

Week 5

A-Level Chemistry

Homework bundle for this week

本周作业合集

exercises.lol

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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

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.

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.

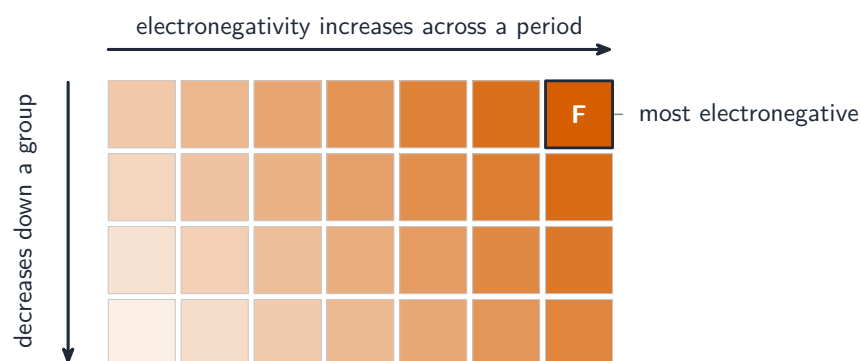
You must be able to:

- 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

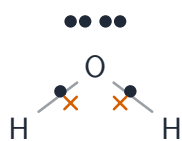
1 Worked examples

■ 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



H₂O: 2 bonding pairs
+ 2 lone pairs



N₂: triple bond
(3 shared pairs)

● electron from one atom
× electron from the other

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.

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 • electrostatic attraction 静电引力
• magnesium oxide 氧化镁

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 CaF₂

1 Worked examples

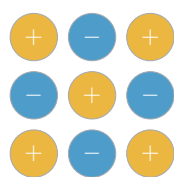
■ Ionic bonding

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.

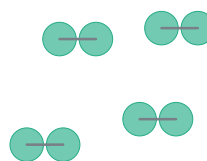


sodium gives its outer electron (×) to chlorine

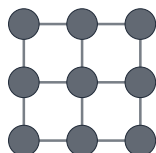
NaCl: sodium transfers its outer electron to chlorine, giving Na⁺ and a full-octet Cl⁻



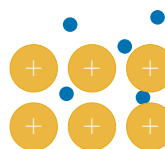
giant ionic
NaCl: high m.p.,
conducts when molten



simple molecular
I₂, ice: low m.p.,
weak forces between



giant covalent
diamond, SiO₂: very
high m.p., insulating



giant metallic
copper: high m.p.,
conducts (solid + molten)

Ionic bonding builds a giant lattice

In **MgO**, Mg gives two electrons to O → Mg²⁺ and O²⁻. In **CaF₂**, Ca gives one electron to each of two F atoms → Ca²⁺ 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.

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.

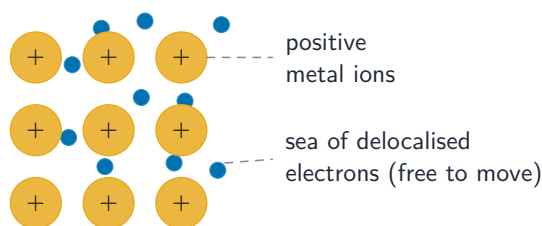
You must be able to:

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

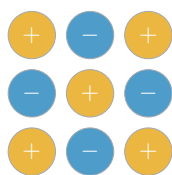
1 Worked examples

■ Metallic bonding

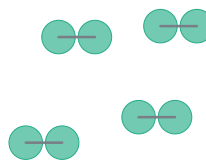
Metallic bonding 金属键 is the **electrostatic attraction** between positive metal ions and a “sea” of **delocalised electrons** 离域电子.



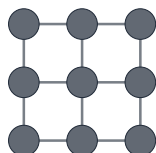
Positive metal ions sit in a sea of delocalised electrons that are free to move



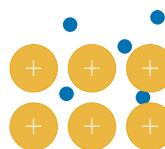
giant ionic
NaCl: high m.p.,
conducts when molten



simple molecular
I₂, ice: low m.p.,
weak forces between



giant covalent
diamond, SiO₂: very
high m.p., insulating



giant metallic
copper: high m.p.,
conducts (solid + molten)

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.

3.4 Covalent bonding and coordinate (dative covalent) bonding

Name: _____ Class: _____ Date: _____

Total: 11 marks

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

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.

