MEGA LECTURE

IB DP CHEMISTRY

Structure 2 · Models of Bonding and Structure

MARK SCHEME

S2.1 The Ionic Model · S2.2 The Covalent Model

S2.3 The Metallic Model · S2.4 From Models to Materials

Standard and Higher Level · Syllabus 2025

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Original exam-style questions written for the IB Diploma Programme Chemistry course (first assessment 2025)

Section A — Multiple Choice: Answer Key (15 marks)

1. Answer: **C** — An ionic bond is the electrostatic attraction between oppositely charged ions, formed when one or more electrons are transferred from a metal atom to a non-metal atom.

2. Answer: **B** — In the NaCl (rock-salt) lattice structure, each Na^+ ion is surrounded by 6 Cl^- ions and vice versa, giving a 6:6 coordination.

3. Answer: **B** — Ionic compounds have high melting points due to strong electrostatic forces throughout the giant lattice, and conduct electricity only when the ions are free to move (molten or dissolved), not as a rigid solid.

4. Answer: **C** — Mg^{2+} and O^{2-} have both a higher ionic charge and a smaller ionic radius than the singly charged alternatives, so this pair gives the strongest electrostatic attraction and greatest lattice enthalpy.

5. Answer: **B** — A covalent bond consists of a shared pair of electrons that is attracted electrostatically to the nuclei of both bonded atoms.

6. Answer: **C** — With 4 bonding electron domains and no lone pairs around the central carbon atom, the electron domains (and the resulting molecular shape) arrange themselves tetrahedrally to minimize repulsion, giving bond angles of 109.5°.

7. Answer: **B** — A polar covalent bond arises when two bonded atoms have different electronegativities, causing unequal sharing of the bonding electrons and small partial charges (δ^+ and $\delta-$) on each atom.

8. Answer: **B** — With 4 total electron domains (2 bonding + 2 lone pairs) arranged tetrahedrally, but only the 2 bonded atoms considered for molecular shape, the resulting molecular shape is bent (V-shaped), as seen in H_2O.

9. Answer: **B** — In metallic bonding, metal atoms lose their outer (valence) electrons to form a lattice of cations; these delocalized electrons form a 'sea' that is electrostatically attracted to the surrounding cations, holding the structure together.

10. Answer: **B** — The delocalized electrons in the 'sea' of a metallic lattice are free to move throughout the structure, allowing metals to conduct electricity when a potential difference is applied.

11. Answer: **A** — Because metallic bonding is non-directional, layers of cations can slide over one another when a force is applied, while the delocalized electron sea continues to hold the structure together, giving metals their malleability.

12. Answer: **B** — An alloy is a mixture (not a compound) of a metal with one or more other elements, typically produced to improve properties such as strength or resistance to corrosion compared with the pure metal.

13. Answer: **B** — SiO_2 forms a giant covalent (macromolecular) structure in which every silicon atom is covalently bonded to four oxygen atoms in a continuous 3-D network, giving it a very high melting point.

14. Answer: **B** — In graphite, each carbon atom forms 3 covalent bonds within a layer, leaving one electron per atom delocalized across the layer; these mobile electrons allow graphite to conduct electricity along its layers, unlike diamond where all 4 electrons per carbon are localized in covalent bonds.

15. Answer: **B** — In simple molecular substances, strong covalent bonds hold atoms together within each molecule, but only weak intermolecular forces (e.g. London dispersion forces) act between molecules; melting requires overcoming only these weak forces, giving much lower melting points than giant covalent structures.

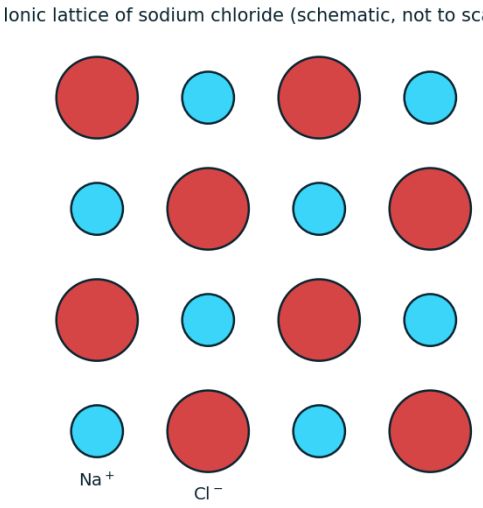
Section B — Structured Questions: Worked Solutions (66 marks)

Mark-scheme notation: M = method mark, A = accuracy mark (dependent on the preceding M mark), R = reasoning mark.

