

HKDSE

CHEMISTRY

A Modern View

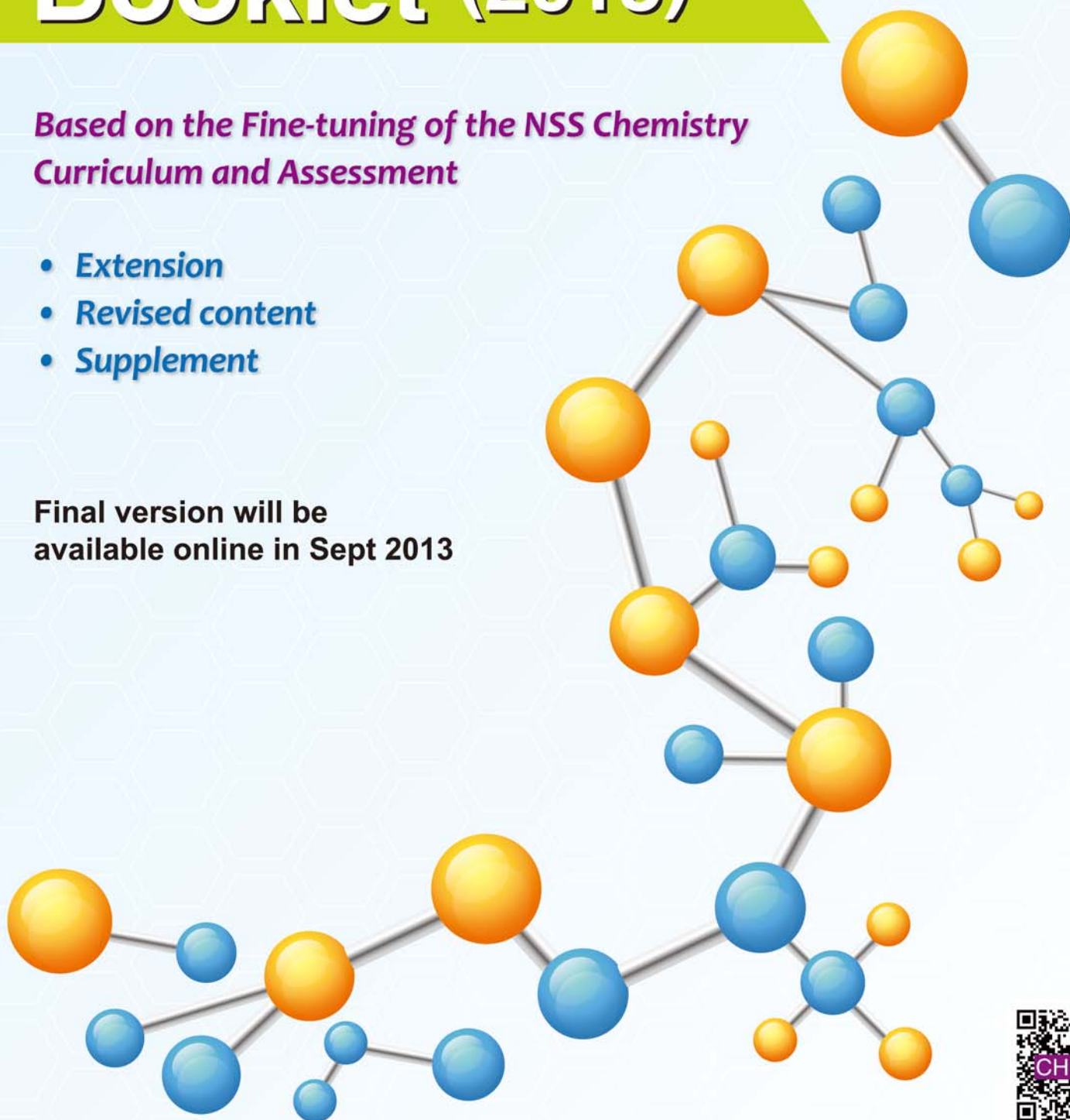
Amendment Booklet (2013)

Sample

Based on the Fine-tuning of the NSS Chemistry Curriculum and Assessment

- Extension
- Revised content
- Supplement

Final version will be available online in Sept 2013



Softcopy can be downloaded from the companion website
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HKDSE Chemistry A Modern View Amendment Booklet (2013)

© 2013 Aristo Educational Press Ltd.
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This Edition May, 2013



Preface

'HKDSE CHEMISTRY — A Modern View' Amendment Booklet (2013) is written in response to the **Fine-tuning of the NSS Chemistry Curriculum and Assessment** to be released in 2013.

Because of the fine-tuning, there are some necessary amendments to the textbooks. These amendments can be divided into three categories:

- *Extension* — refers to the topics or parts that are '**not expected**' according to the explanatory notes clarifying the breadth and depth of the curriculum.
- *Revised content* — rewritten content based on the specific requirements mentioned in the explanatory notes.
- *Supplement* — additional content that enhances understanding of concept and knowledge.

This booklet serves as a reference for teachers and students. It is important that teachers implement the curriculum based on your professional judgement and by considering students' interests and abilities.

