Surfing

NATIONAL CHEMISTRY

Unit 2 Molecular Interactions and Reactions

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Contents

Introduction

Each book in the *Surfing* series contains a summary, with occasional more detailed sections, of all the mandatory parts of the syllabus, along with questions and answers.

All types of questions – multiple choice, short response, structured response and free response – are provided. Questions are written in exam style so that you will become familiar with the concepts of the topic and answering questions in the required way.

Answers to all questions are included.

A topic test at the end of the book contains an extensive set of summary questions. These cover every aspect of the topic, and are useful for revision and exam practice.

Words To Watch

account, account for State reasons for, report on, give an account of, narrate a series of events or transactions.

analyse Interpret data to reach conclusions.

annotate Add brief notes to a diagram or graph.

apply Put to use in a particular situation.

assess Make a judgement about the value of something.

calculate Find a numerical answer.

clarify Make clear or plain.

classify Arrange into classes, groups or categories.

comment Give a judgement based on a given statement or result of a calculation.

compare Estimate, measure or note how things are similar or different.

construct Represent or develop in graphical form.

contrast Show how things are different or opposite.

create Originate or bring into existence.

deduce Reach a conclusion from given information.

define Give the precise meaning of a word, phrase or physical quantity.

demonstrate Show by example.

derive Manipulate a mathematical relationship(s) to give a new equation or relationship.

describe Give a detailed account.

design Produce a plan, simulation or model.

determine Find the only possible answer.

discuss Talk or write about a topic, taking into account different issues or ideas.

distinguish Give differences between two or more different items.

draw Represent by means of pencil lines.

estimate Find an approximate value for an unknown quantity.

evaluate Assess the implications and limitations.

examine Inquire into.

explain Make something clear or easy to understand.

extract Choose relevant and/or appropriate details.

extrapolate Infer from what is known.

hypothesise Suggest an explanation for a group of facts or phenomena.

identify Recognise and name.

interpret Draw meaning from.

investigate Plan, inquire into and draw conclusions about.

justify Support an argument or conclusion.

label Add labels to a diagram.

list Give a sequence of names or other brief answers.

measure Find a value for a quantity.

outline Give a brief account or summary.

plan Use strategies to develop a series of steps or processes.

predict Give an expected result.

propose Put forward a plan or suggestion for consideration or action.

recall Present remembered ideas, facts or experiences.

relate Tell or report about happenings, events or circumstances.

represent Use words, images or symbols to convey meaning.

select Choose in preference to another or others.

sequence Arrange in order.

show Give the steps in a calculation or derivation.

sketch Make a quick, rough drawing of something.

solve Work out the answer to a problem.

state Give a specific name, value or other brief answer.

suggest Put forward an idea for consideration.

summarise Give a brief statement of the main points.

synthesise Combine various elements to make a whole.

Intermolecular Forces and Gases

1 Intramolecular forces

You already know that:

- Atoms of covalent substances share electrons.
- Most covalent substances exist as molecules.
- The atoms within these molecules are held together by strong covalent bonds – intramolecular forces.
- The whole molecules are attracted to each other by weaker forces – intermolecular forces.

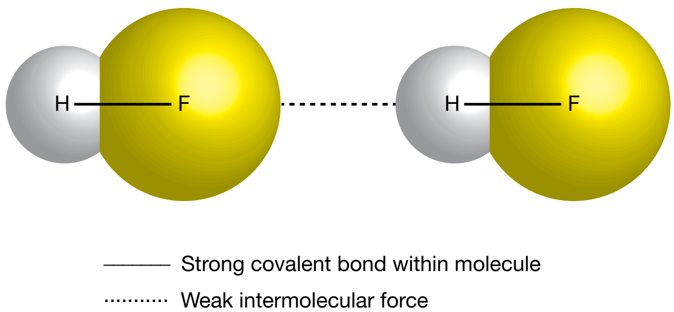


Figure 1.1 Forces within and between molecules of hydrogen fluoride.

The nature and strength of intermolecular forces can explain observable properties such as vapour pressure, melting point, boiling point and solubility. However, to understand this, **we first need to look at bonds and forces in more detail**.

Covalent bonds – non-polar bonds

When atoms share valence electrons to form covalent bonds the shared pair of electrons orbits both nuclei. Covalent bonds are very strong bonds.

In a molecule such as hydrogen (H_2), the two atoms are identical hydrogen atoms, so the bonding electrons are attracted equally to each nucleus and shared equally between them. The two shared electrons spend equal time in the area around each atom and this sharing of electrons forms the covalent bond which holds the atoms together. Such bonds are referred to as **non-polar covalent bonds**. Figure 1.2 shows two different types of models which can be used to illustrate this.

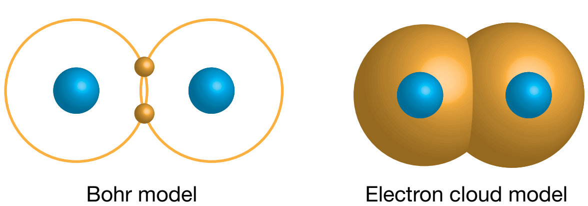


Figure 1.2 Molecule of hydrogen gas with non-polar bond between its atoms.

Covalent bonds – polar bonds

When the atoms joining together are not identical, then their electrons are not shared equally. For example, this happens when hydrogen and chlorine atoms share a pair of electrons to form hydrogen chloride (HCl). The shared electrons are more strongly attracted to the chlorine nucleus than to the hydrogen nucleus, so the electrons spend more time near the chlorine. This makes the chlorine end slightly more negative than the hydrogen end. A bond with unequal sharing of electrons is called a **polar covalent bond**. The bond forms a **dipole**, a structure with two oppositely charged ends and it can be shown as follows.

$$\delta^+ \mathbf{H} - \mathbf{Cl}\ \delta^-$$

The Greek lower case delta (δ) is used to represent a partial (small amount of) electrical charge, a much smaller charge than that on a proton or an electron. δ^+ and δ^- symbols indicate the areas on the molecule surface which are slightly more positive or slightly more negative due to unequal sharing of electrons.

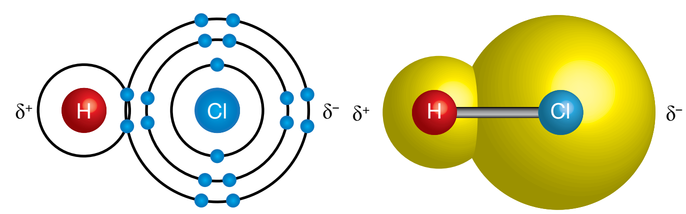


Figure 1.3 Electron dot and electron cloud diagrams to illustrate the polar bond in hydrogen chloride.

Polar and non-polar molecules

A molecule which consists of **only non-polar bonds**, e.g. H_2, Cl_2 will be called a **non-polar molecule**.

A molecule which contains **one polar bond**, e.g. HCl will have an unequal charge distribution so it will form a **polar molecule** with a slightly positive area at one end and a slightly negative area at the other end.

A molecule which has **more than one polar bond** may have **polar molecules**, such as in water H_2O, or **non-polar molecules** such as in tetrachloromethane CCl_4 (which used to be called carbon tetrachloride). Molecules such as CCl_4 have non-polar molecules because the way that the polar bonds are arranged means that they cancel each other out.

