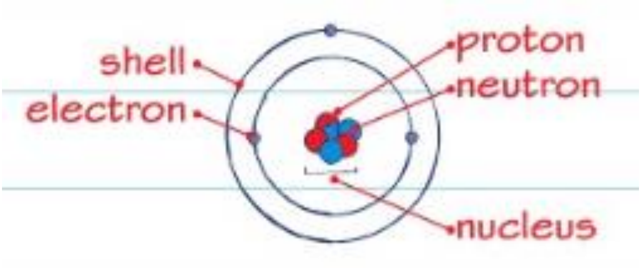
Atomic Structure

C3/C4

Structure of an atom



Relative Charges and Relative Mass

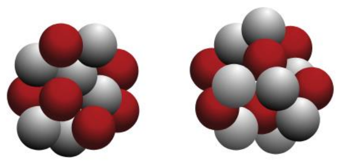
- Most of the mass of an atom is found within the nucleus.
- The overall charge of an atom is neutral due to the atom having the same number of positively charged protons and negatively charged electrons.

Particle	Relative charge	Relative mass
Proton	+1	1
Neutron	0	1
Electron	-1	1/1836

Mass Number: protons + neutrons
(Does not have to be an integer)
Atomic Number: the number of protons
(also equal to the number of electrons and unique to the element)

5 B 10.811	6 C 12.01115	7 N 14.0067	8 O 15.9994	9 F 18.9984
13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453

Isotope: A different atom of the same element. The atomic number is the same but the mass number is different due to the change in the number of neutrons.



C^{12}_{6} C^{14}_{6}

Using atomic number and mass number

- Complete the table:

Element	Symbol	Mass number	Atomic number	No. of protons	No. of neutrons	No. of electrons
Lithium	Li	7	3	3	7 - 3 = 4	3
Sodium						
	Pb					
	Ti					
					0	
				72		
						15

Mendeleev arranged the elements, known at that time, in a periodic table by using properties of these elements and their compounds.

Mendeleev used his table to predict the existence and properties of some elements not yet discovered.

Mendeleev thought that he had arranged elements in order of increasing mass number but this was not always true because of the abundance of isotopes of some pairs in the periodic table.

If the elements are arranged in terms of increasing atomic number then Mendeleev's **pair reversals** are explained.

Iodine should be placed before Tellurium according to its mass number, but after Tellurium using the atomic number. Mendeleev did not know about atomic structure.

Relative atomic mass	128	127
Element symbol	Te	I
Atomic number	52	53

In the periodic table elements are arranged in order of increasing atomic number, in **rows called periods**. Elements with similar properties are placed in the same vertical **columns called groups**.

Metals can be found on the **left hand side** and in the centre. **Non-metals** are on the **right hand side**.

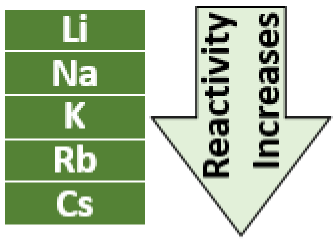
Electrons fill their shells 2.8.8.2 You must be able to draw and write the electronic configurations of the fist 20 elements in the periodic table.

<u>Groups in the periodic table</u>

Paper 4

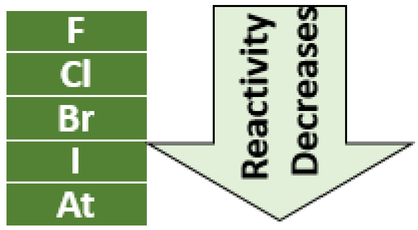
Group 1 – Alkali Metals

- Soft metals
- Low melting points
- Lithium, sodium and potassium all react **vigorously** with water.
  When you add them to water, the metal **floats**, **moves** around and **fizzes**.
  e.g. **potassium + water → potassium hydroxide + hydrogen**
  $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
- Stored in oil to keep air and water away.
- Li, Na and K are less dense than water so they float on water.
- The reactivity of the alkali metals increases down the group. This is because the outer electrons become further from the nucleus each time and the force of attraction between the positive nucleus and negative electron become weaker. This allows for the outer electron to be lost more easily as we go down the group.



Group 7 – The Halogens

- Melting point increase down the group
- Boiling point increases down the group
- The intermolecular forces, going down the group become stronger.
- **Chlorine** is a **yellow/green gas**
- **Bromine** is a **red/ brown liquid**
- **Iodine** a **grey solid**.
- As elements, the halogens exist as **diatomic** molecules **F_2, Cl_2, Br_2, I_2 and At_2**



- **Reactivity:** When Group 7 elements react, the atoms **gain** an electron in their outermost shell. Going down the group, the outermost shell's electrons get **further away** from the attractive force of the nucleus, so it is **harder** to **attract** and **gain** an extra electron.

