Lesson Four
Structures and Properties

Aims

By the end of this lesson you should be able to:

- understand why compounds with giant ionic lattices have high melting and boiling points
- know that ionic compounds do not conduct electricity when solid, but do so when molten and in aqueous solution explain why substances with simple molecular structures are gases or liquids, or solids with low melting and boiling points
- explain why the melting and boiling points of substances with simple molecular structures increase with increasing molecular mass
- explain why substances with giant covalent structures are solids with high melting and boiling points
- explain how the structures of diamond, graphite and C60 fullerene influence their physical properties, including electrical conductivity and hardness
- know that covalent compounds do not usually conduct electricity
- **understand metallic bonding in terms of electrostatic attractions, and know how to represent a metallic lattice by a 2-D diagram**
- **explain the typical physical properties of metals, including electrical conductivity and malleability**

Context

This lesson covers Sections 1.42 – 1.43 and 1.47 – 1.54C of the Edexcel IGCSE Chemistry specification.

Edexcel International GCSE (9-1) Chemistry Student Book, pages 81-83, 92-96 and 98-100.

Introduction

In Lesson Three we saw how atoms react together to form two different sorts of bond: ionic bonds and covalent bonds.

In this lesson we look at the structures of the substances formed as a result of bonding, and the effect this has on the physical properties of these substances. Physical properties include:

- melting point and boiling point
- electrical conductivity
- hardness / softness
- malleability: the ability to bend without breaking

Bonding and physical properties

Melting and boiling points

You will remember from Lesson One that the particles of a solid are held together in fixed positions, while the particles of a liquid have broken free and are able to move around.

Now, if the attractive forces holding the particles of a solid together are *strong*, then you must give them a lot of energy before they can break free and start to move around. This means the substance will have a *high* melting point: it will turn from solid to liquid at a high temperature.

But if the attractive forces holding the particles of a solid together are *weak*, then they only need a little energy before they will break free and start to move around. This means the substance will have a *low* melting point: it will turn from solid to liquid at a low temperature.

Much the same also applies to boiling points. When a liquid boils, its particles must overcome the attractive forces holding them close together so that they can become widely separated. This also needs an input of energy.

So:

- strong forces --> high melting and boiling points

- weak forces --> low melting and boiling points

Electrical conductivity

A current of electricity is *a flow of charged particles*. These can be free electrons or ions. It follows that a substance will conduct electricity if and only if it has either electrons or ions *that are free to move* throughout the substance rather than being tied to one place.

Giant ionic lattices

Ionic substances, like sodium chloride and lithium fluoride, have particles which are positive and negative ions. When solid, these ions are held together by electrostatic attractive forces (ionic bonds) between the oppositely-charged ions into a regular structure called a **giant ionic lattice**. See Figure 7.18 on page 82 of the textbook. These attractive forces are strong.

Two things follow:

- Because all of the particles are bound together into one giant structure by strong attractive forces, these substances have high melting points and boiling points. For example, the melting point of sodium chloride (table salt) is 801^{0}C, hotter than can be achieved using a Bunsen burner (about 700^{0}C) or a gas cooker. Its boiling point is even higher: 1413^{0}C.

- Because the ions are in fixed positions, ionic solids do not conduct electricity. But when ionic substances are heated and become liquid, or when they are dissolved in water, the separated ions are free to move to carry an electric current. So ionic substances do conduct electricity when molten or in solution.

Twig

Log on to Twig and look at the film titled: **Crystal**

www.ool.co.uk/455yp

A solid consisting of atoms, molecules, or ions arranged in a highly ordered, repetitive, three-dimensional lattice.

Simple molecular structures

Substances whose atoms are joined together by covalent bonds to form molecules are called **simple molecular**

structures. These are substances like water, oxygen and candle wax.

In this case, the covalent bonds holding the atoms together *inside* the molecule (***intra*molecular** forces) are also strong. But the forces *between* the molecules, holding one molecule to another (***inter*molecular** forces) are weak. Only these weak intermolecular forces need to be overcome to melt the solid, or boil the liquid, so these substances have low melting and boiling points. These substances tend to be gases or liquids at room temperature, or are solids with low melting points like candle wax.

simple molecular structure --> weak intermolecular forces --> low melting and boiling point

The larger the molecules, the more surface area to form (weak) bonds with neighbouring molecules, and the more energy needed to make the molecules break free. So the melting and boiling points of simple covalent substances increase as the molecular mass increases. Thus candle wax (large molecules) is a solid at room temperature, whereas methane (small molecules, but with a similar content in terms of elements) is a gas.