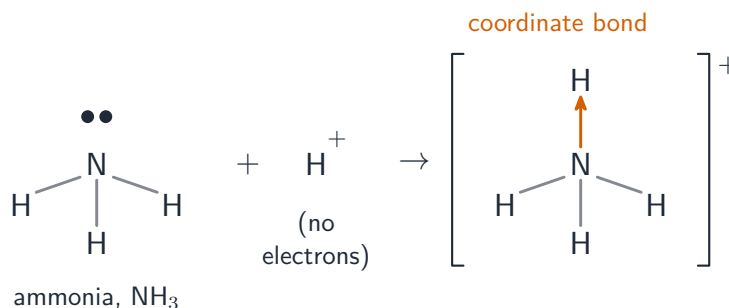
You must be able to:

- 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

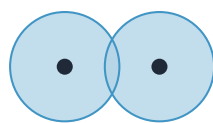
■ Covalent and coordinate bonds

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.

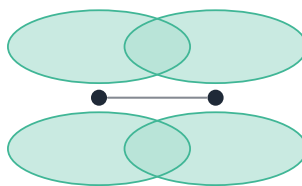


A coordinate bond: N's lone pair forms the fourth N-H bond, both electrons coming from N $\rightarrow NH_4^+$

■ σ and π bonds



σ bond:
head-on overlap



π bond: sideways overlap
of p orbitals

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

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 $\text{C}\equiv\text{C}$ triple bond is shorter and stronger than a $\text{C}=\text{C}$ 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 $\text{H}-\text{N}-\text{H}$ bond angle increases.

Which statements are correct?

- A** both 1 and 2
B 1 only
C 2 only
D neither 1 nor 2

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 • lone pair 孤对电子 • bonding pair 成键电子对
• pyramidal 三角锥形

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.

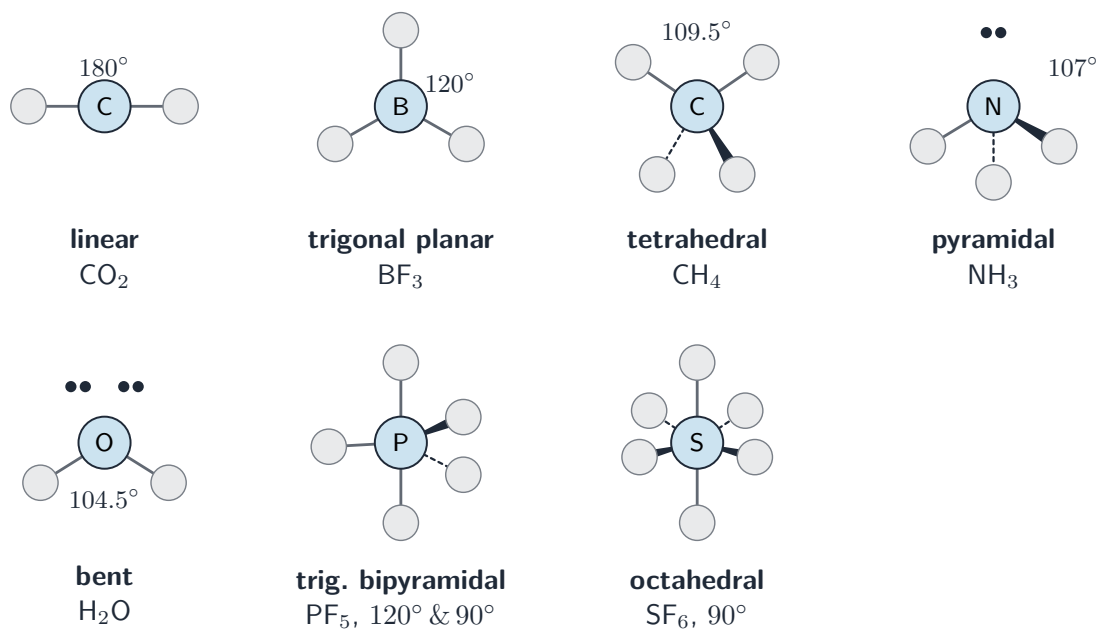
You must be able to:

- 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

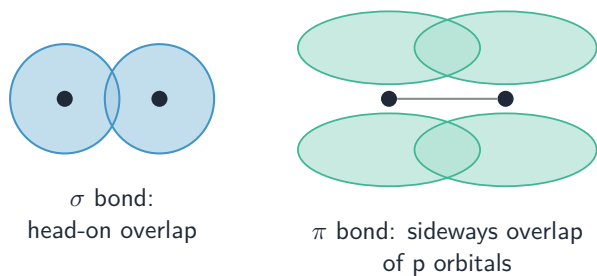
1 Worked examples

■ 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 NH₃ and H₂O 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 ₂ O	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° .

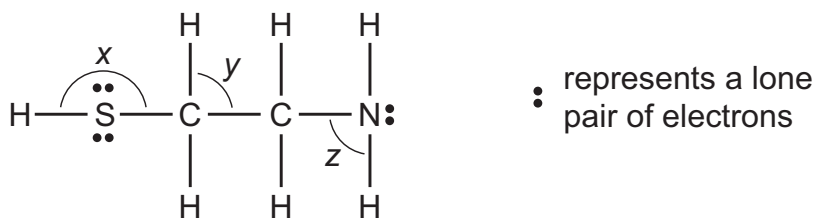
3.2 B. Oxygen's two lone pairs repel more, squeezing the angle.

