

Bonding, structure, and the properties of matter

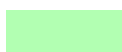
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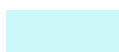
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These are the summary notes only. The full edition of this booklet also carries the worked examples, the required practicals and the formula sheets, with every topic in one place.

Colour Codes



Definitions



Diagrams



Notes and warnings



Sub-topic headings

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Define Ionic bonding

- the strong electrostatic force of attraction between oppositely charged ions, acting in all directions in the lattice.

Define Delocalised electron

- an electron that is not attached to one atom and is free to move through the whole structure.

Define Intermolecular forces

- the weak forces **between** molecules, overcome on melting or boiling (the covalent bonds are not broken).

Define Giant covalent structure

- a solid in which all atoms are joined to other atoms by strong covalent bonds throughout the structure.

Define State symbols

- (s) solid, (l) liquid, (g) gas, (aq) aqueous solution.

Define Empirical formula

- the simplest whole-number ratio of the ions or atoms in a compound.

Define Graphene

- a single layer of graphite.

Define CHEMISTRY ONLY Surface area to volume ratio

- surface area ÷ volume; it increases as a particle gets smaller.

Define Covalent bond

- a shared pair of electrons between two atoms.

Define Metallic bonding

- the strong attraction between positive metal ions and the shared delocalised electrons.

Define Giant ionic lattice

- the regular structure of a large number of oppositely charged ions held by ionic bonds.

Define Polymer

- a substance with very large molecules made of many repeating units joined by strong covalent bonds.

Define Alloy

- a mixture of a metal with one or more other elements, harder than the pure metal.

Define Fullerene

- a molecule of carbon atoms with a hollow shape, based on hexagonal rings that may also contain rings of five or seven carbon atoms.

Define CHEMISTRY ONLY Nanoscience

- structures that are 1–100 nm in size, of the order of a few hundred atoms.

2.1 Chemical bonds, ionic, covalent and metallic

2.1.1 Chemical bonds

There are **three types of strong chemical bond**: ionic, covalent and metallic.

Bond	The particles are...	Occurs in
Ionic	Oppositely charged ions	Compounds formed from metals combined with non- metals
Covalent	Atoms which share pairs of electrons	Most non-metallic elements and compounds of non- metals
Metallic	Atoms which share delocalised electrons	Metallic elements and alloys

You must be able to explain chemical bonding in terms of **electrostatic forces** and the transfer or sharing of electrons.

