IGCSE CHEMISTRY 0620 · CHAPTER 2 NOTES · ~18 MIN READ

Atoms, elements, compounds & bonding

Inside the atom, how electrons are arranged, and why atoms join up — ionic, covalent and metallic bonding — with the structures and properties that follow from each.

Atomic structure | Isotopes | Electronic configuration | Ionic bonding | Covalent bonding | Giant structures | Metallic bonding

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No tag = **Core** (everyone) EXTENDED Extended only highlight = comes up again and again

Test yourself at the end

IN 30 SECONDS

- **Elements** contain one type of atom. **Compounds** are different atoms **chemically bonded**. **Mixtures** are not bonded at all.
- An atom has a tiny **nucleus** (protons + neutrons) with **electrons** in shells around it. Electrons decide the chemistry.
- Atoms bond to get a **full outer shell**. Metals **give** electrons → **ionic**. Non-metals **share** → **covalent**. Metals also **pool** electrons → **metallic**.
- The type of bonding decides the **structure**, and the structure decides the **properties**. EXT

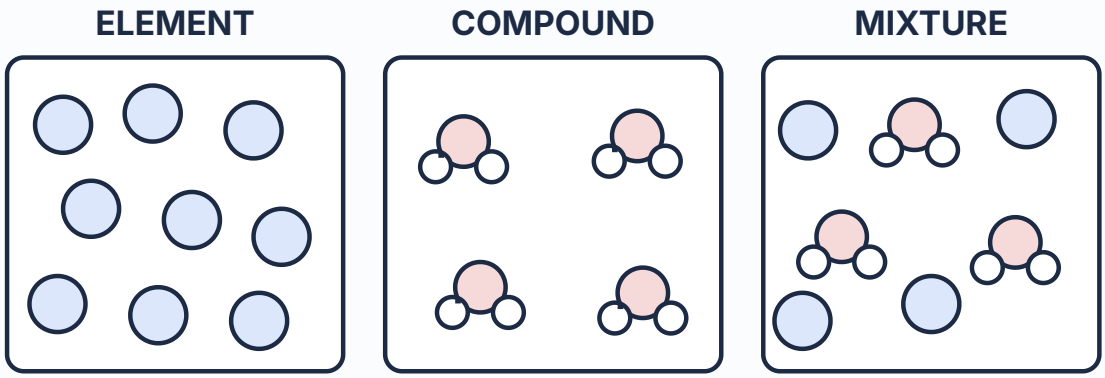
BY THE END YOU CAN

- Tell elements, compounds and mixtures apart.
- Describe an atom, and work out protons, neutrons and electrons from the nuclide symbol.
- Write electronic configurations for the first 20 elements and link them to groups and periods.
- Define isotopes, and calculate relative atomic mass from isotope data. EXT
- Draw dot-and-cross diagrams for ionic and covalent compounds.
- Link structure and bonding to melting point, conductivity and uses.

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1 Elements, compounds and mixtures

Three kinds of substance



one kind of atom (argon, Ar) | different atoms chemically joined (water, H_2O) | different substances not joined (argon + water)

compound: fixed ratio, new properties · mixture: any proportions, easy to separate

Left: in a compound the atoms are bonded in a fixed ratio. Right: in a mixture the particles are just jumbled together.

	Element	Compound	Mixture
made of	one type of atom	two or more elements chemically bonded	two or more substances not bonded
formula	fixed (Fe, O_2)	fixed (H_2O, NaCl)	no formula — amounts can vary
separation	cannot be broken down chemically	only by a chemical reaction	by physical methods (filtering, distilling …)
properties	those of the element	very different from the elements in it	properties of each part are kept
example	iron, oxygen	water, salt	air, sea water, brass

- **Iron + sulfur:** a mixture can be separated with a magnet (iron is magnetic). Heat the mixture strongly and it reacts to form the compound **iron(II) sulfide**, which is not magnetic and has new properties.

