2.2 Bonding and Structure

AQA 7405 · Calculator allowed · about 255 minutes · Write-in edition

Total Marks

Name: _____ Date: ____ / ____ / _____

Answer ALL questions. Show all your working.

1 This question tests key definitions used throughout the bonding and structure topic.

(a) Define electronegativity. **(1)**

(b) Define a coordinate (dative covalent) bond. **(1)**

(c) State what is meant by a permanent dipole. **(1)**

(d) State what is meant by van der Waals (London/dispersion) forces. **(1)**

(Total for Question 1 is 4 marks)

2 This question uses ionic radii to compare the strength of ionic bonding in sodium chloride, NaCl, and magnesium oxide, MgO. Ionic radii: Na^+ = 102 pm, Cl^- = 181 pm, Mg^{2+} = 72 pm, O^{2-} = 140 pm. The electrostatic force of attraction between two ions is proportional to (Q1 x Q2) / r^2, where Q1 and Q2 are the sizes of the ionic charges and r is the distance between the ion centres (the sum of their ionic radii).

(a) State two factors that affect the strength of ionic bonding (the magnitude of lattice enthalpy). **(2)**

(b) Show that the value of (Q1 x Q2) / r for MgO is more than five times greater than the value for NaCl, where r is the sum of the two ionic radii (in pm). **(4)**

(c) Explain how the calculation in part (b) accounts for the fact that magnesium oxide has a higher melting point than sodium chloride. **(2)**

(d) Going down Group 2, from MgO to CaO to SrO to BaO, explain how you would expect the lattice enthalpy and melting point of the Group 2 oxide to change, and why. **(2)**

(Total for Question 2 is 10 marks)

3 This question is about covalent bonding, including dative covalent (coordinate) bonding.

(a) Define the term covalent bond. **(1)**

(b) Describe the covalent bonding present in a molecule of methane, CH_4. **(2)**

(c) Define the term dative covalent (coordinate) bond. **(1)**

(d) Ammonia, NH_3, reacts with a hydrogen ion, H^+, to form the ammonium ion, NH_4^+. Explain how this bond forms, identifying which atom donates the electron pair. **(3)**

(e) State the shape of the ammonium ion, NH_4^+, and its bond angle. **(2)**

(Total for Question 3 is 9 marks)

4 This question is about electronegativity and bond polarity.

(a) Define the term electronegativity. **(1)**

(b) State and explain the general trend in electronegativity across Period 3, from sodium to chlorine. **(3)**

(c) Using these electronegativity values, H = 2.1, Cl = 3.0, F = 4.0, I = 2.5, rank the hydrogen halides HF, HCl and HI in order of increasing bond polarity, showing your working. **(3)**

(d) A student states: "Carbon dioxide, CO_2, has polar C=O bonds, so it must be a polar molecule." Evaluate this statement. **(3)**

(Total for Question 4 is 10 marks)

5 This question is about predicting the shapes of molecules and ions using electron-pair repulsion theory.

(a) State the electron-pair repulsion theory used to predict the shapes of molecules and ions. **(2)**

(b) Predict the shape and bond angle of a molecule of boron trifluoride, BF_3, explaining your answer. **(3)**

(c) Predict the shape and bond angle of a molecule of ammonia, NH_3, explaining why its bond angle differs from that in methane, CH_4 (109.5 degrees). **(3)**

(d) Predict the shape and bond angle of a molecule of water, H_2O, explaining the effect of having two lone pairs. **(3)**

(Total for Question 5 is 11 marks)

6 This question is about metallic bonding and the properties of metals and alloys.

(a) Describe the structure and bonding present in a typical metal. **(2)**

(b) Explain why metals are good conductors of electricity in both the solid and molten states. **(3)**

(c) Explain why metals are malleable. **(2)**

(d) Explain, in terms of structure, why alloys are generally harder than the pure metals from which they are made. **(3)**

(Total for Question 6 is 10 marks)

7 This question is about intermolecular forces: London (dispersion) forces and permanent dipole-dipole forces.

(a) State what is meant by a London (dispersion) force. **(2)**

(b) Explain why the boiling points of the noble gases (He, Ne, Ar, Kr, Xe) increase down the group. **(3)**

(c) Hydrogen chloride, HCl (Mr = 36.5), has a higher boiling point than argon, Ar (Mr = 40), even though Ar has a slightly greater Mr. Explain why. **(3)**

(d) Given that the boiling point of argon is -186 C and the boiling point of HCl is -85 C, use this data to support your answer to part (c). **(2)**

(Total for Question 7 is 10 marks)

8 This question is about hydrogen bonding and its effect on the properties of water.

(a) State the three conditions that must be met for hydrogen bonding to occur between molecules. **(3)**

(b) Describe how a hydrogen bond forms between two water molecules. **(2)**

(c) Explain, in terms of hydrogen bonding, why ice is less dense than liquid water. **(3)**

(d) The density of ice is 0.917 g/cm3 and the density of liquid water is 1.000 g/cm3 at 0 C. Calculate the percentage increase in volume when a fixed mass of water freezes to form ice. **(4)**

(Total for Question 8 is 12 marks)

9 This question is about the giant covalent structures of diamond, graphite and graphene, all forms of carbon.

(a) Describe the structure of diamond. **(2)**

(b) Explain why diamond has a very high melting point. **(2)**

(c) Describe the structure of graphite and explain why graphite conducts electricity while diamond does not. **(4)**

(d) Explain why graphite is used as a lubricant. **(2)**

(e) Graphene is a single, isolated layer of graphite, one atom thick. Suggest one property of graphene that would make it useful in electronics, referring to its structure. **(2)**

(Total for Question 9 is 12 marks)

10 This question is about predicting the shapes and bond angles of molecules and ions using electron-pair repulsion theory.

