Chem Factsheet

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How to Answer AS Exam Questions On Bonding

You need to know and understand all related ideas before you can gain maximum credit when answering exam questions on any particular topic. To help you, the following provides a summary of the essential ideas under the topic heading, "**bonding**". However, **applying** these correctly to exam questions is really what this Factsheet is about

Ionic, Covalent, Dative and Metallic Bonding

Bonding = The force that binds atoms / ions together
Structure = The arrangement of these atoms / ions into a particular shape – the result of bonding!

Three main types of bonding:
1. **Ionic**: metal + non–metal atoms. If a **compound** contains a **metal**, it will nearly always be **ionic**.
2. **Covalent** (including **dative covalent**): non–metal + non–metal atoms.
3. **Metallic**: Metal atoms only.

Note how the 3 main types (including dative bonding) all involve the ***electrostatic attraction*** between positively and negatively charged particles.

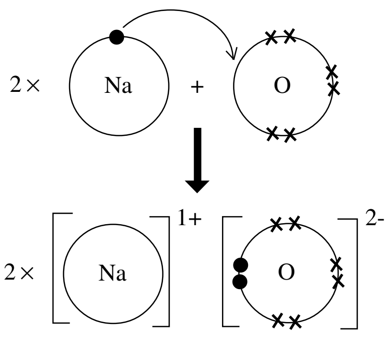
1. ***Ionic***: Electrostatic attraction between oppositely charged *ions*

2. ***Covalent***: Electrostatic attraction between **shared** electron **pair(s)** and the nuclei in the bond
   - ***Dative (co–ordinate) covalent***: one of the bonded atoms supplies *both* shared electrons in the shared electron pair

3. ***Metallic***: Electrostatic attraction between **delocalised** outer shell electrons and the lattice of positive metal ions

1. Ionic Bonding (Dot-and-cross Diagram)

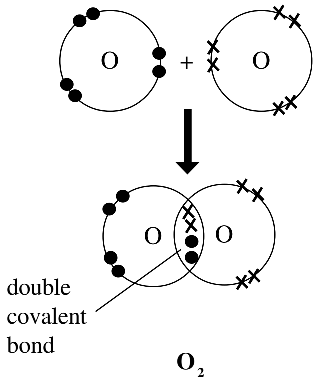
- Transfer of outer shell electrons from a metal to a non-metal to form oppositely charged ions.
- Each ion usually has a full, stable outer shell
- The resulting electrostatic attractions between positive and negative ions are very strong.



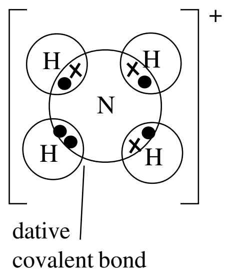
Na_2O

2. Covalent Bonding (Dot-and-cross Diagram)

- Pairs of outer shell electrons can be shared between 2 non-metal atoms (each atom usually donates half of the electrons in the shared pair(s))
- The covalent bond formed is strong.
- Each atom usually ends up with a full outer shell



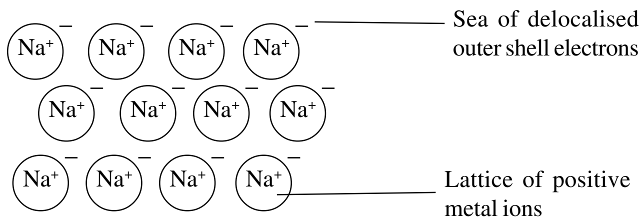
O_2



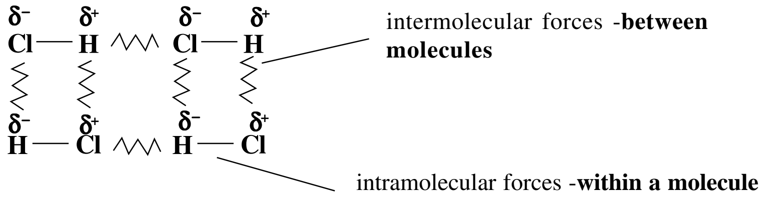
NH_4^+

3. Metallic Bonding

The **delocalised** outer shell electrons can move freely throughout the lattice of positive ions. Hence the electrons can carry charge (and energy) making metals good electrical (and thermal) conductors.



Intermolecular Forces



Three types of **intermolecular** forces, in order of **increasing** strength for similar sized molecules:

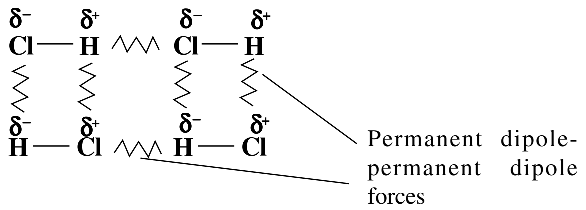
1. Van der Waals (Instantaneous Dipole – Induced Dipole interactions)

- They act between **all molecules**, polar or non polar.
- Caused by attractions between very small temporary dipoles in molecules.

