Mastering Chemistry

- Book 2C
- Topic 6 Microscropic World II

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Content

24.1 Sharing of bonding electrons in a covalent bond (p.23)

- In a covalent bond, two atoms are held together because both nuclei are attracted to the bonding electrons.
- The electron pair in a covalent bond is shared equally if the two bonding atoms come from the same element.

24.1 Sharing of bonding electrons in a covalent bond (p.23)

- If the two bonding atoms come from different elements, the nucleus of one of the atoms is likely to attract the bonding electrons more strongly than the nucleus of the other atom.
- The electron pair in the covalent bond will be unequally shared. A **polar (極性的)** bond results.

24.1 Sharing of bonding electrons in a covalent bond (p.23)

- The ability of an atom in a covalent bond to attract the bonding electrons towards itself is measured by the **electronegativity (電負性)** of the element.

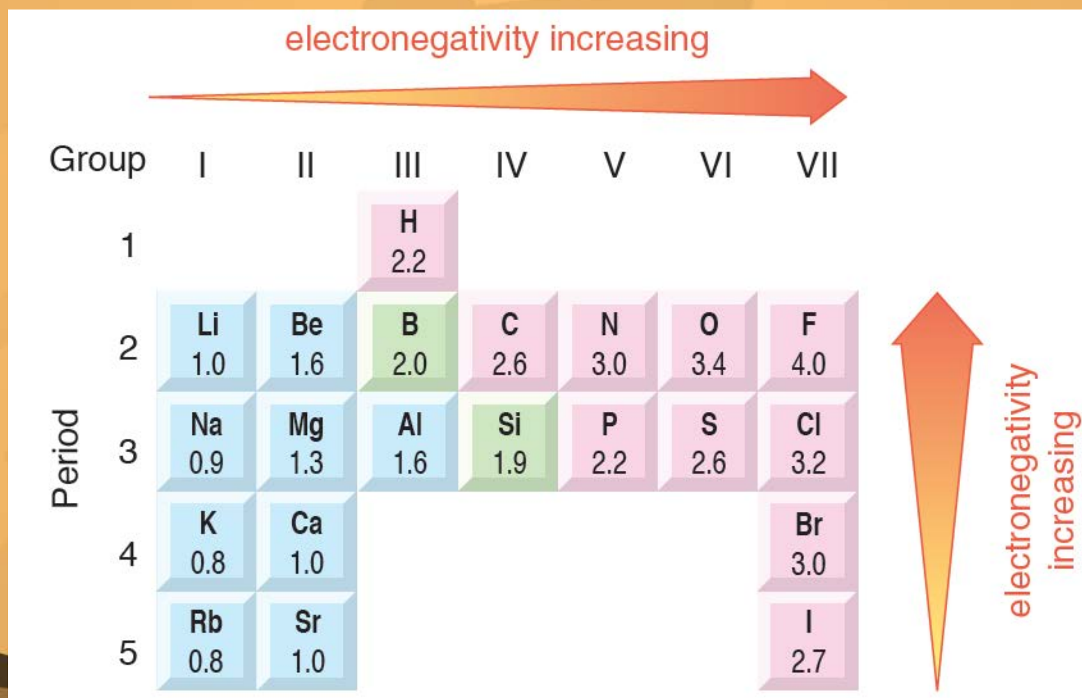
Electronegativity of an element is a measure of the ability of an atom of the element to attract the bonding electrons in a covalent bond towards itself.

24.2 Electronegativity and the Periodic Table (p.24)

- There are various scales of electronegativity. The most commonly used scale of electronegativity is the one developed by Linus Pauling.
- On this scale, the most electronegative element, fluorine, has a value of 4.0. The least electronegative element, francium, has a value of 0.7. Notice that electronegativity values are relative values with no units.
- The greater the electronegativity value of an element, the greater the ability of its atom to attract shared electrons towards itself.

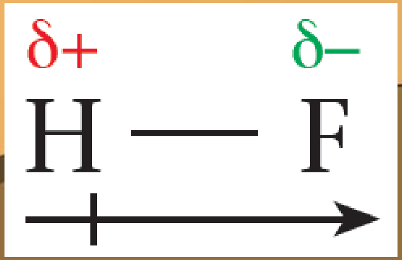
24.2 Electronegativity and the Periodic Table (p.24)

- Electronegativity increases across a period;
- Electronegativity decreases down a group.
- Noble gases have no electronegative values because they do not form covalent bonds.



24.3 Covalent bond between two different atoms (p.25)

- When a covalent bond is formed between atoms of two elements with different electronegativities, the bonding electrons will not be shared equally between the atoms.
- Hydrogen has an electronegativity value of 2.2 and fluorine of 4.0. This means that the electrons in the covalent bond will be attracted more by the fluorine atom than by the hydrogen atom.
- F: partial negative charge (δ–) due to a greater share of the bonding electrons;
  H: positive charge (δ+) due to loss of its share of the bonding electrons.
  The greater the difference in electronegativity of the bonding atoms, the more polar the covalent bond is.



24.3 Covalent bond between two different atoms (p.25)

Practice 24.1

Based on the electronegativity values shown in Fig. 24.3, mark the partial charges carried by atoms in the covalent bonds below, using the symbols δ+ and δ–.

a) $\overset{\delta-}{C}$–$\overset{\delta+}{H}$

b) $\overset{\delta+}{B}$–$\overset{\delta-}{F}$

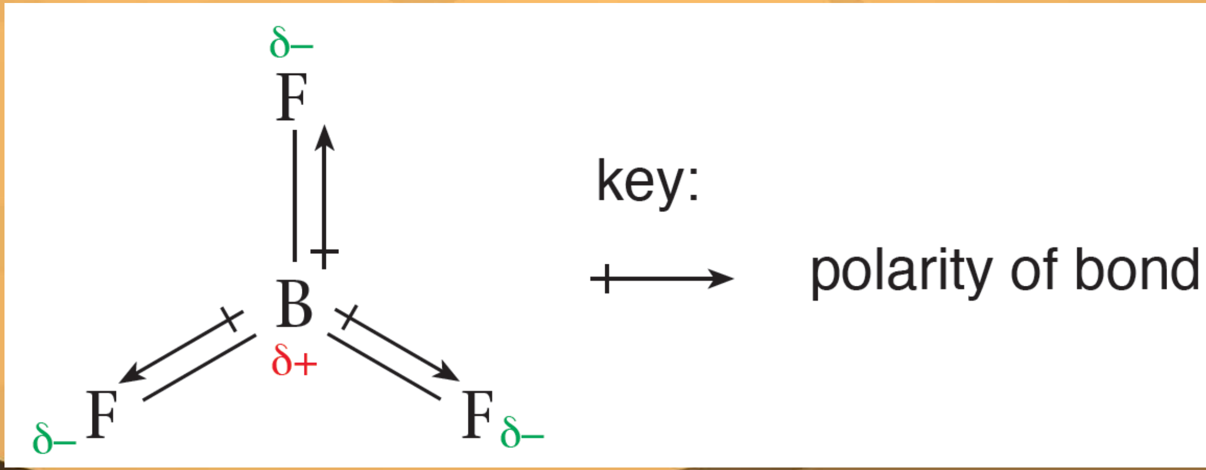
c) $\overset{\delta-}{N}$–$\overset{\delta+}{H}$

d) $\overset{\delta-}{O}$–$\overset{\delta+}{H}$

e) $\overset{\delta+}{P}$–$\overset{\delta-}{F}$

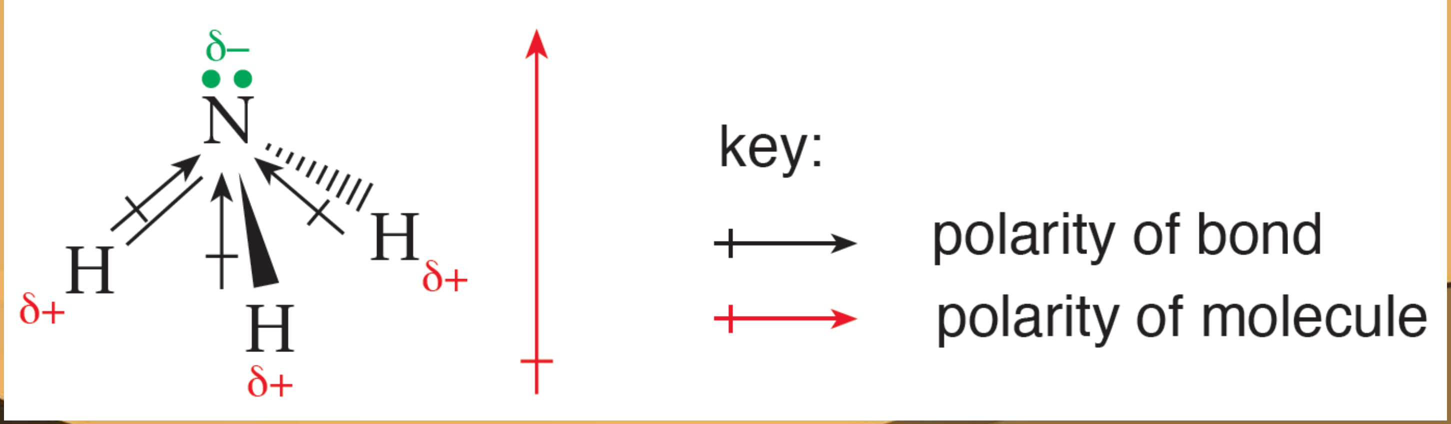
24.4 Polarity of molecules (p.26)

- Condiser BF_3.
- F is more electronegative than B. So each B–F bond is polar.
- BF_3 molecule is trigonal planar. The three B–F bonds are symmetrically arranged on the same plane around B.
- The polarities cancel each other out and so the molecule is **non-polar (非極性的)**.



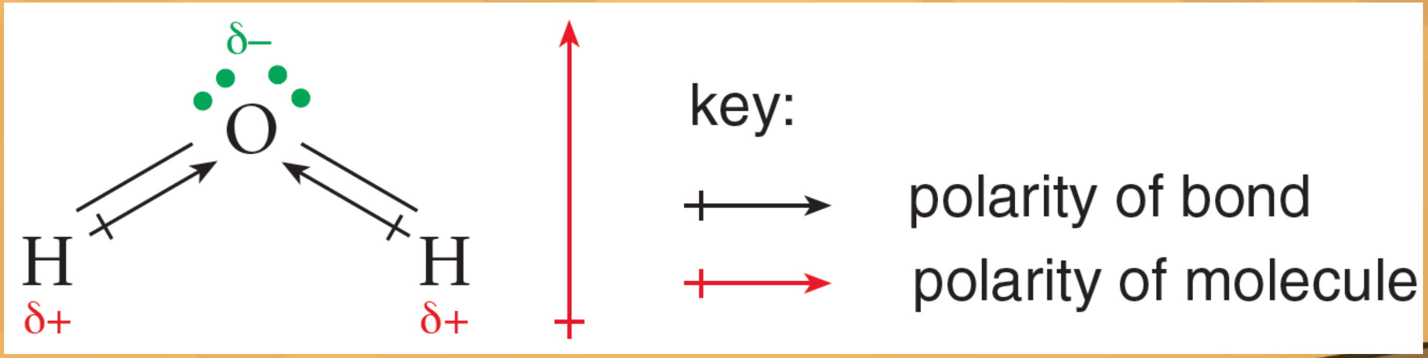
24.4 Polarity of molecules (p.26)

- Condiser NH_3.
- N is more electronegative than H. So each N–H bond is polar.
- NH_3 molecule is trigonal pyramidal. The three N–H bonds are not on the same plane.
- The polarities do not cancel out and so the molecule is polar.



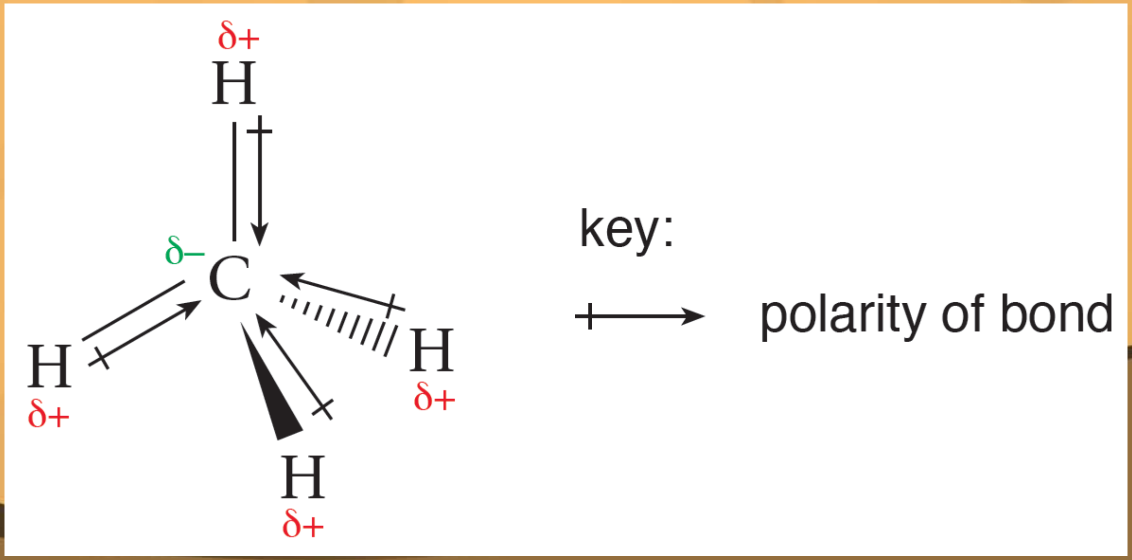
24.4 Polarity of molecules (p.26)

- Condiser H_2O.
- O is more electronegative than H. So each O–H bond is polar.
- H_2O molecule is bent. O is opposite to the other two H atoms.
- The polarities of the two O–H bonds do not cancel out and so the molecule is polar.



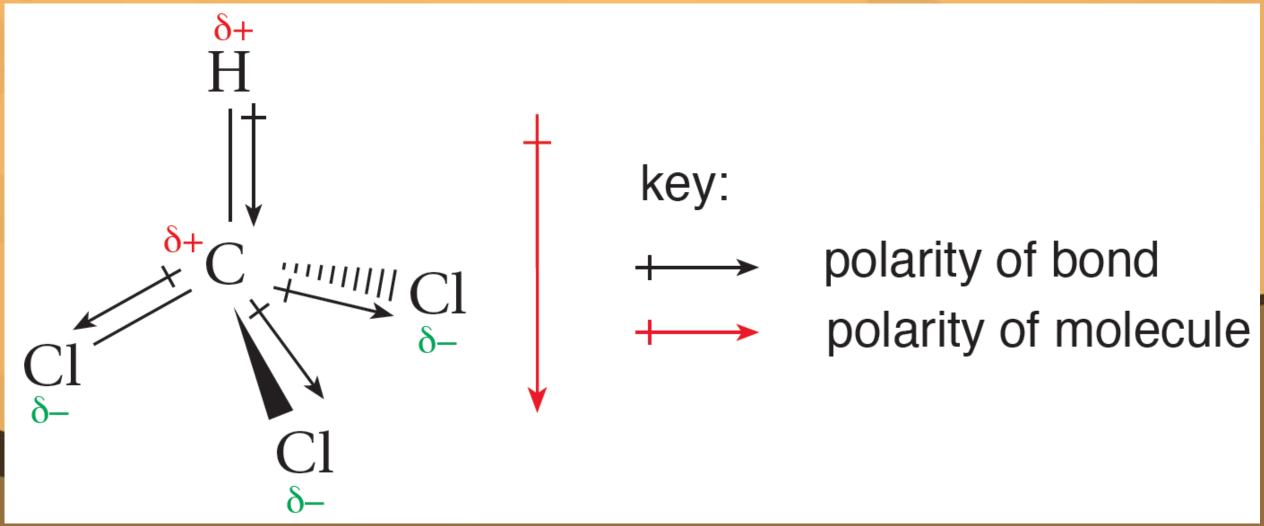
24.4 Polarity of molecules (p.26)

- Condiser CH_4.
- C is more electronegative than H. So each C–H bond is polar.
- CH_4 molecule is tetrahedral. The four polar C–H bonds are symmetrically arranged around C.
- The polarities of the four C–H bonds cancel each other out and so the molecule is non-polar.



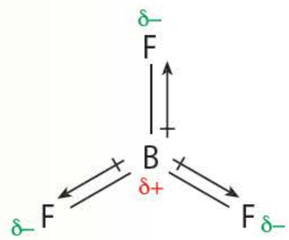
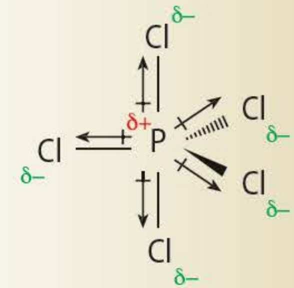
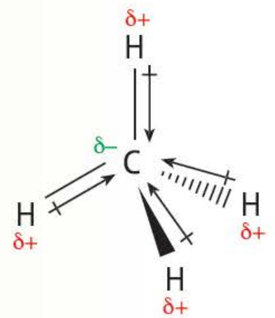
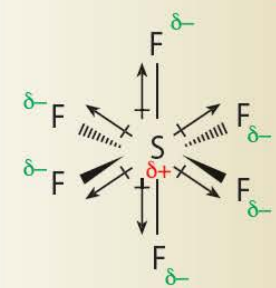
24.4 Polarity of molecules (p.26)

- Condiser CCl_3.
- Cl is more electronegative than C. So each C–Cl bond is polar. C is more electronegative than H. So the C–H bond is polar.
- CCl_3 molecule is tetrahedral. However, the polarity of a C–Cl bond is different from the polarity of a C–H bond.
- Hence the polarities of the bonds do not cancel out. So the molecule is polar.



