MEGA LECTURE · A LEVEL CHEMISTRY

Chapter 5: Electrons, Bonding and Structure

OCR A Level Chemistry A (H432) — Structured Practice Worksheet

SPECIFICATION	TOTAL MARKS	SUGGESTED TIME
2.1.1 Electron configuration · 2.1.2 Ionic bonding · 2.1.3 Covalent bonding & shapes · 2.1.4 Metallic bonding	104	2 hours

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Electron configuration · Ionic, covalent & metallic bonding · Shapes of molecules	104	2 hours

Instructions to candidates

- Answer **all** questions in the spaces provided.
- Draw dot-and-cross diagrams showing outer-shell electrons only, unless told otherwise.
- Marks are shown in brackets [ ] at the end of each part.
- You may use the periodic table on the OCR Data Sheet.

Question 1 — Electron configuration

2.1.1

a. Write the full electron configuration (using sub-shell notation) for: (i) a chlorine atom, (ii) a Ca^{2+} ion. **[4]**

b. Write the electron configuration of an Fe^{3+} ion, and explain why electrons are removed from the 4s sub-shell before the 3d sub-shell. **[2]**

c. Define the terms *shell*, *sub-shell* and *orbital*. **[3]**

d. State the maximum number of electrons that can occupy an s sub-shell and a p sub-shell. **[2]**

e. Sketch the shape of an s orbital and the shape of a p orbital. **[2]**

Sketch both orbitals

Question 1 total: 13 marks

Question 2 — Ionisation energy and the periodic table

2.1.1

a. Define *first ionisation energy*. **[2]**

b. Successive ionisation energies of an element are: 590, 1145, 4912, 6474, 8144 kJ mol^{-1}. Explain the large increase between the second and third values, and deduce the group of the periodic table to which the element belongs. **[3]**

c. Explain why the first ionisation energy generally increases across Period 3 (Na to Ar). **[3]**

d. Explain why the first ionisation energy of aluminium is lower than that of magnesium, despite aluminium having a greater nuclear charge. **[3]**

e. Explain why the first ionisation energy decreases down Group 2. **[2]**

Question 2 total: 13 marks

Question 3 — Ionic bonding

2.1.2

a. Define *ionic bonding*. **[1]**

b. Draw a dot-and-cross diagram to show the bonding in magnesium oxide, MgO. Show outer-shell electrons only and the charges on each ion. **[3]**

Draw the dot-and-cross diagram

c. Draw a dot-and-cross diagram to show the bonding in calcium chloride, $CaCl_2$. **[3]**

Draw the dot-and-cross diagram

d. Describe the structure of a giant ionic lattice such as sodium chloride. **[2]**

e. Explain why ionic compounds have high melting points but conduct electricity only when molten or in aqueous solution, not as solids. **[3]**

Question 3 total: 12 marks

Question 4 — Covalent bonding and dot-and-cross diagrams 2.1.3

a. Define *covalent bond*. **[1]**

b. Draw dot-and-cross diagrams, showing outer-shell electrons only, for: (i) CO_2, (ii) NH_3, (iii) CH_4. **[6]**

Draw all three dot-and-cross diagrams

c. Draw a dot-and-cross diagram for the ammonium ion, NH_4^+, and use it to define the term *dative covalent (coordinate) bond*. **[3]**

Draw NH_4^+

d.

