JABchem

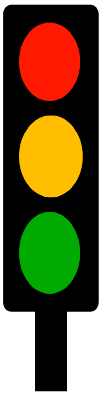
Name: ……………………………………………………………………………………

Class: ………………

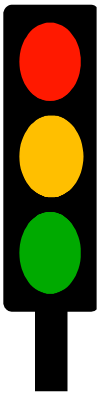
Teacher: ………………………………

Self-Evaluation

HIGHER CHEMISTRY



Unit 1



Chemical Changes & Structure

Section	Title	Completed
1.1	Periodicity	
1.2a	Types of Chemical Bond	
1.2b	Intermolecular Forces	
1.3	Oxidising and Reducing Agents	

Higher Chemistry Self-Evaluation
Unit 1a Periodicity

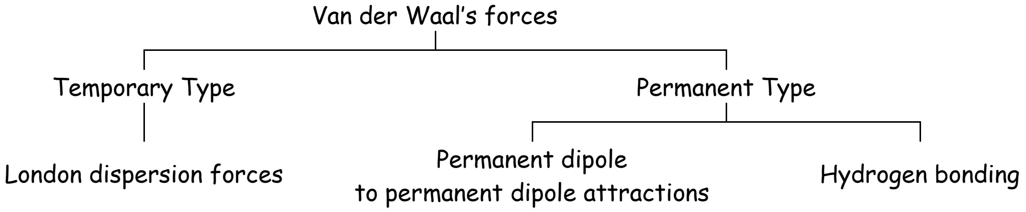
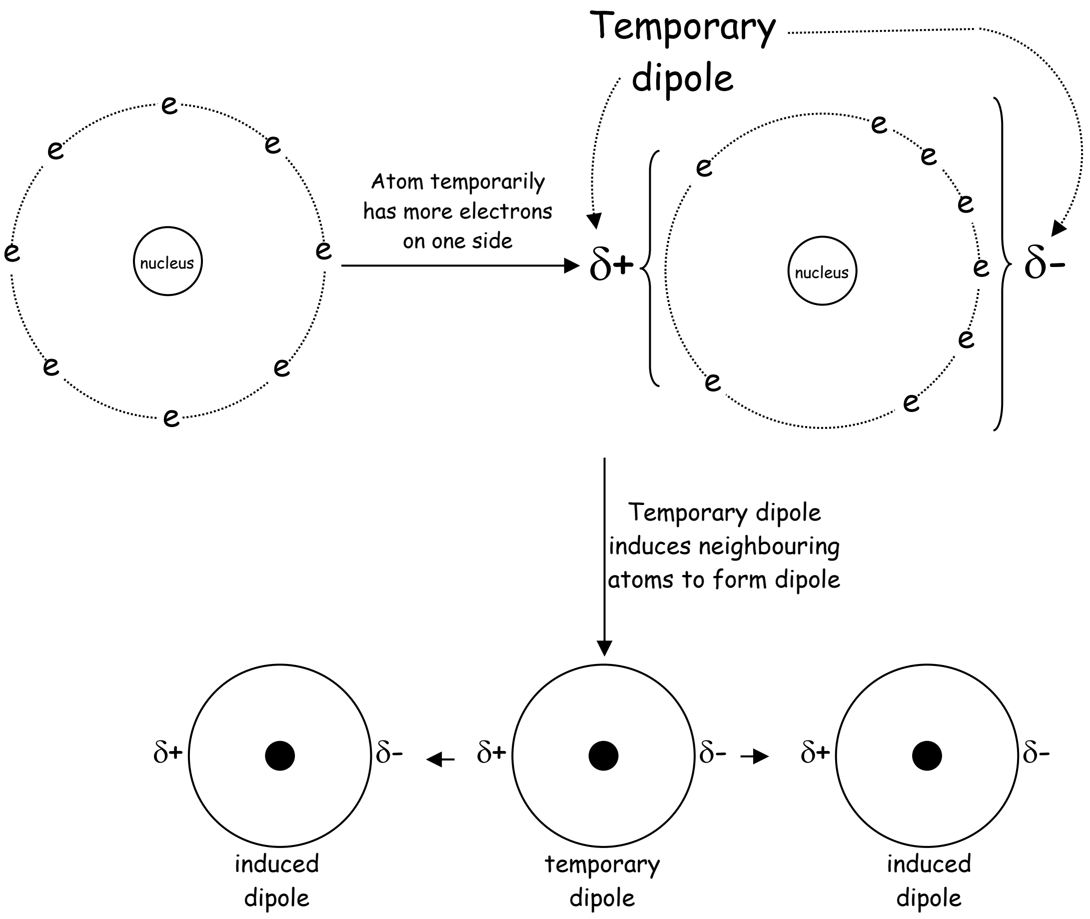
<table>
<tr><th rowspan="2"></th><th rowspan="2"></th><th rowspan="2">Page</th><th colspan="3">Traffic Light</th></tr>
<tr><th>Red</th><th>Amber</th><th>Green</th></tr>
<tr><td>1</td><td>In the Periodic Table, elements are arranged in order of increasing atomic number with elements with similar chemical properties in the same (vertical) group.</td><td>12</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>2a</td><td>The chemical and physical properties of an element can be predicted from its position in Periodic Table
• the *Alkali Metals* in group 1 are reactive elements where reactivity increases down the group
• the *Halogens* in Group 7 are reactive elements where reactivity decreases down the group
• the *Noble Gases* in Group 0 are unreactive elements
• the *Transition Metals* are in the middle section of the Periodic Table between Groups 2&3.</td><td>13
22</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>2b</td><td>Vertical columns on the Period Table are called **groups**.
• elements in the same group have the same number of outer electrons
○ elements with the same number of outer electrons have similar chemical properties
○ the number of outer electrons is the same as the group number for groups 1-7</td><td>14
16
17</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>2c</td><td>Horizontal rows on the periodic table are called **periods**.
Rows of elements are arranged in order of increasing atomic number (number of protons)
• increasing atomic number leads to an increasing number of electrons in the outer shell going across a period from left to right
• as the number of electrons in outer shell increases, the elements move from metallic characteristics to non-metal characteristics</td><td>15
18</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>3</td><td>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>H_2, N_2, O_2</td><td>C (diamond)</td><td>He</td></tr>
<tr><td>Na, Mg, Al</td><td>F_2, Cl_2</td><td>C (graphite)</td><td>Ne</td></tr>