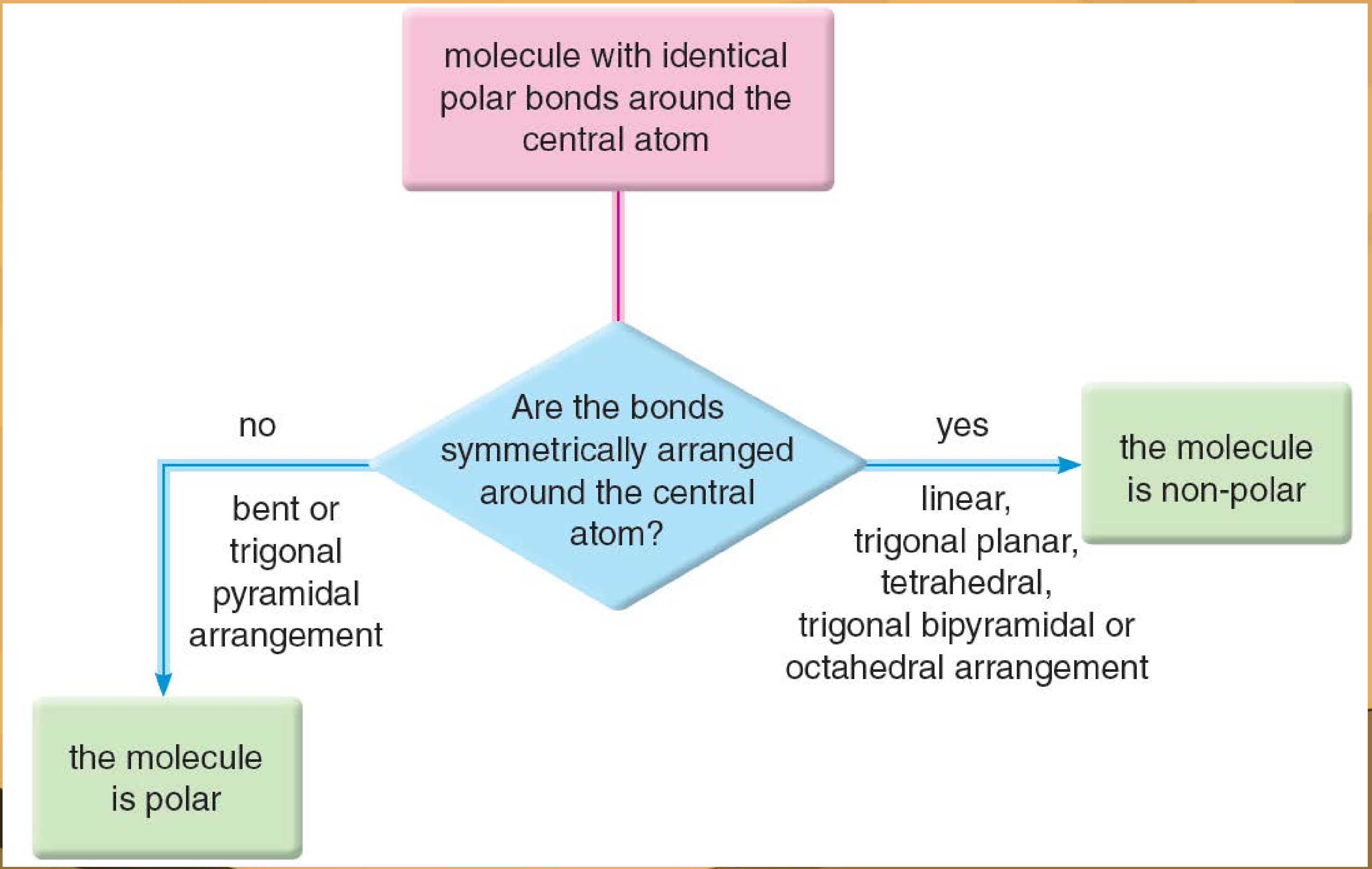
24.4 Polarity of molecules (p.26)

Identifying non-polar and polar molecules

Molecule	Shape	Polarities of bonds	Molecule	Shape	Polarities of bonds
Boron trifluoride (BF_3)	trigonal planar		Phosphorus pentachloride (PCl_5)	trigonal bipyramidal	
Methane (CH_4)	tetrahedral		Sulphur hexafluoride (SF_6)	octahedral	

DSE 2012 Paper 1A Q5
DSE 2017 Paper 1A Q12

24.4 Polarity of molecules (p.26)



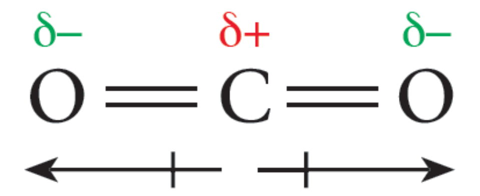
24.4 Polarity of molecules (p.26)

Q (Example 24.1)

Decide whether each of the molecules is polar.
a) CO_2
b) OF_2

A

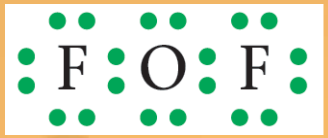
a) Oxygen is more electronegative than carbon. So each C=O bond in the molecule is polar.
Carbon dioxide has a linear shape. The two polar C=O bonds are symmetrically arranged on opposite sides of the carbon atom.
The polarities of the two C=O bonds cancel each other out.



So the molecule is non-polar.

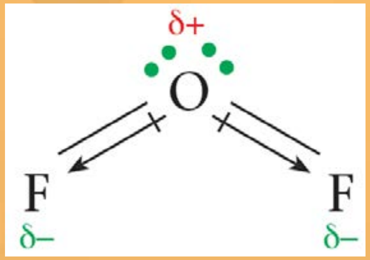
24.4 Polarity of molecules (p.26)

b) The electron diagram of OF_2 is shown below (showing electrons in the outermost shells only):



There are two bond pairs and two lone pairs in the outermost shell of the central oxygen atom. As there are only two fluorine atoms bonded to the oxygen atom, the shape of the molecule is bent or V-shaped.

Fluorine is more electronegative than oxygen. So each O–F bond in the molecule is polar.

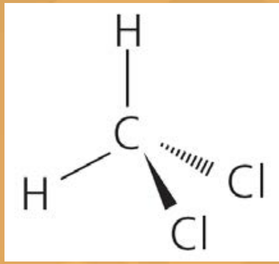


The molecule is not 'entirely' symmetrical and the polarities of the two O–F bonds do not cancel out. So the molecule is polar.

24.4 Polarity of molecules (p.26)

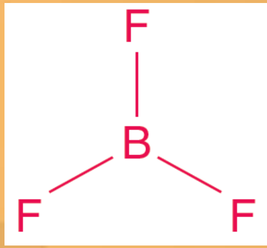
Practice 24.2

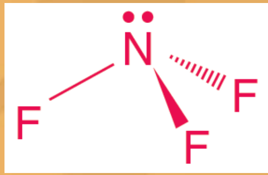
1 The three-dimensional structure of CH_2Cl_2 is shown. Identify, with explanation, the polar bond(s) in CH_2Cl_2.



C–H and C–Cl bonds are polar.
C and H / C and Cl have different electronegativities.

2 Both BF_3 and NF_3 exist as simple molecules.
For each of these molecules,
a) draw its three-dimensional structure;
b) explain whether or not it is polar.





The BF_3 molecule is non-polar.
The three polar B–F bonds are symmetrically arranged on the same plane around the central boron atom. The polarities of the three B–F bonds cancel each other out.
The NF_3 molecule is polar.
The central nitrogen atom has a lone pair in its outermost shell.
The three polar N–H bonds are not on the same plane. The polarities of the three N–H bonds do not cancel out.

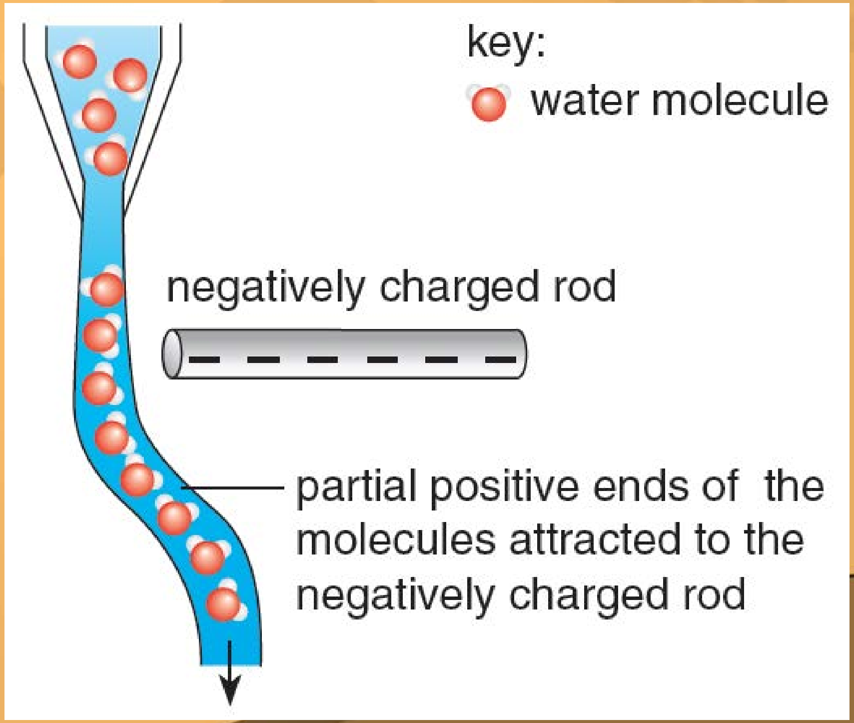
24.5 Testing whether or not a liquid is polar (p.32)

- A polar liquid is a liquid consisting of polar molecules.
- A simple test to find whether or not a liquid is polar is to run a stream of the liquid from a burette and bring a rod with *electrostatic charge* close to the liquid stream.
- A polar liquid will be deflected from its vertical path towards the charged rod whereas a non-polar liquid will not be deflected. The greater the polarity of the liquid, the greater the deflection is (under the same experimental conditions).

Testing liquids to find out if their molecules are polar *Ref.*

24.5 Testing whether or not a liquid is polar (p.32)

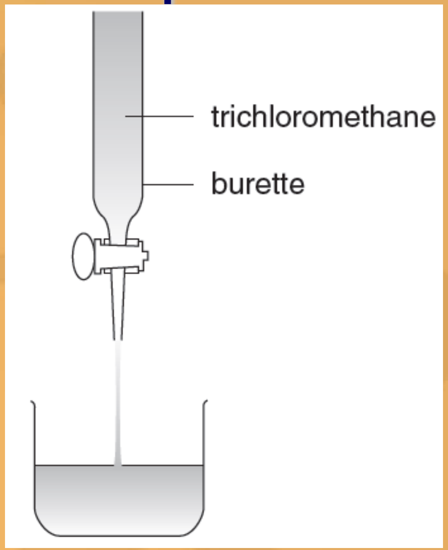
- When the negatively charged rod is brought close to the water stream, partial positive ends of the water molecules are attracted towards the rod.



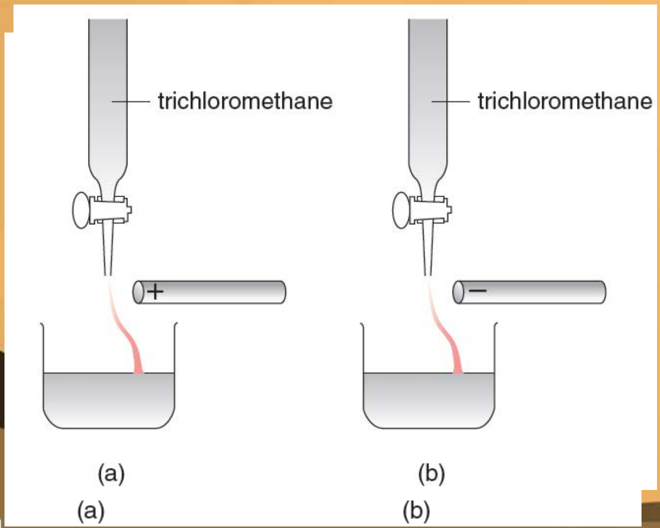
24.5 Testing whether or not a liquid is polar (p.32)

Practice 24.3

A jet of trichloromethane ($CHCl_3$) is run from a burette as shown.



A positively charged rod and a negatively charged rod are placed near the jet of trichloromethane as shown in diagrams (a) and (b). Show and explain your expected observation in each case.



A trichloromethane molecule is polar.
The jet of trichloromethane is deflected by the positively charged rod.
The negative ends of the molecules are attracted towards the rod.
The jet of trichloromethane is deflected by the negatively charged rod.
The positive ends of the molecules are attracted towards the rod.

24.6 Forces of attraction between non-polar molecules (p.34)

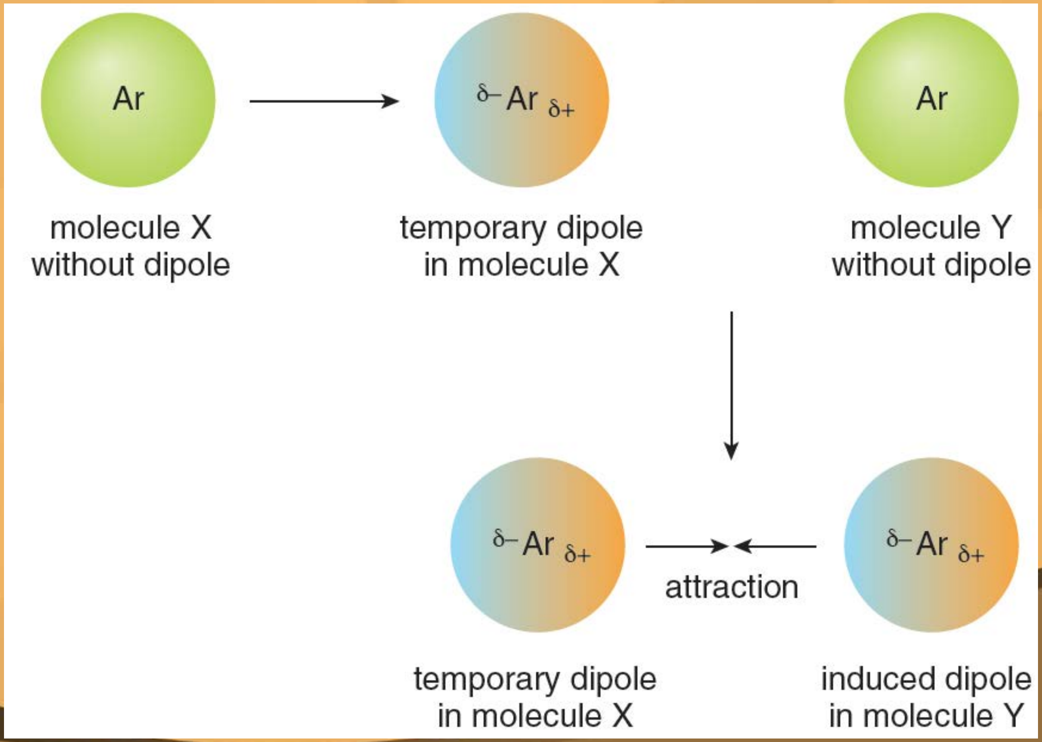
- Substances composed of non-polar molecules, such as halogens and noble gases can all be liquified and then solidified by cooling.
- This observation suggests that there must be weak forces of attraction between their molecules, holding them together in liquid and solid states.
- These are **intermolecular forces (分子間引力)**; 'intermolecular' means 'between molecules'. So how do such forces of attraction arise?

24.6 Forces of attraction between non-polar molecules (p.34)

- It is easier to describe the interaction between argon molecules by considering two neighbouring molecules X and Y.
- Each molecule as a whole is neutral but the electrons are constantly moving.
- If at any instant there are more electrons on one side of molecule X than the other, one end of the molecule has more negative charge than the other. A **temporary dipole (暫時偶極)** is set up. A dipole is a separation of opposite electrical charges.

24.6 Forces of attraction between non-polar molecules (p.34)

- The temporary dipole in molecule X affects the electron distribution in the neighbouring molecule Y, resulting in the formation of another dipole. This process is termed *induction*, and the newly formed dipole is an **induced dipole (誘發偶極)**.



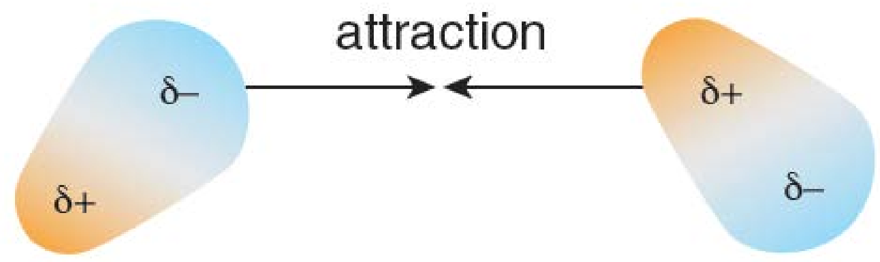
24.6 Forces of attraction between non-polar molecules (p.34)

- As a result, there are forces of attraction between the δ+ end of the dipole in molecule X and the δ– end of the dipole in molecule Y.
- The induced dipole may induce further dipoles on neighbouring molecules, which then attract one another.
- These dipoles are forming and disappearing continuously due to electron movement. The forces between these dipoles are always attractive forces.
- Attractive forces between non-polar molecules are termed **van der Waals' forces (范德華力)**. Such forces are present between all molecules in substances in solid and liquid states.

24.7 Forces of attraction between polar molecules (p.35)

- Because of the difference in electronegativities between hydrogen and chlorine, hydrogen chloride molecules are polar.
- In liquid hydrogen chloride, the molecules are constantly moving. Sometimes, the moleucles line up so that the δ+ end of one molecule is next to the δ– end of another molecule, causing attraction.
- Sometimes, the δ+ ends or δ– ends of two molecules are next to each other, causing repulsion.

24.7 Forces of attraction between polar molecules (p.35)



(a) Attraction arises when opposite partial charges are next to each other

(b) Repulsion arises when like partial charges are next to each other

24.7 Forces of attraction between polar molecules (p.35)

- Overall, there are more attractions than repulsions between the molecules. This is the reason why the molecules stay together in the liquid state.

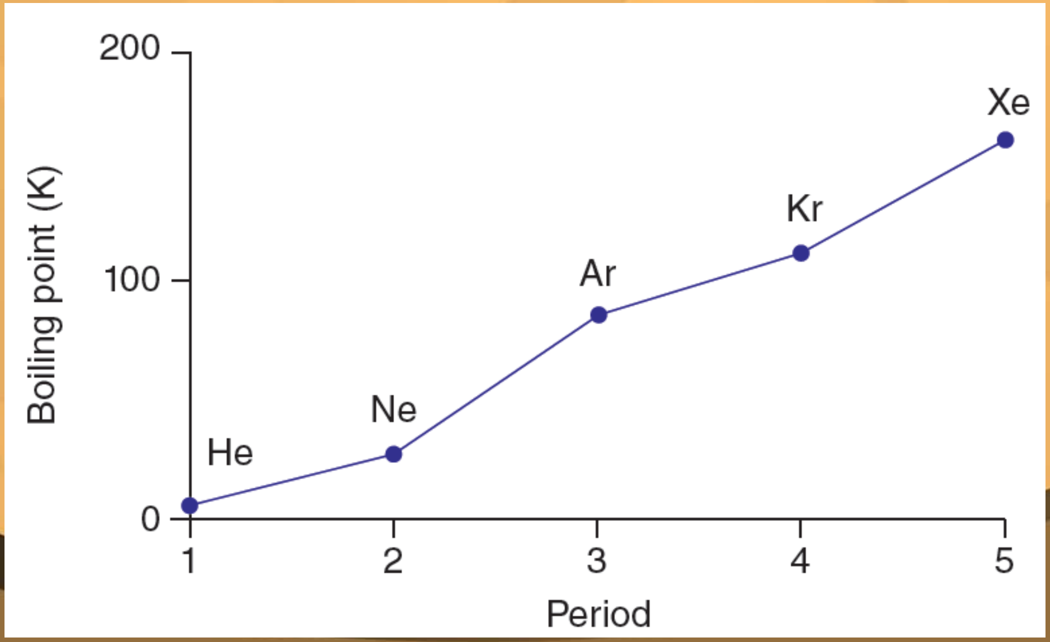
Van der Waals' forces is the collective name given to the attractive forces between molecules, including
- **attractive forces between non-polar molecules; and**
- **attractive forces between polar molecules.**

- When a molecular substance is in solid or liquid state, van der Waals' forces are present between the molecules. Once these forces are overcome, the molecules move apart from each other and the substance becomes a gas.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

- The strength of van der Waals' forces is affected by two factors:
  - molecular size (or the number of electrons per molecule);
  - molecular shape.

Molecular size



The boiling points of noble gases increase down Group 0.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

- Down the group, the outermost electrons are further away from the nucleus and hence less tightly bound.
  - Consequently, the electron distribution in a molecule can be distorted increasingly easily.
  - This favours the occurrence of temporary dipole and the creation of induced dipoles, resulting in stronger van der Waals' forces down the group.
  - More heat is needed to separate the molecules, so the boiling points of noble gases increase down the group.
- For substances with small non-polar molecules such as Cl_2, CO_2, CH_4 and CCl_4, the van der Waals' forces are very weak. So these substances tend to be gases or volatile liquids.

24.8 Factors affecting the strength of van der Waals’ forces (p.36)

Q (Example 24.2)

Suggest why, under room temperature and pressure, carbon dioxide (CO_2) is a gas but carbon disulphide (CS_2) is a liquid.

A

The intermolecular forces between CO_2 or CS_2 molecules are van der Waals' forces.