PICTURE IT

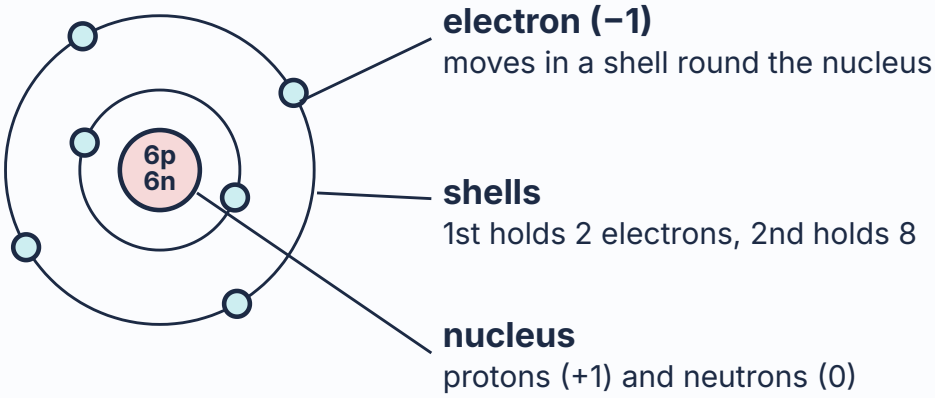
A compound is a cake — you cannot pick the flour back out. A mixture is a fruit salad — you can pick out each fruit.

EXAM TIP

Compounds have a **fixed ratio** (water is always H_2O). Mixtures have **no fixed ratio** (air can have more or less water vapour).

2 Inside the atom

Protons, neutrons, electrons



carbon atom: 6 protons, 6 electrons → electronic configuration 2,4

An atom. The nucleus holds the protons and neutrons. The electrons move in shells outside it.

Particle	Relative charge	Relative mass	Where
proton	+1	1	nucleus
neutron	0	1	nucleus
electron	−1	1/1836 (almost nothing)	shells

- An atom is **neutral**: the number of protons equals the number of electrons.
- Almost all the mass is in the **nucleus** — it is tiny but very dense. Almost all the volume of the atom is empty space.

Proton number (atomic number, Z) — the number of protons in the nucleus. It identifies the element.

Nucleon number (mass number, A) — the total number of protons and neutrons in the nucleus.

$$\text{neutrons} = A - Z$$

nucleon number minus proton number

- The nuclide symbol shows both numbers: ${}^{23}_{11}\text{Na}$ has **A = 23** (top) and **Z = 11** (bottom).

WORKED EXAMPLE

How many protons, neutrons and electrons are in an atom of ${}^{27}_{13}\text{Al}$?

1. protons = Z = **13**
2. electrons = protons = **13** (the atom is neutral)
3. **neutrons = A − Z = 27 − 13 = 14**

WORKED EXAMPLE

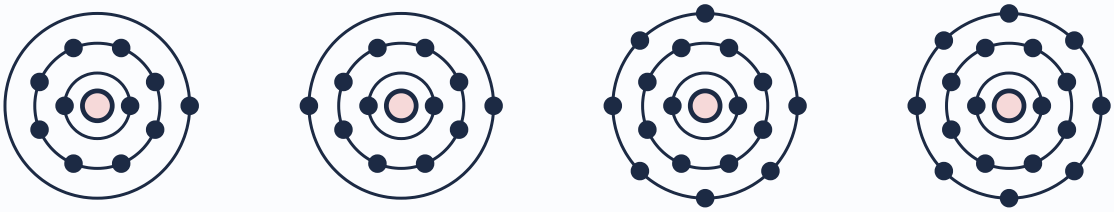
EXT **How many protons, neutrons and electrons are in the ion ${}^{24}_{12}\text{Mg}^{2+}$?**

1. protons = Z = **12**
2. neutrons = 24 − 12 = **12**
3. **the ion has a 2+ charge, so it has lost 2 electrons: 12 − 2 = 10 electrons**

COMMON MISTAKE

Mixing up which number is on top. The **bigger** number (A) is on top.

Electronic configuration



sodium
2,8,1
Group I
Period 3

magnesium
2,8,2
Group II
Period 3

chlorine
2,8,7
Group VII
Period 3

argon
2,8,8
Group VIII
Period 3

outer electrons = group number · number of shells = period number

Electrons fill the shells from the inside outwards. Sodium, magnesium, chlorine and argon are shown — all in Period 3, so all have 3 shells.

Filling the shells

- Electrons sit in **shells** (energy levels). Fill them from the **inside out**:
- shell 1 holds up to **2** electrons; shell 2 holds up to **8**; shell 3 holds up to **8** (for the first 20 elements).
- Write it as numbers separated by commas: sodium (11 electrons) = **2,8,1**; chlorine (17) = **2,8,7**.
- The electrons in the **outer shell** decide how the element reacts.

