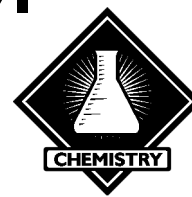




Dalziel High School

Chemistry

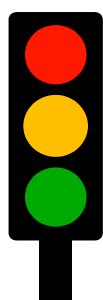


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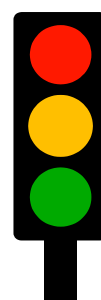
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Self-Evaluation









CE Higher Unit 1 Chemical Changes & Structure










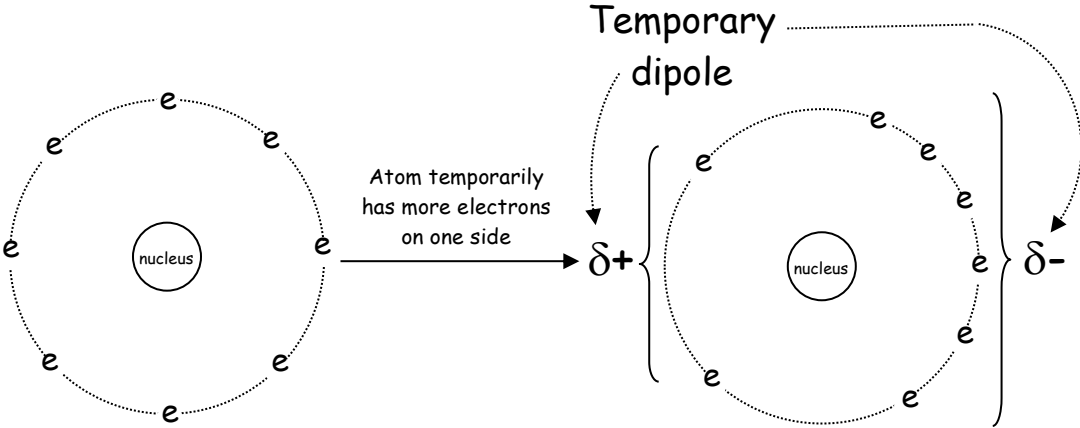
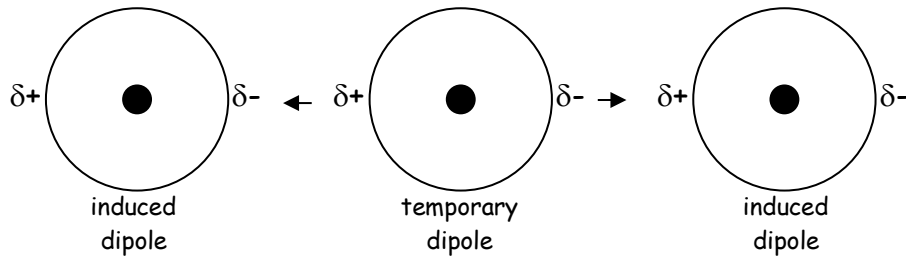








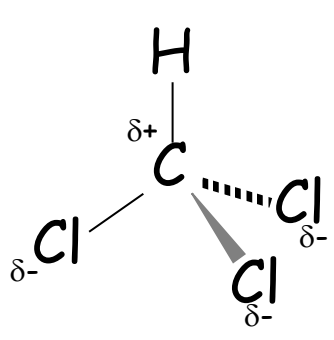
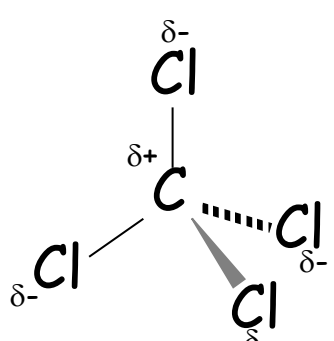
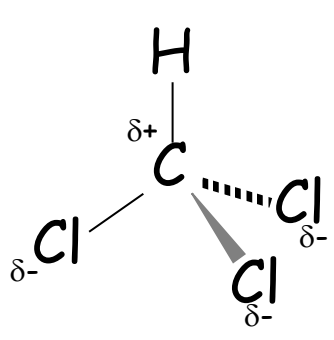
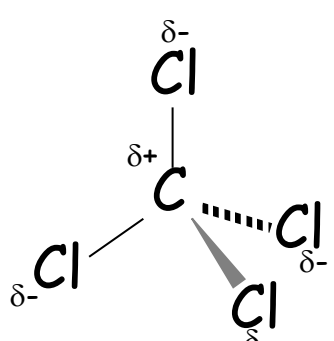
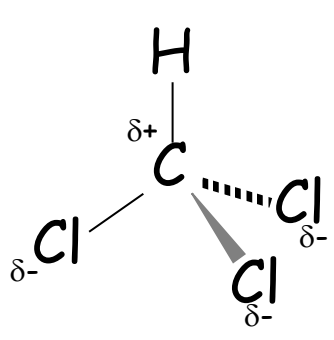
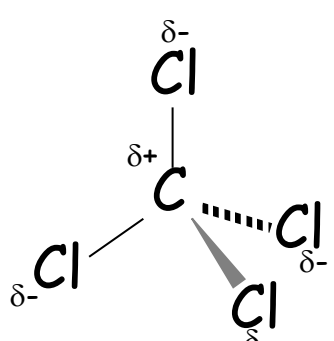
Section	Title	Completed
1.1a	Collision Theory	
1.1b	Reaction Profiles	
1.1c	Temperature & Kinetic Energy	
1.1d	Catalysts	
1.2a	Arrangement of Elements in Periodic Table	
1.2b	Bonding & Structure of 1 st 20 Elements	
1.2c	Periodic Trends in I.E. and Cov Radii	
1.2d	Periodic Trends in Electronegativity	
1.2e	Polar Covalent Bonds	
1.2f	Bonding Continuum	
1.3a	Van der Waals' Forces	
1.3b	London Dispersion Forces	
1.3c	Permanent Dipole – Permanent Dipole Interactions	
1.3d	Hydrogen Bonding	
1.3e	Relating Properties to Intermolecular Forces	
1.3f	Predicting Solubility From Solute and Solvent Properties	

