

ESAT

Chemistry

Bonding, Structure and Particles

Free lesson notes

Covers	C6.1 to C6.7 and C15.1 to C15.2
Format	Lesson, worked examples, self-check
Price	Free

About this lesson

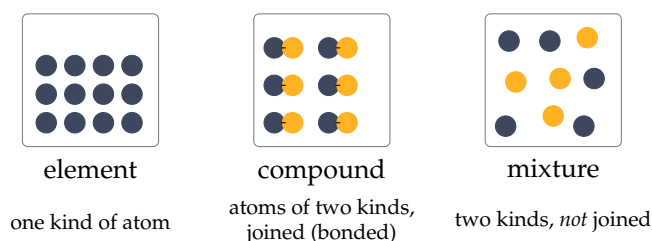
Why is salt a brittle solid that melts at 801°C , while the oxygen you breathe is a gas at -183°C ? Why does copper conduct electricity but diamond, also an element, does not? The answer to every question of this kind is the same: it depends on *how the particles are held together*. Chemists recognise just three types of strong bond (ionic, covalent and metallic) plus a weaker attraction between separate molecules, and from the type of bonding you can predict a substance's melting point, hardness, conductivity and solubility.

This lesson builds the whole picture from scratch. We start by sorting substances into elements, compounds and mixtures, then explain *why* atoms bond at all (they are heading for a full outer shell). We then take ionic, covalent and metallic bonding in turn, with dot-and-cross diagrams, the rules for writing formulae, and the structures of the key giant lattices. We look at the weak forces between molecules, and finish with the particle model of solids, liquids and gases and what really happens when a substance melts or boils. Specification codes: C6.1–C6.7 (bonding and structure) and C15.1–C15.2 (particle theory). Questions in the real test are quick (about 90 seconds each), so the aim here is to make the reasoning automatic: identify the type of structure, then read off the properties.

Heurisco is an independent practice provider and is not affiliated with, endorsed by, or connected to UAT-UK, Pearson VUE, the UCAT Consortium, or any university. All questions are original.

1 Elements, compounds and mixtures

Three kinds of substance. An **element** contains only one kind of atom (oxygen, iron, carbon); it cannot be broken down into anything simpler by chemical means. A **compound** contains two or more different elements *chemically bonded* together in a fixed ratio (water, H_2O , always has two hydrogen atoms for each oxygen atom). A **mixture** contains two or more substances that are *not* chemically joined, in any proportions you like (air, sea water, brass).



How they differ. In a compound the properties are quite different from those of the elements it contains: sodium is a violently reactive metal and chlorine is a poisonous gas, yet sodium chloride is the harmless salt on your chips. A compound can only be broken back into its elements by a chemical reaction. The parts of a mixture keep their own properties and can be separated by *physical* methods such as filtering, distilling or using a magnet (this is the subject of a later lesson).

Worked example: classifying

A clear liquid boils at a single fixed temperature, and when it is passed through an electric current it splits into two different gases. Is it an element, a compound or a mixture?

A single fixed boiling point suggests a pure substance, and it was split into two different gases, so it contains at least two kinds of atom and is a *compound*. (An element could not be split by electricity into different substances, and a mixture would boil over a range of temperatures.)

Common mistakes

- Calling air or sea water a compound. They are mixtures, because the parts are not bonded and can vary in proportion.
- Thinking that O_2 is a compound because it has two atoms. Both atoms are oxygen, so it is an element (a molecule

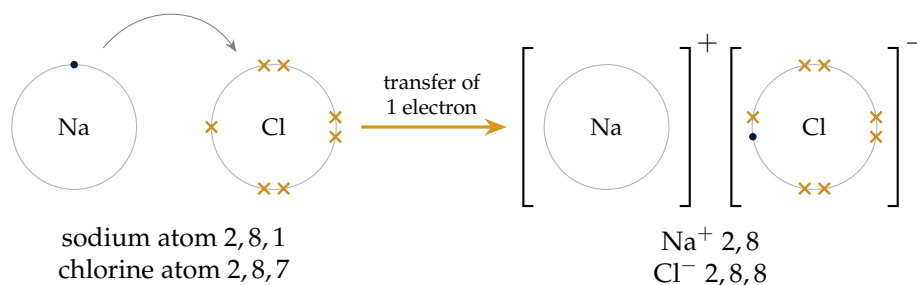
of an element).

- Saying that a compound “contains” the properties of its elements. The properties change completely.

