AC2D Bonding and Structure: Depth and Exam Drill

AQA 7405 · Calculator allowed · about 155 minutes

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)

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2 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 PF5, and hence give its shape and bond angles. **(3)**

(b) SF6 has 6 bonding pairs and no lone pairs around the central sulfur atom. State its shape and bond angle. **(2)**

(c) The ion ICl4- 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, NO2-, 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 2 is 9 marks)

3 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 3 is 10 marks)

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4 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 4 is 12 marks)

5 This question is about electronegativity, bond polarity and molecular dipoles.

(a) Pauling electronegativity values: H = 2.1, C = 2.5, N = 3.0, O = 3.5, F = 4.0, Na = 0.9, Mg = 1.2, Cl = 3.0. Calculate the electronegativity difference for each of the following bonds: Na-Cl, C-Cl, C-H, O-H. **(4)**

(b) Using the general guide that a difference below 0.4 indicates an essentially non-polar covalent bond, a difference of roughly 0.4-1.8 indicates a polar covalent bond, and a difference above 1.8 indicates a bond that is best described as ionic, classify each of the four bonds in (a). **(4)**

(c) Methane, CH4, and tetrachloromethane, CCl4, both have no overall (net) molecular dipole moment, even though CCl4 contains polar C-Cl bonds. Explain why. **(1)**

(Total for Question 5 is 9 marks)

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6 This question uses boiling point data for the halogens to explore trends in intermolecular forces and the limitations of extrapolation.

(a) Boiling points: F2 (Mr = 38) = -188 degrees C, Cl2 (Mr = 71) = -34 degrees C, Br2 (Mr = 160) = 59 degrees C, I2 (Mr = 254) = 184 degrees C. State and explain the trend in boiling point from F2 to I2. **(3)**

(b) At2 has Mr = 420 (approximately). Using the gradient between the Br2 and I2 data points, estimate the boiling point of At2. **(3)**

(c) Suggest why extrapolating this trend to estimate the boiling point of At2 may be unreliable. **(2)**

(Total for Question 6 is 8 marks)

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

(a) Boiling points: NH3 = -33 degrees C, PH3 = -88 degrees C. Explain why NH3 has a significantly higher boiling point than PH3, even though PH3 has the larger Mr. **(3)**

(b) Ethanol, C2H5OH, and dimethyl ether, CH3OCH3, are structural isomers with the same molecular formula, C2H6O (Mr = 46), yet ethanol boils at 78.5 degrees C while dimethyl ether boils at -24 degrees 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 7 is 9 marks)

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8 Both ice (solid H2O) and dry ice (solid CO2) are solids formed from small covalent molecules, yet ice melts at 0 degrees C while dry ice sublimes at -78 degrees 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

9 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 degrees 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 degrees C. Calculate the percentage uncertainty in this single reading. **(2)**

(c) The sample finishes melting at 184.0 degrees 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 9 is 8 marks)

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10 This question is about ionic polarisation and covalent character in Group 2 chlorides.

(a) State Fajans' rules for predicting when an ionic bond will have significant covalent character, in terms of the cation and the anion involved. **(3)**

(b) Ionic radii: Be^{2+} = 45 pm, Mg^{2+} = 72 pm, Cl^- = 181 pm. Using the data given and Fajans' rules, predict which chloride, $BeCl_2$ or $MgCl_2$, has the greater degree of covalent character in its bonding. Explain your answer. **(3)**

(c) Predict, giving a reason, which of $BeCl_2$ and $MgCl_2$ has the lower melting point. **(3)**

(Total for Question 10 is 9 marks)

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11 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 degrees C. Silicon dioxide, SiO_2, is a solid that melts at about 1710 degrees 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 11 is 10 marks)

Mark scheme · AC2D Bonding and Structure: Depth and Exam Drill

Question 1

- (a) **B1** the power (ability) of an atom in a covalent bond to attract the bonding (shared) pair of electrons towards itself, oe
- **(a) Answer: The power of an atom in a covalent bond to attract the bonding pair of electrons towards itself.**
- (b) **B1** a covalent bond in which both of the shared electrons in the bonded pair are provided by the same atom, oe
- **(b) Answer: A covalent bond in which both shared electrons come from the same atom.**
- (c) **B1** 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
- **(c) Answer: A permanent small charge separation across a bond, delta+ on one atom and delta- on the other, caused by unequal electronegativities.**
- (d) **B1** weak intermolecular forces caused by instantaneous (temporary) dipoles inducing dipoles in neighbouring molecules or atoms, oe
- **(d) Answer: Weak intermolecular forces caused by instantaneous dipoles inducing dipoles in neighbouring particles.**