The authors
May 2013



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Book 3 List of amendments

Part VI Microscopic world II

Chapter 25 Simple molecular substances with non-octet structures and shapes of simple molecules

Coursebook page	Item	Note
p.7	Supplement	Refer to 'N25.1 Shapes of molecules with multiple bonds obeying octet rule' on p.6–7 of this book.
	Extension	Shape of beryllium fluoride
p.13	Extension	Table 25.1 Example of beryllium fluoride
p.16	Revised content	Summary, pt.6 '...the simple molecules such as BeF_2 , ...'
p.17	Extension	BeF_2 in concept map
p.E1–E3	Extension	For those questions asking about bond angles in Q7–11; Q19, 20, 22, 23(b), 24

Chapter 26 Bond polarity

Coursebook page	Item	Note
p.25–28	Revised content	The content of 'Polarity of molecules' in 'Section 26.2 Polar and non-polar molecules' has been revised. Refer to 'N26.1 Polarity of molecules' on p.8–11 of this book.
p.30	Extension	Key terms, pt.1
		Check your progress, Q7
p.31	Extension	Summary, pts.6, 7, 9
p.32	Extension	Dipole moment in concept map
p.E4	Extension	Chapter exercise, Q4
	Revised content	Chapter exercise, Q12, Option B: It is a polar molecule without dipole moment. with all bond polarities cancelled out

Chapter 27 Intermolecular forces

Coursebook page	Item	Note
p.35–42	Revised content	The content in 'Section 27.2 Van der Waals' forces' has been revised. Refer to ' N27.1 Van der Waals' forces ' on p.12–18 of this book.
p.44	Revised content	Check your concept: dipole-dipole forces and dispersion forces van der Waals' forces
p.47	Revised content	The content in 'Example 27.1 Comparing the boiling points of propane, methoxymethane and ethanol in terms of van der Waals' forces and hydrogen bonding' has been revised. Refer to ' Example N27.1 Comparing the boiling points of propane, methoxymethane and ethanol in terms of van der Waals' forces and hydrogen bonding ' on p.19 of this book.
p.50–51	Revised content	The content in 'Section 27.4 Comparing the strengths of van der Waals' forces, hydrogen bonding and covalent bonding' has been revised. Refer to ' N27.2 Comparing the strengths of van der Waals' forces, hydrogen bonding and covalent bonding ' on p.20–21 of this book.
p.52	Extension	Key terms, pts.1, 2, 4, 5, 7
		Check your progress, Q3–5
p.53	Extension	Summary, pts.3–5
p.54	Extension	Dipole-dipole forces and dispersion forces in concept map
p.E6	Extension	Chapter exercise, Q3, 4
	Revised content	Chapter exercise, Q12 Statement (2): dipole-dipole forces hydrogen bonds Statement (3): dispersion forces van der Waals' forces

Part Exercise

Coursebook page	Item	Note
p.E12–E14	Extension	Q3, 5, 13(b),(d), 17(a),(b)

Public examination questions

Coursebook page	Item	Note
p.E16–E17	Extension	Q1, 6, 7, 8(b)(ii)

Part VII Redox reactions, chemical cells and electrolysis

Chapter 29 Chemical cells in daily life

Coursebook page	Item	Note
p.E19	Extension	Chapter exercise, Q17, 19, 20, 22

Chapter 32 Redox reactions in chemical cells

Coursebook page	Item	Note
p.141	Extension	Section 32.4 Chemistry of a lead-acid accumulator
p.145	Supplement	Refer to ' N32.1 Redox reactions in more complicated chemical cells ' on p.22–24 of this book.
p.154	Extension	Section 32.4 Chemistry of a lead-acid accumulator
p.158	Extension	Class practice 32.4
p.159	Extension	Key terms, pt.5
		Check your progress, Q8, 9
p.160	Extension	Summary, pt.6
p.E32–E33	Extension	Chapter exercise, Q4, 16

Chapter 34 Importance of redox reactions in modern ways of living

Coursebook page	Item	Note
p.206–207	Extension	The content of 'Different types of fuel cells and their applications' in 'Section 34.2 Fuel cells and their applications'
p.209	Extension	Class practice 34.1
p.213	Extension	Key terms, pts.1, 8–11
		Check your progress, Q2
p.214	Extension	Summary, pt.4
p.215	Extension	Common types of fuel cell and their applications in concept map
p.E41	Extension	Chapter exercise, Q3
p.E43	Extension	Chapter exercise, Q15

Part VIII Chemical reactions and energy

Chapter 36 Standard enthalpy change of combustion, neutralization, solution and formation

Coursebook page	Item	Note
p.235	Revised content	Bullet pt.1 '...combustion, neutralization, solution and formation'
p.244–245	Extension	The content of 'Standard enthalpy change of solution' in 'Section 36.1 Standard enthalpy change of combustion, neutralization, solution and formation'
p.248	Extension	Class practice 36.3 Q2(c)
p.255–256	Extension	The content of 'Determination of enthalpy change of solution' in 'Section 36.2 Simple calorimetric methods'
p.258	Extension	Key terms, pt.10
		Check your progress, Q8

Coursebook page	Item	Note
p.259	Extension	Summary, pt.5
p.261	Extension	Standard enthalpy change of solution in concept map
p.E58	Extension	Chapter exercise, Q5
p.E60	Extension	Chapter exercise, Q19

Chapter 37 Hess's Law

Coursebook page	Item	Note
p.263	Revised content	Bullet pt.1 apply Hess's Law to construct simple enthalpy change cycles and enthalpy level diagrams
p.266–267	Extension	The content of 'Enthalpy level diagram' in 'Section 37.1 Hess's Law'
	Extension	Class practice 37.1 Part (b)
p.269	Extension	Class practice 37.2 Part (b)
p.277	Extension	Key terms, pt.2
		Check your progress, Q2
p.278	Extension	Summary, pt.3
p.279	Extension	Enthalpy level diagram in concept map
p.E61	Extension	Chapter exercise, Q2, 6

Part Exercise

Coursebook page	Item	Note
p.E63–E64	Extension	Q5, 11
p.E66	Extension	Q19

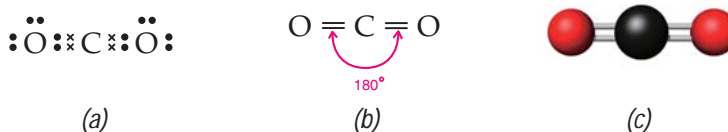
N25.1 Shapes of molecules with multiple bonds obeying octet rule

We can predict the shape of simple molecules containing multiple bonds using the Valence Shell Electron Pair Repulsion (VSEPR) theory. The theory states that a **multiple bond** (i.e. double bond and triple bond) is treated as an electron pair.