		Higher Chemistry Self-Evaluation Unit 1a Periodicity		Page	Traffic Light																		
					Red	Amber	Green																
1	In the Periodic Table, elements are arranged in order of increasing atomic number with elements with similar chemical properties in the same (vertical) group.			12	☹	☹	☺																
2a	The properties of an element can be predicted using its position in the Periodic Table <ul style="list-style-type: none">the Alkali Metals in group 1 are reactive elements where reactivity increases down the groupthe Halogens in Group 7 are reactive elements where reactivity decreases down the groupthe Noble Gases in Group 0 are unreactive elementsthe Transition Metals are in the middle section of the Periodic Table between Groups 2&3.			13 22	☹	☹	☺																
2b	Vertical columns on the period table are called groups . <ul style="list-style-type: none">elements in the same group have the same number of outer electrons<ul style="list-style-type: none">elements with the same number of outer electrons have similar chemical propertiesthe number of outer electrons is the same as the group number for groups 1-7			14 16 17	☹	☹	☺																
2c	Horizontal rows on the periodic table are called periods . Rows of elements are arranged in order of increasing atomic number (number of protons) <ul style="list-style-type: none">increasing atomic number leads to an increasing number of electrons in the outer shell going across a period from left to rightas the number of electrons in outer shell increases, the elements move from metallic characteristics to non-metal characteristics			15 18	☹	☹	☺																
3	The first 20 elements in the Periodic Table can be categorised according to bonding and structure: <table><tr><th>Metallic Bonding</th><th>Covalent Molecular</th><th>Covalent Network</th><th>Monotomic</th></tr><tr><td>Li, Be</td><td>H₂, N₂, O₂</td><td>C (diamond)</td><td>He</td></tr><tr><td>Na, Mg, Al</td><td>F₂, Cl₂</td><td>C (graphite)</td><td>Ne</td></tr><tr><td>K, Ca</td><td>P₄, S₈, C₆₀ (fullerene)</td><td>B, Si</td><td>Ar</td></tr></table>			Metallic Bonding	Covalent Molecular	Covalent Network	Monotomic	Li, Be	H ₂ , N ₂ , O ₂	C (diamond)	He	Na, Mg, Al	F ₂ , Cl ₂	C (graphite)	Ne	K, Ca	P ₄ , S ₈ , C ₆₀ (fullerene)	B, Si	Ar	26	☹	☹	☺
Metallic Bonding	Covalent Molecular	Covalent Network	Monotomic																				
Li, Be	H ₂ , N ₂ , O ₂	C (diamond)	He																				
Na, Mg, Al	F ₂ , Cl ₂	C (graphite)	Ne																				
K, Ca	P ₄ , S ₈ , C ₆₀ (fullerene)	B, Si	Ar																				
3<	Metals are at the right hand side of the Periodic Table while non-metals are at the left hand side <ul style="list-style-type: none">Metallic characteristics decrease as you go across a periodThe gaseous elements are at the top right hand side of the Periodic Table.<ul style="list-style-type: none">hydrogen is the exception and is at the top left hand side of the Periodic Tablenitrogen, oxygen, fluorine, chlorine, helium, neon, argon, krypton and xenon are at the top right hand side of the Periodic Table			19 20 21																			
4	Changes in covalent radius across a period and down a group can be explained in terms of changes in the number of occupied shells and the nuclear charge. <ul style="list-style-type: none">covalent radius/radii is a measure of the size of an atom.atoms increase in size <u>down a group</u><ul style="list-style-type: none">down a group there is an additional shell of electronsadditional shell of electrons ∴ atoms increase in atomic sizeatoms decrease in size <u>across a period</u><ul style="list-style-type: none">across a period the same shell of electrons is being filled upgreater positive charge in the nucleus pulls electron shells in towards nucleusouter shell pulled towards nucleus ∴ atoms decrease in atomic size			27 28	☹	☹	☺																
5	Ionisation is the process where electrons are removed from gaseous atoms: <ul style="list-style-type: none">the first ionisation energy is the energy required to remove one mole of electrons from one mole of atoms in the gaseous statethe second and subsequent ionisation energies refer to the energies required to remove further moles of electrons <table><tr><td>1st ionisation Energy of sodium Na(g) → Na⁺(g) + e⁻</td><td>2nd ionisation energy of carbon C⁺(g) → C²⁺(g) + e⁻</td><td>3rd ionisation energy of boron B²⁺(g) → B³⁺(g) + e⁻</td></tr></table> <ul style="list-style-type: none">Each ionisation energy is the removal of one mole of electrons.<ul style="list-style-type: none">Any equation which involves the removal of more than 1mol of electrons, ionisation energies must be added together to calculate total energy required			1 st ionisation Energy of sodium Na(g) → Na ⁺ (g) + e ⁻	2 nd ionisation energy of carbon C ⁺ (g) → C ²⁺ (g) + e ⁻	3 rd ionisation energy of boron B ²⁺ (g) → B ³⁺ (g) + e ⁻	29 30	☹	☹	☺													
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6	Changes in ionisation energies can be explained by atomic size and screening effect: <ul style="list-style-type: none">Changes to Ionisation Energy <u>Down a Group</u><ul style="list-style-type: none">atoms get bigger down a group as there is an addition shell of electrons each time.outer electrons are further from the nucleusouter electrons are also shielded from the full effect of the nucleus by the inner electrons shells			31 32	☹	☹	☺																