Question 2

- (a) **B1** 5 bonding pairs and 0 lone pairs around P
- (a) **B1** trigonal bipyramidal
- (a) **B1** bond angles of 90 degrees (axial to equatorial) and 120 degrees (equatorial to equatorial)
- **(a) Answer: 5 bonding pairs, 0 lone pairs; trigonal bipyramidal; 90 degrees and 120 degrees.**
- (b) **B1** octahedral
- (b) **B1** 90 degrees
- **(b) Answer: Octahedral; 90 degrees.**
- (c) **B1** square planar
- (c) **B1** 90 degrees
- **(c) Answer: Square planar; 90 degrees.**
- (d) **B1** bent (V-shaped, non-linear)
- (d) **B1** 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
- **(d) Answer: Bent, approximately 115 degrees.**

Question 3

- (a) **B1** the magnitude increases as the product of the ionic charges increases, oe
- (a) **B1** the magnitude increases as the sum of the ionic radii decreases (the ions get closer together), oe
- **(a) Answer: Magnitude increases with greater ionic charge and with smaller sum of ionic radii.**
- (b) **B1** MgO: 212 pm
- (b) **B1** CaO: 240 pm
- (b) **B1** SrO: 258 pm
- **(b) Answer: MgO = 212 pm, CaO = 240 pm, SrO = 258 pm.**
- (c) **M1** 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
- (c) **A1** correct order: MgO (most exothermic) > CaO > SrO (least exothermic)
- (c) **B1** 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
- **(c) Answer: MgO more exothermic than CaO, which is more exothermic than SrO.**
- (d) **M1** correct substitution: 212/240
- (d) **A1** 0.88 (cao, 2 sf), so the lattice enthalpy of CaO is predicted to be about 88% of the magnitude of that of MgO
- **(d) Answer: 0.88; CaO's lattice enthalpy magnitude is predicted to be about 88% of MgO's.**

Question 4

- (a) **M1** convert 564.0 pm to cm: 5.640 x 10^{-8} cm
- (a) **M1** cube the edge length: (5.640 x 10^{-8})^3
- (a) **A1** cso: 1.79 x 10^{-22} cm^3 (3 sf), as required
- **(a) Answer: 1.79 x 10^{-22} cm^3 (shown)**
- (b) **M1** correct identification of Z = 4 and M = 58.5 g/mol in the formula
- (b) **M1** correct substitution: (4 x 58.5) / (6.02 x 10^{23} x 1.79 x 10^{-22})
- (b) **A1** 2.17 g/cm^3 (accept 2.16-2.17), cao
- **(b) Answer: 2.17 g/cm^3**
- (c) **M1** percentage uncertainty in a = (2/564.0) x 100
- (c) **M1** recognise percentage uncertainty in density = 3 x percentage uncertainty in a
- (c) **A1** 1.1% (accept 1.0-1.1%), cao
- **(c) Answer: Approximately 1.1%**
- (d) **M1** absolute uncertainty = (1.1/100) x 2.17
- (d) **A1** 0.02 g/cm^3 (accept 0.02-0.03)
- (d) **B1** density = (2.17 ± 0.02) g/cm^3, with the uncertainty quoted to the same number of decimal places as the density
- **(d) Answer: (2.17 ± 0.02) g/cm^3**

Question 5

- (a) **B1** Na-Cl: 2.1
- (a) **B1** C-Cl: 0.5
- (a) **B1** C-H: 0.4
- (a) **B1** O-H: 1.4
- **(a) Answer: Na-Cl = 2.1, C-Cl = 0.5, C-H = 0.4, O-H = 1.4.**
- (b) **B1** Na-Cl: ionic
- (b) **B1** C-Cl: polar covalent
- (b) **B1** C-H: essentially non-polar covalent (accept weakly polar covalent)
- (b) **B1** O-H: polar covalent
- **(b) Answer: Na-Cl ionic; C-Cl polar covalent; C-H (essentially) non-polar covalent; O-H polar covalent.**
- (c) **B1** both molecules are symmetrical (tetrahedral), so the individual bond dipoles are arranged symmetrically and cancel out (the vector sum of the bond dipoles is zero), oe
- **(c) Answer: The bond dipoles are arranged symmetrically in the tetrahedral molecules and cancel out.**