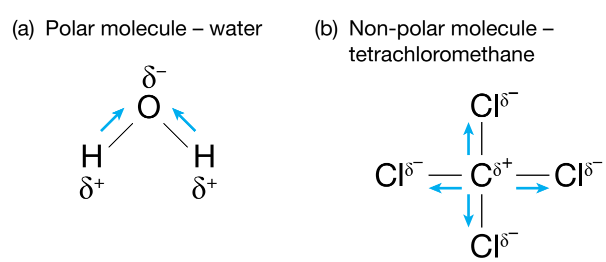


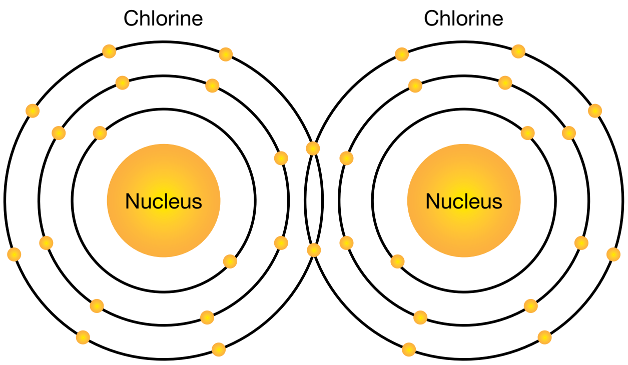
Figure 1.4 Polar and non-polar molecules.

With these non-polar molecules, the charge distribution is even over the surface of the molecules even though polar covalent bonds are present.

The shape of a molecule helps to determine whether or not polar bonds cancel each other out and thus whether the molecule will be polar or non-polar. You will be learning about shapes of molecules soon.

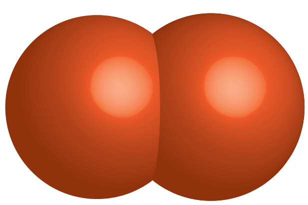
QUESTIONS

1. A molecule of chlorine is drawn below.
   (a) Identify the shared electrons in this molecule.
   (b) Explain why the chlorine molecule is said to contain a non-polar covalent bond.

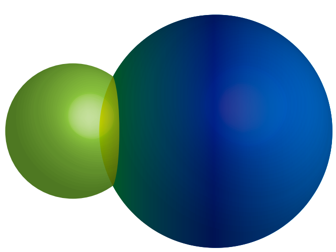


2. The following diagrams are said to illustrate the difference between an ionic bond, a non-polar covalent bond and a polar covalent bond. Identify which is which and justify your decisions.

(a)



(b)



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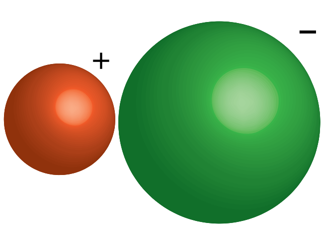
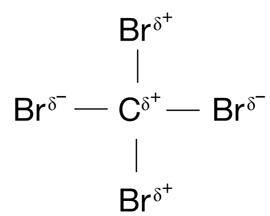


3. Copy and complete the table to show whether or not each of the substances listed has polar or non-polar bonds within its molecules.

Substance	Formula	Polar/non-polar bonds
Hydrogen		
Ammonia		
Hydrogen bromide		
Water		
Carbon dioxide		
Argon		

4. The word ‘polar’ can be applied to a bond, a molecule and a substance. Discuss this statement.
5. Explain the difference in polarity between a hydrogen molecule and a hydrogen fluoride molecule.
6. Why is the tetrabromomethane molecule (see below) non-polar, even though it contains polar bonds.



7. Identify each of the following statements as true or false and justify your answer.
   (a) Oxygen molecules are non-polar.
   (b) Sodium chloride contains molecules of NaCl.
   (c) Chlorine contains polar covalent bonds.
   (d) Covalent bonds are strong forces of attraction within molecules.
8. Check your knowledge with this quick quiz.
   (a) A polar bond is an attraction between two atoms which are (the same/different).
   (b) The covalent bond in a bromine molecule is a (polar/non-polar) bond.
   (c) The covalent bond in a hydrogen fluoride molecule is (polar/non-polar).
   (d) Covalent bonds occur when electrons are (shared/donated/received).

2 Intermolecular Forces

This chapter and the next look at the types of forces that exist between covalent molecules.

Intermolecular forces (also called van der Waals forces) are the attractive electrostatic forces holding covalent molecules together.

These have to be broken for a covalent molecular substance to dissolve or change state. When a covalent substance dissolves, melts or evaporates, the molecule itself does not break up. Instead, whole molecules move away from each other.

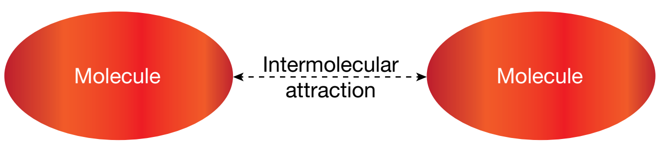


Figure 2.1 Intermolecular forces.

There are three main **types of intermolecular forces** – dispersion forces, dipole-dipole forces and hydrogen bonds. We will look at the first two in this chapter, and then in Chapter 3 we will compare them to hydrogen bonding.

Dispersion forces

Dispersion forces are very weak, constantly changing, electrostatic forces of attraction between covalent molecules.

Dispersion forces occur between **all covalent molecules**. In non-polar substances, dispersion forces are the only forces between the molecules. Polar substances have other types of intermolecular forces as well as dispersion forces.

Dispersion forces are caused by changes in the electron distribution of atoms in nearby molecules. This is due to the shifting of electron clouds as electrons orbit the nuclei of atoms within the molecule.

At any one instant all the electrons might be on one side of the atom, at the next instant they might be evenly spread, or they might be at the other end. These constant changes in charge distribution within molecules cause constantly fluctuating forces of attraction. As the electron clouds are distorted, they form – just for an instant – a temporary dipole.

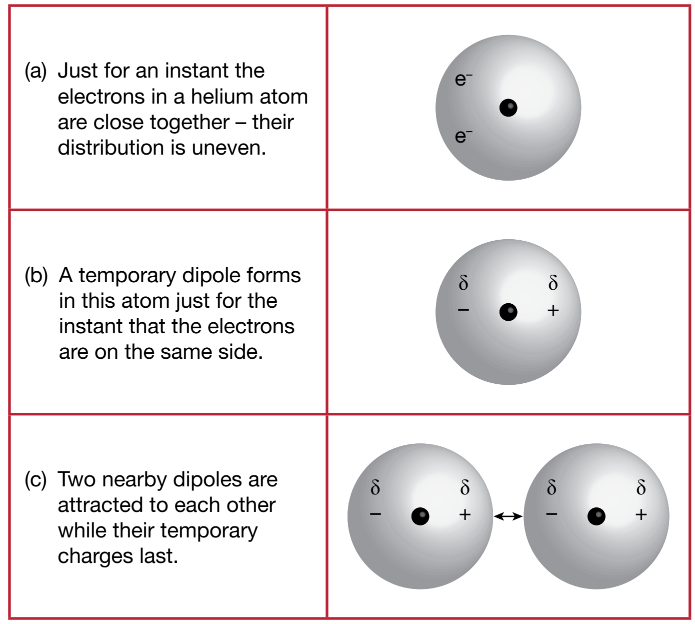


Figure 2.2 Electrons are constantly moving.

When one of these temporary dipoles comes close to another molecule it can induce a charge in it also.