- The halogens react with **hydrogen** to form **hydrogen halides**. The reactions with hydrogen become **less** reactive as you go **down** the group.
  E.g. **fluorine + hydrogen → hydrogen fluoride**
  $F_2(g)$ + $H_2(g)$ → $2HF(g)$

- The halogens also react with **metals**. The halogen atoms **gain** a single electron to give them a stable arrangement of electrons. They form **ionic compound**.
  E.g. **sodium + chlorine → sodium chloride**
  $2Na(s)$ + $Cl_2(g)$ → $2NaCl(s)$

- **Test for Chlorine:** The test for chlorine uses **litmus paper**. When **damp** litmus paper is put into chlorine gas, the litmus paper is bleached and turns **white**.

Group 0 – Noble Gases

- They are **unreactive/inert** because their atoms have stable arrangements of electrons.
- The atoms have **eight** electrons in their **outermost shell**, apart from helium which has just two but still has a complete outer shell.
- The **stable electronic structure** explains why they exist as single atoms; they have **no tendency to react** to form molecules.
- The **boiling points** of the noble gases **get higher going down the group**.

Uses:

- Helium: lifting gas in party balloons and air ships. Less dense than air and inflammable.
- Argon, Krypton, Xenon: filling gas in filament lamps. The filament becomes hot enough to glow, the gas prevents it burning away.
- Argon: Shielding gas during welding. Argon is denser than air so keeps air from the metal and prevents oxidisation of the metal.

Ion: is an atom or group of atoms with a positive or negative charge

Cation: a positively charged ion. Usually formed from Groups 1 & 2

Anion: a negatively charged ion. Usually formed from Groups 7 & 6

Draw 3 ions from each of the above groups

Ionic Bonding

Papers 3 & 4

Ionic Bonding is the total **swapping of electrons**

Formation and naming of Ionic Bonds**:**

Formula	Name	
H^+	hydrogen	
Li^+	lithium	group 1
Na^+	sodium	
K^+	potassium	
Mg^{2+}	magnesium	group 2
Ca^{2+}	calcium	
Ba^{2+}	barium	
Al^{3+}	aluminium	group 3
Ag^+	silver	transition metals
Cu^{2+}	copper	
Zn^{2+}	zinc	
Fe^{2+}	iron(II)	
Fe^{3+}	iron(III)	
NH_4^+	ammonium	compound ion

Positively charged ions formed from hydrogen or metal atoms take the name of the element.

Formula	Name	
F^-	fluoride	group 7
Cl^-	chloride	
Br^-	bromide	
I^-	iodide	
O^{2-}	oxide	group 6
S^{2-}	sulfide	
NO_3^-	nitrate	compound ions
CO_3^{2-}	carbonate	
SO_4^{2-}	sulfate	
OH^-	hydroxide	

OH^- does not follow the -ate rule.

Negatively charged ions formed from a single non metal atom take the name of the element, but end in **-ide**

Negatively charged ions in compounds containing three or more elements, one of which is oxygen, end in **-ate**

Properties of Ionic Compounds

- Bonds are strong due to strong electrostatic forces
- Bonds form a lattice structure which has a regular arrangement of ions
- Ionic compounds usually have high melting points and high boiling points
- Solid at room temperature
- No delocalised (free) electrons, therefore don't conduct electricity in solid state
- Soluble in water
- Once dissolved in aqueous solution electrons are free to move so they conduct electricity

Covalent Bonding: is formed when a pair of electrons are shared between 2 atoms

Covalent bonds are:
- Strong
- Form between non-metal atoms
- Often produce molecules, which can be elements or compounds

Draw a dot cross diagram of the following covalent substances:
Hydrogen, hydrogen chloride, water, methane, oxygen, carbon dioxide

Simple Molecular Substances
- A simple molecule consists of a few atoms joined by strong covalent bonds
- Have low melting points
- Have low boiling points
- Usually a gas or liquid at room temperature
- Do not conduct electricity in any state, because they are not electrically charged and do not contain any delocalised electrons.
- They don't conduct in a solution of water as they are insoluble in water
- Can conduct when dissolved in an acid.

Giant Molecular Substance
- Contain many atoms rather than a few
- Strong covalent bonds
- Regular lattice structure
- High melting point
- High boiling point
- Solid at room temperature

Examples: Diamond and Graphite (similarity and differences)

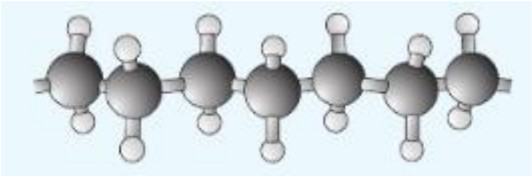
Diamond	Graphite
Carbon bond Where each atom is bonded to four others	Carbon bond Where each atom is bonded to three others
Covalently bonded	Covalently bonded
Lattice structure	Layer structure
Strong intermolecular forces	Weak intermolecular forces between the layers
No delocalised electrons	Delocalised electrons
Cannot conduct electricity	Can conduct electricity
Used in jewellery and cutting tools. Drill bits are diamond tipped	Used in pencils and as a lubricant in industrial engines as the layers can slide over one another. Also used for electrodes as it can conduct electricity due to delocalised electrons

Examples: Graphene and Fullerenes

Graphene	Fullerene
Single layer of graphite	Sheet of graphene rolled to form a buckyball
Carbon atom bonded to three others	C_{60} Buckminsterfullerene has carbon atoms arranged in pentagons and hexagons
Regular lattice structure	Hollow sphere
Conducts electricity due to delocalised electrons	Conducts electricity due to delocalised electrons
Strong and flexible	Soft when solid due to weak intermolecular forces

A **polymer** is a large molecule made from many smaller molecules, called monomers, joined together.