S2.1 The Ionic Model

Question 1 [7 marks]

The diagram shows part of the ionic lattice structure of sodium chloride, NaCl.



Ionic lattice of sodium chloride (schematic, not to scale).

(a) Using the diagram, state the coordination number of the Na^+ ion in this lattice (in the full three-dimensional structure). **[2]**

Coordination number = **6** (each Na^+ ion is surrounded by 6 nearest-neighbour Cl^- ions in the full 3-D lattice, shown partially in the 2-D schematic). **A1 A1**

(b) Explain, in terms of electron transfer, how the ionic bond between sodium and chlorine forms. **[3]**

A sodium atom loses one electron from its outer shell to form a positively charged Na^+ ion; this electron is transferred to a chlorine atom, which gains it to form a negatively charged Cl^- ion. Both ions attain a full outer shell (noble gas electron configuration). The oppositely charged ions then attract each other electrostatically, forming the ionic bond. **R1 R1 A1**

(c) Deduce the empirical formula of sodium chloride from the ratio of ions shown in the lattice. **[2]**

The lattice contains Na^+ and Cl^- ions in a 1:1 ratio, so the empirical formula is **NaCl**. **A1 A1**

Question 2 [6 marks]

Ionic compounds such as sodium chloride share several characteristic physical properties.

(a) State and explain why ionic compounds typically have high melting points. **[2]**

Ionic compounds have **high melting points** because a large amount of energy is required to overcome the strong electrostatic forces of attraction acting throughout the giant ionic lattice. **A1 R1**

(b) Explain why solid sodium chloride does not conduct electricity, but molten or aqueous sodium chloride does. **[2]**

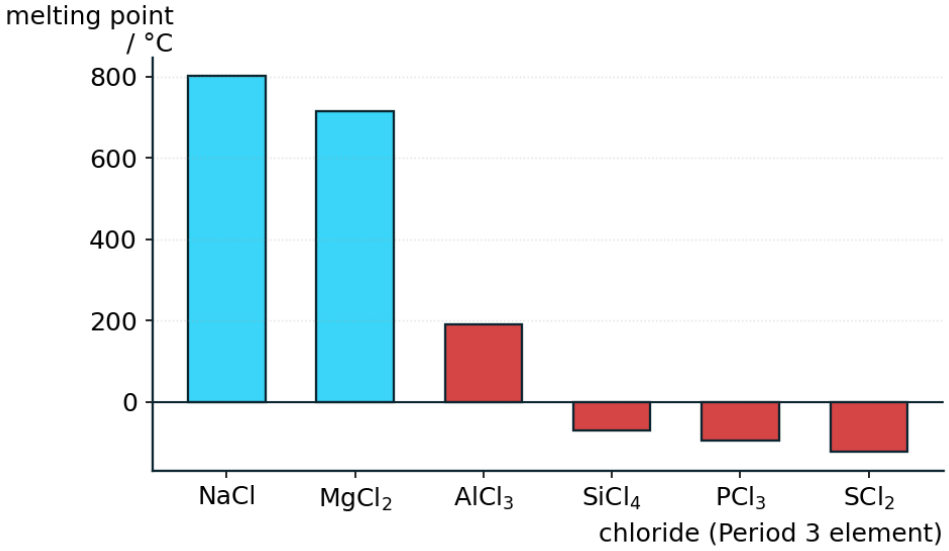
In the solid, the ions are held in fixed positions in the lattice and are not free to move, so there are no mobile charge carriers. When molten or dissolved in water, the ions are free to move and can carry electric charge through the liquid, so it conducts. **R1 R1**

(c) Explain why ionic crystals are brittle and tend to shatter when struck, rather than bending. **[2]**

A sharp blow can shift one layer of ions relative to the next, bringing ions of the **same charge** next to each other. The resulting strong electrostatic **repulsion** between like-charged ions causes the layers to split apart, so the crystal shatters. **R1 R1**

Question 3 [7 marks]

The graph shows the melting points of the chlorides of the Period 3 elements, from sodium to sulfur.



Melting points of Period 3 chlorides (given data).