Period	I	II	III	IV	V	VI	VII	VIII
1	H 1	–	–	–	–	–	–	He 2
2	Li 2,1	Be 2,2	B 2,3	C 2,4	N 2,5	O 2,6	F 2,7	Ne 2,8
3	Na 2,8,1	Mg 2,8,2	Al 2,8,3	Si 2,8,4	P 2,8,5	S 2,8,6	Cl 2,8,7	Ar 2,8,8
4	K 2,8,8,1	Ca 2,8,8,2	–	–	–	–	–	–

Groups, periods and noble gases

- **Group number = number of outer-shell electrons** (Groups I to VII). Group VIII (also called Group 0) has a **full** outer shell.
- **Period number = number of shells** that contain electrons.

Noble gases — the Group VIII elements (He, Ne, Ar ...). They have a full outer shell, so they are very unreactive. Atoms of other elements bond to reach this stable arrangement.

PICTURE IT

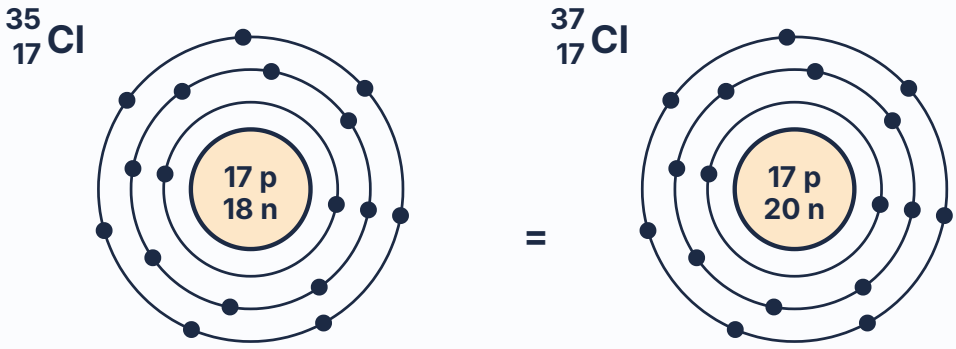
Sodium (2,8,1) has one lonely electron in its outer shell, so it is easy to lose. Chlorine (2,8,7) is one short of a full set, so it is eager to gain one. Put them together and both get what they want.

EXAM TIP

Sodium is in **Group I** (1 outer electron) and **Period 3** (3 shells) — you can read both straight from 2,8,1.

Isotopes

Isotopes — atoms of the same element with the same number of protons but different numbers of neutrons.



chlorine-35
17 protons, 18 neutrons
17 electrons (2,8,7)

chlorine-37
17 protons, 20 neutrons
17 electrons (2,8,7)

same electrons = same chemistry · different neutrons = different mass

Two isotopes of chlorine. Both have 17 protons and 17 electrons. Chlorine-35 has 18 neutrons; chlorine-37 has 20.

- Isotopes have the **same proton number** but **different nucleon numbers**. Examples: carbon-12 and carbon-14; chlorine-35 and chlorine-37; hydrogen-1, hydrogen-2 and hydrogen-3.
- EXT Isotopes of an element have the **same chemical properties**, because chemistry depends on electrons and they have the same number. They have different **physical** properties that depend on mass, for example density.

EXT $A_r = \frac{\Sigma \text{ (isotope mass} \times \% \text{ abundance)}}{100}$

relative atomic mass is the weighted average of the isotope masses

WORKED EXAMPLE

EXT **Chlorine is 75% ^{35}Cl and 25% ^{37}Cl. Calculate its relative atomic mass.**

1. $A_r = (75 \times 35 + 25 \times 37) \div 100$
2. $= (2625 + 925) \div 100$
3. **$= 3550 \div 100 = 35.5$**

WORKED EXAMPLE

EXT **Magnesium is 79% ^{24}Mg, 10% ^{25}Mg and 11% ^{26}Mg. Calculate its relative atomic mass to 1 decimal place.**

1. $(79 \times 24) + (10 \times 25) + (11 \times 26) = 1896 + 250 + 286 = 2432$
2. $2432 \div 100 = 24.32$
3. **$A_r = 24.3$**

EXAM TIP

Always check that your answer lies **between** the lightest and heaviest isotope, and closer to the one that is most abundant. That is why chlorine is 35.5, not 36.

3 Ions and ionic bonding

How ions form

Ion — a charged particle formed when an atom (or group of atoms) loses or gains electrons.

