

Structure and bonding: ionic, covalent and metallic (O Level Chemistry notes)

These notes cover the bonding and structure topic for Cambridge O Level Chemistry (5070) and Cambridge IGCSE Chemistry (0620). You need to be able to describe how ionic, covalent and metallic bonds form, draw dot-and-cross diagrams for simple compounds, describe the main giant structures, and use the type of structure to explain melting point, electrical conductivity and solubility. Diagrams in this topic are described in words and tables here; practise drawing each one on paper as you read.

1. Why atoms bond

1.1 Stable electron arrangements

The noble gases (Group 18, also called Group 0 or Group VIII) are very unreactive. Their atoms have a full outer shell of electrons: helium has 2, and neon and argon have 8 in their outer shell. Atoms of other elements react so that they end up with the same electronic configuration as a noble gas. They can do this in three ways:

- **transferring electrons** from a metal atom to a non-metal atom, which forms ions (ionic bonding);
- **sharing pairs of electrons** between non-metal atoms (covalent bonding);
- **releasing outer electrons** into a shared "sea" in a metal (metallic bonding).

1.2 Electronic configurations you need

For the first 20 elements, shells fill in the order 2, 8, 8, 2. The number of outer-shell electrons equals the group number for Groups I to VII.

Element	Proton number	Configuration	Outer electrons	Usual change
Hydrogen, H	1	1	1	shares 1
Carbon, C	6	2,4	4	shares 4
Nitrogen, N	7	2,5	5	shares 3 or gains 3
Oxygen, O	8	2,6	6	shares 2 or gains 2
Fluorine, F	9	2,7	7	shares 1 or gains 1
Sodium, Na	11	2,8,1	1	loses 1
Magnesium, Mg	12	2,8,2	2	loses 2
Aluminium, Al	13	2,8,3	3	loses 3
Chlorine, Cl	17	2,8,7	7	shares 1 or gains 1



Calcium, Ca	20	2,8,8,2	2	loses 2
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Exam tip

Metals lose electrons to form positive ions. Non-metals gain electrons to form negative ions, or share electrons with other non-metals. A compound of a metal and a non-metal is usually ionic; a compound of two non-metals is covalent.

2. Ionic bonding

2.1 Forming ions

An **ion** is an atom or group of atoms that carries an electric charge because the number of electrons is not equal to the number of protons. When a sodium atom (2,8,1) loses its outer electron it becomes a sodium ion, Na^+ , with configuration 2,8, the same as neon. When a chlorine atom (2,8,7) gains one electron it becomes a chloride ion, Cl^- , with configuration 2,8,8, the same as argon. The protons in the nucleus do not change, so the charge on the ion is the difference between protons and electrons.

Learn word for word: ionic bond

An ionic bond is the strong electrostatic attraction between oppositely charged ions.

Ionic bonding happens between a metal and a non-metal. The metal atom transfers its outer electrons to the non-metal atom. The total number of electrons lost by the metal atoms must equal the total number gained by the non-metal atoms, which is why the formula of an ionic compound has no overall charge.

2.2 Common ions

Positive ions	Negative ions
Na^+ , K^+ , Li^+ (Group I)	F^- , Cl^- , Br^- , I^- (Group VII, the halide ions)
Mg^{2+} , Ca^{2+} (Group II)	O^{2-} , S^{2-} (Group VI)
Al^{3+} (Group III)	N^{3-} (Group V)
NH_4^+ (ammonium)	OH^- , NO_3^- , SO_4^{2-} , CO_3^{2-}

2.3 Drawing ionic dot-and-cross diagrams

In a dot-and-cross diagram, the electrons of one atom are shown as dots and those of the other atom as crosses, so you can see where each electron came from. For ionic compounds, the examiner usually wants the **ions** drawn, not the atoms.

Exam tip: rules for an ionic dot-and-cross diagram

1. Draw each ion separately inside square brackets.
2. Write the charge at the top right outside the bracket, for example $[\]^+$ or $[\]^{2-}$.
3. Show the outer shell of each ion. The metal ion usually has an empty outer shell after transfer, so show its new outer shell (the full shell below).
4. The transferred electron(s) on the non-metal ion are drawn with the metal's symbol (crosses if

the metal used crosses).

