AS Chemistry worksheet: dot-and-cross diagrams, bonding and shapes of molecules

How to use this worksheet

This worksheet covers the chemical bonding section of Cambridge International AS Chemistry (9701): ionic and covalent dot-and-cross diagrams, dative covalent bonds, molecules that do not obey the octet rule, shapes and bond angles from electron-pair repulsion, bond polarity and intermolecular forces. There are six sets of questions, numbered continuously, worth 50 marks in total. Marks are shown in square brackets, and a full answer block follows each set.

Draw every diagram you are asked for on paper first, then compare it with the written description in the answers. Each answer describes the diagram precisely: which atoms are shown, how many electrons are in each outer shell, which pairs are shared, which pairs are lone pairs, and where brackets and charges go. Only outer-shell electrons are shown in the diagrams in this worksheet.

> **Exam tip: what a complete dot-and-cross diagram must show.** Use dots for the electrons of one element and crosses for the other, so the examiner can see where each electron came from. Show every outer-shell electron, including lone pairs. A shared pair sits in the overlap of the two outer shells. For an ion, put square brackets around the whole species and write the charge at the top right outside the bracket. For a dative bond, both electrons in the shared pair carry the same symbol, because both came from the same atom.

Set 1: ionic compounds

1. Sodium chloride is formed from sodium atoms (electron configuration 2.8.1) and chlorine atoms (2.8.7). Describe the dot-and-cross diagram for sodium chloride, stating which electrons are transferred and the charge on each ion. [3]

2. Describe the dot-and-cross diagram for magnesium oxide, MgO, formed from magnesium (2.8.2) and oxygen (2.6). [2]

3. Calcium chloride has the formula $CaCl_2$. Explain, with reference to the electrons transferred, why two chloride ions are needed for each calcium ion, and describe the diagram. [3]

> **1.** The single outer-shell electron of the sodium atom (shown as a dot) is transferred to the outer shell of the chlorine atom, which already holds seven electrons (shown as crosses) [1]. The diagram shows two ions: the sodium ion is drawn with an empty outer shell, or with its full inner shell of eight only, in square brackets with a + charge, because it has lost one electron [1]. The chloride ion is drawn with eight outer electrons, seven crosses and one dot, in square brackets with a - charge, because it has gained one electron [1]. Both ions now have the electron configuration of a noble gas: Na^+ is 2.8 and Cl^- is 2.8.8.
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> **2.** Magnesium loses both of its outer-shell electrons (two dots) and the oxygen atom gains both, adding them to its six outer electrons (crosses) [1]. The diagram shows $[Mg]^{2+}$ with an empty outer shell and $[O]^{2-}$ with eight outer electrons, six crosses and two dots, each species in square

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brackets with its charge [1].

3. A calcium atom (2.8.8.2) has two outer-shell electrons and must lose both to reach a stable configuration, but a chlorine atom (2.8.7) can only accept one electron to complete its outer shell [1]. Each of the two electrons therefore goes to a different chlorine atom, so one Ca^{2+} ion is formed for every two Cl^- ions, giving a neutral compound with the formula $CaCl_2$ [1]. The diagram shows one calcium ion in square brackets with a 2+ charge and an empty outer shell, and two separate chloride ions, each in square brackets with a - charge and eight outer electrons, seven crosses and one dot [1].

Set 2: simple covalent molecules

4. Describe the dot-and-cross diagram for a water molecule, H_2O. [2]

5. Describe the dot-and-cross diagram for carbon dioxide, CO_2, and explain how the diagram shows that each carbon-oxygen bond is a double bond. [3]

6. Describe the dot-and-cross diagram for a nitrogen molecule, N_2. [2]

7. Ethene has the formula C_2H_4. Describe its dot-and-cross diagram and state the total number of electrons in the outer shell of each carbon atom. [3]

4. The oxygen atom is drawn in the centre with its six outer electrons (crosses) and a hydrogen atom on each side with one electron (dot). Two of the oxygen electrons each pair with a hydrogen electron, giving two shared pairs, each drawn as one cross and one dot in the overlap between the oxygen and hydrogen shells [1]. The remaining four oxygen electrons form two lone pairs, drawn as two pairs of crosses on the oxygen shell away from the hydrogens. Oxygen now has eight outer electrons and each hydrogen has two [1].

5. The carbon atom is drawn in the centre with four outer electrons (dots) and an oxygen atom (six crosses) on each side [1]. Each carbon-oxygen bond is a double bond, so the overlap region between carbon and each oxygen contains four electrons: two dots from carbon and two crosses from oxygen, drawn as two shared pairs [1]. Each oxygen keeps four electrons as two lone pairs on the side away from the carbon. Carbon has eight electrons in its outer shell (four dots shared with four crosses) and each oxygen also has eight (two shared pairs plus two lone pairs) [1].