<tr><td>K, Ca</td><td>P_4, S_8, C_{60} (fullerene)</td><td>B, Si</td><td>Ar</td></tr>
</table></td><td>26</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>4</td><td>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.
• covalent radius/radii is a measure of the size of an atom.
• atoms increase in size down a group
○ down a group there is an additional shell of electrons
○ additional shell of electrons ∴ atoms increase in atomic size
• atoms decrease in size across a period
○ across a period the same shell of electrons is being filled up
○ greater positive charge in the nucleus pulls electron shells in towards nucleus
○ outer shell pulled towards nucleus ∴atoms decrease in atomic size</td><td>27
28</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>5</td><td>Ionisation is the process where electrons are removed from gaseous atoms:
• the first ionisation energy is the energy required to remove one mole of electrons from one mole of atoms in the gaseous state
• the second and subsequent ionisation energies refer to the energies required to remove further moles of electrons
<table>
<tr><th>1st ionisation Energy of sodium</th><th>2nd ionisation energy of carbon</th><th>3rd ionisation energy of boron</th></tr>
<tr><td>$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-}$</td><td>$C^{+}_{(g)} \rightarrow C^{2+}_{(g)} + e^{-}$</td><td>$B^{2+}_{(g)} \rightarrow B^{3+}_{(g)} + e^{-}$</td></tr>
</table>
• Each ionisation energy is the removal of one mole of electrons.
○ Any equation which involves the removal of more than 1mol of electrons requires ionisation energies to be added together to calculate total energy required</td><td>29
30</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>6</td><td>Changes in ionisation energies can be explained by atomic size and screening effect:
• Changes to Ionisation Energy Down a Group
○ atoms get bigger down a group as there is an addition shell of electrons each time.
○ outer electrons are further from the nucleus
○ outer electrons are also shielded from the full effect of the nucleus by the inner electron shells
○ outer electron is easier to remove ∴ ionisation energy decreases.
• Changes to Ionisation Energy Across A Period
○ 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.</td><td>31
32</td><td>☹</td><td>😐</td><td>☺</td></tr>
</table>

7	Atoms of different elements have different attractions for bonding electrons:
	• 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
	  ○ electronegativity values increase across a period
	  ○ electronegativity values decrease down a group
		33
		34
			☹	😐	☺

