Chem Factsheet

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Common Confusions : Structure and Bonding - When, and When Not, to Use IMF

Structure and Bonding - Likely Candidate Errors

The aim of this Factsheet is to enable 'A' level candidates to score full marks when answering questions on structure and bonding. To write answers that are precise and that do not contain conflicting or contradictory information. Such information always results in loss of mark(s) previously gained! There are usually alternative answers that would gain credit but these are often not given in mark schemes because they are often less precise.

The approach taken is to consider some typical types of 'A' level questions and answers that candidates might provide. The merits of those answers will then be considered with faults and omissions discussed. However first of all, here are the meanings of the terms "structure" and "bonding".

Structure

This describes the arrangement in which atoms, ions or molecules are held together in a particular substance. Terms such as "lattice", "molecule", "giant", "ion", dipole" etc *may* be applicable here.

Bonding

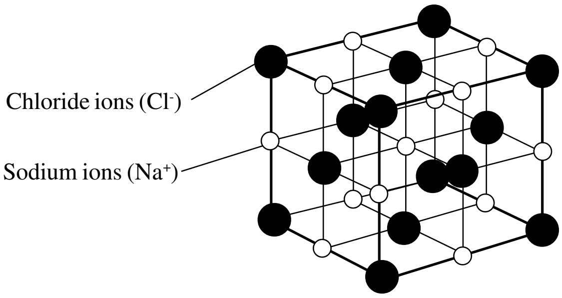
This describes the "forces" between the atoms, ions or molecules in a particular substance to create and maintain the resulting structure. Terms such as "covalent", "dative", "van der Waals", "metallic", "dipole-dipole" etc *may* be applicable here.

Note: *Structure* is what results from bonding! Structure is measurable directly by experiment but bonding is theoretical, although based on observations.

1. Consider questions of the type, "**Name and describe the BONDING in a** ***particular substance***".

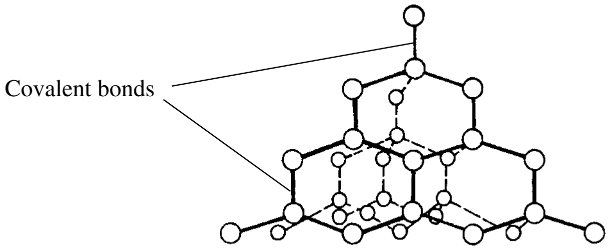
(a) Sodium chloride

Ionic bonding. The (electrostatic) attraction between oppositely charged ions / Na^+ and Cl^- ions. Don't miss out, "oppositely charged". Don't explain how the ions are formed from the atoms.



(b) Diamond / silicon dioxide

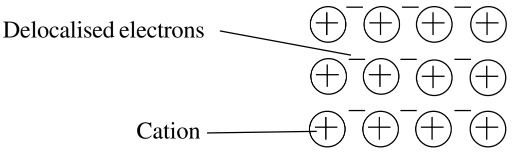
Covalent bonding. A shared pair of electrons (between two atoms). Remember to say that the electrons are "shared" and that there is a "pair".



Note The covalent bond is the net electrostatic attraction between the nuclei of the two atoms and the shared pair of electrons between them.

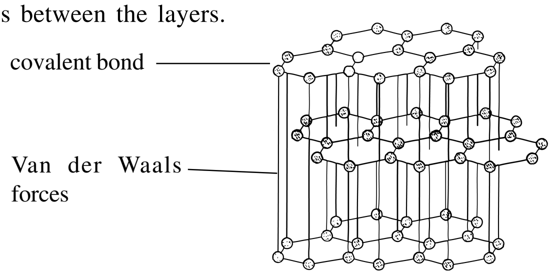
(c) Sodium

Metallic bonding. The electrostatic attraction between positive ions / cations and the delocalised (outer) electrons / sea of delocalised electrons. Don't just say "electrons" / "outer electrons".



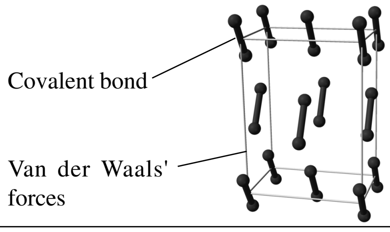
(d) Graphite

Covalent bonding between the carbon atoms within the layers. Plus delocalised electrons across layers. Plus van der Waals' forces between the layers.