5. Show the right number of each ion, for example two Cl^- ions for one Mg^{2+} .

Unless the question asks for all shells, outer shells only are accepted.

Worked example 1: sodium chloride, NaCl

Sodium atom 2,8,1 (electrons as crosses). Chlorine atom 2,8,7 (electrons as dots).

Sodium transfers its one outer electron to chlorine.

Sodium ion: $[\text{Na}]^+$, configuration 2,8. Its outer shell is now the second shell with 8 crosses.

Chloride ion: $[\text{Cl}]^-$, configuration 2,8,8. The outer shell holds 7 dots and 1 cross.

Ratio 1 : 1, so the formula is NaCl.

Worked example 2: magnesium oxide, MgO

Magnesium atom 2,8,2 transfers two electrons to an oxygen atom 2,6.

Ions: $[\text{Mg}]^{2+}$ (2,8) and $[\text{O}]^{2-}$ (2,8). The oxide ion outer shell holds 6 dots and 2 crosses.

Ratio 1 : 1, formula MgO. The ions carry double charges, so the attraction is stronger than in NaCl.

Worked example 3: calcium chloride, CaCl_2

Calcium (2,8,8,2) must lose 2 electrons, but each chlorine (2,8,7) can accept only 1.

So one calcium atom gives one electron to each of two chlorine atoms.

Ions: one $[\text{Ca}]^{2+}$ (2,8,8) and two separate $[\text{Cl}]^-$ ions (2,8,8), each with 7 dots and 1 cross in the outer shell.

Formula CaCl_2 .

Other examples you should be able to draw: lithium fluoride LiF, potassium oxide K_2O (two K^+ and one O^{2-}), magnesium chloride MgCl_2 , and aluminium oxide Al_2O_3 (two Al atoms lose 6 electrons in total to form two Al^{3+} , and three O atoms gain 6 to form three O^{2-}).

2.4 Practice: ionic dot-and-cross

Draw the dot-and-cross diagram for each compound, outer shells only, and check it against the description in the answers.

1. Lithium fluoride, LiF
2. Potassium oxide, K_2O
3. Magnesium chloride, MgCl_2

Practice answers

1. $[\text{Li}]^+$ with 2 crosses (configuration 2, like helium) and $[\text{F}]^-$ with 7 dots and 1 cross (2,8). Ratio 1 : 1.
2. Two separate $[\text{K}]^+$ ions, each 2,8,8 with 8 crosses shown, and one $[\text{O}]^{2-}$ with 6 dots and 2 crosses (one cross from each potassium). Ratio 2 : 1.
3. One $[\text{Mg}]^{2+}$ (2,8, 8 crosses) and two $[\text{Cl}]^-$, each with 7 dots and 1 cross. Ratio 1 : 2.

3. Giant ionic lattices and their properties

3.1 The lattice structure

Ionic compounds do not exist as small separate molecules. The ions are packed together in a regular, repeating three-dimensional arrangement called a **giant ionic lattice**. In sodium chloride each Na^+ ion is surrounded by 6 Cl^- ions and each Cl^- ion by 6 Na^+ ions, giving a cubic structure. The electrostatic attractions act in all directions between every ion and its oppositely charged neighbours.

Learn word for word: giant ionic lattice

A giant ionic lattice is a regular, repeating three-dimensional arrangement of alternating positive and negative ions held together by strong electrostatic attraction.

Because the formula only shows the ratio of ions, NaCl is called a **formula unit**, not a molecule.

3.2 Properties of ionic compounds

Property	Explanation
High melting and boiling points; solid at room temperature	Many strong electrostatic attractions between oppositely charged ions in a giant lattice. A lot of energy is needed to overcome them.
Do not conduct electricity when solid	Ions are held in fixed positions in the lattice and cannot move.
Conduct electricity when molten or in aqueous solution	The lattice breaks down and the ions are free to move and carry charge. (Electrolysis then decomposes the compound.)
Often soluble in water, insoluble in organic solvents	Water molecules are attracted to the ions and separate them from the lattice. Not all ionic compounds are soluble, for example calcium carbonate and silver chloride.
Hard but brittle	A blow can shift a layer so that ions of the same charge line up and repel, and the crystal splits.

