Week 5

A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

Contents 目录

A-Level Chemistry Name: _____ Score: _____

1. The power of an atom to attract the electrons in a covalent bond is its _____.

   _____

2. Ionic bonding is the electrostatic attraction between oppositely charged ions.

   ☐ True ☐ False

3. Metallic bonding is the attraction between:

   ☐ positive metal ions and a sea of delocalised electrons

   ☐ two oppositely charged ions

   ☐ a shared pair of electrons

   ☐ two molecules

4. A covalent bond is:

   ☐ a shared pair of electrons attracted to both nuclei

   ☐ a transfer of electrons

   ☐ a sea of delocalised electrons

   ☐ a force between molecules

5. A pi (π) bond forms by:

   ☐ the sideways overlap of two p orbitals, above and below the axis

   ☐ head-on overlap between the atoms

   ☐ electron transfer

   ☐ a sea of delocalised electrons

6. VSEPR theory says the shape of a molecule is set by:

   ☐ electron pairs around the central atom repelling and spreading as far apart as possible

   ☐ the mass of the atoms

   ☐ the number of neutrons

   ☐ the colour of the molecule

7. Compared with the bonds inside substances, intermolecular forces are:

   ☐ much weaker

   ☐ much stronger

☐ exactly the same

☐ never present

8. Metals conduct electricity because of their _____ (delocalised) electrons.

_____

9. In an ionic solid the ions are packed into a regular giant _____, held by attraction in every direction.

_____

10. Metallic bonding helps explain why metals are strong.

☐ True ☐ False

3. Chemical bonding

Answers · 答案

A-Level Chemistry

1. **electronegativity**

   It increases across a period and up a group (fluorine is the highest).

2. **True**

   It forms by transferring electrons from a metal to a non-metal.

3. **positive metal ions and a sea of delocalised electrons**

   In a metal, positive ions sit in a sea of delocalised electrons that hold the structure together.

4. **a shared pair of electrons attracted to both nuclei**

   In a covalent bond two atoms share a pair of electrons, attracted to both nuclei.

5. **the sideways overlap of two p orbitals, above and below the axis**

   A σ bond is head-on overlap; a π bond is the sideways overlap of p orbitals.

6. **electron pairs around the central atom repelling and spreading as far apart as possible**

   Like charges repel, so the electron pairs push apart to minimise repulsion, fixing the shape.

7. **much weaker**

   Intermolecular forces (between molecules) are far weaker than ionic, covalent or metallic bonds within substances.

8. **free**

   A sea of delocalised electrons carries charge.

9. **lattice**

   Each ion is surrounded by oppositely charged neighbours, so the structure is very strong.

10. **True**

    The strong attraction between the positive ions and the electron sea makes metals strong.

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.1 Electronegativity and bonding

Name: _____ Class: _____ Date: _____ **Total: 9 marks**

KEY WORDS electronegativity 电负性 • electron 电子 • ion 离子 • shielding 屏蔽 • atomic radius 原子半径

Objective

Build the skills to answer exam questions on **electronegativity and bonding** — defining **electronegativity** 电负性, its trends, and using **Pauling** 鲍林 values to predict bond type.

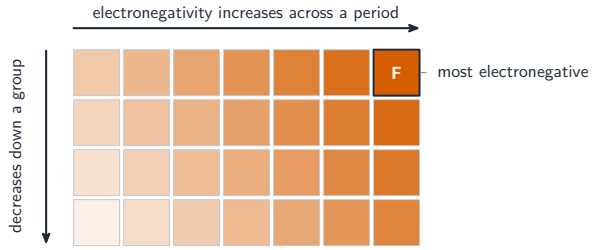
You must be able to:

- define electronegativity and explain the factors that affect it
- state and explain the trends across a period and down a group
- use electronegativity differences to predict ionic or covalent bonding

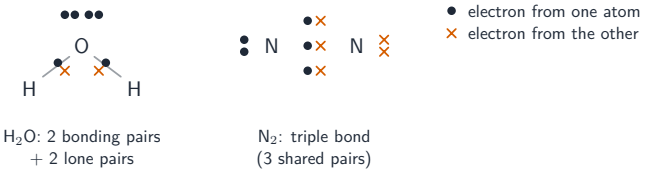
1 Worked examples

■ Electronegativity and its trends

Electronegativity is the power of an atom to attract the **electrons** in a bond towards itself. It depends on **nuclear charge** (more → higher), **atomic radius** (smaller → higher) and **shielding** (more → lower).