(a) Using the graph, identify which chlorides are ionic and which are simple molecular covalent, giving a reason based on melting point. **[2]**

NaCl and $MgCl_2$ have very high melting points and are **ionic**. $AlCl_3$, $SiCl_4$, PCl_3 and SCl_2 have much lower melting points and are **simple molecular covalent** substances. **A1 A1**

(b) Explain the sharp fall in melting point between $MgCl_2$ and $AlCl_3$ in terms of bonding type. **[3]**

$MgCl_2$ is ionic, held together by strong electrostatic forces throughout a giant lattice, requiring a large amount of energy to melt. $AlCl_3$ bonding has significant covalent character (Al^{3+} strongly polarizes the Cl^- ions) and exists as a simple molecular structure, held together between molecules only by weak London (dispersion) forces, which require much less energy to overcome. **R1 R1 A1**

(c) Explain why $SiCl_4$ has a very low melting point. **[2]**

$SiCl_4$ is a **simple molecular** substance: strong covalent bonds hold the atoms together within each molecule, but only weak London (dispersion) forces act between separate $SiCl_4$ molecules. Only these weak intermolecular forces need to be overcome to melt the substance, giving a low melting point. **A1 R1**

S2.2 The Covalent Model

Question 4 [7 marks]

Ammonia, NH_3, is a simple covalent molecule with the nitrogen atom as the central atom.

(a) State the number of bonding electron pairs and the number of lone pairs on the central nitrogen atom in NH_3. **[2]**

Nitrogen has 5 valence electrons; 3 are used to form bonding pairs with the 3 hydrogen atoms, leaving 1 **lone pair**. So NH_3 has **3 bonding pairs** and **1 lone pair** on the central N atom. **A1 A1**

(b) Using VSEPR theory, deduce the molecular shape and approximate bond angle of NH_3. **[3]**

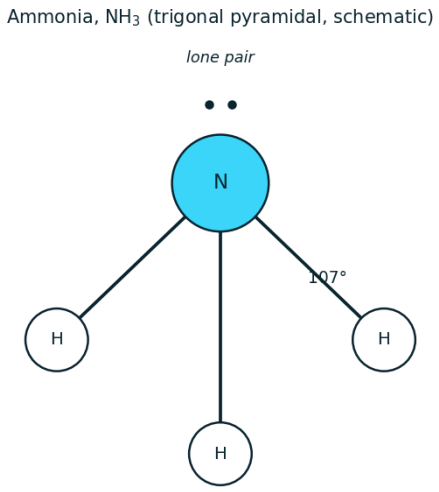
4 electron domains (3 bonding + 1 lone pair) arrange tetrahedrally to minimize repulsion, but since only the 3 bonded H atoms define the molecular shape, NH_3 is **trigonal pyramidal**, with a bond angle of approximately **107°**. **M1 A1 A1**

(c) Explain why the H_3N bond angle (107°) is smaller than the ideal tetrahedral angle of 109.5°. **[2]**

A lone pair occupies more space and repels adjacent bonding pairs more strongly than a bonding pair does (lone pair−bonding pair repulsion > bonding pair−bonding pair repulsion). This compresses the H−N−H bond angle slightly below the ideal tetrahedral angle. **R1 R1**

Question 5 [6 marks]

The diagram shows the shape of an ammonia molecule, NH_3, determined using VSEPR theory.



Ammonia, NH_3 (trigonal pyramidal, schematic).

(a) Identify the molecular shape and bond angle shown in the diagram. **[2]**

Molecular shape = **trigonal pyramidal**; bond angle ≈ **107°**. **A1 A1**

(b) Explain the difference between 'electron domain geometry' and 'molecular geometry' for a molecule such as NH_3. **[2]**

Electron domain geometry describes the arrangement of **all** electron domains (bonding pairs and lone pairs) around the central atom, which is tetrahedral for NH_3. Molecular geometry describes only the positions of the **bonded atoms** (ignoring lone pairs), which for NH_3 is trigonal pyramidal. **R1 R1**

(c) Predict the molecular shape of a water molecule, H_2O, which has 2 bonding pairs and 2 lone pairs on the central oxygen atom. **[2]**

With 2 bonding pairs and 2 lone pairs (4 electron domains, tetrahedral electron domain geometry), the molecular shape of H_2O is **bent (V-shaped)**, with a bond angle of approximately 104.5° (further compressed by 2 lone pairs). **A1 A1**

Question 6 [7 marks]

Consider the carbon-chlorine bond in tetrachloromethane, CCl_4. (Pauling electronegativities: C = 2.55, Cl = 3.16.)

