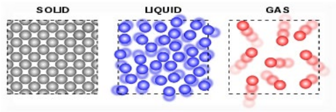
CC1/2 States of Matter & Separation Techniques

1. States of Matter



All substances are made up of particles. Solids (s), liquids (l) and gases (g) are made up of particles. Particle arrangement, movement and energy levels determine the state of matter.

State	Arrangement	Movement	Energy levels
Solid	Fixed, regular pattern, tightly packed	Vibrate about fixed positions	Least
Liquid	Irregular pattern, most touching	Slide over one another	-
Gas	Randomly	Freely, in all directions	Most

2. Changes of State

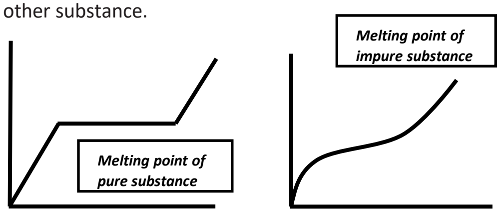
When particles gain energy, their movement and arrangement changes.

More energy causes more forces of attraction between particles to break. The amount of energy needed for a state change depends on the strength of the forces between particles.

Melting and freezing happen at melting point. Boiling and condensing happen at boiling point.

3. Pure vs Impure

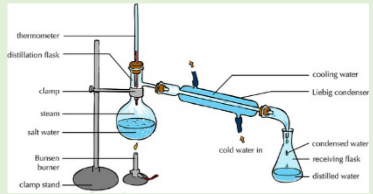
A pure substance is a single element or compound, not mixed with any other substance.



Pure substances melt and boil at specific temperatures. Heating graphs can be used to distinguish pure substance from impure substances.

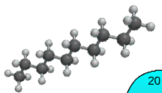
4. Simple Distillation

This technique is used to separate a mixture of liquids. During distillation, the mixture gets heated causing one liquid at a time to evaporate and then condense.



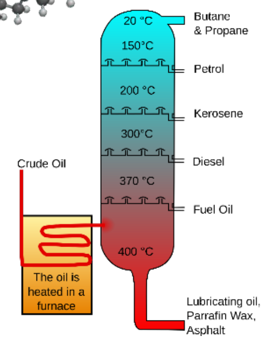
Each liquid has a different boiling point. This enables the liquids to be separated.

5. Fractional Distillation



Crude oil is a mixture of hydrocarbons. These hydrocarbons have different boiling points. Each fraction contains molecules with a similar number of carbon atoms in them. The process used to do this is called fractional distillation.

Crude oil is heated and hydrocarbons boil and condense at certain temperatures. Small hydrocarbon chains boil at low temperatures and long hydrocarbon chains boil at high temperatures.



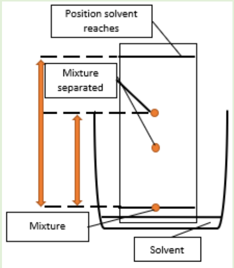
6. Chromatography

Chromatography is used to separate mixtures and help identify substances. The solvent (mobile phase) separates substances on chromatography paper (stationary phase) due to their solubility.

An Rf value can be calculated to show the ratio of the distance moved by a compound to the distance moved by the solvent:

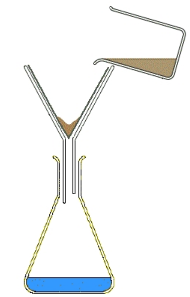
$$R_f = \frac{\text{distance moved by substance}}{\text{distance moved by solvent}}$$

Mixtures or impure substances produce multiple spots. Pure substances produce a single spot.



7. Filtration

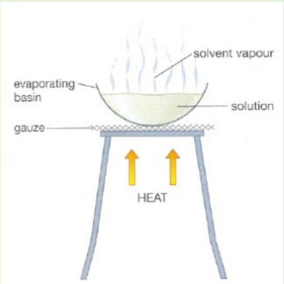
Filtration can be used to separate substances that are insoluble in a solvent from those that are soluble. An example is sand and water. Large sand particles collect in the filter paper (residue) and the water will pass through the filter paper (filtrate).