(a) Determine the number of bonding pairs and lone pairs of electrons around the central phosphorus atom in PF_5, and hence give its shape and bond angles. **(3)**

(b) SF_6 has 6 bonding pairs and no lone pairs around the central sulfur atom. State its shape and bond angle. **(2)**

(c) The ion ICl_4^- has 4 bonding pairs and 2 lone pairs of electrons around the central iodine atom. State its shape and bond angle. **(2)**

(d) In the nitrite ion, NO_2^-, there are 2 bonding regions and 1 lone pair of electrons around the central nitrogen atom. State the shape of the ion and estimate its bond angle, explaining why the angle differs from that of a molecule with 3 bonding pairs and no lone pairs. **(2)**

(Total for Question 10 is 9 marks)

11 Table 2 shows the boiling points of the hydrides of Group 14 and Group 16.

Table 2: boiling points of Group 14 and Group 16 hydrides

Group 14 hydride	Mr	Boiling point / C
CH_4	16	-162
SiH_4	32	-112
GeH_4	77	-88
SnH_4	123	-52

Group 16 hydride	Mr	Boiling point / C
H_2O	18	100
H_2S	34	-60
H_2Se	81	-41
H_2Te	130	-2

Explain, using ideas about intermolecular forces, why the boiling points of the Group 14 hydrides increase steadily down the group, but the boiling point of H_2O is anomalously high compared with the trend shown by the other Group 16 hydrides.

(Total for Question 11 is 6 marks)

12 This question is about ionic polarisation and the bonding in aluminium chloride, $AlCl_3$.

(a) Explain, using the idea of ionic polarisation, why aluminium chloride, $AlCl_3$, has significant covalent character rather than being purely ionic. **(4)**

(b) Predict, with a reason, whether aluminium fluoride, AlF_3, or aluminium chloride, $AlCl_3$, has the greater ionic character. **(2)**

(c) State the type of structure adopted by solid aluminium chloride as a result of its covalent character. **(1)**

(d) Explain how the dative covalent bonds form within the Al_2Cl_6 dimer. **(2)**

(Total for Question 12 is 9 marks)

13 Boron trifluoride, BF_3, reacts with ammonia, NH_3, to form the adduct $F_3B\text{-}NH_3$, in which a new dative covalent bond forms from the nitrogen atom to the boron atom.

(a) Suggest, with reasoning, the shape around the boron atom in the adduct $F_3B\text{-}NH_3$, given that BF_3 itself is trigonal planar. **(3)**

(b) Explain why BF_3 must have an incomplete octet (an empty orbital) on boron in order to accept the dative bond from nitrogen. **(2)**

(c) In the adduct $F_3B\text{-}NH_3$, the nitrogen atom's lone pair is now used in the dative bond to boron. Suggest how this affects the ability of $F_3B\text{-}NH_3$ molecules to hydrogen bond to one another, and hence compare the expected volatility of the adduct with that of a simple molecular compound of similar Mr that has no hydrogen bonding capability. **(2)**

(d) If ammonia, NH_3, were replaced with trimethylamine, $N(CH_3)_3$, in this reaction, a dative bond to boron would still form. Suggest why the melting point of the resulting adduct, $F_3B\text{-}N(CH_3)_3$, might differ from that of $F_3B\text{-}NH_3$, and in which direction you would expect this change. **(2)**

(Total for Question 13 is 9 marks)

14 This synoptic question compares the structure, bonding and thermal behaviour of carbon dioxide and silicon dioxide, both Group 14 oxides.

(a) Carbon dioxide, CO_2, is a gas at room temperature and sublimes at -78°C. Silicon dioxide, SiO_2, is a solid that melts at about 1710°C. Describe the structure and bonding of solid CO_2 (dry ice) and of SiO_2. **(2)**

(b) Bond enthalpy data: C=O (in CO_2) = 805 kJ/mol; Si-O (in SiO_2) = 466 kJ/mol. Explain, using this data and the structures described in (a), why SiO_2 has a vastly higher melting point than the sublimation point of CO_2, even though the individual Si-O bond enthalpy is lower than the individual C=O bond enthalpy. **(4)**

(c) In the giant covalent lattice of SiO_2, there are 4 mol of Si-O bonds per mole of SiO_2 formula units. Using the Si-O bond enthalpy given, estimate the enthalpy needed to break all the covalent bonds in 1 mole of solid $SiO_2(s)$ into gaseous atoms, and state one limitation of using an average bond enthalpy for this estimate. **(4)**

(Total for Question 14 is 10 marks)

15 This question uses ionic radii to compare the lattice enthalpies of the Group 2 oxides MgO, CaO and SrO.

(a) State how the magnitude of the lattice enthalpy of an ionic compound depends on (i) the charges of the ions and (ii) the sum of their ionic radii. **(2)**

(b) Ionic radii: Mg^{2+} = 72 pm, Ca^{2+} = 100 pm, Sr^{2+} = 118 pm, O^{2-} = 140 pm. Calculate the sum of ionic radii (r+ + r-) for MgO, CaO and SrO. **(3)**

(c) The product of the ionic charges (2 x 2 = 4) is the same for MgO, CaO and SrO. Using your answers to (b), predict the order of lattice enthalpy magnitude (most exothermic to least exothermic) for these three oxides, explaining your reasoning. **(3)**

(d) Calculate the ratio (r+ + r- for MgO) / (r+ + r- for CaO), giving your answer to 2 significant figures, and state what this ratio suggests about the relative magnitude of the lattice enthalpy of CaO compared with MgO. **(2)**

(Total for Question 15 is 10 marks)

16 This question is about determining the density of solid sodium chloride from its unit cell dimensions, and the associated measurement uncertainty.