(a) Calculate the electronegativity difference between carbon and chlorine, and use it to explain the polarity of a single C−Cl bond. **[2]**

ΔEN = 3.16 − 2.55 = **0.61**. Since chlorine is more electronegative than carbon, it attracts the shared bonding electrons more strongly, making the C−Cl bond **polar**, with δ^+ on C and $\delta-$ on Cl. **M1 A1**

(b) CCl_4 is a tetrahedral molecule with four identical, symmetrically arranged C−Cl bonds. Deduce whether the CCl_4 molecule as a whole is polar, explaining your reasoning. **[3]**

Although each individual C−Cl bond is polar, the four bond dipoles are arranged symmetrically (tetrahedrally) around the central carbon atom, so they **cancel out exactly** in vector sum. The overall molecule is therefore **non-polar**, despite containing polar bonds. **R1 R1 A1**

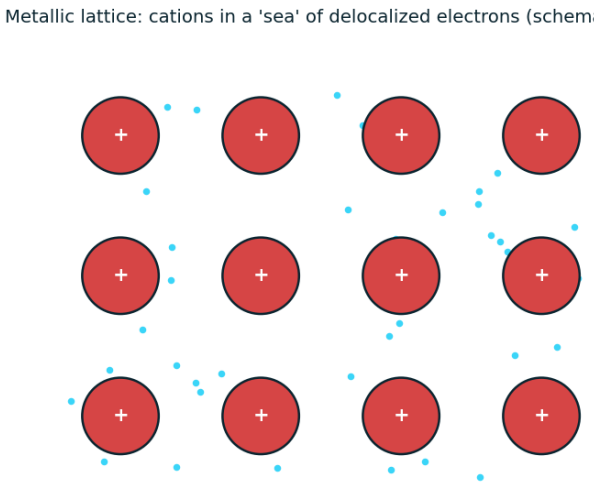
(c) Compare the boiling points expected for CH_4 and CCl_4, explaining your answer in terms of intermolecular forces. **[2]**

Both CH_4 and CCl_4 are non-polar, so only London (dispersion) forces act between molecules. CCl_4 has many more electrons than CH_4 (larger molar mass), giving stronger London forces. CCl_4 therefore has a **higher boiling point** than CH_4. **R1 A1**

S2.3 The Metallic Model

Question 7 [6 marks]

The diagram shows a simplified model of the structure of a typical metal.



Metallic lattice: cations in a 'sea' of delocalized electrons (schematic).

(a) Describe the type of bonding shown in the diagram. **[2]**

The diagram shows **metallic bonding**: a regular lattice of positively charged metal cations, held together by electrostatic attraction to a 'sea' of delocalized (mobile) valence electrons that belong to the structure as a whole rather than to individual atoms. **A1 A1**

(b) Explain why metals are good conductors of electricity, using the diagram to support your answer. **[2]**

The delocalized electrons shown surrounding the cations are **free to move** throughout the lattice. When a potential difference is applied, these mobile electrons can flow through the structure, allowing the metal to conduct electricity. **R1 A1**

(c) Explain why metals are malleable (can be bent or hammered into shape without breaking). **[2]**

Metallic bonding is **non-directional** (not localized between specific atoms), so layers of cations can slide over one another when a force is applied, without breaking any bonds. The delocalized electron sea continues to hold the shifted layers together, so the metal deforms rather than shattering. **R1 R1**

Question 8 [7 marks]

Sodium (Na) and magnesium (Mg) are both Period 3 metals with the giant metallic structure.

(a) Magnesium has a considerably higher melting point than sodium. Explain this difference in terms of metallic bonding. **[3]**

Magnesium ions (Mg^{2+}) have a **greater positive charge** and contribute **2 delocalized electrons** per atom (compared with Na^{+} with charge 1+ and 1 delocalized electron per atom). Mg^{2+} ions are also smaller than Na^{+} ions. Both factors increase the electrostatic attraction between the cations and the delocalized electron sea, making the metallic bonding in magnesium stronger, so more energy (a higher temperature) is required to melt it. **R1 R1 A1**

(b) Explain why alloys, such as bronze (copper and tin), are generally harder than the pure metals from which they are made. **[2]**

Alloying introduces atoms of a **different size** into the metallic lattice. This distorts the regular arrangement of layers, making it more difficult for the layers of cations to slide over each other, so the alloy is **harder** and less malleable than the pure metal. **R1 A1**

