

PHYSICAL SCIENCES REVISION: INTERMOLECULAR FORCES

Questions 1 – 10 with Comprehensive Marking Memorandum

PART A: PRACTICE QUESTIONS (QUESTIONS 1 – 10)

QUESTION 1

[12 Marks]

The table below shows the boiling points of group 4 compounds and hydrogen chloride:

Compound	Formula	Boiling Point (°C)
A	CH ₄	-161.6
B	SiH ₄	-112.0
C	GeH ₄	-88.5
D	SnH ₄	-52.0
E	HCl	X

- 1.1 Define *boiling point*. (2)
- 1.2 Write down the phase of compounds **A** to **D** at room temperature. (1)
- 1.3 Explain the trend in the boiling points of compounds **A** to **D**. (3)
- 1.4 Which compound in the table (**A** to **D**) will have the highest vapour pressure? Give a reason for your answer by referring to the data in the table. (2)
- 1.5 How will the value of **X** compare to the boiling point of CH₄? Choose from **HIGHER THAN**, **LOWER THAN** (1) or **EQUAL TO**.
- 1.6 Fully explain your answer to QUESTION 1.5. (3)

QUESTION 2**[12 Marks]**

The table below shows the relationship between the melting points of three molecules:

Molecule	Melting Point (°C)
CH ₄	-182.5
CF ₄	-150.0
CCl ₄	-23.0

2.1 Define the term *melting point*. (2)

2.2 Explain the trend in melting points of the molecules in the above table by referring to the intermolecular forces and energy involved. (4)

2.3 Which molecule in the table will have the highest vapour pressure at a given temperature? Explain the answer by referring to the data in the table. (2)

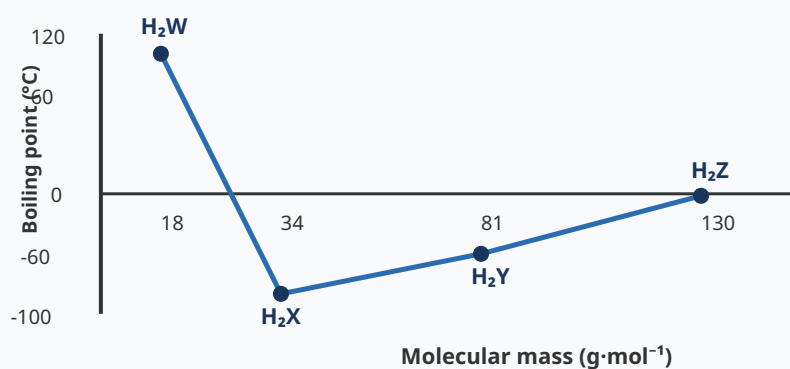
2.4 Water (H₂O) is a smaller molecule than CCl₄, but water has a higher melting point than CCl₄. Explain this observation by referring to the type of intermolecular forces involved. (2)

2.5 Write down the name of the intermolecular force that will exist in a mixture of H₂O and CCl₄. (2)

QUESTION 3

[12 Marks]

The graph below shows the relationship between the boiling points of the hydrides of Group 16 elements and molecular mass:



3.1 Define the term *boiling point*. (2)

3.2 Consider the Group 16 compounds, excluding H₂W:

3.2.1 State the relationship between the molecular mass and the boiling points of the compounds. (2)

3.2.2 Explain this trend in boiling points. Refer to molecular mass, intermolecular forces and energy in your answer. (3)

3.3 H₂W does not follow the trend in boiling point for the Group 16 elements:

3.3.1 Identify the compound H₂W. (1)

3.3.2 Name the intermolecular forces present in this compound. (2)

3.3.3 Explain why the boiling point is higher than expected. (2)

QUESTION 4**[16 Marks]**

4.1 The boiling points in the table below were obtained during an investigation carried out to compare the boiling points of group 7 hydrides. The hydrides HCl and HBr are labelled as compounds **A** and **B** respectively:

Compound	Formula	Molecular Mass ($\text{g}\cdot\text{mol}^{-1}$)	Boiling Point ($^{\circ}\text{C}$)
A	HCl	36.5	-85
B	HBr	81.0	-66