Carbon dioxide (CO₂)

In a carbon dioxide molecule, the carbon atom forms a double bond with each of the two oxygen atoms (Figure N25.1a). As a double bond is treated as an electron pair according to the VSEPR theory, the central carbon atom can be treated as having **two electron pairs** (all bond pairs) in its outermost shell (Figure N25.1b). To minimize the repulsion between them, the two electron pairs lie on the opposite sides of the carbon atom. Thus, the molecule is **linear** in shape. The O=C=O bond angle is 180° (Figure N25.1c).

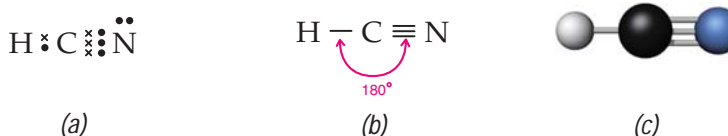
Figure N25.1 (a) Electron diagram of carbon dioxide, CO₂. (b) The CO₂ molecule is linear in shape. (c) The ball-and-stick model of a CO₂ molecule.



Hydrogen cyanide (HCN)

In a hydrogen cyanide molecule, the carbon atom forms a single bond with a hydrogen atom and a triple bond with a nitrogen atom (Figure N25.2a). As a triple bond is treated as an electron pair according to the VSEPR theory, the central carbon atom is treated as having **two electron pairs** (all bond pairs) in its outermost shell (Figure N25.2b). The two electron pairs lie on the opposite sides of the carbon atom so as to minimize their repulsion. Thus, the molecule is **linear** in shape. The H—C≡N bond angle is 180° (Figure N25.2c).

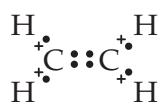
Figure N25.2 (a) Electron diagram of hydrogen cyanide, HCN. (b) The HCN molecule is linear in shape. (c) The ball-and-stick model of a HCN molecule.



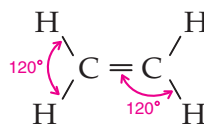
Ethene (C_2H_4)

In an ethene molecule, there are two carbon atoms. For each carbon atom, it forms two single bonds with two hydrogen atoms and a double bond with another carbon atom (Figure N25.3a). As the double bond is treated as an electron pair according to the VSEPR theory, each carbon atom is treated as having **three electron pairs** (all bond pairs) in its outermost shell (Figure N25.3b). To minimize the repulsion between them, the three electron pairs take up the **trigonal planar** arrangement.

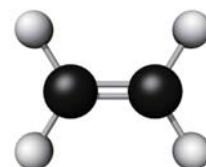
Thus, the ethene molecule consists of two trigonal planar units joined together. All the six atoms lie on the same plane. The molecule has a **planar** structure. The bond angles of all $\text{H}-\text{C}-\text{H}$ and $\text{C}=\text{C}-\text{H}$ are approximately 120° (Figure N25.3c).



(a)



(b)



(c)

Figure N25.3 (a) Electron diagram of ethene, C_2H_4 . (b) The C_2H_4 has a planar structure. (c) The ball-and-stick model of a C_2H_4 molecule.



Class practice N25.1

For each of the following molecules:

- (a) methanal (HCOH)
- (b) ethyne (C_2H_2)
 - (i) Draw an electron diagram (showing electrons in the outermost shells only) of molecule.
 - (ii) Draw a three-dimensional diagram of molecule.
 - (iii) Deduce its shape.

N26.1 Polarity of molecules

Hydrogen chloride is *diatomic*. Its molecule has only *one polar bond*, H—Cl, which is formed by bonding a hydrogen atom and a chlorine atom together. Thus, a hydrogen chloride molecule is *polar*. See Figure N26.1.



Figure N26.1 HCl is a polar molecule which has a partial positive end and a partial negative end.

If a molecule has more than one bond, how can we decide whether it is polar or not? In such case, we should consider its *shape* and *polarity of each bond*. These two factors affect the overall polarity of a molecule.

Polar molecules

Water molecule



Learning tip

Water is an ideal solvent for most ionic compounds because of the high polarity of its molecules.

A water molecule has two polar O—H bonds, in which the shared electrons are attracted more closely to the oxygen atom. As the water molecule is *V-shaped*, the polarities of the polar O—H bonds *cannot cancel out* each other. Thus, a water molecule is *polar*. See Table N26.1.



Think about

Why is an ammonia molecule polar?

Trichloromethane molecule

A trichloromethane molecule has three polar C—Cl bonds (in which the shared electrons are attracted more closely to the chlorine atom) and one polar C—H bond (in which the shared electrons are attracted more closely to the carbon atom). As the molecule is *tetrahedral* in shape, the polar C—Cl and C—H bonds are *symmetrically* arranged. However, as the polarities of the four polar bonds in the molecule are *not identical*, the polarities of the polar bonds *cannot cancel out* each other. Thus, a trichloromethane molecule is *polar*. See Table N26.1.

Some common examples of **polar molecules** are given in Table N26.1 below.

Formula	Shape	Spatial arrangement of bonds ($\delta^+ \rightarrow \delta^-$ = polarity of individual bond)	Indication of the molecular polarity
HF	Linear	$\delta^+ \text{H} \rightarrow \delta^- \text{F}$	$\delta^+ \rightarrow \delta^-$
H ₂ O	V-shaped		δ^- δ^+
NH ₃	Trigonal pyramidal		δ^- δ^+
CHCl ₃	Tetrahedral		δ^+ δ^-

Table N26.1 Some common examples of polar molecules.



Think about

Why are H₂, O₂ and Cl₂ molecules non-polar?

Non-polar molecules

Molecules without polar bonds are of course non-polar. On the other hand, many molecules with polar bonds can also be non-polar. Why? The polar bonds in these molecules are of equal polarity and are arranged *symmetrically*. As a result, all the polarities of the polar bonds *cancel out*.

Some examples of **non-polar molecules** (with polar bonds) are given in Table N26.2.