Electronegativity rises across a period and falls down a group — fluorine is the most electronegative



H_2O: 2 bonding pairs + 2 lone pairs

N_2: triple bond (3 shared pairs)

Electronegativity sets where a bond sits on the ionic-covalent scale

■ Predicting bond type

A **large** difference in Pauling values → **ionic**; a **small** difference → **covalent**.

2 Practice

2.1 Define **electronegativity**. [1]

2.2 State the most electronegative element. [1]

3 Exam-style questions

3.1 Electronegativity increases across Period 3 because: [1]

- **A** the atomic radius increases
- **B** the nuclear charge increases while shielding stays similar
- **C** more shells are added
- **D** the atoms gain electrons

3.2 Two elements have a **large** difference in electronegativity. The bond between them is likely to be: [1]

- **A** metallic
- **B** ionic
- **C** pure covalent
- **D** a hydrogen bond

3.3 Explain why electronegativity decreases down Group 17. [3]

3.4 Using electronegativity, explain why the bond in HCl is polar. [2]

Solutions

2.1 the power of an atom to attract the electrons in a bond towards itself.

2.2 fluorine.

3.1 B. Rising nuclear charge with similar shielding pulls bonding electrons more strongly.

3.2 B. A large electronegativity difference gives an ionic bond.

3.3 down the group each element has more shells, so a larger atomic radius and more shielding; the nucleus attracts the bonding electrons less strongly, so electronegativity falls.

3.4 chlorine is more electronegative than hydrogen, so it pulls the bonding electrons closer, giving Cl a $\delta-$ and H a $\delta+$ — a polar bond.

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.2 Ionic bonding

Name: _____ Class: _____ Date: _____ **Total: 10 marks**

KEY WORDS ionic bonding 离子键 • anion 阴离子 • cation 阳离子 • sodium chloride 氯化钠 • electrostatic attraction 静电引力

Objective

Build the skills to answer exam questions on **ionic bonding** — defining it and describing it in **sodium chloride** 氯化钠, **magnesium oxide** 氧化镁 and **calcium fluoride** 氟化钙.

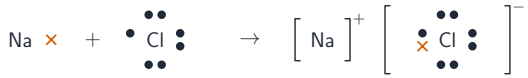
You must be able to:

- define ionic bonding as the electrostatic attraction between oppositely charged ions
- describe ionic bonding using NaCl, MgO and CaF_2

1 Worked examples

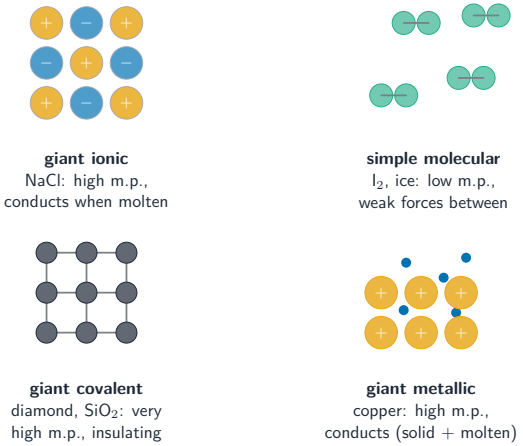
■ Ionic bonding

Ionic bonding is the **electrostatic attraction** 静电引力 between oppositely charged **ions** — positive **cations** 阳离子 and negative **anions** 阴离子. It forms when a metal gives electrons to a non-metal; the ions pack into a giant **lattice** 晶格.



sodium gives its outer electron (×) to chlorine

NaCl: sodium transfers its outer electron to chlorine, giving Na^+ and a full-octet Cl^-



Ionic bonding builds a giant lattice

In **MgO**, Mg gives two electrons to O → Mg^{2+} and O^{2-}. In **CaF_2**, Ca gives one electron to each of two F atoms → Ca^{2+} and two F^-.

2 Practice

2.1 Define **ionic bonding**. [2]

2.2 State the charges on the ions in magnesium oxide. [1]

3 Exam-style questions

3.1 Ionic bonding is best described as: [1]

- **A** a shared pair of electrons
- **B** the electrostatic attraction between oppositely charged ions
- **C** positive ions in a sea of electrons
- **D** weak forces between molecules

3.2 Which compound contains ions with charges 2+ and 2−? [1]

- **A** NaCl
- **B** MgO
- **C** CaF_2
- **D** HCl

3.3 Describe, in terms of electron transfer, how the ionic bond forms in calcium fluoride, CaF_2. [3]

3.4 Explain why magnesium oxide has a higher melting point than sodium chloride. [2]

Solutions

2.1 the electrostatic attraction between oppositely charged ions (cations and anions).

