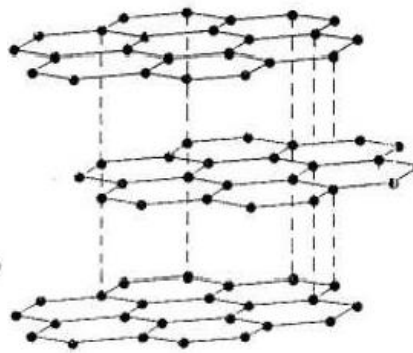
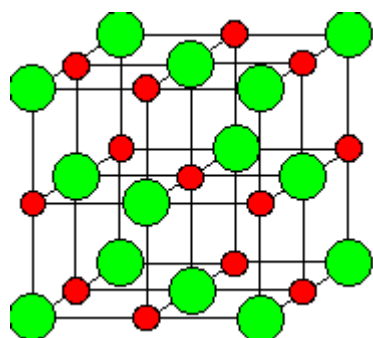


UNIT 2

PARTICLES, BONDING AND STRUCTURE

PART 3 – CHEMICAL STRUCTURES



Contents

1. Bonding and Structure
2. Giant Ionic Structures
3. Giant Metallic Structures
4. Simple Molecular Structures
5. Giant Covalent Structures
6. Summary of Chemical Structures
7. Structure and Bonding in the Periodic Table
8. Elements, Compounds and Mixtures

Key words: atom, molecule, ion, ionic bond, metallic bond, covalent bond, intermolecular forces, Van der Waal's force, hydrogen bond, lattice, element, compound, mixture, law of constant composition, law of multiple proportions

Units which must be completed before this unit can be attempted:

Unit 1 – Atomic Structure and the Periodic Table

1) Bonding and Structure

All matter is made of up of particles. These particles can be:

- **Atoms**
- **Molecules** (small groups of atoms held together by covalent bonds)
- **Ions** (atoms or molecules with a charge)

The way in which individual atoms, ions and molecules are attracted to each other is called **bonding**.

Oppositely charged ions are held together by **ionic** bonding.

Some electropositive atoms are held together by **metallic** bonding.

Most other atoms are held together by **covalent** bonding.

Molecules are held together by **intermolecular** bonding (either **Van Der Waal's forces** or **hydrogen bonds**).

The way in which atoms, molecules or ions arrange themselves on a larger scale is called **structure**.

Most structures in the solid state consist of a regular, repeating pattern, usually in three dimensions. This regular, repeating pattern is called a **lattice**.

The physical and chemical properties of a material depend on:

- The type of particles they contain (atoms, molecules or ions)
- The type of bonding between the particles (ionic, covalent, metallic or intermolecular)
- The type of **structure** which the material adopts

There are four main types of structure:

- (a) **Giant ionic lattice structure** – this is the structure adopted by oppositely charged ions held together by ionic bonds
- (b) **Giant metallic lattice structure** – this is the structure adopted by atoms when they are held together by metallic bonds
- (c) **Simple molecular structure** – this is the structure adopted when atoms combine with covalent bonds to form molecules, which are held together by intermolecular bonds
- (d) **Giant covalent (or macromolecular) lattice structure** – this is the structure adopted when atoms combine with covalent bonds but cannot form stable molecules

2) Giant Ionic Structures

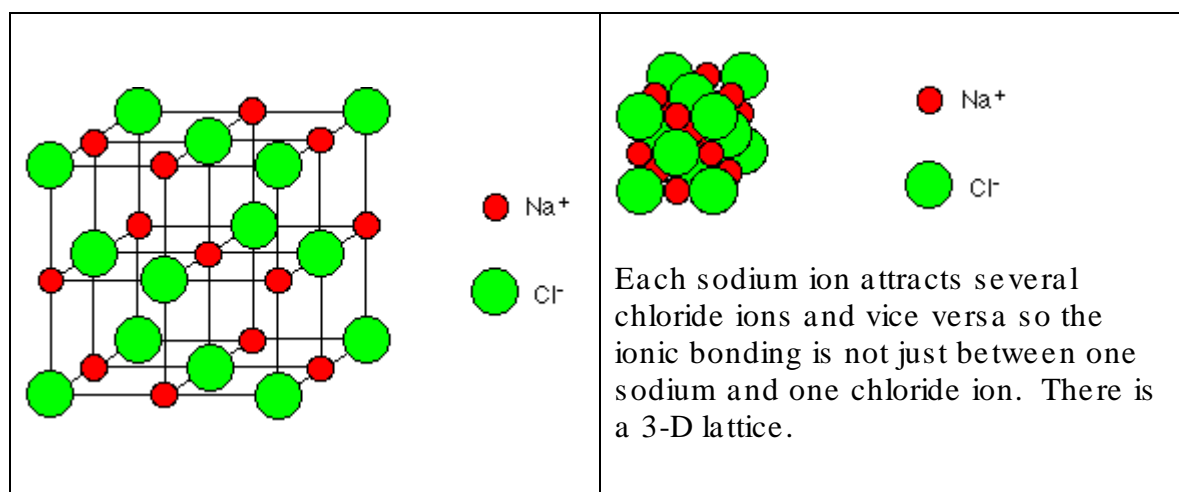
(a) Description of Structure

An ionic bond is an attraction between oppositely charged ions. After the ions are formed they all come together to form a **lattice**. All of the anions are surrounded by cations and all of the cations are surrounded by anions.