Formula	Shape	Spatial arrangement of bonds (' \rightarrow ' = polarity of individual bond)
CO ₂	Linear	$\begin{array}{c} \delta^- \leftarrow \delta^+ \rightarrow \delta^- \\ \text{O} = \text{C} = \text{O} \end{array}$
BF ₃	Trigonal planar	
CH ₄	Tetrahedral	

Table N26.2 Some examples of non-polar molecules.

**Think about**

Why is a methane molecule non-polar?

A boron trifluoride molecule is *non-polar*. It has three polar B–F bonds in which shared electrons are attracted more closely to the more electronegative fluorine atoms. As the molecule is *trigonal planar* in shape, the polar B–F bonds are *symmetrically* arranged. The polarities of the three B–F bonds *cancel out* each other.

The *polarity of covalent bonds* and the *shape of the molecule* affect the *overall polarity* of a molecule.

In conclusion, we can determine whether a molecule is polar or non-polar by answering the questions in Figure N26.2.

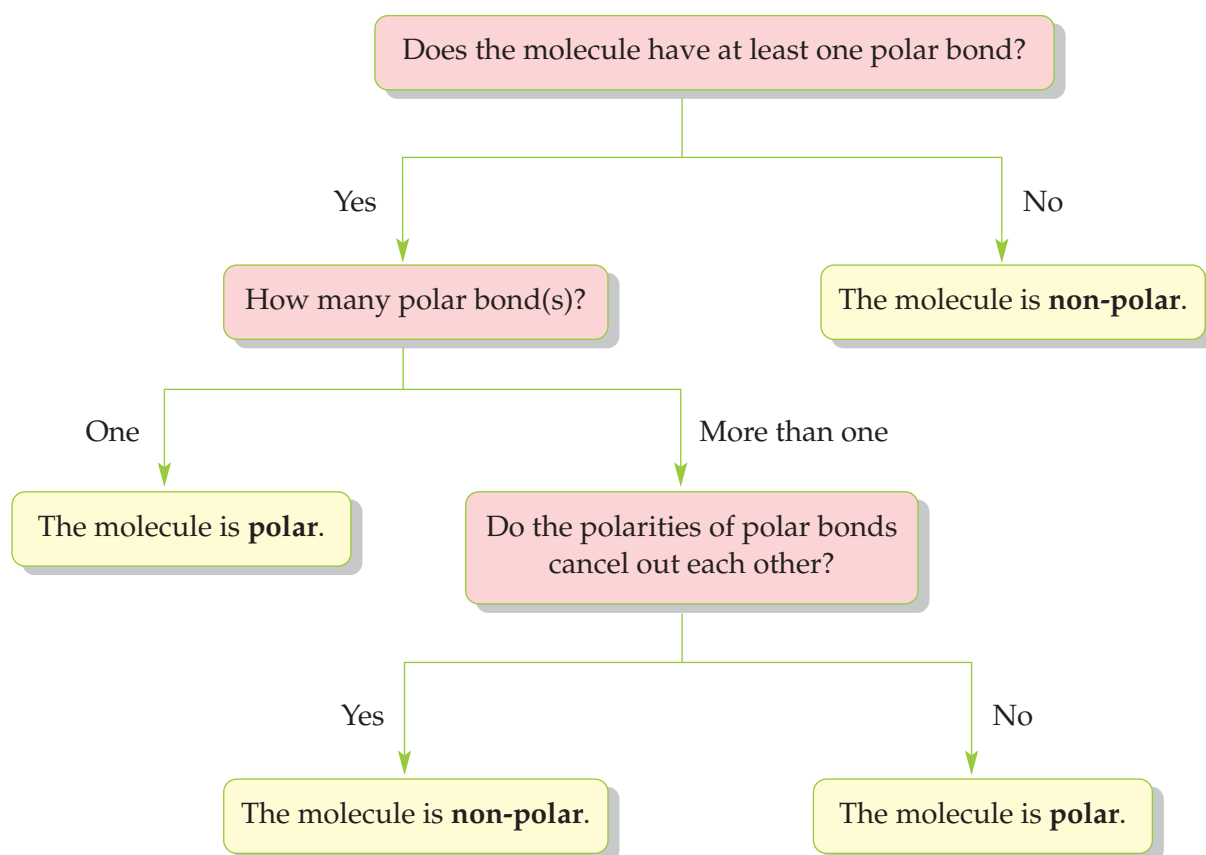


Figure N26.2 A flow chart for determining whether a molecule is polar or not.



Everyday chemistry

Why does oil not mix with water?

We all know that oil and water do not mix. However, do you know that this phenomenon is related to the polarity of molecules? There is an expression 'like dissolves like'. Polar compounds dissolve in polar compounds and non-polar compounds dissolve in non-polar compounds.



Oil is composed of non-polar molecules while water is composed of polar molecules. They do not dissolve in each other. When we mix oil and water, two layers are formed. Oil floats on water because of its lower density.

N27.1 Van der Waals' forces

Van der Waals' forces are named after the Dutch scientist Johannes van der Waals (Figure N27.1). He studied the attractive forces holding molecules together. Van der Waals' forces exist between **all molecules**, no matter the molecule is polar or non-polar. Let us take hydrogen chloride, argon and chlorine as examples, to explain the existence of van der Waals' forces between molecules.



Figure N27.1 Johannes van der Waals (1837–1923).

Van der Waals' forces between polar molecules

In a hydrogen chloride molecule, the hydrogen atom *has a partial positive charge* ($\delta+$) and the chlorine atom *has a partial negative charge* ($\delta-$). The positive end of an HCl molecule will be attracted to the negative end of another HCl molecule. The attractive forces between HCl molecules are *van der Waals' forces* (Figure N27.2).

(Note: The green part in the molecule denotes electron density. The darker the colour, the higher is the electron density.)

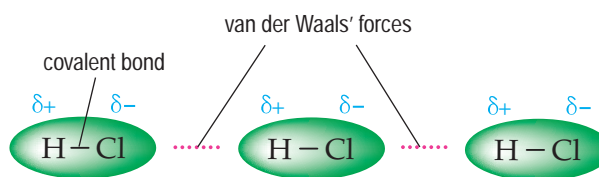


Figure N27.2 Van der Waals' forces exist between hydrogen chloride molecules.