4.1.1 Define the term *boiling point*. (2)

4.1.2 In what phase are compounds **A** and **B** at 0°C and 101.3 kPa of external pressure? (1)

4.1.3 Name the type of intermolecular force that exists between molecules of both compounds **A** and **B** due to the polar nature of these molecules. (1)

4.1.4 Which ONE of the compounds (**A** or **B**) has STRONGER London forces (dispersion forces)? Give a reason for your answer. (3)

4.1.5 Which compound (**A** or **B**) will have the higher vapour pressure? Give a reason for the answer by referring to the data in the table. (2)

Compound **C** (HF), not shown in the table, has a smaller molecular size than both compounds **A** and **B**, but has a relatively higher boiling point of 19.5°C .

4.1.6 Explain why the boiling point of compound **C** is HIGHER than that of compounds **A** and **B** by referring to the TYPE and STRENGTH of intermolecular forces involved. (3)

4.2 Solid iodine (I_2) is added to equal volumes of carbon tetrachloride (CCl_4) and water in separate test tubes.

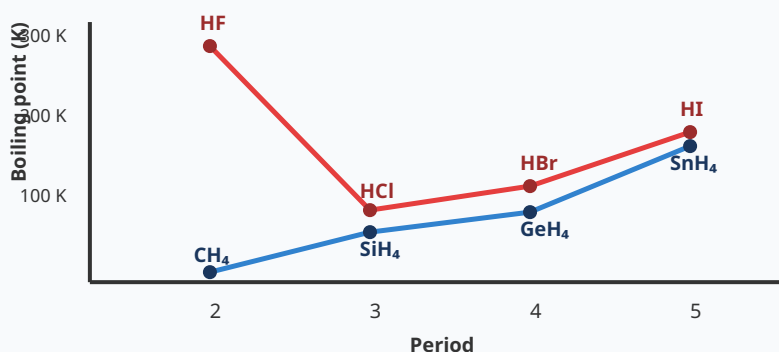
4.2.1 In which liquid (CCl_4 or H_2O) will the iodine dissolve? (1)

4.2.2 Explain the answer to QUESTION 4.2.1 above by referring to the intermolecular forces involved. (3)

QUESTION 5

[11 Marks]

The boiling points of the hydrogen halides and group 4 hydrogen compounds are compared in the graph below:



5.1 Define *boiling point*. (2)

5.2 Write down the boiling point of HCl from the graph. (1)

5.3 Explain why the boiling points of the hydrogen halides are higher than those of corresponding group 4 hydrides from period 3 to 5, by referring to the type of intermolecular forces present in these compounds and energy involved. (4)

HF is the halide with the HIGHEST boiling point.

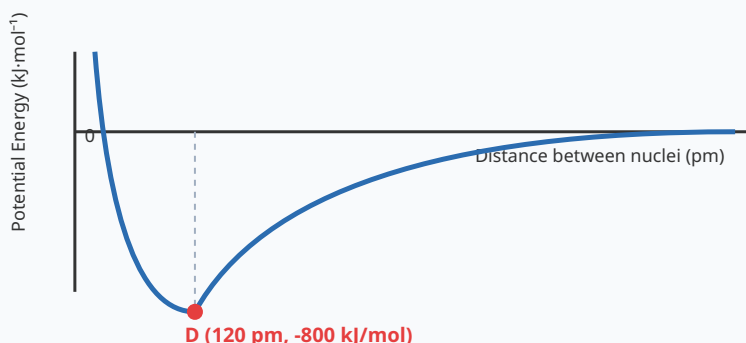
5.4 Write down the name of the intermolecular force present in HF responsible for the high boiling point. (2)

5.5 Which one of HBr and GeH₄ will have the highest vapour pressure? Give a reason for the answer by referring to data in the graph. (2)

QUESTION 6

[17 Marks]

The graph below shows the change in potential energy that takes place when an oxygen atom approaches a carbon atom during the formation of a C=O bond:



