### Student Notes

**Topic 1:**

**Periodicity, Structure & Bonding**

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The pace of this course has to be fast-moving if the total content of each unit is to be covered in the time available.

You have to be prepared to take much more responsibility for your own learning and be actively involved - learning is not a passive activity, it is not done to you, it is something you achieve.

It starts with each new lesson - listen carefully to what your Teacher tells you read carefully what is in your Study Guide watch carefully any videos or animations carry out all practical work thoughtfully

The questions in your Study Guide have been carefully put together to help you think and reason correctly. It is very important that you use these opportunities. Discuss your thought with your Group and help each other achieve the necessary understanding.

Only fill in your Notes once you have completed the relevant part of your Study Guide. Filling in Notes alone will not teach you Chemistry.

Self-Check questions are aimed directly at the ideas met in that particular lesson. You should attempt them as soon after finishing that lesson as possible and should not proceed to next lesson until you have shown you understood the previous lesson.

Research clearly shows that your brain benefits from a ‘top-up’ a few hours after a lesson. You should, therefore, try and complete these questions the same evening after a lesson is completed.

By the end of a Topic, you should be able to answer questions that require understanding from several lessons. These questions will be more like exam questions.

You must be prepared to, firstly, maximise the opportunities provided in school - use your teacher! - but also be prepared to continue working at home:—

finishing off work not completed in class reading over Notes
completing work missed through absence Self-Checks
filling in Notes reading Text Book Home Practices
1.1 Metallic & Monatomic Elements

This first lesson looks at bonding in metals, and how this helps explain some of the typical properties of a metal, and the monatomic elements and the very weak type of bonding force that exists between atoms of these elements.

Metallic Bond

The inner elec of a metal atom are tightly held by the nuc. They are not able to escape from the atom. They are said to be loc.

The outermost elec are also strongly attracted by the nuc, but they are free to move from atom to atom. They are said to be deloc.

Each delocalised electron can be attr to the nuc of more than one neighbouring atom.

The metallic bond is the electrostatic attr between the neg charged outermost electrons and the pos charged nuclei of neighbouring atoms.

The metallic bond is usually strong and acts in all directions around each atom.

In a solid metal, the atoms are arra neatly in a simple pattern known as a lattice.

Physical Properties

The Physical properties of a substance are those properties which do not involve chemical reactions.

Metals make good cond.

Since the outermost electrons are deloc, it is easy to make them all move in the same direction.

Once a volt is applied across the metal, the elec will move away from the neg electrode towards the pos electrode.

Met bonding is very ‘fluid’. This makes it easy for metal ato to be slid around each other and yet remain strongly bon together. Metals are described as malleable.
Metals can be drawn into thin wires by stretching. This property is called ductility.

Again, it is because it is easy for metal atoms to slide around each other while staying strongly bonded to each other. Metals are described as ductile.

**Boiling Points**

To boil a metal, and allow its atoms to move around freely as a gas, the metallic bond must be overcome.

The Boiling Points of metals are a guide to the relative strengths of the metallic bonds holding the atoms together.

As you go down Group 1, the outermost electron is further and further away from the nucleus. The strength of attraction between electrons and nuclei will, therefore, decrease. This makes it easier for the metal atoms to move further apart as shown by the decrease in Boiling points.
As you go across the fourth period, the outermost electron are always to be found in the fourth shell, at a similar distance from the nucleus. The number of electrons is, however, increasing.

Since the Boiling points are increasing, the bond strength must increase as the number of outermost electrons in a metal increases.

The strength of the Metallic bond decreases as the outermost electrons become further from the nucleus. It increases as the number of outermost electrons increase.

At very high temperature, metals can be vaporised and the atoms will be too far apart to allow delocalised electrons to move between the now single atoms. Attractions still exist, but they are of a much weaker type.

**Very Weak Attractions**

The Noble gases are the only elements normally found as single atoms, monatomic.

At extremely low temperature, elements like Neon become liquid or even solid, like every other substance. At these low temperatures, there must be some weak forces holding these atoms together. Very little energy is needed to overcome these weak forces, as the low Melting and Boiling Points for the Noble gases demonstrate. These very weak bonding forces are called London Dispersion Forces.

Normally, the electrons are spread evenly within the set of electron clouds (orbits) that make up an electron shell. The centre of negative charge and the centre of positive charge will coincide at the centre of the atom.

Uneven distribution of the electrons causes the electron cloud to ‘wobble’ and the centre of negative charge no longer coincides with the positive charge. One side of the atom is temporarily slightly negative (δ−), while the other side becomes temporarily positive (δ+). We call this a temporary dipole.

The dipole is only temporary as electron repulsions quickly act to ensure a more even distribution. However, the brief wobbling on one atom can affect the electron clouds on any nearby atoms.
This causes, or induces, a temporary dip in the nearby atom. Short lived attr will then exist between the slightly neg (δ−) side of one atom and the slightly pos (δ+) side of the other atom.

**London Dispersion Forces** are the shortlived attractions that exist between a slight charge on one atom (temporary dipole) and the opposite charge set up on a nearby atom (induced dipole).

**Noble Gas Series**

Every atom has a nuc and clouds of elec around it. So all atoms have London Dispersion forces acting between them (as long as they are close enough to each other).