(e) Iodine

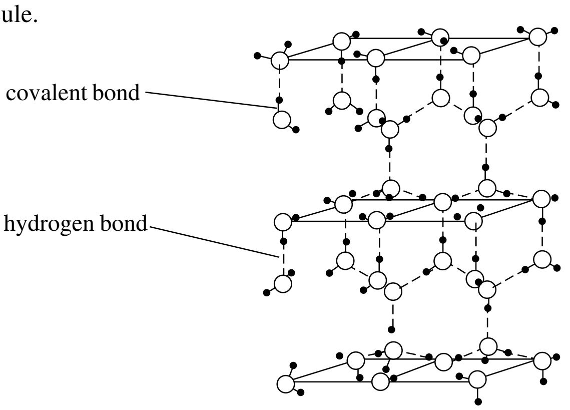
Covalent bonding (a shared pair of electrons between two atoms. Remember to say that the electrons are "shared" and that there is a "pair") between the iodine atoms of the iodine molecule (I_2) / within the molecule. van der Waals' forces / dispersion forces / London forces / temporary dipole – induced dipole forces are between the iodine molecules.



Note: An "instantaneous dipole" occurs when the electron density in an atom is unsymmetrical. This dipole induces a dipole of opposite polarity in neighbouring molecules / atoms. The opposite electrical charges attract.

(f) Ice

Covalent bonding between O and H atoms within the water molecule. Hydrogen bonding between the water molecules. Don't say "Hydrogen" for "Hydrogen bonding". The HB is drawn from the O lone pair. The $O^{\delta-}$**:** **- - -** $^{\delta+}H$ — $^{\delta-}O$ is linear (180^0) . The electrostatic attraction between the lone pair of oxygen in one water molecule and the electron deficient hydrogen atom / $^{\delta+}H$ of an adjacent molecule.



(g) Solid hydrogen bromide

Covalent bonding between the H and Br atoms of the molecule (HBr). Dipole-dipole attractions <u>and</u> van der Waals' forces between the molecules.

Don't just say "dipole". Dipoles arise because bromine is more electronegative than hydrogen. Hence the covalent bond is polar and the oppositely charged dipoles in adjacent molecules attract.

$H^{\delta+}$— $Br^{\delta-}$ $H^{\delta+}$— $Br^{\delta-}$ $H^{\delta+}$— $Br^{\delta-}$ $H^{\delta+}$— $Br^{\delta-}$
$Br^{\delta-}$— $H^{\delta+}$ $Br^{\delta-}$— $H^{\delta+}$ $Br^{\delta-}$— $H^{\delta+}$ $Br^{\delta-}$— $H^{\delta+}$

Don't just say "bromine is electronegative" – make a *comparison* of the electronegativities of H and Br.

In d) – g) make it clear if the bonds are "within" (covalent) or "between" (VW, DP-DP, H-bonds) molecules.

Always be carefully of the name of the type of particle being bonded; don't confuse the terms atom, ion and molecule.

2. Consider questions of the type, "**Describe, with the aid of a diagram, the STUCTURE in a** ***particular substance".***

Note : appropriate diagrams have already been given in part 1.

a) Sodium chloride

Giant structure of ions. There is a 3D lattice of sodium ions each surrounded by an octahedral arrangement of (six) chloride ions and vice versa. The overall structure is cubic.

b) i) Diamond

Giant structure of atoms / macromolecular. There is a 3D lattice of carbon atoms each tetrahedrally surrounded by 4 carbon atoms.

b) ii) Silicon dioxide

Giant structure of atoms / macromolecular. There is a 3D lattice of silicon atoms each surrounded by 4 oxygen atoms and each O atom is bonded to two Si atoms.

c) Sodium

Giant metallic structure. There is a 3D regular lattice of positively charged ions in a "sea" of delocalised electrons.

In all these "giant structures" (a-c) never mention "molecules" or any type of intermolecular force. This would result in loss of previous mark(s) for previous correct statements.

d) Graphite

Giant structure of atoms in 2D / macromolecular. There are layers of fused hexagonal rings of carbon atoms. The layers are far apart compared to the C-C distance in the layers.

e) Iodine, ice and solid hydrogen bromide

All are (simple) molecular structures. The molecules are arranged in a 3D lattice structure.