Example – sodium chloride

In sodium chloride, each sodium ion is surrounded by six chloride ions and vice versa.

The diagram below shows the structure of sodium chloride. The pattern repeats in this way and the structure extends (repeats itself) in all directions throughout a single crystal. (Remember - this diagram represents only a tiny part of a single sodium chloride crystal)



(b) Physical Properties

(i) Melting and boiling point

The attractions between the oppositely charged ions (ie the ionic bonds) are very strong. A lot of energy is thus required to overcome them. The melting point and boiling point of ionic compounds is very high.

Compound	NaCl	MgO
Melting point/°C	801	2852

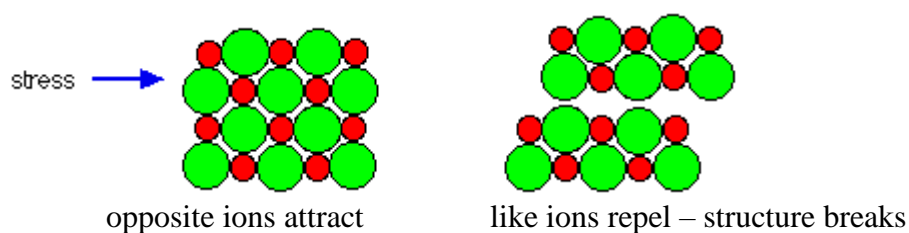
The higher the charge on the ions, and the smaller they are, the stronger the attraction between them will be and the higher the melting and boiling points. In MgO, the ions have a 2+ and 2- charge and thus the attraction between them is stronger than in NaCl, so the melting and boiling points are higher.

(ii) Electrical Conductivity

Since ionic solids contain ions, they are attracted by electric fields and ions will, if possible, move towards the electrodes and thus conduct electricity. In the solid state, however, the ions are not free to move since they are tightly held in place by each other. Thus ionic compounds do not conduct electricity in the solid state – they are insulators. In the liquid state, the ions are free to move and so can move towards their respective electrodes. Thus ionic compounds can conduct electricity in the liquid state. This is also true of ionic compounds dissolved in a solvent such as water.

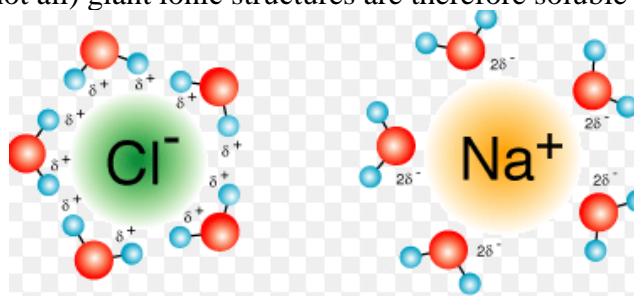
(iii) Mechanical properties

Since ions are held strongly in place by the other ions, they cannot move or slip over each other easily and are hence hard and brittle.



(iv) Solubility in water and non-polar solvents

Water is a polar molecule. Cations are attracted to the negative dipole on a water molecule and anions are attracted to the positive dipole on a water molecule. Sometimes this attraction is enough to break up the ionic lattice, and many (but not all) giant ionic structures are therefore soluble in water:



Non-polar solvents, such as ether and hexane, are not able to attract anions and cations in the same way and so giant ionic structures are generally insoluble in non-polar solvents.

(c) Formulae and state symbols

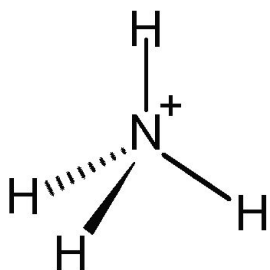

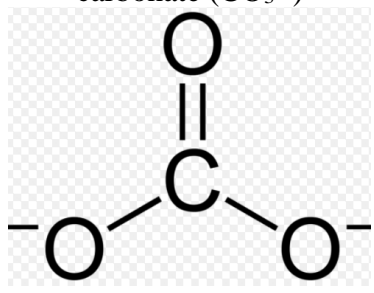
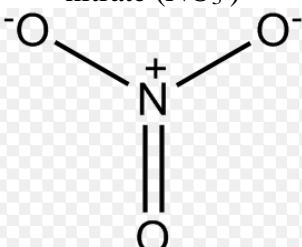
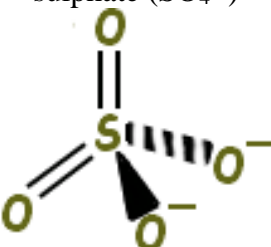
Ionic compounds are represented by their **unit formula**. The unit formula of an ionic compound is the simplest whole number ratio of the ions in the compound, shown by writing the cation followed by the anion:

