times 0) \\ &= (-700 - 594) - (-412 - 0) \\ &= -1294 - (-412) \\ &= -882 \text{ kJ mol}^{-1} \end{aligned}$
	Substance	ΔH° _f (kJ mol ⁻¹)													
	2ZnS _(s)	-206													
	O _{2(g)}	0													
ZnO _(s)	-350														
2SO _{2(g)}	-297														
118	Entropy (S) is a measure of the degree of disorder of a system			😊	😐	😞									
119	• the greater the degree of disorder, the greater the entropy														
120	• solids have lower values of entropy than gases.														
121	• entropy increases as temperature increases														
	• there is a large change in entropy at a substance's melting and boiling point <ul style="list-style-type: none">no change in temperature as state changes but large increase in entropy/disorder as solids turn into liquids or liquids turn into gas														
122	Second law of thermodynamics states that the total entropy of a reaction system and its surroundings always increases for a spontaneous process.			😊	😐	😞									
123	When heat is released by a reaction system to the surroundings there is an increase in the entropy (disorder) of the surroundings.			😊	😐	😞									
124	• when heat is absorbed by a reaction system to the surroundings there is a decrease in the entropy (disorder) of the surroundings.														
125	Third law of thermodynamics states that the entropy of a perfect crystal at 0 K is zero.			😊	😐	😞									
126	The standard entropy of a substance is the entropy content of one mole of a substance at 1atm pressure and 298K			😊	😐	😞									
127	I can calculate the change in standard enthalpy of a reaction using: $\Delta S^\circ = \sum \Delta S^\circ (\text{products}) - \sum \Delta S^\circ (\text{reactants})$ e.g. calculate ΔS° for the following reaction: 2ZnS _(s) + 3O _{2(g)} → 2ZnO _(s) + 2SO _{2(g)}			😊	😐	😞									
	<table><tr><th>Substance</th><th>ΔS° (J K⁻¹ mol⁻¹)</th></tr><tr><td>2ZnS_(s)</td><td>58</td></tr><tr><td>O_{2(g)}</td><td>205</td></tr><tr><td>ZnO_(s)</td><td>44</td></tr><tr><td>2SO_{2(g)}</td><td>248</td></tr></table>	Substance	ΔS° (J K ⁻¹ mol ⁻¹)				2ZnS _(s)	58	O _{2(g)}	205	ZnO _(s)	44	2SO _{2(g)}	248	$\begin{aligned} \Delta S^\circ &= \sum \Delta S^\circ (\text{products}) - \sum \Delta S^\circ (\text{reactants}) \\ &= (2 \times 44) + (2 \times 248) - (2 \times 58) + (3 \times 205) \\ &= 88 + 496 - 116 + 615 \\ &= 1083 \text{ J K}^{-1} \text{ mol}^{-1} \end{aligned}$
	Substance	ΔS° (J K ⁻¹ mol ⁻¹)													
	2ZnS _(s)	58													
	O _{2(g)}	205													
ZnO _(s)	44														
2SO _{2(g)}	248														
128	I know that the change in free energy for a reaction is related to the enthalpy and entropy changes by: $\Delta G^\circ = \Delta H^\circ - T \Delta S^\circ$ Free Energy (kJ mol ⁻¹) = Enthalpy Change (kJ mol ⁻¹) - Temperature (K) x Entropy Change (kJ K ⁻¹ mol ⁻¹) e.g. calculate ΔG° for the following reaction: 2ZnS _(s) + 3O _{2(g)} → 2ZnO _(s) + 2SO _{2(g)} at 7000K			😊	😐	😞									
	$\Delta G^\circ = \Delta H^\circ - T \times \Delta S^\circ$														
	$\Delta G^\circ = -882 - 7000 \times \frac{-147}{1000}$														
	$\Delta G^\circ = -882 - 5000 \times -0.147$														
	$\Delta G^\circ = -882 - (-735)$														
	$\Delta G^\circ = -147 \text{ kJ mol}^{-1}$														
	Please note: entropy values of ΔS° are usually given in J K ⁻¹ mol ⁻¹ so must be divided by 1000 to get them into kJ K ⁻¹ mol ⁻¹ when using this equation.														