	<ul style="list-style-type: none"> ○ outer electron is easier to remove ∴ the ionisation energy decreases. • Changes to Ionisation Energy Across A Period <ul style="list-style-type: none"> ○ Same shell of electrons is filled up across a period ○ outer electrons are closer to the nucleus as they are more attracted to the increased positive charge of nucleus ○ outer electrons are more difficult to remove ∴ the ionisation energy increases. 				
7	<p>Atoms of different elements have different attractions for bonding electrons:</p> <ul style="list-style-type: none"> • electronegativity is the attraction an atom involved in a bond has for the electrons of the bond • the higher the electronegativity value, the stronger the attraction of the atom for the shared electrons in a covalent bond <ul style="list-style-type: none"> ○ electronegativity values increase across a period ○ electronegativity values decrease down a group 	33 34	☹	☺	☺
8	<p>Changes in electronegativity down a group and across a period can be explained using covalent radius, nuclear charge and screening effect due to inner shells of electrons.</p> <ul style="list-style-type: none"> • <u>Down a group</u> the atoms increase in size <ul style="list-style-type: none"> ○ bonding electrons in outer shell are further away from the nucleus ○ bonding electrons in outer shell are screened from the full effect of the nucleus by the inner shells of electrons ○ This causes the electronegativity value to decrease. • <u>Across a period</u> the atoms decrease in size <ul style="list-style-type: none"> ○ bonding electrons are closer to the positive nucleus and are more strongly attracted to the nucleus ○ this causes the electronegativity value to increase 	35 36	☹	☺	☺
9	<p>In metals the outer electrons are delocalised around positively charged ions</p> <ul style="list-style-type: none"> • delocalised electrons can jump from atom to atom to conduct electricity • positively charged ions consist of the nucleus and the inner electron shells 	24	☹	☺	☺
10	<p>Covalent bonding can exist in two different arrangements:</p> <ul style="list-style-type: none"> • Covalent molecular <ul style="list-style-type: none"> ○ discrete molecules with definite number of atoms held together by covalent bonds ○ strong covalent bonds inside molecules ○ weak Van der Waals' forces between molecules • Covalent network <ul style="list-style-type: none"> ○ network structures have a vast and indefinite number of atoms bonded together ○ strong covalent bonds hold the network together and has high melting point 	25	☹	☺	☺
			☹	☺	☺