2.2 Mg^{2+} and O^{2-}.

3.1 B. Ionic bonding is the electrostatic attraction between oppositely charged ions.

3.2 B. MgO has Mg^{2+} and O^{2-}.

3.3 each calcium atom loses 2 electrons to form Ca^{2+}; each of two fluorine atoms gains 1 electron to form F^-; the oppositely charged ions attract in a lattice.

3.4 MgO has ions with higher charges (2+/2− vs 1+/1−), so a stronger electrostatic attraction needing more energy to break.

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.3 Metallic bonding

Name: _____ Class: _____ Date: _____ **Total: 8 marks**

KEY WORDS delocalised electrons 离域电子 • metallic bonding 金属键 • electron 电子 • ion 离子 • electrostatic attraction 静电引力

Objective

Build the skills to answer exam questions on **metallic bonding** — defining it and using it to explain the properties of metals.

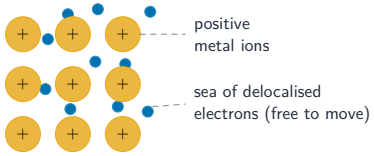
You must be able to:

- define metallic bonding as the attraction between positive metal ions and delocalised electrons
- use it to explain electrical conductivity and strength

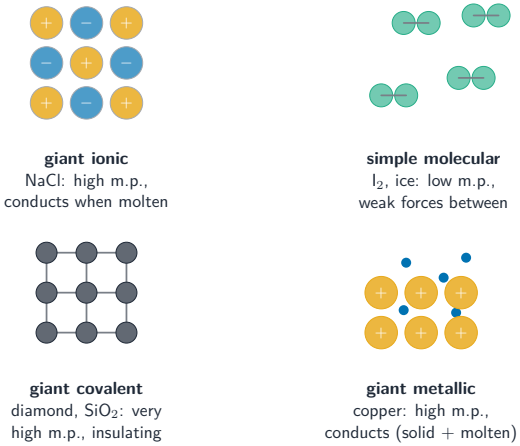
1 Worked examples

■ Metallic bonding

Metallic bonding 金属键 is the **electrostatic attraction** between positive metal ions and a "sea" of **delocalised electrons** 离域电子.



Positive metal ions sit in a sea of delocalised electrons that are free to move



The metallic lattice among the four structures

The free electrons explain why metals **conduct electricity** (the electrons move and carry charge) and are **strong** (the attraction acts throughout the lattice).

2 Practice

2.1 Define **metallic bonding**. [2]

2.2 Name the particles that carry charge when a metal conducts electricity. [1]

3 Exam-style questions

3.1 Metals conduct electricity because they contain: [1]

- **A** oppositely charged ions
- **B** delocalised electrons that are free to move
- **C** shared pairs of electrons
- **D** weak intermolecular forces

3.2 Explain, in terms of metallic bonding, why metals are good conductors of electricity. [2]

3.3 Magnesium has a higher melting point than sodium. Suggest why, in terms of metallic bonding. [2]

Solutions

2.1 the electrostatic attraction between positive metal ions and delocalised (sea of) electrons.

2.2 delocalised electrons.

3.1 B. Delocalised electrons are free to move and carry charge.

3.2 the delocalised electrons are free to move through the metal; they carry charge when a voltage is applied.

3.3 magnesium ions have a 2+ charge (vs 1+ for sodium) and contribute more delocalised electrons; the stronger attraction needs more energy to break, so a higher melting point.

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.4 Covalent bonding and coordinate (dative covalent) bonding

Name: _____ Class: _____ Date: _____ **Total: 11 marks**

KEY WORDS coordinate bond 配位键 • covalent bonds 共价键 • bond energy 键能 • hybridisation 杂化 • bond length 键长

Objective

Build the skills to answer exam questions on **covalent and coordinate bonding** — **covalent bonds** 共价键, **coordinate (dative)** 配位 bonds, **σ and π bonds** σ 和 π 键, **hybridisation** 杂化, and bond energy/length.