Simple molecular substances do not conduct electricity, even when liquid. Their molecules are uncharged, so there are no free charged particles to carry the current.

Giant covalent structures

A few substances formed by covalent bonds do not form simple molecules. Important examples are diamond and graphite, both forms of the element **carbon**.

Melting point

In **diamond** all of the atoms in the crystal are joined to each other by strong covalent bonds. This forms, in effect, one giant molecule. Figure 8.26 on page 93 of the textbook shows how it works. This means that a large number of strong bonds must be broken for the atoms to break free and the crystal to melt, so diamond has a very high melting point indeed (about 3550^0C).

A lump of **graphite** is formed from many flat sheets of carbon atoms, but within the sheets each atom is bound to the others by strong covalent bonds (see Figure 8.30 on page 94 of the textbook). Once again many strong covalent bonds must be broken for the substance to melt, so graphite also has a very high melting point (still over 3000^{0}C).

Diamond and graphite are examples of **giant covalent structures**. Giant covalent structures all have high melting points.

Activity 1

Here are four substances and their melting points in ^{0}C:

A: 36 B: -125 C: 1263 D: 590

(a) Arrange these in order of their melting points with the lowest first

(b) Which will be solids at room temperature?

(c) Which is/are most likely to be:

(i) giant covalent structures?

(ii) simple molecular structures?

(iii) ionic substances?

Log on to Twig and look at the film titled: **Intro to Carbon**

www.ool.co.uk/1358dx

Carbon occurs naturally in four different forms. Discover how they differ and why?

Diamond

Each carbon atom has the electronic structure 2,4, so it needs to gain four extra electrons to fill its outside shell and become stable. In diamond, each atom does this by forming *four* bonds with neighbouring carbon atoms.

The four neighbouring carbon atoms are spaced in a **tetrahedral** arrangement in three dimensions, as shown below. The structure goes on and on in all directions beyond those atoms shown:

Because all of the atoms are held in fixed positions by strong covalent bonds, diamond has a very high melting point (see above) and is also very *hard*. In fact, it is so hard that it is able to cut or scratch most other substances. It is therefore used in glass-cutters and drill bits for cutting glass and rock respectively.

Get it right! The terms "hard" and "soft" refer to how difficult it is to scratch or cut a substance. "Brittle" means a substance will crack or break if hit. Diamond is extremely hard, but quite brittle: if you hit it with a hammer its crystals will break.

Log on to Twig and look at the film titled: **Synthetic Diamonds**

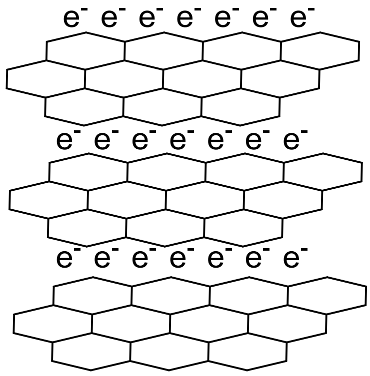
www.ool.co.uk/1359bw

Carbon occurs naturally in four different forms. Discover how they differ and why.

Graphite

Graphite is the black material used (mixed with clay) in normal (non-coloured) pencils. Unlike diamond, it is soft and black. But it too is made of just carbon atoms.

In graphite, each carbon atom forms a covalent bond with just *three* neighbouring carbon atoms. This produces flat sheets of atoms, as shown below and in Figure 8.30 on page 94 of the textbook. Notice the six-sided rings formed, like the cells of a honeycomb.



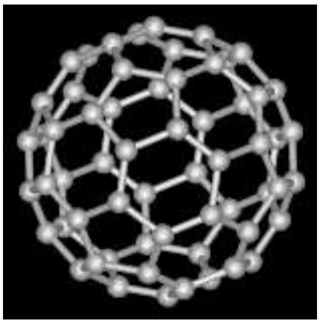
Although the atoms *within* a sheet are very tightly held together, the forces *between* neighbouring sheets of atoms are weak. The sheets can therefore side over each other, which makes graphite slippery to feel, a bit like soap. It is therefore sometimes used as a **lubricant**. It melts at a very high temperature, so it is useful as a lubricant under conditions where oil would boil away.

Because each carbon atom only forms three covalent bonds, it has a spare electron in its outside shell. These electrons are called **delocalised electrons** – they are able to wander around throughout the graphite sheet instead of staying attached to their atoms. This means that, unlike diamond, graphite can conduct electricity.

Activity 2	Put the search term "carbon allotropes" into the search box at www.youtube.com to see some videos on the different forms ("allotropes") of carbon.