Explain, in terms of electron pair repulsion, why a dative covalent bond is identical in every way to a normal covalent bond once formed. **[2]**

Question 4 total: 12 marks

Question 5 — Shapes of molecules (VSEPR) 2.1.3

a. State the basic principle of electron pair repulsion theory used to predict molecular shapes. **[2]**

b. For each species, state the shape and bond angle, and draw a 3D diagram showing the shape. **[12]**

Species	Number of bonding pairs	Number of lone pairs	Shape	Bond angle
BF_3				
CH_4				
NH_3				
H_2O				
PCl_5				
SF_6				

c. Explain why the bond angle in NH_3 (107°) is smaller than in CH_4 (109.5°). **[3]**

d. Explain why the bond angle in H_2O (104.5°) is smaller than in NH_3. **[2]**

Question 5 total: 19 marks

Question 6 — Electronegativity and bond polarity 2.1.3

a. Define *electronegativity*. **[2]**

b. State and explain the general trend in electronegativity across Period 3. **[2]**

c. Explain why the H–F bond is polar, using the symbols δ+ and δ− and an arrow to show the bond dipole. **[3]**

d. CO_2 contains polar C=O bonds but is a non-polar molecule overall. Explain why, referring to the shape of the molecule. **[3]**

e. Explain why $CHCl_3$ is a polar molecule. **[2]**

Question 6 total: 12 marks

Question 7 — Metallic bonding 2.1.4

a. Describe the structure and bonding present in a metal. **[2]**

b. Explain why metals conduct electricity in both the solid and liquid state. **[2]**

c. Explain why magnesium has a higher melting point than sodium. **[2]**

d. Explain why metals are malleable, while ionic solids are brittle and shatter when struck. **[2]**

Question 7 total: 8 marks

Question 8 — Bonding continuum and comparing structures

2.1.2 – 2.1.4

a. Explain what is meant by the statement that "bonding is a continuum between ionic and covalent" and describe how ionic character can increase in a covalent bond. **[3]**

b. Aluminium chloride, $AlCl_3$, has significant covalent character. Explain, using ionic charge and radius, why this is the case. **[3]**

c. Complete the table, stating the type of bonding/structure and the type of particle present. **[3]**

Substance	Bonding / structure type	Particles present
Sodium chloride		
Magnesium		
Iodine (I_2)		

Question 8 total: 9 marks

Question 9 — Extended response

Synoptic

Explain, with reference to structure and bonding, why sodium chloride has a much higher melting point than water, but a lower melting point than silicon dioxide (SiO_2, a giant covalent structure). Your answer should refer to the type of particle, the forces holding the particles together, and the energy needed to overcome them. **[6]**

This question is marked using a level of response.

Question 9 total: 6 marks

PAPER TOTAL: 104 MARKS

MARK SCHEME — Chapter 5: Electrons, Bonding and Structure

Award marks independently unless stated.

Q1 (a) [4]

- Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$. (2)
- Ca^{2+}: $1s^2 2s^2 2p^6 3s^2 3p^6$ (loses the two 4s electrons). (2)

Q1 (b) [2]

- Fe^{3+}: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$. (1)
- 4s electrons are removed first because, once occupied, the 4s sub-shell is at a higher energy than 3d in the transition metal ion. (1)

Q1 (c) [3]

- Shell: a group of sub-shells with the same principal quantum number, n. (1)
- Sub-shell: a group of orbitals of the same energy within a shell (s, p, d, f). (1)
- Orbital: a region of space that can hold up to 2 electrons (with opposite spins). (1)

Q1 (d) [2]

- s sub-shell: maximum 2 electrons. (1)
- p sub-shell: maximum 6 electrons. (1)

Q1 (e) [2]

- s orbital: sphere. (1)
- p orbital: "dumb-bell" / figure-of-eight shape. (1)

Q2 (a) [2]

- The energy required to remove one electron from each atom in one mole of gaseous atoms... (1)
- ...to form one mole of gaseous 1+ ions. (1)

Q2 (b) [3]

- The large jump shows the third electron is removed from a shell closer to the nucleus (inner shell). (1)
- This electron experiences less shielding and greater nuclear attraction, so much more energy is needed. (1)
- Two electrons are easily removed (outer shell) — the element is in Group 2. (1)

Q2 (c) [3]

- Nuclear charge increases across the period (more protons). (1)
- Shielding stays approximately constant (electrons added to the same outer shell). (1)
- Increased nuclear attraction on the outer electrons, so more energy is needed to remove one. (1)

