

Atomic Structure

(a) describe protons, neutrons and electrons in terms of relative charge and relative mass;

	Proton	Neutron	Electron
Relative Charge	+1	0	-1
Relative Mass	1	1	1/2000

(b) describe the distribution of mass and charge Within an atom;

**Positively charged Nucleus containing most of the mass, surrounded by atomic shells with orbiting electrons of negative charge and negligible mass.**

(c) describe the contribution of protons and neutrons to the nucleus of an atom, in terms of atomic (proton) number and mass (nucleon) number.

**The atomic number of the nucleus, also the proton number, shows the number of protons in the nucleus. It also defines the element. The number of neutrons added to the number of protons is the nucleon number or mass number.**

**Eg.  $^{12}\text{C}_6$  - 12 is the nucleon number, and 6 is the proton number. The number of neutrons can be found by subtracting the atomic number from the nucleon number.**

(d) deduce the numbers of protons, neutrons and electrons in:

**The number of protons is equal to the atomic number, coincidentally so is the number of electrons in most elements. The number of neutrons can be found by subtracting the proton number from the mass number.**

(i) an atom given its atomic and mass number,

**The element is always given by its proton number.**

**$^{12}\text{C}_6$  is Carbon, however  $^{14}\text{C}_6$  is also carbon, albeit an Isotope of carbon.**

(e) explain the term isotopes.

**Atoms with the same number of protons, although a different number of neutrons.**

*(f) state that  $^{12}\text{C}$  is used as the standard measurement of relative masses;*

*(g) define the terms relative isotopic mass and relative atomic mass, based on the  $^{12}\text{C}$  scale;*

**Relative isotopic mass – Mass of an atom of an isotope compared with one twelfth of the mass of an atom of Carbon 12**

**Relative atomic mass – Average weighted mass of an atom of an element compared with one twelfth of the mass of an atom of Carbon 12.**

### Relative Masses

*(a) explain the terms:*

*(i) amount of substance,*

**A quantity whose unit is the mole, used as a means of counting the number of atoms.**

*(ii) mole (symbol 'mol'), as the unit for amount of substance,*

*(iii) the Avogadro constant,  $N_A$ , as the number of particles per mole ( $6.02 \times 10^{23} \text{ mol}^{-1}$ );*

*(b) define and use the term molar mass (units  $\text{g mol}^{-1}$ )*

### **Mass per mole of a substance**

### Empirical and Molecular Formula

*(c) explain the terms:*

*(i) empirical formula as the simplest whole number ratio of atoms of each element present in a compound,*

*(ii) molecular formula as the actual number of atoms of each element in a molecule;*

*(d) calculate empirical and molecular formulae, using composition by mass and percentage compositions;*

$$\text{Empirical formula} = \frac{\text{Mass/Percentage of Composition}}{\text{Molar Mass}} : \frac{\text{Mass/Percentage of Composition}}{\text{Molar Mass}}$$

**Then divide by the smallest number of moles to find out the ratio.**

### Chemical Equations

*(e) construct balanced chemical equations for reactions studied and for unfamiliar reactions given reactants and products;*

**Things to remember:**

- 1. Group 7, oxygen, nitrogen and hydrogen are diatomic**
- 2. State symbols should usually be included in the answer, even if not asked for.**
- 3. Answers should always be balanced, again, even if not asked for.**

### Calculation of reacting masses, moles, concentrations and volumes of gases

*(f) carry out calculations, using amount of substance in mol, involving:*

*(i) mass,*

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass}}$$

*(ii) gas volume,*

$$\text{Number of moles} = \frac{\text{Volume(dm}^3\text{)}}{24}$$

*(iii) solution volume and concentration;*

$$\text{Number of moles} = \text{Concentration} \times \text{Volume (dm}^3\text{)}$$

*(g) deduce stoichiometric relationships from calculations;*

**Things to remember:**

1. Stoichiometric value determines the ratio between the number of moles of a substance.
2. This ratio does not apply to any other calculation apart from working out the number of moles. After the number of moles is worked out for each substance, the ratio can be ignored.