6.1 Define the term *bond length*. (2)

6.2 From the graph, write down the:

6.2.1 Bond length of the C=O bond. (1)

6.2.2 Bond energy required to break the C=O bond. (2)

6.2.3 Name of potential energy represented at point D. (2)

6.3 How will the bond length of a C=S bond compare to that of the C=O bond? Write down **EQUAL TO**, **SHORTER THAN** or **LONGER THAN**. Give a reason for your answer. (3)

6.4 What can be said about the forces between the two atoms at point D? (2)

6.5 The table below shows the bond length and bond energy for different bonds between two carbon atoms:

Bond	Type	Length (pm)	Energy (kJ·mol ⁻¹)
A	C–C	154	348
B	C=C	134	614
C	C≡C	120	839

6.5.1 Describe the relationship between bond length and bond energy as shown in the table. (2)

6.5.2 Which ONE of the bonds (**A**, **B** or **C**) will be the weakest? Explain. (3)

QUESTION 7

[27 Marks]

The boiling points of four compounds of hydrogen at standard pressure are given in the table below:

Formula	Boiling Point (°C)
CH ₄	-164
NH ₃	-33
H ₂ O	100
SiH ₄	-112

7.1 Define the term *boiling point*. (2)

7.2 Fully explain the difference in boiling points between CH₄ and:

7.2.1 NH₃ (3)

7.2.2 SiH₄ (3)

7.3 Explain why the boiling points of NH₃ and H₂O differ by referring to electronegativity, molecular shapes, and intermolecular forces. (4)

7.4 Select substances from the table above which are classified as:

7.4.1 Non-polar molecules (2)

7.4.2 Polar molecules (2)

7.5 Consider the diagram showing an arrangement of water molecules in the liquid phase:

[Diagram Representation]

Water molecule H–O–H where **X** represents the intramolecular covalent bond inside H₂O, and **Y** represents the dotted intermolecular hydrogen bond between neighboring H₂O molecules.

7.5.1 Explain the term *interatomic bond*. (2)

7.5.2 Name the specific type of interatomic bond represented by letter **X** in the diagram. (2)

7.5.3 Define the term *intermolecular force*. (2)

7.5.4 Name the specific type of intermolecular force represented by letter **Y** in the diagram. (2)

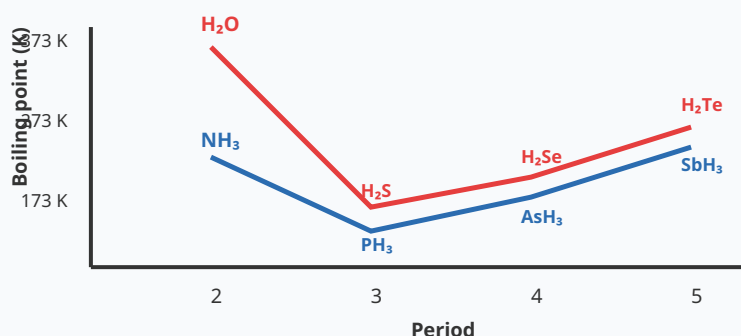
7.5.5 State TWO properties of the oxygen atom that make this type of intermolecular force **Y** possible. (2)

7.6 What is the partial charge on the hydrogen atom in the water molecule? (1)

QUESTION 8

[15 Marks]

The graph below shows the boiling points of different substances from groups 15 and 16 on the periodic table:



8.1 Define the term *boiling point* in words. (2)

8.2 Explain, in terms of the strengths of intermolecular forces, why the boiling points of H₂O and NH₃ are much higher than those of the rest of group 15 and 16. (3)

8.3 The boiling points increase from H₂S to H₂Te:

8.3.1 Identify the strongest type of intermolecular forces present. (1)

8.3.2 Explain the trend that can be seen in their boiling points. (2)

8.3.3 Give the formula of one of the above mentioned compounds that will be a liquid at 0 °C (273 K). (1)

8.4 In the laboratory you are provided with the following chemicals: Cl₂, SO₂ and NaBr.

8.4.1 Identify the intermolecular forces / chemical bonding in each of the above compounds. (3)

8.4.2 Which ONE of the above chemicals will have the highest vapour pressure? Explain the answer. (3)

QUESTION 9**[18 Marks]**

Consider the list of six substances with their formulae and boiling points in the table below:

Name	Formula	Boiling Point (°C)
Water	H ₂ O	100.0
Ethanol	CH ₃ CH ₂ OH	78.0
Bromine	Br ₂	58.8
Iodine	I ₂	184.3
Ammonia	NH ₃	-33.3
Phosphine	PH ₃	-87.7

9.1 Define the term *boiling point*. (2)

9.2 Explain why the boiling point of iodine is so much higher than that of bromine, by referring to the intermolecular forces present in each substance. (4)

9.3 Compare ammonia and phosphine and explain the difference in their boiling points by referring to the intermolecular forces present in each substance. (4)

9.4 Define the term *vapour pressure*. (2)

9.5 Which substance in the table above will have the lowest rate of evaporation? Give a reason for your answer. (3)

9.6 Explain why iodine is not soluble in water by referring to the intermolecular forces in each substance. (3)

QUESTION 10**[8 Marks]**

The table below shows the boiling points of ammonia and water respectively:

Name of Substance	Boiling Point (°C)
Ammonia (NH ₃)	-33 °C
Water (H ₂ O)	100 °C

10.1 Define *boiling point*. (2)

10.2 Explain the difference in boiling points between ammonia and water. (4)

10.3 Which substance will have a lower vapour pressure? Give a reason for the answer. (2)

PART B: MARKING GUIDELINES / MEMORANDUM

MEMORANDUM – QUESTION 1

1.1 The temperature at which the vapour pressure of a liquid equals atmospheric (external) pressure. ✓✓ (2)

1.2 Gas phase ✓ (1)

1.3 From A to D:

- The molecular size / surface area increases. ✓
- The strength of the London forces / dispersion forces / induced dipole forces increases. ✓
- More energy is needed to overcome the intermolecular forces. ✓

1.4 Compound A (CH_4) ✓ — Lowest boiling point. ✓ (2)(3)

1.5 HIGHER THAN ✓ (1)

1.6

- HCl has dipole-dipole forces between molecules. ✓
- CH_4 has London / dispersion / induced dipole forces between molecules. ✓
- Dipole-dipole forces are stronger than London forces, therefore more energy is required to overcome intermolecular forces in HCl. ✓

(3)

MEMORANDUM – QUESTION 2

2.1 The temperature at which the solid and liquid phases of a substance are in equilibrium. ✓✓ (2)

2.2

- Molecular size / mass increases from CH_4 to CCl_4 . ✓
- All three molecules have London / dispersion forces. ✓
- Strength of London forces increases with an increase in molecular size / mass. ✓
- More energy is required to overcome London forces in molecules with higher melting point. ✓

2.3 CH_4 ✓ — Has the lowest melting point / weakest intermolecular forces. ✓ (2)(4)

2.4

- CCl_4 has only London forces. ✓
- H_2O has hydrogen bonds (and London forces). Hydrogen bonds are stronger than London forces. ✓

2.5 Dipole-induced dipole forces. ✓✓ (2)(2)

MEMORANDUM – QUESTION 3

3.1 The temperature at which the vapour pressure of a liquid equals atmospheric pressure. ✓✓ (2)

3.2.1 As the molecular mass increases, the boiling point increases. ✓✓ (2)

3.2.2

- As molecular mass increases, the strength of the intermolecular (London) forces increases. ✓
- Therefore more energy is needed to overcome / weaken the intermolecular forces. ✓
- Thus the boiling point increases. ✓

3.3.1 H_2O / Water ✓ (1)(3)

3.3.2 Hydrogen bonds ✓✓ (2)

3.3.3 Hydrogen bonds are stronger than dipole-dipole forces ✓, therefore more energy is needed to overcome/weaken the forces. Thus boiling point is higher than expected. ✓ (2)