8. Crystallisation

Crystallisation can be used to separate a soluble substance from a solvent by evaporation.

The heat energy causes liquid particles to move further apart and become randomly arranged, moving freely in all directions.



An example of crystallisation is producing sodium chloride from a salt solution.

9. Potable Water

Human drinking water containing low levels of dissolved salts and microbes is safe to drink, clean and cook with. This water is known as potable water.

Most of the UK's water collects in lakes and rivers. Sterilising agents such as chlorine, ozone and UV can be used to treat water.

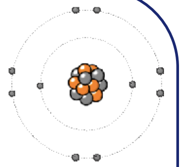
4 steps to produce potable water:

1. Select water source
2. Sedimentation
3. Filtration
4. Sterilisation/chlorination

In countries where fresh water is limited, desalination of sea water provides potable water. Distillation or reverse osmosis are used BUT require large amounts of energy!

CC3/4 Atomic structure and periodic table

1. Atoms



The smallest part of an element that can exist

Have a radius of around 0.1 nanometres and have no charge (0). The nucleus is very small compared to the overall size of the atom.

Name of Particle	Relative Charge	Relative Mass
Proton	+1	1
Neutron	0	1
Electron	-1	Very small
Atoms contain equal numbers of protons and electrons in order to have an overall neutral charge.		

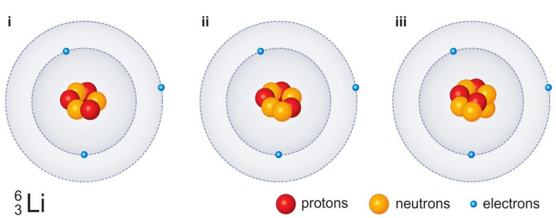
2. Reading the periodic table

For each element on the periodic table there are two numbers. The top number is the mass number, the bottom number is the atomic or proton number

7 Li 3			
	Mass number	*The sum of the protons and neutrons in the nucleus*	
	Atomic number	*The number of protons in the atom*	Number of electrons = number of protons
	Elements	*All atoms of a certain element had the same number of protons*	This number of protons is unique to that element.

3. Isotopes

Atoms of the same element with the same number of protons and different numbers of neutrons are called isotopes



A Isotopes of the same element are chemically identical because they have the same number of protons and electrons.

5. Calculating the average relative atomic mass of all isotopes

All elements exist as mixtures of isotopes. We use this idea to calculate an elements relative atomic mass (RAM or A_r).

A relative atomic mass is the mean mass of an atom of an element compared with carbon-12.

RAMs are not whole numbers 9e.g. Chlorine RAM is 35.5 we can calculate this using the abundances of each isotope. See below **^{35}Cl (75%) and ^{37}Cl (25%)**

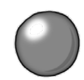
Relative abundance =

(% isotope 1 x mass isotope 1) + (% isotope 2 x mass isotope 2) ÷ 100

e.g. (25 x 37) + (75x 35) ÷ 100 = 35.5

4. History of the atom

Pre 1900



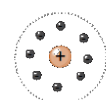
Tiny solid spheres that could not be divided

Before the discovery of the electron**, John Dalton** said the solid sphere made up the different elements.

1897 'plum pudding'

A ball of positive charge with negative electrons embedded in it

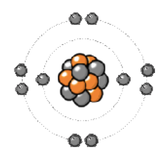
JJ Thompson 's experiments showed that showed that an atom must contain small negative charges (discovery of electrons).



1909 nuclear model

Positively charge nucleus at the centre surrounded negative electrons

Ernest Rutherford's alpha particle scattering experiment showed that the mass was concentrated at the centre of the atom.



1913 Bohr model

Electrons orbit the nucleus at specific distances

Niels Bohr proposed that electrons orbited in fixed shells; this was supported by experimental observations

6. Mendeleev's periodic table 1869

By 1869 there was 63 elements discovered.