*(h) use the terms concentrated and dilute as qualitative descriptions for the concentration of a solution.*

**Concentration: the amount of solute, in mol, per  $1\text{dm}^3$  of solution**

**Concentrated: a high amount of solute per  $\text{dm}^3$  of solution**

**Dilute: a low amount of solute per  $\text{dm}^3$  of solution.**

**Acids and Bases**

*(a) explain that an acid releases  $\text{H}^+$  ions in aqueous solution;*

**When acids are added to water,  $\text{H}^+$  ions are released into the solution. An acid is therefore a proton donor.**

*(b) state the formulae of the common acids: hydrochloric, sulfuric and nitric acids;*

**Hydrochloric Acid:  $\text{HCl}$**

**Sulphuric Acid:  $\text{H}_2\text{SO}_4$**

**Nitric Acid:  $\text{HNO}_3$**

*(c) state that common bases are metal oxides, metal hydroxides and ammonia;*

**Common Metal Oxides:  $\text{MgO}$  and  $\text{CuO}$**

**Common Metal Hydroxides:  $\text{NaOH}$  and  $\text{Mg}(\text{OH})_2$**

**Ammonia –  $\text{NH}_3$**

**Amines –  $\text{CH}_3\text{NH}_2$**

**Bases are proton acceptors and they neutralise acids.**

(d) state that an alkali is a soluble base that releases  $\text{OH}^-$  ions in aqueous solution;

**$\text{OH}^-$  ions are added to neutralise  $\text{H}^+$  ions, hence why bases neutralise acids.**

(e) state the formulae of the common alkalis: sodium hydroxide, potassium hydroxide and aqueous ammonia;

**Sodium Hydroxide: NaOH**

**Potassium Hydroxide: KOH**

**Ammonia:  $\text{NH}_3$**

**Alkali's are special bases that dissolve in water forming aqueous hydroxide ions ( $\text{OH}^-$ )**

e.g  $\text{NaOH}_{(s)} + \text{aq} \rightarrow \text{Na}^+_{(aq)} + \text{OH}^-$

**Hydroxide ions from alkalis neutralise the protons from acids, forming water:**

$\text{H}^+_{(aq)} + \text{OH}^-_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)}$

**Salts**

(f) explain that a salt is produced when the  $\text{H}^+$  ion of an acid is replaced by a metal ion or  $\text{NH}_4^+$ ;

(g) describe the reactions of an acid with carbonates, bases and alkalis, to form a salt;

**Carbonates:**

**Acid + Carbonate  $\rightarrow$  Salt + Carbon Dioxide + Water**

**Bases:**

**Acid + Base  $\rightarrow$  Salt + Water**

**Alkalis**

**Acid + Alkali  $\rightarrow$  Salt + Water**

**Ammonium Salts**

**Acid + Aqueous Ammonia  $\rightarrow$  Ammonium Salt**  
 $\text{NH}_3 + \text{HNO}_3 \rightarrow \text{NH}_4\text{NO}_3$

(h) explain that a base readily accepts  $H^+$  ions from an acid: eg  $OH^-$  forming  $H_2O$ ;  $NH_3$  forming  $NH_4^+$

(i) explain the terms anhydrous, hydrated and water of crystallisation;

**Anhydrous:** Substance containing no water molecules.

**Hydrated:** Crystalline compound containing water molecules.

**Water of Crystallisation:** water molecules that form an essential part of the crystalline structure of a compound.

(j) calculate the formula of a hydrated salt from given percentage composition, mass composition or experimental data;

1. Mass of hydrated – Mass of anhydrous = Mass of  $H_2O$
2. Calculate number of moles of anhydrous substance
3. Calculate number of moles of water
4. Look at the molar ratio and divide by the smallest number of find the formula of the hydrated salt.

### Oxidation Number

(a) apply rules for assigning oxidation number to atoms in elements, compounds and ions;

### Oxidation Rules

<u>Species</u>	<u>Oxidation Number</u>	<u>E.gs</u>
<u>Uncombined Element</u>	<u>0</u>	<u>C / Na / <math>O_2</math> / <math>P_4</math></u>
<u>Combined oxygen</u>	<u>-2</u>	<u><math>H_2O</math> / CaO</u>
<u>Combined hydrogen</u>	<u>+1</u>	<u><math>NH_3</math> / <math>H_2S</math></u>
<u>Simple Ion</u>	<u>Same as charge on ion</u>	<u><math>Na^+ = 1</math> / <math>Mg^{2+} = 2</math></u>
<u>Combined Fluorine</u>	<u>-1</u>	<u>NaF / <math>CaF_2</math></u>

(b) describe the terms oxidation and reduction in terms of:

(i) electron transfer,

If the number of electrons increases a reduction occurs  
If the number of electrons decreases an oxidation occurs

(ii) changes in oxidation number;