			Higher Chemistry Self-Evaluation Unit 1b(i) Types of Chemical Bond		Page	Traffic Light																	
						Red	Amber	Green															
9			A covalent bond is a pair of shared electrons between two atoms <ul style="list-style-type: none">covalent bonds are usually formed between non-metal atomsthe atoms in a covalent bond are held together as the two positive nuclei and their common attraction for the shared pair of electrons		23 37	☹	☹	☺															
10a			Polar covalent bonds are formed when the attraction of the atoms for the pair of bonding electrons is different. Electronegativity values can distinguish between pure covalent & polar covalent bonds <ul style="list-style-type: none">the polarity of a covalent bond depends on the difference in electronegativity within the bond <table border="1"><tr><td><i>Small or No Difference</i> in Electronegativity ∴ Pure Covalent e.g. elements and hydrocarbons</td><td><i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H₂O, NH₃</td><td><i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na⁺Cl⁻, Mg²⁺(Cl⁻)₂</td></tr></table>	<i>Small or No Difference</i> in Electronegativity ∴ Pure Covalent e.g. elements and hydrocarbons	<i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H ₂ O, NH ₃	<i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na ⁺ Cl ⁻ , Mg ²⁺ (Cl ⁻) ₂		38	☹	☹	☺												
<i>Small or No Difference</i> in Electronegativity ∴ Pure Covalent e.g. elements and hydrocarbons	<i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H ₂ O, NH ₃	<i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na ⁺ Cl ⁻ , Mg ²⁺ (Cl ⁻) ₂																					
10b			Electronegativity values in the data book can be used to assign δ+ and δ- partial charges on atoms <ul style="list-style-type: none">the atom with the higher electronegativity value will have a greater share of the bonding electrons ∴ atom will have a δ- chargethe atom with the lower electronegativity value will have a lower share of the bonding electrons ∴ atom will have a δ+ charge.atoms in a polar covalent bond have a δ+ and δ- partial charges on atoms called a permanent dipole		39 40	☹	☹	☺															
11			Ionic formulae can be written giving the simplest ratio of each type of ion in the substance. <ul style="list-style-type: none">Ionic bonds are the electrostatic attraction between positive and negative ions.Ionic compounds form lattice structures of oppositely charged ions.		new	☹	☹	☺															
12a			The type of bonding of a substance is defined by electronegativity difference of the elements: <table border="1"><tr><td><i>Small or No Difference</i> in Electronegativity ∴ Pure covalent e.g. elements and hydrocarbons</td><td><i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H₂O, NH₃</td><td><i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na⁺Cl⁻, Mg²⁺(Cl⁻)₂</td></tr><tr><td>low ionic character high covalent character</td><td align="center">↔</td><td>high ionic character low covalent character</td></tr></table>	<i>Small or No Difference</i> in Electronegativity ∴ Pure covalent e.g. elements and hydrocarbons	<i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H ₂ O, NH ₃	<i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na ⁺ Cl ⁻ , Mg ²⁺ (Cl ⁻) ₂	low ionic character high covalent character	↔	high ionic character low covalent character		41	☹	☹	☺									
<i>Small or No Difference</i> in Electronegativity ∴ Pure covalent e.g. elements and hydrocarbons	<i>Medium Difference</i> in Electronegativity ∴ Polar Covalent e.g. H-Cl, H ₂ O, NH ₃	<i>Large Difference</i> in Electronegativity ∴ Ionic Bonding e.g. Na ⁺ Cl ⁻ , Mg ²⁺ (Cl ⁻) ₂																					
low ionic character high covalent character	↔	high ionic character low covalent character																					
12b			If the difference in electronegativity values is large then the electrons within the bond transfer from the element with lower electronegativity to the element with higher electronegativity resulting in the formation of ions		new	☹	☹	☺															
13			Compounds formed between metals and non-metals are often, but not always, ionic. Physical properties of a compound should be used to deduce the type of bonding and structure in the compound a) state at room temperature and melting & boiling point <ul style="list-style-type: none">All ionic compounds have high melting points ∴ All ionic compounds are solid at room tempCovalent Networks have high melting points ∴ All covalent networks are solid at room tempCovalent Molecular can have low boiling pts ∴ All gases/liquid compounds are covalent molecular b) Solubility <ul style="list-style-type: none">Ionic compounds and polar covalent substances are more likely to be soluble in water as they have charges inside the substanceNon-polar substances are less likely to be soluble in a polar solvent like water but will dissolve in non-polar solvents c) electrical conductivity, <table border="1"><tr><th>Type of Bonding</th><th>Conduction as a Solid</th><th>Conduction as a Liquid</th><th>Conduction as a Solution</th></tr><tr><td>Metallic (Metals only)</td><td align="center">✓</td><td align="center">✓</td><td>metals do <u>not</u>  dissolve in water</td></tr><tr><td>Covalent (Non-metals only)</td><td align="center">✗</td><td align="center">✗</td><td align="center">✗</td></tr><tr><td>Ionic (Metals + Non-metals)</td><td align="center">✗</td><td align="center">✓</td><td align="center">✓</td></tr></table>	Type of Bonding	Conduction as a Solid	Conduction as a Liquid	Conduction as a Solution	Metallic (Metals only)	✓	✓	metals do <u>not</u>  dissolve in water	Covalent (Non-metals only)	✗	✗	✗	Ionic (Metals + Non-metals)	✗	✓	✓	42 43	☹	☹	☺
Type of Bonding	Conduction as a Solid	Conduction as a Liquid	Conduction as a Solution																				
Metallic (Metals only)	✓	✓	metals do <u>not</u>  dissolve in water																				
Covalent (Non-metals only)	✗	✗	✗																				
Ionic (Metals + Non-metals)	✗	✓	✓																				