(a) Sodium chloride crystallises with a face-centred cubic unit cell of edge length a = 564.0 pm. Show that the volume of the unit cell is 1.79×10^{-22} cm^3. **(3)**

(b) Each unit cell contains 4 formula units of NaCl. Using density = (Z x M) / (NA x V), where Z is the number of formula units per unit cell, M is the molar mass of NaCl (58.5 g/mol) and NA = 6.02×10^{23} mol^{-1}, calculate the density of solid NaCl. Give your answer to 3 significant figures. **(3)**

(c) The edge length was measured as a = 564.0 pm ± 2 pm. Calculate the percentage uncertainty in a, and hence the percentage uncertainty in the density calculated in (b), given that density is proportional to $1/a^3$ (so the percentage uncertainty in density is 3 times the percentage uncertainty in a). **(3)**

(d) Express the density found in (b) together with its absolute uncertainty, quoted to an appropriate number of significant figures. **(3)**

(Total for Question 16 is 12 marks)

17 This question is about hydrogen bonding and its effects on the physical properties of substances.

(a) Boiling points: NH_3 = -33°C, PH_3 = -88°C. Explain why NH_3 has a significantly higher boiling point than PH_3, even though PH_3 has the larger Mr. **(3)**

(b) Ethanol, C_2H_5OH, and dimethyl ether, CH_3OCH_3, are structural isomers with the same molecular formula, C_2H_6O (Mr = 46), yet ethanol boils at 78.5°C while dimethyl ether boils at -24°C. Explain this large difference in boiling point. **(4)**

(c) Ice is less dense than liquid water. State the structural feature of ice that is responsible for this. **(2)**

(Total for Question 17 is 9 marks)

18 Both ice (solid H_2O) and dry ice (solid CO_2) are solids formed from small covalent molecules, yet ice melts at 0°C while dry ice sublimes at -78°C at atmospheric pressure. Compare the structure and bonding in ice and dry ice, and use this to explain the large difference between their melting point and sublimation point.

(Total for Question 18 is 6 marks)

19 This question is about uncertainty in temperature measurements made using a melting point apparatus.

(a) A student determines the melting point of a solid organic compound using a melting point apparatus with a thermometer marked in divisions of 1.0°C. State the uncertainty in a single temperature reading from this thermometer, given that the uncertainty in an analogue scale reading is taken as half the smallest scale division. **(1)**

(b) The student records the sample beginning to melt at 182.0°C. Calculate the percentage uncertainty in this single reading. **(2)**

(c) The sample finishes melting at 184.0°C. Calculate the melting range (the difference between the two readings) and its absolute uncertainty. **(3)**

(d) Calculate the percentage uncertainty in the melting range, and comment on what this shows about using this apparatus to determine a small melting range precisely. **(2)**

(Total for Question 19 is 8 marks)

Mark scheme · 2.2 Bonding and Structure

Question 1

(a) the power (ability) of an atom in a covalent bond to attract the bonding (shared) pair of electrons towards itself, oe **B1**

(a) Answer: The power of an atom in a covalent bond to attract the bonding pair of electrons towards itself.

(b) a covalent bond in which both of the shared electrons in the bonded pair are provided by the same atom, oe **B1**

(b) Answer: A covalent bond in which both shared electrons come from the same atom.

(c) a small, permanent charge separation across a bond caused by a difference in electronegativity between the two bonded atoms, giving one atom a partial positive charge (delta+) and the other a partial negative charge (delta-), oe **B1**

(c) Answer: A permanent small charge separation across a bond, delta+ on one atom and delta- on the other, caused by unequal electronegativities.

(d) weak intermolecular forces caused by instantaneous (temporary) dipoles inducing dipoles in neighbouring molecules or atoms, oe **B1**

(d) Answer: Weak intermolecular forces caused by instantaneous dipoles inducing dipoles in neighbouring particles.

Question 2

(a) the size of the ionic charge (a greater charge gives stronger electrostatic attraction) oe **B1**

(a) the ionic radius (a smaller radius means ions are closer together, giving stronger attraction) oe **B1**

(a) Answer: Ionic charge and ionic radius.

(b) sum of ionic radii: NaCl, r = 102 + 181 = 283 pm; MgO, r = 72 + 140 = 212 pm **M1**

(b) charge product: NaCl, Q1 x Q2 = 1 x 1 = 1; MgO, Q1 x Q2 = 2 x 2 = 4 **M1**

(b) value for NaCl = 1/283 = 3.53 x 10^{-3} (pm^{-1}); value for MgO = 4/212 = 1.887 x 10^{-2} (pm^{-1}) (allow awrt) **M1**

(b) ratio = 1.887x10^{-2} / 3.53x10^{-3} = 5.3 (accept 5.3-5.4), confirming the value for MgO is more than five times greater than for NaCl, cso **A1**

(b) Answer: Ratio approx 5.3: the MgO value is about 5.3 times greater than the NaCl value, confirming the statement.

(c) the larger calculated value for MgO (ft from (b)) shows a stronger electrostatic force of attraction between its ions, meaning MgO has a greater (more exothermic) lattice enthalpy oe **M1**

(c) more energy is required to overcome these stronger ionic bonds, so a higher temperature is needed to melt MgO than NaCl oe **A1**

(c) Answer: The larger value for MgO shows stronger ionic attraction (greater lattice enthalpy), so more energy (a higher temperature) is needed to melt MgO, giving it a higher melting point than NaCl.

(d) the ionic radius of the Group 2 cation increases down the group, while its charge stays the same (2+) oe **B1**

(d) the ions become further apart, so electrostatic attraction weakens: lattice enthalpy becomes less negative (decreases in magnitude) and melting point decreases down the group oe **B1**

(d) Answer: Lattice enthalpy (in magnitude) and melting point both decrease going down Group 2, because increasing cation radius (same charge) weakens the electrostatic attraction between ions.

Question 3

(a) a shared pair of electrons, attracted to the nuclei of both atoms oe **B1**

(a) Answer: A shared pair of electrons, attracted to the nuclei of both bonded atoms.