**If the oxidation number of an element increases, it is oxidised.**

**If the oxidation number of an element decreases, it is reduced.**

(c) use a Roman numeral to indicate the magnitude of the oxidation state of an element, when a name may be ambiguous, eg nitrate(III) and nitrate(V);

**The roman numerals state the oxidation state of an ambiguous element, usually an element from the transition metals.**

**I = 1**

**II = 2**

**III = 3**

**IV = 4**

**V = 5**

**VI = 6**

**VII = 7**

**VIII = 8**

**E.g – Sodium Chlorate (V) = NaClO<sub>3</sub>**

+1 ? -2x3 = -6

**-6 + +1 = -5, therefore the chlorine needs to be +5 in order to make the overall oxidation number 0.**

### Redox Reactions

(e) explain that:

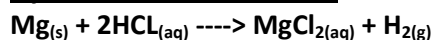
(i) metals generally form ions by losing electrons with an increase in oxidation number to form positive ions,

(ii) non-metals generally react by gaining electrons with a decrease in oxidation number to form negative ions;

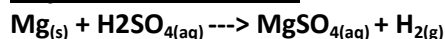
(f) describe the redox reactions of metals with dilute hydrochloric and dilute sulphuric acids;

**In redox reactions of dilute hydrochloric acid and dilute sulphuric acid, a salt of the acid plus hydrogen are made:**

#### Hydrochloric acid reaction:



#### Sulphuric Acid reaction:



## Ionisation Energy

*(a) Define the terms first ionisation energy and successive ionisation energy;*

**First ionisation energy** – Energy required to remove one electron from each atom in one mole of gaseous atoms.

**Successive ionisation energy** – energy required to remove each electron in turn.

*(b) Explain that ionisation energies are influenced by nuclear charge, electron shielding and the distance of the outermost electron from the nucleus;*

**Nuclear Charge** – The greater the nuclear charge, the greater the attractive force of the outer electrons. Therefore the more energy needed to remove an electron from the outer shell. This increases across a period.

**Electron Shielding** – more inner electron shells shield the nuclear charge from the outer electron, so the electron is easier to remove. Shielding increases down a group.

**Atomic Radius** – in larger atoms, the outer electrons are further from the nucleus due to the number of shells, this means that nuclear attraction is lower so the energy needed to remove an electron is much lower. This also increases down a group.

*(c) predict from successive ionisation energies of an element:*

All group 1,2 and 3 elements have relatively low ionisation energies, although there is a large jump in all atoms when another shell is entered when the electrons in the most outer shell are all lost. The next electron is closer to the nucleus so atomic radius and electron shielding decrease, so nuclear attraction is also greater, meaning more energy is required to remove an electron from the outer shell.

*(i) the number of electrons in each shell of an atom,*

Although not strictly true, electrons are often seen as having a 2,8,8,8 etc formation; 2 electrons in the first shell, followed by 8 in the succeeding shells.



Electrons: electronic energy levels, shells, sub-shells, atomic orbital's, electron Configuration

*(d) state the number of electrons that can fill the first four shells;*

**1<sup>st</sup> shell – 2**  
**2<sup>nd</sup> shell – 8**  
**3<sup>rd</sup> shell – 18**  
**4<sup>th</sup> shell - 32**

*(e) describe an orbital as a region that can hold up to two electrons, with opposite spins;*

*(f) describe the shapes of s and p orbitals;*

**s orbitals are spherical and can hold two electrons in total**  
**p orbitals are in the shape of a dumb-bell, and can also hold two electrons in each orbital.**

*(g) state the number of:*

*(i) orbitals making up s-, p- and d-subshells,*

**In an s subshell, there is only 1 orbital, which can hold 2 electrons of opposite spin.**  
**In a p subshell, there are 3 orbitals which can, again, each hold two electrons of opposite spin.**  
**In a d orbital, there are 5 orbitals, each holding two electrons of opposite spin.**

*(ii) electrons that occupy s-, p- and d-subshells;*

**Each orbital can hold two electrons:**  
**A s subshell has 1 orbital, so can hold 2 electrons**  
**A p subshell has 3 orbitals, so can hold 6 electrons.**  
**A d subshell has 5 orbitals, so can hold 10 electrons.**  
**A f subshell had 7 orbitals, so can hold 14 electrons.**

*(h) describe the relative energies of s-, p- and d orbitals for the shells 1, 2, 3 and the 4s and 4p orbitals;*

**The 1s subshell has the lowest relative energy so the electrons must fill this shell first.**  
**The 2s subshell has the second lowest relative energy, so must therefore be filled second, followed by the 2p subshell.**  
**In the 3<sup>rd</sup> shell, the 3s subshell has the lowest relative energy so must be filled first, followed by the 3p subshell. After this, the 4s subshell is filled next as it has a lower relative energy than the 3d subshell. The 3d subshell is filled afterwards, followed by the 4p, then 4d, then 4f.**

(j) classify the elements into s, p and d blocks.