129	When the change in free energy (ΔG°) between reactants and products is negative , a reaction may occur and the reaction is said to be feasible .			😊	😐	😞									
130	A feasible reaction is one that tends towards the products rather than the reactants. This does not give any indication of the rate of the reaction.			😊	😐	😞									
131	The standard free energy change for a reaction can be calculated using: $\Delta G^\circ = \sum \Delta G^\circ (\text{products}) - \sum \Delta G^\circ (\text{reactants})$			😊	😐	😞									

132 135	The feasibility of a chemical reaction under standard conditions can be predicted from the calculated value of the change in standard free energy (ΔG°).						
	When $\Delta G^\circ < 0$ Reaction is feasible	When $\Delta G^\circ = 0$ Reaction is just feasible (reaction is in equilibrium)	When $\Delta G^\circ > 0$ Reaction is not feasible				
133	<p>I can estimate the temperatures at which a reaction may be feasible by considering the range of values of T for which $\Delta G^\circ < 0$.</p> <p>e.g. calculate the temperature when the reaction becomes feasible:</p> $2\text{ZnS(s)} + 3\text{O}_{2\text{(g)}} \rightarrow 2\text{ZnO(s)} + 2\text{SO}_{2\text{(g)}}$ $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = 0 \quad \therefore T = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{-882 \times 1000 \text{ J mol}^{-1}}{-147 \text{ J K}^{-1} \text{ mol}^{-1}} = 6000\text{K}$						
134	Any reaction is feasible if ΔG is negative, even under non-standard conditions						
136	A reversible reaction will proceed spontaneously until the composition is reached where $\Delta G = 0$.						

<div></div> <div>AH Chemistry: Physical Chemistry</div> <div>Section 2c: Kinetics</div>		<div>JAB</div> <div>chem</div>		Traffic Light																																											
				red	amber	green																																									
137	Chemical reactions normally depend on the concentration of reactants																																														
138 139	<div>The order of reaction with respect to each reactant can be determined by changing the concentration of each reactant individually.</div> <div><ul style="list-style-type: none">The order of reaction relates each reacting species to the rate of reaction.</div> <div>For the reaction: $A + B + C \longrightarrow D$</div> <div><ul style="list-style-type: none">Each reactant is varied one at a time and the reaction rate is measured:</div> <table><thead><tr><th>Experiment</th><th>[A] (mol l⁻¹)</th><th>[B] (mol l⁻¹)</th><th>[C] (mol l⁻¹)</th><th>Initial Rate of Reaction (mol l⁻¹ s⁻¹)</th></tr></thead><tbody><tr><td>1</td><td>1.0</td><td>1.0</td><td>1.0</td><td>2.0</td></tr><tr><td>2</td><td>2.0</td><td>1.0</td><td>1.0</td><td>4.0</td></tr><tr><td>3</td><td>1.0</td><td>2.0</td><td>1.0</td><td>2.0</td></tr><tr><td>4</td><td>1.0</td><td>1.0</td><td>2.0</td><td>8.0</td></tr></tbody></table> <div><ul style="list-style-type: none">Each reactant is compared to determine the effect of change of that reactant on the reaction rate</div> <table><thead><tr><th>Experiments Compared</th><th>Change in Conditions</th><th>Effect of Change on Rate</th><th>Order of Reactant</th></tr></thead><tbody><tr><td>1+2</td><td>[A] x2</td><td>rate x2</td><td>[A]¹</td></tr><tr><td>1+3</td><td>[B] x2</td><td>no change in rate</td><td>[B]⁰</td></tr><tr><td>1+4</td><td>[C] x2</td><td>rate x4</td><td>[C]²</td></tr></tbody></table> <div><ul style="list-style-type: none">The order of the reactant can be assessing the effect of changing concentration on rate</div>			Experiment	[A] (mol l ⁻¹)	[B] (mol l ⁻¹)	[C] (mol l ⁻¹)	Initial Rate of Reaction (mol l ⁻¹ s ⁻¹)	1	1.0	1.0	1.0	2.0	2	2.0	1.0	1.0	4.0	3	1.0	2.0	1.0	2.0	4	1.0	1.0	2.0	8.0	Experiments Compared	Change in Conditions	Effect of Change on Rate	Order of Reactant	1+2	[A] x2	rate x2	[A] ¹	1+3	[B] x2	no change in rate	[B] ⁰	1+4	[C] x2	rate x4	[C] ²			