Name of ionic compound	Formula	Meaning of formula
Sodium chloride	NaCl	Ratio of Na ⁺ to Cl ⁻ ions is 1:1
Sodium oxide	Na ₂ O	Ratio of Na ⁺ to O ²⁻ ions is 2:1
Magnesium chloride	MgCl ₂	Ratio of Mg ²⁺ to Cl ⁻ ions is 1:2
Aluminium oxide	Al ₂ O ₃	Ratio of Al ³⁺ to O ²⁻ ions is 2:3

The charges on the ions always cancel out, so the formula has no overall charge.

Some ions contain two or more atoms, held together by covalent bonds, but they also have a charge. These are known as polyatomic ions. You do not need to know how they are formed.

Common examples are:

ammonium (NH_4^+) 	hydroxide (OH^-) 	carbonate (CO_3^{2-}) 
nitrate (NO_3^-) 	sulphate (SO_4^{2-}) 	

The compounds formed from these ions are also giant ionic structures. If a unit formula contains more than one of these ions, the formula of the ion is written inside brackets and the number of ions is written outside the brackets:

Name of ionic compound	Formula	Meaning of formula
sodium carbonate	Na_2CO_3	Ratio of Na^+ to CO_3^{2-} ions is 2:1
ammonium sulphate	$(\text{NH}_4)_2\text{SO}_4$	Ratio of NH_4^+ to SO_4^{2-} ions is 2:1
ammonium nitrate	NH_4NO_3	Ratio of NH_4^+ to NO_3^- ions is 1:1
magnesium nitrate	$\text{Mg}(\text{NO}_3)_2$	Ratio of Mg^{2+} to NO_3^- ions is 1:2
aluminium sulphate	$\text{Al}_2(\text{SO}_4)_3$	Ratio of Al^{3+} to SO_4^{2-} ions is 2:3

Test Your Progress: Topic 2 Part 3 Exercise 1

- Deduce the unit formula of the following giant ionic structures:
 - ammonium chloride
 - aluminium nitrate
 - potassium nitrate
 - magnesium hydroxide
 - magnesium sulphate

At room temperature, giant ionic structures are found either in the solid state or dissolved in water. The physical state of the structure can be shown by the corresponding state symbol after the formula:

$\text{NaCl}(\text{s})$ means a solid sample of sodium chloride

$\text{NaCl}(\text{aq})$ means a sample of sodium chloride dissolved in water

3) Giant Metallic Structures

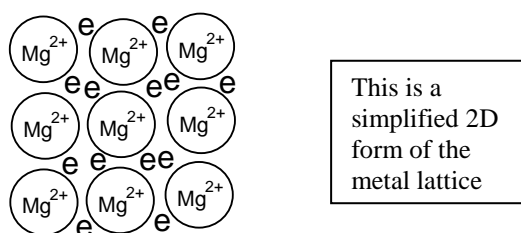
(a) Description of Structure

Metallic bonding is the attraction between cations and a sea of delocalised electrons. The cations are arranged to form a lattice, with the electrons free to move between them. This is a giant metallic structure.

A material which adopts a giant metallic structure is called a **metal**. A material which adopts any other structure is called a **non-metal**.

A simplified form of a magnesium lattice is shown below:

Example - magnesium



(b) Physical Properties

(i) Electrical conductivity

Since the electrons in a metal are delocalised, they are free to move throughout the crystal in a certain direction when a potential difference is applied and metals can thus conduct electricity in the solid state. The delocalised electron system is still present in the liquid state, so metals can also conduct electricity well in the liquid state.

(ii) Melting and boiling point

Metallic bonds (the attractions between cations and delocalised electrons) are relatively strong, although usually not as strong as ionic bonding. As a result the melting and boiling points of metals are relatively high.

Metal	Na	K	Be	Mg
Melting point/ °C	98	64	1278	649

Smaller ions, and those with a high charge, attract the electrons more strongly and so have higher melting points than larger ions with a low charge. Na has smaller cations than K so has a higher melting and boiling point. Mg cations have a higher charge than Na cations so has Mg a higher melting and boiling point than Na.

(iii) Mechanical properties

Since the bonding in metals is non-directional, it does not really matter how the cations are oriented relative to each other. The metal cations can be moved around and there will still be delocalized electrons available to hold the cations together. The metal cations can thus slip over each other fairly easily. As a result, metals tend to be soft, malleable and ductile.