2 Why atoms bond, and ionic bonding

The noble gas goal. The noble gases (helium, neon, argon and so on) have a full outer shell of electrons (2, or 2, 8, or 2, 8, 8) and are almost completely unreactive. Every other atom is, in effect, trying to reach the same kind of full outer shell, and it can do so in three ways: by *transferring* electrons (ionic bonding), by *sharing* electrons (covalent bonding) or by pooling them (metallic bonding). The energy released when the new arrangement forms is what drives the reaction.

Ionic bonding: transfer. A metal atom has few outer electrons and loses them; a non-metal atom has nearly full shells and gains electrons. An atom that has lost or gained electrons is an **ion**. Sodium (2, 8, 1) gives its single outer electron to chlorine (2, 8, 7). Sodium becomes Na^+ (2, 8, the neon arrangement) and chlorine becomes Cl^- (2, 8, 8, the argon arrangement). The two oppositely charged ions then attract each other very strongly. That electrostatic attraction between oppositely charged ions is the **ionic bond**.

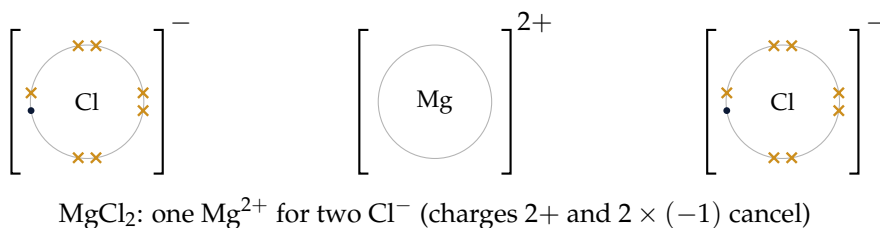


In a **dot-and-cross diagram** the electrons of one atom are drawn as dots and those of the other as crosses (they are identical really; the symbols just let you follow where each came from). After transfer we show only the outer shell, in square brackets with the charge at the top right.

Predicting the charge. For the main groups the charge follows from the group number, because an atom loses or gains the fewest electrons to reach a full shell:

group	1	2	13 (aluminium)	15	16	17
ion formed	1+	2+	3+	3-	2-	1-
example	Na^+	Mg^{2+}	Al^{3+}	N^{3-}	O^{2-}	Cl^-

Writing the formula. A compound has no overall charge, so the positive and negative charges must cancel. Magnesium chloride needs two Cl^- to cancel one Mg^{2+} , giving MgCl_2 .



A quick rule: swap the charges (ignoring signs) and use them as subscripts, then cancel any common factor. For Al^{3+} and O^{2-} the swap gives Al_2O_3 (two 3+ and three 2- both make 6). For Mg^{2+} and O^{2-} the swap gives Mg_2O_2 , which cancels to MgO .

Worked example: writing formulae

Give the formulae of (a) calcium chloride, (b) aluminium oxide, (c) potassium oxide.

(a) Ca^{2+} and Cl^- : one calcium needs two chlorides, CaCl_2 . (b) Al^{3+} and O^{2-} : Al_2O_3 ($2 \times 3 = 3 \times 2 = 6$). (c) K^+

and O^{2-} : two potassium ions for one oxide ion, K_2O .

Compound ions. Some ions are groups of atoms bonded covalently that carry a charge together. You should know these five: hydroxide OH^- , nitrate NO_3^- , carbonate CO_3^{2-} , sulfate SO_4^{2-} and ammonium NH_4^+ . When more than one is needed the ion goes in brackets: magnesium hydroxide is $Mg(OH)_2$ and ammonium sulfate is $(NH_4)_2SO_4$.

Metals with more than one charge. The transition metals can form ions of different charges, so the name carries a *Roman numeral* giving the charge: iron(II) is Fe^{2+} and iron(III) is Fe^{3+} . Hence iron(III) oxide, with Fe^{3+} and O^{2-} , is Fe_2O_3 , and copper(II) sulfate is $CuSO_4$.

Worked example: compound ions and Roman numerals

Write the formula of iron(III) sulfate.