Experiment	[A] (mol l ⁻¹)	[B] (mol l ⁻¹)	[C] (mol l ⁻¹)	Initial Rate of Reaction (mol l ⁻¹ s ⁻¹)																																											
1	1.0	1.0	1.0	2.0																																											
2	2.0	1.0	1.0	4.0																																											
3	1.0	2.0	1.0	2.0																																											
4	1.0	1.0	2.0	8.0																																											
Experiments Compared	Change in Conditions	Effect of Change on Rate	Order of Reactant																																												
1+2	[A] x2	rate x2	[A] ¹																																												
1+3	[B] x2	no change in rate	[B] ⁰																																												
1+4	[C] x2	rate x4	[C] ²																																												
140 141 142	<div>The rate equation for a reaction can be written using the individual orders for each reactant.</div> <div>$\text{Rate} = k [A]^1 \times [B]^0 \times [C]^2$$\text{Rate} = k [A] \times 1 \times [C]^2$$\text{Rate} = k [A] \times [C]^2$</div>																																														
143	<div>The overall rate of a reaction can be determined from the rate equation.</div> <div>$\text{Overall Order} = 1 + 0 + 2 = 3$</div>																																														
144	The order of a reaction can only be determined from experimental data.																																														
145	<div>Using the rate equation and data of initial rate in the results table, the value of the rate constant and units of the rate constant can be calculated.</div> <div><ul style="list-style-type: none">reactants can be zero, first, second or third order</div> <div>$\begin{aligned} \text{Rate} &= k [A] \times [C]^2 \\ 2.0 \text{ mol l}^{-1} \text{ s}^{-1} &= k [1.0 \text{ mol l}^{-1}] \times [1.0 \text{ mol l}^{-1}]^2 \\ k &= \frac{2.0 \text{ mol l}^{-1} \text{ s}^{-1}}{1.0 \text{ mol l}^{-1} \times 1.0 \text{ mol}^2 \text{ l}^{-2}} \\ k &= \frac{2.0 \text{ mol l}^{-1} \text{ s}^{-1}}{1.0 \text{ mol}^3 \text{ l}^{-3}} \\ k &= 2.0 \text{ l}^2 \text{ mol}^{-2} \text{ s}^{-1} \end{aligned}$</div>																																														
146 147	<div>Reactions usually occur by a series of steps called a reaction mechanism. The rate of the reaction is dependent on the slowest step called the rate determining step.</div>																																														
148	<div>Reaction mechanisms can be worked out from experimentally determined rate equations</div> <div>For reaction: $\text{CH}_3\text{CH}_2\text{Br} + \text{OH}^- \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{Br}^-$</div> <div>Experiments show reactants have order $[\text{OH}^-]^1$ and $[\text{CH}_3\text{CH}_2\text{Br}]^1$</div> <div></div> <div>Both reactants are first order so one molecule of each takes part in the slow rate determining step</div> <div>For the reaction: $\text{CH}_3\text{C}(\text{CH}_3)_2\text{Br} + \text{OH}^- \rightarrow \text{CH}_3\text{C}(\text{CH}_3)_2\text{OH} + \text{Br}^-$</div> <div>Experiments show reactants have order $[\text{OH}^-]^0$ and $[\text{CH}_3\text{C}(\text{CH}_3)_2\text{Br}]^1$</div> <div></div> <div>Only $\text{CH}_3\text{C}(\text{CH}_3)_2\text{Br}$ is first order so is the only chemical involved in the slow rate determining step</div>			