			Higher Chemistry Self-Evaluation Unit 1b(ii) Intermolecular Forces		Page	Traffic Light		
						Red	Amber	Green
14		All substances will freeze at sufficiently low temperatures. For the physical change of freezing from a liquid to a gas there must be at least one form of attractive forces between the molecules or discrete atoms. <div><div><u>Molecular Elements</u> e.g. H₂, N₂, O₂</div><div><u>Molecular Compounds</u> e.g. HCl, H₂O, NH₃</div><div><u>Monatomic Elements</u> e.g. He, Ne, Ar, Kr, Xe</div></div>						
15		There are more than one type of Van der Waal's intermolecular forces found between molecules: <ul style="list-style-type: none">There are temporary and permanent types of Van der Waals' forces <div><div>Van der Waal's forces</div><div><div>Temporary Type</div><div>Permanent Type</div></div><div><div>London dispersion forces</div><div>Permanent dipole to permanent dipole attractions</div><div>Hydrogen bonding</div></div></div>			45			
16		London Dispersion Forces are caused by a temporary uneven distribution of electrons within atoms and molecules <div><div></div><div>Atom temporarily has more electrons on one side</div><div>Temporary dipole</div><div>Temporary dipole induces neighbouring atoms to form dipole</div><div><div></div><div>induced dipole temporary dipole induced dipole</div></div></div>			47			
17		The strength of London dispersion forces is decided by the number of electrons within the atom or molecule <ul style="list-style-type: none">the greater the number of electrons the stronger the London dispersion forces.			49			