Mendeleev a Russian chemistry arranged these elements into order. This was the first periodic table. His table:

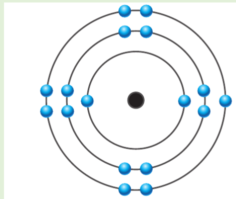
- Arranged elements in rows according to their chemical properties (i.e. lithium, sodium, potassium)
- The columns he arranged in mass number.
- Mendeleev used gaps in his table to make predictions about the properties f undiscovered elements.

One main difference to our modern periodic table is Mendeleev ordered by mass number where we now order by atomic number. The reason for this is in 1869 the proton had not been discovered yet. Therefore no atomic number

7. Electron configuration

In an atom electrons occupy electron Shells arranged around the nucleus.

The way in which an atoms electrons are arranged is called its electron configuration



C The electronic configuration of chlorine shows three occupied shells.

- The first shell can fit 2 electrons
- The second and third shells can contain up to eight electrons.
- You fill a shell before moving to the final shell.

Chlorine has 17 electrons (1st shell 2, second shell 8, third shell 7)
Or 2,8,7

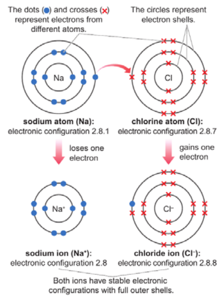
8. Todays periodic table structure

- Elements in a row or periodic are in order of increasing atomic number.
- Each row has the same number of electron shells
- Elements with similar properties are in the same column or group
- Each group has the same number of electrons on their outer shell
- Non metals are on the right of the table
- Metals on the left.

CC5-6 Ionic and covalent Bonding

1. Ionic bonds

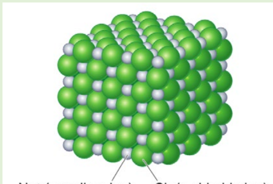
Atoms are more stable if they have an outer electron shell that is full. This can happen by atoms losing or gaining electron(s). When this happens the atoms form an ion.
Metal atoms **lose electrons, forming positive ions**.
Non metal atoms **gain electrons, forming negative ions**.
There are forces of attraction between positive and negative objects. These are called electrostatic forces and these are what hold ionic substances together.



C Dot and cross diagrams can be used to show what happens when ions are formed.

2. Ionic lattices

These strong ionic bonds allow 'billions' of ions to be packed together in a regular repeating arrangement called a lattice structure.



3. Forming ionic compounds

Ionic compounds are electrically neutral. The formula of an ionic compound contains the same number of positive and negative charges. To work out the ionic formulae we need to balance the positive ions and the negative ions.

Positive ion	Ion formula	Negative ion	Ion formula
sodium	Na^+	fluoride	F^-
lithium	Li^+	chloride	Cl^-
potassium	K^+	bromide	Br^-
magnesium	Mg^{2+}	oxide	O^{2-}
calcium	Ca^{2+}	sulfide	S^{2-}
aluminium	Al^{3+}	phosphide	P^{3-}

Polyatomic ion name	Ion formula
ammonium	NH_4^+
nitrate	NO_3^-
hydroxide	OH^-
carbonate	CO_3^{2-}
sulfate	SO_4^{2-}
sulfite	SO_3^{2-}

Magnesium oxide contain Mg^{2+} and O^{2-}ions. As there is the same number of positive charges to negative charges then they are balanced. Therefore, the formula is MgO.

Sodium sulfide contain Na^+ and S^{2-}. Therefore two Na^+ ions are needed to balance the S^{2-}. The formula is Na_2S

4. Properties of ionic compounds

High melting and boiling points
Most ionic substances are solid at room temperature. The electrostatic forces of attraction create strong bonds which need lots of energy to break.

Conducting electricity
Ionic substances do not conduct electricity when they are a solid. When they are a liquid (molten) or when they dissolve in water they do conduct electricity.

This is due to the delocalised electrons being free to carry a charge.