Cation — a positive ion, formed when an atom **loses** electrons. Metals form cations.

Anion — a negative ion, formed when an atom **gains** electrons. Non-metals form anions.

- Atoms lose or gain electrons to get a **full outer shell**, like a noble gas.
- **Group I** atoms lose 1 electron (+1); **Group II** lose 2 (2+); **Group III** lose 3 (3+). **Group VII** gain 1 (−1); **Group VI** gain 2 (2−); **Group V** gain 3 (3−).

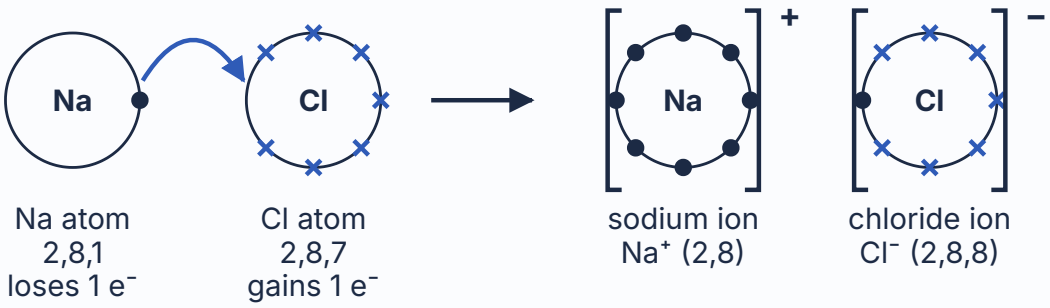
PICTURE IT

A sodium atom has one outer electron it does not need. A chlorine atom is one electron short. Sodium hands over the electron and both end up with a full outer shell — sodium has a + charge and chlorine has a − charge, and they stick together.

Ionic bond — the strong electrostatic attraction between oppositely charged ions.

Dot-and-cross diagrams for ionic compounds

● electron from Na × electron from Cl



oppositely charged ions attract each other
this electrostatic attraction is the ionic bond

Sodium chloride. Sodium loses one electron to chlorine. Both ions end up with a full outer shell. Draw the outer shell only.

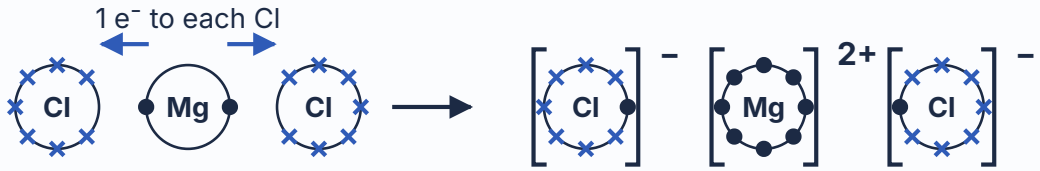
- **Step 1:** draw the outer shell of each atom. Use dots for one atom's electrons and crosses for the other's.
- **Step 2:** move the electron(s) from the metal to the non-metal.
- **Step 3:** draw each ion in **square brackets** with its **charge** at the top right.
- Na^+ has 2,8 and Cl^- has 2,8,8 — both now have a noble-gas arrangement.

Magnesium compounds

EXTENDED



magnesium oxide MgO: one Mg^{2+} and one O^{2-} (ratio 1 : 1)



magnesium chloride $MgCl_2$: one Mg^{2+} and two Cl^- (ratio 1 : 2)

Top: magnesium oxide, one Mg^{2+} and one O^{2-}. Bottom: magnesium chloride, one Mg^{2+} and two Cl^- — each chlorine takes one electron.

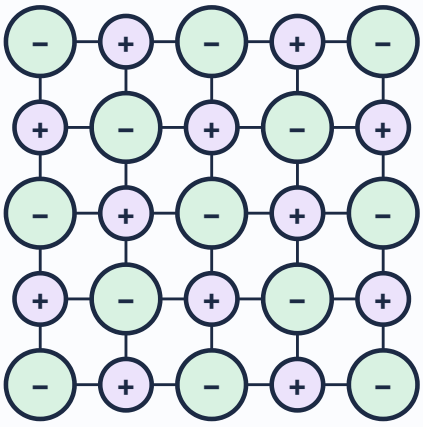
- The number of electrons lost must equal the number gained, so **the total charge is zero**.
- Mg loses 2 electrons. Each Cl can only gain 1, so you need **two** chlorine atoms: $MgCl_2$.