Worked example 4: comparing melting points

Magnesium oxide melts at a much higher temperature than sodium chloride. Explain why. Both are giant ionic lattices. Mg^{2+} and O^{2-} carry charges of 2, while Na^+ and Cl^- carry charges of 1. The attraction between doubly charged ions is stronger, so more energy is needed to overcome the electrostatic attractions in MgO .

Exam tip

Say "ions are free to move" for conduction in ionic compounds. Never say "electrons move" for a molten or dissolved ionic compound; that loses the mark.

4. Covalent bonding in simple molecules

4.1 The covalent bond

When two non-metal atoms bond, neither can easily lose electrons, so they share pairs of electrons. Each shared pair is attracted to the nuclei of both atoms, and this holds the atoms together. Each atom gains a share of enough electrons to reach a noble gas configuration.

Learn word for word: covalent bond

A covalent bond is formed when a pair of electrons is shared between two atoms, leading to noble gas electronic configurations.

One shared pair is a **single bond** (drawn as one line, C-H). Two shared pairs form a **double bond** (C=O). Three shared pairs form a **triple bond** (N≡N). A pair of outer electrons that is not shared is a **lone pair**.

The number of covalent bonds an atom usually forms is the number of electrons it needs to fill its outer shell: H forms 1, Cl 1, O 2, N 3, C 4.

4.2 Drawing covalent dot-and-cross diagrams

Exam tip: rules for a covalent dot-and-cross diagram

1. Draw overlapping circles for the outer shells of bonded atoms (or no circles, but keep pairs clear).
2. Put the shared pairs in the overlap: one dot and one cross per pair.
3. Show all the non-bonding outer electrons (lone pairs) too.
4. Count: every atom except hydrogen should have 8 outer electrons; hydrogen has 2.
5. No brackets and no charges for neutral molecules.

The table below describes the molecules on the syllabus. "Dots" belong to the first atom listed; "crosses" to the other atom(s).

Molecule	Shared pairs	Lone pairs	Description of the diagram
Hydrogen, H ₂	1 (single bond)	none	One dot and one cross in the overlap. Each H has 2 electrons.
Chlorine, Cl ₂	1	3 on each Cl	One dot-cross pair shared; each Cl also has 6 of its own electrons as 3 lone pairs, total 8.
Hydrogen chloride, HCl	1	3 on Cl	One shared pair; Cl has 3 lone pairs.
Water, H ₂ O	2	2 on O	O in the centre sharing one pair with each H; O keeps 2 lone pairs.
Ammonia, NH ₃	3	1 on N	N shares one pair with each of three H atoms and keeps 1 lone pair.
Methane, CH ₄	4	none	C shares one pair with each of four H atoms.
Oxygen, O ₂	2 (double bond)	2 on each O	Two dot-cross pairs in the overlap; each O has 2 lone pairs.
Carbon dioxide, CO ₂	4 (two double bonds)	2 on each O	C in the centre; two shared pairs with each O; each O keeps 2 lone pairs. C has no lone pairs.
Nitrogen, N ₂	3 (triple bond)	1 on each N	Three dot-cross pairs in the overlap; each N has 1 lone pair.

Ethene, C ₂ H ₄	6 (one C=C, four C-H)	none	Two pairs between the carbons; each C also shares one pair with each of two H atoms.
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Worked example 5: checking a diagram for ammonia

Nitrogen is 2,5, so it has 5 outer electrons (dots). Each hydrogen has 1 electron (cross).

Nitrogen needs 3 more electrons, so it forms 3 shared pairs, one with each H.

Count around N: 3 shared pairs (6 electrons) + 1 lone pair (2 electrons) = 8. Count around each H: 2. Correct.

Worked example 6: carbon dioxide

Carbon (2,4) needs 4 more electrons; each oxygen (2,6) needs 2 more.

Carbon shares 2 pairs with each oxygen, forming two C=O double bonds.

Count around C: 4 pairs = 8. Count around each O: 2 shared pairs + 2 lone pairs = 8.

4.3 Practice: covalent dot-and-cross

1. Hydrogen chloride, HCl
2. Methane, CH₄
3. Nitrogen, N₂
4. Ethene, C₂H₄

Practice answers

1. H and Cl share one dot-cross pair. Cl also shows 3 lone pairs (6 of its own electrons). H has 2 outer electrons, Cl has 8.
2. C in the centre with four overlaps, each containing one dot (C) and one cross (H). No lone pairs. C has 8, each H has 2.
3. Three dot-cross pairs in the overlap and one lone pair on each N. Each N has $3 \times 2 + 2 = 8$.
4. Two dot-cross pairs between the two carbons. Each carbon shares one pair with each of its two hydrogens. Each C has $2 + 2 + 2 + 2 = 8$ electrons; each H has 2.