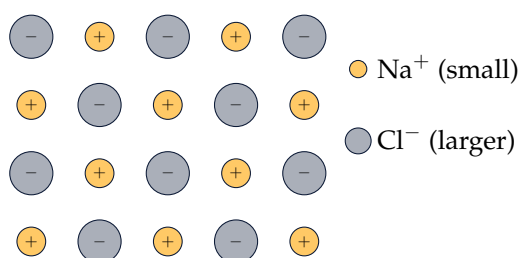
Iron(III) is Fe^{3+} ; sulfate is SO_4^{2-} . Swapping charges gives $Fe_2(SO_4)_3$ (two $3+$ and three $2-$ both make 6). The sulfate is in brackets because we need three of the whole group.

Common mistakes

- Putting the charges in the formula as subscripts, e.g. Mg_2Cl . The charges decide the *ratio*, the subscript is the number of ions.
- Writing Mg_2O_2 instead of MgO . Always cancel to the simplest ratio for an ionic compound.
- Forgetting brackets: $Mg(OH)_2$ is not $MgOH_2$.

3 Structure and properties of ionic compounds

A giant lattice. An ionic compound does not form separate molecules. The ions pack into a regular *giant ionic lattice* in which every ion is surrounded by ions of the opposite charge, held by strong electrostatic attraction in all directions.



giant ionic lattice: every ion is held by strong attraction to its oppositely charged neighbours

This one structure explains all the properties you must know:

- High melting and boiling points.** Melting means freeing the ions from their fixed positions, which requires overcoming many strong attractions, so a lot of energy is needed (NaCl melts at $801^\circ C$).
- Does not conduct when solid; conducts when molten or dissolved in water.** Electricity is carried by charged particles that are free to move. In the solid the ions are locked in the lattice, but when molten or dissolved the ions can move and carry the current.
- Brittle.** If a force slides one layer by one ion, ions of the same charge come next to each other and repel, so the crystal splits.

Stronger bonds, higher melting point. The attraction is bigger when the ions carry larger charges and are smaller (closer together). So MgO (Mg^{2+} , O^{2-}) melts at about $2800^\circ C$, far higher than $NaCl$ (Na^+ , Cl^-) at $801^\circ C$.

Worked example: explaining a property

Explain why solid sodium chloride does not conduct electricity but molten sodium chloride does.

In the solid, the Na^+ and Cl^- ions are held in fixed positions in the lattice, so there are no mobile charged particles. When molten, the lattice has broken down and the ions can move, carrying charge. (No free *electrons* are involved in either case.)

Common mistakes

- Saying that ionic compounds conduct because of free electrons. It is free *ions*.
- Saying that “the ionic bond is broken” when salt dissolves or melts without mentioning that the ions are separated from the lattice.
- Saying that you melt ionic compounds easily because the bond is weak: the bonds are strong, so the melting point is high.

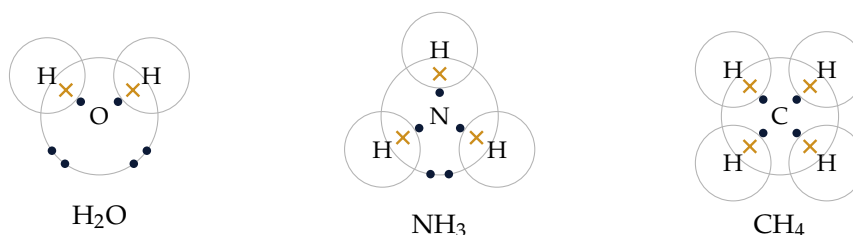
4 Covalent bonding and simple molecules

Sharing. Two non-metal atoms both need electrons, so neither can give one away. Instead they *share*. A **covalent bond** is a shared pair of electrons, held between the two nuclei, which attracts both. Each atom counts the shared electrons towards its own outer shell, so both reach a full shell. Hydrogen needs 2 in its shell (like helium); most other non-metals need 8.

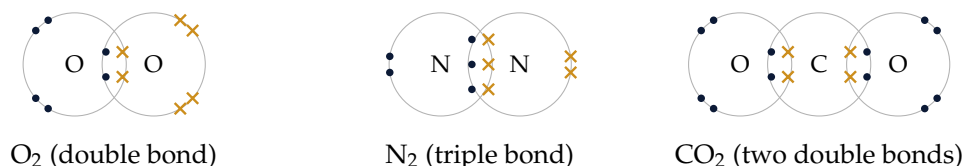


Count it up for chlorine: each Cl has seven outer electrons. In Cl_2 one pair is shared, so each atom has $6 + 2 = 8$. The six unshared electrons on each atom form three **lone pairs**.

More atoms, more bonds. The number of bonds an atom forms is the number of electrons it is short of a full shell: H forms 1 bond, O forms 2, N forms 3 and C forms 4. This gives the molecules below.