They are such sm forces that they can usually be ignored, especially if other, much stro bonds, bonding forces are present, e.g. met bonds.

However, for the noble gases, London Dispersion forces are the only type of bonding force present. So the melting points and boiling points of the noble gases depend entirely on the strength of the London Dispersion forces between their atoms.

The lar an atom is, the fur out its outermost elec, the more likely it is that the electron clouds will be able to ‘wobble’ and set up a temp dipole.

The lar the dipole set up, the stro the London Dispersion forces that will exist. As the attr between atoms increase, the ene needed to move the atoms further apart will inc as shown by the inc in Boiling points.

London Dispersion forces are very small, but they are larger between bigger atoms.
**SELF CHECK**

**Q1.** Which of the following represents a metallic lattice?

![Metallic Lattice Options](image)

- A. positioned entirely within one atom
- B. fixed in position between two atoms
- C. repelled by positive charges
- D. free-moving across a number of atoms

**Q2.** Delocalised electrons are

- A. positioned entirely within one atom
- B. fixed in position between two atoms
- C. repelled by positive charges
- D. free-moving across a number of atoms

**Q3.** The atoms in a metallic lattice are held together by

- A. positive metal ions attracting negative metal ions
- B. the common attraction of delocalised electrons for neighbouring positive cores
- C. the strong electrostatic attraction between fixed electrons and delocalised electrons
- D. sharing pairs of electrons

**Q4.** Which of the following is **not** a physical property of a metal?

- A. reactivity with oxygen
- B. melting point
- C. conductivity
- D. malleability

**Q5.** Solid metals are able to conduct electricity because

- A. their outermost electrons are free to move
- B. their inner electrons are free to move
- C. they have ions which can move in the lattice
- D. their atoms can move about in the lattice

**Q6.** Metals are malleable and ductile because the metallic bond

- A. is easily broken
- B. forms again quickly
- C. is fixed in one direction
- D. can move around as an atom is moved

**Q7.** The strength of a metallic bond increases as the

- A. charge on the electron increases
- B. outer electrons get further from the nucleus
- C. number of delocalised electrons increases
- D. size of the inner core gets larger

**Q8.** On going down the group 1 metals the boiling points

- A. increase because the sizes of their atoms become larger
- B. increase because the number of energy levels increases.
- C. decrease because there are fewer delocalised electrons
- D. decrease because the outermost electron is further from the nucleus
Q1. Noble gases are described as monatomic because they all
   A form molecules in which every atom is the same
   B exist as single uncombined atoms
   C have very stable electron arrangements
   D belong to the same column in the periodic table

Q2. London Dispersion forces are best described as
   A very weak repulsions
   B quite strong repulsions
   C very weak attractions
   D quite strong attractions

Q3. London Dispersion forces are caused by
   A electrons changing level in an atom
   B electrons moving from atom to atom
   C the nucleus of an atom changing charge
   D electron clouds in atoms wobbling

Q4. Which diagram shows correctly the presence of London Dispersion forces?
   A
   B
   C
   D

Q5. A temporary dipole
   A a momentary separation of charge
   B a fleeting attraction between atoms
   C an induced positive charge
   D an uneven spread of electrons

Q6. The sign δ+ means
   A a permanent positive charge
   B a temporary positive charge
   C an induced positive charge
   D a slight positive charge

Q7. What effect does the size of an atom have on London Dispersion forces?
   A the larger the atoms the smaller the force
   B the larger the atoms the greater the force
   C size makes no difference at all
   D the strongest force is between the smallest atoms

Q8. On going down the noble gases in the periodic table, the boiling points
   A decrease because electron clouds wobble less in larger atoms
   B decrease because the number of dipoles decreases with atomic size
   C increase because electron clouds wobble more in larger atoms
   D increase because the number of dipoles increases with atomic size
Q1. The diagram illustrates the nature of the metallic bond.

<table>
<thead>
<tr>
<th>negatively</th>
<th>mobile</th>
<th>electrostatic</th>
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<tbody>
<tr>
<td>positively</td>
<td>ductile</td>
<td>neighbouring</td>
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Select the appropriate words to complete the following sentence description.

The metallic bond is the ... (a) ... attraction between the ... (b) ... charged ... (c) ... electrons and the ... (d) ... charged nuclei of ... (e) ... atoms

Q2. Trends in boiling points of metals can be related to their bonding and position in the periodic table.

Select the appropriate word from *increases* / *decreases* to complete the following sentence about alkali metals.

On going down column 1, the boiling point of these metals ... (a) .... because the distance of the outermost electron from the nucleus ... (b) .... and therefore the strength of the metallic bond ... (c) ....

Q3. Gold is a good example of a metal which is malleable.

a) What does malleable mean?  

b) How is it possible for gold atoms to move without them becoming separated?