4.4 Simple molecular structures and their properties

Substances made of small covalent molecules, such as water, methane, carbon dioxide, iodine and ethanol, are **simple molecular** substances. There are two kinds of force to think about:

- **covalent bonds** inside each molecule, which are strong;
- **intermolecular forces** between separate molecules, which are weak.

When a simple molecular substance melts or boils, only the weak intermolecular forces are overcome. The covalent bonds do not break. That is the key idea for every property below.

Property	Explanation
Low melting and boiling points; many are gases or liquids at room temperature	Weak intermolecular forces need little energy to overcome.

Melting and boiling points rise with molecular size	Larger molecules have stronger intermolecular forces, for example the boiling points increase from methane to ethane to propane, and from chlorine to bromine to iodine.
Do not conduct electricity (solid, liquid or most solutions)	The molecules have no overall charge and there are no free electrons or ions to carry a current.
Usually insoluble in water, often soluble in organic solvents	Many molecules are not attracted strongly to water molecules. There are exceptions: ethanol, ammonia and hydrogen chloride dissolve well. HCl and NH ₃ react with water to form ions, so their solutions conduct.

Exam tip

"Methane has a low boiling point because its covalent bonds are weak" is wrong and scores zero. The C-H bonds are strong. Write: "there are weak forces of attraction between the molecules, which need little energy to overcome".

5. Giant covalent structures

In a giant covalent (macromolecular) structure, a very large number of atoms are joined by covalent bonds in a continuous network throughout the whole structure. There are no separate molecules, so to melt the substance many strong covalent bonds must be broken.

Learn word for word: giant covalent structure

A giant covalent structure is a three-dimensional network of a very large number of atoms in which every atom is joined to its neighbours by strong covalent bonds.

5.1 Diamond

- A form (allotrope) of carbon.
- Each carbon atom forms **four** covalent bonds to four other carbon atoms, arranged tetrahedrally, in a rigid three-dimensional network.
- Very high melting point: many strong covalent bonds must be broken.
- Very hard: the rigid network of strong bonds in every direction resists scratching. Used in cutting tools and drill tips.
- Does not conduct electricity: all four outer electrons of each carbon are used in bonding, so there are no free electrons.
- Insoluble in water and organic solvents.

5.2 Graphite

- Also an allotrope of carbon.
- Each carbon atom forms **three** covalent bonds to three other carbon atoms, making flat layers of hexagonal rings.
- The fourth outer electron of each carbon is **delocalised**: it is free to move along the layers.
- The layers are held together by weak forces, so they can slide over each other.
- Very high melting point: strong covalent bonds within the layers must be broken.

- Soft and slippery: layers slide, so it is used as a lubricant and in pencil "leads".
- Conducts electricity: delocalised electrons move along the layers and carry charge. Used for electrodes in electrolysis.

5.3 Silicon(IV) oxide, SiO₂

- Found as sand and quartz.
- Each silicon atom is covalently bonded to **four** oxygen atoms, and each oxygen atom to **two** silicon atoms, giving a ratio of 1 : 2 and the formula SiO₂.
- The structure is a three-dimensional tetrahedral network similar to diamond.
- Very high melting point, hard, insoluble in water, and does not conduct electricity (no free electrons or ions).

Worked example 7: diamond and graphite

Explain why graphite conducts electricity but diamond does not.

In graphite each carbon atom forms only three covalent bonds, so one outer electron per atom is delocalised and can move through the layers, carrying charge. In diamond each carbon atom forms four covalent bonds, so all outer electrons are held in bonds and none are free to move.

Worked example 8: similar properties

Suggest why silicon(IV) oxide and diamond have similar physical properties.

Both are giant covalent structures with each atom bonded tetrahedrally in a 3D network. Both have very many strong covalent bonds that need a lot of energy to break (high melting point, hard), and neither has free electrons or ions (no conduction).