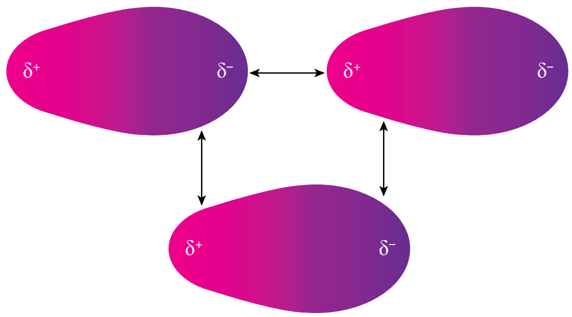
These temporary dipoles, and the dispersion forces they cause, are constantly being formed and broken. They attract each other briefly and then disappear as the electrons move on and form different dipoles.

Generally, not much energy is needed to break dispersion forces, as they are often quite weak. Non-polar substances, with only dispersion forces holding their molecules together, e.g. hydrogen, methane and nitrogen, have very low melting and boiling points. This is because their only intermolecular forces are dispersion forces which are weak and easy to break.

However, dispersion forces do **vary in strength** depending on the **size and shape of the molecule**. The bigger the molecular mass, the more electrons in the molecule, and the greater the distance they can move, the bigger the dispersion forces. Also long thin molecules have greater dispersion forces than short fat ones with the same number of electrons – they develop bigger dipoles and are more strongly attracted as they can get closer together.

Dipole-dipole forces

Dipole-dipole forces are stronger, permanent, electrostatic forces of attraction between **polar molecules**. Polar molecules are dipoles – they have permanently charged positive and negative ends. The slightly positive end of one molecule attracts the slightly negative end of another molecule. This electrostatic attraction between opposite charges is the dipole-dipole force. Because they are stronger attractive forces, more energy is needed to break dipole-dipole forces than to break dispersion forces.



Dipole-dipole force between molecules

Figure 2.3 Polar molecules are attracted to each other by dipole-dipole forces between opposite charges.

Sometimes a polar molecule induces a charge in a non-polar molecule and they are then attracted while they are close to each other.

(a) A non-polar molecule and a polar molecule come near each other.

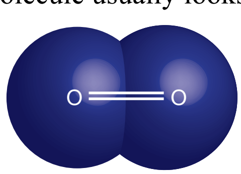
(b) A temporary charge is induced in the non-polar molecule, so it forms a temporary dipole.

Figure 2.4 Inducing dipoles.

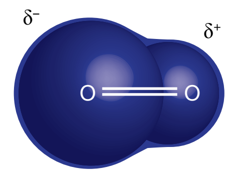
In the next chapter we will compare these dispersion forces and dipole-dipole forces with the third type of intermolecular force – hydrogen bonding.

QUESTIONS

1. Dispersion forces are present between all molecules.
   (a) Describe the formation of dispersion forces between molecules.
   (b) Identify what determines the strength of dispersion forces.
2. Outline two differences between dispersion forces and dipole-dipole forces.
3. (a) Define a dipole.
   (b) Where would you expect to find a dipole?
   (c) What is a dipole-dipole force?
4. An oxygen molecule usually looks like this:



In a textbook it was shown looking like this:



Suggest a reason for it being drawn this way.

5. Would sodium chloride have any type of intermolecular forces? Justify your answer.
6. The following substances exist as molecules. Divide them into two groups based on the nature of their intermolecular forces.
   Chlorine, ethane, nitrogen, sulfur dioxide.
7. Explain how a molecule can be non-polar when it has polar covalent bonds.
8. Research the relevance of dispersion forces to lizards called geckoes.

9. Identify whether each statement is true or false and justify your decision.
   (a) Dispersion forces are permanent forces between all molecules.
   (b) Dispersion forces occur between all molecules, both polar and non-polar molecules.
10. Check your knowledge with the following quick quiz.
   (a) Intermolecular forces are those (within/between) molecules.
   (b) Identify the intermolecular forces that occur between all covalent molecules.
   (c) The temporary forces of attraction between molecules are called .... forces.
   (d) What do we call permanent attractive forces between molecules?
   (e) Which are weaker forces, dispersion forces or dipole-dipole forces?
   (f) Dipole-dipole forces occur between (polar/non-polar) molecules.

3 More Intermolecular Forces – Hydrogen Bonds

In Chapter 2 you looked at two types of forces between molecules (intermolecular forces) – dispersion forces and dipole-dipole forces.

The third type of electrostatic attractive force that can occur between covalent molecules is the hydrogen bond.

Hydrogen bonds are strong electrostatic attractions between the slightly positive **hydrogen** atoms of one molecule and a highly electronegative atom (**oxygen, nitrogen or fluorine)** in another nearby molecule. Hydrogen bonding is sometimes regarded as a special type of dipole-dipole force.

Just having a hydrogen atom in a molecule is not enough for hydrogen bonds to form between molecules. Hydrogen bonds only occur in highly polar molecules – those which contain at least one hydrogen atom and *also* either oxygen, nitrogen or fluorine. They are found in compounds such as water, ammonia, hydrogen fluoride and ethanol.

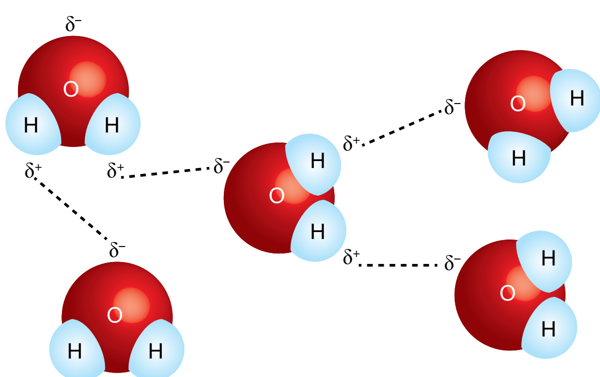


Figure 3.1 Hydrogen bonds between water molecules. (*Note*: Dispersion forces and dipole-dipole forces also exist between these molecules but are not shown here.)

Remember that dispersion forces, dipole-dipole forces and hydrogen bonds are all **intermolecular forces** – they occur between molecules.

Intermolecular forces vary in strength. However, they are not the strongest forces that exist in the atomic world. None of the intermolecular forces, not even hydrogen bonds, are nearly as strong as the covalent forces between atoms inside the molecule, nor are they as strong as the forces between positive and negative ions.

QUESTIONS

1. Compare the three main types of intermolecular forces. Tabulate your answer.

2. (a) Describe what is meant by a hydrogen bond.
   (b) Use a diagram to show hydrogen bonding between ammonia molecules.
   (c) A student wrote that, 'Hydrogen bonds hold the hydrogen atoms together in a molecule of ammonia gas.' For this answer the student received no marks. Explain why the teacher marked the answer with a zero.

3. Copy and complete the table to summarise the types of intermolecular forces between molecules of the substances listed.

Substance	Dispersion forces	Dipole-dipole forces	Hydrogen bonds
Helium He			
Oxygen O_2			
Water H_2O			
Methane CH_4			
Hydrogen sulfide H_2S			
Ammonia NH_3			
Hydrogen fluoride HF			
Acetic acid CH_3COOH			
Carbon dioxide CO_2			
Hydrogen chloride HCl			
Bromine liquid Br_2			
Ethanol C_2H_5OH			

4. Identify each of the following statements as true or false and justify your decision.
   (a) Hydrogen bonding occurs in a molecule of water between the hydrogen and the oxygen atoms.
   (b) Dipole-dipole forces are the weakest intermolecular forces.