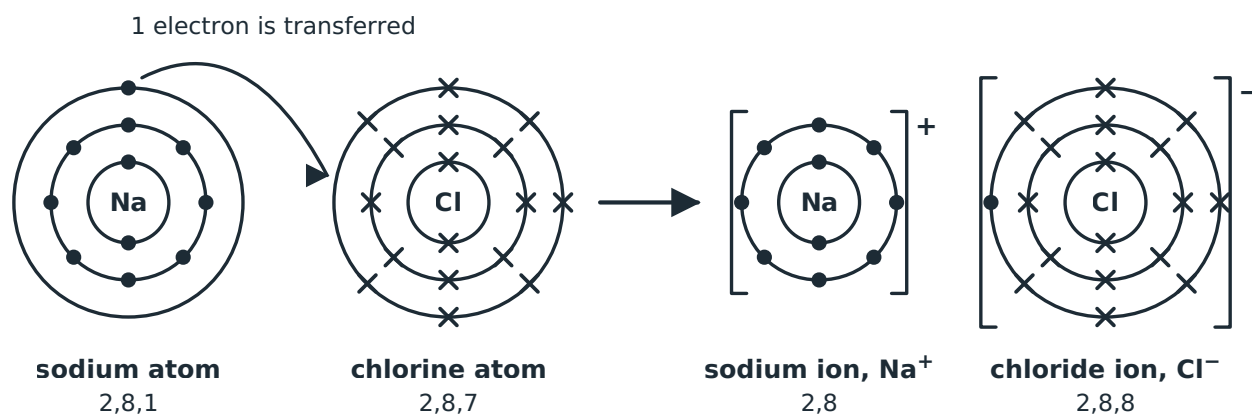
2.1.2 Ionic bonding

- 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 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).
- The electron transfer can be represented by a dot and cross diagram, eg for sodium chloride. Draw the outer shells only, with the metal's electrons as dots and the non-metal's as crosses, an arrow showing the transfer, and each ion drawn in square brackets with its charge outside.
- You should be able to draw dot and cross diagrams for ionic compounds formed by metals in Groups 1 and 2 with non-metals in Groups 6 and 7.
- The charge on the ion relates to the group number.

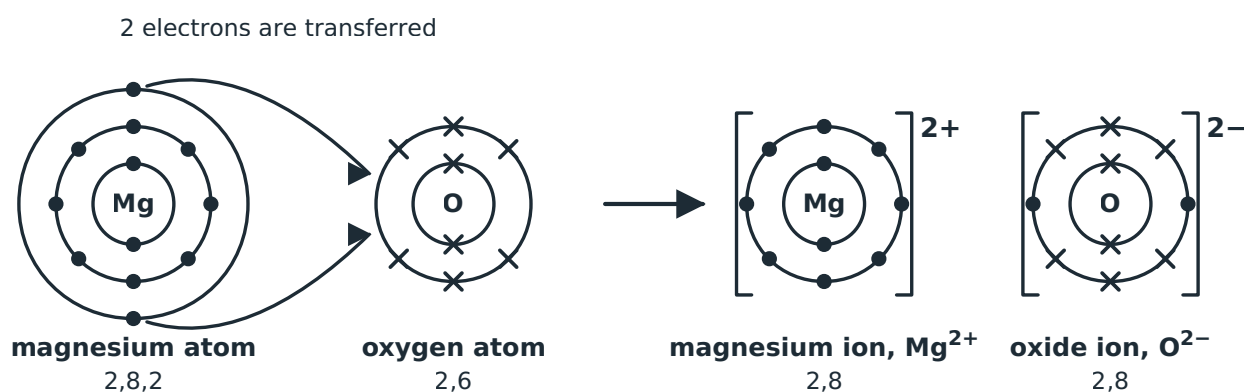
ANALOGY

Ionic vs covalent bonding. Ionic bonding is a **transfer of money** — one atom hands electrons over completely, and the two are then held together because one is positive and the other negative. Covalent bonding is two people sharing a bag of sweets they both hold on to — neither lets go, and the shared grip is the bond. The analogy breaks down in that ionic attraction acts in **all directions**, not just between the two ions that swapped electrons.

Dot and cross diagrams for ionic bonding: sodium chloride and magnesium oxide



Sodium chloride, NaCl — a Group 1 metal with a Group 7 non-metal



Magnesium oxide, MgO — a Group 2 metal with a Group 6 non-metal

● = an electron from the metal atom × = an electron from the non-metal atom

Metal atoms lose electrons to form positively charged ions; non-metal atoms gain electrons to form negative ions. Both ions get the electronic structure of a noble gas.

Ions you must be able to work out:

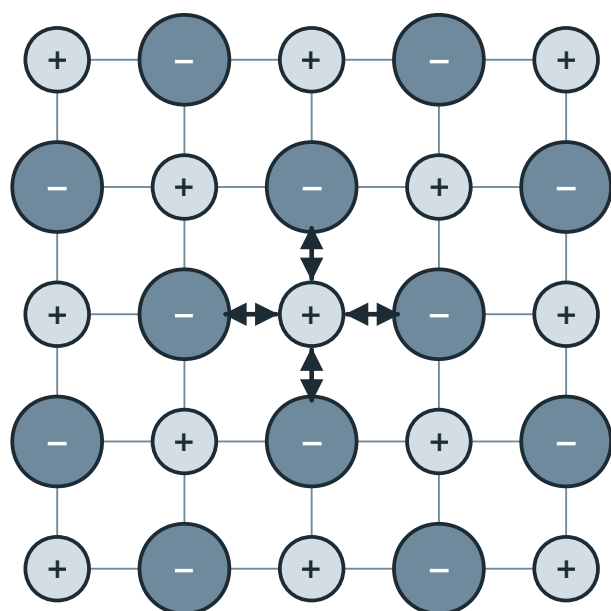
Group	Example element	Electrons lost/gained	Ion formed
1	Sodium	Loses 1	Na ⁺
2	Magnesium	Loses 2	Mg ²⁺
6	Oxygen	Gains 2	O ²⁻
7	Chlorine	Gains 1	Cl ⁻