Think about

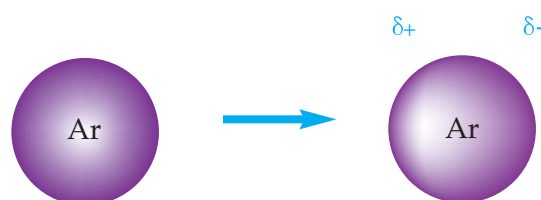
Why are van der Waals' forces much weaker than ionic and covalent bonds?

Van der Waals' forces are *electrostatic* in nature like ionic bonds or covalent bonds, though much weaker.

Van der Waals' forces between non-polar molecules

In an argon atom, the electrons are in *constant motion*. At any given instant, the electron distribution may be *uneven* and there are more electrons on one side of the atom than on the other side. The side with more electrons would *form a partial negative charge* (δ^-), and the side with fewer electrons would *form a partial positive charge* (δ^+). Obviously, the charges are *not permanent*. They continue to arise and disappear all the time (Figure N27.3).

(Note: The purple part in the atom denotes electron density. The darker the colour, the higher is the electron density.)



At any given instant, there are more electrons on one side of the atom than on the other side. As a result, a pair of opposite charges would be formed on two sides of the atom. But the charges are not permanent.

Figure N27.3 An argon atom would have a partial positive charge on one side and a partial negative charge on the other side at any given instant.

At any given instant, the partial charges on one atom can affect the electron distributions in neighbouring atoms. They induce a similar separation of charges on the neighbouring atoms. As a result, the atoms are attracted to each other. These attractive forces are also *van der Waals' forces* (Figure N27.4).

At any given instant, the partial negative charge on one atom repels electrons on the second atom. This gives rise to a region of partial positive charge on the second atom. The two partial charges attract each other.

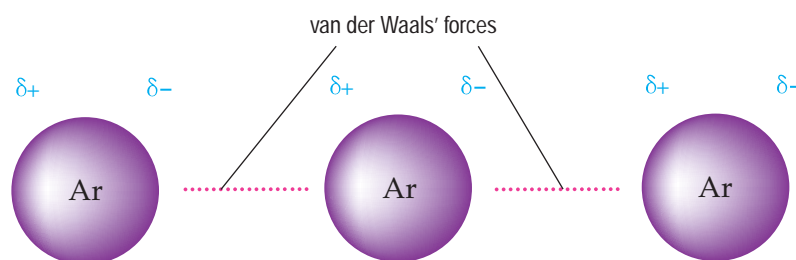


Figure N27.4 Van der Waals' forces exist between argon atoms.

In the case of chlorine, van der Waals' forces exist between molecules as illustrated in Figure N27.5 below.

(Note: The yellow part in the molecule denotes electron density. The darker the colour, the higher is the electron density.)

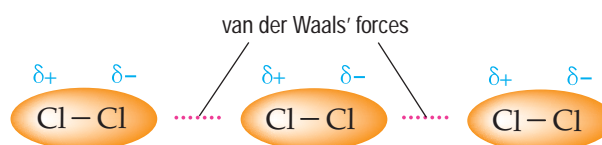


Figure N27.5 Van der Waals' forces exist between chlorine molecules.

Van der Waals' forces exist between *all* molecules and are *electrostatic* in nature.

Factors affecting the strength of van der Waals' forces between molecules

In general, the strength of van der Waals' forces depends on the following three factors.

1. Molecular size

In general, the *larger* the size of a molecule, the *greater* is the number of electrons in the molecule. This results in a *greater chance* of the uneven distribution of electrons in the molecule. In other words, the strength of van der Waals' forces *increases* with molecular size.

Boiling points of simple molecules depend on the strength of intermolecular forces. Let us consider halogens and noble gases as examples. As the size of the molecules *increases* down the group, the *van der Waals' forces* between the molecules become *stronger*. More energy is required to separate the molecules. As a result, their boiling points *increase* down the group (Table N27.1).

Group number	Molecule	Boiling point (°C)
VII	Fluorine (F ₂)	-188
	Chlorine (Cl ₂)	-35
	Bromine (Br ₂)	59
	Iodine (I ₂)	184
0	Helium (He)	-269
	Neon (Ne)	-246
	Argon (Ar)	-186
	Krypton (Kr)	-152

Table N27.1 Boiling points of Group VII and 0 elements increase down the group.

2. Molecular shape

Long, thin molecules have a *larger contact surface area* with one another. This allows the partial charges of molecules produced at any given instant to *interact more strongly*. As a result, there are *stronger* van der Waals' forces.



Learning tip

Refer to the discussion of straight-chain and branched-chain structures learnt in Chapter 22.

For example, both pentane and 2,2-dimethylpropane have the molecular formula of C_5H_{12} . In pentane, the carbon atoms are arranged in a *straight chain*. In 2,2-dimethylpropane, the carbon atoms are arranged in a *branched chain*. Pentane has a higher boiling point than 2,2-dimethylpropane (Table N27.2).

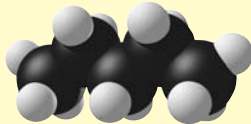
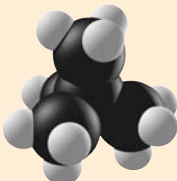
Molecule	Molecular formula	Structural formula and space-filling model	Boiling point (°C)
Pentane	C_5H_{12}	$ \begin{array}{ccccccc} & H & H & H & H & H & \\ & & & & & & \\ H & -C & -C & -C & -C & -C & -H \\ & & & & & & \\ & H & H & H & H & H & \end{array} $ 	36
2,2-dimethylpropane	C_5H_{12}	$ \begin{array}{ccccc} & & H & & \\ & & & & \\ H & & H-C-H & & H \\ & & & & \\ H-C & - & C & - & C-H \\ & & & & \\ & & H-C-H & & H \\ & & & & \\ & & H & & \end{array} $ 	9

Table N27.2 Structural formulae and boiling points of pentane and 2,2-dimethylpropane.