The s-block of the periodic table consists of the first two groups, plus hydrogen and helium.

The p block of the periodic table is from group 3 to group 0 minus helium

The d block are the transition metals.

### Ionic Bonding

(a) describe the term ionic bonding as electrostatic attraction between oppositely charged ions;

(b) construct 'dot-and-cross' diagrams, to describe ionic bonding;

(c) predict ionic charge from the position of an element in the Periodic Table;

The ionic charge of an atom can be found by looking at what group it is in. If it is in group 1, the charge is usually +, if in group 2 the charge is usually – etc. But then it gets to group 4 the charge can be either +4 or -4. At group 5 the charge is usually 3- and so on.

Be, B, C and Si do not normally form ions and too much energy is needed to transfer the outer-shell electrons to form ions.

(d) state the formulae for the following ions:  $\text{NO}_3^-$ ,  $\text{CO}_3^{2-}$ ,  $\text{SO}_4^{2-}$  and  $\text{NH}_4^+$ ;

$\text{NO}_3^-$  - Nitrate

$\text{CO}_3^{2-}$  - Carbonate

$\text{SO}_4^{2-}$  - Sulphate

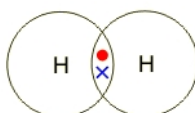
$\text{NH}_4^+$  - Ammonium

### Covalent bonding and dative covalent (coordinate) bonding

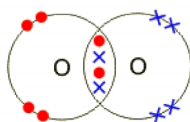
(e) describe the term **covalent bond** as a **shared pair of electrons**;

(f) construct 'dot-and-cross' diagrams to describe:

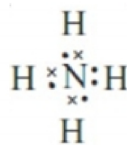
(i) single covalent bonding, eg as in  $\text{H}_2$ ,  $\text{Cl}_2$ ,  $\text{HCl}$ ,  $\text{H}_2\text{O}$ ,  $\text{NH}_3$ ,  $\text{CH}_4$ ,  $\text{BF}_3$  and  $\text{SF}_6$ ,



(ii) multiple covalent bonding, eg as in  $O_2$ ,  
 $N_2$  and  $CO_2$ ,



(iii) dative covalent (coordinate) bonding,  
eg as in  $NH_4^+$ ,



(iv) molecules and ions analogous to those  
specified in (i), (ii) and (iii);

### Shapes of simple molecules and ions

(g) explain that the shape of a simple molecule is  
determined by repulsion between electron  
pairs surrounding a central atom;

(h) state that lone pairs of electrons repel more  
than bonded pairs;

(i) explain the shapes of, and bond angles in,  
molecules and ions with up to six electron  
pairs (including lone pairs) surrounding a  
central atom, eg as in:

(i)  $BF_3$  (trigonal planar), -  $120^\circ$

(ii)  $CH_4$  and  $NH_4^+$  (tetrahedral), -  $109.5^\circ$

(iii)  $SF_6$  (octahedral), -  $90^\circ$

(iv)  $NH_3$  (pyramidal), -  $107^\circ$

(v)  $H_2O$  (non-linear), -  $104.5^\circ$

(vi)  $CO_2$  (linear); -  $180^\circ$

(j) predict the shapes of, and bond angles in,  
molecules and ions analogous to those  
specified in (i);

### Electronegativity and bond polarity

(k) describe the term electronegativity as the  
ability of an atom to attract the bonding  
electrons in a covalent bond;

*(l) explain that a permanent dipole may arise when covalently-bonded atoms have different electro negativities, resulting in a polar bond;*

If an atom is more electronegative than another atom, it will have a greater attraction for the bonding pair of electrons than the other atom, so the bonding electrons will be closer to it. This creates a small charge difference which is a permanent dipole. Since the electrons are not evenly distributed, the bond is now polar.

### Intermolecular forces

*(m) describe intermolecular forces based on permanent dipoles, as in hydrogen chloride, and induced dipoles (van der Waals' forces), as in the noble gases;*

#### Permanent Dipoles:

Polar molecules have permanent dipoles. The permanent dipole of one molecules attracts another permanent dipole to form a weak permanent dipole-dipole force.



