

Revision Checklist : 4.2 Bonding and Structure

States of Matter

The three states of matter are solid, liquid and gas.

Melting and freezing between solid and liquid take place at the **melting point**.
Boiling and condensing between liquid and gas take place at the **boiling point**

The amount of energy needed to change state from solid to liquid and from liquid to gas depends on the strength of the forces between the particles of the substance. The nature of the particles involved depends on the type of bonding and the structure of the substance. The **stronger** the forces between the particles the **higher the melting point** and boiling point of the substance.

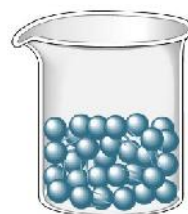
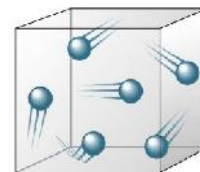
In chemical equations, the three states of matter are shown as (s), (l) and (g), with (aq) for aqueous solutions.

Pure elements and compounds melt and boil at specific temperatures.

Melting point and boiling point data can be used to distinguish pure substances from mixtures.

- A pure substance will melt or boil at a fixed temperature.
- A mixture will melt over a range of temperatures and not a sharp melting point.

Gas **particles** are widely spaced and in constant quick motion. Collisions are frequent and elastic. Weak forces between particles.



Liquids **particles** are closely spaced but still in constant motion, and therefore are constantly colliding. Forces between particles less than in solid.

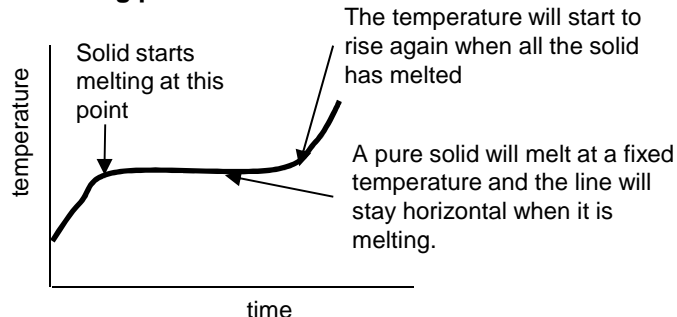
Solid particles can only **vibrate** in a **fixed position**



The three states of matter can be represented by a simple model. In this model, particles are represented by small solid spheres. Particle theory can help to explain melting, boiling, freezing and condensing.

Limitations of the simple model include that there are **no forces between the spheres**, and that atoms, molecules and ions are **solid spheres**. These are not true.

Melting point curve



The temperature does not rise when the solid is melting because the heat is absorbed to break the bonds between the solid particles

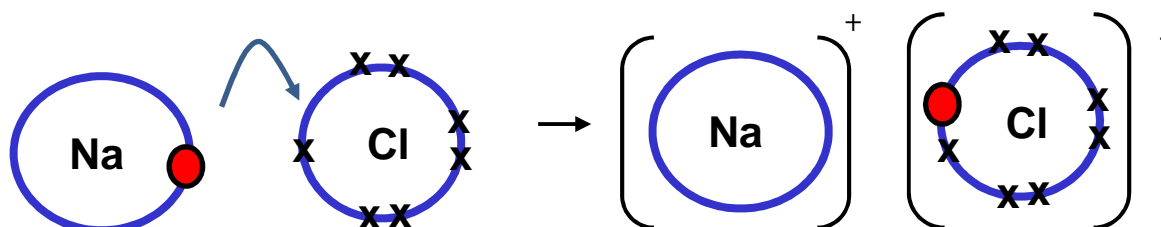
Ionic bonding

Ionic bonding occurs in compounds formed from metals combined with non-metals.

When a **metal atom** reacts with a **non-metal atom** **electrons** in the outer shell of the metal atom are **transferred**.

Metal atoms **lose** electrons to become positively charged ions. Non-metal atoms **gain** electrons to become negatively charged ions.

The electron transfer during the formation of an ionic compound can be represented by a dot and cross diagram e.g. for sodium chloride.



Just showing outer shell electrons

In sodium chloride NaCl, the sodium atom **loses one** electron forming a +1 sodium ion. The chlorine **gains the one** electron to become a -1 chloride ion. The ions formed have full shells of electrons. The ionic bond is the force of attraction between the oppositely charged ions.

The ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 have the electronic structure of a noble gas (Group 0).

Magnesium ion $\text{Mg}^{2+} = [2,8]^{2+}$

Calcium ion $\text{Ca}^{2+} = [2,8,8]^{2+}$

Chloride ion $\text{Cl}^- = [2,8,8]^-$

Oxide ion $\text{O}^{2-} = [2,8]^{2-}$

Describing electron transfer in other compounds

Be able to describe electron transfer for other substances.

For example in potassium oxide K_2O .

The potassium atom **loses one** electron forming a +1 **potassium ion**. The oxygen atom **gains two** electrons to become a -2 **oxide ion**. **Two potassium atoms react with one oxygen atom**. The ions formed have full shells of electrons.

Working out Formulae for Ionic Compounds

Common charges of ions

The charge on the ions produced by metals in Groups 1 and 2 and by non-metals in Groups 6 and 7 relates to the group number of the element in the periodic table

Group 1 always form +1 ions e.g. Na^+

Group 2 always form +2 ions e.g. Mg^{2+}

Group 6 always form -2 ions e.g. O^{2-}

Group 7 always form -1 ions e.g. F^-

Using the ions in group 1,2,6,7 you should be able to work out the formulae of compounds. You may also be given other ions to use.

To do this ions are combined together so that the total positive charge of the ions is the same as the total negative charge of the ions – giving a neutral compound.

For sodium chloride the ions are Na^+ and Cl^- . Here the +1 charge cancels the -1 charge so the formula is **NaCl**

For magnesium fluoride the ions are Mg^{2+} and F^- . Here we need two F ions to cancel out the +2 charge on the Mg, so the formula is **MgF₂**

For lithium oxide the ions are Li^+ and O^{2-} . Here we need two Li^+ ions to cancel out the -2 charge on the O, so the formula is **Li₂O**

Some common groups of elements have a charge. These are called **compound ions**. (These will be given in the exam)

Carbonates CO_3^{2-} Sulfates SO_4^{2-} have a -2 charge.

Hydroxides OH^- Nitrates NO_3^- have a -1 charge

Ammonium NH_4^+ has a +1 charge

When to use brackets in a formula

If there are two or more of a compound ion in a formula we put the ion in brackets and put the number outside the bracket

e.g. Copper nitrate is $\text{Cu}(\text{NO}_3)_2$

Calcium hydroxide is $\text{Ca}(\text{OH})_2$

Ammonium sulfate is $(\text{NH}_4)_2\text{SO}_4$

For Magnesium nitrate the ions are Mg^{2+} and NO_3^- . Here we need two NO_3^- ions to cancel out the +2 charge on the Mg, so the formula is **$\text{Mg}(\text{NO}_3)_2$** . Note: we have to use brackets as the nitrate ion is a compound ion. When more than one compound ion is used brackets must be used.

Giant Ionic Structure

An ionic compound is a giant structure of ions. Ionic compounds are held together by **strong electrostatic forces of attraction** between **oppositely charged ions**. These forces act in all directions in the lattice and this is called ionic bonding.

Properties of Ionic Substances

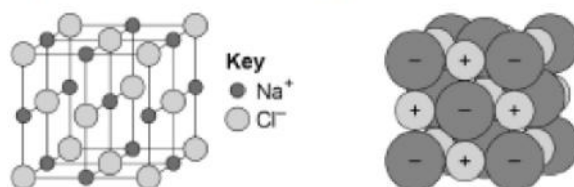
These compounds have **high melting points** and **high boiling points** because of the **large amounts of energy** needed to break the **many strong electrostatic forces of attraction** between **oppositely charged ions**.

When in **solid form** ionic compounds **do not conduct electricity** because the **ions** are fixed in place

When **melted** or **dissolved in water**, ionic compounds **conduct electricity** because the **ions** are **free to move** and so **charge can flow**.