8	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.
	• Down a group the atoms increase in size
	  ○ 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 decreases.
	• Across a period the atoms decrease in size
	  ○ 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
			☹	😐	☺

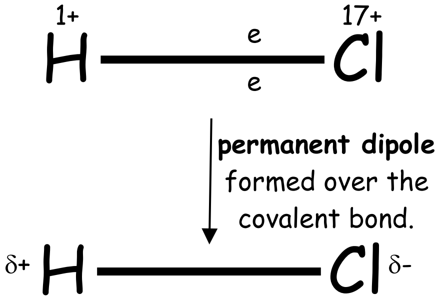
Higher Chemistry Self-Evaluation
Unit 1.2a Types of Chemical Bond

<table>
<tr><th></th><th></th><th>Page</th><th colspan="3">Traffic Light</th></tr>
<tr><th></th><th></th><th></th><th>Red</th><th>Amber</th><th>Green</th></tr>
<tr><td>9</td><td>A covalent bond is a pair of shared electrons between two atoms
covalent bonds are usually formed between non-metal atoms
the atoms in a covalent bond are held together as the two positive nuclei and their common attraction for the shared pair of electrons</td><td>23
37</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>10a</td><td>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
the polarity of a covalent bond depends on the difference in electronegativity within the bond
<table>
<tr><td><i>Small</i> or <i>No</i> Difference in Electronegativity
∴ Pure Covalent
e.g. elements and hydrocarbons</td><td><i>Medium</i> Difference in Electronegativity
∴ Polar Covalent
e.g. H-Cl, H_2O, NH_3</td><td><i>Large</i> Difference in Electronegativity
∴ Ionic Bonding
e.g. Na^+Cl^-, $Mg^{2+}(Cl^-)_2$</td></tr>
</table></td><td>38</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>10b</td><td>Electronegativity values in the data book can be used to assign δ+ and δ- partial charges on atoms
the atom with the higher electronegativity value will have a greater share of the bonding electrons ∴ this atom will have a δ- charge
the atom with the lower electronegativity value will have a smaller share of the bonding electrons ∴ this 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</td><td>39
40</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>11</td><td>Ionic formulae can be written giving the simplest ratio of each type of ion in the substance.
Ionic bonds are the electrostatic attraction between positive and negative ions.
Ionic compounds form lattice structures of oppositely charged ions.</td><td>new</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>12a</td><td>The type of bonding of a substance is defined by electronegativity difference of the elements:
<table>
<tr><td><i>Small</i> or <i>No</i> Difference in Electronegativity
∴ Pure covalent
e.g. elements and hydrocarbons</td><td><i>Medium</i> Difference in Electronegativity
∴ Polar Covalent
e.g. H-Cl, H_2O, NH_3</td><td><i>Large</i> Difference in Electronegativity
∴ Ionic Bonding
e.g. Na^+Cl^-, $Mg^{2+}(Cl^-)_2$</td></tr>
<tr><td>low ionic character
high covalent character</td><td>⟷</td><td>high ionic character
low covalent character</td></tr>
</table></td><td>41</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>12b</td><td>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</td><td>new</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>13</td><td>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
All ionic compounds have high melting points ∴ All ionic compounds are solid at room temp
Covalent Networks have high melting points ∴ All covalent networks are solid at room temp
Covalent Molecular can have low boiling pts ∴ All gases/liquid compounds are covalent molecular
b) Solubility
Ionic compounds and polar covalent substances are more likely to be soluble in water as they have charges inside the substance
Non-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>
<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>✓</td><td>✓</td><td>metals do <u>not</u>
dissolve in water</td></tr>
<tr><td>Covalent
(Non-metals only)</td><td>×</td><td>×</td><td>×</td></tr>
<tr><td>Ionic
(Metals + Non-metals)</td><td>×</td><td>✓</td><td>✓</td></tr>
</table></td><td>42
43</td><td>☹</td><td>😐</td><td>☺</td></tr>
</table>