Total (10)
**Home Practice 1.1b**

**Q1.** This diagram shows a neutral atom in which a small dipole has arisen.

- **a)** What do the $\delta^+$ and $\delta^-$ signs stand for? 1
- **b)** What does the term **dipole** mean? 1
- **c)** What causes the dipole to arise? 1
- **d)** Why is this kind of dipole said to be temporary? 1

**Q2.** A momentary dipole in one atom can induce a momentary dipole in a neighbouring atom.

- **a)** What effect does the $\delta^+$ side of the first atom have on its neighbour? 1
- **b)** Draw the diagram of the two atoms to show the induced dipole. 1
- **c)** How does this lead to attraction between the two atoms? 1
- **d)** What is this type of attractive force called? 1

**Q3.** Of the monatomic elements, radon has the largest atoms.

How does this explain why radon has the highest melting point of these elements? 2

**Total** (10)
1.2 Molecular & Network Elements

This lesson is about the elements which exist normally as covalent discrete molecules or joined together with many strong covalent bonds to form vast networks.

### Diatomic elements

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<tbody>
<tr>
<td>$H_2$</td>
<td>$N_2$</td>
<td>$O_2$</td>
<td>$F_2$</td>
<td>$Cl_2$</td>
</tr>
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</table>

Molecular elements are elements made up of individual molecules i.e. distinct groups of atoms joined together by **covalent** bonds.

The simplest molecular elements are **diatomic** i.e. each molecule is made of just **two atoms**.

Substances which are molecular have **more than one type** of bonding present.

**Within** each molecule the atoms are **strongly** held together by a **covalent** bond.

A **covalent bond** is the electrostatic attraction between a **shared pair** of electrons and the **nuclei** of the two sharing atoms.

Usually, the **covalent bond** is represented by a line.

**Between** the **molecular** units of fluorine there are the same **weaker forces** that exist between atoms - **London Dispersion forces**.

Though very weak, these **intermolecular forces** can still determine **physical properties** such as **melting** and **boiling points**.

![Graph showing melting points of diatomic elements](image)
The larger a molecule is, the further out the outermost electrons, the more likely it is that the electron clouds will be able to ‘wobble’ and set up a temporary dipole. The larger the dipole set up, the stronger the London Dispersion forces that will exist.

As the attraction between molecules increase, the energy needed to move the molecules further apart will increase as shown by the increase in Melting points.

**Larger Covalent Molecules**

The size of atoms has an effect on the strength of London Dispersion attractions between molecules. The number of atoms can also have an effect.

Chlorine molecules have the formula \( \text{Cl}_2 \) and a melting point = \(-101 \) °C.

Phosphorus molecules have formula = \( \text{P}_4 \) and a melting point = \( 44 \) °C.

Sulphur molecules have formula = \( \text{S}_8 \) and a melting point = \( 113 \) °C.

Chlorine, phosphorus and sulphur atoms are of a similar size and have similar numbers of electrons so are likely to produce similar dipoles, and similar London Dispersion forces of attraction. However, with more atoms in a molecule, a larger number of London Dispersion forces are possible.

In all these examples, it is important to remember that there is no need to break the very strong covalent bonds within these molecules.

Larger molecules like fullerene will have even higher melting points, though, they are still relatively low compared to other substances which have stronger bonds than London Dispersion forces to be broken.

The strength of London Dispersion forces increases as the size of the atom increases.

increases as the number of atoms in the molecule increases.
Diamond & Graphite Structures

This activity is about the network structures of the diamond & graphite forms of carbon and how these explain their different properties and uses.

The arrangement of carbon atoms can be described as tetra because of the position of the 4 carbon atoms around a central carbon atom.

Rings of 6 carbon atoms, hexa, can also be seen.

Dias are not malleable - the carbon atoms cannot slide around each other without breaking strong covalent bonds.

Diamonds are strong and hard since many strong covalent bonds would need to be overcome if a diamond is to be cut or broken.

Unusually, car has two net structures - dia and gra. Their properties are very different, and so must their structures be.

In gra, each carbon atom is only bonded to 3 other carbon atoms by strong covalent bonds. This forms flat she of 6 membered rings (gons). The fourth outermost electrons are free to move (delocalised) between all the carbon atoms in the same sheet or layer.

Between the layers are the much weaker London Dispersions forces that exist between all atoms.

The distance between the layers is greater than the distance between carbon atoms within the same layer.

We can use the structure of graphite to explain its properties.

Conductivity gra is able to conduct along layers because it has some delocalised electrons. (Conductivity between layers is not possible).

Lubrication powdered graphite makes an excellent solid lubricant because the weak London Dispersions forces allow the layers to slide over each other.
Other Networks

The only other covalent networks in the first 20 elements are boron and silicon (which has a diamond structure like carbon).

Not surprisingly they are all high melting point solids.

A lot of strong covalent bonds have to be broken to melt these elements.

Melting point of boron = °C
Melting point of silicon = °C
Melting point of carbon = °C

Boron and silicon are also very hard substances, though not as hard as diamond, of course.

The atoms in these substances still wobble but London Dispersion forces can be ignored as the many many strong covalent bonds determine both the physical and chemical properties.

Elements 1 - 20 Structures

Higher pupils will be expected to know in detail the bonding structure of the first 20 elements and be able to describe what kind of bonds are involved in both physical and chemical changes.