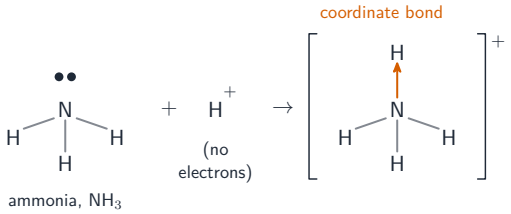
You must be able to:

- define covalent and coordinate bonding and describe an expanded octet
- describe σ and π bonds (single, double, triple) and sp/sp^2/sp^3 hybridisation
- define bond energy and bond length and relate them to reactivity

1 Worked examples

Covalent and coordinate bonds

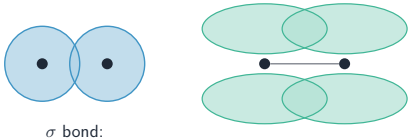
Covalent bonding is the electrostatic attraction between the nuclei of two atoms and a **shared pair** of electrons. Period 3 atoms can **expand the octet** (e.g. SO_2, PCl_5, SF_6). A **coordinate (dative)** bond is a covalent bond where **both** electrons come from one atom.



ammonia, NH_3

A coordinate bond: N's lone pair forms the fourth N–H bond, both electrons coming from $N \rightarrow NH_4^+$

σ and π bonds



σ = head-on overlap; π = sideways p-orbital overlap. Single = 1σ; double = 1σ+1π; triple = 1σ+2π

Hybridisation mixes orbitals: sp (linear), sp^2 (e.g. C_2H_4, planar), sp^3 (e.g. CH_4, tetrahedral). **Bond energy** = energy to break one mole of a bond (gas); a shorter bond is usually stronger (triple > double > single).

2 Practice

2.1 Define a **coordinate (dative covalent) bond**. [1]

2.2 State how many σ and π bonds are in a nitrogen molecule, N_2 (triple bond). [2]

3 Exam-style questions

3.1 A double bond consists of: [1]

- **A** two σ bonds
- **B** one σ bond and one π bond
- **C** two π bonds
- **D** one σ bond and two π bonds

3.2 Which molecule contains an atom with an expanded octet? [1]

- **A** CH_4
- **B** NH_3
- **C** SF_6
- **D** H_2O

3.3 (a) Describe how a coordinate bond forms when ammonia reacts with H^+ to form NH_4^+. [2]

(b) Explain why a C≡C triple bond is shorter and stronger than a C=C double bond. [2]

3.4 Explain, in terms of σ and π bonds, the bonding in an ethene molecule, C_2H_4. [2]

Solutions

2.1 a covalent bond in which both shared electrons come from the same atom.

2.2 1 σ bond and 2 π bonds.

3.1 B. A double bond is one σ plus one π bond.

3.2 C. SF_6 has six bonds around sulfur — an expanded octet.

3.3 (a) the lone pair on the nitrogen atom is donated to the H^+ ion; both electrons come from N, forming the fourth N–H (coordinate) bond.
(b) a triple bond has more shared electron pairs, so a greater attraction pulling the nuclei closer; the shorter bond is harder to break — higher bond energy.

3.4 each C–H and the C–C σ bond form by direct (head-on) overlap; the C=C also has a π bond from sideways overlap of p orbitals above and below.

6 NH_3 and HCN react together to form NH_4CN, an ionic compound.

Which row states the number of coordinate bonds and the number of π bonds in one formula unit of NH_4CN?

	number of coordinate bonds	number of π bonds
A	0	2
B	1	2
C	0	3
D	1	3

24 When dry ammonia and hydrogen chloride gases are mixed, a solid white ionic compound is formed.

Two statements are listed.

1 The formation of the ionic compound is a redox reaction.

2 During the formation of the ionic compound, the H–N–H bond angle increases.

Which statements are correct?

A both 1 and 2

B 1 only

C 2 only

D neither 1 nor 2

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.5 Shapes of molecules

Name: _____ Class: _____ Date: _____ **Total: 11 marks**

KEY WORDS vsepr theory 价层电子对互斥理论 • bond angle 键角 • tetrahedral 四面体形 • trigonal planar 平面三角形 • lone pair 孤对电子

Objective

Build the skills to answer exam questions on **shapes of molecules** — using **VSEPR theory** 价层电子对互斥理论 to predict shapes and **bond angles** 键角.