#### Van der Waals forces:

1. The movement of electrons in shells unbalances the distribution of the charge in the electron shells.
2. An instantaneous dipole is formed due to the unbalanced distribution of electrons.
3. The instantaneous dipole induces a dipole in neighbouring molecules, which then in turn induce further molecules.
4. The induced dipoles attract each other, forming weak intermolecular bonds.

Having a greater number of electrons increases Van der Waals forces, it causes a greater attraction force between molecules.

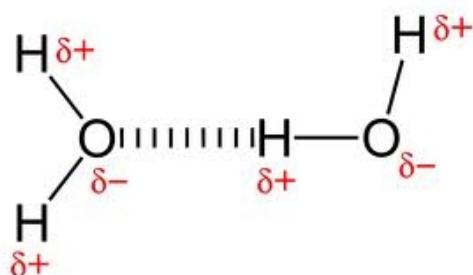
**Furthermore**, as the number of electrons increases so does van der Waals' forces which will increase the boiling point in gases, such as the Noble gases.

(n) describe hydrogen bonding, including the role of a lone pair, between molecules containing  $-OH$  and  $-NH$  groups, ie as in  $H_2O$ ,  $NH_3$  and analogous molecules;

**A hydrogen bond is a strong dipole-dipole interaction between an electron deficient hydrogen atom and a lone pair of electrons on a highly electronegative atom. (Usually oxygen or nitrogen)**

**Molecules containing O-H or N-H bonds are polar with permanent dipoles, which are particularly strong.**

**The permanent dipole-dipole interaction between O-H and N-H bonds is a Hydrogen bond.**



(o) describe and explain the anomalous properties of  $H_2O$  resulting from hydrogen bonding, eg:

(i) the density of ice compared with water,

**Ice has an open lattice with hydrogen bonds holding the water molecules apart, when ice melts the rigid hydrogen bonds collapse, allowing the water molecules to move closer together. The distance the molecules are held apart is what makes ice less dense than water.**

(ii) its relatively high freezing point and boiling point;

**Hydrogen bonds are much stronger than Van der Waals' forces, this means more energy is needed to break hydrogen bonds, hence the higher melting point and boiling point. (Freezing point and melting point are theoretically the same as just above the melting point, something will melt, although just below a melting point a substance will freeze)**

(p) describe *metallic bonding* as the attraction of positive ions to delocalised electrons;

### Bonding and physical properties

(q) describe structures as:

(i) giant ionic lattices, with strong ionic bonding, ie as in NaCl,

Formed by the attraction of oppositely charged ions.

1. Each ion is surrounded by oppositely charged ions.
2. The ions attract each other to form a giant ionic lattice.

(ii) giant covalent lattices, ie as in diamond and graphite,

A three-dimensional structure of atoms bonded together by strong intramolecular, covalent bonds.

Examples are Diamond (C), Graphite (C) or Silicon dioxide (SiO<sub>2</sub>)

(iii) giant metallic lattices,

1. Contains ionised atoms.
2. Positive ions occupy fixed positions in a lattice.
3. Delocalised outer shell electrons.
4. The delocalised electrons are spread throughout the metallic structure.
5. The electrons can move within the structure.
6. The charges balance over the whole structure.

(iv) simple molecular lattices, ie as in I<sub>2</sub> and ice;

A three dimensional structure of molecules, bonded together by weak intermolecular forces.

1. The molecules are held together by weak forces between them.
2. The atoms within each molecule are bonded strongly together by covalent bonds.

When a simple molecular structure is broken, it is the weak Van der Waals' and not the covalent bonds that are broken.

*(r) describe, interpret and/or predict physical properties, including melting and boiling points, electrical conductivity and solubility.*

### Metallic structures:

1. High melting and boiling point
  - a. The electrons can move within the structure, but the ions cannot.
  - b. The attraction between the ions and electrons is strong.
  - c. High temperatures are needed to break the metallic bonds and dislodge the ions from their positions in the lattice.
2. Good Electrical conductivity
  - a. The delocalised electrons can move freely within the lattice.
  - b. This allows the metal to conduct in solid state.
3. Malleable and Ductile
  - a. The delocalised electrons can move, giving the structure a degree of 'give' which allows atoms or layers to slide past each other.