Uneven distribution of electrons at an instant of time... → ...causes an instantaneous dipole... → ...which induces a dipole in neighbouring molecules → This induced dipole induces a dipole in other molecules → Weak intermolecular interactions occur

- Van der waals forces increase in strength with:
  - o Increasing number of electrons. The boiling point increases due to increasing van der Waals forces.
  - o Increasing chain length. There is a greater surface area so greater points of contact between molecules hence greater surface interactions / surface charges. Also, there are more electrons. This results in a higher boiling point.
  - *o Decreasing* branching in molecule – greater surface area and hence greater surface interactions between molecules.

2. Permanent Dipole – Permanent Dipole Interactions



Occur in non-symmetrical molecules with atoms of different electronegativities. Permanent dipole – permanent dipole forces increase with:

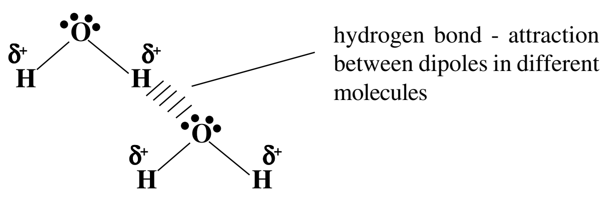
1. Greater size of permanent dipole determined by electronegativity differences
2. Smaller molecules since a closer approach between the molecules is possible

3. Hydrogen Bonding

Hydrogen bonds (a special type of dipole- dipole interaction) only exist if 2 conditions are met:

(a) A lone pair on a highly electronegative atom – N, O or F ("NOF")

(b) A hydrogen atom covalently bonded to a highly electronegative atom in a different molecule – H-N, H-O or H-F ("NOF" again!).



- Special Properties of Water Arising From Hydrogen Bonding
  (a) The solid (ice) is less dense than the liquid (water) so it floats on water.
  - Ice has an open lattice type structure. On average, the hydrogen bonds are longer in ice than in liquid water, so ice is less dense.

  (b) Water has a relatively high freezing point and boiling point than expected, compared to other group 6 hydrides like H_2S.
  - The hydrogen bonds are stronger than dipole-dipole and van der Waals forces, so much more energy is required to break these bonds and separate water molecules

Improving Your Performance – Typical Exam Questions on Bonding and Intermolecular Forces

The following are typical exam questions on bonding and intermolecular forces. Have a go at them. In the answers, common mistakes made by candidates are highlighted and tips to maximise your marks are given. If the answer is not explicitly given, it can be found in the ideas section of this FactSheet.

Question 1

(a) Explain what is meant by ionic, covalent and dative bonds.

(b) Draw "dot-and-cross" diagrams to show bonding in sodium oxide, Na_2O, oxygen, O_2, and the ammonium ion, NH_4^+. Outer shell electrons only need be shown.

Answers Question 1

(a) Just the *definitions* of the different types of bonding are needed – see earlier.

> **Tip**: Remember that bonding is the force that holds particles together. Examiners try to think of different ways to phrase this type of question (e.g. "Define…" or "Describe…"), but if you've learnt the definitions ***accurately***, you will get full marks.

(b) **Mistake**: It is a very common mistake to fail to identify whether a substance is ionic or covalent. Na is a metal, and O a non–metal, so Na_2O is *ionic*. O_2 is covalent since it consists of 2 non–metal atoms. NH_4^+ consists of 2 non-metals – hence *covalent* bonding but it also involves a charge!
See earlier for the required diagrams.

> **Tip**: NH_4^+ is often used as an example of a particle which has a dative covalent bond (H_3O^+ is another possibility).

> **Tip**: For covalently bonded substances, make sure the total number of outer shell electrons is correct e.g. in O_2, each O atom contributes six electrons (O is in group 6) so there should be 12 in total. For NH_4^+, N = 5, 4 × H = 4, but, since NH_4^+ has a 1+ charge, 8 electrons overall.

Question 2

This question concerns the bonding and properties of metals.

(a) Describe the bonding in metals such as sodium. You should illustrate your answer with a diagram.

(b) Explain why metals are good at conducting electricity in the solid and liquid state.

(c) Suggest why magnesium has a higher melting point than sodium.

Answers Question 2:

(a) A typical opener. For "Describe…" read "Define…".

> **Tip**: Make sure you label the lattice of positive metal ions and the delocalised electrons on your diagram – see earlier.

(b) A common follow up to part (a) – see earlier for the answer.

> **Tip**: Make sure you make it clear that the delocalised electrons can move throughout the lattice and carry charge.

(c)

> **Tip**: Melting points and boiling points depend on the *strength* of the forces between particles. Here, the examiner is trying to determine whether you realise this and can relate it to metallic bonding. Hence, ———

Magnesium has two outer shell electrons as opposed to sodium which only has one. So magnesium has more delocalised outer shell electrons and the ions in its lattice have 2+ charges as opposed to 1+. Hence the forces of attraction between ions and delocalised electrons are stronger and therefore the melting point is higher.