> <u>Note</u>. These structures are "molecular"; the particles are "molecules".

3. Consider questions of the type, **"explain *physical properties* in terms of structure and or bonding".**

> **NB** You must **relate** the property to the bonding and / or the structure, NOT merely describe the bonding.

a) In terms of structure and bonding explain why magnesium has a higher melting point than sodium.

Both metals have giant metallic structures with metallic bonding. Na atoms have one and Mg atoms have two delocalised electrons. Mg^{2+} (+2) has a higher ionic charge than Na^+ (+1). So the electrostatic attraction between the ions and delocalised electrons increases from Na to Mg. Hence the melting point increases. ("Metallic bond is stronger in Mg" is not explaining. Positive ions does not mean "ionic bonds"!)

b) In terms of structure and bonding explain why silicon has a much higher melting point than phosphorus.

Silicon has a macromolecular structure. There are many covalent bonds between Si atoms (4 per Si). These have to be broken to melt it. (Don't use "molecule" and don't mention intermolecular forces.)

Phosphorus has a (simple) molecular structure. (P_4) molecules are held in the lattice by van der Waals' forces. These have to be broken to melt phosphorus.

The covalent bonds between Si atoms are **much** stronger than the van der Waals' forces between phosphorus molecules. Hence Si-Si covalent bonds require far more energy that P_4 - - P_4 intermolecular forces to break. Hence Si has a **much** higher melting point.

c) In terms of structure and bonding explain why sulfur has a higher melting point than chlorine.

Both elements have (simple) molecular structures. The intermolecular forces are van der Waals' forces in both cases. Sulfur (S_8) has more electrons / is a bigger molecule / more polarisable than chlorine (Cl_2). Hence the van der Waals' forces are strong**er** between sulfur molecules than between chlorine molecules. (Use "strong**er**" not "strong".) Hence S_8 — S_8 inter-molecular forces require more energy that Cl_2 — Cl_2 intermolecular forces to break. Hence sulphur has a higher melting point.

d) In terms of structure and bonding explain why strontium has a higher melting point than barium.

Structures: giant metallic. Bonding: metallic. There are two delocalised electrons per atom. The ions each have 2+ charges. Sr^{2+} is a small**er** ion than Ba^{2+}. So Sr^{2+} has the high**er** charge density than Ba^{2+}. Hence there is greater (electrostatic) attraction between it and the delocalised electrons. Hence, more energy is required to overcome the metallic bonding in Sr. Hence, Sr has a higher melting point.

e) In terms of appropriate bonding forces explain why these boiling points differ:

NH_3 -33^{0}C, PH_3 -88^{0}C.

These are both (simple) molecular substances. Hence, the "appropriate" bonding forces *are inter-molecular forces*!

NH_3 has hydrogen bonding between molecules. PH_3 has dipole-dipole attractions / van der Waal forces. Hydrogen bonding is strong**er** and therefore requires more energy to break it. Hence, NH_3 has the higher boiling point.

f) In terms of appropriate bonding forces explain why these boiling points differ:

H_2S -60^{0}C, H_2Se -41^{0}C.

Again, these are both (simple) molecular substances. Hence, the "appropriate" bonding forces *are inter-molecular forces*!

The boiling point of H_2S is lower than that of H_2Se. The van der Waals' forces between H_2S molecules are weak**er** than those between H_2Se molecules. H_2S has less electrons / is a smaller molecule / less polarisable than H_2Se molecules. NB Because S is more electronegative than Se, dipole-dipole attractions get weaker from H_2S to H_2Se and hence don't explain the boiling point difference (-60 is lower than -41!).

g) Explain how the ability of the diamond, graphite, potassium bromide, magnesium and iodine to conduct electricity can be explained in terms of structure and bonding.

Conductivity or non-conductivity MUST be explained in terms of the presence or absence of mobile charged particles – *electrons or ions* are appropriate.

Solid diamond has a macromolecular structure with all outer electrons in covalent bonds. There are *no free electrons* so diamond cannot conduct.

Solid graphite has a macromolecular structure. The carbon atoms use only three electrons to form covalent bonds. The *fourth electron of each carbon atom becomes delocalised* allowing conduction. Solid KBr has a giant ionic structure with the *ions in fixed positions* in a lattice so it does not conduct. On melting / dissolving in water the *ions are free to move* and hence conduction occurs.