MEMORANDUM – QUESTION 4

4.1.1 The temperature at which the vapour pressure of a liquid equals atmospheric pressure. ✓✓ (2)

4.1.2 Gas ✓ (1)

4.1.3 Dipole-dipole forces ✓ (1)

4.1.4 Compound B (HBr) ✓ — Compound B has a larger molecular size / higher molecular mass than compound A ✓✓. (3)

4.1.5 Compound A (HCl) ✓ — Lower boiling point ($-85\text{ }^\circ\text{C}$) means weaker intermolecular forces and higher vapour pressure. ✓ (2)

4.1.6

- Compound C (HF) has hydrogen bonds between molecules. ✓
- HCl and HBr have dipole-dipole forces between molecules. ✓
- Hydrogen bonds are stronger than dipole-dipole forces, so more energy is required to overcome intermolecular forces in HF. ✓

4.2.1 CCl_4 ✓ (1)(3)

4.2.2

- CCl_4 and I_2 are both non-polar and have London forces only. ✓
- H_2O is polar and has hydrogen bonds. ✓
- Intermolecular forces between CCl_4 and I_2 molecules are of comparable magnitude ("like dissolves like"). ✓

(3)

MEMORANDUM – QUESTION 5

5.1 The temperature at which the vapour pressure of a liquid equals external (atmospheric) pressure. ✓✓ (2)

5.2 188 K (Accept range 180 K – 190 K) ✓ (1)

5.3

- Group 4 hydrogen hydrides have London / dispersion forces between molecules. ✓
- Hydrogen halides have dipole-dipole forces between molecules. ✓
- Dipole-dipole forces are stronger than London / dispersion forces. ✓
- More energy is required to overcome dipole-dipole forces in hydrogen halides. ✓

5.4 Hydrogen bonds ✓✓ (2)(4)

5.5 GeH_4 ✓ — It has a lower boiling point than HBr , which indicates weaker intermolecular forces and higher vapour pressure. ✓ (2)

MEMORANDUM – QUESTION 6

6.1 The average distance between the nuclei of two bonded atoms. ✓✓ (2)

6.2.1 120 pm ✓ (1)

6.2.2 $800 \text{ kJ}\cdot\text{mol}^{-1}$ (Accept range 780 – 800) ✓✓ (2)

6.2.3 Bond energy / Potential energy minimum ✓✓ (2)

6.3 LONGER THAN ✓ — The sulfur atom has a larger atomic radius than the oxygen atom ✓; therefore the bond length will be longer as electrostatic attraction between nuclei and electrons is less. ✓ (3)

6.4 At point D there is a balance of forces (electrostatic attractive forces equal repulsive forces / net force is zero). ✓✓

6.5.1 The shorter the bond length, the higher the bond energy. ✓✓ (2)(2)

6.5.2 Bond A (C–C) ✓ — It has the lowest bond energy ($348 \text{ kJ}\cdot\text{mol}^{-1}$) ✓, requiring the least energy to break the single covalent bond. ✓ (3)

MEMORANDUM – QUESTION 7

- 7.1 Temperature at which the vapour pressure of a liquid equals atmospheric pressure. ✓✓ (2)
- 7.2.1 NH_3 has hydrogen bonds between molecules, while CH_4 has London forces. ✓ Hydrogen bonds are stronger than London forces. ✓ More energy is required to overcome the stronger forces in NH_3 , resulting in a higher boiling point. ✓ (3)
- 7.2.2 SiH_4 has a larger molecular size/mass than CH_4 . ✓ The London forces in SiH_4 are stronger than in CH_4 . ✓ More energy is required to overcome the London forces in SiH_4 . ✓ (3)
- 7.3 Both NH_3 and H_2O are polar molecules with hydrogen bonds. Electronegativity difference between O and H ($3.5 - 2.1 = 1.4$) is higher than N and H ($3.0 - 2.1 = 0.9$). ✓ H_2O is angular while NH_3 is trigonal pyramidal. ✓ Water can form up to two hydrogen bonds per molecule compared to one on average in NH_3 . ✓ Intermolecular forces in H_2O are stronger and require more energy to overcome. ✓ (4)
- 7.4.1 Non-polar: CH_4 and SiH_4 ✓✓ (2)
- 7.4.2 Polar: NH_3 and H_2O ✓✓ (2)
- 7.5.1 A chemical bond that exists within a molecule between atoms. ✓✓ (2)
- 7.5.2 Polar covalent bond ✓✓ (2)
- 7.5.3 Force of attraction between molecules or between atoms of noble gases. ✓✓ (2)
- 7.5.4 Hydrogen bond ✓✓ (2)
- 7.5.5 High electronegativity ✓ and small atomic size / radius ✓ (with lone pairs of electrons). (2)
- 7.6 Partial positive charge ($\delta+$) ✓ (1)