5. List the following substances in order of decreasing strength of intermolecular forces: hydrogen sulfide, ethane, water, argon.

6. Check your knowledge with this quick quiz.
   (a) What do we call the strongest of the intermolecular forces of attraction?
   (b) Name a substance which would have hydrogen bonding.
   (d) Name two atoms that must be present in two molecules for hydrogen bonding to occur between the molecules.

4 Vapour Pressure

The nature and strength of forces between the molecules of a substance (intermolecular forces) can explain observable properties such as vapour pressure.

Vapour pressure

A liquid consists of particles which are constantly moving and some of these particles that are at the surface will have enough kinetic energy to break any forces holding them and escape. This is what we call evaporation. If a liquid is in a closed container, any escaping particles will be trapped inside.

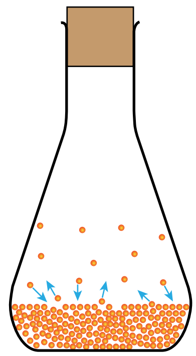


Figure 4.1 Liquid in a closed container.

The particles that escape from the liquid move around in the space above it and frequently collide with the walls of the container. When they hit the walls they exert pressure on the walls. This pressure is called **vapour pressure** – it is the pressure of a gas above a liquid and it can vary in size. Pressure is defined as the force acting per unit area. The standard unit is the pascal (Pa).

Table 4.1 Some typical vapour pressures at 20°C.

Substance	Vapour pressure (kPa)
Octane	1.47
Water	2.34
Ethanol	5.85
Acetone (in nail polish remover)	24.6

Vapour pressure and volatility

Liquids with relatively high vapour pressure are described as being **volatile**, they evaporate rapidly. This indicates that the intermolecular forces holding their molecules together in the liquid are relatively small – it is easy to break these intermolecular forces so that the molecules escape easily. Substances with low vapour pressure have strong forces holding their molecules together – their intermolecular forces are high so they are not volatile (they do not change from a liquid to a gas easily).

Temperature and vapour pressure

Vapour pressure is affected by the temperature. As the temperature increases, the vapour pressure of a substance increases.

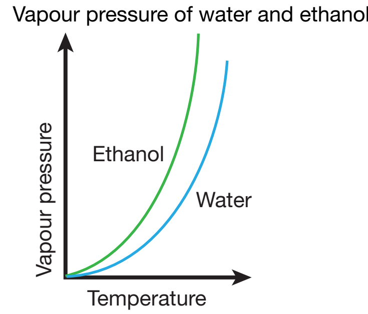


Figure 4.2 Vapour pressure and temperature.

This increase in vapour pressure with temperature occurs because at higher temperatures, more particles have enough energy to escape from the surface, as shown in Figure 4.3. As gas particles lose energy, they return to the liquid phase.

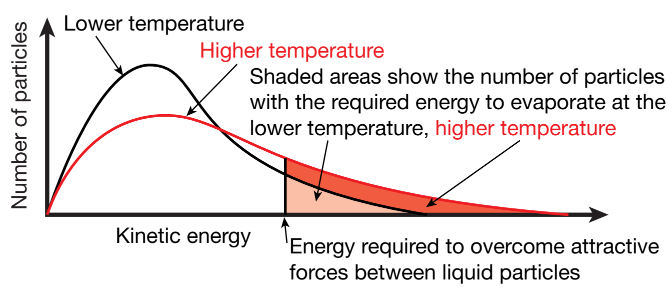


Figure 4.3 Kinetic energy and particles.

QUESTIONS

1. What is meant by the following terms?
   (a) Evaporation.
   (b) Volatile.
   (c) Vapour pressure.
2. Two liquids have vapour pressures of 2.34 kPa and 5.85 kPa at the same temperature. Which liquid probably has stronger intermolecular forces? Justify your answer.
3. Research the differences between evaporation and boiling.
4. Check your answer with this quick quiz.
   (a) Identify the term for changing from a liquid to a gas.
   (b) What do we call the pressure of a gas above a liquid?
   (c) Does vapour pressure increase or decrease when the temperature rises?
   (d) Does a volatile substance have a relatively high or low vapour pressure?
   (e) Would a high vapour pressure indicate weak or strong intermolecular forces.

5 Melting and Boiling Points

Melting and boiling points are determined by the forces/bonds between the particles of a substance.

Table 5.1 Melting and boiling points of different types of substances.

Type of substance	Type of bonds	Melting and boiling points
Ionic substances	Strong ionic bonds between ions in lattice.	High.
Metals	Strong metallic bonds between metal ions and mobile electrons in a lattice.	High.
Covalent network	Strong covalent bonds between atoms in lattice.	Very high
Covalent molecular	Weak forces between molecules.	Low.

Table 5.2 compares the strength of a covalent bond within molecules (the O–H bond) with the three types of intermolecular forces.

Table 5.2 Relative strength of forces.

Force	Typical energy (kJ mol^{-1})
O–H covalent bond	463
Hydrogen bond	20
Dipole-dipole force	10
Dispersion force	2

In this chapter we are focusing on melting and boiling points of covalent molecular substances.

Covalent molecular substances

The nature and strength of forces between the molecules of a covalent substance (intermolecular forces) can explain observable properties such as melting and boiling points.

Strong bonds between the molecules of a covalent molecular substance need more energy to break them so the melting and boiling points are higher.

If we look at a number of covalent molecular elements and compounds we find that they all have low melting and boiling points, compared to other types of substances. However, there is still a lot of variation within covalent molecular substances. For example, see Table 5.3.

Table 5.3 Melting and boiling points of covalent molecular substances.

Substance	Melting point (°C)	Boiling point (°C)
Neon (Ne)	−249	−246
Oxygen (O_2)	−219	−183
Methane (CH_4)	−183	−162
Ethanol (C_2H_5OH)	−114	78
Phosphorus trichloride (PCl_3)	−91	74
Water (H_2O)	0	100
Hydrogen sulfide (H_2S)	−83	−62
Ammonia (NH_3)	−78	−33

To explain the differences in melting and boiling points of these covalent molecular substances we need to look more closely at the forces between their molecules because melting and boiling points of covalent substances are determined by the strength of their intermolecular forces.

Strength of intermolecular forces

The strength of intermolecular forces depends on the following factors.

- The type of intermolecular force present.
- The size of molecules.
- The polarity of molecules.
- The shape of molecules.

The type of intermolecular force present. Covalent substances which are **non-polar** have only weak dispersion forces between their molecules, so they have the lowest melting and boiling points, e.g. hydrogen and methane.

Polar covalent substances have dispersion forces and also dipole-dipole forces, so more energy is needed to break the intermolecular forces – they have higher boiling and melting points. The more polar they are, the higher their melting and boiling points.

Substances that also have **hydrogen bonds** (compounds containing hydrogen as well as O, N or F) have even higher melting and boiling points, e.g. water.

The size of molecules. Table 5.4 compares the boiling points of group 7 halogens, whose molecules are only held together by dispersion forces. As the size of their molecules increases, the boiling point increases. With increasing mass, the dispersion forces increase, so more energy is needed to bring about a change in state.

Table 5.4 Boiling points of halogens.