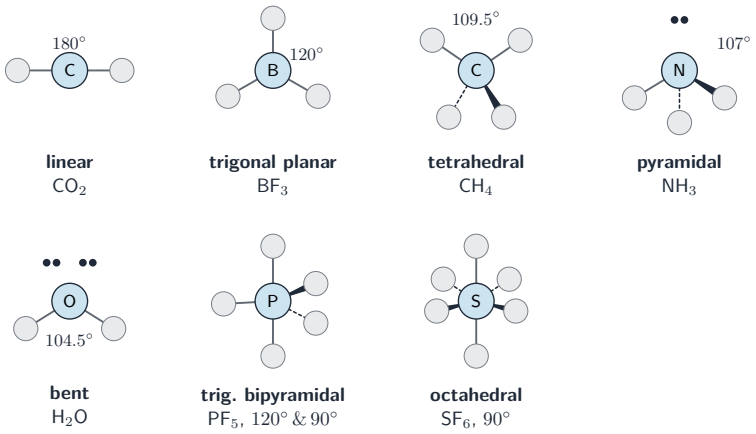
You must be able to:

- state and explain the shapes and bond angles of the standard molecules
- predict the shapes of analogous molecules and ions
- explain the effect of lone pairs on bond angle

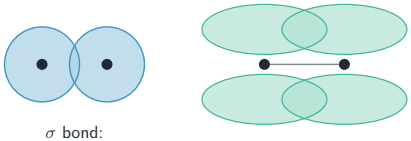
1 Worked examples

VSEPR theory

Electron pairs around the central atom repel and push as far apart as possible. A **lone pair** 孤对电子 repels more than a **bonding pair**, squeezing the bond angle by about 2.5° each.



The seven shapes; lone pairs on NH_3 and H_2O squeeze the angle below 109.5°



Bonding pairs set the shape; sigma and pi bonds

molecule	shape	angle
CO_2	linear	180°
BF_3	trigonal planar	120°
CH_4	tetrahedral	109.5°
NH_3	pyramidal	107°
H_2O	bent	104.5°
PF_5	trigonal bipyramidal	120° & 90°
SF_6	octahedral	90°

2 Practice

2.1 State the shape and bond angle of a methane molecule, CH_4. [2]

2.2 State the shape of an ammonia molecule, NH_3. [1]

3 Exam-style questions

3.1 What is the bond angle in a carbon dioxide molecule, CO_2? [1]

- **A** 104.5°
- **B** 109.5°
- **C** 120°
- **D** 180°

3.2 Why is the bond angle in water (104.5°) smaller than in methane (109.5°)? [1]

- **A** water has more bonds
- **B** the two lone pairs on oxygen repel more, squeezing the angle
- **C** oxygen is bigger than carbon
- **D** water has a double bond

3.3 (a) Using VSEPR theory, explain the shape of a BF_3 molecule. [2]

(b) Predict the shape and bond angle of an NH_4^+ ion. [2]

3.4 Explain why ammonia (107°) has a smaller bond angle than methane (109.5°). [2]

Solutions

2.1 tetrahedral; 109.5°.

2.2 pyramidal (trigonal pyramidal).

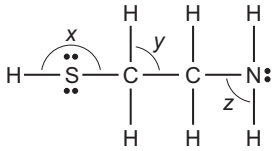
3.1 D. CO_2 is linear, 180°.

3.2 B. Oxygen's two lone pairs repel more, squeezing the angle.

3.3 (a) boron has 3 bonding pairs and no lone pairs; they repel equally to a trigonal planar shape, 120°.
(b) tetrahedral; 109.5° (4 bonding pairs, no lone pairs).

3.4 ammonia has one lone pair on nitrogen; the lone pair repels the bonding pairs more strongly, squeezing the angle from 109.5° to 107°.

6 Three bond angles are labelled on the molecule shown.



: represents a lone pair of electrons

What is the order of **decreasing** size of the bond angles *x*, *y* and *z*?

	largest →		smallest
A	*x*	*y*	*z*
B	*x*	*z*	*y*
C	*y*	*z*	*x*
D	*z*	*y*	*x*

5 In which set do all the molecules have all their atoms arranged in one plane?

A $AlCl_3$, BF_3, PH_3

B $AlCl_3$, CO_2, NH_3

C BF_3, C_2H_4, C_3H_6

D C_2H_4, CO_2, H_2O

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.6 Intermolecular forces, electronegativity and bond properties

Name: _____ Class: _____ Date: _____ **Total: 31 marks**

KEY WORDS dipole 偶极 • intermolecular forces 分子间作用力 • hydrogen bonding 氢键 • permanent dipole 永久偶极 • induced dipole 诱导偶极

Objective

Build the skills to answer exam questions on **intermolecular forces** — **bond polarity** 键极性 and dipoles, **van der Waals' forces** 范德华力, **hydrogen bonding** 氢键, and the anomalous properties of water.