### Giant Ionic Compounds

1. High melting point and boiling point
  - a. Compounds are solid at room temperature
  - b. The ions are held together by strong electrostatic forces, so a large amount of energy is needed to break the bonds.
  - c. The charge also affects the melting point. The greater the charge, the stronger the electrostatic force between ions, so more energy is needed to break the bonds.
    - i. E.g – NaCl = Na<sup>+</sup> and Cl<sup>-</sup> has a lower melting point than MgO = Mg<sup>2+</sup> and O<sub>2</sub><sup>-</sup>
2. Doesn't conduct when solid
  - a. The ions are in a fixed position and cannot move
3. Does conduct when molten.
  - a. The ions are free to move
4. Can dissolve in polar solvents, such as water.
  - a. Water molecules break down the lattice by surrounding each ion, forming a solution.
    - i. In NaCl, water molecules attract Na<sup>+</sup> and Cl<sup>-</sup> ions
    - ii. The lattice dissolves and breaks down.
    - iii. Water molecules surround the ion.
    - iv. Na<sup>+</sup> attracts δ<sup>-</sup> charges on O atoms of the water molecules.
    - v. Cl<sup>-</sup> attractions δ<sup>+</sup> charges on the H atoms of the water molecules.

### Simple molecular structures

1. Low melting and boiling point.
  - a. Weak Van der Waals' forces so a small measure of energy is needed to break the bonds.
2. Non conductors of energy
  - a. No charged particles free to move.
3. Soluble in Non-polar substances
  - a. Van der Waals' forces form between simple molecular structures and the non polar solvent
  - b. The formation of Van der Waals' forces, however, weakens the lattice structure.

**Giant Covalent Structures:**

1. High melting point and boiling point.
  - a. High temperatures needed to break the strong covalent bonds in the lattice.
2. Non conductors of electricity
  - a. No free charged particles (except in Graphite – See below)
3. Insoluble in both polar and non-polar solvents.
  - a. Covalent bonds in the lattice are too great to be broken by both solvents.

**Examples****1. Diamond**

- a. Poor conductivity
  - i. No delocalised electrons, all outer shell electrons are used for the covalent bonds.
- b. Strong
  - i. Tetrahedral shape allows external forces to be spread throughout the lattice

**2. Graphite**

- a. Good conductivity
  - i. Delocalised electrons between layers.
  - ii. Electrons are free to move parallel to the layers where a voltage is applied.
- b. Soft
  - i. Bonding with each layer is strong
  - ii. Weak Van der Waals' forces between layers allow them to slide easily

*(s) deduce the type of structure and bonding present from given information.*

	<i>Metallic bonding</i>	<i>Ionic compound</i>	<i>Simple molecular</i>	<i>Giant covalent</i>
<i>High Melting point</i>	✓	✓	X	✓
<i>Can conduct electricity</i>	✓	✓ (Only when molten)	X	X (except in graphite)
<i>Soluble in polar substances</i>	X	✓	X	X
<i>Soluble in non-polar substances</i>	X	X	✓	X
<i>Formed with a metal and a non metal</i>	X	✓	X	X



The structure of the Periodic Table in terms of groups and periods

*(a) describe the Periodic Table in terms of the arrangement of elements:*

- (i) by increasing atomic (proton) number,**
- (ii) in periods showing repeating trends in physical and chemical properties,**
- (iii) in groups having similar physical and chemical properties, due to having the same number of electrons in the outer shell;**

*(b) describe periodicity in terms of a repeating pattern across different periods;*

*(c) explain that atoms of elements in a group have similar outer shell electron configurations, resulting in similar properties;*

Periodicity of physical properties of elements

*(d) describe and explain the variation of the first ionisation energies of elements shown by:*

*(i) a general increase across a period, in terms of increasing nuclear charge,*

- 1. As you go across a period, the number of protons increases, so there is more attraction towards the electrons so more energy is needed to remove an electrons from the outer shell.**
- 2. Across a period, electrons are added to the outer shell so it is drawn inwards slightly, but the electron shielding hardly changes.**

*(ii) a decrease down a group in terms of increasing atomic radius and increasing electron shielding outweighing increasing nuclear charge;*

- 1. Down a group, the number of shells increases. This increases the atomic radius so therefore the attraction between the nucleus and the electrons is less. So ionisation energy decreases.**
- 2. Down a group, there are also more inner shells so the shielding effect of in the outer electrons increases, so there is, again, less attraction between the nucleus and the energy.**
- 3. The nuclear charge also increases, but this doesn't have much of an effect on the ionisation energy.**

*(e) for the elements of Periods 2 and 3:*

*(i) describe the variation in electron configurations, atomic radii, melting points and boiling points,*

1. Period 2 elements all have 2 outer shells, whereas period 3 elements all have 3 outer shells.
2. Period 3 had an extra electron shell than period 2, so the atomic radius is larger.
3. Boiling points in group 1 to 4 in both periods are fairly high, this is because group 1 to 4 have strong forces holding the structures together.
4. Between group 4 and 6, there is a sharp decrease in boiling point, this marks the change between the strong forces between atoms, to the weak forces between molecules.
5. In group 5 to 8, the boiling points are all comparatively low due to the weak forces.