CAUTION! When describing ionic structures and their properties always use the word **IONS**. **Never use** the terms delocalised electrons, molecules, intermolecular forces

The structure of sodium chloride can be represented in the following forms:



The ball and stick model shows how the ions are arranged relative to the neighbours but there are not in reality sticks or gaps between ions

This model shows relative sizes of ions and the regular pattern but only shows outer layers

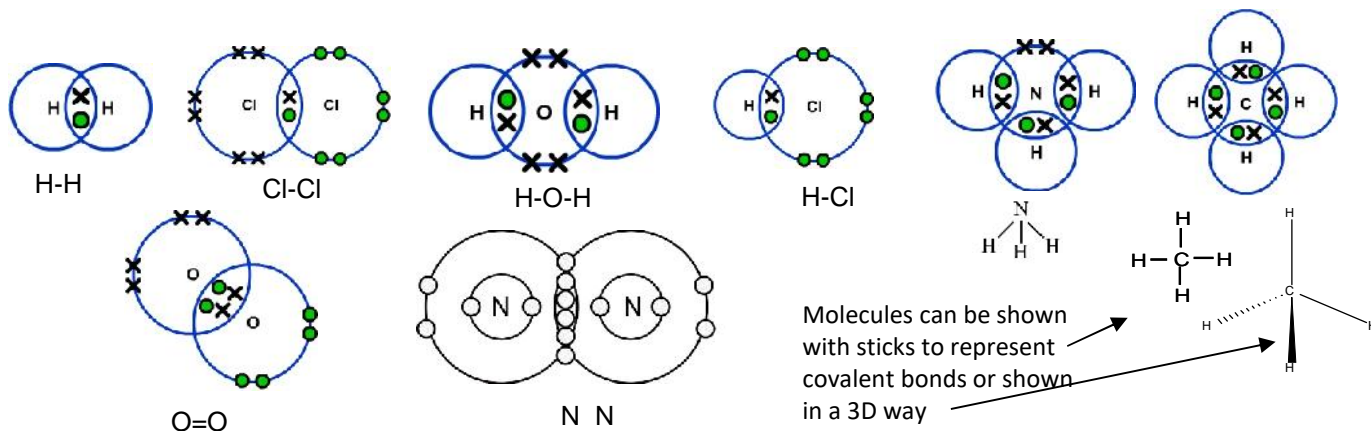
Empirical formula. The empirical formula is the **simplest ratio** of ions in a compound. In the diagrams above we can see there are equal numbers of sodium and chloride ions so the empirical formula is NaCl (1:1 ratio)

Covalent Bonding

Covalent bonding occurs in non-metallic elements and in compounds of non-metals.

When atoms **share pairs** of electrons, they form covalent bonds. Covalent bonds between atoms are **strong**.

Covalently bonded substances may consist of small molecules, such as H_2 , Cl_2 , O_2 , N_2 , HCl , H_2O , NH_3 and CH_4 .



Limitations of models

The different models have different limitations:

- none are to scale
- some do not show the 3 dimensional arrangement in space
- some do not show the electrons or all the shells

Structure and properties of simple molecular covalent substances

Substances that consist of **small molecules** are usually gases or liquids that have relatively **low** melting points and boiling points.

These substances do not conduct electricity because the molecules do not have an overall electric charge.

These substances have only **weak forces** between the molecules (**intermolecular forces**). It is these **intermolecular forces that are overcome, not the covalent bonds**, when the substance melts or boils. These require little energy to overcome.

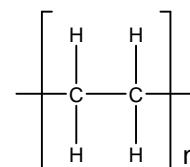
Effect of size of molecule on intermolecular force

Intermolecular forces increase with the size of the molecules. So larger molecules need more energy to break the intermolecular forces and have higher melting and boiling points. This explains why the boiling points of the elements in group 7 increase down the group.

You should **only** use the word **intermolecular forces** with this type of bonding

Polymers

Polymers have very large molecules. The atoms in the polymer molecules are linked to other atoms by strong covalent bonds. The intermolecular forces between polymer molecules are relatively strong and so these substances are solids at room temperature.



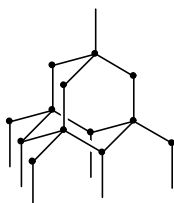
Polymers can be represented in the form above, where n is a large number.

Structure and Properties of Giant Covalent Substances

Substances that consist of **giant covalent** structures are solids with **very high melting points**. All of the atoms in these structures are linked to other **atoms by strong covalent bonds**. These bonds must be overcome to melt or boil these substances.

Diamond and graphite (forms of carbon) and silicon dioxide (silica) are examples of giant covalent structures.