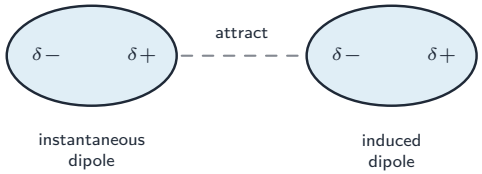
You must be able to:

- explain bond polarity and dipole moments using electronegativity
- describe London (id–id) and permanent dipole (pd–pd) forces, and hydrogen bonding
- use hydrogen bonding to explain the anomalous properties of water

1 Worked examples

Polarity and van der Waals' forces

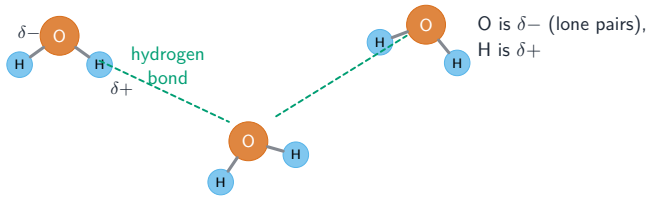
A bond between atoms of different electronegativity has a **dipole** 偶极 ($\delta-$ / $\delta+$). If the dipoles do not cancel, the molecule is **polar**. **Van der Waals' forces** is the general term for all intermolecular forces.



London (id–id) force: a momentary dipole induces one in a neighbour — acts between all molecules, stronger with more electrons

- **London (id–id)** force — acts between all molecules; stronger when there are more electrons.
- **permanent dipole (pd–pd)** force — between always-polar molecules.

Hydrogen bonding and water



Hydrogen bonding: a $\delta+$ H (bonded to N, O or F) is attracted to a lone pair on a $\delta-$ N, O or F

Hydrogen bonding (a strong pd–pd case) explains water's anomalies: high melting/boiling point (many H-bonds to break); high surface tension; and **ice is less dense** than water (H-bonds hold an open structure), so ice floats.

The three forces, in the order you should check for them

Every molecule has **van der Waals' forces** (induced dipole–induced dipole). A polar molecule has those **plus** permanent dipole–dipole. A molecule with H bonded to N, O or F has those **plus** hydrogen bonding. So the forces stack; the question is which one dominates.

Force	Between	Strength
van der Waals' 范德华力	all molecules; arise from instantaneous, fluctuating dipoles inducing dipoles in neighbours	weakest, but grows with the number of electrons and with surface contact
permanent dipole–dipole 永久偶极	polar molecules, from a difference in electronegativity across a bond	intermediate
hydrogen bonding 氢键	a lone pair on N, O or F and an H attached to N, O or F	strongest of the three

A polar bond does not make a polar molecule. CO_2 has two strongly polar bonds, but they are opposite and equal in a linear molecule, so the dipoles cancel and the molecule is non-polar. **Check the shape before deciding.**

Reading a boiling-point trend — the two-part answer the mark scheme wants:

1. Name the force being broken.
2. Explain why it is stronger or weaker here.

So along the alkanes: van der Waals' forces increase with more electrons and greater surface contact, so boiling point rises. A branched isomer boils lower than its straight-chain isomer because branching reduces the surface contact, weakening the van der

Waals' forces — **the covalent bonds are never broken on boiling**, and saying so loses the mark.

Worked example. Explain why H_2O boils at 100 °C but H_2S at −60 °C, even though H_2S has more electrons.

1. Both are polar bent molecules with dipole–dipole forces and van der Waals' forces.
2. In water, H is bonded to the highly electronegative, small oxygen atom, which also has lone pairs, so **hydrogen bonding** occurs; sulfur is not electronegative enough for it.
3. Hydrogen bonds are much stronger than the dipole–dipole and van der Waals' forces in H_2S, so far more energy is needed to separate water molecules — outweighing the larger number of electrons in H_2S.

Ice is less dense than water for the same reason: hydrogen bonds hold the molecules in an **open tetrahedral lattice** with gaps, so the same number of molecules occupies more volume.