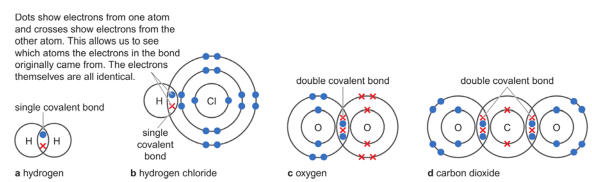
Ionic compound	Melting point (°C)	Boiling point (°C)
sodium bromide, NaBr	747	1390
sodium chloride, NaCl	801	1413
magnesium oxide, MgO	2852	3600

5. Covalent bonds

Molecular substances contain groups of atoms that are held together by strong bonds called **covalent bonds**. The number of atoms of each element bonded together in a simple **molecule** is shown by its **molecular formula.**

Covalent bonds are usually formed between non-metal atoms and are produced by sharing pairs of electrons. By forming the bond the atoms become more stable, because they can use the shared electrons to complete their **outer electron shells**. The reason why noble gases are so stable is because they have full outer electron shells.

The **dot and cross diagrams** in diagram B show how covalent bonds are formed. Counting the shared electrons, each atom now has a complete outer shell of electrons. Sometimes atoms share more than one pair of electrons to fill their outer shells. In oxygen and carbon dioxide the atoms share two pairs of electrons, to form **double bonds**.



B Dot and cross diagrams can be used to explain how covalent bonds are formed.

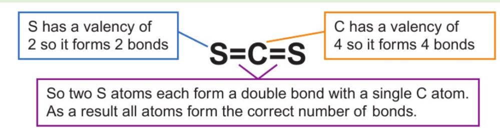
6. Working out molecular (covalent) formulae

Working out molecular formulae

The numbers of covalent bonds formed by atoms of different elements are shown in table D. This is called the **valency** of the element. It is the same as the number of electrons needed to obtain a complete outer shell.

Group number	Examples	Outer electrons	Bonds formed	Valency
4	C and Si	4	4	4
5	N and P	5	3	3
6	O and S	6	2	2
7	F and Cl	7	1	1

D valencies of some elements



E working out the formula of carbon sulfide

7. Properties of covalent compounds

Low melting and boiling points
Water is an example of a covalent compound. There are strong forces of attraction between the atoms. However, the intermolecular forces between the molecules are weak. These require little energy to break.

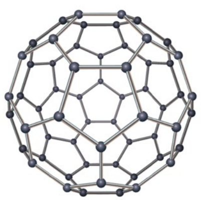
Do not conduct electricity
As there are no charged particles there is nothing to carry the electric current.

The electrons are shared between the atoms so are not delocalised (free to move) and cannot carry a charge.

CC7 Bonding models

2. Allotropes of carbon

The element carbon can form a number of different molecules. Different forms of the same element are called allotropes. The structure and bonding will influence the properties and uses.



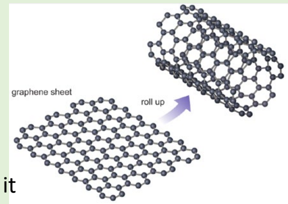
A Buckminsterfullerene is a simple molecule.

Fullerenes

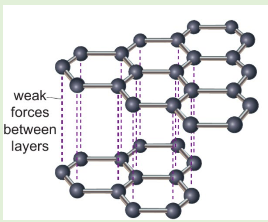
- A simple molecule
- 3 bonds between each carbon
- Either as a tube (nanotubes) or buckminsterfullerene (C_{60}) bucky ball
- Low melting point

Graphene

- A giant covalent molecule
- 3 bonds between each carbon
- A sheet is one atom thick therefore the lightest known material.
- As it only has 3 covalent bonds it conducts electricity



B Graphene can be a sheet or can be rolled into a tube.

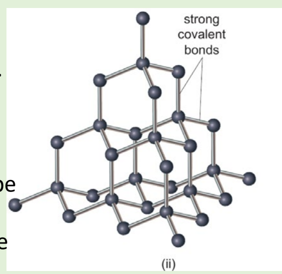


Graphite

- A giant covalent molecule
- Layers of graphene stacked on to of each other.
- Weak forces between layers means it can be used as a carbon based lubricant 3 bonds between each carbon
- Like graphene it can conduct electricity

Dimond

- A giant covalent molecule
- Each carbon atom has 4 bonds. Means it has no free charged particles therefore cannot conduct electricity.
- The tetrahedral (pyramid) shape means it can be shaped to be very sharp and strong so can be used in cutting tools).



