Chem Factsheet

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AS Chemical bonding: Intramolecular bonds

This fact sheet is designed to help you check your understanding about intramolecular bonding – bonds that form **within** molecules. The areas covered are:-

- Bonding: why do bonds form?
- Intramolecular bonds: the extremes and "in-betweens" –e.g. ionic, covalent, polar covalent
- Bonding in metals – a special case
- Drawing "dot-and-cross" diagrams for ionic and covalent bonds
- Exam questions and answers on intramolecular bonds

Exam questions and answers are included. Don't forget to complete your understanding of chemical bonding by reading Chem Factsheet 106 on intermolecular bonds.

Bonding: why do bonds form?

Check your understanding by trying this question.

Methane
Natural gas is mainly methane, CH_4.
Explain as clearly as you can why carbon and hydrogen form molecules with the formula CH_4 rather than CH_3, CH_2 or CH.

The full answer is given at the end of the FactSheet.

Bonds form because *making bonds lowers the overall energy of the particles* involved. For example, methane molecules exist as CH_4 because this is the most *energetically favoured* arrangement for 1 carbon atom bonding with hydrogen atoms. So, a molecule with the formula CH_4 has less potential energy than any other suggested combination. Bonds *don't* form because atoms "want" to form bonds, or because, in this case "carbon needs 4 more electrons" or to "satisfy the Octet Rule".

Intramolecular bonds

The extremes: ionic and covalent

There are two "extreme" types of chemical bond: ionic and covalent, as shown in the table below:

	An ionic bond	**A covalent bond**
Forms when..	transfer of one or more electrons between atoms i*s* energetically favourable	electron transfer is *not* energetically favourable
Is attraction between…	ions of opposite charge: cation +ve and anion –ve	a shared electron pair and two atomic nuclei
Examples	Compounds of Group I and Group II metals, e.g. sodium chloride NaCl lithium fluoride LiF calcium chloride $CaCl_2$ magnesium oxide MgO	**Compounds of 'p' block elements, e.g.** Carbon dioxide CO_2 Boron trichloride BCl_3 Methane CH_4 **Atoms of 'p' block elements, e.g.** Carbon $C_{(n)}$ Nitrogen N_2 Oxygen O_2 Sulfur S_8 Chlorine Cl_2
Examiners dislike…	• Sodium chloride exists as "molecules".	• Covalent bonds can be "disturbed" or "weakened". • Covalent bonds are "weaker" than ionic bonds.
Examiners like…	• (For example) "NaCl" represents the ratio of sodium to chloride ions in the ionic lattice, i.e. 1:1. • The ions are packed alternately in as energetically favourable way as possible. • Ionic bonds are multi-directional. • Ionic bonds are strong.	• Covalent bonds can be broken. • Covalent bonds exist in simple molecules such as iodine and water and in giant structures such as diamond, graphite and silicon dioxide. • Covalent bonds are linear. • Covalent bonds are strong.

The "in-betweens": polar covalent and ionic bonds with covalent character

In most bonds distortion of electron charge clouds surrounding atomic nuclei occurs. Hence, bonds are neither 100% covalent or ionic, but in-between, having a degree of "ionic" or "covalent" character.

Polar covalent bonds

In a covalent bond, electrons may spend more time around the nucleus of one atom in the pair. This leaves the other nucleus more exposed creating a partial positive charge labelled "δ^+". The "other side", where there is a greater chance of finding electrons, is labelled "δ^-". The bond has a "permanent dipole" (PD) and is described as "polar covalent".

Examples	H-Cl bond in hydrogen chloride	$^{\delta+}$H - Cl$^{\delta-}$
	H-F bond in hydrogen fluoride	$^{\delta+}$H - F$^{\delta-}$
	H-O bond in water (H_2O)	$^{\delta+}$H - O$^{\delta-}$
	C=O bond in carbon dioxide (CO_2	$^{\delta+}$C = O$^{\delta-}$
	N-H bond in ammonia (NH_3)	$^{\delta+}$N - H$^{\delta+}$

Atoms of nitrogen, oxygen and fluorine *attract the bonding electrons in a covalent bond significantly*. These elements are highly "*electronegative*".