(c) Formulae and state symbols

Almost all giant metallic structures are elements. They are therefore represented simply by the atom.

Name of giant metallic structure	Formula	Meaning of formula
Magnesium	Mg	Only Mg atoms present
Aluminium	Al	Only Al atoms present

Giant metallic structures are almost always solids at room temperature. To distinguish between an individual atom of magnesium and a sample of solid magnesium metal, the solid state symbol is used:

Mg (s) means a sample of solid magnesium metal.

4) Simple Molecular Structures

(a) Description of Structure

A covalent bond is a pair of electrons shared between two atoms. The electron pair in the covalent bond is attracted to the nuclei of both atoms, and this holds the atoms together. These bonds can be very strong.

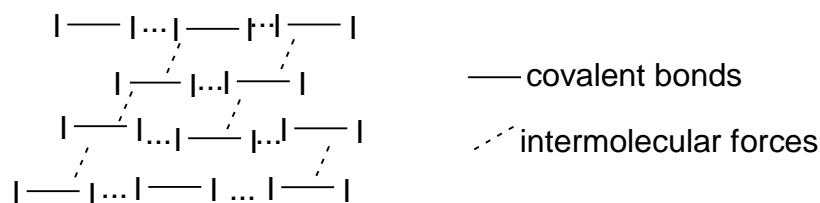
In most cases, atoms bond with one or a few other atoms to form a molecule, after which the bonding capacity of each atom has been reached. In such cases, the structure of the material consists of a large number of individual molecules with no strong bonds holding the molecules together.

The molecules are held together by **intermolecular forces**, which are much weaker than covalent bonds but are often strong enough to keep the substance in the solid or liquid state.

Structures consisting of molecules held together by intermolecular forces are known as simple molecular structures. In some cases the intermolecular forces are strong enough to form a solid lattice, but usually simple molecular structures are liquids or gases at room temperature.

There are many examples of simple molecular substances: CH_4 , I_2 , He, S_8 , P_4 , O_2 , H_2O , NH_3 .

Example - Iodine



There are attractive forces between these molecules, known as intermolecular forces (in this case Van der Waal's forces), but they are weak. In the gaseous state, the intermolecular forces are broken but the bonds within the molecule remain intact. The gas phase therefore consists of molecules, not atoms.

Noble gases do not form any bonds at all and so do not form molecules. The individual atoms are held together by Van der Waal's forces, so the structure is called **simple atomic** but the materials behave in a manner similar to simple molecular structures.

(b) Physical Properties**(i) Melting and boiling point**

The melting and boiling points of simple molecular structures are generally low, since intermolecular forces are weak. Most simple molecular structures are liquids or gases at room temperature.

The greater the number of electrons in a molecule, and the greater the surface area of the molecule, the stronger the Van der Waal's forces and the higher the boiling point of the substance:

Eg halogens

Substance	F ₂	Cl ₂	Br ₂	I ₂
Number of electrons	18	34	70	106
Boiling point/°C	-188	-35	59	184

Eg alkanes:

Substance	CH ₄	C ₂ H ₆	C ₃ H ₈	C ₄ H ₁₀
Number of electrons	10	18	26	34
Boiling point/°C	-164	-88	-42	0

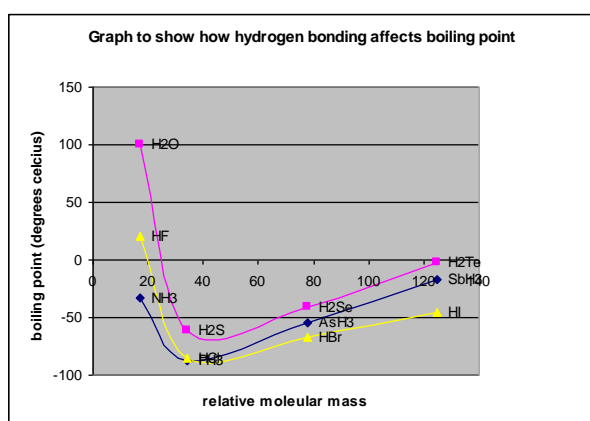
The presence of hydrogen bonding significantly increases the total strength of the intermolecular force between molecules. Molecules which form hydrogen bonds, such as HF, H₂O and NH₃, therefore have much higher boiling points than similarly sized molecules which do not have hydrogen bonding. This is most clearly seen in the boiling points of the hydrides of elements in groups V, VI and VII of the periodic table:

Group V: NH₃, PH₃, AsH₃, SbH₃

Group VI: H₂O, H₂S, H₂Se, H₂Te

Group VII: HF, HCl, HBr, HI

The boiling points of these graphs are shown graphically below:



The general increase in boiling points of the hydrides on descending the group results from the greater number of electrons in the molecules, which causes stronger Van der Waal's forces.