*(ii) explain variations in melting and boiling points in terms of structure and bonding;*

In Period 2, the first two groups are metals so the bonding is metallic, which is a strong force between the molecules. Group 3 and 4 are then giant covalent structures, so still have strong forces between the atoms, so the boiling point remains high. After group 4 in Period 2, the forces change to weak Van der Waals' forces as the elements are no longer metals. This causes a sharp decrease in melting and boiling point as less energy is needed to break the bonds between the molecules.

This is the same in period 3, apart from group 3 is also metallic bonding.

*(f) interpret data on electron configurations, atomic radii, first ionisation energies, melting points and boiling points to demonstrate periodicity.*

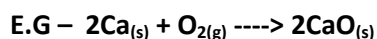
1. If elements have the same number of electrons in the outer shell, they usually have similar properties, this repeats down a group.
2. The atomic radius increases down a group, which causes a decrease in ionisation energy but also an increase in reactivity due to it becoming easier to lose an outer electrons. This also repeats down a group.
3. The number of protons increasing across a period which increases nuclear attraction, therefore increasing the ionisation energy across a period. This is also a repeating pattern.
4. The melting and boiling point also increase up to group 4, and then decrease sharply afterwards, staying relatively low until the next period is reached. This is due to the different types of bonds the atoms and molecules experience. This is also repeated down a period.

### Redox reactions of Group 2 metals

*(a) describe the redox reactions of the Group 2 elements Mg - Ba:*

*(i) with oxygen,*

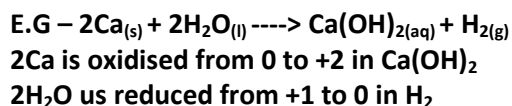
**Group 2 elements react vigorously with oxygen. A redox reaction occurs where an ionic oxide is produced with the formula MO (where M is the group 2 metal)**



**The oxidation number of calcium changes from 0 in 2Ca to +2 in 2CaO, this is an oxidation  
The oxidation number of oxygen changes from 0 in O<sub>2</sub> to -2 in 2CaO, this is a reduction.**

*(ii) with water;*

**Group 2 elements react with water to form hydroxides, with the general formula M(OH)<sub>2</sub>  
Hydrogen gas is also formed.**



*(b) explain the trend in reactivity of Group 2 elements down the group due to the increasing ease of forming cations, in terms of atomic size, shielding and nuclear attraction;*

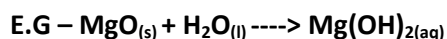
**Group 2 elements are both reactive metals and strong reducing agents.**

- 1. Group 2 elements are oxidised in reactions. Each atom loses two electrons from its outer s subshell to form a 2<sup>+</sup> ion.**
- 2. Reactivity increases down the group, due to increased atomic radius and nuclear shielding the forces of attraction between the electrons and the nucleus is lower, so it is easier to lose electrons down the group.**

### Reactions of Group 2 compounds

*(c) describe the action of water on oxides of elements in Group 2 and state the approximate pH of any resulting solution;*

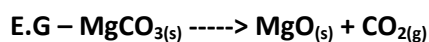
**Group 2 oxides react with water to form a solution of metal hydroxide.**



**The typical pH of these solutions is between 10-12**

*(d) describe the thermal decomposition of the carbonates of elements in Group 2 and the trend in their ease of decomposition;*

**Group 2 carbonates are decomposed by heat, forming a metal oxide and carbon dioxide.**



**This is a thermal decomposition.**

**The carbonates become more difficult to decompose with heat as you go down a group.**

*(e) interpret and make predictions from the chemical and physical properties of Group 2 elements and compounds;*

- 1. Group 2 elements are more reactive as you go down the group.**
- 2. Group 2 carbonates decompose at higher temperatures down the group.**
- 3. Hydroxides are more soluble in water, and the resulting solutions become more alkaline down the group.**

*(f) explain the use of  $\text{Ca(OH)}_2$  in agriculture to neutralise acid soils; the use of  $\text{Mg(OH)}_2$  in some indigestion tablets as an antacid.*

- 1. Calcium Hydroxide ( $\text{Ca(OH)}_2$ ) is used as 'lime' by farmers and gardeners to neutralise acidic soils, using too much makes the soil too alkaline.**
- 2. Magnesium hydroxide ( $\text{Mg(OH)}_2$ ) is used in 'milk of magnesia' to relieve indigestion. It neutralises excess acid in the stomach.**

Characteristic physical properties

(a) explain, in terms of van der Waals' forces, the trend in the boiling points of  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$ ;

As you move down group 7, the boiling point increases.