As CS_2 has a greater molecular size than CO_2, the van der Waals' forces between CS_2 molecules are stronger than those between CO_2 molecules.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

Practice 24.4

The table below lists the boiling points of three halogens.

Halogen	Boiling point (K)
Fluorine (F_2)	83
Chlorine (Cl_2)	239
Bromine (Br_2)	331

a) Name the type of forces between halogen molecules.
Van der Waals' forces

b) Describe the trend in boiling points of these halogens down the group.
The boiling points of the halogens increase down the group.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

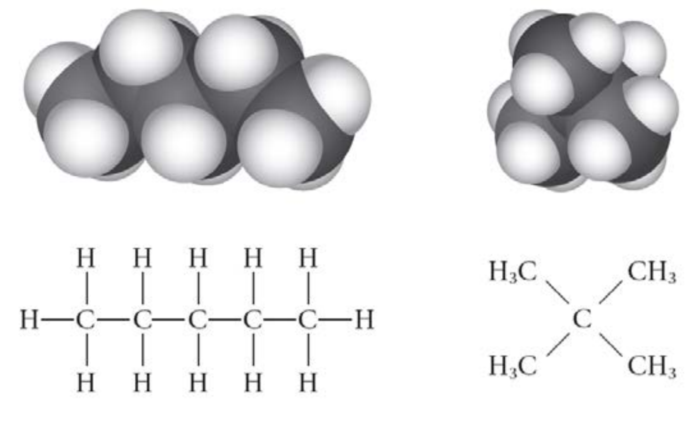
c) Explain the trend in these boiling points.

The molecular size / number of electrons per molecule increases from fluorine to bromine.
Thus, the strength of van der Waals' forces among halogen molecules increases from fluorine to bromine.
More heat is needed to separate the molecules, so the boiling points of halogens increase down the group.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

Molecular shape

- Molecular shape is also an important factor in determining the strength of van der Waals' forces.
- The pentane molecule is longer and somewhat spreadout. The molecule has a larger surface area, allowing a greater area of contact with neighbouring molecules.

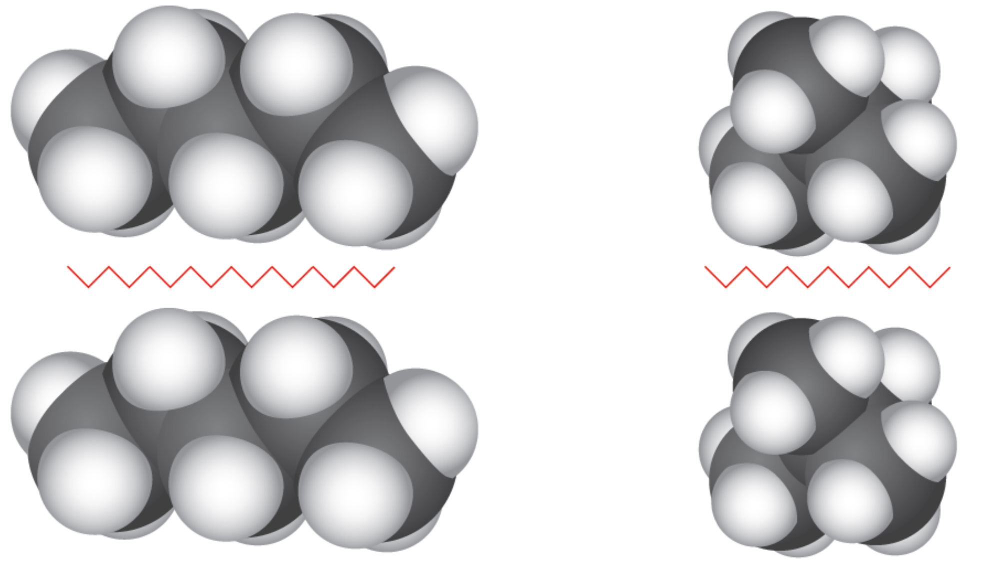


(a) (b)

- In contrast, the 2,2-dimethylpropane molecule is more compact, adopting a roughly spherical shape. The molecule has a smaller surface area for coming into contact with neighbouring molecules.

24.8 Factors affecting the strength of van der Waals' forces (p.36)

- The van der Waals' forces between pentane molecules are thus stronger. As a result, the boiling point of pentane is higher than that of 2,2-dimethylpropane (10 °C).



(a) pentane

(b) 2,2-dimethylpropane

key:

contact area between neighbouring molecules

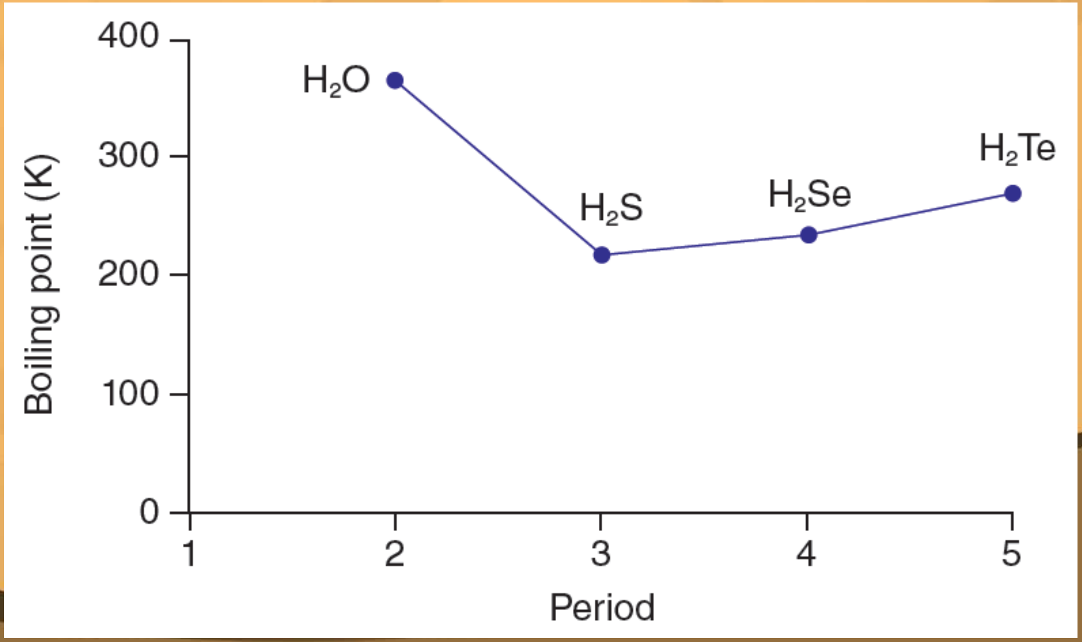
24.8 Factors affecting the strength of van der Waals' forces (p.36)

Van der Waals' forces increases with:

- **increasing molecular size (or increasing number of electrons per molecule); and**
- **increasing surface area of a molecule (allowing a greater contact area with neighbouring molecules).**

24.9 Hydrogen bonds (p.40)

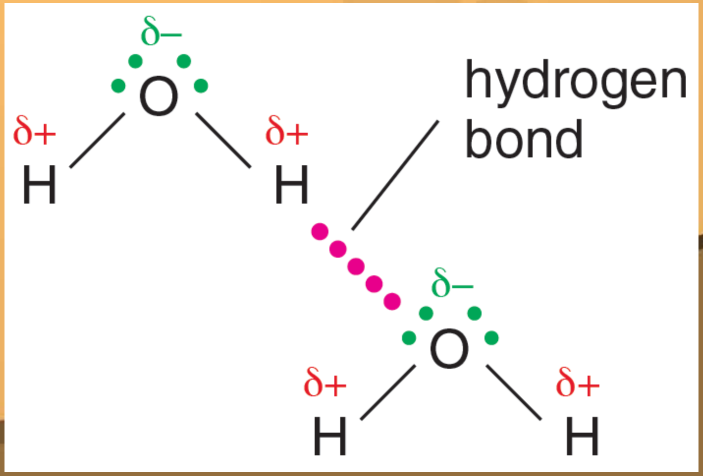
- The figure below shows the boiling points of the hydrides of Group VI elements. The steady increase in boiling points from H_2S to H_2Te is the result of the increasing molecular size, which in turn results in an increase in van der Waals' forces.
- However, water (H_2O) has a much higher boiling point than would be expected from its molecular size.



24.9 Hydrogen bonds (p.40)

- This suggests that there must be intermolecular forces other than van der Waals' forces between water molecules, and the intermolecular forces are stronger than van der Waals' forces.

- a) An oxygen atom is highly electronegative.
  b) In a water molecule, it attracts the shared electrons in each O–H bond strongly towards itself, making the O–H bond exceptionally highly polar.
  c) The hydrogen atom carries an unusually high partial positive charge.



24.9 Hydrogen bonds (p.40)

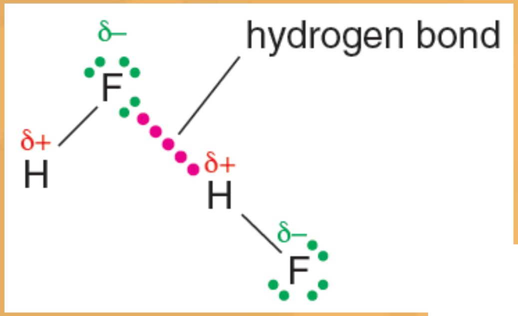
- Such a hydrogen atom is strongly attracted to the lone pair of electrons on the oxygen atom of an adjacent water molecule. This strong intermolecular force is called a **hydrogen bond (氫鍵)**。

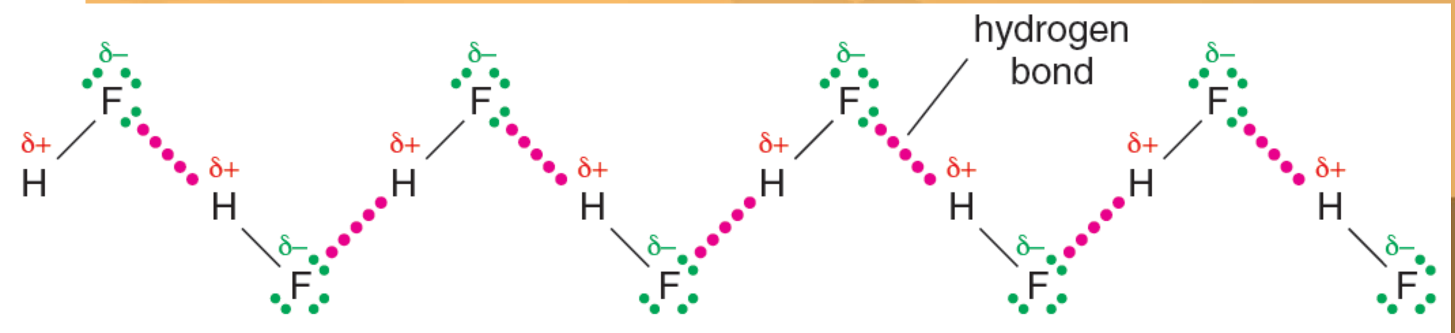
A hydrogen bond is the strong attractive force between a hydrogen atom attached to a highly electronegative atom and the lone pair on another electronegative atom.

24.9 Hydrogen bonds (p.40)

Hydrogen bonding in liquid hydrogen fluoride

- In liquid hydrogen fluoride, the hydrogen atom in one molecule forms a hydrogen bond with the lone pair of electrons on the fluorine atom of an adjacent molecule. The result is a zigzag chain of hydrogen fluoride molecules。

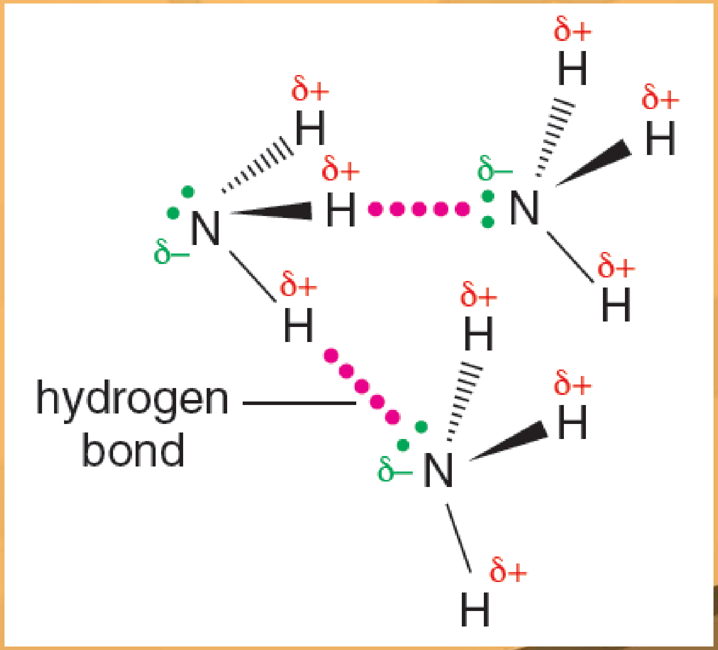




24.9 Hydrogen bonds (p.40)

Hydrogen bonding in liquid ammonia

- In liquid ammonia, the hydrogen atom in one molecule forms a hydrogen bond with the lone pair of electrons on the nitrogen atom of an adjacent molecule.



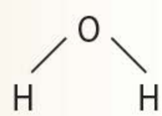
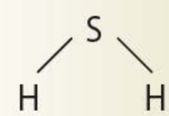
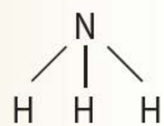
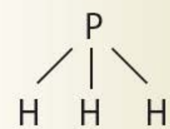
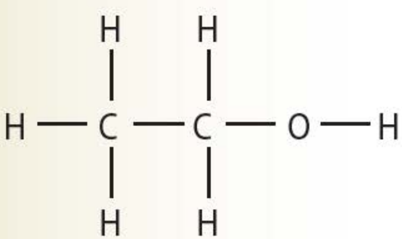
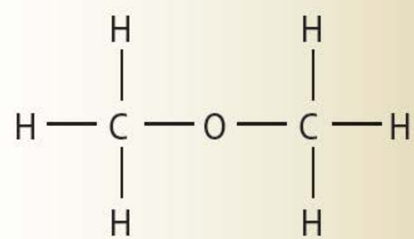
24.9 Hydrogen bonds (p.40)

For hydrogen bonding to occur between two molecules,

- one molecule should have a hydrogen atom covalently bonded to F, O or N (the three most electronegative elements); and
- an adjacent molecule should have an atom of F, O or N with an available lone pair of electrons.

24.9 Hydrogen bonds (p.40)

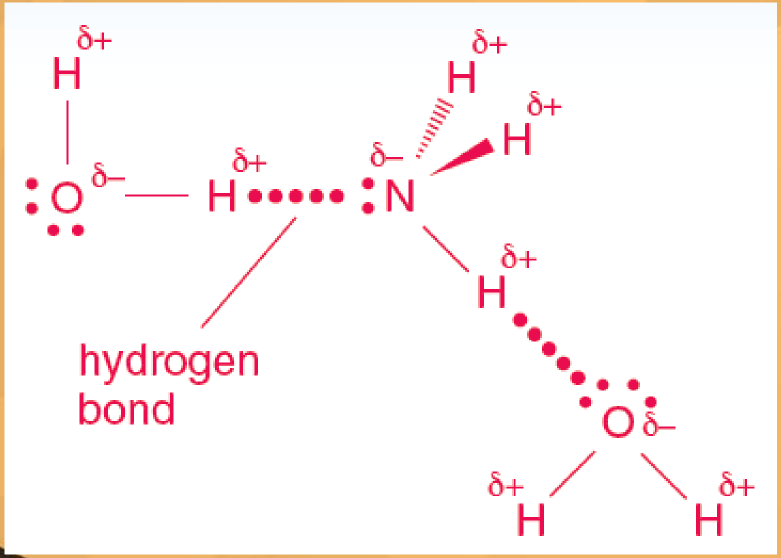
Table 24.2 Hydrogen bonding or no hydrogen bonding?

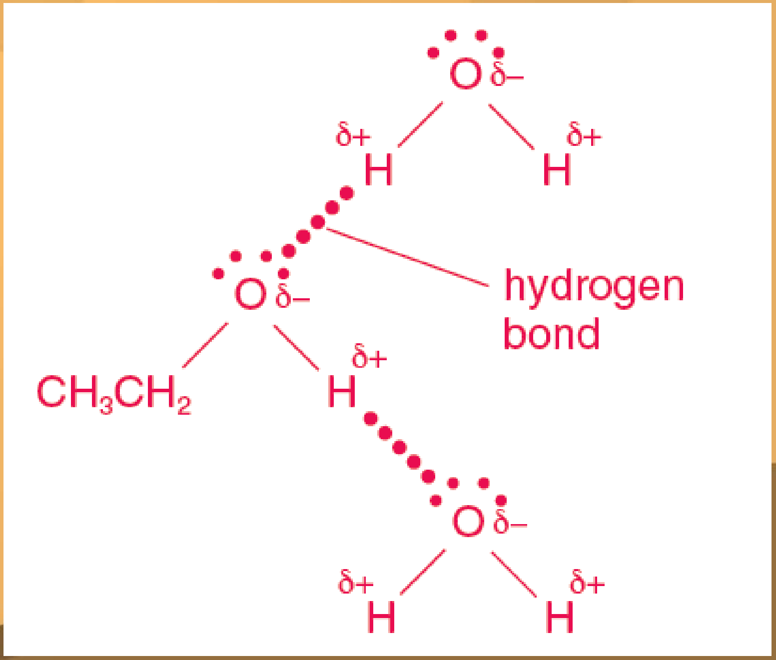
Hydrogen bonding between molecules		No hydrogen bonding between molecules	
hydrogen fluoride	H–F	hydrogen chloride	H–Cl
water		hydrogen sulphide	
ammonia		phosphine	
ethanol		methoxymethane	

24.9 Hydrogen bonds (p.40)

Practice 24.5

1 Hydrogen bonding can occur between different molecules in mixtures. Draw diagrams to show hydrogen bonding between water molecules and
a) ammonia molecules in a solution of ammonia (NH_3);
b) ethanol molecules in a solution of ethanol (CH_3CH_2OH).
Include partial charges and all lone pairs.





24.9 Hydrogen bonds (p.40)

2 The table below lists the boiling points of three substances.

Substance	Boiling point (K)
Hydrogen (H_2)	20
Fluorine (F_2)	83
Hydrogen fluoride (HF)	293

a) Explain, by comparing the forces involved, why fluorine has a higher boiling point than hydrogen.

Molecules of H_2 and F_2 are held together by van der Waals' forces.
The van der Waals' forces among F_2 molecules are stronger than those among H_2 molecules because F_2 has a larger size than H_2. / A F_2 molecule has more electrons than a H_2 molecule.

b) Explain, by comparing the forces involved, why hydrogen fluoride has a higher boiling point than fluorine.

Hydrogen bonding exists among HF molecules.
Hydrogen bonding is stronger than van der Waals' forces.