Metallic Lattices (Networks)

Intramolecular = metal bonds (delocalised electrons)

Intermolecular = metal bonds (delocalised electrons)

Physical properties - determined by strength of metal bonds

Chemical Properties - determined by strength of metal bonds

Covalent Networks

Intramolecular = covalent bonds (shared pairs of electrons)

Intermolecular = covalent bonds (shared pairs of electrons)

Physical properties - determined by strength of many many strong covalent bonds

Chemical Properties - determined by strength of many many strong covalent bonds
**Chemical Changes & Structure**

**Topic 1 - Bonding & Periodicity**

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**Covalent Molecular**

- **Intramolecular** = cov bonds (sh pairs of electrons)
- **Intermolecular** = Lon Disp Forces (wob electrons)

**Physical properties** - determined by strength of Lon Disp forces

**Chemical Properties** - determined by strength of cov bonds

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**Single Atoms (Monatomic)**

- **Intramolecular** = none
- **Intermolecular** = Lon Disp Forces (wob electrons)

**Physical properties** - determined by strength of very weak Lon Disp forces

**Chemical Properties** - determined by strength of very weak Lon Disp forces

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**Structures of Elements 1 - 20**
Chemical Changes & Structure

Self Check

Q1. In which of the following are all three examples of molecular elements?
   A. Hydrogen, helium, lithium
   B. Potassium, sulphur, fluorine
   C. Chlorine, sulphur, phosphorus
   D. Carbon, silicon, boron

Q2. Covalently bonded atoms are held together by electrostatic forces of attraction between
   A. a shared pair of electrons and the nuclei of the two atoms
   B. the oppositely charged nuclei and delocalised outer electrons
   C. pairs of electrons in the outer energy level of each atom
   D. the nuclei of the two atoms and the inner core of electrons

Q3. In a molecular element
   A. all molecules are joined to each other by covalent bonds
   B. London dispersion forces link atoms within each molecule
   C. covalent bonds only exist within distinct molecules
   D. all the atoms in the structure form a giant covalent lattice.

Q4. An intermolecular force acts between
   A. atoms of the same kind inside a molecule.
   B. one molecule and another
   C. any two parts of the same molecule
   D. atoms which are different within a molecule.

Q5. Phosphorus is easily melted because
   A. only London dispersion forces are acting between its molecules
   B. its molecules do not have any covalent bonds
   C. the covalent bonding in molecular elements is very weak
   D. the London dispersion forces only act within the molecule

Q6. The melting point of iodine is greater than that of chlorine because iodine has
   A. more covalent bonds
   B. larger atoms in its structure
   C. stronger covalent bonds
   D. more London dispersion forces between its molecules

Q7. The melting point of sulphur is greater than that of chlorine because sulphur has
   A. more covalent bonds
   B. larger atoms in its structure
   C. stronger covalent bonds
   D. more London dispersion forces between its molecules

Q8. Which diagram illustrates the molecular structure of a fullerene?
**Self Check**

Q1. In a covalent network
   A molecules are linked by weak forces
   B covalent bonds extend throughout the structure
   C covalent bonds exist within distinct molecules
   D covalent bonds link ions into one network

Q2. Three elements which have covalent network structures are
   A lithium, carbon, and nitrogen
   B boron, carbon, and nitrogen
   C carbon, silicon, and lead
   D carbon, silicon, and boron

*Questions 3 and 4 refer to these diagrams*

Q3. Which diagram shows the structure of silicon?

Q4. Which diagram shows the structure of graphite?

Q5. Why is the melting point of diamond so high?
   A all its atoms are bonded to other atoms by covalent bonds
   B it has covalent bonds which are much stronger than normal covalent bonds
   C its atoms have a much greater attraction for bonding electrons than in other elements.
   D delocalised electrons make it much more difficult to separate its atoms.

Q6. Diamond is very hard because it has
   A covalent bonds throughout in all directions
   B covalent bonds within layers only
   C atoms which are very hard
   D molecules which interlock with each other

Q7. Graphite conducts electricity because it has electrons which
   A move freely along layers
   B are shared in covalent bonds
   C move freely from layer to layer
   D form covalent bonds in all directions

Q8. Graphite is used as a lubricant because it
   A conducts electricity
   B has delocalised electrons
   C has layers which can slide
   D is made entirely of carbon atoms
### SELF CHECK

**Q1.** Which type of bonding structure is present in the element *magnesium*?
- A covalent network
- B covalent molecular
- C metallic lattice
- D monatomic

**Q2.** Which type of bonding structure is present in the element *silicon*?
- A covalent network
- B covalent molecular
- C metallic lattice
- D monatomic

**Q3.** Which type of bonding structure is present in the element *argon*?
- A covalent network
- B covalent molecular
- C metallic lattice
- D monatomic

**Q4.** Which type of bonding structure is present in the element *phosphorus*?
- A covalent network
- B covalent molecular
- C metallic lattice
- D monatomic

**Q5.** Which type of bond would explain the difference in melting points between sulfur and phosphorus?
- A metallic
- B covalent
- C ionic
- D London dispersion forces

**Q6.** Which type of bond would explain the difference in reactivity between sulfur and phosphorus?
- A metallic
- B covalent
- C ionic
- D London dispersion forces

**Q7.** Which type of bond would explain the difference in melting points between boron and carbon?
- A metallic
- B covalent
- C ionic
- D London dispersion forces

**Q8.** Which type of bond would explain the difference in reactivity between boron and carbon?
- A metallic
- B covalent
- C ionic
- D London dispersion forces
HOME PRACTICE 1.2A

Q1. Sulphur is an example of a molecular solid.
   a) What is the molecular formula of the sulphur molecule?  
   b) Draw a sketch of the sulphur molecule 
   c) What type of bonding holds the atoms together within the molecule?  
   d) What type of bonding holds the molecules together within the solid?  