1. The number of electrons increases, leading to an increase in Van der Waals forces between molecules.
2. Therefore more energy is needed to break the bonds, so hence a higher boiling point.

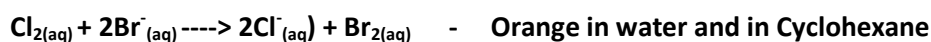
Redox reactions and trends in reactivity of Group 7 elements and their compounds

(b) describe the redox reactions, including ionic equations, of the Group 7 elements  $\text{Cl}_2$ ,  $\text{Br}_2$  and  $\text{I}_2$  with other halide ions, in the presence of an organic solvent, to illustrate the relative reactivity of Group 7 elements;

Halogens form solutions of different colours.

Halogen	Water	Cyclohexane
$\text{Cl}_2$	Pale green	Pale green
$\text{Br}_2$	Orange	Orange
$\text{I}_2$	Brown	Violet

Any colour change will show whether a redox reaction has taken place. The mixture is usually shaken with an organic solvent, Cyclohexane, to distinguish between bromine and iodine.

Chlorine oxidises both Bromide ions and Iodide ions:Bromine oxidises Iodide ions only:

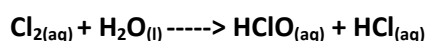
(c) explain the trend in reactivity of Group 7 elements down the group from the decreasing ease of forming negative ions, in terms of atomic size, shielding and nuclear attraction;

As you go down group 7, they become less reactive.

1. Atomic radius increases
2. Electron shielding increases
3. The nuclear attraction between the electrons and the nucleus therefore increases, making it harder for the outer shell to gain an electron into the p-subshell.

(d) describe the term disproportionation as a reaction in which an element is simultaneously oxidised and reduced, illustrated by:

(i) the reaction of chlorine with water as used in water purification,



(ii) the reaction of chlorine with cold, dilute aqueous sodium hydroxide, as used to form bleach,



The above two equations are both example of disproportionation reactions as the chlorine is both oxidised and reduced in both.

(e) interpret and make predictions from the chemical and physical properties of the Group 7 elements and their compounds;

Group 7 all share the same characteristics

<b>Bonding</b>	<b>Covalent diatomic molecules (I<sub>2</sub>, F<sub>2</sub> etc)</b>
<b>Structure</b>	<b>Simple molecular Held together by van der Waals' forces between molecules.</b>
<b>Redox Character</b>	<b>Oxidising agents. Halogen + electron ----&gt; Halide ion</b>
<b>Reactivity</b>	<b>Decreases down the group.</b>

Some Halides compounds are used in products as most are very stable.

1. NaCl is common salt
2. NaF and SnF<sub>2</sub> are fluoride compounds added to toothpaste to prevent tooth decay.
3. CaF<sub>2</sub> or fluorite is used to makes lenses to focus infrared light.

*(f) contrast the benefits of chlorine use in water treatment (killing bacteria) with associated risks (hazards of toxic chlorine gas and possible risks from formation of chlorinated hydrocarbons);*

**Chlorine makes water safer to drink by sterilising it through killing bacteria. However, chlorine reacts with organic matter to form chlorinated hydrocarbons, which may cause cancer. Chlorine is also toxic.**

### Characteristic reactions of halide ions

*(g) describe the precipitation reactions, including ionic equations, of the aqueous anions Cl<sup>-</sup>, Br<sup>-</sup> and I<sup>-</sup> with aqueous silver ions, followed by aqueous ammonia;*

1. An unknown halide is dissolved in water.
2. Aqueous silver nitrate (AgNO<sub>3</sub>) is added.
3. The Ag<sup>+</sup> ions react with the halide ions to form coloured precipitates.
4. If unsure about the colour, adding aqueous ammonia (NH<sub>3</sub>) may help distinguish as different precipitates have different solubility in the ammonia.

### Chloride



### Bromide



### Iodide



**These types of reactions are called precipitation reactions. They take place in aqueous solutions when aqueous ions react together to form a solid precipitate.**