(b) carbon shares one electron with each of four hydrogen atoms, forming four single covalent bonds oe **B1**

(b) each shared (bonding) pair consists of one electron from carbon and one electron from the hydrogen atom, giving carbon a share in 8 electrons (a full outer shell/octet) oe **B1**

(b) Answer: Carbon forms four single covalent bonds, each a shared pair (one electron from C, one from H), giving carbon a share of 8 outer electrons.

(c) a shared pair of electrons in which both electrons are provided (donated) by only one of the two bonded atoms oe **B1**

(c) Answer: A covalent bond in which both shared electrons come from only one of the two atoms.

(d) the nitrogen atom in NH_3 has a lone pair of electrons oe **B1**

(d) the H^+ ion has an empty (1s) orbital and no electrons of its own to contribute oe **B1**

(d) nitrogen donates its lone pair into the empty orbital on H^+, forming a new dative covalent bond oe **B1**

(d) Answer: The nitrogen lone pair in NH_3 is donated into the empty orbital of H^+, forming a new N-H dative covalent bond.

(e) tetrahedral **B1**

(e) 109.5 degrees **B1**

(e) Answer: Tetrahedral, with a bond angle of 109.5 degrees.

Question 4

(a) the power (ability) of an atom to attract the bonding (shared) pair of electrons in a covalent bond oe **B1**

(a) Answer: The ability of an atom to attract the bonding pair of electrons in a covalent bond.

(b) electronegativity increases across Period 3 from Na to Cl **M1**

(b) nuclear charge (number of protons) increases across the period while the number of occupied shells (and shielding by inner shells) stays approximately constant oe **A1**

(b) atomic radius decreases, so there is a greater attraction for the bonding pair of electrons oe **A1**

(b) Answer: Electronegativity increases from Na to Cl because nuclear charge increases with roughly constant shielding, so atomic radius decreases and the attraction for bonding electrons increases.

(c) calculates the electronegativity difference for each: HF = 4.0-2.1 = 1.9; HCl = 3.0-2.1 = 0.9; HI = 2.5-2.1 = 0.4 **M1**

(c) correct order of increasing polarity: HI < HCl < HF **A1**

(c) correct justification: a greater electronegativity difference means a more polar bond oe **A1**

(c) Answer: Order of increasing bond polarity: HI < HCl < HF.

(d) the statement is not fully correct; the shape of the molecule must also be considered oe **M1**

(d) CO_2 is a linear, symmetrical molecule oe **M1**

(d) the two C=O bond dipoles are equal in size and point in exactly opposite directions, so they cancel out, giving no overall (net) molecular dipole: CO_2 is non-polar oe **A1**

(d) Answer: The statement is incorrect. CO_2 has polar bonds, but because the molecule is linear and symmetrical the bond dipoles cancel, so CO_2 itself is a non-polar molecule.

Question 5

(a) electron pairs (bonding and lone pairs) around a central atom repel one another and arrange themselves as far apart as possible, to minimise repulsion oe **B1**

(a) the order of repulsion strength is: lone pair-lone pair > lone pair-bonding pair > bonding pair-bonding pair oe **B1**

(a) Answer: Electron pairs around a central atom repel each other and arrange as far apart as possible; lone pair-lone pair repulsion is strongest, then lone pair-bonding pair, then bonding pair-bonding pair.

(b) boron has three bonding pairs and no lone pairs of electrons around it **B1**

(b) shape is trigonal planar **B1**

(b) bond angle is 120 degrees, since the three bonding pairs repel equally and spread out as far apart as possible oe **B1**

(b) Answer: Trigonal planar, bond angle 120 degrees.

(c) shape is trigonal pyramidal **B1**

(c) bond angle is 107 degrees **B1**

(c) the lone pair on nitrogen repels the three bonding pairs more strongly than a bonding pair repels another bonding pair, pushing the N-H bonds closer together and reducing the angle from 109.5 to 107 degrees oe **B1**

(c) Answer: Trigonal pyramidal, bond angle 107 degrees; the extra repulsion from the lone pair compresses the angle below the tetrahedral 109.5 degrees.

(d) shape is non-linear/bent (V-shaped) **B1**

(d) bond angle is 104.5 degrees **B1**

(d) two lone pairs on oxygen cause even greater repulsion than one lone pair, pushing the two O-H bonding pairs even closer together, reducing the angle further than in NH_3 (107 to 104.5 degrees) oe **B1**

(d) Answer: Non-linear (bent), bond angle 104.5 degrees; two lone pairs cause greater repulsion than one, compressing the angle further than in NH_3.

Question 6

(a) a giant metallic lattice of positive metal ions arranged in a regular, close-packed structure oe **B1**

(a) surrounded by a "sea" of delocalised electrons that are free to move throughout the whole structure oe **B1**

(a) Answer: A giant lattice of positive metal ions in a regular arrangement, surrounded by a sea of delocalised electrons.

(b) the delocalised electrons are free to move throughout the metallic structure and can carry charge oe **B1**

(b) metallic bonding (and the delocalised electrons) is present in both the solid and molten states, since only the regular arrangement of ions is lost on melting, not the bonding itself oe **B1**

(b) so mobile charge carriers (delocalised electrons) are available to conduct electricity in both states oe **B1**

(b) Answer: Delocalised electrons are free to move and carry charge in both solid and molten metal, since metallic bonding persists on melting.

(c) the layers of positive metal ions can slide over one another when a force is applied oe **B1**

(c) because metallic bonding is non-directional, the delocalised electrons continue to attract and hold the ions together as the layers move, so the metal bends/changes shape without breaking oe **B1**

(c) Answer: Layers of metal ions can slide over each other under force, while the non-directional sea of delocalised electrons keeps holding the structure together, so metals bend rather than shatter.

(d) an alloy contains atoms of one or more other elements, which are usually a different size from the atoms of the pure metal oe **M1**

(d) this distorts the regular arrangement of layers of ions in the lattice oe **M1**

(d) the layers can no longer slide over each other as easily, making the alloy harder/stronger than the pure metal oe **A1**

(d) Answer: Alloys contain atoms of different sizes, which distort the regular layered structure and prevent layers sliding easily, making alloys harder than pure metals.