Halogen	Molecular weight (amu)	Boiling point (K)
Fluorine F_2	38.0	85.1
Chlorine Cl_2	71.0	238.6
Bromine Br_2	159.8	332.0
Iodine I_2	253.8	457.6

The polarity of molecules. For molecules of approximately equal mass and size, the strengths of attraction between their molecules increases as the polarity increases.

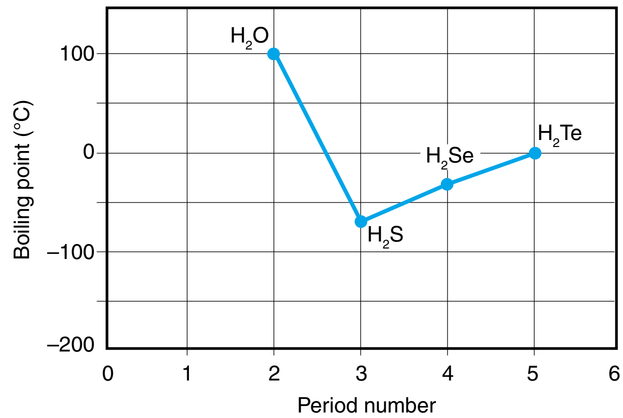
The shape of molecules also affects the size of intermolecular forces. Molecules which are spread out rather than round can have more surface contact and they have greater intermolecular forces between them.

QUESTIONS

1. Identify the substances in Table 5.3 which would be liquids at room temperature.
2. Earlier in the year you learned about ionic and metallic substances and also covalent network substances. Explain the high melting points of these substances compared to the lower melting points of covalent molecular substances.
3. The following table contains melting and boiling point data for the first eight alkanes. Graph the boiling point data, then explain the trends evident in the data.

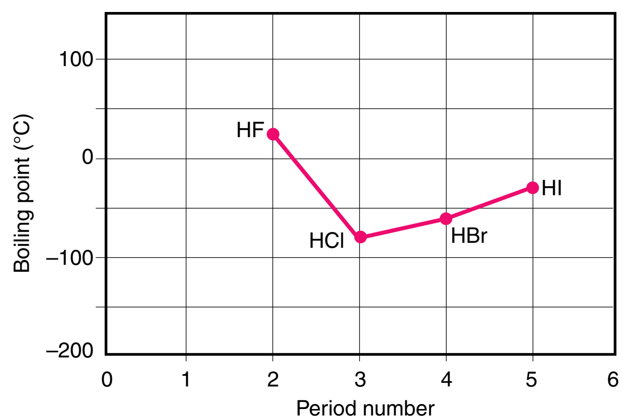
Name of alkane	Melting point (°C)	Boiling point (°C)
Methane	−183	−162
Ethane	−183	−89
Propane	−188	−42
Butane	−138	−1
Pentane	−130	36
Hexane	−95	69
Heptane	−91	98
Octane	−57	126

4. The graph shows boiling points of compounds formed between hydrogen and some elements in group 6.



(a) Group 6 elements react with hydrogen and form compounds which exist as covalent molecules. Based on this information, would you expect such compounds to have high or low boiling points?
(b) Describe any relationship between the boiling point and size of molecules illustrated by this graph.
(c) The boiling point of water does not fit this trend. Account for its higher than expected boiling point.

5. The graph below shows the boiling points of some compounds of group 7 elements with hydrogen. State as many facts and inferences as you can based on the information in this graph.



6. Copy and complete the following statement.
(a) Covalent molecular substances all have relatively (low/high) boiling points because they have (intermolecular/covalent) bonds that are (weak/strong).
(b) Intermolecular forces include .... forces, ... forces and .... bonds.
(c) Hydrogen bonds occur (between/within) molecules and they make boiling points of covalent molecular substances (higher/lower).

6 Solubility

The nature and strength of forces between the particles of a substance can explain observable properties such as solubility. In the case of covalent molecules, the relevant forces are intermolecular forces.

When one substance dissolves in another:

- The **solute** is the substance that gets dissolved.
- The **solvent** does the dissolving.
- A **solution** is formed – this is a homogeneous mixture.

Some solvents are polar, e.g. water, and others are non-polar, e.g. kerosene and petrol. Water is such a good polar solvent that it is sometimes called the **universal solvent**.

A substance dissolves when attractive forces form between its particles and particles of the solvent. In general, ionic compounds and polar molecular compounds tend to be soluble in **polar solvents** such as water. Non-polar molecular substances tend to be soluble in **non-polar solvents**, e.g. petrol.

To help you remember this, you can think that ‘like dissolves like’. However, this cannot be used as an explanation of why substances dissolve.

Solubility of covalent substances in water

Most covalent molecules that are non-polar, both elements and compounds, are insoluble in water or only slightly soluble.

Some elements such as **oxygen gas** (O_2) are partly soluble in water; the water molecules becoming attached to oxygen molecules by weak dispersion forces.

Sucrose (table sugar, $C_{12}H_{22}O_{11}$) however is a polar covalent compound and it dissolves in water. The intermolecular forces in the sugar crystal break, as do the hydrogen bonds between water molecules. New hydrogen bonds then form – this time between water molecules and sugar molecules, so sugar is soluble.

Some very polar covalent compounds such as hydrogen chloride, hydrogen bromide and hydrogen sulfide react with water molecules to form ions. This is called **ionisation**. (H_3O^+ is called a hydronium ion.)

$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ or

$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$

Notice that hydrogen chloride gas is covalent, but it forms ionic hydrochloric acid in solution.

Covalent network substances such as silicon dioxide are insoluble in water. Strong covalent bonds exist throughout the silicon dioxide network and water molecules cannot break these strong covalent bonds, so covalent network substances such as silicon dioxide are insoluble (cannot dissolve) in water.

Macromolecules, e.g. in polymers such as polyethylene are also insoluble. The very large molecules of polyethylene have very strong dispersion forces between their molecules. Water molecules cannot separate these large molecules so polymers are insoluble in water.

QUESTIONS

1. (a) Distinguish between the terms solute, solvent and solution.
   (b) Explain why water is called the ‘universal solvent’.
2. (a) Explain why sodium chloride is soluble in water and silicon dioxide is not.
   (b) To part (a) one student answered, ‘… because like dissolves like.’ Explain why this answer scored no marks.
3. Hydrogen chloride and sucrose are both molecular compounds and both dissolve in water. Explain why a solution of hydrogen chloride will conduct electricity and a solution of sugar will not.
4. Copy and complete the table.

Type of chemical	Example	Solubility
Polar molecular compound		
Molecular element		
Highly polar molecular compound		
Non-polar molecular compound		
Covalent network structure		
Macromolecules		

5. Research the solvent properties of ethanol (C_2H_5OH).
6. Check your knowledge with this quick quiz.
   (a) Is water a polar or non-polar solvent?
   (b) Polar solvents dissolve (non-polar/polar) solutes.
   (c) Polar molecules are (soluble/insoluble) in water.
   (d) Non-polar molecules are (soluble/insoluble) in water.

7 Shapes of Molecules

The shapes of molecules are determined by the arrangement of its atoms and this is determined by the electrons in the outer valence shell of the central atom. We will start by looking at the shapes of two very common molecules, carbon dioxide and water.

The carbon dioxide molecule

The Lewis electron dot structure of carbon dioxide (Figure 7.1) shows two pairs of bonding electrons on each side of the carbon atom, forming a double bond with each oxygen atom. The pairs of electrons are negatively charged, so they repel each other and try to get as far apart as possible. The atoms end up arranged in a straight line making a **linear molecule**.