Q2. Chlorine and iodine are both molecular elements belonging to the halogen family.
   a) What term is applied to molecules that have only two atoms?  
   b) Why does chlorine have a lower boiling point than iodine?  
   c) Why does chlorine also have a much lower boiling point than sulphur?  

Q3. Molecular forms of carbon have now been discovered. The diagram shows part of a carbon molecule.
   a) What name has been applied to the various molecular forms of carbon?  
   b) What is the molecular formula of the molecule which is partly illustrated in the diagram?  
   c) What makes this type of carbon have a low density?

Q4. The diagram illustrates the bonding structure in diamond.
   a) What type of bonding holds the atoms together in diamond?  
   b) Why is it not possible to identify discrete molecules in the structure of diamond?  
   c) Why is diamond not malleable like metals?  
   d) Why is diamond not able to conduct electricity?  
   e) What property of diamond makes it useful as a glass cutter?  

Total (15)
**HOME PRACTICE 1.2b**

**Q1.** The diagram illustrates the bonding structure in graphite.

- **a)** What element is graphite?  
- **b)** What type of bonding holds the layers together in the structure?  
- **c)** Why is graphite able to conduct electricity?  
- **d)** Why are the layers in graphite able to slide easily?  
- **e)** What two properties of graphite make it useful for the electric contacts to the rotating parts of electric motors?

**Q2.** The diagram shows a periodic arrangement for elements 1 to 20.

From this table, name three:

- **a)** metallic elements  
- **b)** monatomic elements  
- **c)** molecular elements which are not diatomic  
- **d)** diatomic elements  
- **e)** covalent network elements  
- **f)** elements whose physical properties are determined by covalent bonds  
- **g)** elements whose physical properties are determined by London dispersion forces

Total: (12)
1.3 Periodic Trends & Properties

This lesson looks more closely at the size of atoms and the way they vary within the Periodic Table.

Atomic Radius

We are used to thinking of atoms as spheres, and of drawing them as circles, so the size of an atom could be defined by its radius the distance from the centre of the atom to its edge.

Unfortunately, outside the nucleus is mainly empty space and there is no obvious edge.

What can be detected (by, for example, X-rays) are the nuclei of atoms. This allows the distance between the nuclei of neighbouring atoms to be measured, and this can be used to ‘calculate’ the atomic radius of individual atoms.

The radius can be calculated by assuming it is half the distance between the nuclei of two atoms. However, there are two ways that this can be done.

Covalent Radius

When two atoms form a covalent bond they overlap. (Also metallic bonding)

Half the distance between the two nuclei would give a value slightly smaller than the true size of the atom.

In the case of iodine, I₂,

\[ \text{covalent radius} = \frac{266 \text{ pm}}{2} = 133 \text{ pm} \]

(\[ \text{pm} = \text{picometre} = 1 \times 10^{-12} \text{ m} \])

Van der Waals Radius

Between some atoms and many molecules, there exist weak London Dispersion forces which do not involve any overlap.

Half the distance between the two nuclei of atoms in neighbouring molecules would give a value closer to the true size of the atom.

In the case of iodine, I₂,

\[ \text{van der waals radius} = \frac{430 \text{ pm}}{2} = 215 \text{ pm} \]

(\[ \text{pm} = \text{picometre} = 1 \times 10^{-12} \text{ m} \])

Unfortunately, there are very few atoms whose size could be determined using elements, in the solid state, held together by just London Dispersion forces. Therefore, Covalent Radii are the values most often quoted in Data Books and these are the values you will use most often.
As you go do a Group (column) there is an increase in the size of the atoms (covalent radii).

- Li 2,1 covalent radius = pm
- Na 2,8,1 covalent radius = pm
- K 2,8,8,1 covalent radius = pm

This is not surprising, as each atom has an extra energy level, further out from the nucleus.

More surprisingly is that:
as you go across a Period (row) there is a decrease in the size of atoms, (covalent radii).

The number of energy levels is the same, but the increasing size of the nuclear charge causes increased attraction between the outermost electrons and the nucleus. This causes the outer shell to be pulled closer and closer to the nucleus, and the size of the atoms decrease.
First Ionisation Energy

The first ionisation energy is the energy required to remove one electron from every atom in a mole of free atoms.

It is measured in kJ mol\(^{-1}\), i.e. kilojoules per mole

\[
\text{Na}_\text{(g)} \rightarrow \text{Na}^{+} \text{(g)} + \text{e}^{-} \quad \text{kJ mol}^{-1}
\]

The sodium atoms have to be free, i.e. not bonded to any other atom. This can only be done by heating the sodium until it is a gas. The electron lost appears as a product, on the right hand side of the equation.

All atoms, including non-metal atoms, can be forced to lose an electron. This results in the formation of a positive ion, whereas in normal reactions, non-metals will usually gain electrons and form negative ions.

\[
\text{Cl}_\text{(g)} \rightarrow \text{Cl}^{+} \text{(g)} + \text{e}^{-} \quad \text{kJ mol}^{-1}
\]

The chlorine atoms have to be free, i.e. not bonded to any other atom. This can only be done by breaking the chlorine molecules apart. The electron lost appears as a product, on the right hand side of the equation.