- Polar covalent bonds are favoured when there is a large difference in electronegativity between the atoms in the bond. The most electronegative elements are nitrogen, oxygen and fluorine.
- Electronegativity increases across a period due to more protons being present in the nucleus while the corresponding electrons go into the same shell.
- This means the distortion power increases left to right.
- Electronegativity decreases down a group due to the increased number of electron shells which shield the atomic nucleus.

Permanent dipoles in bonds do not always create molecules with an **overall** dipole. When the molecule is symmetrical the dipoles effectively cancel each other out. BCl_3 is an example – the B-Cl bond has a dipole, but the molecule overall does not.

Ionic bonds with covalent character

The charge clouds surrounding ions in an ionic lattice may be distorted when the cation is small and highly positive – it has *high charge density* while the anion has *low charge density* in comparison. The electron charge cloud surrounding the anion is drawn towards the cation, so that electrons stand some chance of being found close to the cation nucleus. The bond is not completely ionic, but has "covalent character", as the two electrons are partially shared between the two nuclei. The bond can be so "covalent" that it is referred to as a covalent bond, as in:-

Beryllium iodide BeI_2
Aluminium chloride $AlCl_3$

Cations have high *polarising power* if they are small and highly charged.

You can spot ionic bonds with covalent character by looking for ions like Al^{3+}, B^{3+}, Be^{2+} and Fe^{3+}, especially when combined with relatively large anions like Cl^- and I^-.

Bonding in metals – a special case

In metals, each atom shares one or more electrons between many others – the electrons are *delocalised*. The "atoms" have lost these electrons, so strictly are *positively charged ions*, bonded together by the delocalised "*electron sea*". The metallic bond is the attraction between electrons and ions. The electrons are mobile, which means metals can conduct electricity. The positive ions are packed as energetically favourably as possible in a lattice structure.

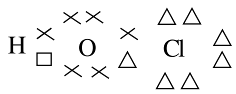
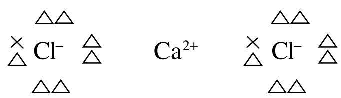
Drawing "dot-and-cross" diagrams for ionic and covalent bonds

Examiners ask you to draw ionic and covalent bonds using "dot-and-cross" diagrams. These show the locations of electrons in the bonds. Each dot or cross represents one electron. The dot-and-cross system enables us to see the origin of each electron in the bond. The "rules" for drawing these are:-

- Use dots for electrons from one atom and crosses for the other atom
- Different symbols to · and + are allowed, but be consistent and sensible in your choices
- When three different atoms are involved, a third symbol, e.g. Ä or can be used for the third atom
- Show the outermost electrons because only these are involved in bonding, but label the diagram "highest energy level electrons shown only"
- Circles are not essential – here we show one with circles and the other without.
- The diagrams score only 1 or 2 marks – so don't spend ages getting circles perfect

Exam questions and answers on intramolecular bonds

These are questions examiners have asked about bonding in recent AS exams. Make the most of this section by covering up the "Answer" column and writing your own answers, taking note of the number of marks available. All the answers can be found in this Factsheet.