The unusually high boiling point of the Period 2 hydrides (HF, H₂O and NH₃) results from the ability of these molecules to form hydrogen bonds with each other.

(ii) Electrical conductivity

There are no ions and no delocalised electrons in simple molecular substances, so these substances do not show electrical conductivity in either solid or liquid state

(iii) Mechanical properties

Intermolecular forces are weak and generally non-directional, so most simple molecular substances are soft, crumbly and not very strong.

(iv) Solubility in water and non-polar solvents

Polar molecules can form stronger intermolecular forces with other polar molecules (because they can also form permanent dipole-dipole interactions or hydrogen bonds) than with non-polar molecules.

Polar molecules are therefore usually soluble in water and other polar solvents, but insoluble in non-polar solvents such as ether or hexane.

The opposite is true of non-polar molecules, which tend to be soluble in other non-polar solvents but insoluble in water.

(c) Formulae and state symbols

Simple molecular structures are represented by their **molecular formula**. This shows the actual number of atoms of each element in one molecule of the substance. The more electropositive atom is often, but not always, shown first.

Name of simple molecular structure	Formula	Meaning of formula
Chlorine	Cl ₂	Each molecule contains 2 Cl atoms
Oxygen	H ₂	Each molecule contains 2 O atoms
Water	H ₂ O	Each molecule contains 2 H atoms and 1 O atom
Ethane	C ₂ H ₆	Each molecule contains 2 C atoms and 6 H atoms
Carbon Dioxide	CO ₂	Each molecule contains 1 C atom and 2 O atoms
Ammonia	NH ₃	Each molecule contains 1 N atom and 3 H atoms
Glucose	C ₆ H ₁₂ O ₆	Each molecule contains 6 C, 12 H and 6 O atoms

The physical states of simple molecular structures can vary widely and the appropriate state symbol is often used to clarify the physical state in which the substance exists.

H₂O (s) means ice

H₂O (g) means steam or water vapour

I₂ (s) means solid iodine

I₂ (g) means iodine vapour

I₂ (aq) means iodine dissolved in water

5) Giant Covalent Structures

(a) Description of Structure

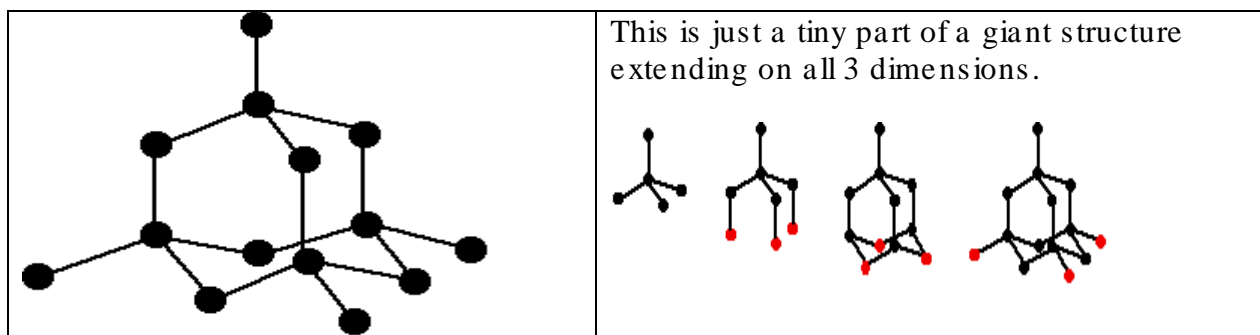
In some structures involving covalent bonds, it is not possible to satisfy the bonding capacity of the atoms by forming a simple molecule; the covalent bonds between atoms therefore continue indefinitely, and a large lattice is formed, with covalent bonds existing between all adjacent atoms.

Such substances are called **giant covalent** (or **giant atomic** or **macromolecular**) substances, and the most important examples are C, B, Si and SiO₂. The giant structure can be two dimensional or three dimensional.

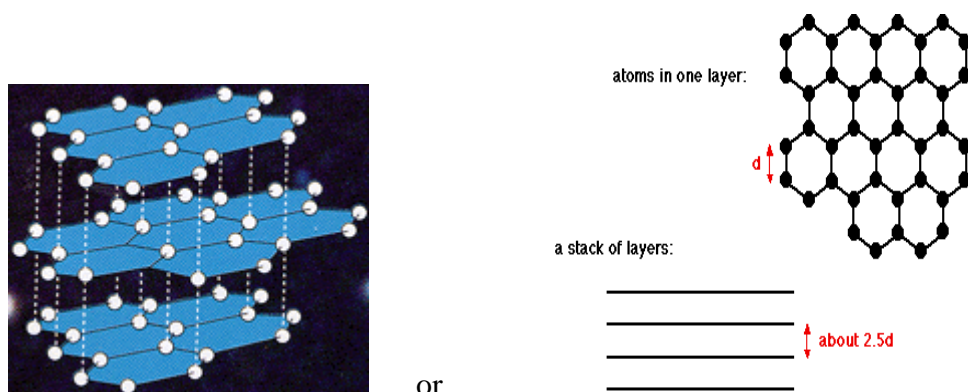
Examples – diamond and graphite

The element carbon can exist in two different forms (different forms of the same element are called **allotropes**) – diamond and graphite. Both form giant covalent structures:

In diamond, each C atom forms four covalent bonds with adjacent atoms. The bonds arrange themselves in a tetrahedral shape. The network of covalent bonds stretches throughout the crystal in three dimensions.



In graphite, each C atom forms three covalent bonds with adjacent atoms and bonds arrange themselves in a planar shape to form layers, which stretch throughout the crystal in two dimensions. The spare electron is delocalized and occupies the space in between the layers. All atoms in the same layer are held together by strong covalent bonds, and the different layers are held together by intermolecular forces.



In graphite, each carbon atom is bonded to three others. The spare electron is delocalized and occupies the space in between the layers. All atoms in the same layer are held together by strong covalent bonds, and the different layers are held together by intermolecular forces.

(b) Physical Properties**(i) Melting and boiling point**

In giant covalent substances, covalent bonds must be broken before the lattice can break down and the atoms separated. This is true of both three-dimensional lattices (like diamond) and two-dimensional lattices (like graphite). Covalent bonds are usually very strong and breaking them requires a lot of energy, so giant covalent substances have high melting and boiling points. The precise melting and boiling points depend on the number of bonds formed by each atom and the bond strength.

Substance	C	Si	B	SiO ₂
Melting point /°C	3550	1410	2300	1510
Boiling point /°C	4827	2355	2550	2230

(ii) Electrical conductivity

Giant covalent substances contain no ions. Most also contain no delocalised electrons, so there is no electrical conductivity in either solid or liquid state. Diamond and silicon dioxide (sand) do not conduct electricity.

Graphite is an important exception – the delocalised electrons in graphite make it a good conductor of electricity.

(iii) Mechanical Properties

The mechanical properties of giant covalent substances depend on whether the giant lattice is three-dimensional or two-dimensional. Three dimensional giant covalent lattices are hard, strong and brittle. Diamond is in fact the hardest substance known to man. For this reason it is used in drills, glass-cutting and styluses for turntables. 2D giant covalent substances (like graphite) form layers which can slip over easily, as the layers are only held together by Van der Waal's forces. Graphite is therefore soft, and can be used in pencils and lubricants.

(c) Formulae and state symbols

Giant covalent compounds, like ionic compounds, are represented by their **unit formula**. The unit formula of a giant covalent compound is the simplest whole number ratio of the atoms in the compound. The more electropositive atom is often, but not always, shown first.

Name of giant covalent compound	Formula	Meaning of formula
silicon dioxide	SiO ₂	Ratio of Si to O atoms is 1:2

Giant covalent elements are represented simply by the atom:

Name of giant covalent element	Formula	Meaning of formula
boron	B	Only B atoms present

Giant covalent substances are almost always found in the solid state. This is often used to distinguish between the element and the atom:

Si (s) means a solid sample of silicon.

6) Summary of different types of chemical structure and their properties

Bond Type	Type of Structure and Description	Physical properties
IONIC	GIANT IONIC STRUCTURE (eg NaCl) Attraction between oppositely charged ions. Infinite lattice of oppositely charged ions in three dimensions	High mpt, bpt Good conductors in liquid state Poor conductors in solid state Hard, strong, brittle Often soluble in water but insoluble in non-polar solvents
METALLIC	GIANT METALLIC STRUCTURE (eg NaCl) Attraction between cations and delocalised electrons. Infinite lattice of cations in three dimensions, with delocalised electrons in the spaces	High mpt, bpt Good conductors in solid state Good conductors in liquid state Strong, malleable
COVALENT	Either SIMPLE MOLECULAR STRUCTURE (Eg I ₂) Discrete molecules. Atoms in molecule held together by covalent bonds. Weak intermolecular forces between molecules.	Low mpt, bpt Poor conductors in solid state Poor conductors in liquid state Soft, weak, powdery Polar molecules are soluble in water; non-polar molecules are soluble in non-polar solvents
	Or GIANT COVALENT STRUCTURE (eg diamond and graphite) Infinite lattice of atoms linked by covalent bonds in two or three dimensions.	Very high mpt, bpt Poor conductors in solid state (unless delocalised electrons present) Poor conductors in liquid state Hard, strong, brittle (if 3D); Soft (if 2D) Insoluble in water and non-polar solvents