As pentane molecules are *rod-shaped*, they have a *larger contact surface area* with one another. Thus, the *van der Waals' forces* between the molecules are *stronger*. On the other hand, 2,2-dimethylpropane molecules are *spherical in shape*. They have a *smaller contact surface area* with one another, so the van der Waals' forces between them are *weaker*. This explains why pentane has a *higher boiling point* than 2,2-dimethylpropane. See Figure N27.6.

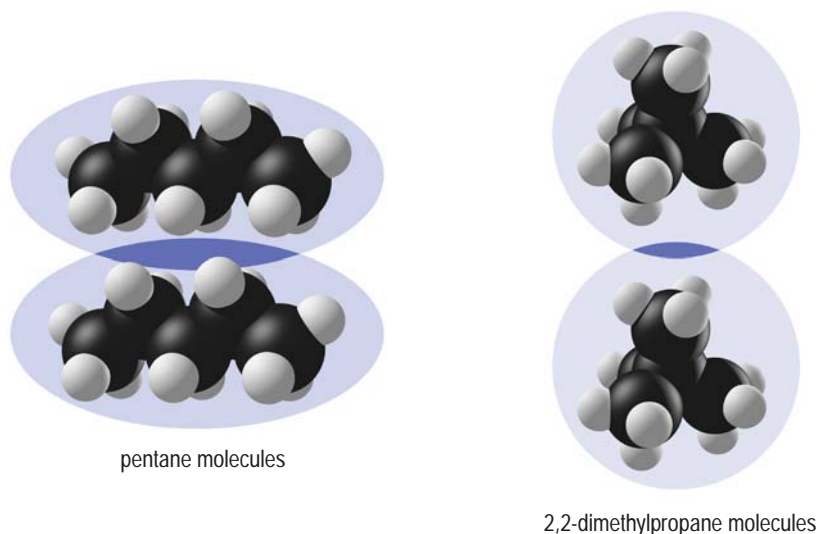
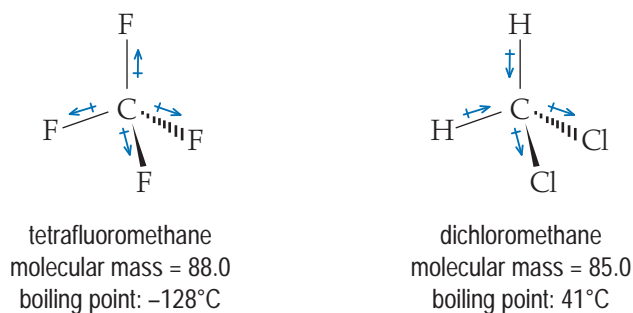


Figure N27.6 Rod-shaped pentane molecules have larger contact surface area with one another than spherical 2,2-dimethylpropane molecules. (The representation of the contact surface area is symbolic.)

3. Polarity of molecules

Let us consider tetrafluoromethane and dichloromethane as examples.



There are four polar C—F bonds in a tetrafluoromethane molecule. As the four polar bonds are symmetrically arranged, the polarities of the polar bonds cancel out each other. Thus, the tetrafluoromethane molecule is non-polar in nature. However, the polarities of the four polar bonds in the dichloromethane molecule cannot cancel out each other. Thus, it is *polar* in nature.

For molecules with *comparable* sizes and shapes, van der Waals' forces between polar molecules are generally stronger than those between non-polar ones. Therefore, dichloromethane has a higher boiling point than tetrachloromethane.

The **van der Waals' forces** between the molecules are *stronger* if:

- the *molecular size* is larger,
- the *contact surface area* between the molecules is greater,
- the molecules are *more polar*.



Do you know?

Scientists have discovered that spiders, like geckos, stick to ceilings or walls by van der Waals' forces. The forces are so strong that spiders could carry over 170 times their own body weight while standing on the ceiling.





Class practice N27.1

1. The information about some hydrocarbons is given below:

Hydrocarbon	Structural formula	Boiling point (°C)
Propane (C ₃ H ₈)	<pre> H H H H — C — C — C — H H H H </pre>	–42.0
Butane (C ₄ H ₁₀)	<pre> H H H H H — C — C — C — C — H H H H H </pre>	4.6
2-methylpropane (C ₄ H ₁₀)	<pre> H H H H — C ——— C ——— C — H H H — C — H H H </pre>	–11.7
Pentane (C ₅ H ₁₂)	<pre> H H H H H H — C — C — C — C — C — H H H H H H </pre>	36.0

- State whether the above hydrocarbons consist of polar or non-polar molecules.
 - Name the intermolecular forces between these hydrocarbon molecules.
 - Based on the above table, comment on the following statements:
 - The boiling points of straight-chain hydrocarbons increase with their molecular masses.
 - The intermolecular forces between branched-chain hydrocarbons are stronger than those between straight-chain hydrocarbons with similar molecular masses.
2. Account for the difference in boiling points for each of the following pairs of substances:
- Ethane (C₂H₆): b.p. –88°C and fluoromethane (CH₃F): b.p. –78°C
 - Hydrogen chloride (HCl): b.p. –85°C and chlorine (Cl₂): b.p. –34°C

Example N27.1 Comparing the boiling points of propane, methoxymethane and ethanol in terms of van der Waals' forces and hydrogen bonding



Example N27.1

Comparing the boiling points of propane, methoxymethane and ethanol in terms of van der Waals' forces and hydrogen bonding

Consider the following three compounds with *comparable molecular masses*. Explain the difference in their boiling points in terms of intermolecular forces.