COMMON MISTAKE

Forgetting the brackets and the charge on each ion, or drawing the shared electrons between the ions. In ionic bonding electrons are **transferred**, not shared.

The giant ionic lattice

EXTENDED



sodium chloride, NaCl

+ Na^+ sodium ion

− Cl^- chloride ion

each ion is held by strong electrostatic attraction in every direction → giant ionic lattice

solid: ions fixed, no conduction
molten or dissolved: ions move, so it conducts

Part of a sodium chloride crystal. Every ion is attracted to all of its neighbours with the opposite charge.

- An ionic solid is not made of molecules. It is a huge, regular, **giant ionic lattice** of alternating positive and negative ions.
- Each ion is held by strong electrostatic attraction in **every direction**.

Properties of ionic compounds

Property	Why
high melting and boiling points (NaCl melts at 801 °C)	lots of energy is needed to overcome the strong attraction between ions in the whole lattice
hard but brittle	hitting the crystal shifts the layers, so like charges line up, repel and the crystal splits
do not conduct when solid	ions are held in fixed positions
conduct when molten or dissolved in water	the ions are free to move and carry the charge

- Many ionic compounds dissolve in water, which breaks the lattice apart into separate ions.

EXAM TIP

"Why can't solid sodium chloride conduct electricity?" → the **ions** cannot move. Never say "no electrons". Conduction in ionic compounds is by **ions**, not electrons.

4 Covalent bonding

Sharing electrons

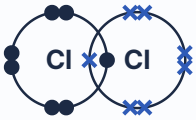
Covalent bond — a pair of electrons shared between two atoms. It forms between **non-metal** atoms.

- Each atom shares electrons so that it gets a **full outer shell**. The shared pair is attracted to the nuclei of both atoms, and this holds the atoms together.
- In a dot-and-cross diagram, the shared electrons sit in the **overlap** of the two outer shells.
- Number of bonds an atom makes: **H = 1, Cl = 1, O = 2, N = 3, C = 4**. (It is just the number of electrons it needs for a full shell.)

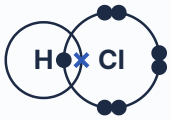
● electron from one atom × electron from the other



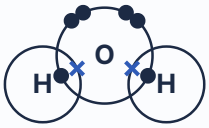
hydrogen H_2



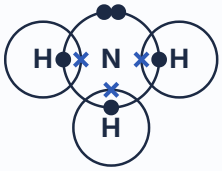
chlorine Cl_2



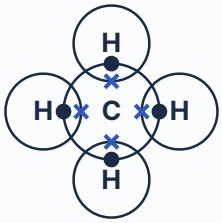
hydrogen chloride HCl



water H_2O



ammonia NH_3



methane CH_4

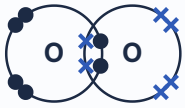
Core covalent molecules: H_2, Cl_2, HCl, H_2O, NH_3 and CH_4. Only the outer electrons are drawn. Pairs that are not shared are called lone pairs.

- Lone pairs (not shared) are drawn on the atom. Oxygen in water has **2** lone pairs; nitrogen in ammonia has **1**.

Double and triple bonds, and larger molecules

EXTENDED

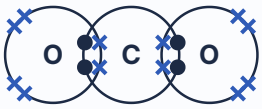
● electron from one atom × electron from the other



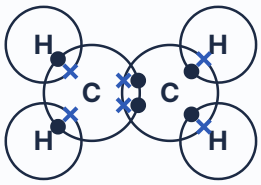
oxygen O_2 (double bond)



nitrogen N_2 (triple bond)



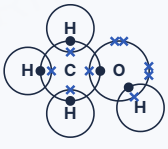
carbon dioxide CO_2



ethene C_2H_4 (C=C double bond)

O_2 has a double bond (2 shared pairs). N_2 has a triple bond (3 shared pairs). CO_2 has two double bonds. Ethene C_2H_4 has a C=C double bond.

Check your diagram



methanol, CH_3OH

C forms 4 bonds
O forms 2 bonds and has 2 lone pairs
H forms 1 bond each

● and × = electrons from different atoms

Methanol, CH_3OH. Carbon makes 4 bonds, oxygen makes 2 and keeps 2 lone pairs, each hydrogen makes 1.