Question 6

- (a) **B1** boiling point increases from F2 to I2
- (a) **B1** the number of electrons (and hence the size of the electron cloud) increases as Mr increases down the group, making the molecules more easily polarised, oe
- (a) **B1** this gives stronger van der Waals (London/induced dipole-dipole) forces between molecules, so more energy is required to separate them, raising the boiling point, oe
- **(a) Answer: Boiling point rises from F2 to I2 because increasing numbers of electrons give stronger van der Waals forces.**
- (b) **M1** correct method to find the gradient between the Br2 and I2 points: (184-59)/(254-160), oe
- (b) **M1** correct use of this gradient to extrapolate from I2 (Mr = 254, bp = 184 degrees C) to Mr = 420
- (b) **A1** answer consistent with the student's own extrapolation, e.g. approximately 400-410 degrees C, ecf
- **(b) Answer: Approximately 405 degrees C (accept any value consistent with a valid linear extrapolation).**
- (c) **B1** the relationship between boiling point and Mr for the halogens is not perfectly linear (the increase in van der Waals forces per unit increase in Mr is not constant), so extending a straight-line trend beyond the measured data may not be accurate, oe
- (c) **B1** astatine is extremely rare and highly radioactive, so only minute quantities have ever been studied, meaning there is little reliable experimental data to confirm the trend continues, oe
- **(c) Answer: Extrapolation beyond the data assumes an unproven linear trend, and At2 data are extremely scarce.**

Question 7

- (a) **B1** 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
- (a) **B1** 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
- (a) **B1** 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
- **(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) **B1** 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
- (b) **B1** 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
- (b) **B1** since both compounds have the same Mr, their van der Waals forces are of similar strength, oe
- (b) **B1** 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
- **(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) **B1** in ice, each water molecule forms hydrogen bonds to (up to) four neighbouring water molecules, oe
- (c) **B1** 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
- **(c) Answer: The open, hydrogen-bonded hexagonal lattice of ice holds molecules further apart than in liquid water.**

Question 8

- **Level 3** (5-6): 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.
- **Level 2** (3-4): 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.
- **Level 1** (1-2): 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.
- **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 degrees C) than the sublimation point of dry ice (-78 degrees 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 9

- (a) **B1** ± 0.5 degrees C
- **(a) Answer: ± 0.5 degrees C**
- (b) **M1** correct substitution: (0.5/182.0) x 100
- (b) **A1** 0.27% (accept 0.27-0.28%), cao
- **(b) Answer: 0.27%**
- (c) **M1** melting range = 184.0 - 182.0 = 2.0 degrees C
- (c) **M1** absolute uncertainties are added when two readings are subtracted: total uncertainty = 0.5 + 0.5 = ± 1.0 degrees C
- (c) **A1** melting range = (2.0 ± 1.0) degrees C
- **(c) Answer: (2.0 ± 1.0) degrees C**
- (d) **M1** correct substitution: (1.0/2.0) x 100 = 50%
- (d) **B1** 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
- **(d) Answer: 50%; the apparatus gives a very large percentage uncertainty for a small temperature difference.**

Question 10

- (a) **B1** a cation with a high charge density (a small ionic radius and/or a high charge) has strong polarising power, oe
- (a) **B1** an anion with a large ionic radius is easily distorted and is described as highly polarisable, oe
- (a) **B1** when a small, highly charged cation strongly polarises (distorts the electron cloud of) a large, polarisable anion, electron density is drawn into the region between the ions, giving the bond significant covalent character, oe
- **(a) Answer: A small, highly charged cation strongly polarising a large, polarisable anion gives a bond significant covalent character.**
- (b) **M1** Be^{2+} has a smaller ionic radius than Mg^{2+}, for the same ionic charge (+2)
- (b) **B1** so Be^{2+} has a higher charge density and is more strongly polarising than Mg^{2+}
- (b) **B1** Be^{2+} polarises the large, polarisable Cl^- ion more strongly than Mg^{2+} does, drawing more electron density into the bond, so $BeCl_2$ has the greater degree of covalent character, oe
- **(b) Answer: $BeCl_2$ has the greater degree of covalent character.**
- (c) **M1** $BeCl_2$ has the lower melting point
- (c) **B1** $BeCl_2$ has more covalent character, so it exists as discrete molecules/small units held together only by weaker intermolecular forces (e.g. van der Waals forces), oe
- (c) **B1** $MgCl_2$ has more ionic character, forming a giant ionic lattice held together throughout by strong electrostatic forces of attraction between oppositely charged ions, requiring more energy to melt, oe
- **(c) Answer: $BeCl_2$ has the lower melting point.**

Question 11

- (a) **B1** 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
- (a) **B1** 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
- **(a) Answer: CO_2 is simple molecular; SiO_2 is giant covalent.**
- (b) **B1** 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
- (b) **B1** melting SiO_2 requires breaking a very large number of strong covalent Si-O bonds throughout the giant covalent lattice, oe
- (b) **B1** 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
- (b) **B1** 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
- **(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) **M1** recognise that 4 mol of Si-O bonds must be broken per mole of SiO_2
- (c) **M1** correct substitution: 4 x 466
- (c) **A1** 1864 kJ/mol (accept 1860-1870), cao
- (c) **B1** 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
- **(c) Answer: Approximately 1864 kJ/mol; the value uses an average bond enthalpy, which assumes every Si-O bond is identical.**