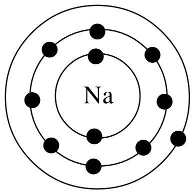
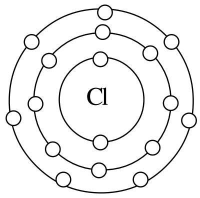
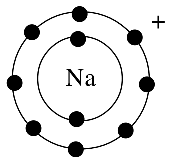
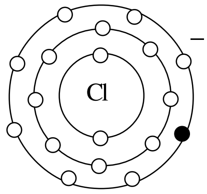
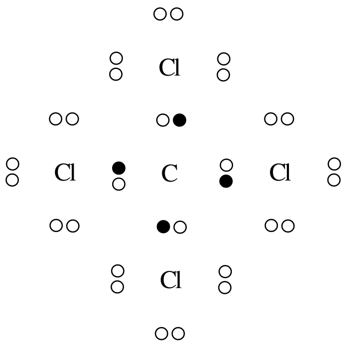
Question	Marks	Answer
Explain how the ions in $MgCl_2$ are formed.	2	Two electrons transfer from each magnesium atom (1) Each chlorine atom gains one electron (1)
State what is meant by an "ionic bond".	1	Electrostatic attraction between (1) oppositely charged ions (1)
State what is meant by a "covalent bond".	2	A pair of electrons shared (1) between two atoms (1)
Define "electronegativity".	2	Power of an atom in a covalent bond (1) to attract the bonding electrons (1)
Explain why electronegativity increases across a period in the Periodic Table.	2	The nuclear charge increases (1) Shielding stays the same (1)
Explain why electronegativity values decrease down group II (Be-Ba).	1	The number of electron shells increases(1) OR The shielding of the nucleus increases (1)
What is meant by "polarised"?	2	Electron cloud (1) around an anion is distorted (1)
Why is a chloride ion polarised more by an aluminium ion than a magnesium ion?	2	Al^{3+} is smaller (1) and has a higher charge (1)
The B-Cl bond is polar. Explain why BCl_3 is not a polar molecule.	2	The molecule is symmetrical (1) so the bond polarities cancel (1)
State and explain the effect a single Be^{2+} ion would have on a Cl- ion. Explain how this effect would lead to formation of a covalent bond.	2 1	Be^{2+} ion has high charge density (1) The Be^{2+} ion polarises the chloride ion (1) Electrons from a chloride ion are shared by Be^{2+} to form each covalent bond (1)
The bonding in aluminium oxide is intermediate between ionic and covalent bonding. Explain why.	4 max 3	Al^{3+} ions are small (1) highly charged (1) O^{2-} ions are large (1) O^{2}- ions are polarised / distorted (1)
Show the polarity of the bonds: H-N, F-B and H-I	2	$^{\delta+}$H-N$^{\delta-}$, $^{\delta-}$F-B$^{\delta+}$ and $^{\delta+}$H-I$^{\delta-}$ 3 correct scores 2, 2 correct scores 1
State what is meant by "metallic bonding". Draw a "dot-and-cross" diagram of a molecule of HOCl	3 2	The attraction (1) between positive ions (1) and a "sea" of delocalised electrons (1) Covalent bonds shown correctly (1) All lone pairs shown on chlorine and oxygen atoms (1) 
Draw a "dot-and-cross" diagram of calcium chloride.	2	Ca $^{2+}$ ion shown correctly – either 8 or no electrons **with no overlap** (1) $2Cl^-$ ions shown correctly – dot and crosses required (1) 

Answer to introductory question

Methane

Stability is associated with the formation of covalent bonds by which electron orbitals are filled by sharing a pair of electrons between two atoms. CH_4 is the most stable formula of those listed. This formula confers the greatest stability on all atoms as their outer electron shells are filled by sharing electrons.

Possible incorrect responses: Carbon needs four bonds / 4 more electrons; Carbon has 4 bonding pairs; Carbon wants to form 4 bonds; Because there are 4Hs and 1C.

What the diagram shows	Dot-and-cross diagram for sodium chloride, NaCl	What to do in an exam
The outermost electron (3s) of a sodium atom transfers to a chlorine atom. The sodium atom is ionised, becoming a positively charged sodium ion, as the nucleus has 11 protons and only 10 electrons remain. Each chlorine atom accepts an electron, creating a negative charge as the nuclei have 17 protons each, but 18 electrons. The atom becomes a chloride ion. An ionic bond forms due to electrostatic attraction between the two oppositely charged ions. This process is repeated millions of times to create a giant ionic lattice of sodium and chloride ions in a 1:1 ratio.	 Sodium atom 2,8,1  Chlorine atom 2,8,7  Sodium ion 2,8  Chloride ion 2,8,8	Only show the atoms in this first stage if the question asks for them. Check the question and the number of marks available. e.g. *"Show how an ionic bond forms between X and Y" (4 marks)* implies draw the atoms and the ions, giving the commentary shown left. *"Draw a dot-and-cross diagram of XYide" (2 marks)* implies draw the ions only, showing *all* electrons in the outermost shells. No commentary is needed. Vital components are:- • the correct number of ions • the correct charges on the ions • all outermost electrons shown So, check the formula of the compound!
Four bonding pairs are created between the carbon atom and the chlorine atoms. The chlorine atom electrons are shared in the outermost shell of the carbon atom. All atoms in the molecule have an noble gas electron configuration. The molecule is symmetrical.		Make sure the dots and crosses are aligned neatly. Check the formula of the molecule – make sure you have included all atoms. Check you have got the right number of electrons in *all* orbitals – here it is easy to forget the other 3 electron pairs in the outermost shells of the chlorine atoms!

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