6. Two nitrogen atoms, each with five outer electrons (one drawn with dots, one with crosses), share three pairs: the overlap region contains six electrons, three dots and three crosses, which is a triple bond [1]. Each nitrogen keeps one lone pair on the side away from the bond, so each has eight outer electrons: six shared plus two in its lone pair [1].

7. The two carbon atoms are drawn side by side, each with four outer electrons, with two hydrogen atoms attached to each carbon [1]. The overlap between the two carbons contains four electrons (two from each carbon), a double bond; each carbon-hydrogen overlap contains two electrons, one from the carbon and one from the hydrogen [1]. Each carbon atom therefore has eight outer electrons: four in the double bond and two in each of its two C-H bonds. There are no lone pairs in ethene [1].

Set 3: dative covalent bonds

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8. An ammonium ion, NH_4^+, is formed when an ammonia molecule reacts with a hydrogen ion. Describe the dot-and-cross diagram of the ammonium ion and explain what is meant by a dative covalent bond. [3]

9. Ammonia reacts with boron trifluoride to form the compound $H_3N{\rightarrow}BF_3$. Explain why this reaction occurs and describe the dot-and-cross diagram of the product. [3]

10. At moderate temperatures aluminium chloride vapour exists as the dimer Al_2Cl_6. Explain why the dimer forms and describe the bonding in it, including the number of electrons around each aluminium atom. [3]

8. A dative covalent (co-ordinate) bond is a covalent bond in which both electrons of the shared pair come from the same atom [1]. In ammonia the nitrogen atom has three bonding pairs with hydrogen and one lone pair. A hydrogen ion has no electrons, so the nitrogen lone pair is donated into the empty space around the H^+ and becomes the fourth bonding pair [1]. The diagram shows nitrogen in the centre with four hydrogen atoms around it: three overlaps each contain one cross (nitrogen) and one dot (hydrogen), and the fourth overlap contains two crosses, because both electrons came from nitrogen. The whole species is in square brackets with a + charge, and there are no lone pairs left on nitrogen [1]. Once formed, all four N-H bonds are identical in length and strength.

9. In BF_3 the boron atom has only three outer electrons and forms three bonding pairs, so it has just six electrons around it; it is electron deficient and can accept a lone pair. The nitrogen in ammonia has a lone pair to donate, so a dative covalent bond forms from N to B [1]. The diagram shows nitrogen bonded to three hydrogens (each overlap one cross from N and one dot from H) and boron bonded to three fluorines (each overlap one electron from B and one from F); each fluorine keeps three lone pairs [1]. Between nitrogen and boron there is an overlap containing two crosses, both from nitrogen. In a displayed formula this bond is drawn as an arrow from N to B. Both nitrogen and boron now have eight outer electrons [1].

10. A single $AlCl_3$ molecule has only six electrons around the aluminium atom (three bonding pairs), so it is electron deficient [1]. Two molecules join through two bridging chlorine atoms: a lone pair on a chlorine atom of one $AlCl_3$ is donated into the aluminium of the other molecule, forming a dative covalent bond, and the same happens in the opposite direction [1]. Each aluminium is now bonded to four chlorines, two terminal and two bridging, and has eight outer electrons: three ordinary covalent bonding pairs plus one dative pair. Each bridging chlorine forms one ordinary covalent bond and one dative bond and keeps two lone pairs; each terminal chlorine forms one bond and keeps three lone pairs [1].

Learn this. A covalent bond is a shared pair of electrons between two atoms. A dative covalent (co-ordinate) bond is a covalent bond in which both electrons in the shared pair are provided by one of the atoms. The donor must have a lone pair and the acceptor must have an empty orbital able to receive it.

Set 4: molecules that do not obey the octet rule

11. Describe the dot-and-cross diagram of boron trifluoride, BF_3, and explain why boron is said to have an incomplete octet. [2]

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12. Sulfur hexafluoride, SF_6, is a stable gas. Describe its dot-and-cross diagram and state how many electrons surround the sulfur atom. [2]

13. Phosphorus forms PCl_5 but nitrogen does not form NCl_5. Explain why. [2]

11. Boron is drawn in the centre with three outer electrons (dots) and three fluorine atoms around it, each with seven outer electrons (crosses). Each B-F overlap contains one dot and one cross, and each fluorine keeps three lone pairs [1]. Boron therefore has only six electrons in its outer shell, two short of an octet, which is why BF_3 readily accepts a lone pair, as in question 9 [1].

12. Sulfur has six outer electrons ($3s^2\ 3p^4$). Each of the six electrons pairs with one electron from a fluorine atom, so the diagram shows sulfur in the centre with six overlaps, each containing one dot and one cross, and each fluorine with three lone pairs [1]. Sulfur is surrounded by twelve electrons, six bonding pairs; its octet has been expanded [1].