18 20	<p>A molecule is described as polar if it contains a permanent dipole</p> <ul style="list-style-type: none">Permanent dipoles are formed in a molecule where electronegativity difference is large enough within the bondElectrons within the bond spend more time at the more electronegative end of the bond <div style="display: flex; align-items: center; justify-content: space-around;"><div style="text-align: center;">Electronegativity of hydrogen = 2.2</div><div style="text-align: center;">$\overset{1+}{\text{H}} \text{---} \overset{17+}{\text{Cl}}$ $\text{e} \quad \text{e}$</div><div style="text-align: center;">Electronegativity of chlorine = 3.0</div></div> <p style="text-align: center;">↓ permanent dipole formed over the covalent bond.</p> <div style="text-align: center;">$\delta^+ \text{H} \text{---} \text{Cl} \delta^-$</div> <ul style="list-style-type: none">The permanent dipoles formed in each molecule are attracted to each otherpermanent dipole-permanent dipole interactions are additional electrostatic forces of attraction between polar molecules (on top of London dispersion forces already present) <div style="display: flex; align-items: center; justify-content: space-around;"><div style="text-align: center;">$\delta^+ \text{H} \text{---} \text{Cl} \delta^-$ ↑ intramolecular force covalent bond</div><div style="text-align: center;">.....</div><div style="text-align: center;">$\delta^+ \text{H} \text{---} \text{Cl} \delta^-$ ↑ intermolecular force permanent dipole to permanent dipole attraction</div><div style="text-align: center;">.....</div><div style="text-align: center;">$\delta^+ \text{H} \text{---} \text{Cl} \delta^-$</div></div>	51	☹	☺	☺						
19	<p>The spatial arrangement of polar covalent bonds is important in deciding whether a molecule is polar or non-polar.</p> <table border="1" style="width: 100%; border-collapse: collapse;"><tr><td style="width: 15%; text-align: center; vertical-align: middle;">Polar Molecule</td><td style="width: 30%; text-align: center; vertical-align: middle;">$\delta^+ \text{:C} \equiv \text{O} \delta^-$ Carbon monoxide CO</td><td style="width: 55%; text-align: center; vertical-align: middle;"> Trichloromethane CHCl3</td></tr><tr><td style="text-align: center; vertical-align: middle;">Non-polar Molecule</td><td style="text-align: center; vertical-align: middle;">$\delta^- \text{O} = \text{C} = \text{O} \delta^-$ Carbon Dioxide CO2 Non-polar as polarity cancels out due to linear shape</td><td style="text-align: center; vertical-align: middle;"> Tetrachloromethane CCl4 Non-polar as molecule has 4 delta- ends and polarity cancels out</td></tr></table>	Polar Molecule	$\delta^+ \text{:C} \equiv \text{O} \delta^-$ Carbon monoxide CO	 Trichloromethane CHCl3	Non-polar Molecule	$\delta^- \text{O} = \text{C} = \text{O} \delta^-$ Carbon Dioxide CO2 Non-polar as polarity cancels out due to linear shape	 Tetrachloromethane CCl4 Non-polar as molecule has 4 delta- ends and polarity cancels out	52	☹	☺	☺
Polar Molecule	$\delta^+ \text{:C} \equiv \text{O} \delta^-$ Carbon monoxide CO	 Trichloromethane CHCl3									
Non-polar Molecule	$\delta^- \text{O} = \text{C} = \text{O} \delta^-$ Carbon Dioxide CO2 Non-polar as polarity cancels out due to linear shape	 Tetrachloromethane CCl4 Non-polar as molecule has 4 delta- ends and polarity cancels out									
21 22c	<p>Permanent dipole-permanent dipole interactions are stronger than London dispersion forces: Hydrogen bonding is weak compared to covalent bonding but is stronger than permanent dipole to permanent dipole attractions:</p>	53 56	☹	☺	☺						

	Ionic and Covalent Bonding	Hydrogen Bonding	Permanent Dipole to Permanent Dipole Forces	London Dispersion Forces																								
22a 22b	<p>Bonds consisting of a hydrogen atom bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen are highly polar.</p> <ul style="list-style-type: none">Hydrogen bonds are electrostatic forces of attraction between molecules that contain highly polar -OH, -NH or H-F groupsMolecules with N, O or F directly bonded to H will have hydrogen bonding between molecules. <div><div>H—F</div><div>—O—H</div><div>—N—H</div></div> <div>Found in water, alcohols and carboxylic acids</div> <div>Found in ammonia and amines</div>				55	☹	☹	☺																				
23 24	<p>The type and strength of intermolecular bonds affects melting point, boiling point and viscosity:</p> <ul style="list-style-type: none">The higher the melting & boiling point of a substance, the stronger the intermolecular forcesCovalent compounds with hydrogen bonding between molecules have higher melting points, boiling points and viscosity than those with only London dispersion forcesPolar molecules will have stronger intermolecular forces than non-polar molecules when comparing molecules with similar numbers of electronsthe greater the number of electrons the stronger the London dispersion forces between molecules. <table><tr><th>Type of Van der Waals' Forces in Substance</th><th>Hydrogen Bonding</th><th>Permanent Dipole to Permanent Dipole Bonding</th><th>London Dispersion Forces</th></tr><tr><td>Strength of bonding</td><td>Strongest</td><td>←————→</td><td>Weakest</td></tr><tr><td>Melting Point</td><td>Higher</td><td>←————→</td><td>Lower</td></tr><tr><td>Boiling Point</td><td>Higher</td><td>←————→</td><td>Lower</td></tr><tr><td>Viscosity</td><td>Thicker</td><td>←————→</td><td>Thinner</td></tr></table>				Type of Van der Waals' Forces in Substance	Hydrogen Bonding	Permanent Dipole to Permanent Dipole Bonding	London Dispersion Forces	Strength of bonding	Strongest	←————→	Weakest	Melting Point	Higher	←————→	Lower	Boiling Point	Higher	←————→	Lower	Viscosity	Thicker	←————→	Thinner	57 58 59	☹	☹	☺
Type of Van der Waals' Forces in Substance	Hydrogen Bonding	Permanent Dipole to Permanent Dipole Bonding	London Dispersion Forces																									
Strength of bonding	Strongest	←————→	Weakest																									
Melting Point	Higher	←————→	Lower																									
Boiling Point	Higher	←————→	Lower																									
Viscosity	Thicker	←————→	Thinner																									
25	<p>Hydrogen bonding has an effect on boiling point, melting point, viscosity and solubility (miscibility)</p> <ul style="list-style-type: none">Hydrogen bonding between the molecules and water molecules makes the substance soluble or miscible with water.				61	☹	☹	☺																				
26	<p>The boiling points of ammonia, water and hydrogen fluoride are higher than expected given the number of electrons present in the molecules</p> <ul style="list-style-type: none">NH₃, H₂O and HF have hydrogen bonding between their molecules				60	☹	☹	☺																				
27	<p>Hydrogen bonding causes ice to be less dense than water at low temperatures</p> <ul style="list-style-type: none">Arrangement of H₂O molecules in solid ice takes up more space than in liquid waterSame mass but increased volume of H₂O causes density to be lower (density = ^{mass}/_{volume})Water is unlike most substances as its solid floats in its liquid.				62	☹	☹	☺																				
28	<p>The type of solvent (polar or non-polar) decides the type of substance which will dissolve in it:</p> <table><tr><th>Type of Solvent</th><th colspan="2">Type of Substance which dissolves in Solvent</th></tr><tr><td>Polar</td><td>Polar covalent compounds</td><td>Ionic compounds</td></tr><tr><td>Non-polar</td><td>Non-polar covalent substances</td><td></td></tr></table> <p>Polar substances are insoluble in non-polar solvents</p> <p>Non-polar substances are insoluble in polar solvents e.g. water</p>				Type of Solvent	Type of Substance which dissolves in Solvent		Polar	Polar covalent compounds	Ionic compounds	Non-polar	Non-polar covalent substances		63 64	☹	☹	☺											
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29	<p>Solubility of a compound can be predicted using the following key structural features:</p> <ul style="list-style-type: none">molecules with O-H and N-H groups have hydrogen bonding between moleculesthe O^{δ-} or N^{δ-} in one molecule attracts the H^{δ+} in a neighbouring moleculesspatial arrangement of polar covalent bonds resulting in a permanent dipole				65 66	☹	☹	☺																				