The diagram shows a section of polyethene a simple polymer. It consists of large molecules containing chains of carbon atoms. The atoms are joined to each other, and to hydrogen atoms, by covalent bonds.



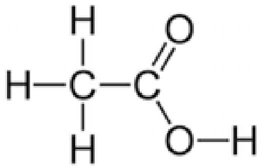
Metals

Property	Metals	Non-metals
Appearance	Shiny	dull
Electrical conduction	Good conductors (delocalised electrons)	Bad conductors
Density	High density	Low density
Melting Point	High melting point (strong electrostatic forces)	Low melting point
Boiling Point	High boiling point (strong electrostatic forces)	Low boiling point
Malleable	Malleable (can be pressed into shape without shattering)	brittle
Solubility in water	Insoluble in water (however, some are soluble in water group 1)	Soluble in water

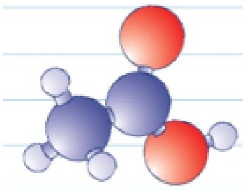
Limitations of models

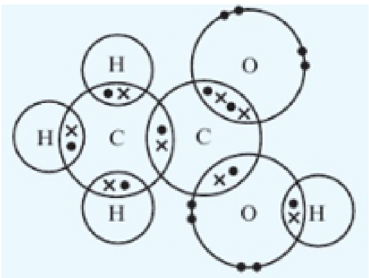
The structure and bonding of different substances are represented using models.

Drawn structures do not show: the 3D shape or the bonding and non-bonding electrons



Ball and Stick models: show how each atom is bonded to the other atoms and its 3D shape. They do not show the bonding and non-bonding electrons or the elements chemical symbol





Dot and Cross Diagram: shows the symbol for each atom in the molecule. It shows how each atom is bonded to other atoms. The pairs of electrons in the covalent bond are shown by the dots and crosses. Non-bonding pairs of electrons in the outer shells are included.

It does not show the 3D shape of the molecule.

Calculations involving Masses

Papers 3 & 4

Hydrogen (H_2)
Water (H_2O)
Potassium Carbonate (K_2CO_3)
Calcium Hydroxide ($Ca(OH)_2$)
Ammonium Sulfate ($(NH_4)_2SO_4$)

The symbol for **relative formula mass is $\boldsymbol{M_r}$**
To get M_r add together the relative atomic masses of all atoms.
Work out the M_r of the following compounds

An **empirical formula** is the simplest whole number ratio of atoms of each element in a compound.
Example: A 10g sample of compound X contains 8g of carbon an 2g of hydrogen.
Calculate the **empirical formula** of compound x

Steps
A 10g sample of compound X contains 8g of carbon an 2g of hydrogen.
Step 1: Write the symbol of each element
Step 2: Write down the mass of each in g
Step 3: Write the atomic mass number of each
Step 4: for each work out the mass divided by atomic mass
Step 5: Divide each answer by the smallest answer
Step 6: Make all numbers whole (integers)

You do: What is the empirical formula of a compound that contains 7.83g of iron and 3.37g of oxygen?

Find the **percentage mass** of Sulphur in Sulfuric Acid (H_2S0_4)

Formula = $\frac{Atomic\ Mass\ x\ no\ of\ element}{M_r} x\ 100$

$\boldsymbol{M_r}$ of Sulphur =
How many Sulphurs =
$\boldsymbol{M_r}$ of Sulfuric Acid =

Example: The empirical formula of Compound G is CH_3
The relative mass (M_r) of compound G is 30
Find the **molecular formula** of compound G

Step 1: Calculate the M_r of the empirical formula
Step 2: Divide M_r of G by empirical M_r
Step 3: Multiply answer by each amount in the empirical formula

<u>Conservation of mass</u>
The total mass of the reactants and products must stay constant during a chemical reaction.
The total mass before a reaction equals the total mass after a reaction.
Closed System: is when no substance can enter of leave the during the reaction.
Non-closed system: is when substances can leave or enter a reaction.

A **solution** is a mixture of a solute in a solvent.
The **solute** is the substance that dissolves.
The **solvent** Is the substances that the solute dissolves in.

The **concentration of a solution** is measured in g/dm^3 or g dm^{-3}

Note: To change cm^3 to dm^3 divide it by 1000

Formula concentration of a solution (g/dm^3) = $\frac{\text{mass of solute (g)}}{volume\ of\ solution\ (dm^3)}$

Example: 2.50g of sodium hydroxide is dissolved in 250cm^3 of water. Calculate the concentration of the solution formed in g dm^{-3}

You do:

1 Calculate the concentrations of the following solutions formed:

(a) 0.40 g of glucose dissolved in 0.50 dm^3 of water.

(b) 1.25 g of copper chloride dissolved in 100 cm^3 of water.

2 Calculate the mass of sodium hydroxide needed to make 150 cm^3 of a 40 g dm^{-3} solution.