	$ \begin{array}{c} \text{H} \quad \text{H} \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \text{H} \quad \text{H} \end{array} $	$ \begin{array}{c} \text{H} \quad \quad \text{H} \\ \quad \quad \\ \text{H}-\text{C}-\text{O}-\text{C}-\text{H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $	$ \begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}-\text{C}-\text{C}-\text{O}-\text{H} \\ \quad \\ \text{H} \quad \text{H} \end{array} $
Compound	propane	methoxymethane	ethanol
Boiling point	-42.0°C	-24.8°C	78.5°C
Relative molecular mass	44.0	46.0	46.0

Solution

Hydrogen bonds in addition to van der Waals' forces exist between ethanol molecules. Hence, ethanol has the highest boiling point.

Methoxymethane and propane have much lower boiling points because there are *only van der Waals' forces* between their molecules. Methoxymethane molecules are *polar* while propane molecules are *non-polar*. As a result, van der Waals' forces between methoxymethane molecules are stronger than those between propane molecules. Therefore, methoxymethane has a higher boiling point than propane.

➡ Try Chapter Exercise Q21

N27.2 Comparing the strengths of van der Waals' forces, hydrogen bonding and covalent bonding

Both hydrogen bonding and van der Waals' forces are intermolecular forces. Figure N27.7 shows a systematic way of identifying the type of intermolecular forces in a particular substance.

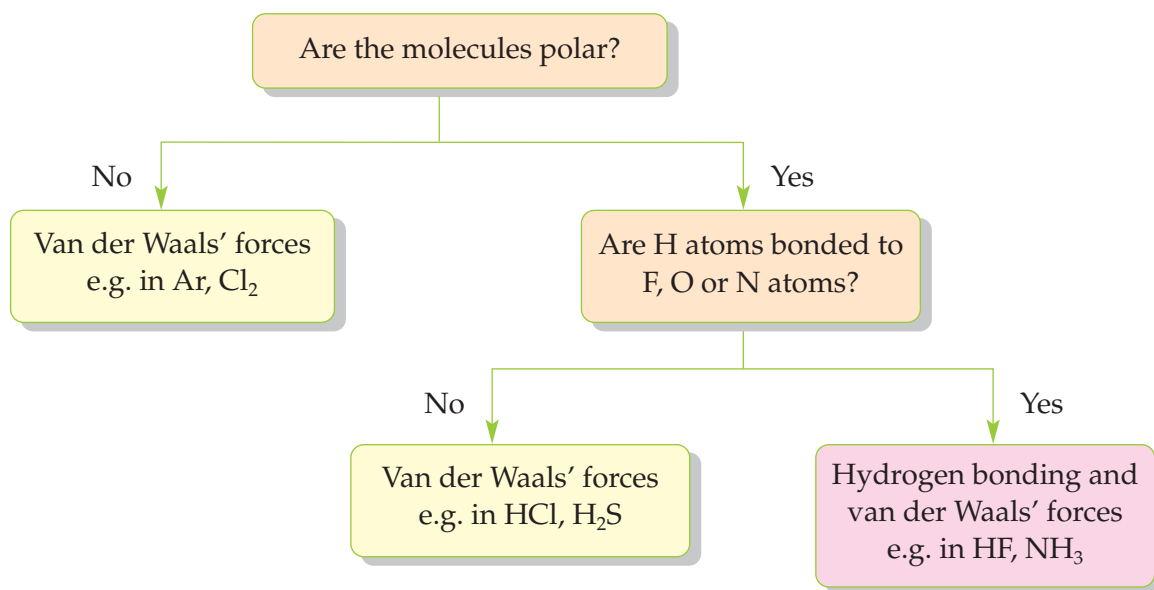


Figure N27.7 A flow chart for identifying the type of intermolecular forces in a substance.

None of the above intermolecular forces are as strong as ionic or covalent bonds. Although hydrogen bonds are the strongest type of intermolecular forces, they are still much weaker than typical covalent bonds. See Table N27.3.

Type of forces	Relative strength
Van der Waals' forces	0.01 – 50
Hydrogen bonding	20 – 40
Covalent bonding	100 – 1000

Table N27.3 The relative strengths of van der Waals' forces, hydrogen bonding and covalent bonding.

**Example N27.2**

Accounting for the difference in boiling points of some covalent substances in terms of inter-particle forces

The table below lists the boiling points of some covalent substances:

	CO ₂	SO ₂	H ₂ O ₂	SiO ₂
Boiling point (°C)	-78	-10	150	2230

Account for the difference in boiling points of the above substances.

Solution

SiO₂ has the highest boiling point. It has a *giant covalent structure* in which all atoms are held together by *strong covalent bonds*. As a result, a lot of energy is needed to separate the atoms.

On the other hand, CO₂, SO₂ and H₂O₂ all have *simple molecular structures*. H₂O₂ molecules are held together by *hydrogen bonds and van der Waals' forces* while both CO₂ and SO₂ molecules are held together by *van der Waals' forces* only. As the intermolecular forces between H₂O₂ are stronger than those between CO₂ and SO₂, H₂O₂ has a much higher boiling point.

SO₂ molecules are *polar* while CO₂ molecules are *non-polar*. Van der Waals' forces between SO₂ molecules are stronger than those between CO₂ molecules. Therefore, SO₂ has a higher boiling point than CO₂.

**Experiment N27.1**

Experiment Workbook 3

Investigating the evaporation rates of substances with different intermolecular forces

In this experiment, you are going to investigate the evaporation rates of water, ethanol, propanone and pentane, so that you can find out the effect of intermolecular forces on the ease of evaporation.

**Class practice N27.2**

Explain the following phenomena:

- The melting point of diamond is much higher than that of ice.
- The boiling point of water is much higher than that of oxygen.
- The evaporation rate of hydrogen chloride is much higher than that of hydrogen fluoride.