In **diamond**, each carbon atom forms **four covalent bonds** with other carbon atoms in a giant covalent structure, so diamond is **very hard**, has a **very high melting point** and does **not conduct electricity**.



Graphene is a **single layer of graphite** and so is one atom thick.

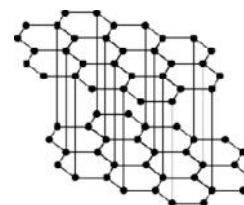
Fullerenes are molecules of carbon atoms with hollow shapes. The structure of fullerenes is based on hexagonal rings of carbon atoms but they may also contain rings with five or seven carbon atoms. The first fullerene to be discovered was Buckminsterfullerene (C_{60}) which has a spherical shape.

Carbon nanotubes are cylindrical fullerenes with very high length to diameter ratios. Their properties make them useful for nanotechnology, electronics and materials.

In graphite, each carbon atom forms **three covalent bonds** with three other carbon atoms, forming **layers of hexagonal rings** and so graphite has a high melting point.

The layers are **free to slide** over each other because there are **no covalent bonds between the layers** and so graphite is soft and slippery.

In graphite, **one electron** from each **carbon atom** is **delocalised**. These delocalised electrons allow graphite to **conduct thermal energy and electricity** through the structure.



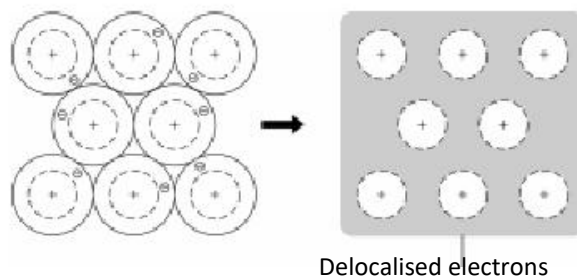
Graphite is similar to metals in that it has delocalised electrons.

Metallic Bonding

Metals consist of giant structures of atoms arranged in a regular pattern. The **electrons** in the **outer shell** of metal atoms are **delocalised** and so are **free to move** through the **whole structure**. The sharing of delocalised electrons gives rise to strong metallic bonds.

This corresponds to a structure of **positive ions** with delocalised electrons between the ions holding them together by **strong electrostatic** attractions.

The bonding in metals may be represented in the following form:



Delocalised electrons

Properties of metals

Strong metallic bonding means that most metals have high melting and boiling points, because lots of energy is needed to break the strong metallic bonds between the positive ions and delocalised electrons.

Metals are good conductors of electricity because the **delocalised electrons** in the metal **carry electrical charge** through the metal. Metals are **good conductors** of thermal energy because thermal **energy is transferred by the delocalised electrons** through the structure

Pure metals and alloys

In pure metals, the atoms are all the same size. The **layers of atoms** are able to **slide over** each other. This means metals can be bent and shaped.

Most metals in everyday use are alloys. **Pure** copper, gold, iron and aluminium are **too soft** for many uses and so are mixed with other metals to make alloys.

The **different sizes of atoms** in an alloy **distort** the **layers** in the structure, making it more **difficult** for them to slide over each other, so **alloys** are **harder** than **pure metals**

Describing electrical conductivity:

Use the term **delocalised electrons** to describe the electrical conductivity in **metals and graphite**.

Use the term 'ions that are free to move' for the conductivity of molten ionic substances

Summary of Properties of Different Structures

Property	Small molecules	Ionic	Giant Covalent	Metallic
boiling and melting points	low- because of weak intermolecular forces between small molecules . These require little energy to overcome.	high- because of giant lattice of ions with strong forces between oppositely charged ions . Takes a lot of energy to break	high- because of many strong covalent bonds in between atoms in giant structure. Takes a lot of energy to break the many strong bonds	high- strong electrostatic forces between positive ions and delocalised electrons . Takes a lot of energy to break
conductivity when solid	poor: no ions or charged particles to conduct	poor: ions can't move	diamond and sand: poor, because electrons can't move graphite : good as free delocalised electrons between layers can move through structure	good: delocalised electrons are free to move through structure
conductivity when molten	poor: no ions	good: ions are free to move	poor	(good)
general description	mostly gases and liquids	crystalline solids	solids	shiny metal solids