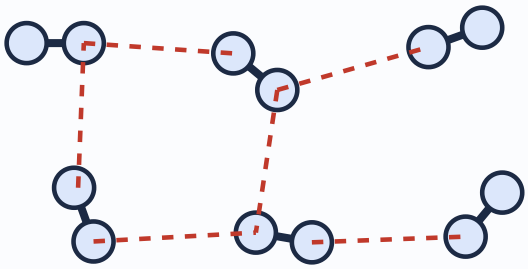
- **Check:** count the electrons around each atom. Every atom except hydrogen should have 8 (**octet**); hydrogen has 2.

COMMON MISTAKE

Forgetting the lone pairs on the outside of O and N atoms.

Properties of simple molecular substances

- **Low melting and boiling points**: many are gases or liquids at room temperature.
- **Do not conduct electricity**: there are no ions and no free electrons.
- EXT The weak forces **between** molecules (intermolecular forces) are overcome when they melt or boil. The strong covalent bonds **inside** the molecules stay intact.



covalent bond
strong; stays intact

forces between
molecules: weak

weak forces → little energy needed → low melting and boiling points

simple molecules (e.g. iodine, water, CO_2)

no free ions or electrons → do not conduct electricity

Strong covalent bonds hold the atoms inside each molecule. Only the weak forces between the molecules are overcome when the substance melts or boils.

EXAM TIP

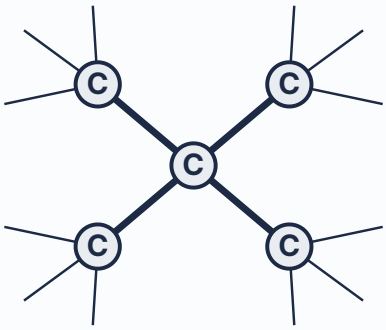
"Why does iodine have a low melting point?" → weak **forces between molecules** need little energy to overcome. **Never** say the covalent bonds break — that is the most common mark lost on this topic.

5 Giant covalent structures

Diamond, graphite and silicon dioxide

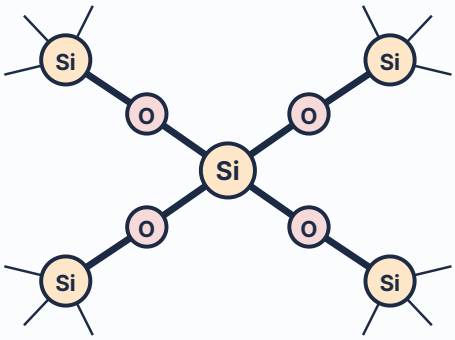
- Some covalent substances do not form small molecules. Millions of atoms join in **one giant network** held together by strong covalent bonds.

diamond (carbon)



each C bonded to 4 other C
strong bonds throughout → very hard

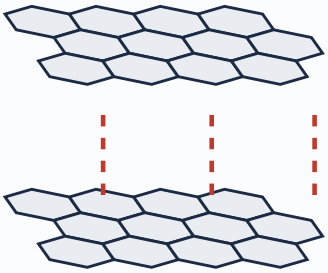
silicon dioxide (silica)



each Si bonded to 4 O, each O to 2 Si
similar structure → very hard

giant covalent: millions of strong bonds, so very high melting points

Diamond: each carbon is bonded to four others. Silicon dioxide: each silicon is bonded to four oxygen atoms, and each oxygen to two silicons. Silicon dioxide is Extended.



graphite (carbon)

strong covalent bonds within each layer (each C bonded to 3 others)

weak forces between the layers → layers slide over each other

soft and slippery: lubricant, pencils · delocalised electrons → conducts: electrodes

Graphite. Each carbon is bonded to three others in flat layers. The layers are held together only by weak forces.

	Diamond	Graphite	Silicon dioxide EXT
structure	each C bonded to 4 others	layers; each C bonded to 3 others	each Si bonded to 4 O; each O to 2 Si
melting point	very high	very high	very high
hardness	very hard	soft and slippery	very hard
conducts electricity?	no — no free electrons	**yes** — one delocalised electron per carbon	no
uses	cutting tools, drill tips	lubricant, pencils, electrodes	glass, sand, abrasives

- Diamond and graphite are both pure **carbon** — different structures give different properties.
- In graphite, the **weak forces between layers** let the layers slide, making it a good lubricant. The **delocalised electrons** are free to move along the layers, so it conducts.

- EXT Silicon dioxide has a structure like diamond, so it has similar properties: hard, high melting point, does not conduct.