		Higher Chemistry Self-Evaluation Unit 1c Oxidising and Reducing Agents		Page	Traffic Light																						
					Red	Amber	Green																				
30	Reduction is a gain of electrons by a reactant in any reaction.																										
31	Oxidation is a loss of electrons by a reactant in any reaction.				☹	☹	☺																				
32	In a redox reaction, reduction and oxidation take place at the same time.																										
33	An oxidising agent is a substance which accepts electrons <ul style="list-style-type: none">oxidising agent oxidises something elseagent itself is reduced and accepts/gains electrons			36a 37a	☹	☹	☺																				
35a	<ul style="list-style-type: none">oxidising agents tend to become more negative e.g. acidified permanganate solution is an example of an oxidising agent which gains electrons $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \longrightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$																										
34	A reducing agent is a substance which donates electrons <ul style="list-style-type: none">reducing agent reduces something elseagent itself is oxidised and loses electrons			36b 37b	☹	☹	☺																				
35b	<ul style="list-style-type: none">reducing agents tend to become more positive e.g. sulphite ions are an example of a reducing agent which loses electrons $\text{SO}_3^{2-} + \text{H}_2\text{O} \longrightarrow \text{SO}_4^{2-} + 2\text{H}^+ + 2\text{e}^-$																										
36	Electronegativity can be used to predict which elements lose or gain electrons when they form ions: <ul style="list-style-type: none">metals have low electronegativity values and tend to lose electrons to become positive ions<ul style="list-style-type: none">metals act as reducing agents as they lose electrons			38 39	☹	☹	☺																				
37	<table border="1"><thead><tr><th>Element</th><th>Metal/Non-metal</th><th>Electronegativity Value</th><th>Equation</th></tr></thead><tbody><tr><td>Potassium</td><td>Metal</td><td>0.8</td><td>$\text{K} \rightarrow \text{K}^+ + \text{e}^-$</td></tr><tr><td>Lithium</td><td>Metal</td><td>1.0</td><td>$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$</td></tr></tbody></table> <ul style="list-style-type: none">non-metals have high electronegativities and tend to gain electrons to become negative ions<ul style="list-style-type: none">non-metals act as oxidising agents as they gain electrons <table border="1"><thead><tr><th>Element</th><th>Metal/Non-metal</th><th>Electronegativity Value</th><th>Equation</th></tr></thead><tbody><tr><td>Chlorine</td><td>Non-metal</td><td>3.0</td><td>$\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$</td></tr><tr><td>Fluorine</td><td>Non-metal</td><td>4.0</td><td>$\text{F} + \text{e}^- \rightarrow \text{F}^-$</td></tr></tbody></table>							Element	Metal/Non-metal	Electronegativity Value	Equation	Potassium	Metal	0.8	$\text{K} \rightarrow \text{K}^+ + \text{e}^-$	Lithium	Metal	1.0	$\text{Li} \rightarrow \text{Li}^+ + \text{e}^-$	Element	Metal/Non-metal	Electronegativity Value	Equation	Chlorine	Non-metal	3.0	$\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$
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38	Group 1 elements are the strongest reducing agents <ul style="list-style-type: none">The strongest oxidising agents are at the bottom of the left-hand column of the electrochemical series.Alkali metals: Lithium, sodium, potassium, rubidium, caesium and francium			40	☹	☹	☺																				
42	Group 7 elements are the strongest oxidising agents <ul style="list-style-type: none">The strongest reducing agents are at the top of the right-hand column of the electrochemical series.Halogens: fluorine, chlorine, bromine, iodine																										
43																											
39a	Hydrogen Peroxide reacts by the following equation: $\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$ <ul style="list-style-type: none">Acidified peroxide is an oxidising agent as it accepts electrons and is reduced itself			44	☹	☹	☺																				
39b	Acidified dichromate solution reacts by the following equation: $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$ <ul style="list-style-type: none">Acidified dichromate is an oxidising agent as it accepts electrons and is reduced itself Acidified permanganate solution reacts by the following equation: $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ <ul style="list-style-type: none">Acidified dichromate is an oxidising agent as it accepts electrons and is reduced itself			43	☹	☹	☺																				
39c	Carbon Monoxide reacts by the following equation: $\text{CO} + \text{H}_2\text{O} \rightarrow \text{CO}_2 + 2\text{H}^+ + 2\text{e}^-$ <ul style="list-style-type: none">Carbon monoxide is a reducing agent as it loses electrons and is oxidised itself			44	☹	☹	☺																				
40	Oxidising agents can be used as a chemical to <ul style="list-style-type: none">Bleach clothes or hair as oxidising agents breaks down coloured compoundsKill fungi and bacteria and inactive viruses			45	☹	☹	☺																				
41	The electrochemical series represents a series of reduction reactions <ul style="list-style-type: none">if the reduction reaction in electrochemical series is reversed it becomes an oxidation reaction				☹	☹	☺																				