So Group 1 → 1+, Group 2 → 2+, Group 6 → 2-, Group 7 → 1-.
The spec limits this to metals in Groups 1 and 2 and non-metals in Groups 6 and 7; transition metal charges such as Fe²⁺ and Fe³⁺ are given in the question, usually in the compound's name.

2.1.3 Ionic compounds

- 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**.
- The structure of **sodium chloride** can be shown as a dot and cross diagram, a ball and stick model, a 2D diagram or a 3D diagram.
- You should be able to deduce that a compound is ionic from a diagram of its structure in one of these forms, and **work out the empirical formula** of an ionic compound from a model or diagram showing the ions.

The giant ionic lattice of sodium chloride, and why it conducts only when molten or dissolved



-  sodium ion, Na^+
(positively charged)
-  chloride ion, Cl^-
(negatively charged)
-  strong electrostatic attraction

The giant ionic lattice

Sodium chloride is a giant ionic lattice: a regular repeating arrangement of oppositely charged ions. The ions are held together by strong electrostatic forces of attraction acting in all directions. This is ionic bonding.

Why the melting point is high

There are many strong bonds to overcome, so a large amount of energy is needed. Melting and boiling points are high.

Conducting electricity

Solid: does not conduct. The ions are held in fixed positions and cannot move.

Molten or dissolved in water: conducts. The ions are free to move, so charge can flow.

Limitation of the diagram

It shows only one layer — the lattice really continues in three dimensions.

Limitations of the models (a very common exam question):

Model	Limitation
Dot and cross	Shows where the electrons came from, but nothing about the 3D arrangement, ion sizes or how many ions there are
Ball and stick	Shows the 3D arrangement, but ions are drawn far apart joined by "sticks" — there are no rods, and the ions really touch
2D diagram	Simple, but a flat drawing cannot show that the forces act in all directions in three dimensions
3D space-filling	Shows relative sizes and that ions touch, but you cannot see into the middle of the structure

All of these show only a tiny part of a lattice containing a huge number of ions.

2.1.4 Covalent bonding

- When atoms share pairs of electrons they form **covalent bonds**. These bonds between atoms are strong.
- Covalently bonded substances may consist of **small molecules** — recognise these from a chemical formula of a few non-metal atoms, eg CO_2 , NH_3 , CH_4 .
- Some have very large molecules, such as polymers; some have giant covalent structures, such as diamond and silicon dioxide.
- Covalent bonds can be shown by dot and cross diagrams, by a line for each single bond, by ball and stick models and by 2D and 3D diagrams. Polymers are drawn as a repeating unit in brackets with n outside, where n is a large number.

You must be able to:

- draw dot and cross diagrams for hydrogen, chlorine, oxygen, nitrogen, hydrogen chloride, water, ammonia and methane
- represent covalent bonds in small molecules, in the repeating units of polymers and in part of giant covalent structures, using a line to represent a single bond
- describe the **limitations** of dot and cross, ball and stick, and 2D and 3D diagrams for molecules and giant structures
- deduce the molecular formula** of a substance from a given model or diagram.