Test Your Progress: Topic 2 Part 3 Exercise 2

1. Name the four main chemical structures and give one example of each.
2. Why does sodium chloride have a high melting point?
3. Why does sodium chloride conduct electricity when molten but not when solid?
4. Why does sodium chloride dissolve in water but not in ether?
5. Why does sodium chloride have a lower melting point than magnesium oxide?
6. Why do metals conduct electricity?
7. Why are metals malleable and ductile?
8. Why does magnesium have a high melting point?
9. Why does magnesium have a higher melting point than sodium?
10. Why does chlorine have such a low boiling point?
11. Why does bromine have a higher boiling point than chlorine?
12. Why does water have a higher boiling point than hydrogen sulphide?
13. Why does hydrogen selenide have a higher boiling point than hydrogen sulphide?
14. Why does ammonia dissolve in water but not in hexane?
15. Why does iodine dissolve in hexane but not in water?
16. Why do diamond and graphite have such high melting points?
17. Why does graphite conduct electricity and why does diamond not do so?
18. Why is diamond hard and why is graphite soft?
19. Why is diamond used in drills?
20. Why is graphite used in pencils?

7) Structure and Bonding in the Periodic Table

Atoms in the periodic table show **periodic** properties (trends in properties which repeat themselves across each period).

Trends across the Periodic Table already discussed include:

- A general increase in ionization energy across a period, and a decrease down a group
- A general increase in electron affinity across a period, and a decrease down a group
- A decrease in atomic and ionic size across a period, and an increase down a group
- An increase in electronegativity across a period, and a decrease down a group

Elements are substances made up of only one type of atom. For each type of atom, therefore, there is a corresponding element. Elements in the Periodic Table show periodic properties just like the atoms do.

(a) Structure and Bonding in Elements

All the atoms in a single element have the same electronegativity, so the bond type cannot be ionic – it has to be either metallic or covalent.

Atoms with low electronegativities form metallic bonds with themselves. The resulting elements therefore have giant metallic structures (ie they are metals).

Atoms with higher electronegativities form covalent bonds with themselves. The resulting elements therefore have either simple molecular structures or giant covalent structures (ie they are non-metals).

Noble gases form no bonds with themselves. The resulting elements therefore have simple atomic structures (ie they are non-metals).

As electronegativity increases across a period, **the elements therefore become less metallic across a period** (sodium and magnesium are metals but chlorine and argon are non-metals).

As electronegativity decreases down a group, **the elements therefore become more metallic down a group** (carbon and silicon are non-metals but tin and lead are metals).

The periodic variation in structure, and hence metallic character, can be shown in the table below:

H (H ₂) simple molecular (non-metal)							He simple atomic (non-metal)
Li giant metallic (metal)	Be giant metallic (metal)	B giant covalent (non-metal)	C giant covalent (non-metal)	N (N ₂) simple molecular (non-metal)	O (O ₂) simple molecular (non-metal)	F (F ₂) simple molecular (non-metal)	Ne simple atomic (non-metal)
Na giant metallic (metal)	Mg giant metallic (metal)	Al giant metallic (metal)	Si giant covalent (non-metal)	P (P ₄) simple molecular (non-metal)	S (S ₈) simple molecular (non-metal)	Cl (Cl ₂) simple molecular (non-metal)	Ar simple atomic (non-metal)

Lithium, beryllium, sodium, magnesium and aluminium have **giant metallic structures** (ie they are metals). They consist of an infinite lattice of cations held together by a sea of delocalised electrons. There is a fairly strong attraction between the cations and the delocalised electrons and as a result metals tend to have fairly high melting points and boiling points.

The melting points increase with increasing charge and decreasing size and thus increase across a period.

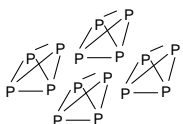
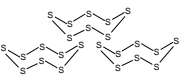
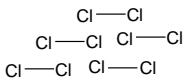
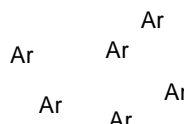
Element	Sodium	Magnesium	Aluminium
Mpt/ $^{\circ}\text{C}$	98	669	680

Boron, carbon (both graphite and diamond) and silicon have **giant covalent structures** and are hence non-metals. Silicon has a 3D tetrahedral structure identical to diamond and boron has a 2D layered structure similar to graphite. In all cases, strong covalent bonds link all atoms in the lattice and these elements have high melting points. With the exception of graphite, these elements do not conduct electricity.

Hydrogen, nitrogen, oxygen, fluorine, phosphorus, sulphur and chlorine form simple molecular structures. There are strong, covalent bonds within the molecule but the different molecules are only held together by weak Van der Waal's forces. Separating these molecules thus requires little energy and the melting and boiling points of these elements are relatively low.