24.10 Relative strength of intermolecular attraction (p.43)

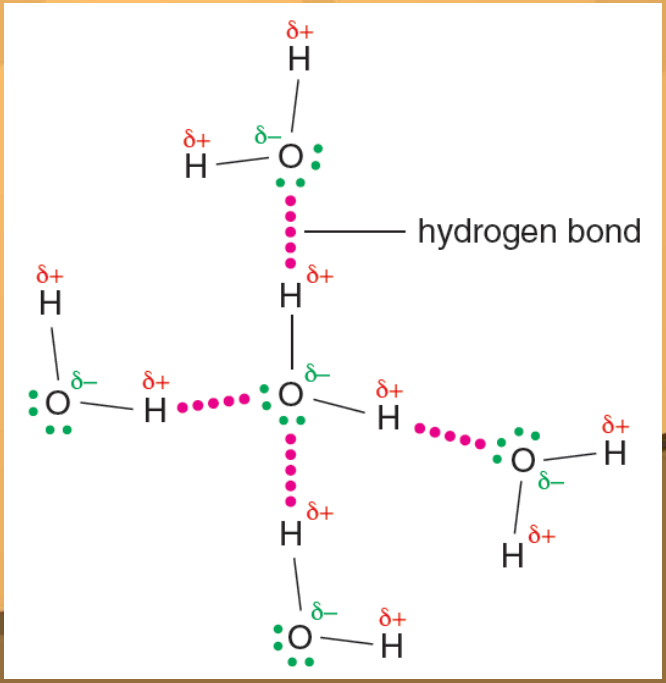
- Van der Waals' forces and hydrogen bonding are the two main types of intermolecular attraction.
- The strengths of these two types of attraction and covalent bonds are in the following order:

  van der Waals' forces < hydrogen bonds < covalent bonds

- Despite being weaker than covalent bonds, van der Waals' forces and hydrogen bonds play an important part in determining the physical properties of many substances.

24.11 Effect of hydrogen bonding on the properties of water (p.44)

- The properties of water are in several ways rather different from those of other simple molecular substances.
- Unlike the molecules of other compounds which are capable of forming hydrogen bonds, a water molecule has two lone pairs of electrons on the oxygen atom together with two hydrogen atoms.



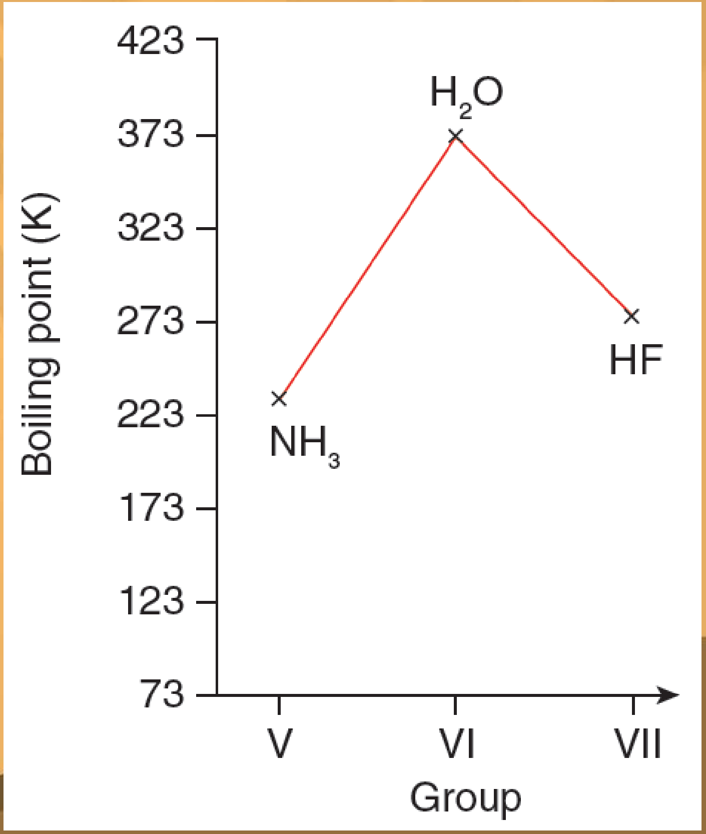
24.11 Effect of hydrogen bonding on the properties of water (p.44)

- This means that it can form two hydrogen bonds per molecule on average, whereas ammonia or hydrogen fluoride can only form one hydrogen bond per molecule.
- This means that the hydrogen bonding in water is more extensive and results in some unexpected properties of water.

24.11 Effect of hydrogen bonding on the properties of water (p.44)

Boiling point of water

- The boiling point of water is much higher than those of ammonia and hydrogen fluoride.

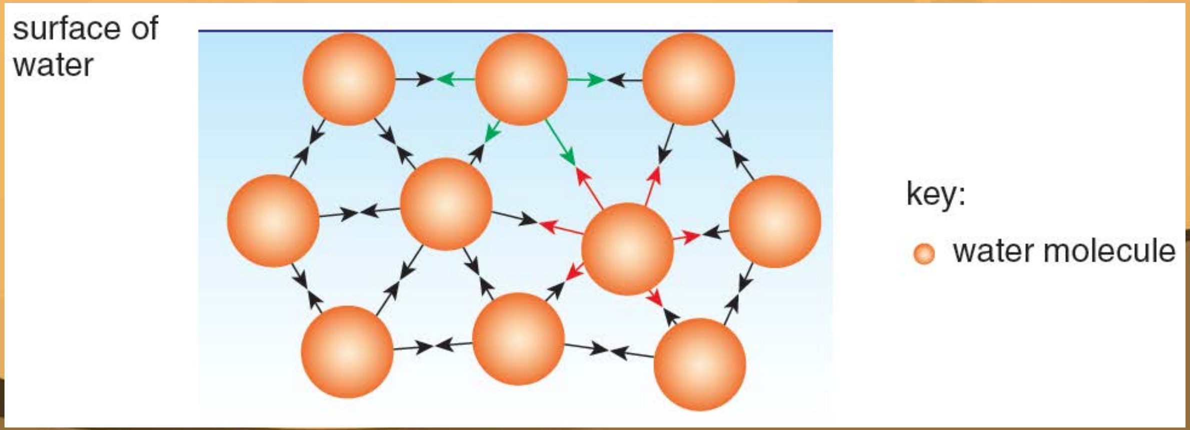


Comparing the surface tension of water and propanone *Ref.*

24.11 Effect of hydrogen bonding on the properties of water (p.44)

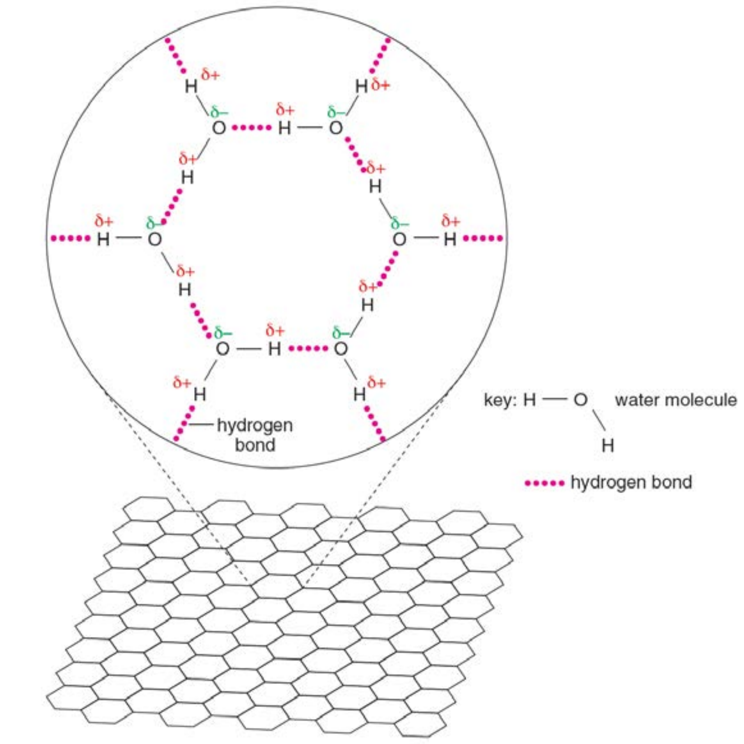
Surface tension of water

- A molecule within a liquid experiences attractions from other molecules in all directions. A molecule at the surface of the water only experiences attractions from the side and below.
- Unbalanced forces on molecules at the surface pull the molecules back into the liquid, causing the surface to tighten like a thin elastic sheet. **Surface tension (表面張力)** results.



24.11 Effect of hydrogen bonding on the properties of water (p.44)

- Liquid water has a high surface tension. Through hydrogen bonding, the molecules form very temporary hexagonal arrays across the surface of water, allowing some objects which you might expect to sink in water, such as pond skaters, to float on water.



24.11 Effect of hydrogen bonding on the properties of water (p.44)

Viscosity of water

- **Viscosity (黏度)** is a quantity that describes a liquid's resistance to flow. It is related to the ease with which molecules can move past each other.

- The viscosity of a liquid depends on:
  - the strength of intermolecular forces;
  - the tendency of molecules to become *entangled* with each other.

- Hydrogen bonding reduces the ability of water molecules to move past each other. Hence the viscosity of water is high.

24.11 Effect of hydrogen bonding on the properties of water (p.44)

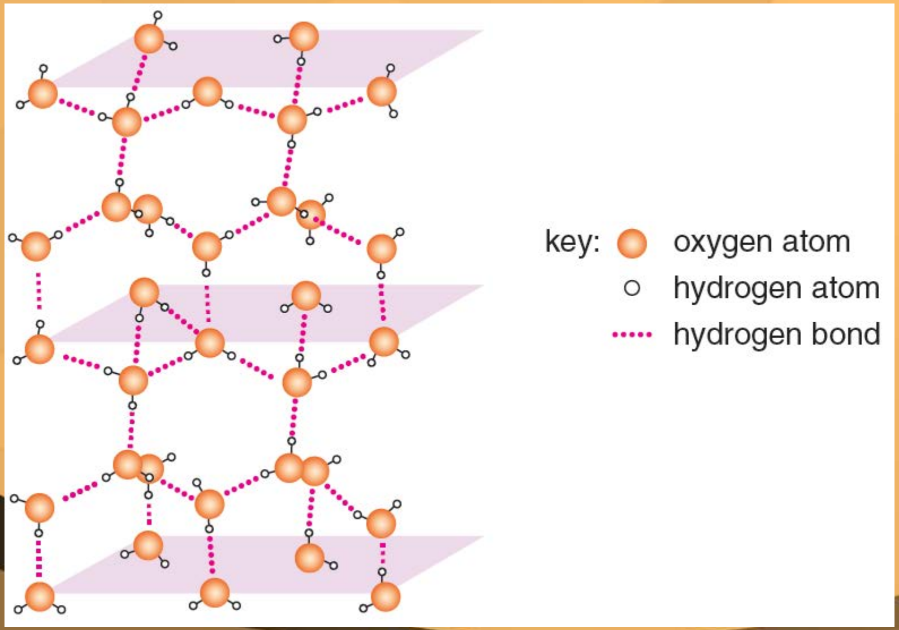
Ice being less dense than water

- Of almost all materials, the solid form is denser than the liquid form.
- Water has the unusual property of its solid form being less dense than its liquid form. Why is it?
- In ice, each water molecule is surrounded tetrahedrally by four other water molecules joined by intermolecular hydrogen bonding. Ice has an ‘open’ structure with water molecules held apart.

Building a model of ice crystal

24.11 Effect of hydrogen bonding on the properties of water (p.44)

- When ice melts, the water molecules in liquid water have relative motion and the open structure of ice collapses. The molecules become more closely packed, resulting in a decrease in volume. Hence the density of ice is lower than that of liquid water at 0 °C.



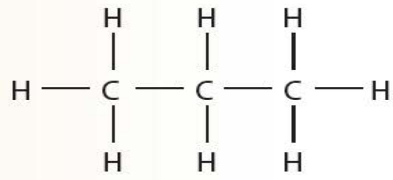
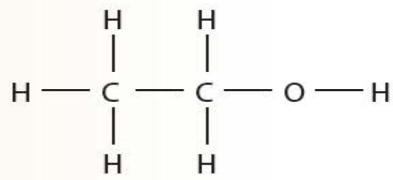
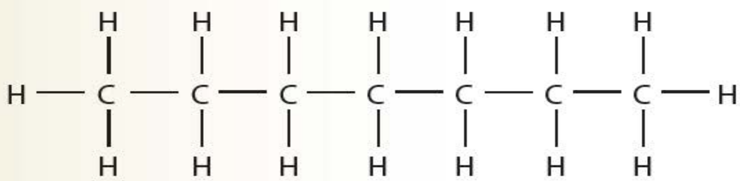
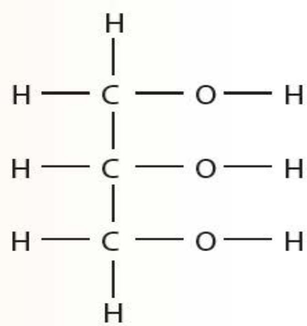
24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Boiling point

- From the table below, you can see that the compounds with hydrogen bonding have higher boiling points than those with similar molecular masses and no hydrogen bonding. This is because more heat is needed to overcome the hydrogen bonds between the molecules during boiling.

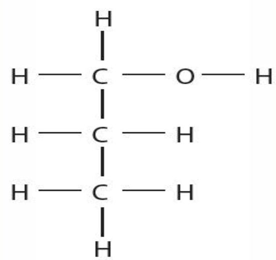
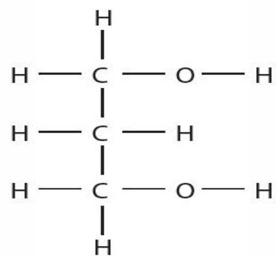
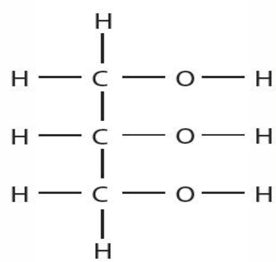
Determining the strength of the hydrogen bond between ethanol molecules

Table 24.3 Types of intermolecular forces in different substances and their boiling points

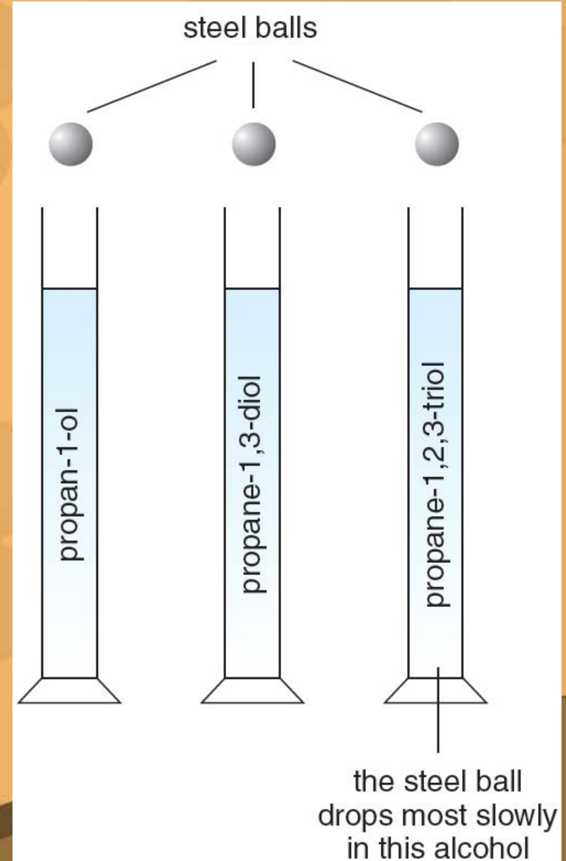
Substance	Structure	Relative molecular mass	Type(s) of intermolecular forces present		Boiling point (K)
			van der Waals' forces	hydrogen bonding	
Propane		44.0	✔		231
Ethanol		46.0	✔	✔	351
Heptane		100.0	✔		371
Propane-1,2,3-triol		92.0	✔	✔	563

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Viscosity

Alcohol	Structure	Relative viscosity at 25 °C
Propan-1-ol		*
Propane-1,3-diol		**
Propane-1,2,3-triol		****

(**** = highest, * = lowest)



24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

- The intermolecular attractions in the alcohols are predominantly hydrogen bonds.

-

	O–H groups per molecule	H-bonds per molecule
Propan-1-ol	1	1
Propane-1,3-diol	2	2
Propan-1,2,3-triol	3	3

Strength of intermolecular attractions:

propan-1-ol < propane-1,3-diol < propane-1,2,3-triol

- The viscosity of propane-1,2,3-diol is the highest because it has the strongest intermolecular attractions and its molecules tend to become the most entangled due to the shape.

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Q (Example 24.3)

It is known that hydrogen bonds exist between ethanol molecules.
An experiment was conducted to investigate the strength of the hydrogen bonds between ethanol molecules. 8.0 g of ethanol were mixed with 16.0 g of cyclohexane (excess). A temperature fall of 3.8 K was recorded.
The following data were provided:

Compound	Ethanol (C_2H_5OH)	Cyclohexane (C_6H_{12})
Molar mass (g mol^{-1})	46.0	84.0
Specific heat capacity (J g^{-1} K^{-1})	2.41	1.83

Amount of heat taken in by a substance
= mass x specific heat capacity x fall in temperature

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

a) Draw a diagram to show the hydrogen bond formed between two ethanol molecules. Include partial charges and all lone pairs.

b) What happened to the hydrogen bonds between ethanol molecules on mixing ethanol with cyclohexane?

c) Why was excess cyclohexane used?

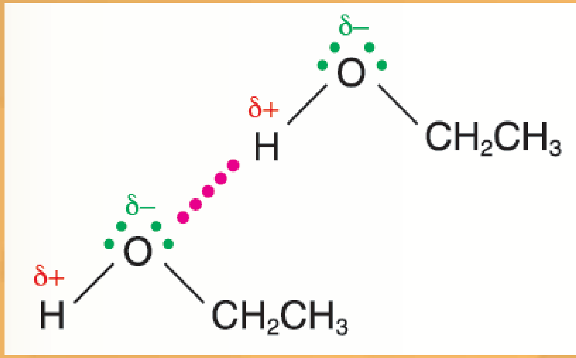
d) Calculate the amount of heat taken in by
 i) ethanol, and
 ii) cyclohexane
 in the experiment.

e) Calculate the strength of the hydrogen bond (in kJ mol^{-1}) between ethanol molecules.

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

A

a)



b) Hydrogen bonds between ethanol molecules were broken.

c) To ensure that all the hydrogen bonds between ethanol molecules were broken.

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

d) i) Amount of heat taken in by ethanol = 8.0 g x 2.41 J g^{-1} K^{-1} x 3.8 K
= 73.3 J

ii) Amount of heat taken in by cyclohexane = 16.0 g x 1.83 J g^{-1} K^{-1} x 3.8 K
= 111 J

e) Number of moles of ethanol used = $\frac{mass}{molar\ mass} = \frac{8.0\ g}{46.0\ g\ mol^{-1}}$ = 0.174 mol

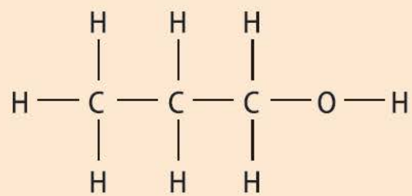
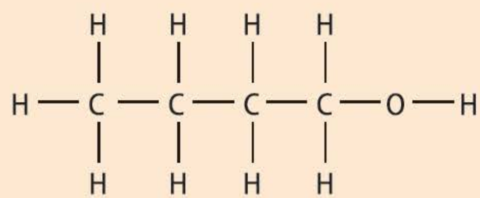
Strength of the hydrogen bond between ethanol molecules

= $\frac{(73.3+111)\ J}{0.174\ mol}$ = 1.06 kJ mol^{-1}

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Q (Example 24.4)

The table below lists the structures and boiling points of two alcohols.

Alcohol	Structure	Boiling point (°C)
Propan-1-ol		97
Butan-1-ol		118

Explain why the boiling point of butan-1-ol is higher than that of propan-1-ol.

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

A

Both alcohols have one O–H group per molecule that can take part in hydrogen bonding. Hydrogen bonds in the two alcohols are of comparable strengths.