	Higher Chemistry Self-Evaluation Unit 1.2b 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. Molecular Elements e.g. H_2, N_2, O_2 Molecular Compounds e.g. HCl, H_2O, NH_3 Monotomic Elements e.g. He, Ne, Ar, Kr, Xe		☹	😐	☺
15	There are more than one type of Van der Waal's intermolecular forces found between molecules: - There are temporary and permanent types of Van der Waals' forces  Van der Waal's forces Temporary Type London dispersion forces Permanent Type Permanent dipole to permanent dipole attractions Hydrogen bonding	45	☹	😐	☺
16	London Dispersion Forces are forces of attraction which happen between all atoms and molecules. - these forces are much weaker than all other types of bonding. - they are caused by a temporary uneven distribution of electrons within atoms and molecules 	47	☹	😐	☺
17	The strength of London dispersion forces is decided by the number of electrons within the atom or molecule - the greater the number of electrons the stronger the London dispersion forces. - Larger molecules have more electrons than smaller molecules so larger molecules have stronger London dispersion forces between them.	49	☹	😐	☺

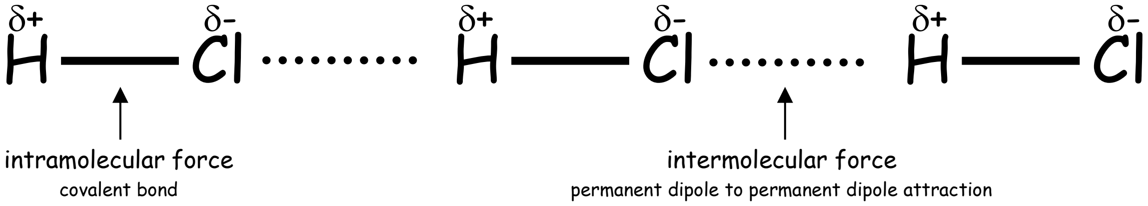
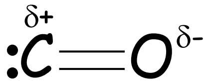
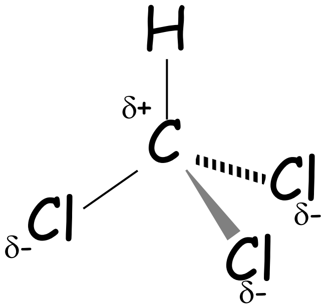
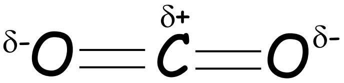
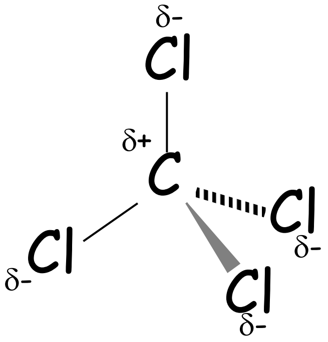
<table>
<tr><td>18
20</td><td>A molecule is described as polar if it contains a permanent dipole

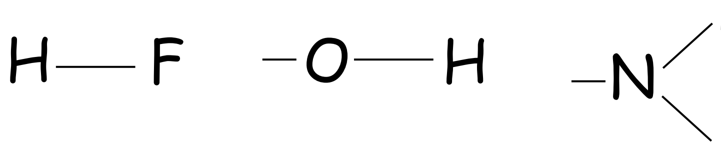
- Permanent dipoles are formed in a molecule where the electronegativity difference is large enough and electrons within the bond are shared unequally.
- Electrons within the bond spend more time at the more electronegative end of the bond

Electronegativity of hydrogen = 2.2


Electronegativity of chlorine = 3.0

- The permanent dipoles formed in each molecule are attracted to each other
- permanent dipole-permanent dipole interactions are additional electrostatic forces of attraction between polar molecules (on top of London dispersion forces already present)


</td><td>51</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>19</td><td>The spatial arrangement of polar covalent bonds is important in deciding whether a molecule is polar or non-polar.
<table>
<tr><td>Polar Molecule</td><td>
Carbon monoxide CO</td><td>
Trichloromethane $CHCl_3$</td></tr>
<tr><td>Non-polar Molecule</td><td>
Carbon Dioxide CO_2
Non-polar as polarity cancels out due to linear shape</td><td>
Tetrachloromethane CCl_4
Non-polar as molecule has 4 δ- ends and polarity cancels out</td></tr>
</table></td><td>52</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>21
22c</td><td>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:
Ionic and Covalent Bonding > Hydrogen Bonding > Permanent Dipole to Permanent Dipole Forces > London Dispersion Forces</td><td>53
56</td><td>☹</td><td>😐</td><td>☺</td></tr>
</table>

<table>
<tr><td>22a
22b</td><td>Bonds consisting of a hydrogen atom directly bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen are highly polar.
- Hydrogen bonds are electrostatic forces of attraction between molecules that contain highly polar –OH, -NH or H-F groups
- Molecules with N, O or F directly bonded to H will have hydrogen bonding between molecules.