13. Phosphorus is in Period 3. Its outer shell (n = 3) contains 3d orbitals as well as 3s and 3p, so it can have more than eight electrons around it: in PCl_5 there are five bonding pairs, ten electrons, around phosphorus [1]. Nitrogen is in Period 2 and its outer shell (n = 2) has only one 2s and three 2p orbitals, a maximum of eight electrons, so nitrogen cannot form five bonds [1].

Set 5: shapes and bond angles

14. State the two main ideas of electron-pair repulsion theory that are used to predict the shape of a molecule. [2]

15. Copy and complete the table. For each molecule give the number of bonding pairs and lone pairs around the central atom, the name of the shape and the bond angle. [6]

Molecule	Bonding pairs	Lone pairs	Shape	Bond angle
CH_4				
NH_3				
H_2O				
BF_3				
CO_2				
SF_6				

16. Methane, ammonia and water each have four pairs of electrons around the central atom, yet their bond angles are different. Explain why. [2]

14. Pairs of electrons in the outer shell of the central atom repel one another and arrange themselves as far apart as possible, so that repulsion is at a minimum [1]. Lone pairs repel more strongly than bonding pairs, so the order of repulsion is lone pair-lone pair, greater than lone pair-bonding pair, greater than bonding pair-bonding pair [1].

15. One mark per correct row (all four entries needed).

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Molecule	Bonding pairs	Lone pairs	Shape	Bond angle
CH_4	4	0	tetrahedral	109.5°
NH_3	3	1	pyramidal (trigonal pyramidal)	107°
H_2O	2	2	non-linear (bent, V-shaped)	104.5°
BF_3	3	0	trigonal planar	120°
CO_2	2 (both double bonds)	0	linear	180°
SF_6	6	0	octahedral	90°

For CO_2, each double bond counts as one region of electron density, so the two regions take up positions 180° apart.

16. In methane all four pairs are bonding pairs, which repel equally, so the angle is the regular tetrahedral angle of 109.5°. Ammonia has one lone pair, which repels the three bonding pairs more strongly than they repel each other, squeezing them closer together to about 107° [1]. Water has two lone pairs, so the two bonding pairs are squeezed even more, to about 104.5°. Each lone pair reduces the bond angle by roughly 2.5° [1].

Set 6: bond polarity and intermolecular forces

17. Define electronegativity, and explain why the bond in HCl is polar while the bond in Cl_2 is not. [2]

18. Both carbon dioxide and water contain polar bonds. Explain why a water molecule has an overall dipole but a carbon dioxide molecule does not. [2]

19. Water boils at 100 °C, whereas hydrogen sulfide, H_2S, which has a larger relative molecular mass, boils at -60 °C. Explain this difference in terms of the intermolecular forces present in each substance. [3]

17. Electronegativity is the ability of an atom to attract the bonding pair of electrons in a covalent bond towards itself [1]. In HCl the chlorine atom is more electronegative than hydrogen, so the shared pair lies closer to chlorine, giving chlorine a partial negative charge (δ-) and hydrogen a partial positive charge (δ+): the bond is polar. In Cl_2 both atoms have the same electronegativity, so the pair is shared equally and the bond is non-polar [1].

18. Carbon dioxide is linear, so its two C=O bond dipoles are equal in size and point in exactly opposite directions; they cancel, and the molecule has no overall dipole [1]. Water is non-linear, so its two O-H bond dipoles do not cancel; they add to give a net dipole, with the oxygen end δ- and the hydrogen end δ+ [1].

19. In water, hydrogen is bonded directly to oxygen, a small and highly electronegative atom with lone pairs, so hydrogen bonds form between the δ+ hydrogen of one molecule and a lone pair on the oxygen of a neighbouring molecule [1]. Sulfur is less electronegative and larger, so H_2S cannot form hydrogen bonds; its molecules are held only by weaker permanent dipole-dipole forces and instantaneous dipole-induced dipole forces [1]. Hydrogen bonds are much stronger than these other intermolecular forces, so far more energy is needed to separate water molecules

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and water has the much higher boiling point, even though H_2S has more electrons and so stronger instantaneous dipole forces [1].

Exam tip. The examiner expects the three types of intermolecular force to be named exactly: instantaneous dipole-induced dipole (id-id, also called London or dispersion forces), permanent dipole-permanent dipole (pd-pd) and hydrogen bonding. Hydrogen bonding needs a hydrogen atom bonded to N, O or F, and a lone pair on an N, O or F atom of a neighbouring molecule. Do not confuse intermolecular forces with the covalent bonds inside the molecule, which do not break when a simple molecular substance boils.

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