Across a Period & Down a Group

<table>
<thead>
<tr>
<th>First Ionisation Energies (kJ mol(^{-1}))</th>
<th>Li</th>
<th>Be</th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
<th>Ne</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>2.5</td>
<td>2.2</td>
<td>2.3</td>
<td>2.4</td>
<td>2.5</td>
<td>2.6</td>
<td>2.7</td>
<td>2.8</td>
</tr>
</tbody>
</table>
As you go across a period the general trend is for the first ionisation energy to increase. (The fluctuations from the general trend require a more advanced explanation that goes beyond this level of chemistry.)

As mentioned earlier, the number of energy levels is the same, but the increasing size of the nuclear charge causes increased attraction between the outermost electron and the nucleus. This makes it harder to remove an electron from the outer shell.

On going down a group, a new energy level (shell) is added each time. Each outer layer is further from the nucleus. This could make it easier to remove an electron.

The outer shell is screened by more and more shells closer to the nucleus. This could make it easier to remove an electron.

On the other hand, the nuclear charge is also increasing. This could make it harder to remove an electron.

As you go down a group the general trend is for the first ionisation energy to decrease.

On going down a group, an extra shell is added each time. Each shell is further from the nucleus. This makes it easier to remove an electron.

The outer shell is screened by more and more shells closer to the nucleus. This makes it easier to remove an electron.

Even though the nuclear charge is increasing, the other two factors mean that it still gets easier to remove an electron as you go down a group.
Further Ionisation Energies

The second ionisation energy is the energy required to remove a second electron from every ‘atom’ in a mole of free ‘atoms’.

\[
\text{(Na}_\text{(g)} \text{ Na}^{+}_\text{(g)} + e^- 502 \text{ kJ mol}^{-1})
\]

\[
\text{Na}^+_\text{(g)} \rightarrow + + \text{kJ mol}^{-1}
\]

Each ionisation will require more energy than the previous one; a dramatic increase occurs whenever an electron has to be removed from the next shell closer to the nucleus.

In the case ofberyllium, Be 2,2, the dramatic increase occurs between the and ionisation energies.

<table>
<thead>
<tr>
<th>Energy (kJ mol(^{-1}))</th>
</tr>
</thead>
<tbody>
<tr>
<td>25000</td>
</tr>
<tr>
<td>20000</td>
</tr>
<tr>
<td>15000</td>
</tr>
<tr>
<td>10000</td>
</tr>
<tr>
<td>5000</td>
</tr>
</tbody>
</table>

**Further Ionisation Energies**

1st 2nd 3rd 4th

### Oxygen

Has electrons in total and an electron arrangement of...

Each ionisation energy is bigger than the previous one but the dramatic increase occurs between the and...

### Silicon

Has electrons in total and an electron arrangement of...

Each ionisation energy is bigger than the previous one but the dramatic increases occur between the and, and then between the and...

### Magnesium

Has electrons in total and an electron arrangement of...

Each ionisation energy is bigger than the previous one but the dramatic increases occur between the and, and then between the and...
**SELF CHECK**

**Q1.** In which of these diagrams does the arrow represent the covalent radius?

A [Diagram A]  
B [Diagram B]  
C [Diagram C]  
D [Diagram D]

**Q2.** What size is a picometre?

A \(1 \times 10^{-12}\) m  
B \(1 \times 10^{-10}\) m  
C \(1 \times 10^{-9}\) m  
D \(1 \times 10^{-6}\) m

**Q3.** What is the covalent radius for an atom of sulphur?

A 26 pm  
B 51 pm  
C 102 pm  
D 204 pm

**Q4.** On going across a period in the periodic table the covalent radius

A increases because the atom has more electrons  
B decreases because the nucleus has a stronger attraction for the outermost electrons  
C increases because the atom has more energy levels for electrons  
D decreases because the mass of the atom is increasing

**Q5.** On going down a group in the periodic table the covalent radius

A increases because the atom has more electrons  
B decreases because the nucleus has a stronger attraction for the outermost electrons  
C increases because the atom has more energy levels for electrons  
D decreases because the mass of the atom is increasing

**Q6.** The alkali metals have particularly low densities because their atoms are

A reactive in water  
B large in size  
C heavy in mass  
D small in volume

**Q7.** First ionisation of an element refers to the

A removal of one electron from each atom  
B addition of one electron to each atom  
C removal of one energy level from each atom  
D completion of one energy level in each atom

**Q8.** Which equation represents the change for the first ionisation of sulphur?

A \(S_{(g)} \rightarrow S^{+}_{(g)} + e^{-}\)  
B \(S_{(g)} \rightarrow S^{2+}_{(g)} + 2 e^{-}\)  
C \(S_{(g)} + e^{-} \rightarrow S^{-}_{(g)}\)  
D \(S_{(g)} + 2 e^{-} \rightarrow S^{2-}_{(g)}\)
Self Check

Q1. The change in ionisation energy across a period is due to

A  the decreasing nuclear attraction for the outer electrons
B  the increasing nuclear attraction for the outer electrons
C  the decreasing screening effect of the inner electrons
D  the increasing screening effect of the inner electrons

Q2. The first ionisation energy of argon is smaller than that of neon because argon has

A  fewer electrons screening its nucleus
B  more electrons screening its nucleus
C  a smaller nuclear charge
D  a greater nuclear charge

Q3. Second ionisation refers to the removal from each atom of

A  a total of two outer electrons
B  a second electron after one has already been removed
C  one electron from the second energy level
D  a second energy level after the removal of the first

Q4. Which of the following statements is true?