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.



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

3.6 Intermolecular forces, electronegativity and bond properties

Name: _____ Class: _____ Date: _____

Total: 12 marks

KEY WORDS hydrogen bonding 氢键 • dipole 偶极 • van der waals' forces 范德华力 • intermolecular forces 分子间力

• polarity 极性

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.

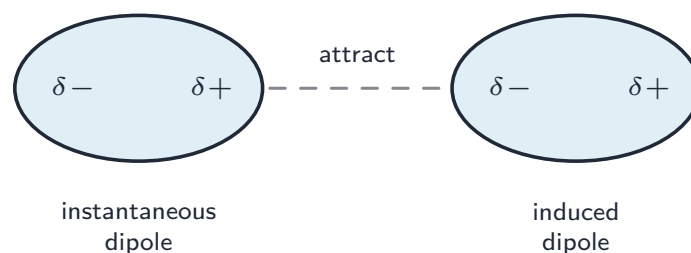
You must be able to:

- 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.

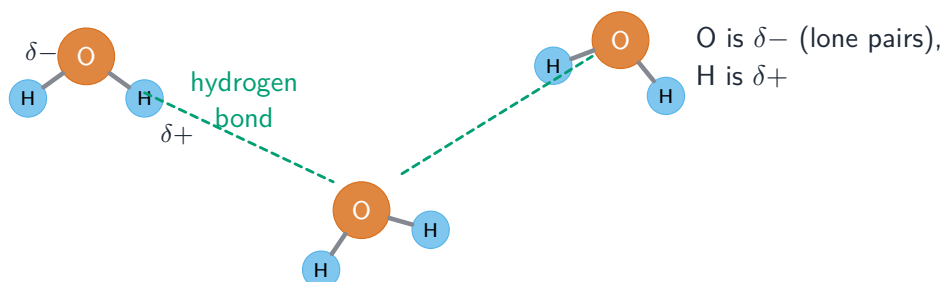


London (id-id) force: a momentary dipole induces one in a neighbour — acts between all molecules, stronger with more electrons

- **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.

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]

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]

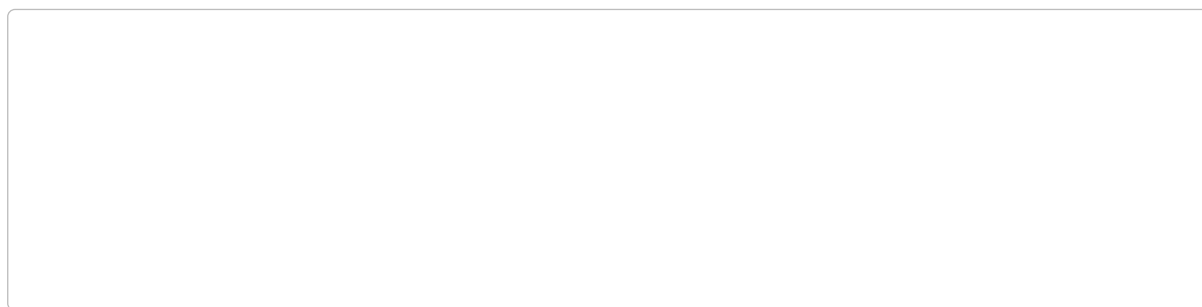
- **A** CH_4
- **B** Cl_2
- **C** H_2O

- D O₂

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]



Solutions

2.1 nitrogen, oxygen (and fluorine).

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

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.

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.

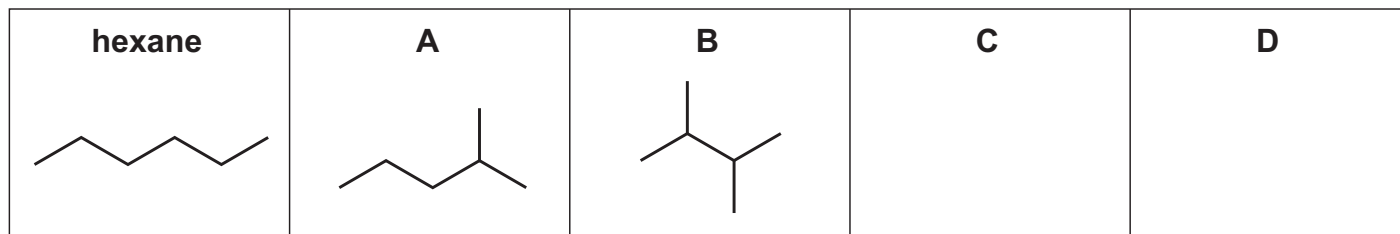


Fig. 3.1

(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.

