Revision Checklist : 4.2 Bonding and Structure Trilogy

**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.

**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.

**Gas particles** are widely spaced and in constant motion. Collisions are frequent and elastic.

**Liquids particles** are closely spaced but still in constant motion, and therefore are constantly colliding.

**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.

**Melting point curve**

- Solid starts melting at this point.
- The temperature will start to rise again when all the solid has melted.
- A pure solid will melt at a fixed temperature and the line will stay horizontal when it is melting.
- 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.

\[ \text{Na}^{+} \text{Cl}^{-} \]

Just showing outer shell electrons

In sodium chloride, the sodium atom gives one electron away forming a +1 ion. The chlorine gains the electron to become a -1 ion. 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).

Mg$^{2+} = [2,8]^{2+}$
Ca$^{2+} = [2,8,8]^{2+}$
Cl$^{-} = [2,8,8]$^{-}
O$^{2-} = [2,8]^{2-}$

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$^{-}$

Working out Formulae for Ionic Compounds
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.

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.

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 MgF$_2$.

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 Li$_2$O.
When to use brackets in a formula

If you need to have 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 Cu(NO$_3$)$_2$

Calcium hydroxide is Ca(OH)$_2$

Ammonium sulphate is (NH$_4$)$_2$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 Mg(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.

Some common groups of elements have a charge. These are called compound ions:

- Carbonates CO$_3^{2-}$
- Sulphates SO$_4^{2-}$
- Hydroxides OH$^-$
- Nitrates NO$_3^-$
- Ammonium NH$_4^+$

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 H₂, Cl₂, O₂, N₂, HCl, H₂O, NH₃, and CH₄.

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 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 substances do not conduct electricity because the molecules do not have an overall electric charge. The intermolecular forces increase with the size of the molecules, so larger molecules have higher melting and boiling points.

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.
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.

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.

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.

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 energy is transferred by the delocalised electrons.

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.

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

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.

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

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

Delocalised electrons

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.
### Summary of Properties of Different Structures

<table>
<thead>
<tr>
<th>Property</th>
<th>Simple Covalent</th>
<th>Ionic</th>
<th>Giant Covalent</th>
<th>Metallic</th>
</tr>
</thead>
<tbody>
<tr>
<td>boiling and melting points</td>
<td>low- because of weak intermolecular forces between molecules</td>
<td>high- because of giant lattice of ions with strong forces between oppositely charged ions.</td>
<td>high- because of many strong covalent bonds between atoms in giant structure</td>
<td>high- strong electrostatic forces between positive ions and delocalised electrons</td>
</tr>
<tr>
<td>conductivity when solid</td>
<td>poor: no ions to conduct</td>
<td>poor: ions can’t move</td>
<td>diamond and sand: poor, because electrons can’t move</td>
<td>good: delocalised electrons are free to move through structure</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>graphite: good as free delocalised electrons between layers can move through structure</td>
<td></td>
</tr>
<tr>
<td>conductivity when molten</td>
<td>poor: no ions</td>
<td>good: ions are free to move</td>
<td>poor</td>
<td>(good)</td>
</tr>
<tr>
<td>general description</td>
<td>mostly gases and liquids</td>
<td>crystalline solids</td>
<td>solids</td>
<td>shiny metal solids</td>
</tr>
</tbody>
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