Molecule	Bonding	Shared pairs
H_2 / Cl_2 / HCl	H-H , Cl-Cl , H-Cl	1 single bond (each Cl also has 3 lone pairs)
O_2	O=O	2 shared pairs (double bond)
N_2	$\text{N}\equiv\text{N}$	3 shared pairs (triple bond)
H_2O	H-O-H	2 single bonds (O has 2 lone pairs)
NH_3	$3 \times \text{N-H}$	3 single bonds (N has 1 lone pair)
CH_4	$4 \times \text{C-H}$	4 single bonds

Dot and cross and displayed formulae for the small molecules you must be able to draw

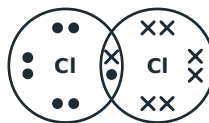
Hydrogen, H_2

displayed formula



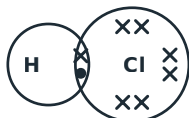
Chlorine, Cl_2

displayed formula



Hydrogen chloride, HCl

displayed formula



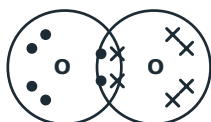
Water, H_2O

displayed formula



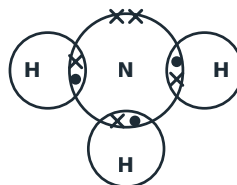
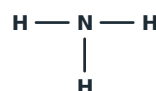
Oxygen, O_2

displayed formula



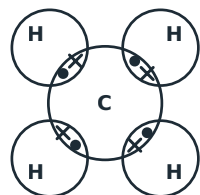
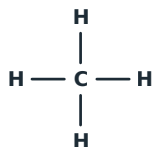
Ammonia, NH_3

displayed formula



Methane, CH_4

displayed formula



Key

• = an electron from one atom

× = an electron from the other atom

Only the outer shell electrons are shown.

Each shared pair of electrons is one covalent bond, shown as one line in the displayed formula.

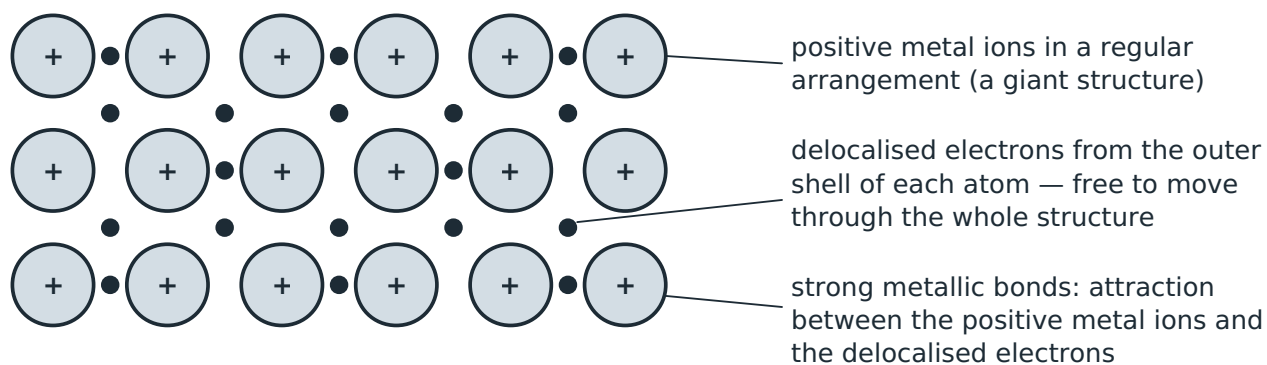
WATCH OUT

When a molecular substance melts or boils, it is the **weak intermolecular forces** that are overcome — not the covalent bonds. Writing "the covalent bonds break when water boils" loses the mark.