Double and triple bonds. Sometimes two atoms share two pairs (a *double bond*) or three pairs (a *triple bond*). Oxygen needs two more electrons, so O_2 has a double bond; nitrogen needs three, so N_2 has a triple bond; in carbon dioxide each oxygen shares two pairs with carbon.



Check. In every molecule above, count the electrons around each atom (own electrons plus the ones shared with it): every atom has 8 (hydrogen 2).

Worked example: drawing a molecule

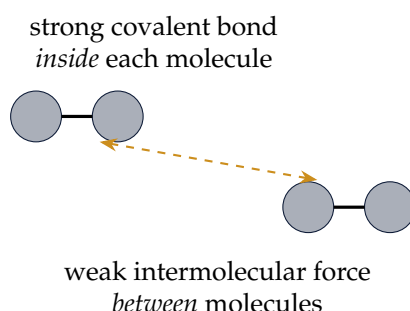
How many bonds and lone pairs are there in a molecule of ammonia, NH_3 ?
 Nitrogen (2, 5) needs 3 more electrons, so it shares one pair with each of three hydrogens: 3 covalent bonds. It used 3 of its 5 outer electrons in bonds, so $5 - 3 = 2$ electrons remain: 1 lone pair. Nitrogen then has $3 \times 2 + 2 = 8$ electrons around it.

Common mistakes

- Forgetting lone pairs: a diagram of H_2O should show two bonding pairs and two lone pairs on oxygen.
- Drawing crosses and dots in the shared region from the same atom. Each shared pair has one electron from each atom.
- Giving hydrogen eight electrons. It is full with two.

5 Simple molecules: weak forces between them

Strong inside, weak between. The atoms inside a small molecule are held by strong covalent bonds, but one molecule is attracted to its neighbours only by **intermolecular forces**, which are much weaker. Compare a molecule with a team: the players are tightly linked to each other, but the teams barely touch.



Properties of simple molecular substances.

- **Low melting and boiling points.** When a simple molecular substance melts or boils, it is the weak intermolecular forces that are overcome, *not* the covalent bonds. Little energy is needed, so many are gases or liquids at room temperature.
- **Do not conduct electricity** (solid, liquid or gas): there are no ions and no free electrons, and the molecules are uncharged.

Bigger molecules, higher boiling point. The intermolecular forces get stronger as the molecules get larger (more electrons). Chlorine (Cl_2) is a gas at room temperature, bromine (Br_2) a liquid and iodine (I_2) a solid. The covalent bond inside the molecule does not matter here: it is not broken when a halogen melts.

Worked example: what is overcome?

Which is overcome when solid iodine, I_2 , sublimes: the covalent bond between the I atoms, or the forces between I_2 molecules?

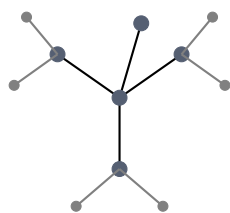
Iodine is a simple molecular substance. On changing state the I_2 molecules move apart, so the *intermolecular forces* are overcome. The I–I covalent bonds stay intact, as shown by the fact that the vapour is still I_2 (purple gas of I_2 molecules).

Common mistakes

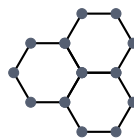
- Saying that boiling water breaks the bonds between hydrogen and oxygen. Steam is still H_2O ; only the forces between molecules are overcome.
- Saying that simple molecules have low boiling points because covalent bonds are weak. Covalent bonds are strong; the *intermolecular* forces are weak.
- Explaining non-conduction by “no electrons”. Molecules contain electrons, but none that are free to move through the substance.

6 Giant covalent structures and metallic bonding

Giant covalent. In some covalent substances the atoms are not in small molecules but form one continuous network of covalent bonds in all directions: a **giant covalent** structure. Melting means breaking strong covalent bonds, so the melting points are very high.



diamond:
each C bonded to 4



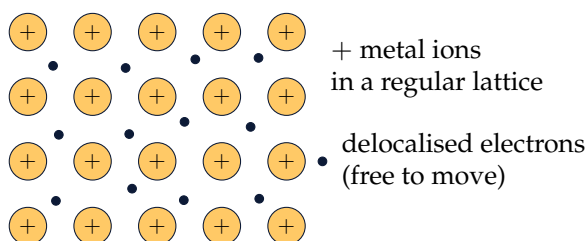
graphite layer:
each C bonded to 3



layers (side view),
slide over each other

Diamond has each carbon bonded to four others in a rigid tetrahedral network: extremely hard, very high melting point, and does not conduct (all four outer electrons are held in bonds). **Graphite** has each carbon bonded to only three others, in flat hexagonal layers. The fourth outer electron of each carbon is *delocalised* (free to move along the layer), so graphite *conducts electricity*. The layers are held together only by weak forces, so they slide over each other and graphite is soft and slippery. **Silicon dioxide**, SiO_2 (sand), has each Si bonded to four O and each O to two Si: hard, with a very high melting point and no conduction.