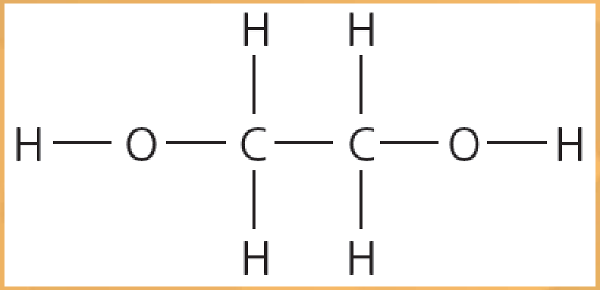
The boiling point difference is due to the difference in the strength of van der Waals' forces in the alcohols. Van der Waals' forces between butan-1-ol molecules are stronger than those between propan-1-ol molecules because of the larger size of a butan-1-ol molecule as compared with a propan-1-ol molecule.

Thus, the boiling point of butan-1-ol is higher than that of propan-1-ol.

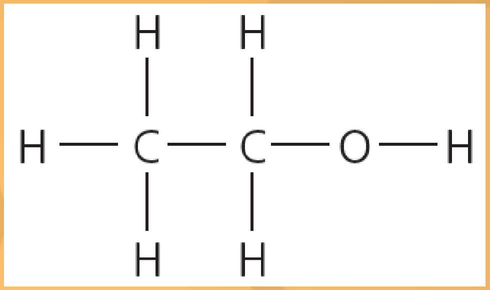
24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Practice 24.4

1 The structures of two alcohols are shown below.



ethane-1,2-diol



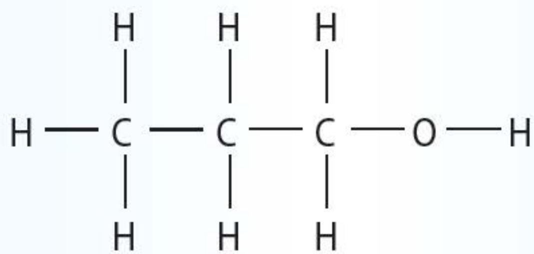
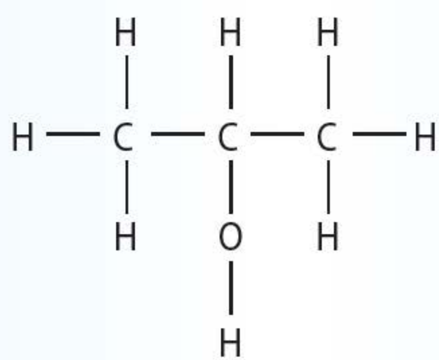
ethanol

Ethane-1,2-diol is much less volatile than ethanol. Suggest why.

Ethane1,2-diol has more O–H groups than ethanol.
Stronger hydrogen bonding exists among ethane-1,2-diol molecules.

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

2 The table below lists the structures and boiling points of two alcohols, A and B. Explain why the boiling point of A is higher than that of B.

Alcohol	Structure	Boiling point (°C)
A		97
B		82

24.12 Effect of hydrogen bonding on the properties of alcohols (p.48)

Both alcohols have one O–H group per molecule that can take part in hydrogen bonding. Hydrogen bonds in the two alcohols are of comparable strengths.

Molecule of A is longer and somewhat spreadout. The molecule has a larger surface area, allowing a greater contact with neighbouring molecules.

In contrast, molecule of B is more compact, adopting a roughly spherical shape. The molecule has a smaller surface area for coming into contact with neighbouring molecules.

Thus, the van der Waals' forces among molecules of A are stronger than those among molecules of B.

Key terms (p.54)

polar	極性的	induced dipole	誘發偶極
electronegativity	電負性	van der Waals' force	范德華力
non-polar	非極性的	hydrogen bond	氫鍵
intermolecular force	分子間引力	surface tension	表面張力
temporary dipole	暫時偶極	viscosity	黏度

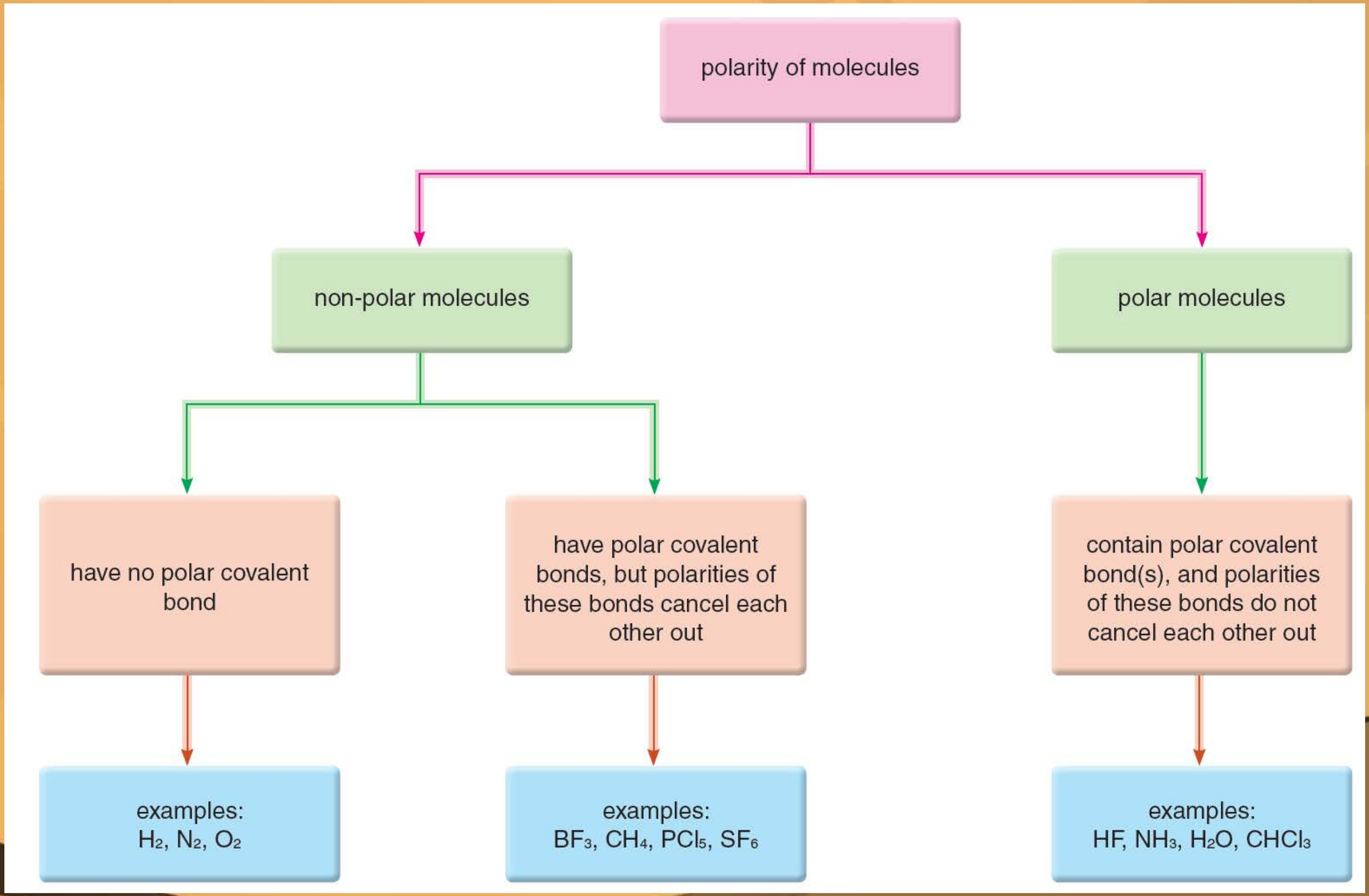
Summary (p.55)

1 Electronegativity of an element is a measure of the ability of an atom of the element to attract the bonding electrons in a covalent bond towards itself.

2 The electronegativity increases across a period and decreases down a group in the Periodic Table.

Summary (p.55)

3 The following diagram summarises the nature of non-polar and polar molecules.



Summary (p.55)

4 Van der Waals' force is the collective name given to the attractive forces between molecules, including
- attractive forces between non-polar molecules; and
- attractive forces between polar molecules.

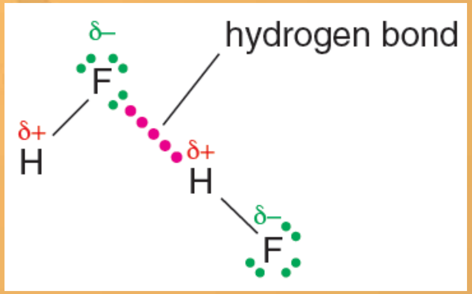
5 Van der Waals' forces increase with:
- increasing molecular size (or increasing number of electrons per molecule); and
- increasing surface area of a molecule (allowing a greater contact area with neighbouring molecules).

Summary (p.55)

6 a) A hydrogen bond is the strong attractive force between a hydrogen atom attached to a highly electronegative atom and the lone pair on another electronegative atom.

b) Molecules should fulfill the following criteria for hydrogen bonding to occur between two molecules:

- one molecule having a hydrogen atom covalently bonded to an atom of fluorine, oxygen or nitrogen (the three most electronegative elements);
- an adjacent molecule having an atom of fluorine, oxygen or nitrogen with an available lone pair of electrons.



c) Hydrogen bonds are stronger than van der Waals' forces.

Summary (p.55)

7 Intermolecular forces affect several physical properties of liquids.

a) A liquid with stronger intermolecular forces have a higher boiling point.

b) A liquid with stronger intermolecular forces have a higher surface tension.

c) The viscosity of a liquid is a measure of a liquid's resistance to flow. The viscosity of a liquid depends on:

- the strength of intermolecular forces;
- the tendency of molecules to become entangled with each other.

8 In ice, each water molecule is surrounded tetrahedrally by four other water molecules joined by intermolecular hydrogen bonding. Ice has an 'open' structure with water molecules held apart. The open structure of ice accounts for the fact that ice is less dense than water at 0 °C.

Unit Exercise (p.57)

Note: Questions are rated according to ascending level of difficulty (from 1 to 5):

 **question targeted at level 3 and above;**

 **question targeted at level 4 and above;**

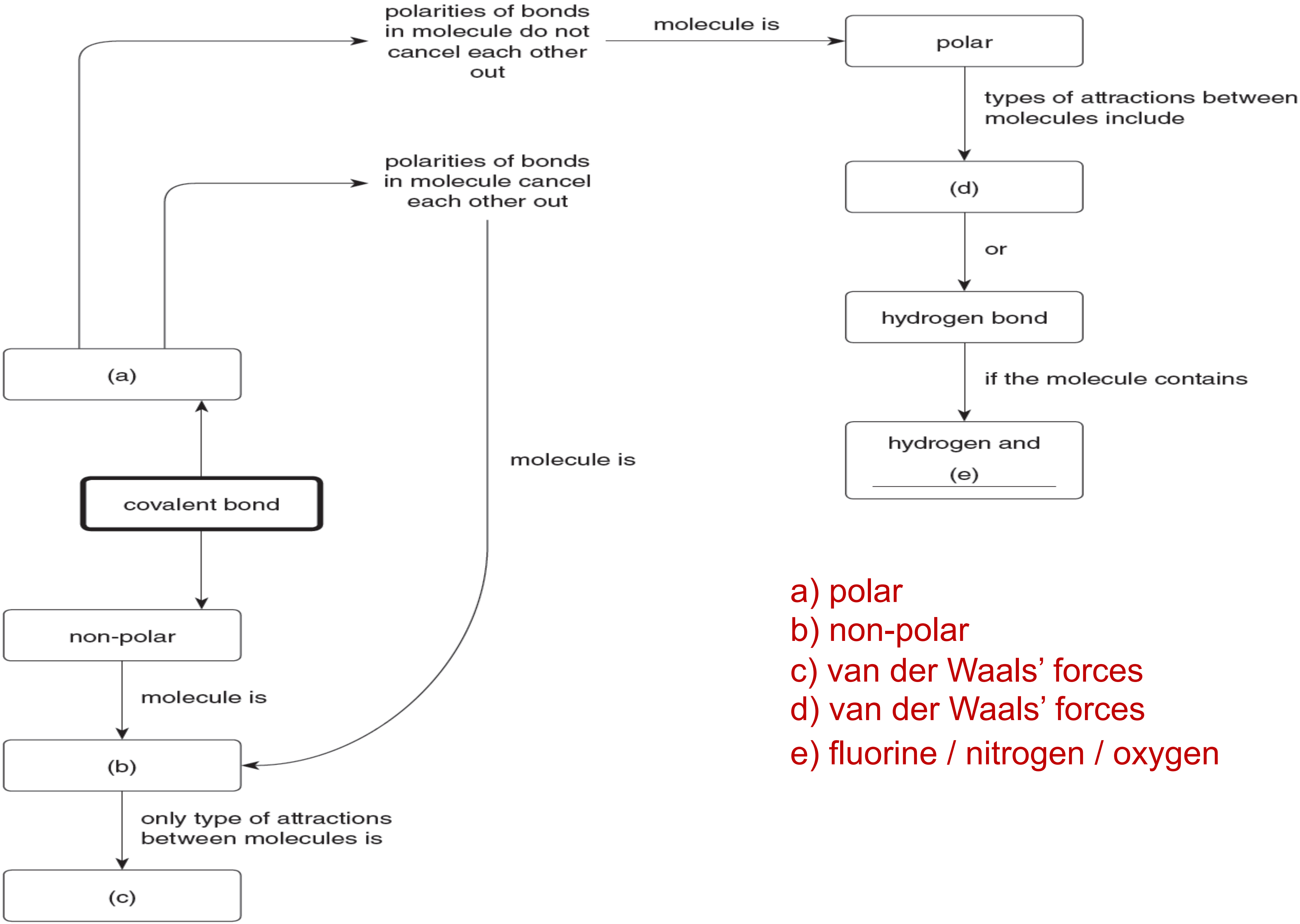
 **question targeted at level 5.**

‘ * ’ indicates 1 mark is given for effective communication.

Unit Exercise (p.57)

PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.



a) polar
b) non-polar
c) van der Waals' forces
d) van der Waals' forces
e) fluorine / nitrogen / oxygen

Unit Exercise (p.57)

PART II MULTIPLE CHOICE QUESTIONS

2 Which of the following bonds is the least polar?

A H–F
B H–Cl
C N–H
D C–H

Answer: D

Unit Exercise (p.57)

3 Which of the following molecules is non-polar?

A H_2S
B $CHCl_3$
C SF_4
D SiF_4

(HKDSE, Paper 1A, 2014, 2)

Answer: D

Explanation:
SiF_4 molecule has a tetrahedral shape. The four polar Si–F bonds are symmetrically arranged around the central silicon atom. The polarities of the four Si–F bonds cancel each other out. So the molecule is non-polar.

Unit Exercise (p.57)

4 Which of the following compounds has the highest boiling point?

A HF
B HCl
C PH_3
D H_2Se

(HKDSE, Paper 1A, 2016, 16)

Answer: A

Unit Exercise (p.57)

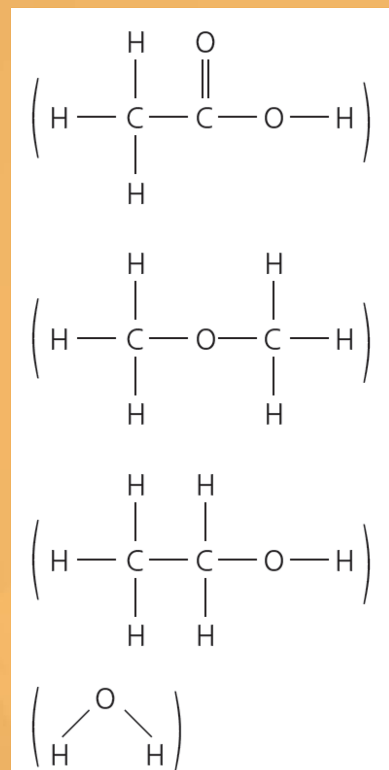
5 Which of the following substances does NOT have intermolecular hydrogen bonds?

A Ethanoic acid

B Methoxymethane

C Ethanol

D Water



Answer: B

Explanation:

The hydrogen atoms in methoxymethane do NOT covalently bond to fluorine / oxygen / nitrogen atoms. Thus, methoxymethane does NOT have intermolecular hydrogen bonds.

Unit Exercise (p.57)

6 Along the series of the Group V hydrides (NH_3, PH_3 and AsH_3), the boiling temperatures

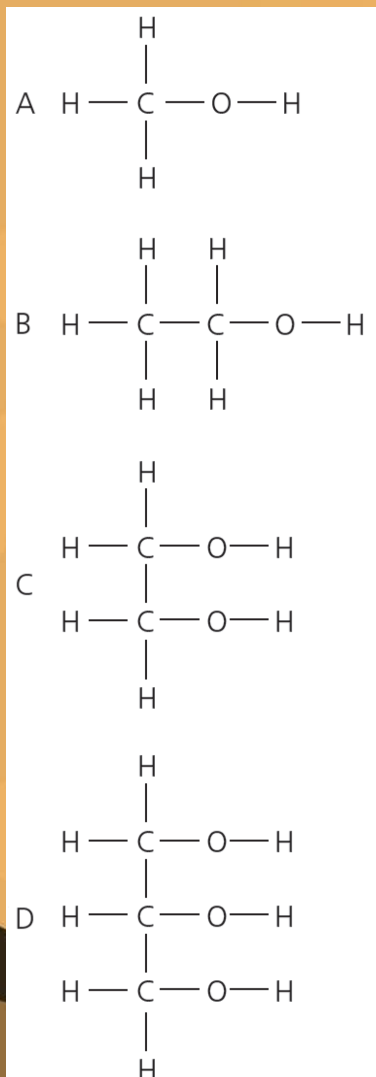
A decrease.
B decrease then increase.
C increase.
D increase then decrease.

(Edexcel Advanced Subsidiary GCE, Unit 2, Jun. 2016, 12)

Answer: B

Unit Exercise (p.57)

7 Which of the following alcohols is the most viscous?



Answer: D

Explanation:

D has three O–H groups per molecule. It forms the most extensive hydrogen bonds.

Unit Exercise (p.57)

8 Which of the following processes involves the cleavage of hydrogen bonds?

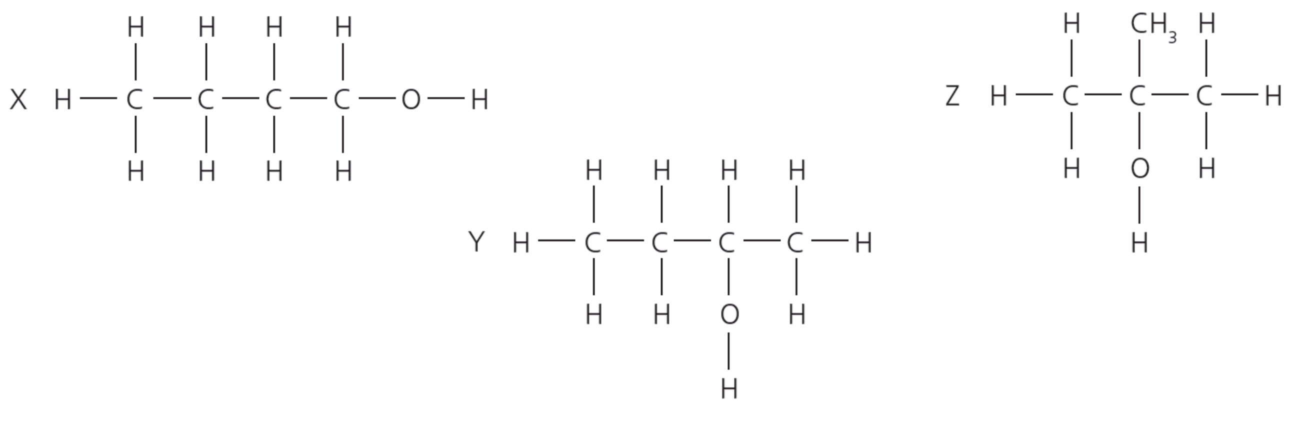
A Sublimation of dry ice
B Melting of ice
C Dissociation of hydrogen molecules
D Condensation of ammonia

Answer: B

Explanation:
In ice, each water molecule is joined to four other water molecules by hydrogen bonding. During the melting of ice, the structure of ice collapses due to the cleavage of some of the hydrogen bonds.