2 Practice

2.1 Name the **two** atoms (besides hydrogen) that must be present for hydrogen bonding in water and ammonia. [2]

2.2 State which intermolecular force acts between all molecules. [1]

2.3 Explain why CO_2 is non-polar despite having polar bonds. [2]

2.4 State the three conditions needed for hydrogen bonding. [2]

2.5 Explain why 2-methylpropane boils at a lower temperature than butane. [2]

3 Exam-style questions

3.1 Why does ice float on water? [1]

- **A** ice is warmer
- **B** hydrogen bonds hold ice in an open structure, so it is less dense
- **C** ice has no hydrogen bonds
- **D** ice contains air

3.2 Which has the **strongest** intermolecular forces? [1]

- **A** CH_4
- **B** Cl_2
- **C** H_2O
- **D** O_2

3.3 (a) Explain why water has a relatively high boiling point. [2]

(b) Explain why the boiling points of the noble gases increase down the group. [2]

3.4 Draw and label a diagram to show hydrogen bonding between two water molecules (show the δ charges and the lone pair). [3]

3.5 The table gives the boiling points of four hydrides.

Hydride	**NH_3**	**PH_3**	**H_2O**	**H_2S**
boiling point / °C	−33	−88	100	−60

(a) Name all the intermolecular forces present in PH_3. [2]

(b) Explain why NH_3 boils at a much higher temperature than PH_3, even though PH_3 has more electrons. [4]

(c) Explain why H_2O boils higher than NH_3, given that both form hydrogen bonds. [2]

(d) State what is broken when H_2O boils, and what is NOT broken. [2]

(e) Explain, in terms of hydrogen bonding, why ice floats on water. [3]

Solutions

2.1 nitrogen, oxygen (and fluorine).

2.2 the London dispersion / instantaneous dipole–induced dipole force.

2.3 The two C = O dipoles are equal in size and point in opposite directions; the molecule is linear, so they cancel and there is no overall dipole.

2.4 A hydrogen atom bonded to N, O or F; and a lone pair on an N, O or F atom of another molecule.

2.5 2-methylpropane is branched, so its molecules have less surface contact with each other; the van der Waals' forces between them are therefore weaker and less energy is needed to separate them.

3.1 B. Hydrogen bonds hold ice in an open lattice, making it less dense than water.

3.2 C. Water has hydrogen bonding, the strongest of the four.

3.3 (a) water molecules are held by hydrogen bonds; much energy is needed to break these, so the boiling point is high.
(b) down the group atoms have more electrons, so stronger London forces, needing more energy to overcome — higher boiling points.

3.4 a dashed hydrogen bond from a $\delta+$ H on one molecule to a lone pair on the $\delta-$ O of the other; both molecules drawn with $\delta+$ on H and $\delta-$ on O; a clear lone pair shown on the accepting oxygen.

3.5 (a) van der Waals' forces; permanent dipole–dipole forces.
(b) NH_3 has hydrogen bonding, because H is bonded to the small, highly electronegative nitrogen, which also carries a lone pair; PH_3 has only van der Waals' and dipole–dipole forces, since phosphorus is not electronegative enough; hydrogen bonds are much stronger, so far more energy is needed to separate ammonia molecules, outweighing PH_3's greater number of electrons.
(c) Oxygen is more electronegative than nitrogen and each water molecule has two hydrogens and two lone pairs, so water forms more hydrogen bonds per molecule, and stronger ones.
(d) The hydrogen bonds and other intermolecular forces between the molecules are broken; the covalent O–H bonds within the molecules are NOT broken.
(e) In ice each molecule is held by hydrogen bonds to four others in a rigid tetrahedral arrangement; this open lattice contains gaps, so the same number of molecules occupies a larger volume; the density is therefore lower than that of liquid water and ice floats.

7 The boiling point of water, H_2O, is higher than that of hydrogen sulfide, H_2S.

Which statement explains this difference in boiling points?

A The bond energy of O–H is greater than the bond energy of S–H.

B The intermolecular forces in H_2O are weaker than the intermolecular forces in H_2S.

C The S–H bond in H_2S is longer than the O–H bond in H_2O.

D There is significant intermolecular hydrogen bonding in H_2O but **not** in H_2S.

3 The alkanes are a homologous series of organic molecules. Alkanes are generally unreactive and are commonly used as fuels.

(a) Define homologous series.

.....

..... [2]

(b) Give two reasons to explain the general unreactivity of alkanes.

1 .....

2 .....

[2]

(c) Alkanes with low relative molecular mass, M_r, are more useful than those found in heavier crude oil fractions.

Name the process that is used to obtain alkanes with low M_r from heavier crude oil fractions.

..... [1]

(d) Hexane, C_6H_{14}, has four structural isomers.

Fig. 3.1 shows hexane and two of its structural isomers.