Question 7

(a) a weak intermolecular force caused by temporary (instantaneous) dipoles, arising from the random movement of electrons in an electron cloud oe **B1**

(a) the temporary dipole induces a dipole in a neighbouring molecule/atom, resulting in a weak force of attraction between them oe **B1**

(a) Answer: A weak attractive force caused by temporary dipoles (from random electron movement) inducing dipoles in neighbouring molecules.

(b) the number of electrons (and atomic size/Mr) increases down the group oe **B1**

(b) this increases the size of the electron cloud, so larger instantaneous and induced dipoles can form, increasing the strength of the London forces oe **B1**

(b) more energy is needed to overcome these stronger forces, so the boiling point increases down the group oe **B1**

(b) Answer: Boiling point increases down the group because more electrons give larger, stronger London forces, requiring more energy to overcome.

(c) HCl is a polar molecule (Cl is more electronegative than H), so it has a permanent dipole; Ar is non-polar oe **B1**

(c) HCl molecules experience permanent dipole-dipole forces in addition to London forces, whereas only London forces act between Ar atoms oe **B1**

(c) for molecules/atoms of similar Mr, the additional dipole-dipole forces are stronger overall than London forces alone, so more energy is needed to separate HCl molecules, giving a higher boiling point oe **B1**

(c) Answer: HCl has permanent dipole-dipole forces (from its polar bond) in addition to London forces, which are stronger overall than the London forces alone acting between non-polar Ar atoms, giving HCl a higher boiling point.

(d) calculates the difference in boiling point: -85 - (-186) = 101 C **M1**

(d) concludes that despite Ar having the greater Mr (and so stronger London forces alone), HCl boils 101 C higher, showing that permanent dipole-dipole forces make a significant additional contribution to intermolecular attraction in HCl oe **A1**

(d) Answer: HCl boils 101 C higher than Ar, showing the extra dipole-dipole forces in HCl contribute significantly more attraction than London forces alone.

Question 8

(a) a hydrogen atom must be covalently bonded directly to a small, highly electronegative atom: nitrogen, oxygen or fluorine (N, O or F) oe **B1**

(a) the electronegative atom (N, O or F) must have at least one lone pair of electrons available oe **B1**

(a) the hydrogen bond forms between the delta+ hydrogen on one molecule and a lone pair on the N, O or F atom of a neighbouring molecule oe **B1**

(a) Answer: H must be bonded directly to N, O or F; that atom must have a lone pair; the hydrogen bond forms between the delta+ H and a lone pair on N/O/F of a neighbouring molecule.

(b) the delta+ hydrogen atom of one water molecule is attracted to a lone pair on the delta- oxygen atom of a neighbouring water molecule oe **B1**

(b) the hydrogen bond acts as a roughly linear, weak attraction (often shown as a dotted/dashed line) between the H and the lone pair, O-H...O oe **B1**

(b) Answer: The delta+ H of one water molecule is attracted to a lone pair on the delta- O of a neighbouring molecule, forming a roughly linear O-H...O hydrogen bond.

(c) in ice, hydrogen bonds hold water molecules in a rigid, open lattice, with each molecule hydrogen bonded to four neighbours oe **M1**

(c) this rigid, open structure creates more empty space between the molecules than in the liquid oe **M1**

(c) in liquid water, hydrogen bonds are constantly breaking and reforming, allowing molecules to pack more closely together, so liquid water is denser than ice oe **A1**

(c) Answer: Ice has a rigid, open hydrogen-bonded lattice with more space between molecules than liquid water, where molecules can pack more closely, making ice less dense.

(d) uses volume = mass / density for a fixed mass (e.g. 1 g): volume of water = 1/1.000 = 1.000 cm^3 **M1**

(d) volume of ice = 1/0.917 = 1.091 cm^3 (awrt 1.09) **M1**

(d) increase in volume = 1.091 - 1.000 = 0.091 cm^3 **M1**

(d) percentage increase = (0.091/1.000) x 100 = 9.1% (accept awrt 9.0-9.1%) cao **A1**

(d) Answer: Approximately 9.1% increase in volume (accept 9.0-9.1%).

Question 9

(a) a giant covalent (macromolecular) structure/lattice oe **B1**

(a) each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral arrangement, extending throughout the whole structure oe **B1**

(a) Answer: A giant covalent lattice in which each carbon atom is bonded to four others tetrahedrally, throughout the whole structure.

(b) diamond has a giant covalent structure containing many strong covalent bonds extending throughout the lattice oe **B1**

(b) a very large amount of energy is needed to break these many strong covalent bonds, giving a very high melting point oe **B1**

(b) Answer: Diamond's giant covalent lattice has many strong covalent bonds; a very large amount of energy is needed to break them, giving a very high melting point.

(c) graphite consists of layers of carbon atoms, each covalently bonded to only three other carbon atoms, arranged in hexagonal rings oe **B1**

(c) there are strong covalent bonds within each layer, but only weak London (van der Waals) forces between layers oe **B1**

(c) each carbon atom in graphite has one delocalised electron (not used in bonding), which is free to move within a layer, allowing graphite to conduct electricity oe **B1**

(c) in diamond, all four outer electrons of every carbon atom are used in (localised) covalent bonds, so there are no delocalised electrons/free charge carriers, and diamond does not conduct oe **B1**

(c) Answer: Graphite has layers of carbon atoms each bonded to 3 others (hexagonal rings), with one delocalised electron per atom free to move within each layer, allowing conduction; diamond has all 4 outer electrons localised in covalent bonds, so it cannot conduct.

(d) the layers in graphite are held together only by weak London (van der Waals) forces oe **B1**

(d) these weak forces allow the layers to slide over one another easily (and flake off), reducing friction between surfaces oe **B1**

(d) Answer: Weak forces between graphite's layers let them slide over each other easily, reducing friction, so graphite acts as a lubricant.