Solid Mg has a giant metallic structure with *mobile delocalised electrons* allowing conduction. Mg will also conduct when molten because there are still mobile delocalised electrons.

Solid I_2 has a simple molecular structure. Molecules have no electrical charges so iodine does not conduct.

Practice Questions

1. Name the strongest type of intermolecular force present in (i) liquid ethane, (ii) liquid fluoroethane and (iii) liquid ethanol.

2. Potassium chloride has the same structure as sodium chloride. Draw a diagram to show the arrangement of the ions in a crystal of potassium chloride. Explain in terms of bonding, why potassium chloride has a high melting point. State and explain what must be done to potassium chloride to make it conduct electricity.

3. (a) Explain how the ions are held together in the crystalline solids, sodium and sodium chloride.
   (b) Why is the melting point of sodium chloride (801^{0}C) much higher than that of sodium (98^{0}C).
   (c) Explain why sodium is malleable.

4. Diamond, graphite and the fullerene (C_{60}) are all forms of carbon.
   (a) State the structures of these three substances and state how the structures of diamond and graphite differ.
   (b) Explain why diamond is very hard and why graphite can be used as a lubricant and as an electrical conductor.
   (c) Predict three physical properties of the fullerene.

5. `The melting points of sodium, bromine and sodium bromide are 98^{0}C, 7 °C and 747 °C respectively.
   (a) For each of the substances sodium, bromine and sodium bromide, state the type of bonding present and explain the nature of the attractive forces holding each solid together.
   (b) Why is the melting point of bromine much lower than that of sodium bromide.
   (c) Which of the substances is soluble in organic solvents such as hexane?

6. Draw a labelled diagram to show a model for metallic bonding.

7. Sodium chloride and pure water do not conduct electricity. Sodium chloride dissolved in water produces a solution that does conduct electricity. Explain these observations.

8. The table shows the boiling points of the elements of Period 3.

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Boiling point /^{0}C	883	1107	2467	2355	280	445	-35	-189
Bonding								
Structure								
Type of intermolecular force								

   (a) Complete the bonding row using M for metallic and C for covalent. Put a dash if there are neither metallic nor covalent.
   (b) Complete the structure row from Na to Cl using G for giant structure and S for simple molecular structure. Write in an appropriate word for the structure of argon.
   (c) Complete the intermolecular force row using HB for hydrogen bonds, DD for dipole-dipole attractions and VW for van der Waals forces. Put a dash if none of these exist.
   (d) Explain why Al has a higher boiling point that Mg.
   (e) Explain the very low boiling point of argon.
   (f) Why is "ionic" NOT an option in part a)?

9. Describe the nature of hydrogen bonding by drawing a diagram showing two appropriate molecules hydrogen bonded. Explain two anomalous properties of water in terms of this bonding.
10. Explain why calcium oxide has a high melting point.
11. Explain why chlorine, bromine and iodine have different physical states at room temperature and pressure.
12. List all intermolecular forces between (i) hydrogen fluoride molecules, (ii) hydrogen bromide molecules.
13. Describe the structure of aluminium and explain why it is a good conductor of electricity.
14. The densities of diamond and graphite are 3.53 and 2.25 g cm^{-3} respectively. From your knowledge of the structures and bonding in diamond and graphite, suggest why diamond is denser than graphite.
15. Explain using structure and bonding why diamond has a higher melting point than iodine.
16. Consider the following equations: $H_2O(l) \rightarrow H_2O(g)$ - - - (1)
    and $H_2O(g) \rightarrow 2H(g) + O(g)$ - - - (2)
    Explain why the heat energy required for reaction (2) is a lot more than for reaction (1).
17. Explain why the heat energy to vaporise one mole of sodium chloride (171 kJ mol^{-1}) is much greater than that to melt one mole of sodium chloride (30 kJ mol^{-1}).