Figure 7.1 Carbon dioxide – Lewis electron dot structure and shape of the molecule.

The water molecule

With the formula H_2O, you might also expect the shape of a water molecule to be linear: H – O – H. But the water molecule is not straight, it is a **bent molecule**. The reason for this bent shape lies in the arrangement of the valence electrons orbiting the central oxygen atom, which we can see in the Lewis electron dot structure.

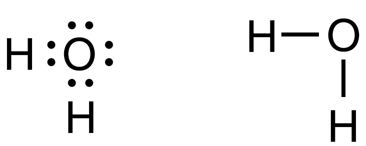


Figure 7.2 Lewis electron dot structure of water and bent molecule.

The outer shell of the oxygen atom contains 4 pairs of electrons: 2 bonding and 2 non-bonding pairs. A non-bonding pair of electrons is called a lone pair. These 4 pairs of electrons are all negatively charged, so they repel each other, arranging themselves as far apart as possible. The pairs of electrons end up arranged in the shape of a tetrahedron (Figure7.3). A **tetrahedron** is a type of pyramid with 4 identical triangular faces, 3 faces meeting at each corner or vertex.

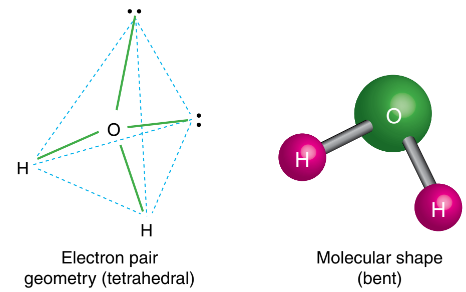


Figure 7.3 Electron pair arrangement and shape of the water molecule.

The shape of the molecule is determined by the positions of the two hydrogen atoms and the oxygen atom. The two hydrogen atoms are at two corners of the tetrahedron. The other two points of the tetrahedron are occupied by the two non-bonding electron pairs. Think of it this way – we cannot 'see' the electron pairs – all that is 'visible' is the bent shape formed by the hydrogen and oxygen atoms. The electron pairs are arranged like a tetrahedron, but the molecule itself is a bent shape.

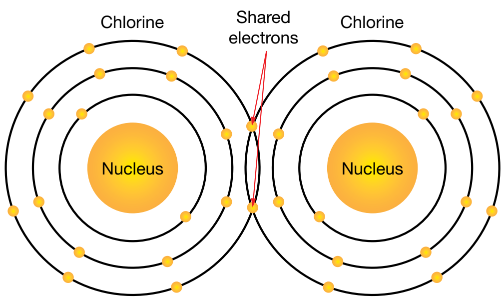
QUESTIONS

1. (a) Write the Lewis electron dot structure for water.
   (b) Identify the most stable geometrical arrangement of the electron pairs in an atom's valence shell.
   (c) Identify the shape of the water molecule.
   (d) Explain why the shape is not described as tetrahedral.
2. (a) Write the Lewis electron diagram for carbon dioxide.
   (b) Illustrate the shape of a carbon dioxide molecule.
3. What is meant by each of the following?
   (a) Linear molecule.
   (b) Tetradedron.
   (c) Lone pair.
4. Is it correct to say that, 'The hydrogen atoms in a molecule of water repel each other forming a bent shape'? Justify your response.
5. Check your knowledge with this quick quiz.
   (a) In the water molecule, the 4 electron pairs around the central oxygen atom (attract/repel) each other.
   (b) The 4 electron pairs around the central oxygen atom in the water molecule are arranged in the shape of a ....
   (c) Identify the shape of the water molecule.
   (d) Identify the shape of the carbon dioxide molecule.

Answers

1 Intramolecular Forces

1. (a)



(b) The bond is the attractive force between the two chlorine atoms and it is formed by the sharing of an electron from each atom. The shared electrons orbit both atoms, holding them together. A bond which involves shared electrons (rather than the electrons being completely donated or taken away) is called a covalent bond. These two chlorine atoms are identical to each other and so they exert the same pull on the two electrons that they share. This means that the two valence electrons are shared equally between the two chorine atoms so the bond is described as non-polar – it does not have a pole (charged end) because the charged particles are distributed evenly throughout the whole molecule.

2. (a) This could be a non-polar covalent bond in a molecule such as Cl_2 or F_2. Electrons are shared equally between the two identical atoms; the electron cloud is symmetrical.
   (b) This could represent a polar bond between atoms which are different to each other, e.g. in HCl or HF. In a polar covalent bond, electrons are shared unequally between two different atoms.
   (c) Diagram (c) could represent the attraction between an anion and a cation, e.g. between sodium and chloride ions in sodium chloride, Na^+Cl^-. In an ionic bond, electrons are not shared, they are transferred from one atom to another – from the sodium to the chlorine. (Later we will see that the difference in electronegativity is big enough for the most electronegative atom (e.g. chlorine) to completely remove the electron from the other atom (e.g. sodium.) The bond is an electrostatic attraction between charged particles (ions). This diagram shows two oppositely charged particles, a cation and an anion.

3.

Substance	Formula	Polar/non-polar bonds
Hydrogen	H_2	Non-polar
Ammonia	NH_3	Polar
Hydrogen bromide	HBr	Polar
Water	H_2O	Polar
Carbon dioxide	CO_2	Non-polar
Argon	Ar	Non-polar

4. Polar refers to having charged ends.
   A *polar bond* is an attraction between atoms where one atom is more electronegative and so attracts the shared pair of electrons more strongly. The electrons spend more time near the most electronegative atom so it develops a slight negative charge, e.g. δ^+ H–Cl δ^-
   Polar molecules refer to molecules which have charged areas. Hydrogen chloride only has the one bond – it is polar so the molecule is polar. Larger molecules are polar if they contain polar bonds which do not cancel each other out, e.g. water molecules.
   A *substance is polar* if it contains polar molecules.

5. A hydrogen molecule is non-polar. It consists of two identical hydrogen atoms which share a pair of electrons equally so no dipole is formed. A hydrogen fluoride molecule is polar – it contains two different atoms; the fluorine attracts electrons more strongly than hydrogen so the shared pair of electrons spends more time near the fluorine, forming a polar covalent bond and giving that end of the molecule a slight negative charge – an HF molecule is a dipole.

6. A tetrabromomethane molecule contains four polar bonds; they are symmetrical around the central carbon atom and cancel each other out.
7. (a) True. Oxygen molecules are made of two identical oxygen atoms so they share the bonding pairs of electrons equally.
   (b) False. Sodium chloride does not contain or consist of molecules. Molecules have covalent bonds due to the sharing of electrons. In sodium chloride there are no shared electrons; an electron is transferred from the sodium atom to the chlorine atom forming oppositely charged ions which are attracted together. Sodium chloride has ionic bonds, not covalent bonds.
   (c) False. Chlorine does contain covalent bonds (it has shared electrons) but they are non-polar as the electrons are shared equally.
   (d) True. Covalent bonds are strong attractive forces within molecules.
8. (a) Different.
   (b) Non-polar.
   (c) Polar.
   (d) Shared.