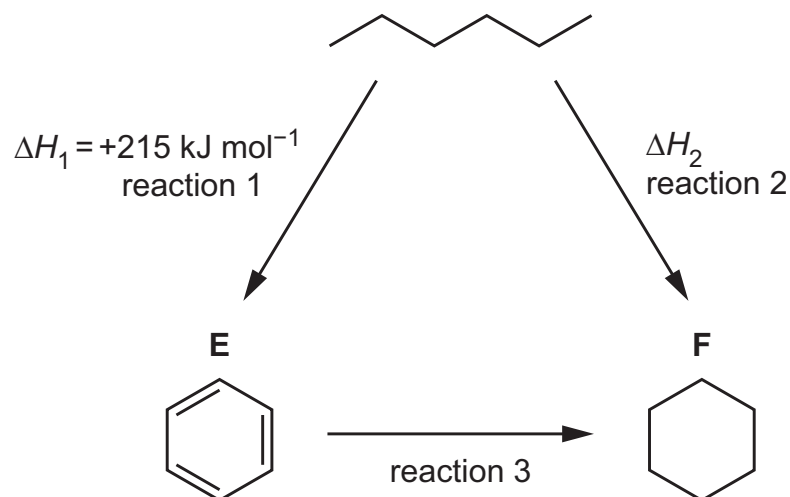
Explain your answer.

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**.

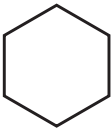
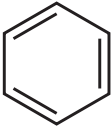


- (i) Identify a suitable reagent for reaction 3.

..... [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, $\Delta H_f^\circ / \text{kJ mol}^{-1}$
	-156
	+48

$\Delta H_2 = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

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

..... [1]

3.7 Dot-and-cross diagrams

Name: _____ Class: _____ Date: _____

Total: 9 marks

KEY WORDS dot-and-cross diagrams 点叉图 • ion 离子 • lone pair 孤对电子 • electron 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.

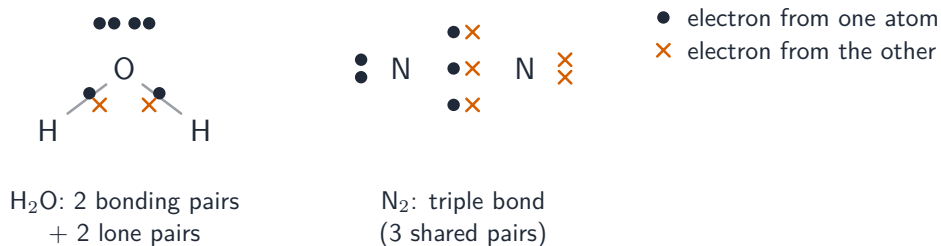
You must be able to:

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

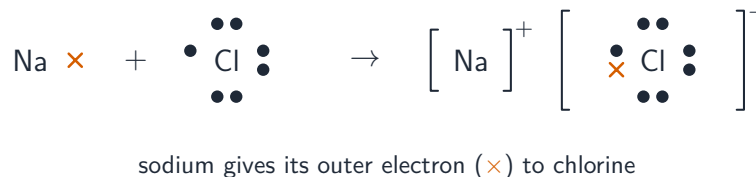
1 Worked examples

■ 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.



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



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
-

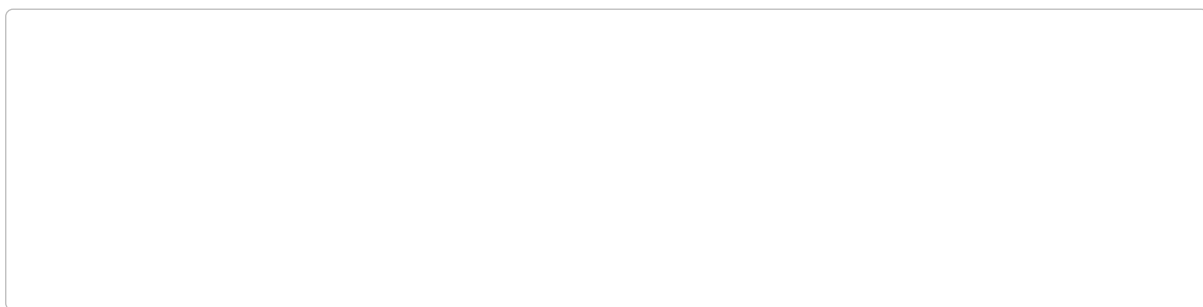
3.2 Draw a dot-and-cross diagram for a molecule of ammonia, NH_3 (show the bonding pairs and the lone pair on nitrogen). [2]

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.