44	<p>Given reactant and product species, ion-electron equations which include $\text{H}^+(\text{aq})$ and $\text{H}_2\text{O}(\text{l})$ can be written.</p> <p>1. Write down the main species involved in the reaction</p> $\text{IO}_3^- \rightarrow \text{I}_2$ <p>2. Balance all atoms except O and H</p> $2\text{IO}_3^- \rightarrow \text{I}_2$ <p>3. Add H_2O to other side to balance O atoms</p> $2\text{IO}_3^- \rightarrow \text{I}_2 + 6\text{H}_2\text{O}$ <p>4. Add H^+ ions to other side to balance H atoms</p> $2\text{IO}_3^- + 12\text{H}^+ \rightarrow \text{I}_2 + 6\text{H}_2\text{O}$ <p>5. Add e^- to most positive side to balance charge</p> $2\text{IO}_3^- + 12\text{H}^+ + 10\text{e}^- \rightarrow \text{I}_2 + 6\text{H}_2\text{O}$	46			
45	<p>Ion-electron equations can be combined to produce redox equations</p> <p>Reduction: $\text{I}_2 + 2\text{e}^- \rightarrow 2\text{I}^-$</p> <p>Oxidation: $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^-$</p> <p>Redox: $\text{I}_2 + 2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{I}^-$</p> <p>Where the electrons do not cancel out, ion-electron equations may have to be multiplied:</p> <p>① $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$</p> <p>② $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$</p> <p>① $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$</p> <p>②x5 $5\text{Fe}^{2+} \rightarrow 5\text{Fe}^{3+} + 5\text{e}^-$</p> <p>add and cancel down</p> $\text{MnO}_4^- + 8\text{H}^+ + \cancel{5\text{e}^-} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ $5\text{Fe}^{2+} \rightarrow 5\text{Fe}^{3+} + \cancel{5\text{e}^-}$ <p>redox $\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$</p>	47			