A  the potassium ion is larger than the potassium atom.
B  the chloride ion is smaller than the chlorine atom
C  the sodium atom is larger than the sodium ion
D  the oxygen atom is larger than the oxide ion

Q5. Which equation represents the change for the second ionisation of magnesium?

A  \( \text{Mg}^+ (s) \rightarrow \text{Mg}^{2+} (s) + e^- \)
B  \( \text{Mg} (s) \rightarrow \text{Mg}^{2+} (s) + 2 e^- \)
C  \( \text{Mg}^+ (g) \rightarrow \text{Mg}^{2+} (g) + e^- \)
D  \( \text{Mg} (g) \rightarrow \text{Mg}^{2+} (g) + 2 e^- \)

Q6. Which equation represents the change for the second ionisation of bromine?

A  \( \text{Br} (g) + e^- \rightarrow \text{Br}^- (g) \)
B  \( \text{Br} (g) + 2 e^- \rightarrow \text{Br}^{2-} (g) \)
C  \( \text{Br}^+ (g) \rightarrow \text{Br}^{2+} (g) + e^- \)
D  \( \text{Br} (g) \rightarrow \text{Br}^{2+} (g) + 2 e^- \)

Q7. What is the first ionisation energy of sulphur with the correct units?

A  2260 mol kJ\(^{-1}\)
B  1010 mol kJ\(^{-1}\)
C  2260 kJ mol\(^{-1}\)
D  1010 kJ mol\(^{-1}\)

Q8. The spike graph shows the variation in successive ionisation energies of an element, Z.

Which group of the Periodic Table is element Z?

A  1  B  3
C  4  D  6
**HOME PRACTICE**  1.3A

**Q1.** The diagram shows a Dalton-type picture of a bromine molecule.

* a) Why is it not easy to define the size of an atom in a molecule? 1

* b) What is the covalent radius of an atom defined as? 1

* c) Illustrate this by drawing a large diagram of a diatomic molecule and clearly marking the covalent radius on it. 1

* d) What is meant by the van der Waals radius of an atom? 1

**Q2.** The covalent radius of sulphur is **104 pm** and the van der Waals radius is **190 pm**

* a) What will be the distance between the two nuclei of neighbouring atoms in the ring. 1

* b) What will be the closest distance between the centres of sulphur atoms in neighbouring molecules. 1

**Q3.** The covalent radii going across a period in the periodic table exhibit a regular trend.

* a) Describe this trend. 1

* b) Explain why this occurs. 1

**Q4.** This graph plots the covalent radii for some of the group 5 elements.

* a) From the graph what value would you predict for the covalent radius of antimony (Sb)? 1

* b) Explain the trend seen in the graph. 1

**Total** (10)
**HOME PRACTICE 1.3b**

**Q1.** Ionisation energies are used by scientists to help understand about the energies of electrons in atoms.

a) What is meant by the first ionisation energy of nitrogen?  

b) Write the symbol equation for this change.  

c) Write the symbol equation for the second ionisation energy of nitrogen.

**Q2.** The diagram shows the target picture for an atom of calcium.

a) What electron is involved in the first ionisation of calcium?  

b) What is meant by electron screening effect?

**Q3.** This simplified diagram shows the relative positions of four elements in the periodic table.

a) From their positions, will the first ionisation energy of neon be higher or lower than that of lithium?  

b) Explain why.  

c) From their positions, will the first ionisation energy of caesium be higher or lower than that of lithium?  

d) Explain why.  

e) Which of the four elements will have the highest ionisation energy?

*Total* (10)
This lesson is about molecular compounds and some of the bonding forces that are present in their structures.

**1.4 Covalent Molecular Compounds**

This activity considers the bonding present in molecular compounds and how this affects properties.

**Carbon tetrafluoride** is a covalent molecular compound. By sharing electrons, both the carbon and the fluorine atoms achieve the extra stability of a full outer shell.

The covalent bonds, however, go no further than the 5 atoms making up each discrete molecule.

**Bonding and Properties**

The intramolecular forces (intra- = within), that join the fluorine atoms to the carbon atom, are strong covalent bonds.

Once cooled enough to solidify, weak London Dispersion forces will exist between molecules (intermolecular forces).

It is only these weak forces that are broken when molecular compounds like CF₄ melt, hence their very low melting points.

Even when solid, substances like CF₄ are not strong - again, this is because only the weak London Dispersion forces need to be broken.