Found in water, alcohols and carboxylic acids
Found in ammonia and amines</td><td>55</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>23
24</td><td>The type and strength of intermolecular bonds affects melting point, boiling point and viscosity:
- The higher the melting & boiling point of a substance, the stronger the intermolecular forces
- Covalent compounds with hydrogen bonding between molecules have higher melting points, boiling points and viscosity than those with only London dispersion forces
- Polar molecules will have stronger intermolecular forces than non-polar molecules when comparing molecules with similar numbers of electrons
- the greater the number of electrons the stronger the London dispersion forces between molecules.
<table>
<tr><td>Type of Van der Waals' Forces in Substance</td><td>Hydrogen Bonding</td><td>Permanent Dipole to Permanent Dipole Bonding</td><td>London Dispersion Forces</td></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></td><td>57
58
59</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>25</td><td>Hydrogen bonding has an effect on boiling point, melting point, viscosity and solubility (miscibility)
- Hydrogen bonding between the molecules and water molecules makes the substance soluble or miscible with water. Both molecules are polar.
- Hydrogen bonding increases the viscosity (thickness) of a substance. Hydrogen bonding brings molecules closer together and makes the substance thicker.
- Hydrogen bonding elevates the melting/boiling point of a substance due to brining the molecules closer together.</td><td>61</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>26</td><td>The boiling points of ammonia, water and hydrogen fluoride are higher than expected given the number of electrons present in the molecules
- NH_3, H_2O and HF have hydrogen bonding between their molecules which raises boiling point
- These higher boiling points are described as anomalous as they are different from the norm.</td><td>60</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>27</td><td>Hydrogen bonding causes ice to be less dense than water at low temperatures
- Expanded structure of ice due to hydrogen bonding between H_2O molecules caused by molecules being spaced further apart than in liquid water
- Arrangement of H_2O molecules in solid ice takes up more space than in liquid water
- Same mass but increased volume of H_2O causes density to be lower (density = $^{mass}/_{volume}$)
- Water is unlike most substances as its solid floats in its liquid.</td><td>62</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>28</td><td>The type of solvent (polar or non-polar) decides the type of substance which will dissolve in it:
<table>
<tr><td>Type of Solvent</td><td colspan="2">Type of Substance which dissolves in Solvent</td></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>
Polar substances are insoluble in non-polar solvents
Non-polar substances are insoluble in polar solvents e.g. water</td><td>63
64</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>29</td><td>Solubility of a compound can be predicted using the following key structural features:
- molecules with O-H and N-H groups have hydrogen bonding between molecules
- the $O^{\delta-}$ or $N^{\delta-}$ in one molecule attracts the $H^{\delta+}$ in a neighbouring molecules
- spatial arrangement of polar covalent bonds resulting in a permanent dipole</td><td>65
66</td><td>☹</td><td>😐</td><td>☺</td></tr>
</table>