C_{60} fullerene

Both diamond and graphite occur naturally. In 1985, a new form of carbon was made artificially by the action of a laser beam on graphite. Called **C_{60} fullerene** (or a variety of other names) each molecule consists of 60 carbon atoms covalently bonded to form a hollow sphere. See below and figures 8.32 and 8.33 on page 95 of the textbook.



Many other similar forms of carbon have since been made artificially, and some have been discovered in nature, for example in soot.

C_{60} fullerene has a high melting point, like graphite and diamond, but otherwise has unusual properties:

- it is usually soft, but becomes hard like diamond when compressed
- other materials can become trapped in the balls, and this can change its properties: if some metals are trapped it becomes a superconductor (having no electrical resistance) at very low temperatures
- unlike the other forms of carbon, it will dissolve in organic solvents like petrol.

Activity 3

Optional activity

Chemists are very excited about possible future uses of C_{60} fullerene and similar forms of carbon. If you are interested, you can find out more about them at www.ool.co.uk/0410ci.

Metallic bonding

Metals, you will remember from Lesson Three, are elements found at the left of the periodic table. They have one, two or sometimes three electrons in their outside shells, and are able to lose these easily to form ionic bonds.

In a piece of metal, these outer-shell electrons are **delocalised** as they are in graphite. They are able to wander around inside the metal, which is the reason that metals too *conduct electricity*. See Figure 9.5 on page 99 of the textbook.

The delocalised electrons also act like a "sea" of negatively charged "glue" sticking a giant structure of positively charged metal ions together by electrostatic attractive forces. See Figure 9.3 and 9.4 on pages 98 and 99 of the textbook. This is known as **metallic bonding**. The strength of the attractive forces varies from metal to metal, determining a metal's melting point: strong forces gives high melting point.

The positive ions are not held at particular angles to each other, and can slide over each other (unlike the carbon atoms in diamond), so their orientation can be changed. See Figure 9.6 on page 100 of the textbook. As a result a metal is **malleable**: it can bend without breaking.

Different metals can mix together to form **alloys**, for example brass which is an alloy of copper and zinc. Alloys are mixtures, not compounds, because:

- the atoms of the different metals are not joined together by ionic or covalent bonds, and
- the metals can be mixed together in variable ratios

Log on to Twig and look at the film titled: **Metallic Bonding**

www.ool.co.uk/1357za

Metals can be strong and hard, and they can be flexible.

Edexcel International GCSE (9-1) Chemistry Student Book, pages 81-83, 92-96 and 98-100.

Keywords

ionic substance
giant covalent structure
lattice
delocalised electron
metallic bonding
simple molecular substance
intermolecular
intramolecular
tetrahedral
conduct
malleable
alloy

Summary

Lesson Four: Structure and properties

Bonding and physical properties

Giant ionic lattices

Simple molecular structures

Giant covalent structures

Metallic bonding

What you need to know

- The meanings of the words in **bold** in this lesson
- That ionic compounds have high melting and boiling points, and the reason for this
- That simple molecular structures have low melting and boiling points, and the reason for this
- that giant molecular structures have high melting points, and the reason for this

What you might be asked to do

- Describe the arrangements of the ions in sodium chloride, and the carbon atoms in diamond, graphite and C_{60} fullerene
- Suggest the type of bonding in a substance given its properties
- Explain how the bonding and structure of a substance produces its properties
- Deduce the type of substance from data about its melting point and electrical conductivity

Self-Assessment Test

Explain the following:

(a) Solid sodium chloride (salt) crystals do not conduct electricity, but sodium chloride solution does.

(b) The molecules of methane (CH_4) and water (H_2O) are about the same size and mass, but water is a liquid while methane is a gas.

(c) The melting point of calcium oxide is much higher than the melting point of sulphur dioxide.

(d) Solid magnesium will conduct electricity but solid carbon dioxide will not.

Suggested Answers to Activities

Activity 1

(a) B, A, D, C
(b) A, C and D
(c) (i) C
(ii) A and B
(iii) D

Suggested Answers to Self-Assessment Test: Lesson Four

(a) Sodium chloride is an ionic solid made of charged particles (ions). In a solution these can move about carrying an electric current. But in a solid they are in fixed positions and cannot.

(b) The intermolecular forces between the water molecules must be greater than those between the methane molecules, making its boiling point higher.

(c) Calcium oxide is an ionic solid. Its ions are all held together by strong ionic bonds. But sulphur dioxide is a simple molecular substance. The intermolecular forces holding its molecules to each other are weak.

(d) Magnesium is a metal. The two electrons in its outer shell are delocalised and can move about carrying an electric current. Carbon dioxide is a simple molecular substance. Its molecules are neutral, so there are no free charged particles at all.