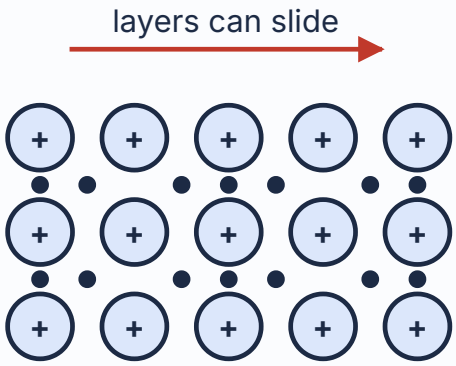
EXAM TIP

"Why is diamond so hard?" → each carbon forms **four strong covalent bonds** in all directions through the whole structure.

6 Metallic bonding

A lattice in a sea of electrons

Metallic bonding — the electrostatic attraction between positive metal ions and a sea of delocalised (free-moving) electrons.



● delocalised electrons

metal ions in a regular lattice

metallic bond: attraction between the + ions and the sea of delocalised electrons

free electrons carry charge and thermal energy: conducts

layers slide without breaking: malleable and ductile [E]

A metal. The positive ions sit in a regular lattice. The outer electrons are free to move through the whole structure.

- Metal atoms lose their outer electrons. These become **delocalised** — they belong to the whole piece of metal, not to one atom.
- The attraction between the + ions and the − electrons holds the metal together.

Properties of metals

EXTENDED

- **Good electrical conductors**: the delocalised electrons carry charge through the metal.
- **Malleable and ductile**: the layers of ions can slide over each other without breaking the bonding, because the electrons keep holding the ions together.
- **High melting points**: the attraction between the ions and the electrons is strong, and it takes a lot of energy to break it.

COMMON MISTAKE

"The metal ions are free to move and carry the current." In a metal the **electrons** carry the current; the ions stay in place.

7 Compare the structures

Which type of substance is it?

Type	Made of	Held together by	Melting point	Conducts?	Examples
ionic	metal + non-metal ions	strong attraction between + and − ions	high	only when molten or dissolved	NaCl, MgO
simple molecular	small molecules	weak forces between molecules	low	never	H_2O, CO_2, I_2
giant covalent	huge network of atoms	strong covalent bonds	very high	no (graphite does)	diamond, SiO_2
metallic	+ ions and electrons	+ ions attracted to a sea of electrons	usually high	yes, solid and liquid	Cu, Fe, Na

- **Quick decision:** metal on its own → metallic. Metal + non-metal → ionic. Two non-metals → covalent (small molecule — unless it is C, Si or SiO_2).

MEMORY TRICK

Ionic = give and take. Covalent = share. Metallic = pool.

Riyaz Sir

8 Test yourself

Test yourself — cover the answers first

QUICK CHECK

1. What is the difference between a compound and a mixture?
2. An atom has 15 protons and a nucleon number of 31. How many neutrons does it have?
3. Write the electronic configuration of calcium (Z = 20).
4. What do all atoms of the same element have in common?
5. Why are noble gases unreactive?
6. Why does graphite conduct but diamond does not?

EXAM-STYLE

7. Draw (or describe) the dot-and-cross diagram for sodium chloride. [3]
8. Explain why sodium chloride has a high melting point but does not conduct when solid. [3]
9. EXT Explain why iodine has a low melting point but diamond has a very high one. [3]
10. EXT Calculate the relative atomic mass of boron: 20% ^{10}B and 80% ^{11}B.

Show answers

1. a compound has atoms chemically bonded in a fixed ratio; a mixture is not bonded and the amounts can vary
2. 31 − 15 = 16
3. 2,8,8,2
4. the same number of protons
5. they have a full outer shell
6. graphite has delocalised electrons; every carbon in diamond uses all 4 outer electrons in bonds
7. Na loses its one outer electron to Cl
   Na^+ has 2,8 and Cl^- has 2,8,8
   each ion shown in brackets with its charge (+ and −)
8. it is a giant lattice of ions with strong electrostatic attraction in all directions
   so a lot of energy is needed to overcome it
   in the solid the ions cannot move, so no charge can flow
9. iodine is simple molecular with weak intermolecular forces between molecules, which need little energy to overcome
   diamond is a giant covalent structure, so every strong covalent bond must be broken
   that needs a lot of energy
10. (20 × 10 + 80 × 11) ÷ 100 = (200 + 880) ÷ 100 = 10.8

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