Question 3

Natural gas can contain significant quantities of alkanes with up to 5 carbon atoms.

(a) Draw a dot–and–cross diagram for methane, showing outer electrons only.

(b) Explain why ethane has a higher boiling point than methane.

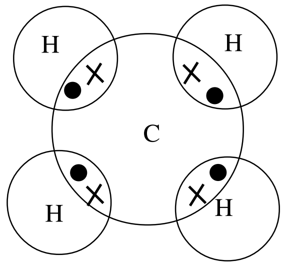
(c) Predict whether butane or 2–methylpropane has the higher boiling point. Explain your answer.

Answers Question 3

(a)

> **Tip**: Again, count the number of outer shell electrons available – 8 in methane - C = 4, 4 × H = 4.

> **Tip**: Note that the *central atom* in a molecule is usually the one present in *smaller* numbers i.e. C is central in CH_4.



CH_4

(b) Ethane and methane are non–polar and therefore only van der Waals forces are present. Ethane has a longer chain with greater surface area so there are greater surface interactions between molecules. Also, ethane has more electrons than methane. So the van der Waals forces between ethane molecules are greater and more energy is required to break these forces.

> **Tip**: Note that boiling points (and melting points) give a good indication of the strength of forces between particles. That is why examiners use these physical properties to test your knowledge of bonding and intermolecular forces.

(c) Butane and 2–methylpropane are non–polar and therefore only van der Waals forces are present. Butane has a higher boiling point. Butane has less branching in its molecule and therefore has a greater surface area and greater surface interactions between molecules.

> **Tip**: Note that the examiner picked these 2 molecules to compare because they are isomers (same molecular formula but different structural formula) and so have the same number of electrons. Therefore, these molecules show how molecular shape affects the strength of van der Waals forces.

Question 4
The following table gives the boiling point of three substances:

	NH_3	NO	N_2
Boiling point / °C	–33	–152	–196

(a) Identify the **dominant** intermolecular force in each substance.
(b) Explain the origin of the intermolecular forces in N_2.
(c) Draw a diagram to illustrate the intermolecular bonding in NH_3. Include any relevant dipoles and lone pairs.

Answers Question 4
(a) NH_3 → **:**N-H present → hydrogen bonding.
NO → O more electronegative than N → permanent dipole – permanent dipole forces.
N_2 → van der Waals only since symmetrical.

Tips: Some molecules such as NH_3 have all 3 types of intermolecular force. As hydrogen bonding is the strongest of these, it is the **dominant** force. You must be able to look at a molecule and decide which intermolecular force dominates. But there are some simple ways to match the intermolecular force with the molecule:

(1) If the molecule is non–polar it is very likely too only have van der Waals (e.g. alkanes, molecular elements). Hydrogen peroxide, H_2O_2 and hydrazine, N_2H_4 , are two common exceptions

(2) If the molecule has a H atom attached to N, O or F ***and*** also a lone pair on a N, O or F, then it must be hydrogen bonding. Typically, examiners focus on H_2O for hydrogen bonding. But as in this case, not always!

Mistake: A common pitfall here is to think a molecule like CH_3CHO (ethanal) or $CHCl_3$ (trichloromethane) can hydrogen bond with itself. But it cannot – the H is attached to a carbon in both cases, not an oxygen atom.

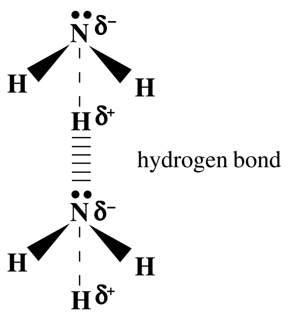
(3) For molecules with more than 2 atoms, you will have to know or work out the shape. To do this, you may even have to draw a dot–and–cross diagram. This is because molecular shape determines whether a molecule has a permanent dipole – bond dipoles cancel if the molecule has a symmetrical shape. e.g. the bonds in CCl_4 are polar, but the overall tetrahedral shape means the 4 bond dipoles cancel each other giving a non-polar molecule.

(4) If boiling points are given, this should help you. Remember that the stronger the intermolecular force, the higher the boiling point. Hydrogen bonding is the strongest and van der Waals the weakest for similar sized molecules.

(b) See page 2 for answer

Tip: It is common for examiners to ask how van der Waals forces arise. That is why "Instantaneous Dipole - Induced Dipole interactions" - the alternative name for van der Waals forces - is useful to remember. It almost describes the origin of these forces.

(c)



Mistake: Hydrogen bonding diagrams showing a hydrogen bond between 2 hydrogens are common. This is impossible since they are both δ+ - they would repel each other!

Mistake: Also, read the question carefully. All too often candidates do not show the lone pairs and dipoles.

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