2 Intermolecular Forces

1. (a) Dispersion forces occur when the electron distribution in atoms of a molecule sets up temporary dipoles. These can then induce temporary dipoles in nearby molecules. Attractive forces between positive and negative ends of these temporary dipoles are the dispersion forces. They are constantly forming, breaking up and re-forming.
   (b) The size and shape of the molecule. The bigger the molecule and thus the more electrons present, the greater the dispersion forces will be. The shape is also important, e.g. long molecules have greater dispersion forces than short fat molecules with the same number of electrons. The long thin molecules develop greater dispersion forces and can also lie closer together – attractive forces are most effective when molecules are very close.
2. Dispersion forces are temporary and dipole-dipole forces are permanent. Dispersion forces are generally considered to be weaker than dipole-dipole forces, although the size of forces does vary with size and shape of the molecule.
3. (a) A dipole is a structure with two oppositely charged areas.
   (b) A dipole occurs when the electrons are shared unevenly between atoms in a polar molecule, when there is a momentary unequal distribution of electrons within an atom due to the movement of electrons, and when a charge is induced in a molecule by a nearby charged particle.
   (c) A dipole-dipole force occurs when two oppositely charged ends of polar molecules are close enough to attract each other.
4. The first diagram shows an oxygen molecule consisting of two identical atoms, with evenly distributed electrons. There is no charge. The second diagram shows an oxygen molecule with a slight positive charge at one end of the molecule and a slight negative charge at the other end. The negatively charged end is shown enlarged due to more electrons being at this end, distorting the electron cloud. This suggests it has come near the positive end of a charged molecule (a dipole) which has attracted its electrons to the end shown as negatively charged, thus the oxygen molecule has temporarily become an induced dipole.
5. No. Sodium chloride is ionic. It does not form molecules at all, so there can be no forces between molecules.

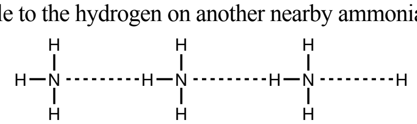
6. Group A – Dispersion forces only – chlorine, nitrogen, ethane.
   Group B – Dispersion and dipole-dipole forces – sulfur dioxide.
7. If the polar bonds are oriented so that they cancel each other out, then the molecule does not contain slightly positive (δ^+) and slightly negative (δ^-) areas so it is a non-polar molecule, e.g. CO_2 (O=C=O). The presence of polar bonds does not necessarily mean the molecule is polar, its overall polarity also depends on the shape of the molecule as polar bonds can cancel each other out leaving a net dipole of zero.
8. Various – geckoes use dispersion forces to cling to and climb over surfaces.
9. (a) False. Dispersion forces are temporary electrostatic attractions between molecules. They develop because of temporary changes in the electron distribution within molecules.
   (b) True. Dispersion forces occur between all covalent substances as electron clouds are moving in all of them.
10. (a) Between. (b) Dispersion forces.
    (c) Dispersion forces. (d) Dipole-dipole forces.
    (e) Dispersion forces. (f) Polar molecules.

3 More Intermolecular Forces – Hydrogen Bonds

1.

Dispersion forces	Dipole-dipole forces	Hydrogen bonds
Temporary	Permanent	Permanent
Caused by constant changes in electron distribution in atoms.	Forces between polar molecules. Slightly positive end of one dipole attracts the slightly negative end of another dipole.	Attractive force between hydrogen atom in one molecule and a nitrogen, fluorine or oxygen atom in a nearby molecule.
Weakest force (but depends on size and shape of molecule).	Stronger than dispersion forces.	Strongest intermolecular force.

2. (a) A hydrogen bond is the electrostatic attraction between the slightly positive hydrogen atoms of one molecule and a highly electronegative atom (such as oxygen, nitrogen or fluorine) in another nearby molecule.
   (b) Hydrogen bonds occur from the nitrogen atom in one ammonia molecule to the hydrogen on another nearby ammonia molecule.



_____ Covalent bond
...... Hydrogen bond

   (c) The student's answer is not correct. Hydrogen bonds do not hold atoms together within a molecule. Hydrogen bonds are forces of attraction between some molecules – they are intermolecular electrostatic forces. Hydrogen bonds occur between a hydrogen atom in one molecule and an electronegative atom in a different molecule. The electronegative atom could be oxygen, fluorine or nitrogen.

3.

Substance	Dispersion forces	Dipole-dipole forces	Hydrogen bonds
Helium He	Yes	No	No
Oxygen O_2	Yes	No	No
Water H_2O	Yes	Yes	Yes
Methane CH_4	Yes	No	No
Hydrogen sulfide H_2S	Yes	Yes	No
Ammonia NH_3	Yes	Yes	Yes
Hydrogen fluoride HF	Yes	Yes	Yes
Acetic acid CH_3COOH	Yes	Yes	Yes
Carbon dioxide CO_2	Yes	No	No
Hydrogen chloride HCl	Yes	Yes	No
Bromine liquid Br_2	Yes	No	No
Ethanol C_2H_5OH	Yes	Yes	Yes

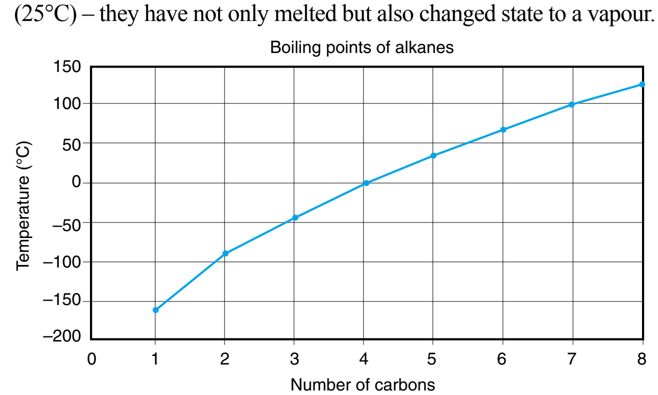
4. (a) False. Hydrogen bonds occur between water molecules, not inside the molecules. Hydrogen bonds occur between a hydrogen atom in one molecule and an oxygen atom in another molecule.
   (b) False. Dispersion forces are the weakest intermolecular attractions. Dipole-dipole forces are stronger than dispersion forces.
5. Water, hydrogen sulfide, ethane, argon.
6. (a) Hydrogen bonds.
   (b) Various, e.g. water, ammonia, hydrogen fluoride.
   (c) Hydrogen and either oxygen, nitrogen or fluorine.

4 Vapour Pressure

1. (a) Evaporation means changing from the liquid state to the gaseous state. This happens when particles at the surface of a liquid have enough energy to break free.
   (b) Volatile means easily able to change from a liquid to a gas – they change state at relatively low temperatures.
   (c) Vapour pressure means the pressure of a gas above a liquid. If the liquid is in a closed container, the pressure is the force per unit area that is exerted on the walls of a container due to the constant movement of particles of a gas.
2. The liquid with the higher vapour pressure of 5.85 kPa would have weaker intermolecular forces. A higher vapour pressure indicates that the substance has weaker intermolecular forces of attraction. Therefore the liquid with the lower vapour pressure of 2.34 kPa has the stronger intermolecular forces – it vaporises less easily.
3. Evaporation occurs only at the surface of a liquid. A substance boils when the vapour pressure of the liquid equals the atmospheric pressure of the surroundings. When a substance boils, continuous change of state takes place throughout the liquid, not just at the surface; bubbles of water vapour form within the water as it becomes hot and boils and rise to the surface.
   Another difference is that evaporation occurs at any temperature whereas boiling occurs at a specific temperature.
4. (a) Evaporating.
   (b) Vapour pressure.
   (c) Increase.
   (d) High.
   (e) Weak.