2.1.5 Metallic bonding

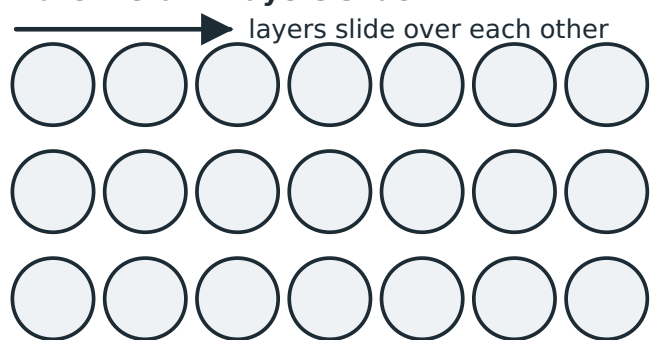
- Metals consist of giant structures of atoms arranged in a regular pattern.
- The electrons in the **outer shell** of metal atoms are **delocalised** and so are free to move through the whole structure.
- The sharing of delocalised electrons gives rise to strong metallic bonds.
- Metallic bonding is drawn as rows of positive ions in a "sea" of delocalised electrons, usually shown as small negative signs between the ions.

Metallic bonding: positive ions in a sea of delocalised electrons, and why alloys are harder



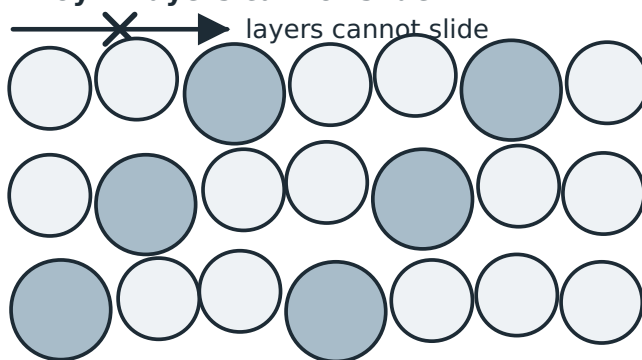
Metals conduct electricity and thermal energy because the delocalised electrons are free to move through the whole structure, carrying charge and energy.

Pure metal — layers slide



The atoms are all the same size, so the layers of atoms can slide over each other. Pure metals are soft and can be bent and shaped.

Alloy — layers cannot slide



The larger shaded atoms are atoms of a different element. They distort the layers, so the layers can no longer slide over each other — the alloy is harder.

ANALOGY

Metallic bonding — the "sea of delocalised electrons". Picture ping-pong balls (the positive metal ions) sitting in a tray of water (the delocalised electrons). The water flows around every ball and holds the whole tray together — which is why layers can slide without the metal snapping, and why moving the "water" carries charge and energy through the metal.

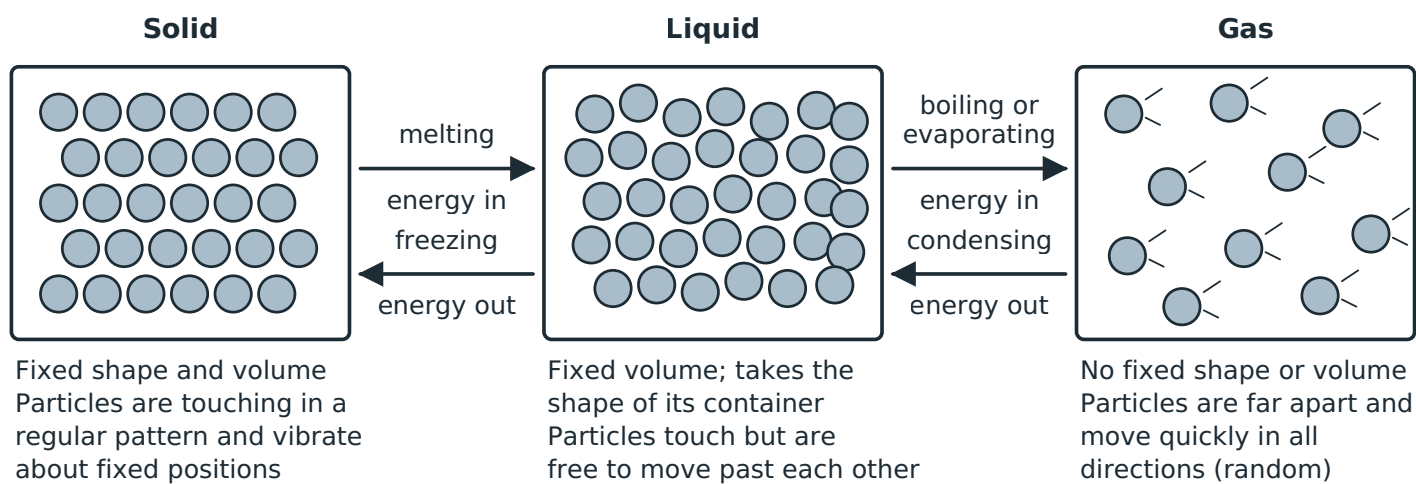
2.2 How bonding and structure are related to the properties of substances