MEMORANDUM – QUESTION 8

- 8.1 The temperature at which the vapour pressure of a substance equals atmospheric pressure. ✓✓ (2)
- 8.2
- NH_3 and H_2O have hydrogen bonds between molecules, while the rest have dipole-dipole forces. ✓
 - Hydrogen bonds are much stronger than dipole-dipole forces. ✓
 - Thus more energy will be needed to overcome hydrogen bonds than dipole-dipole forces. ✓
- 8.3.1 Dipole-dipole forces ✓ (1)(3)
- 8.3.2 Molecular mass increases down the group from H_2S to H_2Te ✓. As molecular mass increases, the strength of the intermolecular forces increases, requiring more energy to overcome, thus boiling point increases. ✓ (2)
- 8.3.3 H_2O (or H_2Te) ✓ (1)
- 8.4.1
- Cl_2 : London / dispersion forces ✓
 - SO_2 : Dipole-dipole forces ✓
 - NaBr : Ionic bonds / Electrostatic forces ✓
- 8.4.2 Cl_2 ✓ — Cl_2 has the weakest intermolecular forces (London forces) ✓, therefore it has the lowest boiling point (3) and highest vapour pressure. ✓ (3)

MEMORANDUM – QUESTION 9

9.1 The temperature at which the vapour pressure of a liquid equals atmospheric pressure. ✓✓ (2)

9.2

- Both iodine (I_2) and bromine (Br_2) have London forces between molecules. ✓
- Iodine molecules are larger / heavier / have a larger surface area than bromine molecules. ✓
- Thus iodine has stronger London forces between molecules. ✓
- More energy will be needed to overcome the forces between iodine molecules. ✓

9.3 (4)

- Intermolecular forces between NH_3 molecules are hydrogen bonds. ✓
- Intermolecular forces between PH_3 molecules are dipole-dipole forces. ✓
- Hydrogen bonds are much stronger than dipole-dipole forces. ✓
- More energy is needed to overcome hydrogen bonds in NH_3 , leading to a higher boiling point. ✓

9.4 The pressure exerted by a vapour at equilibrium with its liquid in a closed system. ✓✓ (2)(4)

9.5 Iodine (I_2) ✓ — It has the highest boiling point ($184.3^\circ C$) and strongest intermolecular forces ✓, resulting in the lowest vapour pressure and lowest rate of evaporation. ✓ (3)

9.6

- Iodine molecules are non-polar and have London forces between molecules. ✓
- Water molecules are polar and have hydrogen bonds between molecules. ✓
- Polar and non-polar substances do not mix / intermolecular forces are not of comparable magnitude. ✓

(3)

MEMORANDUM – QUESTION 10

10.1 The temperature at which the vapour pressure of a substance equals atmospheric pressure. ✓✓ (2)

10.2

- Ammonia has ONE site per molecule available for hydrogen bonding while water has TWO sites for hydrogen bonding. ✓
- Therefore, the total forces of attraction between water molecules are stronger than between ammonia molecules. ✓
- Thus more energy is needed to overcome intermolecular forces in water. ✓
- Therefore water has a significantly higher boiling point. ✓

10.3 Water (H_2O) ✓ — It has a higher boiling point and stronger intermolecular forces, resulting in lower vapour pressure. ✓ (4)
(2)