N32.1 Redox reactions in more complicated chemical cells

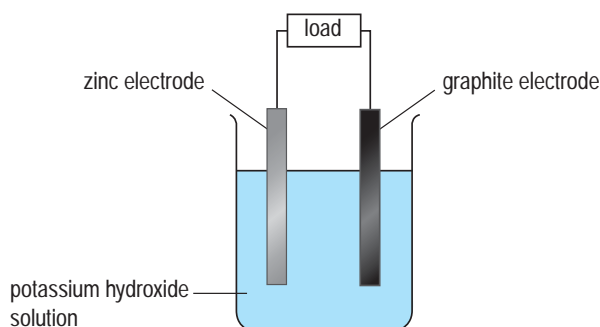
Examples N32.1 and N32.2 below show how we apply the concepts of electrochemistry to solve problems involving more complicated chemical cells.



Example N32.1

Studying the redox reaction in a chemical cell with zinc and graphite electrodes

The following shows a simplified diagram of the chemical cell used in an electric car. During discharge, zinc undergoes oxidation to give zincate ions, $\text{ZnO}_2^{2-}(\text{aq})$. The electrolyte is potassium hydroxide solution. At the graphite electrode, oxygen and water react to produce hydroxide ions.



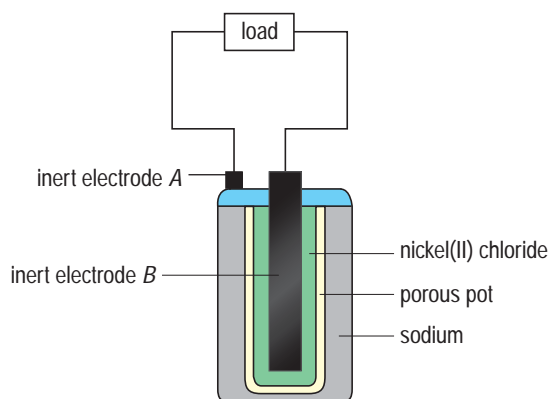
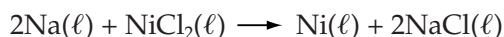
- Write balanced equations for the reactions occurring at
 - the zinc electrode;
 - the graphite electrode.
- Identify the anode and cathode of the chemical cell.
- Write the overall equation for the reaction occurring in the above chemical cell.
- What is the direction of electron flow in the external circuit?

Solution

- $\text{Zn}(\text{s}) + 4\text{OH}^-(\text{aq}) \rightarrow \text{ZnO}_2^{2-}(\text{aq}) + 2\text{H}_2\text{O}(\ell) + 2\text{e}^-$
 - $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$
- Zinc electrode is the anode while graphite electrode is the cathode.
- $2\text{Zn}(\text{s}) + 4\text{OH}^-(\text{aq}) + \text{O}_2(\text{g}) \rightarrow 2\text{ZnO}_2^{2-}(\text{aq}) + 2\text{H}_2\text{O}(\ell)$
- Electrons flow from the zinc electrode to the graphite electrode.


Example N32.2
Studying the redox reaction in a sodium-nickel(II) chloride cell

The following diagram shows a sodium-nickel(II) chloride cell connected to a load. The electrodes are inert and the electrolyte consists of nickel(II) chloride. The cell operates at about 300°C. The overall reaction of the chemical cell is:



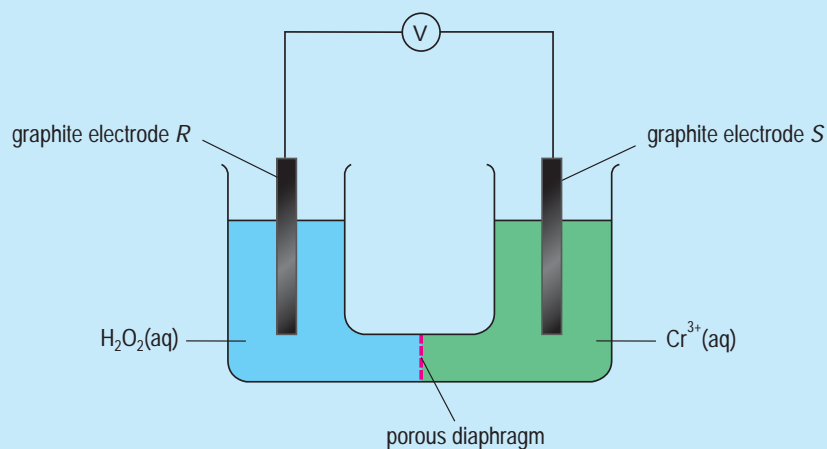
- What is the direction of electron flow in the external circuit when the cell is discharged? Explain briefly.
- Write ionic half-equations for the reactions that occur at electrodes A and B respectively.
- Suggest why a high temperature is needed for the cell to operate.
- Sodium-nickel(II) chloride cell is rechargeable and can be used in electric vehicles. Suggest ONE advantage of using this cell as power source.

Solution

- Electrons flow from electrode A to electrode B because sodium atoms lose electrons when the cell is discharged.
- At electrode A: $\text{Na}(\ell) \longrightarrow \text{Na}^+(\ell) + \text{e}^-$
At electrode B: $\text{Ni}^{2+}(\ell) + 2\text{e}^- \longrightarrow \text{Ni}(\ell)$
- To keep sodium and the electrolyte in molten state.
- No exhaust gas is produced.

Class practice N32.1

The following diagram shows a chemical cell with a porous diaphragm separating the anode and cathode compartments. During discharge, hydrogen peroxide undergoes reduction to give hydroxide ions while chromium(III) ions undergo oxidation in alkaline medium to give chromate(VI) ions ($\text{CrO}_4^{2-}(\text{aq})$).



- (a) (i) Write an ionic half equation for the reaction occurring at electrode R.
- (ii) Explain why the reaction occurring at electrode R is a reduction reaction in terms of change in oxidation number.
- (b) (i) Write an ionic half equation for the reaction occurring at electrode S.
- (ii) Explain why the reaction occurring at electrode S is an oxidation reaction in terms of change in oxidation number.
- (c) Write the overall equation for the reaction occurring in the above chemical cell.
- (d) State the observable change in the anode compartment.
- (e) Suggest what would happen if the porous diaphragm is removed in the above chemical cell.