Unit Exercise (p.57)

9 Consider the alcohols below. They have the same molecular formula.



Which of the following sequences correctly lists the alcohols in the order of increasing boiling point?

A Y < Z < X
B Y < X < Z
C Z < X < Y
D Z < Y < X

Answer: D

Unit Exercise (p.57)

Explanation:
The three alcohols have one O–H group per molecule that can take part in hydrogen bonding. Hydrogen bonds in the three alcohols are of comparable strengths.

A molecule of X is the longest and somewhat spreadout. The molecule has the largest surface area, allowing the greatest contact with neighbouring molecules.

In contrast, a molecule of Z is the most compact, adopting a roughly spherical shape. The molecule has the smallest surface area for coming into contact with neighbouring molecules.

Thus, the strength of van der Waals' forces among molecules is in the order: Z < Y < X

Unit Exercise (p.57)

10 Which of the following molecules is / are non-polar?

(1) BF_3
(2) $CHCl_3$
(3) PCl_5

A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

Answer: C

Explanation:
(3) PCl_5 is non-polar because the five polar P–Cl bonds are symmetrically arranged around the central phosphorus atom. The polarities of the five P–Cl bonds cancel each other out.

Unit Exercise (p.57)

11 Which of the following statements is / are correct?

(1) The density of $H_2O(l)$ is lower than that of $H_2O(g)$.
(2) When ice changes to water, the open structure of ice collapses.
(3) When the temperature of water rises from 10 °C to 30 °C, the average distance between H_2O molecules increases.

A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

Answer: D

(HKDSE, Paper 1A, 2018, 17)

Unit Exercise (p.57)

12 P and Q are two liquids with similar relative molecular masses. Molecules of P attract each other by hydrogen bonds. Molecules of Q attract each other by van der Waals' forces only.

Answer: A

How do the properties of P and Q differ?

(1) P has a higher surface tension than Q.
(2) P is less soluble in water than Q.
(3) P is less viscous than Q.

A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

Explanation:

(2) Molecules of P can form hydrogen bonds with water molecules.
Thus, P should be quite soluble in water.
(3) The viscosity of a liquid depends on the strength of intermolecular forces.
Hydrogen bonds are stronger than van der Waals' forces.
Hence P is probably more viscous than Q.

Unit Exercise (p.57)

PART III STRUCTURED QUESTIONS

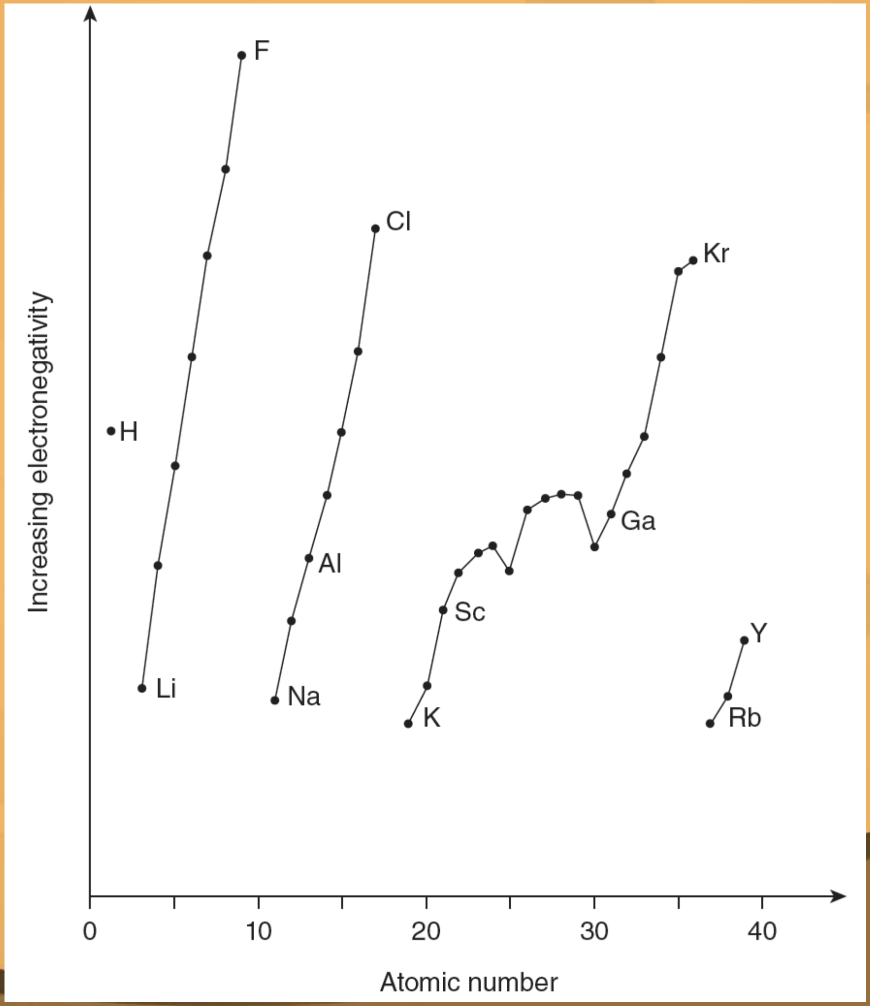
13 Nitrogen reacts with oxygen in car exhausts to produce nitrogen dioxide. Nitrogen dioxide is a gas that can cause air pollution.
Some of the properties of nitrogen dioxide are dependent on the electronegativity of each of the elements it contains.

a) State the meaning of the term electronegativity.

Electronegativity of an element is a measure of the ability of an atom of the element to attract the bonding electrons in a covalent bond towards itself. (1)

Unit Exercise (p.57)

b) The graph shows the variation of electronegativity with atomic number for some of the first 40 elements.



Unit Exercise (p.57)

Describe the trends in electronegativity in the periods and groups of the Periodic Table.

(Pearson BTEC Level 3 Nationals, Applied Science, Unit 1, Sample Assessment Materials, Nov. 2017, 6(a))

Decrease in electronegativity down a group (1)

Increase in electronegativity across a period from left to right (1)

Unit Exercise (p.57)

14 The bond in a HCl molecule is said to be 'polar'.

a) What is the meaning of the term 'polar bond'?

A polar covalent bond is a covalent bond with one bonding atom having a stronger attraction for bonding electrons than the other bonding atom. (1)

b) Explain why the HCl molecule is polar.

H and Cl have different electronegativities. (1)

Unit Exercise (p.57)

15 Carbon dioxide (CO_2) and water (H_2O) both contain polar bonds.

a) Show the polarity of the carbon-oxygen bond and the oxygen-hydrogen bond shown below.

δ+ δ–	δ– δ+
C=O	O–H

b) For each of these molecules, explain whether or not it is polar.

The CO_2 molecule is non-polar.
The two polar C=O bonds are symmetrically arranged on opposite sides of the central carbon atom. The polarities of the two C=O bonds cancel each other out. (1)
The H_2O molecule is polar.
The central oxygen atom has two lone pairs of electrons in its outermost shell. The polarities of the two polar O–H bonds do not cancel out. (1)

Unit Exercise (p.57)

16 Some Pauling electronegativity values for selected elements are given below.

H 2.2						
Li 1.0	Be 1.6	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2

a) How does the electronegativity value change

i) across a period;

Increase (1)

ii) down a group.

Decrease (1)

Unit Exercise (p.57)

b) Using the above values, compare the polarity of the bonds in a molecule of methane (CH_4) and a molecule of silane (SiH_4).

Any two of the following:

Hydrogen is more electronegative than silicon but less electronegative than carbon. (1)

In methane, hydrogen has a δ+ charge while in silane, hydrogen has a δ– charge. (1)

The C–H bonds in methane are more polar than the Si–H bonds in silane. (1)

c) From the table, choose an element which, when covalently bonded to hydrogen, forms a molecule containing bonds that are more polar than those in methane (CH_4). Give the chemical formula of the hydride of your chosen element.

NH_3 / H_2O / HF / HCl (1)

d) Explain why it is possible for the bonds within a molecule to be polar, but for the molecule itself to be non-polar. Give an example of such a molecule.

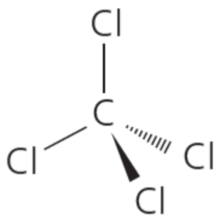
The polar bonds around the central atom are symmetrically arranged and are identical so that the polarities of the bonds cancel each other out. (1)

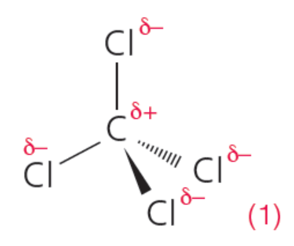
Any suitable example: e.g. CCl_4 (1) / CO_2 (1) / PF_5 (1) / SF_6 (1) / CH_4 (1) / SiH_4 (1)

Unit Exercise (p.57)

17 Chlorine gas reacts with methane. One of the products is tetrachloromethane (CCl_4).

a) The CCl_4 molecule contains polar bonds. Mark all the partial charges on the atoms in the diagram of CCl_4 molecule shown.





b) Explain why the CCl_4 molecule has the partial charges you have shown in (a). Chlorine is more electronegative than carbon. (1)

c) Explain whether or not the CCl_4 molecule is polar.

The CCl_4 molecule is non-polar.

The four polar C–Cl bonds are symmetrically arranged around the central carbon atom.

The polarities of thefour C–Cl bonds cancel each other out. (1)

d) Describe how you might conduct an experiment to test whether CCl_4 is a polar liquid.

Rub a polythene rod with fur.

Bring the charged rod close to a stream of CCl_4 from a burette and notice if any deflection occurs. (1)

A polar liquid would be deflected. (1)

Unit Exercise (p.57)

18 Both BF_3 and NH_3 exist as simple molecules.

a) For each of these molecules, draw its three-dimensional structure.
- BF_3
- NH_3

b) For each of these molecules, explain whether or not it is polar.

c) BF_3 reacts with NH_3 to give F_3BNH_3. Describe the bond formation between BF_3 and NH_3.

(HKDSE, Paper 1B, 2013, 2)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.57)

19 The table below lists the electronegativities of halogens.

Halogen	Electronegativity
Fluorine	4.0
Chlorine	3.2
Bromine	3.0
Iodine	2.7

a) Define the term 'electronegativity'.

Electronegativity of an element is a measure of the ability of an atom of the element to attract the bonding electrons in a covalent bond towards itself. (1)

b) Describe the trend in electronegativity when going down the group.

Decrease (1)

Unit Exercise (p.57)

c) Explain why bromine is less volatile than chlorine.

A bromine molecule has a greater size than a chlorine molecule. / A bromine molecule has more electrons than a chlorine molecule. (1)
Thus, the van der Waals' forces among bromine molecules are stronger than those among chlorine molecules. (1)

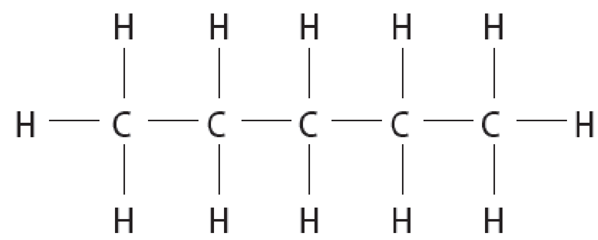
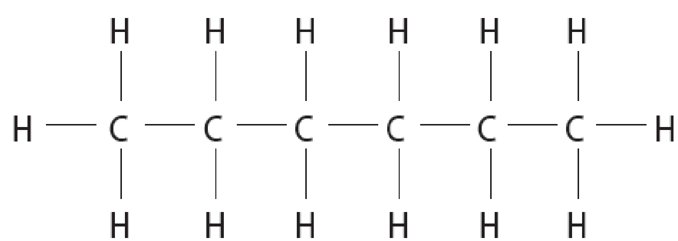
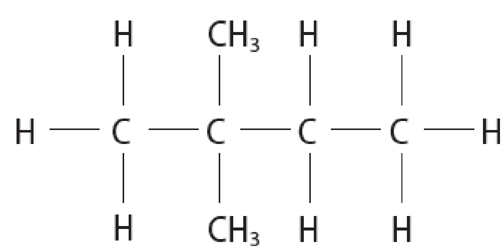
d) The boiling points of hydrogen halides are listed below.
The boiling point increases from hydrogen chloride to hydrogen iodide.
Explain why hydrogen fluoride does NOT follow this trend.

Hydrogen halide	Boiling point (°C)
Hydrogen fluoride	293
Hydrogen chloride	188
Hydrogen bromide	206
Hydrogen iodide	238

There are hydrogen bonds (and van der Waals' forces) between hydrogen fluoride molecules. (1)
Hydrogen bonds are much stronger than van der Waals' forces and require a lot of heat to overcome. (1)

Unit Exercise (p.57)

20 The table below lists the boiling points of three carbon compounds.

Compound	Molecular formula	Structure	Boiling point (°C)
X	C_5H_{12}		36
Y	C_6H_{14}		69
Z	C_6H_{14}		50

Unit Exercise (p.57)

a) Explain why the boiling point of compound Y is higher than that of compound X.

A molecule of Y has a greater size than a molecule of X. / A molecule of Y has more electrons than a molecule of X. (1)

Thus, the van der Waals' forces among molecules of Y are stronger than those among molecules of X. (1)

b) Explain why the boiling point of compound Y is higher than that of compound Z.

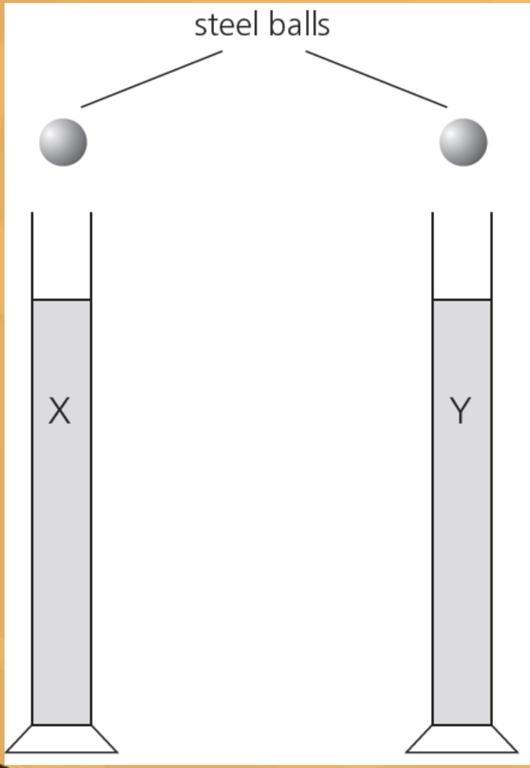
Molecule of Y is longer and somewhat spreadout. The molecule has a larger surface area, allowing a greater contact with neighbouring molecules.

In contrast, molecule of Z is more compact, adopting a roughly spherical shape. The molecule has a smaller surface area for coming into contact with neighbouring molecules. (1)

Thus, the van der Waals' forces among molecules of Y are stronger than those among molecules of Z. (1)

Unit Exercise (p.57)

c) Two identical steel balls are added separately to two identical vertical glass tubes, each containing the same volume of compounds X and Y. In which tube will the steel ball take a longer time to reach the bottom? Explain your answer.



Y takes longer as it has a higher viscosity due to stronger intermolecular forces. (1)

Unit Exercise (p.57)

21 Hydrogen bonds only form when hydrogen atoms are bonded to atoms of certain elements. Name the elements.

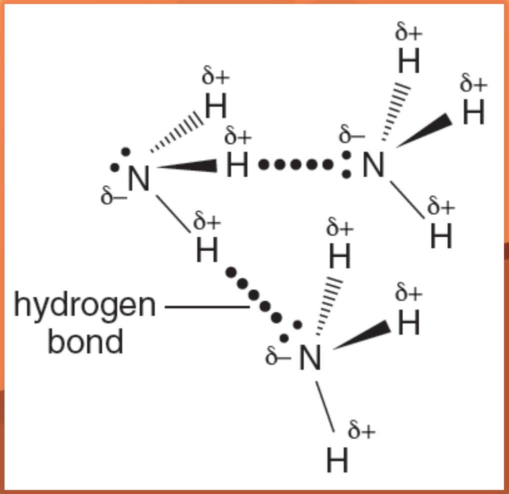
Nitrogen, oxygen and fluorine (1)

Unit Exercise (p.57)

22 Ammonia gas readily condenses to form a liquid when cooled. With the aid of a diagram, explain the formation of hydrogen bonding in ammonia molecules.

A nitrogen atom is highly electronegative. In an ammonia molecule, the nitrogen atom attracts the shared electrons in each N–H bond strongly towards itself, making the N–H bond exceptionally highly polar. The hydrogen atom carries an unusually high partial positive charge. (1)

Such a hydrogen atom is strongly attracted to the lone pair of electrons on the nitrogen atom of an adjacent ammonia molecule. This strong intermolecular force is called a hydrogen bond. (1)



Unit Exercise (p.57)

23 a) Van der Waals' forces exist between all molecules. Explain how these forces arise.

Consider two neighbouring molecules X and Y. The electrons in each molecule are constantly moving. If at any instant there are more electrons on one side of molecule X than the other, one end of the molecule has more negative charge than the other. A temporary dipole is set up. (1)

The temporary dipole in molecule X affects the electron distribution in a neighbouring molecule Y, resulting in the formation of another dipole. (1)

As a result, there are forces of attraction between the δ+ end of the dipole in molecule X and the δ– end of the induced dipole in molecule Y. (1)

Unit Exercise (p.57)

b) The table below shows the boiling points of methanol (CH_3OH) and methanethiol (CH_3SH).

Compound	Boiling point (°C)
Methanol	65
Methanethiol	6

i) Explain, in terms of their intermolecular forces, why the boiling points of these compounds are different.

Hydrogen bonds exist among methanol molecules while only van der Waals' forces exist among methanethiol molecules. (1)

Hydrogen bonds are stronger than van der Waals' forces.

ii) Suggest how a mixture of methanol and methanethiol could be separated.

Fractional distillation / distillation (1)

Unit Exercise (p.57)

c) Suggest why methylselenol (CH_3SeH) has a higher boiling point than methanethiol (CH_3SH).

(AQA Advanced Subsidiary GCE, Unit 1, Jun. 2016, 2(a)–(c))

A methylselenol molecule has a greater size than a methanethiol molecule. / A methylselenol molecule has more electrons than a methanethiol molecule. (1)
Thus, the van der Waals' forces among methylselenol molecules are stronger than those among methanethiol molecules. (1)

Unit Exercise (p.57)

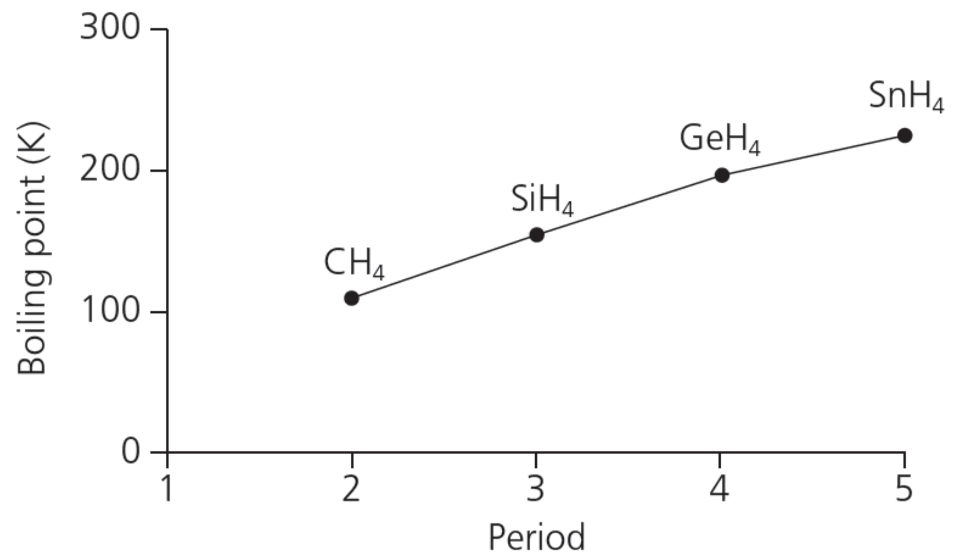
24 Explain, from molecular level, why the density of ice is lower than that of water.