Helium, neon and argon form simple atomic structures with weak Van der Waal's forces holding the atoms together.

The larger the molecule, the greater the magnitude of the temporary and induced dipoles and the higher the melting and boiling points.

Element	Phosphorus	Sulphur	Chlorine	Argon
Structure				
Bpt/ $^{\circ}\text{C}$	280	445	-35	-186
Formula	P ₄ (or P)	S ₈ (or S)	Cl ₂	Ar

(b) Bonding in Compounds – oxides and chlorides

Atoms on the left of the Periodic Table have low electronegativities, and hence a large difference in electronegativity with oxygen and chlorine, both of which have high electronegativities. Atoms on the of the Periodic Table therefore tend to form ionic bonds with oxygen and chlorine, and the resulting compounds have **giant ionic** structures.

As you move from left to right in the Periodic Table, the electronegativities increase, so the difference in electronegativity between the atoms and oxygen or chlorine becomes smaller. So the atoms start to form covalent bonds with oxygen and chlorine, and the resulting compounds have **simple molecular** or **giant covalent** structures.

On moving across a Period from left to right, the type of bond formed with oxygen and chlorine becomes less ionic and more covalent in character.

Type of structure	Examples of chlorides and oxides
Giant Ionic	LiCl, NaCl, MgCl ₂ , Li ₂ O, Na ₂ O, BeO, Al ₂ O ₃
Simple molecular	HCl, H ₂ O, BeCl ₂ , AlCl ₃ , SiCl ₄ , PCl ₃ , SCl ₂ , NO, P ₄ O ₆
Giant covalent	SiO ₂

Test Your Progress: Topic 2 Part 3 Exercise 3

- Why do the elements become less metallic across a Period?
- From the first 20 elements in the Periodic Table, name five metals and five non-metals.
- Explain why sodium chloride has a high melting point and conducts electricity when molten, but aluminium chloride has a low melting point and does not conduct electricity when molten.

8) Elements, Compounds and Mixtures

Materials can be described as **elements, compounds or mixtures** depending on their composition.

(a) Elements

A substance made up of only one type of atom is called an **element**. The name of the element is always the same as the name of the atom, and each atom gives rise to one element. The Periodic Table is a list of known atoms and therefore a list of known elements.

The bonding in elements can be metallic or covalent, and elements are pure substances which cannot be broken down chemically. Elements are often classified as metals and non-metals depending on their structures.

Examples of elements: Hydrogen, Oxygen, Carbon, Chlorine, Zinc

(b) Compounds

A substance made up of two or more different elements bonded together in a fixed proportion is called a **compound**. There are a large number of different compounds.

Compounds can contain ionic or covalent bonding, and compounds are pure substances with a specific chemical formula showing the relative composition of different elements.

There are two laws of chemical combination which specify that compounds have a fixed chemical formula:

(i) The Law of Constant Composition:

A given chemical compound always contains its component elements in fixed ratio (by mass), and this fixed ratio does not depend on its source and method of preparation.

Eg copper oxide contains 80% copper and 20% oxygen by mass, regardless of whether copper oxide is prepared by the direct combination of copper and oxygen, or by the decomposition of copper carbonate, or by the decomposition of copper nitrate. Samples of copper oxide prepared by each of the three methods can be reduced to copper by passing hydrogen gas over the copper oxide. In all three cases, the mass of the solid will drop by exactly 20% (as it is reduced from copper oxide to copper).

(ii) The Law of Constant Composition:

If two elements form more than one compound between them, then the ratios of the masses of the second element which combine with a fixed mass of the first element will be ratios of small whole numbers.

Eg 1.2 g of carbon will combine with 1.6 g of oxygen to form carbon monoxide, or with 3.2 g of oxygen to form carbon dioxide. The ratio of the two masses of oxygen = $3.2:1.6 = 2:1$.

Compounds can be broken down into their constituent elements, but not easily – usually a chemical reaction or electrolytic process is required.

Examples of compounds: sodium chloride (NaCl), water (H₂O), methane (CH₄), ethane (C₂H₆), silicon dioxide (SiO₂).

(c) Mixtures

A substance made up of two or more elements or compounds but not in fixed proportion is called a **mixture**. Mixtures are not pure substances, have no fixed composition and have no single chemical formula. In most cases the different components of the mixture are held together only by intermolecular forces.

Mixtures with giant metallic structures (and hence containing at least one metallic element) are known as **alloys**. Example of alloys are brass, bronze, solder and steel.

Mixtures can take almost any form. Sea water, air, milk, crude oil and most types of earth are all examples of mixtures.

Mixtures can be separated into their pure components (elements and/or compounds) using simple physical separation techniques such as filtration, distillation (simple or fractional), evaporation or chromatography.