Q2 (d) [3]

- Aluminium's outer electron is in a 3p sub-shell, magnesium's is in 3s. (1)
- The 3p sub-shell is at a slightly higher energy / further from the nucleus (more shielded by the 3s electrons). (1)
- So the 3p electron in Al is easier to remove despite the greater nuclear charge. (1)

Q2 (e) [2]

- Down the group, atomic radius increases and extra electron shells add more shielding. (1)
- Both effects outweigh the increased nuclear charge, so the outer electron is easier to remove. (1)

Q3 (a) [1]

- The electrostatic attraction between oppositely charged ions. (1)

Q3 (b) [3]

- Mg loses 2 electrons to form Mg^{2+} (no dots/crosses remaining shown as $[Mg]^{2+}$). (1)
- O gains 2 electrons to form O^{2-}, shown with 8 dots/crosses in square brackets. (1)
- Correct charges shown outside brackets (2+ and 2−). (1)

Q3 (c) [3]

- Ca loses 2 electrons to form Ca^{2+}. (1)
- Two separate Cl atoms each gain 1 electron to form two Cl^- ions (8 dots/crosses each in brackets). (1)
- Correct charges shown: Ca^{2+} and two Cl^-. (1)

Q3 (d) [2]

- A regular, repeating 3D arrangement of alternating positive and negative ions... (1)
- ...held together by strong electrostatic forces acting in all directions. (1)

Q3 (e) [3]

- High melting point: strong electrostatic forces between ions act in all directions, requiring a lot of energy to overcome. (1)
- Solid: ions are fixed in place in the lattice and cannot move to carry charge. (1)
- Molten/aqueous: ions are free to move and can carry charge. (1)

Q4 (a) [1]

- A shared pair of electrons between two atoms. (1)

Q4 (b) [6]

- CO_2: two C=O double bonds, each sharing 2 pairs, linear, all outer electrons shown correctly. (2)
- NH_3: three N–H single bonds plus one lone pair on N shown. (2)
- CH_4: four C–H single bonds, all electrons shown correctly. (2)

Q4 (c) [3]

- NH_4^+: four N–H bonds around N, one shown with both electrons originating from N (arrow or different symbol), overall 1+ charge shown. (2)
- Dative covalent bond: a shared pair of electrons where both electrons are donated by one atom. (1)

Q4 (d) [2]

- Once formed, the bonding pair cannot be distinguished from a normal covalent bond — both electrons are shared equally in the bond. (1)
- Electron pair repulsion treats it identically to any other bonding pair when predicting shape. (1)

Q5 (a) [2]

- Electron pairs (bonding and lone) around a central atom repel each other... (1)
- ...and arrange themselves as far apart as possible to minimise repulsion, determining the molecular shape. (1)

Q5 (b) [12] — 2 marks each (shape + angle; ecf for diagram)

- BF_3: 3 bp, 0 lp — trigonal planar, 120°. (2)
- CH_4: 4 bp, 0 lp — tetrahedral, 109.5°. (2)
- NH_3: 3 bp, 1 lp — pyramidal (trigonal pyramidal), 107°. (2)
- H_2O: 2 bp, 2 lp — bent (non-linear/V-shaped), 104.5°. (2)
- PCl_5: 5 bp, 0 lp — trigonal bipyramidal, 90° and 120°. (2)
- SF_6: 6 bp, 0 lp — octahedral, 90°. (2)

Q5 (c) [3]

- Lone pairs are held closer to the nucleus and are more concentrated / occupy more space than bonding pairs. (1)
- Lone pair–bonding pair repulsion is greater than bonding pair–bonding pair repulsion. (1)
- This pushes the N–H bonds closer together, reducing the angle from 109.5° to 107°. (1)

Q5 (d) [2]