2.2.1 The three states of matter

- The three states of matter are solid, liquid and gas. Melting and freezing take place at the melting point; boiling and condensing take place at the boiling point.
- The three states can be represented by a **simple model** in which particles are represented by **small solid spheres**. Particle theory helps to explain melting, boiling, freezing and condensing.
- 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. The nature of the particles depends on the type of bonding and the structure. The stronger the forces between the particles, the higher the melting point and boiling point.
- Atoms themselves do not have the bulk properties of materials — a single copper atom is not shiny, malleable or a conductor; those are properties of the structure.
- **HT ONLY** Limitations of the simple model: it has no forces between the particles, all particles are represented as spheres, and the spheres are solid. In reality forces exist (they are what must be overcome to change state), particles are not all spherical, and they are not solid inelastic balls.

You must be able to predict the states of substances at different temperatures given appropriate data: solid below the melting point, liquid between melting and boiling point, gas above the boiling point.

The three states of matter and the changes of state between them



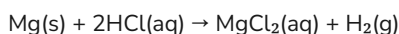
Melting and freezing happen at the melting point. Boiling and condensing happen at the boiling point. The stronger the forces between the particles, the higher the melting point and the boiling point.

HT only — Limitations of the simple particle model

In this model the particles are drawn as small solid spheres, all identical, with no forces between them. Real particles are not solid inelastic spheres, are not all the same, and forces do act between them.

2.2.2 State symbols

In chemical equations, the three states of matter are shown as **(s)**, **(l)** and **(g)**, with **(aq)** for **aqueous solutions** (dissolved in water). You must include appropriate state symbols in chemical equations, eg:



2.2.4 Properties of small molecules

- Substances that consist of small molecules are usually gases or liquids with relatively low melting points and boiling points.
- They 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.
- Intermolecular forces increase with the size of the molecules, so larger molecules have higher melting and boiling points.
- They do not conduct electricity because the molecules do not have an overall electric charge (there are no free ions or delocalised electrons).

2.2.6 Giant covalent structures

- Substances with **giant covalent structures** are solids with very high melting points.
- All of the atoms are linked to other atoms by strong covalent bonds, and these bonds must be overcome to melt or boil the substance.
- Examples: diamond and graphite (forms of carbon) and **silicon dioxide (silica)**.

2.2.8 Metals as conductors

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

Properties of each structure type — the table to learn:

2.2.3 Properties of ionic compounds

- Ionic compounds have **regular structures (giant ionic lattices)** in which there are strong electrostatic forces of attraction in all directions between oppositely charged ions.
- They have high melting points and high boiling points because **large amounts of energy** are needed to break the **many strong bonds**.
- When melted or dissolved in water, ionic compounds conduct electricity because the ions are free to move and so charge can flow. As solids they do not conduct, because the ions are held in fixed positions in the lattice.

2.2.5 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 (the molecules are very long), so these substances are solids at room temperature.
- You should be able to recognise polymers from diagrams showing their bonding and structure — look for a repeating unit in brackets with *n*.