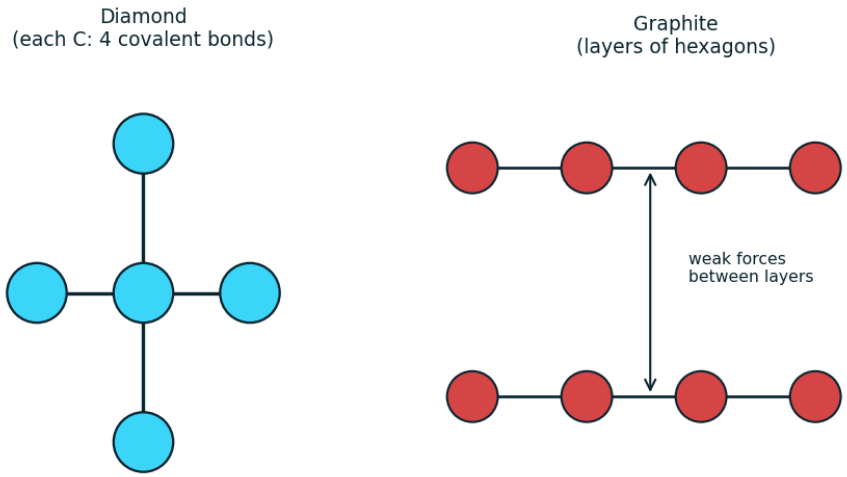
(c) Suggest one practical advantage of using an alloy such as steel (iron and carbon) rather than pure iron in construction. **[2]**

Steel is **stronger and harder** than pure iron (due to the distorted lattice restricting layers from sliding), making it more suitable for load-bearing structures such as beams and reinforcement bars. **A1 R1**

S2.4 From Models to Materials

Question 9 [6 marks]

The diagrams show simplified models of the structures of diamond and graphite, both giant covalent forms of carbon.



Diamond (tetrahedral network) and graphite (layered hexagonal sheets), schematic.

(a) Describe the key structural difference between diamond and graphite shown in the diagrams. **[2]**

In **diamond**, each carbon atom is covalently bonded to **4** other carbon atoms in a rigid three-dimensional tetrahedral network. In **graphite**, each carbon atom is covalently bonded to only **3** other carbon atoms, forming flat hexagonal layers that are stacked on top of one another. **A1 A1**

(b) Explain why graphite conducts electricity but diamond does not. **[2]**

In graphite, only 3 of each carbon's 4 valence electrons are used in covalent bonds; the 4th electron is **delocalized** and free to move within each layer, allowing conduction. In diamond, all 4 valence electrons of each carbon are used in localized covalent bonds, so there are no free (mobile) electrons and diamond does not conduct. **R1 R1**

(c) Explain why graphite is soft and slippery, making it useful as a lubricant, while diamond is extremely hard. **[2]**

The covalent bonds within each layer of graphite are strong, but only weak **London (dispersion) forces** act between adjacent layers, allowing the layers to slide over each other easily, making graphite soft and slippery. In diamond, every atom is held by strong covalent bonds in all three dimensions, so there are no weak layers to slide, making diamond extremely hard. **R1 R1**

Question 10 [7 marks]

Silicon dioxide (SiO_2, found in sand and quartz) and carbon dioxide (CO_2) are both oxides of Group 14 elements, but have very different structures and properties.

(a) Classify the structure of SiO_2 and explain why it has a very high melting point (1710°C) and does not conduct electricity. **[3]**

SiO_2 has a **giant covalent (macromolecular)** structure, in which every Si atom is covalently bonded to 4 O atoms in a continuous 3-D network. Melting requires breaking many strong covalent bonds throughout the giant lattice, requiring a very large amount of energy and giving a very high melting point. All valence electrons are held in localized covalent bonds, so there are no mobile charge carriers and SiO_2 does not conduct electricity. **A1 R1 R1**

(b) CO_2 is a simple molecular covalent substance with a melting point of −78°C. Explain this very large difference in melting point compared with SiO_2. **[2]**

CO_2 exists as small, discrete, simple molecules. Only weak **London (dispersion) forces** act between separate CO_2 molecules, and these require far less energy to overcome than the many strong covalent bonds that must be broken throughout SiO_2's giant covalent lattice, giving CO_2 a much lower melting point. **R1 A1**

(c) Graphene (a single layer of graphite) is an example of a nanostructured material. Suggest one property of graphene and a related application based on this property. **[2]**

Graphene has very **high electrical conductivity** (due to delocalized electrons within its single hexagonal carbon layer), making it a candidate material for use in flexible electronic devices and high-performance transistors. (Other valid properties/applications, e.g. extremely high tensile strength for composite materials, are also accepted.) **A1 A1**