(e) identifies a valid property, e.g. very high electrical conductivity (or extremely high strength/thinness/flexibility) oe **B1**

(e) links the property to its structure/usefulness, e.g. every delocalised electron is at the surface (none are "wasted" between internal layers as in graphite), giving very low resistance/fast electron movement, useful in very thin, flexible electronic components oe **B1**

(e) Answer: Graphene has very high electrical conductivity, since all its delocalised electrons are at the exposed single layer, useful for extremely thin, flexible conducting components.

Question 10

(a) 5 bonding pairs and 0 lone pairs around P	**B1**
(a) trigonal bipyramidal	**B1**
(a) bond angles of 90 degrees (axial to equatorial) and 120 degrees (equatorial to equatorial)	**B1**

(a) Answer: 5 bonding pairs, 0 lone pairs; trigonal bipyramidal; 90 degrees and 120 degrees.

(b) octahedral	**B1**
(b) 90 degrees	**B1**

(b) Answer: Octahedral; 90 degrees.

(c) square planar	**B1**
(c) 90 degrees	**B1**

(c) Answer: Square planar; 90 degrees.

(d) bent (V-shaped, non-linear)	**B1**
(d) bond angle of about 115 degrees, which is less than 120 degrees because a lone pair repels neighbouring electron pairs more strongly than a bonding pair does, compressing the O-N-O angle, oe	**B1**

(d) Answer: Bent, approximately 115 degrees.

Question 11

Level 3 A full and coherent explanation addressing both patterns: correctly identifies that the Group 14 hydrides are non-polar, symmetrical (tetrahedral) molecules held together only by London (dispersion) forces, and correctly explains that these forces strengthen down the group as the number of electrons (and Mr) increases; and correctly identifies that H_2O (unlike H_2S, H_2Se, H_2Te) can form hydrogen bonds, because its H atoms are bonded directly to the small, highly electronegative O atom which has lone pairs available, and explains that hydrogen bonds are much stronger than the London and permanent dipole-dipole forces present in the other Group 16 hydrides, requiring significantly more energy to overcome and giving H_2O an anomalously high boiling point.	**5-6**
Level 2 The explanation addresses most of the required chemistry for both patterns but is incomplete in at least one respect (for example, does not fully explain why H_2O can hydrogen bond and H_2S cannot, or omits reference to the number of electrons/Mr in the Group 14 trend), or contains minor inaccuracies; both the Group 14 trend and the Group 16 anomaly are attempted.	**3-4**
Level 1 A basic or partial explanation with limited use of correct terminology; may address only one of the two patterns (for example, only explains the Group 14 trend, or only states "hydrogen bonding" without further detail linking it to boiling point).	**1-2**
Level 0 No relevant chemistry, or the answer does not relate to the question.	**0**

Indicative content:

- The Group 14 hydrides (CH_4, SiH_4, GeH_4, SnH_4) are non-polar, symmetrical, tetrahedral molecules with no permanent dipole, so the only intermolecular force present between them is London (dispersion) forces.
- The number of electrons (and Mr) increases down Group 14 from CH_4 to SnH_4.
- A greater number of electrons gives a larger electron cloud, allowing larger instantaneous and induced dipoles to form, so the strength of the London forces increases down the group.
- Stronger London forces require more energy to overcome, so the boiling point of the Group 14 hydrides increases steadily down the group.
- The Group 16 hydrides H_2S, H_2Se and H_2Te also show an increasing boiling point down the group, for the same reason (increasing number of electrons strengthens London forces); these three are also weakly polar, adding weak permanent dipole-dipole forces.
- H_2O is different because its hydrogen atoms are bonded directly to oxygen, which is small and highly electronegative and has two lone pairs, satisfying the conditions needed for hydrogen bonding.
- Hydrogen bonds are much stronger than permanent dipole-dipole forces and London forces.
- Significantly more energy is needed to overcome the additional hydrogen bonding between H_2O molecules, giving H_2O a much higher boiling point than the trend (based on Mr/London forces alone) would predict.
- H_2S, H_2Se and H_2Te cannot form hydrogen bonds because S, Se and Te are larger and less electronegative than O, so their H-S, H-Se and H-Te bonds are not polar enough to meet the conditions for hydrogen bonding.

Question 12

(a) Al^{3+} is a small ion with a high (3+) charge, so it has a high charge density (high polarising power) oe **M1**

(a) Cl^- is a relatively large ion, so its electron cloud is polarisable (easily distorted) oe **M1**

(a) the small, highly charged Al^{3+} ion distorts (polarises) the electron cloud of the Cl^- ion, pulling electron density towards itself oe **M1**

(a) this distortion means electron density is shared between the ions, giving the Al-Cl bond significant covalent character oe **A1**

(a) Answer: Al^{3+} has high charge density and strongly polarises the large, polarisable Cl^- ion, distorting its electron cloud and sharing electron density, giving the bond significant covalent character.

(b) AlF_3 has the greater ionic character **B1**

(b) correct reason: F^- is smaller and less polarisable than Cl^- (its electron cloud is held more tightly), so Al^{3+} causes less distortion, and the bonding remains closer to purely ionic oe **B1**

(b) Answer: AlF_3 has the greater ionic character, because F^- is smaller and less polarisable than Cl^-, so it is less distorted by Al^{3+}.

(c) a dimer, Al_2Cl_6 (two $AlCl_3$ units joined together) oe **B1**

(c) Answer: It exists as the dimer Al_2Cl_6.