6. Metallic bonding

6.1 The structure of a metal

In a solid metal, the atoms are packed closely in a regular lattice. Each atom releases its outer-shell electrons, which become **delocalised**, meaning they are not attached to any one atom and are free to move throughout the structure. The atoms become positive ions. The metal is held together by the attraction between the positive ions and the "sea" of delocalised electrons around them.

Learn word for word: metallic bonding

Metallic bonding is the electrostatic attraction between a lattice of positive metal ions and a 'sea' of delocalised electrons.

6.2 Properties of metals

Property	Explanation
Good electrical conductors (solid and liquid)	Delocalised electrons move through the structure and carry the current.
Good thermal conductors	Delocalised electrons transfer energy quickly through the metal.

Malleable (can be hammered into shape) and ductile (can be drawn into wires)	Layers of positive ions can slide over each other without breaking the metallic bond, because the delocalised electrons move with them and keep holding the ions together.
Usually high melting and boiling points	Strong attraction between the positive ions and delocalised electrons needs a lot of energy to overcome. (Group I metals are exceptions with fairly low melting points.)
Insoluble in water	The metallic lattice is not broken up by water (some reactive metals react with water instead of dissolving).

Worked example 9: why magnesium melts higher than sodium

Magnesium releases two electrons per atom and forms Mg^{2+} ions; sodium releases one and forms Na^+ ions. Magnesium therefore has more delocalised electrons and more highly charged ions, so the metallic bonding is stronger and more energy is needed to melt it.

6.3 Alloys

An alloy is a mixture of a metal with other elements, usually other metals or carbon. The different-sized atoms disrupt the regular layers, making it harder for layers to slide, so alloys are usually harder and stronger than the pure metal. Examples: brass (copper and zinc) and steel (iron with carbon).

7. How structure explains properties

7.1 Summary table

Structure	Particles	Forces broken on melting	Melting point	Conducts when solid?
Giant ionic	positive and negative ions	strong ionic bonds	high	no
Simple molecular	small molecules	weak intermolecular forces	low	no
Giant covalent	atoms	strong covalent bonds	very high	no (graphite yes)
Metallic	positive ions and delocalised electrons	strong metallic bonds	usually high	yes

Structure	Conducts when molten/aqueous?	Soluble in water?
Giant ionic	yes (ions move)	often
Simple molecular	no (usually)	usually not
Giant covalent	not applicable	no
Metallic	yes (liquid metal)	no

7.2 Identifying the structure from data

Exam tip: a four-step method

1. Low melting point (below about 200 °C)? It is simple molecular.
2. High melting point and conducts as a solid? It is a metal (or graphite).
3. High melting point, does not conduct as a solid but conducts when molten or dissolved? It is ionic.
4. Very high melting point, never conducts, insoluble? It is giant covalent.

Worked example 10: using data

Substance X melts at 801 °C. Solid X does not conduct, but an aqueous solution of X does.

Substance Y melts at -23 °C and does not conduct in any state. Identify the structure of each.

X has a high melting point and conducts only when its ions can move: giant ionic.

Y has a low melting point and never conducts: simple molecular.

7.3 Writing a full explanation

Name the structure, name the particles or bonds, and link to energy or moving charge. For example: "Sodium chloride has a giant ionic lattice. There are strong electrostatic attractions between many oppositely charged ions. A lot of energy is needed to overcome these, so the melting point is high."

Quick check

Answer all 15 questions. Marks are shown in brackets.

1. State what is meant by an ionic bond. [1]
2. A magnesium atom reacts with a fluorine atom. (a) State the electronic configurations of a magnesium ion and a fluoride ion. (b) Deduce the formula of magnesium fluoride. [3]
3. Describe, in words, the dot-and-cross diagram for aluminium oxide, Al_2O_3 , showing outer shells only. [3]
4. Explain why solid potassium bromide does not conduct electricity but molten potassium bromide does. [2]
5. State what is meant by a covalent bond. [1]
6. Draw (or describe) the dot-and-cross diagram for a water molecule and state the number of lone pairs on the oxygen atom. [3]
7. A molecule of hydrogen sulfide, H_2S , forms in the same way as water. Sulfur is in Group VI. Deduce the number of shared pairs and lone pairs around the sulfur atom. [2]
8. Methane boils at -162 °C. Explain why the boiling point is low. [2]
9. Describe the structure of diamond and use it to explain why diamond is used in drill tips. [3]
10. Explain why graphite is used as a lubricant and as an electrode. [4]
11. Describe the structure of silicon(IV) oxide and explain why it has a high melting point. [3]
12. Describe metallic bonding and use it to explain why copper is ductile. [3]
13. Explain why magnesium oxide has a higher melting point than sodium chloride. [2]
14. The table gives data for four substances.