Metallic bonding. In a metal the outer electrons leave their atoms and move freely through the whole structure. The result is a lattice of positive metal ions in a "sea" of **delocalised electrons**. The **metallic bond** is the strong attraction between the positive ions and the delocalised electrons.



- **Conduct electricity** (solid or molten), because the delocalised electrons are free to move and carry the charge; **conduct heat** for the same reason.
- **Malleable and ductile:** layers of ions can slide over each other, and the electron sea keeps holding them together, so the metal bends instead of shattering.
- **High melting points**, because the attraction between ions and the electron sea is strong.

Worked example: graphite or diamond?

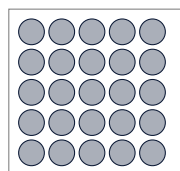
Explain why graphite conducts electricity but diamond does not, although both are pure carbon.
In diamond every carbon uses all four outer electrons in covalent bonds, so no electron is free to move. In graphite each carbon uses only three outer electrons in bonds; the fourth is delocalised between the layers, and can move to carry a current.

Common mistakes

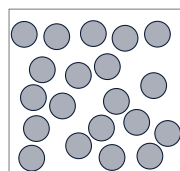
- Saying that diamond has weak forces between its atoms, or that graphite has no covalent bonds. Both are held by strong covalent bonds; graphite's *layers* have weak forces between them.
- Thinking metals conduct because of free *ions*. In a metal the ions are fixed and the electrons move.
- Calling a giant covalent substance a molecule. It has no separate molecules.

7 The particle model: states and changes of state

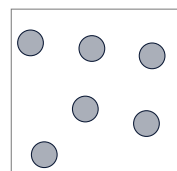
Three states. Everything is made of tiny particles in constant motion. How they are arranged and how fast they move decide the state.



solid: regular, touching, vibrating



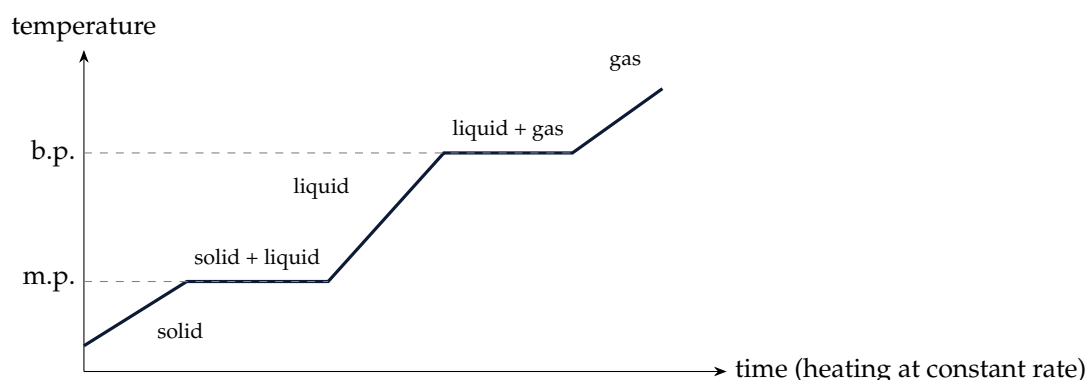
liquid: touching, irregular, moving past



gas: far apart, random, fast

	arrangement	movement	properties
solid	regular, touching	vibrate about fixed positions	fixed shape and volume; cannot be compressed
liquid	irregular, touching	move past one another	fixed volume, takes the shape of the container
gas	far apart, random	fast, free in all directions	fills the container; easily compressed

Changes of state. Heating gives particles more energy, so they vibrate or move faster. At the **melting point** the particles gain enough energy to overcome the forces holding them in position, and the solid becomes a liquid; at the **boiling point** they overcome the forces holding them together and escape as a gas. Cooling reverses these changes (freezing, condensing).