- H_2O has two lone pairs (compared with one in NH_3). (1)
- Lone pair–lone pair repulsion is even greater, compressing the bond angle further to 104.5°. (1)

Q6 (a) [2]

- The ability of an atom (in a covalent bond)... (1)
- ...to attract the bonding (shared) pair of electrons towards itself. (1)

Q6 (b) [2]

- Electronegativity increases across Period 3. (1)
- Nuclear charge increases while shielding stays similar, so the nucleus attracts the bonding pair more strongly. (1)

Q6 (c) [3]

- F is more electronegative than H, so it attracts the bonding pair more strongly. (1)
- δ+ on H, δ− on F. (1)
- Arrow drawn from H towards F along the bond, ending in a bar/cross. (1)

Q6 (d) [3]

- CO_2 is linear (bond angle 180°). (1)
- The two C=O bond dipoles are equal in magnitude and point in exactly opposite directions. (1)
- They cancel out, so there is no overall (net) molecular dipole. (1)

Q6 (e) [2]

- $CHCl_3$ is tetrahedral but not symmetric (three C–Cl and one C–H bond, different dipoles). (1)
- The bond dipoles do not cancel, so there is a net molecular dipole. (1)

Q7 (a) [2]

- A lattice of positive metal ions... (1)
- ...surrounded by / in a "sea of" delocalised electrons. (1)

Q7 (b) [2]

- The delocalised electrons are mobile / free to move throughout the structure in both states. (1)
- They can carry charge/current when a potential difference is applied. (1)

Q7 (c) [2]

- Mg^{2+} has a greater ionic charge than Na^+, and contributes two delocalised electrons per atom instead of one. (1)
- Stronger electrostatic attraction between the ions and the delocalised electrons, so more energy is needed to melt it. (1)

Q7 (d) [2]

- In metals, layers of ions can slide over each other without breaking the (non-directional) metallic bonding — malleable. (1)
- In ionic solids, sliding brings ions of like charge together, causing strong repulsion and the lattice to shatter — brittle. (1)

Q8 (a) [3]

- Real bonds are rarely 100% ionic or 100% covalent — most lie somewhere between the two extremes. (1)
- A very electronegative cation can distort/polarise the electron cloud of a nearby anion. (1)
- This pulls electron density into the region between the ions, giving the bond covalent character. (1)

Q8 (b) [3]

- Al^{3+} has a high charge and small ionic radius, giving it a high charge density. (1)
- This strongly polarises the electron cloud of the Cl^- ions. (1)
- Electron density is distorted towards Al^{3+}, giving the Al–Cl bond significant covalent character. (1)

Q8 (c) [3] — 1 mark each row

- NaCl: giant ionic lattice — ions (Na^+ and Cl^-).
- Mg: giant metallic lattice — metal ions and delocalised electrons.
- I_2: simple molecular, covalent within the molecule — I_2 molecules (held by van der Waals forces).

Q9 [6] — Level of response

- **Level 3 (5–6):** Clear, correct comparison of all three substances covering particle type, bonding/forces, and relative energy required, in a logical structure.
- **Level 2 (3–4):** Correct comparison of two of the three substances, or all three with minor inaccuracies.
- **Level 1 (1–2):** Fragmented statements; only one substance discussed correctly.

Indicative content: *NaCl is a giant ionic lattice held by strong electrostatic forces between ions acting in all directions — high melting point. Water is a simple molecular substance; within each molecule bonding is covalent (strong), but between molecules only (weaker) hydrogen bonds/van der Waals forces need to be overcome to melt it — much lower melting point than NaCl. SiO_2 is a giant covalent (macromolecular) structure in which every atom is joined to its neighbours by strong covalent bonds throughout the lattice; melting requires breaking these strong covalent bonds throughout the giant structure, requiring more energy than breaking the ionic lattice of NaCl — highest melting point of the three.*

END OF MARK SCHEME · 104 marks