Higher Chemistry Self-Evaluation
Unit 1.3 Oxidising and Reducing Agents

<table>
<tr><th rowspan="2"></th><th rowspan="2"></th><th rowspan="2">Page</th><th colspan="3">Traffic Light</th></tr>
<tr><th>Red</th><th>Amber</th><th>Green</th></tr>
<tr><td>30
31
32</td><td>Reduction is a gain of electrons by a reactant in any reaction.
Oxidation is a loss of electrons by a reactant in any reaction.
In a redox reaction, reduction and oxidation take place at the same time.</td><td></td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>33
35a</td><td>An oxidising agent is a substance which accepts electrons
oxidising agent oxidises something else
oxidising agent itself is reduced and accepts/gains electrons
oxidising agents tend to become more negative
e.g. acidified permanganate solution is an example of an oxidising agent which gains electrons
$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$</td><td>36a
37a</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>34
35b</td><td>A reducing agent is a substance which donates electrons
reducing agent reduces something else
agent itself is oxidised and loses electrons
reducing agents tend to become more positive
e.g. sulphite ions are an example of a reducing agent which loses electrons
$SO_3^{2-} + H_2O \longrightarrow SO_4^{2-} + 2H^+ + 2e^-$</td><td>36b
37b</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>36
37</td><td>Electronegativity can be used to predict which elements lose or gain electrons when they form ions:
metals have low electronegativity values and tend to lose electrons to become positive ions
metals act as reducing agents as the lose electrons
<table>
<tr><th>Element</th><th>Metal/Non-metal</th><th>Electronegativity Value</th><th>Equation</th></tr>
<tr><td>Potassium</td><td>Metal</td><td>0.8</td><td>$K \rightarrow K^+ + e^-$</td></tr>
<tr><td>Lithium</td><td>Metal</td><td>1.0</td><td>$Li \rightarrow Li^+ + e^-$</td></tr>
</table>
non-metals have high electronegativities and tend to gain electrons to become negative ions
non-metals act as oxidising agents as they gain electrons
<table>
<tr><th>Element</th><th>Metal/Non-metal</th><th>Electronegativity Value</th><th>Equation</th></tr>
<tr><td>Chlorine</td><td>Non-metal</td><td>3.0</td><td>$Cl + e^- \rightarrow Cl^-$</td></tr>
<tr><td>Fluorine</td><td>Non-metal</td><td>4.0</td><td>$F + e^- \rightarrow F^-$</td></tr>
</table></td><td>38
39</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>38
42
43</td><td>Group 1 elements are the strongest reducing agents
The strongest oxidising agents are at the bottom of the left-hand column of the electrochemical series.
Halogens: fluorine, chlorine, bromine, iodine
Group 7 elements are the strongest oxidising agents
The strongest reducing agents are at the top of the right-hand column of the electrochemical series.
Alkali metals: lithium, sodium, potassium, rubidium, caesium and francium</td><td>40</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>39a</td><td>Hydrogen Peroxide reacts by the following equation:
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$
Acidified peroxide is an oxidising agent as it accepts electrons and is reduced itself</td><td>44</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>39b</td><td>Acidified dichromate solution reacts by the following equation:
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
Acidified dichromate is an oxidising agent as it accepts electrons and is reduced itself
Acidified permanganate solution reacts by the following equation:
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
Acidified dichromate is an oxidising agent as it accepts electrons and is reduced itself</td><td>43</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>39c</td><td>Carbon Monoxide reacts by the following equation:
$CO + H_2O \rightarrow CO_2 + 2H^+ + 2e^-$
Carbon monoxide is a reducing agent as it loses electrons and is oxidised itself</td><td>44</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>40</td><td>Oxidising agents can be used as a chemical to
Bleach clothes or hair as oxidising agents breaks down coloured compounds
Kill fungi and bacteria and inactive viruses</td><td>45</td><td>☹</td><td>😐</td><td>☺</td></tr>
<tr><td>41</td><td>The electrochemical series represents a series of reduction reactions
if the reduction reaction in electrochemical series is reversed it becomes an oxidation reaction</td><td></td><td>☹</td><td>😐</td><td>☺</td></tr>
</table>

44	Given reactant and product species, ion-electron equations which include $H^+(aq)$ and $H_2O(l)$ can be written. 1. **Write down the main species involved in the reaction** $IO_3^- \rightarrow I_2$ 2. **Balance all atoms except O and H** $2IO_3^- \rightarrow I_2$ 3. **Add H_2O to other side to balance O atoms** $2IO_3^- \rightarrow I_2 + 6H_2O$ 4. **Add H^+ ions to other side to balance H atoms** $2IO_3^- + 12H^+ \rightarrow I_2 + 6H_2O$ 5. **Add e^- to most positive side to balance charge** $2IO_3^- + 12H^+ + 10e^- \rightarrow I_2 + 6H_2O$	46	☹	😐	☺
45	Ion-electron equations can be combined to produce redox equations Reduction: $I_2 + 2e^- \rightarrow 2I^-$ Oxidation: $2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2e^-$ Redox: $I_2 + 2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 2I^-$ Where the electrons do not cancel out, ion-electron equations may have to be multiplied: ❶ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ ❷ $Fe^{2+} \rightarrow Fe^{3+} + e^-$ ❶ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ ❷x5 $5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$ add and cancel down $MnO_4^- + 8H^+ + \cancel{5e^-} \rightarrow Mn^{2+} + 4H_2O$ $5Fe^{2+} \rightarrow 5Fe^{3+} + \cancel{5e^-}$ redox $MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}$	47	☹	😐	☺