Why does the temperature stay flat? During a change of state the energy supplied goes into *overcoming the forces between particles*, not into making them move faster. Temperature measures the average kinetic energy of the particles, so it stays constant until the change is complete (the flat parts of the graph). A pure substance has a sharp melting point; a mixture melts over a range.

Link to bonding. The stronger the forces to be overcome, the higher the melting and boiling points. Simple molecular substances overcome only the weak *intermolecular forces* and have low points. Ionic, giant covalent and metallic substances must overcome strong bonds throughout the structure and have high points.

Worked example: reading a heating curve

A pure solid is heated steadily. The temperature rises from 20 °C to 80 °C in 2 minutes, stays at 80 °C for 6 minutes, then rises again. What is happening at 80 °C, and what is the state at 50 °C?

The flat section is the melting point, 80 °C: energy is overcoming the forces between particles as the solid turns to liquid, and the temperature stays constant. At 50 °C, below the melting point, the substance is a solid.

Common mistakes

- Thinking nothing is happening on the flat part. Energy is still being supplied and used to separate the particles.
- Saying particles “expand” or “get bigger” on heating. The particles stay the same size; they move faster and further apart.
- Saying a gas has no forces at all. There are forces, but they are negligible compared with the particles’ energy.

8 Putting it together: identifying a structure from data

The standard ESAT approach is to read the *properties* and decide the structure:

structure	melting point	conducts?	example
simple molecular	low	never	H ₂ O, CO ₂ , I ₂
ionic	high	only if molten or dissolved	NaCl, MgO
giant covalent	very high	no (graphite: yes)	diamond, SiO ₂
metallic	usually high	yes, solid and molten	Cu, Fe, Na

Identify the structure

Decision order. (1) Conducts when solid? Then it is a metal (or graphite). (2) Conducts only when molten or dissolved? Ionic. (3) Low melting point, no conduction? Simple molecular. (4) Very high melting point, no conduction, insoluble? Giant covalent.

Worked example: which structure?

Substance W melts at -20°C , is a liquid at room temperature and does not conduct. Substance Z melts at 1700°C , is hard, insoluble in water, and does not conduct even when molten. Name the structure of each.

W: low melting point and no conduction, so simple molecular. Z: very high melting point, hard and no conduction when molten (so no free ions), so a giant covalent structure. (If Z conducted when molten but not solid it would be ionic.)

Exam tip

A frequent trap is a statement that is true of the wrong thing. When the question says “explain why X has a low melting point”, check that your answer talks about the *weak intermolecular forces* and not the covalent bonds; and for ionic compounds, about free *ions*, not electrons.

Common mistakes

- Deciding by the elements alone: carbon can be a hard insulator (diamond), a soft conductor (graphite), or part of a gas (CO₂).
- Assuming that a high melting point means ionic. Giant covalent and metals are also high.
- Forgetting that metals conduct in the solid state, whereas ionic compounds do not.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. Write the formula of (a) aluminium chloride, (b) calcium nitrate.
2. Draw (describe) the dot-and-cross diagram of H_2O : how many bonding pairs and lone pairs on oxygen?
3. Why does magnesium oxide have a higher melting point than sodium chloride?
4. Why is CO_2 a gas but SiO_2 a solid with a very high melting point?
5. Explain why metals can be hammered into shape.
6. What stays the same on the flat part of a heating curve, and where does the energy go?
7. Name the type of bonding in iron(III) oxide, Fe_2O_3 , and give the charge on iron.
8. A substance conducts when molten and when dissolved but not when solid. What is its structure?

Answers to the self-check

1. (a) AlCl_3 . (b) $\text{Ca}(\text{NO}_3)_2$.
2. Two bonding pairs and two lone pairs.
3. Its ions carry larger charges ($2+$, $2-$ against $1+$, $1-$), so the electrostatic attraction in the lattice is stronger.
4. CO_2 is simple molecular (weak intermolecular forces); SiO_2 is a giant covalent network of strong bonds.
5. Layers of positive ions can slide over each other while the delocalised electrons keep them bonded.
6. Temperature stays the same; the energy is used to overcome forces between the particles.
7. Ionic; Fe^{3+} .
8. Giant ionic lattice.

Next step

Heurisco's *ESAT Chemistry: Bonding, Structure and Particles Practice Set* has twenty original exam-style questions on this lesson, with dot-and-cross diagrams, heating curves and data tables, in order of difficulty, with a full worked solution for each.