Answers

1. (i) van der Waals
   (ii) dipole-dipole attractions – not H bonds because F is directly bonded to C, not H.
   (iii) hydrogen bonding.
2. Diagram: see NaCl: K$^+$ replaces Na$^+$. The oppositely charge ions have a strong (electrostatic) attractions to each other and so require much energy to overcome the forces of attraction. ("strong ionic bonds": is not explaining.)
   Melt / dissolve in water. The ions are now free to move.
   (Note Don't mention "molecules", "atoms", "delocalised electrons","intermolecular forces", or "covalent bonds".)
3. (a) Na : (electrostatic) attraction between positive ions (of the lattice) and the delocalised electrons. NaCl : (electrostatic) attraction between oppositely charged ions / Na$^+$ and Cl$^-$.
   ("Metallic bonding" and "ionic bonding" is not an explanation.)
   (b) The ionic bonding in NaCl is stronger than metallic bonding in Na. (See Note in answer to 2)
   (c) The layers of ions/atoms can slide over each other. (Don't say "molecules")
4. (a) Diamond and graphite are macromolecular / giant covalent structures. Fullerene is (simple) molecular.
   See 1 b), 1 d), 2b) and 2 d) and diagrams
   (b) Hard: strong covalent bonds must be broken. (Not weakened).
   Lubricant: Weak van der Waals' forces between the layers. A weak applied force will break these forces and the layers will slide over each other.
   Conductor: Graphite has delocalised electrons. (Don't use "sea of electrons".) So these electrons can move along the graphite (layers).
   (c) (Relatively) low melting point / boiling point. Non-conductor of electricity. Insoluble in water.
5. (a) Na: metallic bonding. (See 1 c). $Br_2(s)$: van der Waals' between the molecules (See 1 e). Covalent between the atoms in molecules (See 1 b). NaBr: ionic bonding. (See 1 a). (Wrong bonding / particles: no marks.)
   (b) van der Waals' forces are much weaker that ionic bonds.
   (c) Br_2 it as similar imf to the solvent.
6. See 1 c)
7. NaCl(s): ions are not free to move about / fixed in a lattice.
   Water: water molecules are uncharged.
   NaCl(aq) conducts as the ions are free to move.
8.

Element	Na	Mg	Al	Si	P	S	Cl	Ar
Boiling point /^{0}C	883	1107	2467	2355	280	445	-35	-189
Bonding	M	M	M	C	C	C	C	-
Structure	G	G	G	G	M	M	M	Atomic
Type of intermolecular force	-	-	-	-	VW	VW	VW	VW

   (d) Al has 3 delocalised outer electrons and Mg only 2. Al^{3+} is smaller / higher positive charge / greater charge density that Mg^{2+}. Al has stronger metallic bonding.
   (e) Ar atoms have few electrons so the van der Waals' forces are very weak (and easy to break).
   (f) Ionic bonding can only apply to *compounds* (between an element of low electronegativity and one of high electronegativity)
9. See 1 f). (Any diagram showing hydrogen bonding is suitable – H_2O, NH_3 or HF are the best examples)
   Water has a higher melting / boiling point than expected (from its M_r). Intermolecular hydrogen bonds, that need to be broken, are stronger (than intermolecular forces between similar small molecules).
   Ice is less dense than water. Hydrogen bonds hold the water molecules apart.
10. CaO has a giant ionic structure. The (electrostatic) attraction between Ca^{2+} and O^{2-} is very strong because they both have double charges.
11. Cl_2 is a gas, Br_2 is a liquid and I_2 is a solid.
    The boiling points increase from Cl_2 to Br_2 to I_2 since the van der Waals' forces/ temporary dipole-induced dipole forces increase as the number of electrons increases down the group. Stronger IMFs to be broken. (In answers, be careful of using symbols rather than formulas. e.g. Cl is a gas! – best to use names)
12. HF has hydrogen bonding and temporary dipole – induced dipole forces.
    HBr has permanent dipole –permanent dipole forces and temporary dipole – induced dipole forces.
13. See 2c) The delocalised electrons are mobile.
14. The covalent bonds is diamond are shorter than the distance between layers in graphite.
    The atoms in diamond are packed closer together.
15. Diamond has a macromolecular structure with many strong covalent bonds throughout and they require a lot of energy to break.
    Iodine is molecular with covalent bonds between atoms of the I_2 molecule.
    Iodine has weak van der Waals' forces between the molecules which are easily broken.
16. Reaction (1) involves breaking (intermolecular) hydrogen bonds. Reaction (2) involves breaking covalent bonds. Covalent bonds are much stronger than hydrogen bonds.
17. When sodium chloride boils *all* ionic bonds are broken (to produce separate gaseous ions.)

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