2.2.7 Properties of metals and alloys

- Metals have giant structures of atoms with strong metallic bonding, so most metals have high melting and boiling points.
- In **pure metals**, atoms are arranged in layers, which allows metals to be bent and shaped (the layers slide over each other).
- Pure metals are too soft for many uses and so are mixed with other metals to make alloys, which are harder.
- **Why alloys are harder:** the added atoms are a different size, so they **distort the layers** of atoms in the structure of the pure metal. The layers can no longer slide over one another easily.

Structure	Melting/boiling point	Conducts as a solid?	Conducts when molten or in solution?	Why
Giant ionic	High	No	Yes	Ions fixed in the lattice when solid; free to move when melted or dissolved
Small molecules	Low	No	No	Weak intermolecular forces; molecules have no overall charge
Polymers	Solid at room temperature	No	No	Very large molecules, so relatively strong intermolecular forces; no charged particles
Giant covalent	Very high	No (graphite and graphene excepted)	No	Many strong covalent bonds to overcome; no charged particles free to move
Metallic	High	Yes	Yes	Delocalised electrons are free to move and carry charge

2.3 Structure and bonding of carbon

2.3.1 Diamond

In diamond, each carbon atom forms four covalent bonds with other carbon atoms in a **giant covalent structure**, so diamond:

- is very hard — every atom is held by four strong covalent bonds in a rigid 3D network, so it is used for cutting tools
- has a **very high melting point** — a very large amount of energy is needed to overcome the many strong covalent bonds
- **does not conduct electricity** — all four outer electrons of each carbon atom are used in bonding, so there are no delocalised electrons free to carry charge.

ANALOGY

Why graphite conducts but diamond does not. Think of each carbon atom as having four hands. In **diamond** all four are gripping neighbours — nothing is left free to move, so no current. In **graphite** only three grip, so the fourth electron is free to drift through the structure carrying charge. The layers are like sheets of paper in a stack: strong across each sheet, easy to slide apart.

2.3.2 Graphite

- In graphite, each carbon atom forms three covalent bonds with three other carbon atoms, forming **layers of hexagonal rings** which have no covalent bonds between the layers.
- In graphite, one electron from each carbon atom is delocalised.
- Graphite is similar to metals in that it has delocalised electrons.

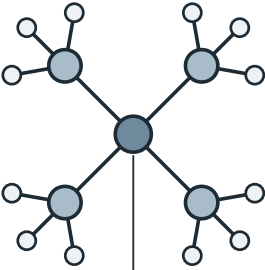
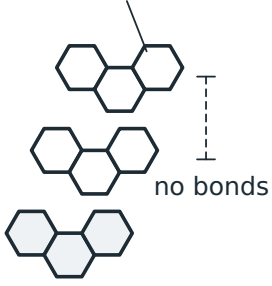
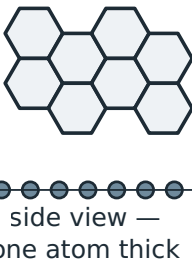
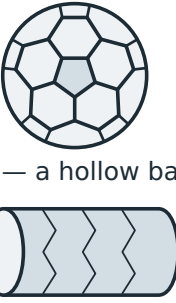
Explaining the properties:

- **Soft and slippery** — only weak forces hold the layers together, so they slide over each other. Used as a lubricant and in pencils.
- **Very high melting point** — the covalent bonds within the layers are strong and many must be overcome.
- **Conducts electricity and thermal energy** — the delocalised electrons move freely through the structure and carry charge. Used for electrodes.

2.3.3 Graphene and fullerenes

- **Graphene** is a **single layer of graphite** and has properties that make it useful in electronics and composites — it is one atom thick, very strong for its mass, and conducts electricity because of the delocalised electrons.
- **Fullerenes** are molecules of carbon atoms with hollow shapes. Their structure 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 — they conduct electricity and have very high tensile strength, so they reinforce composites such as tennis rackets.
- Other uses of fullerenes include delivering drugs into the body, **lubricants** and **catalysts** (large surface area).
- You should be able to recognise graphene and fullerenes from diagrams and descriptions of their bonding and structure.