(d) a chlorine atom on one $AlCl_3$ unit uses one of its lone pairs of electrons to form a new bond, donated to the aluminium atom of the neighbouring $AlCl_3$ unit, which has an empty orbital (only 6 electrons around Al in $AlCl_3$, an incomplete octet) oe **M1**

(d) two such dative bonds form (from two bridging chlorine atoms), joining the two $AlCl_3$ units into the dimer Al_2Cl_6 oe **A1**

(d) Answer: A lone pair from a Cl atom on one $AlCl_3$ unit is donated into the empty orbital on the Al atom of the other unit; two such dative bonds (bridging chlorines) form the Al_2Cl_6 dimer.

Question 13

(a) in the adduct, boron now has four bonding pairs of electrons around it: the three original B-F bonds plus the new dative B<-N bond oe **M1**

(a) boron has four bonding pairs and no lone pairs oe **M1**

(a) the shape around boron becomes tetrahedral (bond angles approximately 109.5 degrees), rather than the trigonal planar (120 degree) shape in BF_3 itself oe **A1**

(a) Answer: Tetrahedral (bond angles approximately 109.5 degrees), because boron now has four bonding pairs and no lone pairs.

(b) boron has only three electrons in its outer shell (Group 13), so after forming three covalent bonds to fluorine it has only six electrons around it, an incomplete octet oe **M1**

(b) this leaves an empty orbital on boron, which can accept the lone pair of electrons donated by the nitrogen atom of NH_3, allowing the dative bond to form oe **A1**

(b) Answer: Boron has only 6 electrons around it in BF_3 (incomplete octet), leaving an empty orbital that can accept the nitrogen lone pair, forming the dative bond.

(c) because nitrogen's lone pair is now used in the dative bond to boron, it is no longer available to accept hydrogen bonds from the N-H hydrogens of neighbouring adduct molecules oe **M1**

(c) so the adduct cannot hydrogen bond extensively between its own molecules, and its volatility (boiling point) would be expected to be similar to other simple molecular compounds of comparable Mr (governed mainly by London and permanent dipole-dipole forces), rather than being anomalously raised as it would be if the lone pair were still free oe **A1**

(c) Answer: With the N lone pair used in the dative bond, the adduct molecules cannot hydrogen bond extensively to each other, so its volatility should resemble that of a similar-Mr compound relying on London/dipole-dipole forces, not an anomalously high-boiling hydrogen-bonded compound.

(d) nitrogen still has a lone pair available in $N(CH_3)_3$ (as in NH_3), so a dative bond to boron can still form oe **M1**

(d) the bulkier CH_3 groups increase the Mr, surface area and number of electrons of the molecule, increasing the strength of the London forces between adduct molecules, which would tend to raise the melting point of $F_3B\text{-}N(CH_3)_3$ compared with $F_3B\text{-}NH_3$ oe **A1**

(d) Answer: The dative bond still forms (N retains its lone pair), but the bulkier methyl groups increase Mr and London forces between molecules, tending to raise the melting point compared with $F_3B\text{-}NH_3$.

Question 14

(a) CO_2 has a simple molecular (covalent) structure: discrete, linear CO_2 molecules held together in the solid by weak van der Waals forces, oe **B1**

(a) SiO_2 has a giant covalent (macromolecular) structure in which every Si atom is covalently bonded to four O atoms (and every O atom to two Si atoms), forming a continuous, extended three-dimensional network, oe **B1**

(a) Answer: CO_2 is simple molecular; SiO_2 is giant covalent.

(b) subliming CO_2 only requires overcoming the weak van der Waals forces between molecules; the strong covalent C=O bonds within each molecule are not broken, oe **B1**

(b) melting SiO_2 requires breaking a very large number of strong covalent Si-O bonds throughout the giant covalent lattice, oe **B1**

(b) it is the type of force overcome (many strong covalent bonds throughout a lattice, versus weak van der Waals forces between separate molecules) that determines the melting/sublimation point, not the strength of one individual bond, oe **B1**

(b) so far more energy overall is required to melt SiO_2 than to sublime CO_2, giving SiO_2 its vastly higher melting point, oe **B1**

(b) Answer: Melting SiO_2 breaks many strong covalent bonds throughout a lattice, while subliming CO_2 only overcomes weak van der Waals forces between molecules.

(c) recognise that 4 mol of Si-O bonds must be broken per mole of SiO_2 **M1**

(c) correct substitution: 4 x 466 **M1**

(c) 1864 kJ/mol (accept 1860-1870), cao **A1**

(c) this is only an estimate because an average bond enthalpy assumes every Si-O bond in the lattice has identical strength/environment, which is not exactly true in the real crystal, oe **B1**

(c) Answer: Approximately 1864 kJ/mol; the value uses an average bond enthalpy, which assumes every Si-O bond is identical.

Question 15

(a) the magnitude increases as the product of the ionic charges increases, oe **B1**

(a) the magnitude increases as the sum of the ionic radii decreases (the ions get closer together), oe **B1**

(a) Answer: Magnitude increases with greater ionic charge and with smaller sum of ionic radii.

(b) MgO: 212 pm **B1**

(b) CaO: 240 pm **B1**

(b) SrO: 258 pm **B1**

(b) Answer: MgO = 212 pm, CaO = 240 pm, SrO = 258 pm.

(c) recognise that, with the charge product constant, lattice enthalpy magnitude is approximately inversely related to r+ + r-, i.e. it decreases as r+ + r- increases **M1**

(c) correct order: MgO (most exothermic) > CaO > SrO (least exothermic) **A1**

(c) MgO has the smallest sum of ionic radii, so its ions can approach most closely, giving the strongest electrostatic attraction between oppositely charged ions and hence the most exothermic lattice enthalpy, oe **B1**

(c) Answer: MgO more exothermic than CaO, which is more exothermic than SrO.

(d) correct substitution: 212/240 **M1**

(d) 0.88 (cao, 2 sf), so the lattice enthalpy of CaO is predicted to be about 88% of the magnitude of that of MgO **A1**

(d) Answer: 0.88; CaO's lattice enthalpy magnitude is predicted to be about 88% of MgO's.