(HKDSE, Paper 1B, 2013, 1(c))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.57)

25 The graph below shows the boiling points of hydrides of Group IV elements.



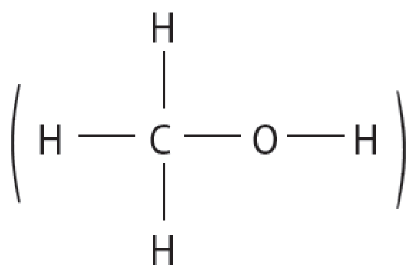
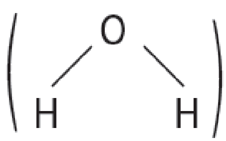
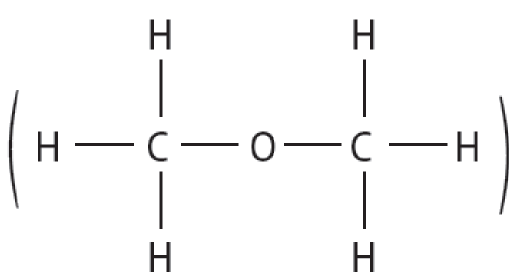
Explain the trend in boiling points from CH_4 to SnH_4. Include the name(s) of all relevant forces.

The boiling points of hydrides of Group IV elements increase from CH_4 to SnH_4.

This is because the strength of van der Waals' forces among the hydride molecules increases from CH_4 to SnH_4. (1)

Unit Exercise (p.57)

26 The boiling points of three substances are listed below.

Substance	Boiling point (°C)
Methanol 	65
Water 	100
Methoxymethane 	34.5

Unit Exercise (p.57)

a) Explain why the boiling point of methanol is higher than that of methoxymethane.

Hydrogen bonds exist among methanol molecules while only van der Waals' forces exist among methoxymethane molecules. (1)
Hydrogen bonds are stronger than van der Waals' forces. (1)

b) Explain why the boiling point of water is higher than that of methanol.

The hydrogen bonding is more extensive in water while it is less extensive in methanol. (1)

Unit Exercise (p.57)

27 a) Explain the following decreasing order of the boiling points of three substances:

$NH_3 > PH_3 > CH_4$

b) Draw a three-dimensional diagram to represent the molecular shape of SF_6.

(HKDSE, Paper 1B, 2018, 3(b)–(c))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Unit Exercise (p.57)

28 The table below lists the boiling points of ammonia, fluorine and bromine.

Substance	Boiling point (°C)
Ammonia, NH_3	240
Fluorine, F_2	83
Bromine, Br_2	331

Explain the difference in boiling points of NH_3, F_2 and Br_2. Include the name(s) of all relevant forces.

Hydrogen bonds and van der Waals' forces exist among ammonia molecules while fluorine and bromine have only van der Waals' forces among molecules. (1)

A bromine molecule has a greater size than a fluorine molecule. / A bromine molecule has more electrons than a fluorine molecule. (1)

Thus, the van der Waals' forces among bromine molecules are stronger than those among fluorine molecules. (1)

Van der Waals' forces in bromine are stronger than the hydrogen bonding and van der Waals' forces in ammonia.

Hydrogen bonding and van der Waals' forces in ammonia is stronger than the van der Waals' forces in fluorine. (1)

Communication mark (1)

Topic Exercise (p.66)

Note: Questions are rated according to ascending level of difficulty (from 1 to 5):

 **question targeted at level 3 and above;**

 **question targeted at level 4 and above;**

 **question targeted at level 5.**

‘ * ’ indicates 1 mark is given for effective communication.

Topic Exercise (p.66)

PART I MULTIPLE CHOICE QUESTIONS

1 Which of the following molecules has a shape different from others?

A PF_3
B NH_3
C PH_3
D BF_3

Answer: D

Topic Exercise (p.66)

2 Which of the following statements about electronegativity is true?

A Non-metals have lower electronegativity than metals.
B Electronegativity decreases across a period in the Periodic Table.
C Electronegativity decreases going down a group in the Periodic Table.
D The bonds between atoms with equal electronegativity are always weak.

(Edexcel Advanced Subsidiary GCE, Unit 2, Jun. 2013, 3)

Answer: C

Topic Exercise (p.66)

3 Which of the following bonds is the most polar?

A Cl–F
B Br–F
C I–F
D F–F

Answer: C

Explanation:
The electronegativity of halogens decreases down the group. Thus, among Cl and F, Br and F, and I and F, the difference in electronegativity between I and F is the greatest.

Topic Exercise (p.66)

4 Which of the following molecules is polar?

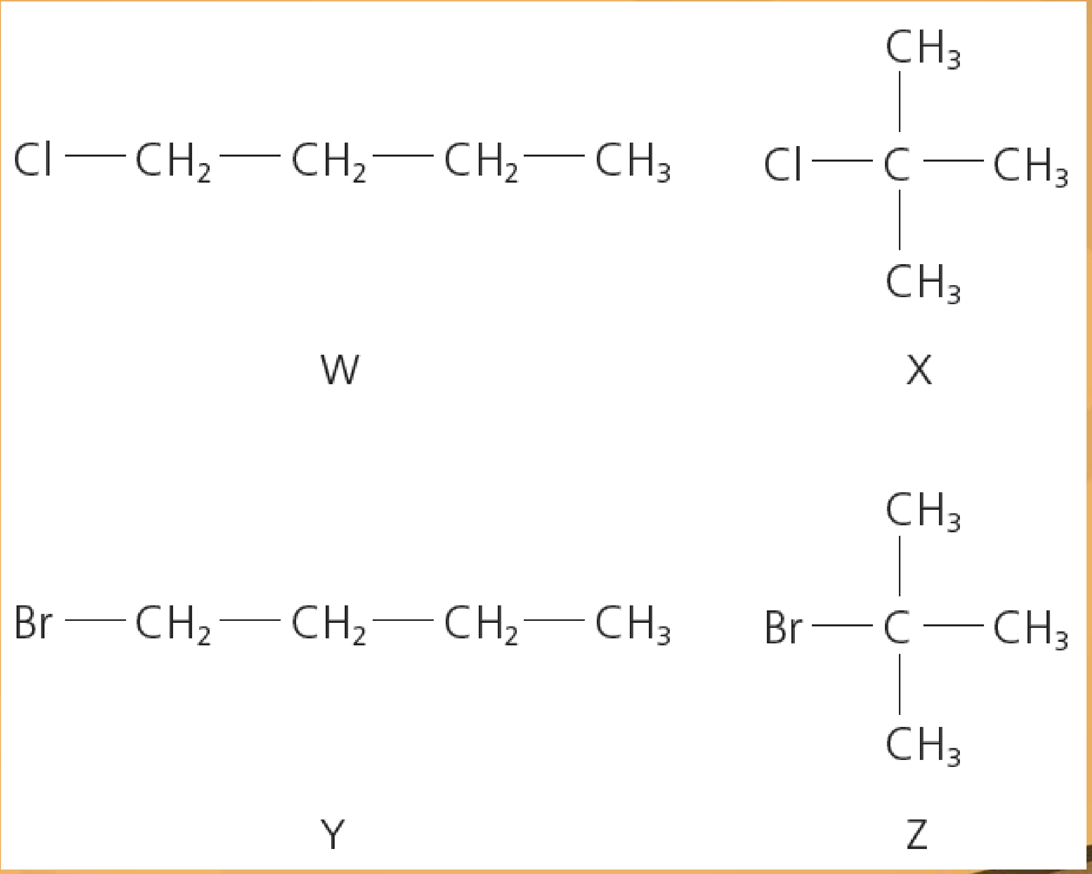
A CO_2
B PCl_3
C SiF_4
D SF_6

(HKDSE, Paper 1A, 2017, 12)

Answer: B

Topic Exercise (p.66)

5 Consider the four carbon compounds (W, X, Y and Z) below.



Topic Exercise (p.66)

Which of the following represents the arrangement of the compounds in increasing order of boiling point?

A X < W < Y < Z
B X < Z < W < Y
C Y < W < X < Z
D Y < Z < W < X

Answer: B

Explanation:
Molecule of W / Y is longer and somewhat spreadout. The molecule has a larger surface area, allowing a greater contact with neighbouring molecules.

In contrast, molecule of X / Z is more compact, adopting a roughly spherical shape. The molecule has a smaller surface area for coming into contact with neighbouring molecules.

Thus, the van der Waals' forces among molecules of W / Y are stronger than those among molecules of X / Z.

A molecule of Y has a greater size than a molecule of W. / A molecule of Y has more electrons than a molecule of W.

Thus, the van der Waals' forces among molecules of Y are stronger than those among molecules of W.

Therefore, the boiling point of X is the lowest while that of Y is the highest.

Topic Exercise (p.66)

6 Which of the following compounds has the highest boiling point?

A HF
B H_2O
C NH_3
D CH_4

Answer: B

Explanation:
A water molecule can form two hydrogen bonds per molecule on average, whereas ammonia or hydrogen fluoride can only form one hydrogen bond per molecule. The hydrogen bonding in water is more extensive.

Topic Exercise (p.66)

7 Which of the following statements can be explained by intermolecular hydrogen bonding?

A Ethanol (CH_3CH_2OH) has a higher boiling point than propane ($CH_3CH_2CH_3$).
B Hydrogen bromide (HBr) has a higher boiling point than hydrogen chloride (HCl).
C Silane (SiH4) has a higher boiling point than methane (CH_4).
D Propanone (CH_3COCH_3) has a higher boiling point than propane ($CH_3CH_2CH_3$).

Answer: A

Topic Exercise (p.66)

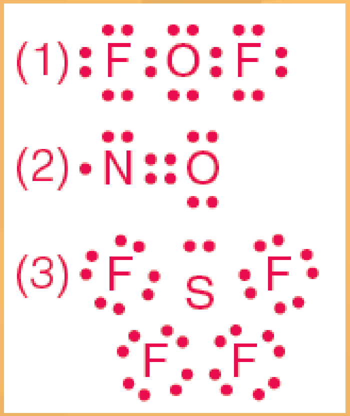
8 Which of the following molecules have non-octet structures?

(1) OF_2
(2) NO
(3) SF_4

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Answer: C

Explanation:
Electron diagrams (showing electrons in the outermost shells only):



Topic Exercise (p.66)

9 Which of the following molecules is / are non-polar?

(1) BCl_3
(2) PCl_3
(3) $CHCl_3$

A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

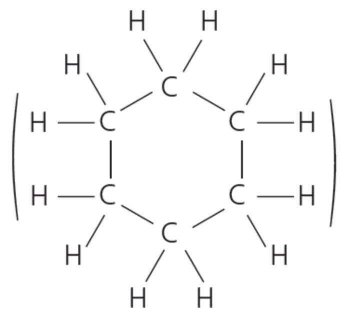
(HKDSE, Paper 1A, 2018, 16)

Answer: A

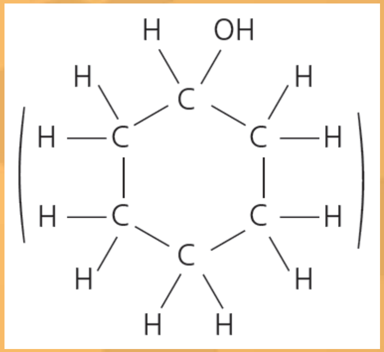
Topic Exercise (p.66)

10 Polar liquids are affected by electric fields. Which of the following liquid jets would be affected by an electric field?

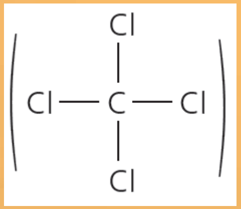
(1) Cyclohexane



(2) Cyclohexanol



(3) Tetrachloromethane



A (1) only
B (2) only
C (1) and (3) only
D (2) and (3) only

Answer: B

Topic Exercise (p.66)

11 Which of the following are characteristics of iodine?

(1) It is more electronegative than bromine.
(2) It is a solid at room temperature and pressure.
(3) Its hydride has a higher boiling point than hydrogen bromide.

A (1) and (2) only
B (1) and (3) only
C (2) and (3) only
D (1), (2) and (3)

Answer: C

Explanation:
(1) Iodine is less electronegative than bromine.

Topic Exercise (p.66)

Directions : Each question (Questions 12–14) consists of two separate statements. Decide whether each of the two statements is true or false; if both are true, then decide whether or not the second statement is a correct explanation of the first statement. Then select one option from A to D according to the following table :
A Both statements are true and the 2nd statement is a correct explanation of the 1st statement. B Both statements are true but the 2nd statement is NOT a correct explanation of the 1st statement. C The 1st statement is false but the 2nd statement is true. D Both statements are false.

Topic Exercise (p.66)

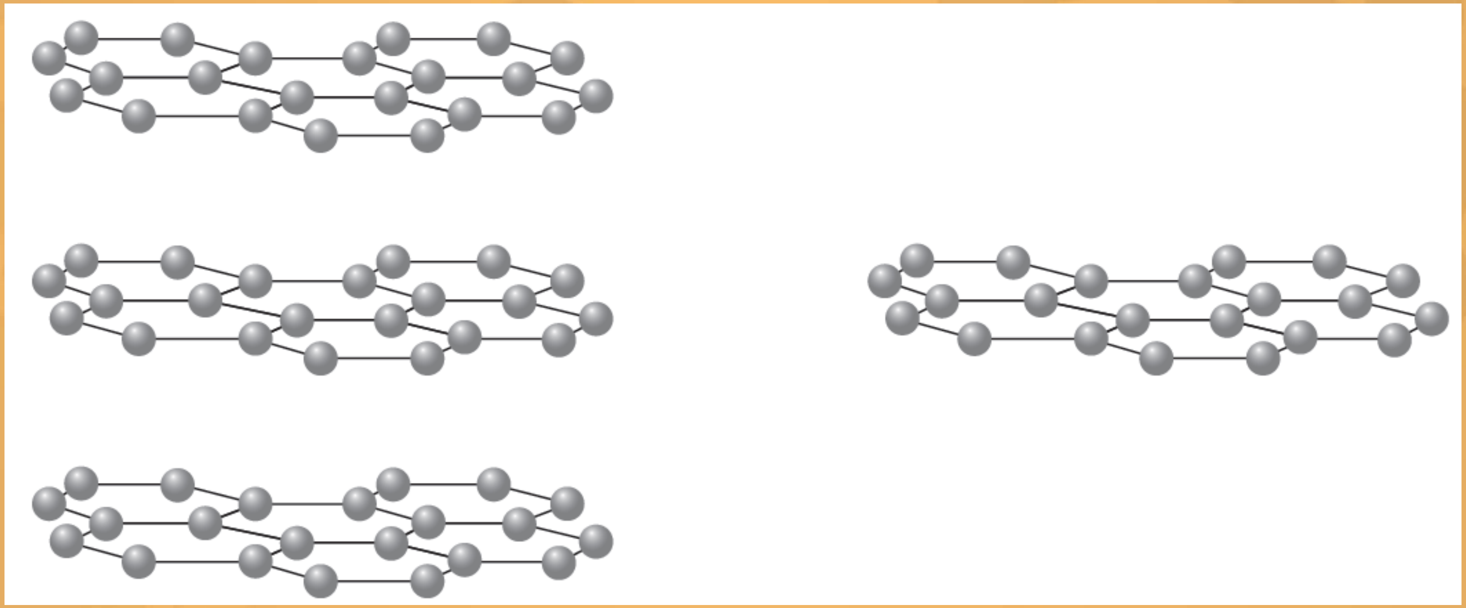
	1st statement	2nd statement	
12	PCl_5 is a polar molecule.	P–Cl bond is polar.	Answer: C
13	The boiling point of H_2O is lower than that of HF.	The electronegativity of oxygen is lower than that of fluorine. *(HKDSE, Paper 1A, 2015, 24)*	Answer: C
14	The boiling point of hydrogen chloride is higher than that of hydrogen fluoride.	The molecular size of hydrogen chloride is greater than that of hydrogen fluoride. *(HKDSE, Paper 1A, 2013, 24)*	Answer: C

Explanation (Q12): The polar P–Cl bonds around the central phosphorus atom are symmetrically arranged so that the polarities of the bonds cancel each other out.

Topic Exercise (p.66)

PART II STRUCTURED QUESTIONS

15 Graphite is a form of carbon and has a layer structure. Graphene is an individual single layer of graphite. Their structures are shown below:



graphite graphene

Topic Exercise (p.66)

a) Thin sheets of graphene can be easily peeled off from graphite using adhesive tape.

i) Explain why graphene can be easily peeled off.

ii) Explain whether graphene can conduct electricity.

iii) Draw an electron diagram for a molecule of the compound formed by complete combustion of graphene, showing electrons in the *outermost shells* only.

b) Based on the fact that graphene can be easily peeled off from graphite, a student concluded that graphite should have a low melting point due to its layer structure. Explain whether you agree with this conclusion.

(HKDSE, Paper 1B, 2014, 1(a)–(b))

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Topic Exercise (p.66)

16 One nitrogen-containing compound formed in the atmosphere is N_2O.

N≡N→O

a) Draw the electron diagram of the N_2O molecule, showing electrons in the *outermost shells* only.

(1)

b) Describe and explain the shape of the N_2O molecule.

Linear shape (1)

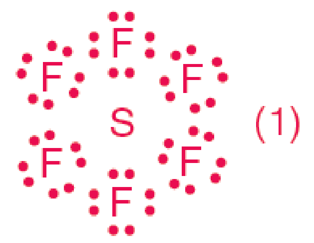
Consider the central nitrogen atom as being surrounded by two electron groups. The two electron groups repel each other and stay as far apart as possible to minimise repulsion and a linear arrangement results. (1)

Topic Exercise (p.66)

17 Sulphur and fluorine combine to form a non-polar molecule sulphur hexafluoride (SF_6).

a) Label the diagram to show the polarity of the S–F bond: $\overset{\delta+}{S}–\overset{\delta-}{F}$ (1)

b) i) Draw the electron diagram of a SF_6 molecule, showing electrons in the *outermost shells* only.

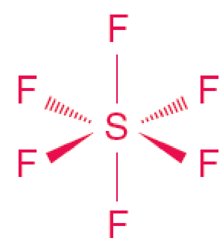


ii) Explain whether the octet rule is applicable to a SF_6 molecule.

The octet rule is not applicable to a SF_6 molecule. There are 12 electrons in the outermost shell of the central sulphur atom. (1)

Topic Exercise (p.66)

c) i) Draw a three-dimensional structure of the SF_6 molecule and explain the shape of the molecule.



(1)

The six bond pairs in the outermost shell of the central sulphur atom stay as far part as possible to minimise repulsion and an octahedral arrangement results. (1)

ii) Suggest why SF_6 is a non-polar molecule, even though it contains polar bonds.