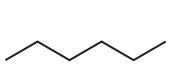
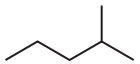
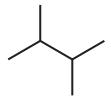
hexane	A	B	C	D
				

Fig. 3.1

(i) Complete Fig. 3.1 by drawing structures for **C** and **D**, the other two structural isomers of hexane. [2]

(ii) **A**, **B** and hexane have different boiling points.

Arrange **A**, **B** and hexane in order of increasing boiling point.

Explain your answer.

lowest ..... < ..... < ..... highest

.....

.....

....

[3]

(e) Hexane can be converted into compounds **E** and **F** at high temperature and pressure. Fig. 3.2 shows the reaction scheme involving hexane, **E** and **F**.

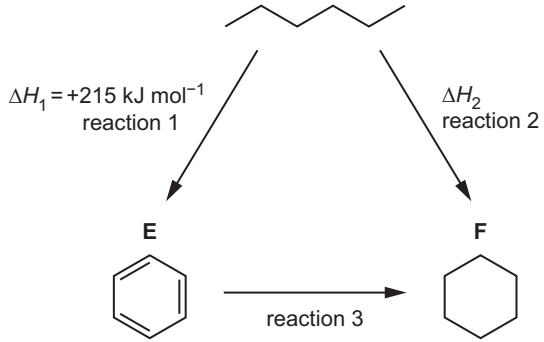


Fig. 3.2

(i) Identify a suitable reagent for reaction 3.

..... [1]

(ii) Use the data in Fig. 3.2 and in Table 3.1 to calculate the enthalpy change of reaction 2, ΔH_2.

Table 3.1

compound	enthalpy change of formation, ΔH_f / kJ mol^{-1}
	−156
	+48

ΔH_2 = ... kJ mol^{-1} [2]

(iii) Write an equation for the complete combustion of hexane.

..... [1]

[Total: 14]

A-LEVEL CHEMISTRY • EXERCISE SHEET

3.7 Dot-and-cross diagrams

Name: _____ Class: _____ Date: _____ **Total: 9 marks**

KEY WORDS dot-and-cross diagrams 点叉图 • electron 电子 • ion 离子 • lone pair 孤对电子 • coordinate bond 配位键

Objective

Build the skills to answer exam questions on **dot-and-cross diagrams** 点叉图 — drawing them for ionic, covalent and coordinate bonding.

You must be able to:

- draw dot-and-cross diagrams for ionic, covalent and coordinate bonding
- include molecules with an expanded octet or an odd number of electrons

1 Worked examples

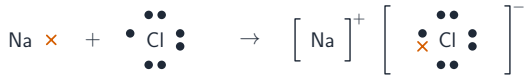
■ Drawing dot-and-cross diagrams

Use **dots** for one atom's outer electrons and **crosses** for the other, so you can see where each bonding electron came from. Show **lone pairs** too. For ionic bonds, draw the ions in brackets with their charge.

H_2O: 2 bonding pairs + 2 lone pairs

N_2: triple bond (3 shared pairs)

Water: two bonding pairs + two lone pairs on O. Nitrogen: three shared pairs (a triple bond), one lone pair each



sodium gives its outer electron (×) to chlorine

Ionic NaCl: the transferred electron is shown, and each ion is drawn in brackets with its charge

2 Practice

2.1 State what dots and crosses represent in a dot-and-cross diagram. [1]

2.2 State how many lone pairs are on the oxygen atom in a water molecule. [1]

3 Exam-style questions

3.1 In a dot-and-cross diagram of methane, CH_4, how many shared (bonding) pairs are shown around the carbon? [1]

- **A** 1
- **B** 2
- **C** 4
- **D** 8

3.2 Draw a dot-and-cross diagram for a molecule of ammonia, NH_3 (show the bonding pairs and the lone pair on nitrogen). [2]

3.3 Draw a dot-and-cross diagram for carbon dioxide, CO_2 (show the two double bonds). [2]

3.4 Draw a dot-and-cross diagram for the ammonium ion, NH_4^+, showing the coordinate bond. [2]

Solutions

2.1 the outer (valence) electrons of each of the two atoms.

2.2 2 lone pairs.

3.1 C. Methane has four shared pairs (four C–H bonds).

3.2 three N–H bonding pairs, each one dot + one cross; one lone pair shown on the nitrogen.

3.3 carbon in the centre with two double bonds to the oxygens (two shared pairs each); two lone pairs shown on each oxygen.

3.4 four N–H bonds around nitrogen; one shown as a coordinate bond (both electrons from N), with the whole ion in brackets and a + charge.