Question 16

(a) convert 564.0 pm to cm: 5.640×10^{-8} cm **M1**

(a) cube the edge length: $(5.640 \times 10^{-8})^3$ **M1**

(a) cso: 1.79×10^{-22} cm^3 (3 sf), as required **A1**

(a) Answer: 1.79×10^{-22} cm^3 (shown)

(b) correct identification of Z = 4 and M = 58.5 g/mol in the formula **M1**

(b) correct substitution: $(4 \times 58.5) / (6.02 \times 10^{23} \times 1.79 \times 10^{-22})$ **M1**

(b) 2.17 g/cm^3 (accept 2.16-2.17), cao **A1**

(b) Answer: 2.17 g/cm^3

(c) percentage uncertainty in a = (2/564.0) x 100 **M1**

(c) recognise percentage uncertainty in density = 3 x percentage uncertainty in a **M1**

(c) 1.1% (accept 1.0-1.1%), cao **A1**

(c) Answer: Approximately 1.1%

(d) absolute uncertainty = (1.1/100) x 2.17 **M1**

(d) 0.02 g/cm^3 (accept 0.02-0.03) **A1**

(d) density = (2.17 ± 0.02) g/cm^3, with the uncertainty quoted to the same number of decimal places as the density **B1**

(d) Answer: (2.17 ± 0.02) g/cm^3

Question 17

(a) NH_3 molecules can form hydrogen bonds with each other, because N is highly electronegative and has a lone pair, and a hydrogen atom is bonded directly to it, oe **B1**

(a) P is not electronegative enough for PH_3 to form hydrogen bonds, so PH_3 molecules are held together only by (weaker) van der Waals forces, oe **B1**

(a) hydrogen bonds are much stronger than van der Waals forces, so despite its smaller Mr (and so weaker van der Waals forces), more energy is needed to separate NH_3 molecules, giving NH_3 the higher boiling point, oe **B1**

(a) Answer: NH_3 can hydrogen bond and PH_3 cannot; hydrogen bonds are much stronger than the van der Waals forces in PH_3.

(b) ethanol has an O-H bond, with a hydrogen atom bonded directly to the highly electronegative oxygen (which has lone pairs), so ethanol molecules can form hydrogen bonds with each other, oe **B1**

(b) dimethyl ether has no O-H bond (all its hydrogen atoms are bonded to carbon), so it cannot form hydrogen bonds between its molecules, oe **B1**

(b) since both compounds have the same Mr, their van der Waals forces are of similar strength, oe **B1**

(b) the additional hydrogen bonding present in ethanol is much stronger, so significantly more energy is needed to separate ethanol molecules, giving it the much higher boiling point, oe **B1**

(b) Answer: Ethanol can hydrogen bond (O-H bond) while dimethyl ether cannot, despite both having the same Mr and similar van der Waals forces.

(c) in ice, each water molecule forms hydrogen bonds to (up to) four neighbouring water molecules, oe **B1**

(c) this produces an open, rigid, hexagonal (cage-like) lattice with more empty space between molecules than in liquid water, so ice is less dense than liquid water, oe **B1**

(c) Answer: The open, hydrogen-bonded hexagonal lattice of ice holds molecules further apart than in liquid water.

Question 18

Level 3 A detailed and accurate comparison of the structure and bonding in ice and dry ice, correctly identifying the intermolecular forces present in each, with a clear and correct explanation of the large difference in melting/sublimation point. The answer is coherent and logically structured, using appropriate scientific terminology throughout. **5-6**

Level 2 A mostly correct comparison of structure and/or bonding in ice and dry ice, with a partially correct explanation of the temperature difference. There may be a minor inaccuracy or an underdeveloped link between structure/bonding and the observed properties. The answer shows some structure. **3-4**

Level 1 Some relevant, briefly stated points about the structure or bonding of ice and/or dry ice are given, but these are not linked into a coherent comparison or explanation of the temperature difference. **1-2**

Indicative content:

- Both ice and dry ice are simple molecular (covalent) lattices; the covalent bonds within each molecule (O-H in water, C=O in CO_2) are not broken when the solid melts or sublimes.
- CO_2 is a linear, non-polar molecule (the two C=O bond dipoles cancel by symmetry), so the only forces between CO_2 molecules in dry ice are relatively weak van der Waals (London/induced dipole-dipole) forces.
- In ice, each water molecule can form hydrogen bonds to up to four neighbouring water molecules (via its two O-H bonds and two oxygen lone pairs), in addition to van der Waals forces.
- Hydrogen bonds are considerably stronger than the van der Waals forces present in dry ice, so significantly more energy is needed to separate water molecules from the ice lattice than to separate CO_2 molecules from the dry ice lattice.
- This explains why ice has a much higher melting point (0°C) than the sublimation point of dry ice (-78°C): far less energy is needed to overcome the weak van der Waals forces in dry ice than to overcome the extensive hydrogen-bonded network (plus van der Waals forces) in ice.

Question 19

(a) ± 0.5°C **B1**

(a) Answer: ± 0.5°C

(b) correct substitution: (0.5/182.0) x 100 **M1**

(b) 0.27% (accept 0.27-0.28%), cao **A1**

(b) Answer: 0.27%

(c) melting range = 184.0 - 182.0 = 2.0°C **M1**

(c) absolute uncertainties are added when two readings are subtracted: total uncertainty = 0.5 + 0.5 = ± 1.0°C **M1**

(c) melting range = (2.0 ± 1.0) °C **A1**

(c) Answer: (2.0 ± 1.0) °C

(d) correct substitution: (1.0/2.0) x 100 = 50% **M1**

(d) even though each individual reading has a small percentage uncertainty (about 0.3%), the percentage uncertainty in a small difference between two similar readings is very large (50%), so this apparatus is unsuitable/unreliable for precisely determining a narrow melting range, oe **B1**

(d) Answer: 50%; the apparatus gives a very large percentage uncertainty for a small temperature difference.