5 Melting and Boiling Points

1. Ethanol, phosphorus trichloride, water.
2. Ionic compounds and metallic elements require strong electrostatic bonds to be broken (ionic bonds and metallic bonds) for them to melt. Covalent network substances have strong covalent bonds extending throughout their networks and a lot of energy is also needed to break these bonds. Thus these types of substances have high melting points. Covalent molecular substances on the other hand only have comparatively weak intermolecular forces that need to be broken in order for them to melt. Their melting points are much lower. In fact, many covalent molecular substances are gases at room temperature (25°C) – they have not only melted but also changed state to a vapour.
3.



   Trend: As the size of the molecule increases, the melting and boiling points increase. This happens because with increase in size, the intermolecular forces also increase, making more energy (a higher temperature) necessary to overcome these forces.
4. (a) Low boiling points.

(b) Three of the compounds, H_2S, H_2Se and H_2Te show an increase in boiling point (from about −60 to −40, to 0°C) with increase in size of the molecule. Water does not fit this trend – its boiling point is much higher than would be expected based on size of the molecules alone.

(c) H_2Te, H_2Se and H_2S all have low boiling points (approximately 0°C, −40°C and −60°C respectively) but water has a much higher boiling point than would be expected (100°C). All four compounds, H_2Te, H_2Se, H_2S and H_2O have dispersion forces and dipole-dipole intermolecular forces attracting their molecules. These are weak forces, so a small rise in heat is able to break these bonds and allow them to move more freely so that a change of state occurs. Water would have not only these forces, but also hydrogen bonds. Hydrogen bonds are attractive forces between a hydrogen atom in one molecule and an oxygen atom in an adjacent molecule. Hydrogen bonding would cause a higher boiling point because to break these stronger hydrogen bonds more energy would be needed – the boiling point would be higher.

5. Various, e.g.
   - The boiling points of these compounds are relatively low – most are below zero Celsius.
   - Boiling points decrease as you move up the group from HI to HBr to HCl.
   - Hydrogen fluoride (HF) has the highest boiling point and is an exception to the trend of boiling point decreasing as you go up the group.
   - As you go up the group, the molecules get smaller. Smaller molecules have weaker dispersion forces.
   - If HF followed the trend, its boiling point would be about −100°C, but instead it is much higher at about 20°C.
   - As the boiling point of HF is higher than expected, its bonding must be stronger than in the other compounds shown in the graph. As hydrogen fluoride contains hydrogen and fluoride, it would have intermolecular hydrogen bonds, between H in one molecule and F in an adjacent molecule. These bonds would be in addition to the dispersion forces and dipole-dipole forces found in all the compounds.
   - The other three compounds, HI, HBr and HCl, contain hydrogen, but they do not have O, F or N atoms, so they cannot form hydrogen bonds. They would only have dispersion forces and dipole-dipole forces between their molecules.
   - This difference in bonding would account for hydrogen fluoride having a higher boiling point than the other compounds on the graph.

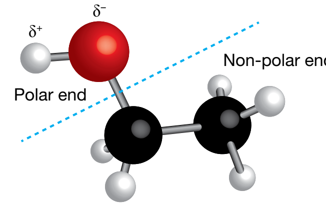
6. (a) Low, intermolecular, weak.
   (b) Dispersion, dipole-dipole, hydrogen.
   (c) Between, higher.

6 Solubility

1. (a) The solute is the substance that gets dissolved, the solvent does the dissolving and the solution is a mixture formed when a solute dissolves in a solvent.
   (b) Water can dissolve a wide variety of substances including ionic and polar covalent substances.
2. (a) Sodium chloride is ionic, and water is polar and has dipoles. The water dipoles can develop attractive forces with the sodium and chloride ions and thus dissolve it. Silicon dioxide is a covalent network structure. It does not form any dipoles and has no ions so there is no attraction to water molecules, it is insoluble.
   (b) 'Like dissolves like' is just a saying to help remember which substances dissolve and which do not. It does not provide an explanation. An acceptable answer must mention the type of bonding in each substance involved and explain either how bonds are formed between a solute and a solvent when dissolving occurs, or else why they cannot form (in the case of an insoluble substance).
3. Hydrogen chloride dissolves in water to form ions in solution, which can then carry a charge – it conducts electricity and is called an electrolyte. Sugar molecules spread out through the water, bonding with the water, but no ions are formed so the solution cannot conduct an electric current (it is a non-electrolyte).
4.

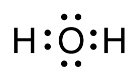
Type of chemical	Example (various)	Solubility
Polar molecular compound	Sucrose	Soluble
Molecular element	Oxygen	Slightly soluble
Highly polar molecular compound	Hydrogen chloride	Soluble
Non-polar molecular compound	Hexane	Insoluble
Covalent network structure	Silicon dioxide	Insoluble
Macromolecules	Polyethylene	Insoluble

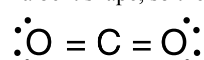
5. Ethanol is able to dissolve both polar and non-polar substances because its molecule has a polar end and a non-polar end. The ethyl (C_2H_5) group at one end is non-polar, allowing it to attract other non-polar molecules. The hydroxyl (OH) group at the other end is polar and the high electronegativity of oxygen allows it to form hydrogen bonding with molecules containing hydrogen. Ethanol is able to be used as a solvent in cosmetics such as perfumes, food colourings and flavourings, medical products and cleaning agents. Methylated spirits is mainly ethanol with the addition of a toxic substance so it cannot be used as a drink.



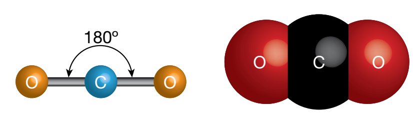
6. (a) Polar. (b) Polar.
   (c) Soluble. (d) Insoluble.

7 Shapes of Molecules

1. (a) Water. 

   (b) In the shape of a tetrahedron.
   (c) Bent.
   (d) The four pairs of electrons around the oxygen atom are arranged in a tetrahedral shape, but two pairs are lone pairs and these are not involved in bonding so there are no atoms attached to them. The 'visible' parts of the molecule are the oxygen atom and the two hydrogen atoms associated with the bonding pairs. The shape of a molecule is determined by the position of its atoms, in this case an oxygen and two hydrogen atoms. These form a bent shape, so the molecule is bent.
2. (a) Carbon dioxide. 

   (b) Carbon dioxide is a linear molecule.



3. (a) A linear molecule has its atoms arranged in a straight line.
   (b) A tetrahedron is a type of pyramid with 4 identical triangular faces, 3 faces meeting at each corner or vertex.
   (c) A lone pair is a pair of outer shell electrons that are not involved in bonding,
4. Not correct. The shape of the water molecule is bent, but it is the negatively charged electron pairs that repel each other, not the hydrogen atoms.
5. (a) Repel. (b) Tetrahedron.
   (c) Bent. (d) Linear.

8 The VSEPR Theory

1. (a) According to the VSEPR – valence shell electron pair repulsion – theory the shapes of molecules are determined by the arrangement of its atoms and the arrangement of the atoms in a molecule is influenced by its outer shell electrons. These electron clouds around each atom repel each other, moving as far away from each other as possible while still being attracted to the positive nucleus. A lone pair can repel more strongly than a bonding pair. These forces determine where electrons will be at any moment and thus they determine the arrangement of atoms and the shape of the molecule.