Substance	Melting point / °C	Conducts as solid	Conducts when molten

A	1083	yes	yes
B	-95	no	no
C	714	no	yes
D	1710	no	no

Identify the type of structure in each substance and give a reason for each. [4]

15. A student says: "When iodine sublimes, the covalent bonds between the iodine atoms break."
Evaluate this statement. [2]

Answers

- The strong electrostatic attraction between oppositely charged ions. [1]
- (a) Magnesium (2,8,2) loses 2 electrons: Mg^{2+} is 2,8. Fluorine (2,7) gains 1 electron: F^- is 2,8. [2]
(b) Mg^{2+} needs two F^- to balance the charge (+2 and 2×-1), so the formula is MgF_2 . [1]
- Each aluminium atom (2,8,3) loses 3 electrons; each oxygen atom (2,6) gains 2. To balance, 2 Al lose 6 electrons and 3 O gain 6. [1]
Diagram: two separate $[\text{Al}]^{3+}$ ions in brackets, each showing an outer shell of 8 crosses (configuration 2,8). [1]
Three separate $[\text{O}]^{2-}$ ions in brackets, each with 6 dots and 2 crosses in the outer shell, charges written outside the brackets. [1]
- In the solid the ions are held in fixed positions in the lattice and cannot move. [1] When molten the lattice breaks down and the ions are free to move and carry charge. [1]
- A pair of electrons shared between two atoms (giving each atom a noble gas configuration). [1]
- Oxygen in the centre with two overlapping circles, one with each hydrogen. [1] Each overlap contains one dot (O) and one cross (H). [1] Oxygen also has 4 non-bonding electrons, which are 2 lone pairs, giving 8 in total. [1]
- Sulfur is in Group VI, so it has 6 outer electrons and needs 2 more. [1] It forms 2 shared pairs (one with each H) and keeps 2 lone pairs, total 8. [1]
- Methane is a simple molecular substance. [1] The forces between methane molecules are weak, so little energy is needed to overcome them (the C-H covalent bonds do not break on boiling). [1]
- Diamond is a giant covalent structure in which each carbon atom is bonded covalently to four other carbon atoms in a tetrahedral, three-dimensional network. [2] There are very many strong covalent bonds in every direction, so diamond is extremely hard and does not wear away when cutting through rock. [1]

10. Lubricant: graphite has layers of carbon atoms with weak forces between the layers [1], so the layers can slide over each other easily. [1]

Electrode: each carbon forms only three bonds, so one electron per atom is delocalised [1] and these electrons move along the layers and carry current. (It also has a high melting point and is fairly unreactive.) [1]

11. A giant covalent network in which each silicon atom is bonded to four oxygen atoms and each oxygen atom to two silicon atoms. [2] Melting requires breaking very many strong covalent bonds, which needs a lot of energy. [1]

12. Metallic bonding is the attraction between a lattice of positive ions and a sea of delocalised electrons. [1] In copper the layers of ions can slide over one another when a force is applied [1], and the delocalised electrons still hold the ions together in their new positions, so the metal can be drawn into a wire without breaking. [1]

13. Both are giant ionic lattices. [1] Mg^{2+} and O^{2-} have charges of 2, larger than the charges of 1 on Na^+ and Cl^- , so the electrostatic attraction is stronger and more energy is needed to overcome it. [1]

14. A: metallic. High melting point and conducts when solid because it has delocalised electrons. [1]

B: simple molecular. Low melting point (weak intermolecular forces) and no charged particles to conduct. [1]

C: giant ionic. High melting point; conducts only when molten because only then are the ions free to move. [1]

D: giant covalent. Very high melting point (many strong covalent bonds) and never conducts because there are no free electrons or ions. [1]

15. The statement is incorrect. [1] Iodine is a simple molecular substance made of I_2 molecules. When it sublimes, only the weak forces between the molecules are overcome; the covalent bond inside each I_2 molecule stays intact, so the vapour is still I_2 molecules. [1]