1. Different models of a molecular substance

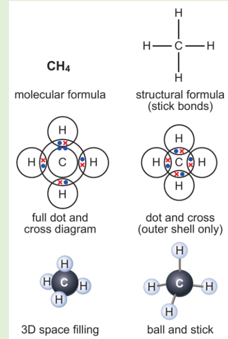
Molecular formula
Shows the number of each element. Dose not show the bonds

Structural formula
Shows how many bonds between each atom. Does not show electron sharing

Dot and cross diagram
Shows electron sharing

3D space filling
Shows relative size of each atom, does not show bonds

Ball and stick
Shows relative size of each atom and bonds



C different representations of methane

3. Summary of different types of bonds

Ionic
Where found: in most compounds containing metal and non-metal atoms.
Bonding: ionic bonds formed by the loss and gain of electrons to produce oppositely charged ions that attract one another.
Structure: billions of ions held together in a lattice structure.
Properties:
- high melting/boiling points
- many are soluble in water
- conduct electricity when liquid or in solution but do not when solid.

Simple molecular (covalent)
Where found: in most non-metal elements and compounds.
Bonding: covalent bonds formed when atoms share pairs of electrons.
Structure: small, distinct groups of atoms.
Properties:
- low melting/boiling points
- a few are soluble in water
- most do not conduct electricity.

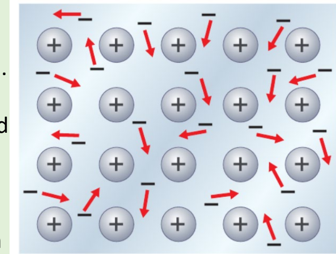
Giant covalent
Where found: in a few non-metal elements and some compounds of non-metals.
Bonding: covalent bonds formed when atoms share pairs of electrons.
Structure: billions of atoms held together in a lattice structure.
Properties:
- high melting/boiling points
- insoluble in water
- most do not conduct electricity (except in carbon as graphite).

Metallic
Where found: in all metals.
Bonding: metallic bonds are the electrostatic attraction between positive metal ions and negative delocalised electrons.
Structure: billions of ions held together in a giant lattice structure of positive ions in a 'sea' of negative delocalised electrons.
Properties:
- high melting/boiling points
- insoluble in water
- conduct electricity when solid or liquid.

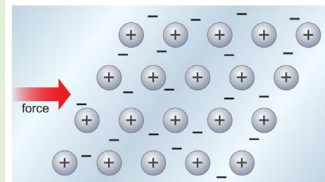
4. Properties of metals

Metallic structure and bonding
The atoms in a metallic element are all the same size and are packed closely together.

Metal atoms have 1,2,3, electrons on the outer shell. These outer shell electrons are lost from each atom and become free to move randomly throughout the metal. This gives a sea of delocalised electrons which move in random directions



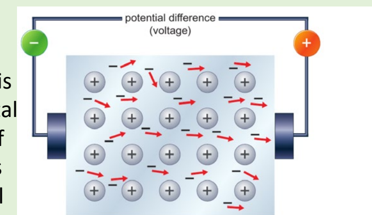
B Metals consist of stacked layers of ions in a 'sea' of delocalised ('free') electrons.



C When hit or bent, the layers of ions in a metal can slide over each other.

Metals are malleable
They can be hammered or rolled into shape without shattering. Layers slide over each other. The sea of electrons hold the ions together.

Metals are good conductors of electricity
They when a charge is passed though a metal this causes the sea of delocalised electrons to carry the electrical flow through the metal.



E When a voltage is applied to a piece of metal, an electrical current flows.

The electrical conductivity of the metals increase as the number of delocalised electrons increase. Each sodium ion has one positive charge. And contributes 1 electron to the 'sea' of delocalised electrons. As Magnesium loses two electrons it has a higher electrical conductivity than sodium.