The polar S–F bonds around the central sulphur atom are symmetrically arranged so that the polarities of the bonds cancel each other out. (1)

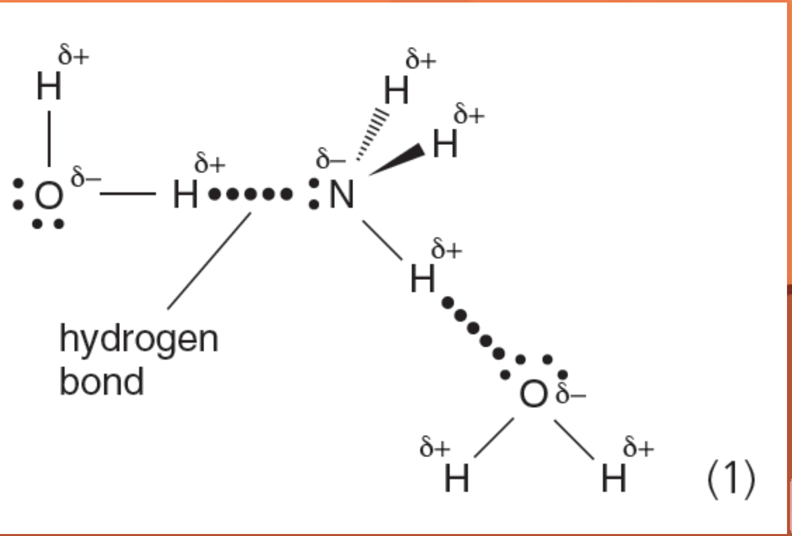
Topic Exercise (p.66)

18 Ammonia molecules contain polar covalent bonds.

a) Explain, using ammonia as an example, the meaning of the term 'polar covalent bond'.

A polar covalent bond is a covalent bond with one bonding atom (in this case N) having a stronger attraction for bonding electrons than the other bonding atom (in this case H). (1)

b) Ammonia is very soluble in water. This is because ammonia molecules can form hydrogen bonds with water molecules.
Draw a diagram to show the hydrogen bond between an ammonia molecule and a water molecule. Include partial charges and all lone pairs of electrons.



Topic Exercise (p.66)

19 Consider the molecules CO_2, CS_2 and CH_2Br_2.

a) For each of the following molecules, draw its three-dimensional structure.

i) CS_2

ii) CH_2Br_2

b) Identify, with explanation, the polar bond(s) in CH_2Br_2.

c) Suggest why, under room temperature and pressure, CO_2 is a gas but CS_2 is a liquid.

(HKDSE, Paper 1B, 2016, 4)

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Topic Exercise (p.66)

20 Sulphur and fluorine react together to give the covalent compounds SF_2, SF_4 and SF_6.

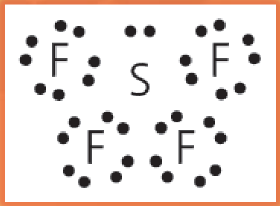
a) Suggest a reason why sulphur can form SF_2, SF_4 and SF_6 whereas oxygen only forms OF_2.

Sulphur can form compounds with more than 8 electrons in the outermost shell of its atom. / Oxygen CANNOT form compounds with more than 8 electrons in the outermost shell of its atom. (1)

b) Explain why the fluorides of sulphur contain polar bonds.

Fluorine and sulphur have different electronegativities. (1)

c) i) Draw the electron diagram of a SF_4 molecule, showing electrons in the *outermost shells* only.



ii) State whether a SF_4 molecule is polar. Explain your answer.

SF_4 molecule is polar.

The central sulphur atom has one lone pair in its outermost shell. The polar S–F bonds around the central sulphur atom are NOT symmetrically arranged so that the polarities of the bonds do NOT cancel out. (1)

Topic Exercise (p.66)

d) i) State the shape of a SF_6 molecule.

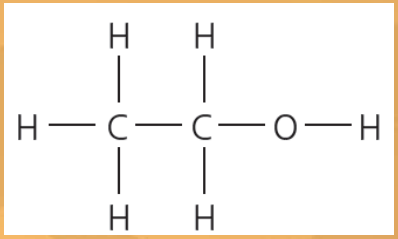
Octahedral (1)

ii) Explain why a SF_6 molecule is non-polar.

The polar S–F bonds around the central sulphur atom are symmetrically arranged so that the polarities of the bonds cancel each other out. (1)

Topic Exercise (p.66)

21 The structure of ethanol is shown below:



a) Name the strongest type of intermolecular attraction between ethanol molecules.
Explain how such intermolecular attraction arises.

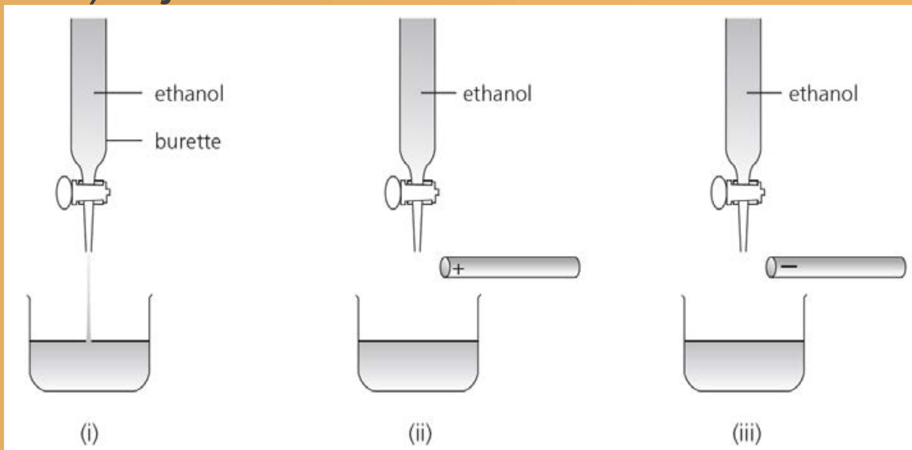
Hydrogen bonds (1)

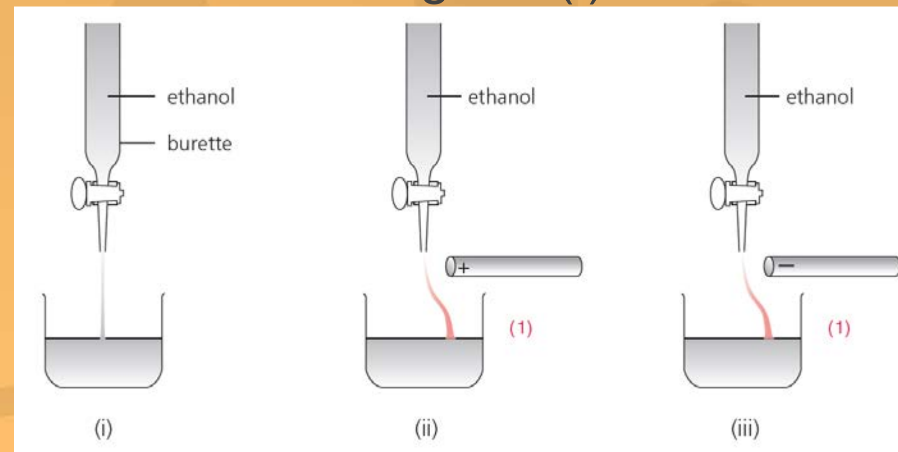
An oxygen atom is highly electronegative. In an ethanol molecule, the oxygen atom attracts the shared electrons in the O–H bond strongly towards itself, making the O–H bond exceptionally highly polar. The hydrogen atom carries an unusually high partial positive charge. (1)

Such a hydrogen atom is strongly attracted to the lone pair of electrons on the oxygen atom of an adjacent ethanol molecule. This strong intermolecular force is called a hydrogen bond. (1)

Topic Exercise (p.66)

b) A jet of ethanol is run from a burette as shown in diagram (i).





A positively charged rod and a negatively charged rod are placed near the jet of ethanol as shown in diagrams (ii) and (iii).

Show and explain your expected observation in each case of (ii) and (iii).

An ethanol molecule is polar. (1)
The jet of ethanol is deflected by the positively charged rod.
The negative ends of the molecules are attracted towards the rod. (1)
The jet of ethanol is deflected by the negatively charged rod.
The positive ends of the molecules are attracted towards the rod. (1)

Topic Exercise (p.66)

22 Both silicon and phosphorus combine with fluorine. Silicon forms silicon tetrafluoride (SiF_4) while phosphorus forms phosphorus trifluoride (PF_3) and pentafluoride (PF_5).

a) In which of the fluoride molecules does the central atom have a non-octet electronic arrangement? Explain your choice. PF_6

For PF_5, there are 10 electrons in the outermost shell of the central phosphorus atom. (1)

b) The shape of a SiF_4 molecule is the same as that of a CH_4 molecule. Explain why it is so.

In both CH_4 and SiF_4 molecules, there are four bond pairs in the outermost shell of the central atom. (1)

These bond pairs stay as far apart as possible to minimise repulsion and a tetrahedral arrangement results. (1)

c) i) Show the polarity of the P–F bond in the phosphorus fluorides using the diagram below.

$\delta+$ $\delta-$
P–F (1)

Topic Exercise (p.66)

c) ii) Explain why the bond is polar.

Phosphorus and fluorine have different electronegativities. (1)

iii) For the molecule of each fluoride of phosphorus (PF_3 and PF_5), explain whether or not it is polar.

A PF_3 molecule is polar.

The central phosphorus atom has one lone pair of electrons in its outermost shell. The polar P–F bonds around the central phosphorus atom are NOT symmetrically arranged so that the polarities of the bonds do NOT cancel out. (1)

A PF_5 molecule is non-polar.

The polar P–F bonds around the central phosphorus atom are symmetrically arranged so that the polarities of the bonds cancel each other out. (1)

Topic Exercise (p.66)

23 The boiling points of the first four alkanes are listed in the table below.

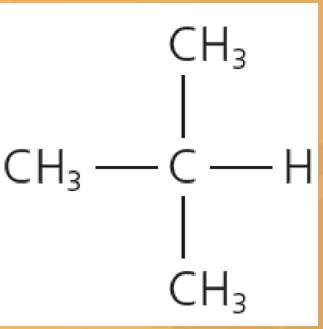
Alkane	Methane CH_4	Ethane CH_3CH_3	Propane $CH_3CH_2CH_3$	Butane $CH_3CH_2CH_2CH_3$
Boiling point (K)	111	184	231	273

a) Suggest an explanation for the increase in boiling point from methane to butane.

The molecular size / number of electrons per molecule increases from methane to butane. (1)
Thus, the strength of van der Waals' forces among molecules increases from methane to butane. (1)

Topic Exercise (p.66)

b) The structure of methylpropane is shown below.



The methylpropane has the same molecular formula as butane. Its boiling point is 261 K.

Explain the difference between this value and the boiling point of butane.

The butane molecule is longer and somewhat spreadout. The molecule has a larger surface area, allowing a greater contact with neighbouring molecules. In contrast, the methylpropane molecule is more compact, adopting a roughly spherical shape. The molecule has a smaller surface area for coming into contact with neighbouring molecules.

Thus, the van der Waals' forces among butane molecules are stronger than those among methylpropane molecules. (1)

Topic Exercise (p.66)

24 The table below shows the electronegativity values of some elements.

Element	H	C	N	O	F
Electronegativity	2.2	2.6	3.0	3.4	4.0

a) State the strongest type of intermolecular attraction in

i) methane (CH_4); Van der Waals' forces (1)

ii) hydrogen fluoride (HF). Hydrogen bonds (1)

b) Use the values in the table to explain how the type of intermolecular attraction stated in (a)(ii) arises between two hydrogen fluoride molecules.

A fluorine atom is highly electronegative. In a hydrogen fluoride molecule, the fluorine atom attracts the shared electrons in the H–F bond strongly towards itself, making the H–F bond exceptionally highly polar. The hydrogen atom carries an unusually high partial positive charge. (1)

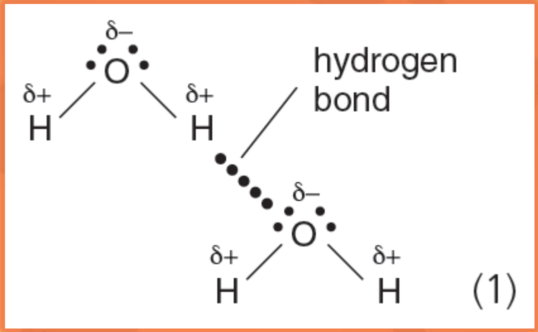
Such a hydrogen atom is strongly attracted to the lone pair of electrons on the fluorine atom of an adjacent hydrogen fluoride molecule.

This strong intermolecular force is called a hydrogen bond. (1)

Topic Exercise (p.66)

25 Water has hydrogen bonding.

a) Draw a labelled diagram to show the hydrogen bonding between two water molecules. Include partial charges and all lone pairs of electrons.



b) Suggest and explain TWO anomalous properties of water caused by hydrogen bonding.

Property 1 and explanation

Ice is less dense than water. (1)

In ice, each water molecule is surrounded tetrahedrally by four other water molecules joined by intermolecular hydrogen bonding. Ice has an 'open' structure with water molecules held apart. (1)

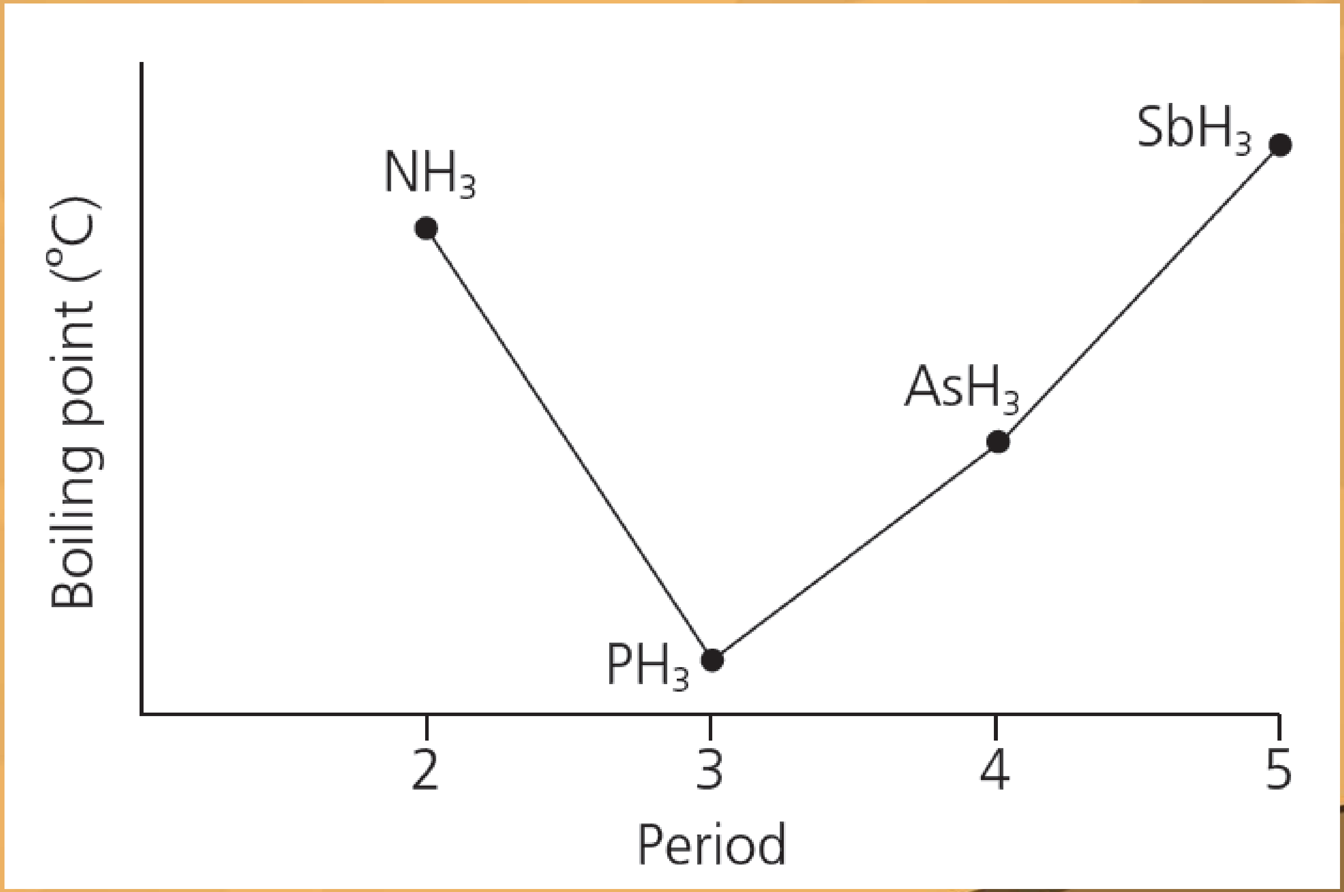
Property 2 and explanation

Ice has a relatively high melting point. (1)

Hydrogen bonds among water molecules are relatively strong. (1)

Topic Exercise (p.66)

26 The trend of the boiling points of Group V hydrides is shown below:



Topic Exercise (p.66)

a) Name the force that has to be overcome in order to boil PH_3. Explain what causes this force.

Van der Waals' forces (1)

Consider two PH_3 molecules. The electrons in each molecule are constantly moving. If at any instant there are more electrons on one side of a PH_3 molecule than the other, then one end of the molecule has more negative charges than the other. A temporary dipole is set up.

The temporary dipole in this PH_3 molecule affects the electron distribution in a neighbouring PH_3 molecule, resulting in the formation of another dipole. (1)

As a result, there are forces of attraction between the δ+ end of the dipole in the first PH_3 molecule and the δ– end of the induced dipole in the neighbouring PH_3 molecule. (1)

b) Explain the trend in the boiling points from PH_3 to SbH_3.

The molecular size / number of electrons per molecule increases from PH_3 to SbH_3. (1)

Thus, the strength of van der Waals' forces among molecules increases from PH_3 to SbH_3. (1)

c) Explain why the boiling point of NH_3 does not follow the trend.

Hydrogen bonds exist among NH_3 molecules and hydrogen bonds are stronger than van der Waals' forces. (1)

Topic Exercise (p.66)

27 Explain the following increasing order of the boiling points of three substances:

$$H_2 < F_2 < HF$$

(HKDSE, Paper 1B, 2017, 5)

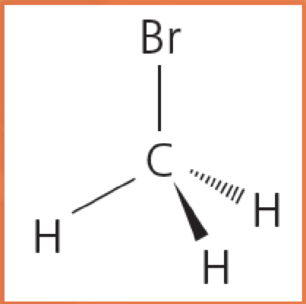
Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

Topic Exercise (p.66)

28 The table lists the boiling points of bromomethane, chloromethane and water.

	Boiling point (K)
Bromomethane (CH_3Br)	277
Chloromethane (CH_3Cl)	249
Water (H_2O)	373

a) Draw the three-dimensional structure of a CH_3Br molecule.



b) Identify, with explanation, the polar bond(s) in CH_3Br.

C–H and C–Br bonds are polar.
C and H / C and Br have different electronegativities. (1)

Topic Exercise (p.66)

c) Explain whether or not a CH_3Br molecule is polar.

A CH_3Br molecule is polar.
In a CH_3Br molecule, the four bonds adopt a tetrahedral arrangement.
However, the polarity of a C–Br bond is different from the polarity of a C–H bond. Hence the polarities of the bonds do not cancel out. (1)

d) Explain why CH_3Br has a higher boiling point than CH_3Cl.

A CH_3Br molecule has a greater size than a CH_3Cl molecule. / A CH_3Br molecule has more electrons than a CH_3Cl molecule.
Thus, the van der Waals' forces among CH_3Br molecules are stronger than those among CH_3Cl molecules. (1)

e) Explain why H_2O has a higher boiling point than CH_3Br.

Hydrogen bonds exist among H_2O molecules while only van der Waals' forces exist among CH_3Br molecules. (1)
Hydrogen bonds are stronger than van der Waals' forces. (1)