Diamond, graphite, graphene and fullerenes compared

Diamond	Graphite	Graphene	Fullerenes
 <p>each C: 4 bonds</p>	 <p>hexagonal rings</p> <p>no bonds</p>	 <p>side view — one atom thick</p>	 <p>C_{60} — a hollow ball</p> <p>carbon nanotube</p>
Each carbon atom forms 4 covalent bonds — giant covalent structure Very hard, very high melting point Does not conduct electricity	Each carbon atom forms 3 covalent bonds, in layers of hexagonal rings No bonds between layers, so it is soft 1 delocalised electron per atom — conducts	A single layer of graphite — one atom thick Very strong for its low mass Conducts electricity Used in electronics and composites	Molecules of carbon with hollow shapes, based on hexagonal rings (some 5- or 7-carbon rings) C_{60} was the first one Nanotubes: cylinders, high length:diameter

WATCH OUT

Fullerenes such as C_{60} are **molecules**, not giant structures. C_{60} has a low melting point compared with graphite because only the weak intermolecular forces between the molecules need to be overcome — the covalent bonds inside each ball stay intact.

2.4 Bulk and surface properties of matter including nanoparticles CHEMISTRY ONLY

2.4.1 Sizes of particles and their properties CHEMISTRY ONLY

CHEMISTRY ONLY The whole of section 2.4 is **chemistry only** — it is not on the Combined Science course.

- **Nanoscience** refers to structures that are **1–100 nm** in size, of the order of a few hundred atoms.
- Nanoparticles are smaller than fine particles (PM2.5), which have diameters between 100 and 2500 nm (1×10^{-7} m and 2.5×10^{-6} m).
- **Coarse particles (PM10)** have diameters between 1×10^{-5} m and 2.5×10^{-6} m, and are often referred to as **dust**.

Compare these with an **atom**, which has a radius of about **0.1 nm** (so a diameter of about 0.2 nm): a nanoparticle is roughly 5 to 500 atoms across.

- As the side of a cube decreases by a factor of 10, the surface area to volume ratio increases by a factor of 10.
- Nanoparticles may have properties different from those of the same materials in bulk because of their high surface area to volume ratio. It may also mean that smaller quantities are needed to be effective than for materials with normal particle sizes.

2.4.2 Uses of nanoparticles CHEMISTRY ONLY

CHEMISTRY ONLY Nanoparticles have many applications:

- in **medicine** — eg delivering drugs to specific cells
- in **electronics** — smaller, faster components
- in cosmetics and sun creams — better coverage and UV protection, without white marks
- as **deodorants** — silver nanoparticles are antibacterial
- as **catalysts** — a huge surface area to volume ratio means a small mass is very effective.

New applications for nanoparticulate materials are an important area of research. You should be able to evaluate the use of nanoparticles for a specified purpose given appropriate information, weighing the advantages (better performance, less material needed) against the disadvantages, and explain that there are possible risks.

Possible risks. Nanoparticles are new, so long-term effects on health are not fully tested. They are small enough to be breathed in and to pass into cells and the bloodstream, may catalyse unwanted reactions in the body, and may build up in the environment once washed into water systems.

HIGHER TIER

NOTE

NOT REQUIRED BY THE SPECIFICATION

- Knowledge of the structures of specific ionic compounds other than sodium chloride is not required.
- For nanoparticles, you do not need to know specific examples or properties other than those specified — the categories above (medicine, electronics, cosmetics and sun creams, deodorants, catalysts) are all you must recall. Any particular product or property will be given in the question for you to evaluate.
- **HT ONLY** The limitations of the simple particle model (no forces, all particles spheres, spheres solid) are HT only.