<table>
<thead>
<tr>
<th>“Dot-and-cross” diagram</th>
<th>Structural formula</th>
<th>Molecular formula and name</th>
</tr>
</thead>
</table>
| ![Diagram](image1)       | F \[\begin{array}{c} | \\ F \end{array} \]
                          | F – C – F           | carbon tetrafluoride      |
| ![Diagram](image2)       | ![Diagram](image3)  | phosphorus trichloride    |

We London Dispersion forces

We London Dispersion forces

Str covalent bonds
Effect of Increasing Atomic Size

This activity looks at the effect that increasing atomic size has on the melting points of the carbon tetrahalides.

As the size of the halogen atom in these compounds, the melt point increases. This is because the electron in larger atoms are more likely to ‘wobble’ leading to stronger London Dispersion forces between molecules.
**Effect of Increasing the Number of Atoms**

This activity looks at the effect that increasing the number of atoms has on the boiling points of the alkanes.

<table>
<thead>
<tr>
<th>Boiling point (°C)</th>
<th>Number of carbon atoms in the main alkane chain</th>
</tr>
</thead>
<tbody>
<tr>
<td>150</td>
<td></td>
</tr>
<tr>
<td>125</td>
<td></td>
</tr>
<tr>
<td>100</td>
<td></td>
</tr>
<tr>
<td>75</td>
<td></td>
</tr>
<tr>
<td>50</td>
<td></td>
</tr>
<tr>
<td>25</td>
<td></td>
</tr>
<tr>
<td>0</td>
<td></td>
</tr>
<tr>
<td>-25</td>
<td></td>
</tr>
<tr>
<td>-50</td>
<td></td>
</tr>
</tbody>
</table>

As the **num** of **car** atoms in the molecules **increases**, the **melt point** **increases**. This is because **lar num** of atoms produce **more temp dip** and, therefore, **mo London Dispersion forces bet molecules. Mo London Dispersion forces means stro London Dispersion forces so mo ene** is needed to **move** these molecules **fur apart**, hence the **hi mel point**.

![Diagram of propane (a gas) and octane (a liquid) with London Dispersion forces shown.](image)
This activity explores the effect that unequal sharing of electrons has on the atoms with a covalent bond and on the intermolecular bonding between such molecules.

In a mol$_2$ like F$_2$, both atoms are exactly the same. They have equal attraction for the bonding pair of elec$_2$. The elec$_2$ are equ sha. This is a pure cov bond.

A fluo atom has a stron attraction for elec than a hydr atom. The bonding pair is pulled clo to the fluo. The fluorine becomes sligh neg (δ−), while the hydrogen becomes sligh pos (δ+). A perm dipole is set up. This is a polar cov bond.

The significant difference is that these dip are perm meaning that the attr are perm. This results in a stro van der Waals force than the Lon Disp forces met so far which rely on temp dip inducing dip on neighbouring particles.

**How Strong An Attraction ?**

This activity is about how different atoms have different strengths of attraction for the shared electrons in a covalent bond.

**Electronegativity** is a measure of the attraction an atom has for the shared electrons in a covalent bond.

**Periodic Pattern: Electronegativity**

<table>
<thead>
<tr>
<th></th>
<th>H</th>
<th>C</th>
<th>O</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li</td>
<td>2.2</td>
<td>2.5</td>
<td>4.0</td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Rb</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cs</td>
<td>0.7</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Transition Metals
As you go across a period the trend is for the electronegativity to increase.

The atoms are ‘shrinking’ in size bringing any shared electrons closer and closer to the nucleus.

As you go down a Group there is a decrease in the electronegativity.

This is not surprising, as each atom has an extra energy level, (shell) further out from the nucleus.

The furthest from the nucleus the shared pair of electrons are, the weaker the attraction.

Electronegativity values can be used to determine the polarity of a covalent bond. The atom with the highest electronegativity value will have a stronger attraction for the shared pair of electrons. This atom will end up slightly negative, while the other atom is slightly positive.

The relative polarity of two bonds can also be compared using electronegativity values. Subtracting one value from the other provides a rough indication. The value for the H—Br bond is 0.6 (2.8 - 2.2). The value for the H—I bond is ( - ), showing that the H—Br bond is more polar.

In general, particularly in simple molecules, the more polar the molecule is, the stronger the intermolecular attractions are likely to be.

H₃C—CH₃  ethane  non-polar
H₃C—NH₂  methylamine
H₃C—OH  methanol
H₃C—Cl  chloromethane  most polar

Bpt (°C)  -6  -24
Electrons  18
Mass  30

In reality, other factors such as number of electrons, mass of molecule, shape of molecule and, as you will see later, the presence of hydrogen atoms on a very electronegative atom (O—H and N—H in particular) can also have a big effect on the intermolecular attractions.
This activity is about polar molecules and the property which can be used to detect when they are present in a liquid.

A polar molecule is a molecule which has a permanent dipole (a permanent slight positive charge on one side and a permanent slight negative charge on the other).

A water molecule is a good example of a polar molecule. It is polar, first of all, because the greater electronegativity of the oxygen atom (3.5) compared to the hydrogen atom (2.2), makes the O—H bonds polar.

Secondly, and just as important; the shape of the water molecule means that one side has a slight negative charge while the other side is slightly positive. This makes water a polar molecule.

Liquids which contain polar molecules can be detected using a charged rod. A fine jet of liquid from a burette will be attracted to the rod if the liquid contains polar molecules.

The O—H bond is polar, and the shapes of water and ethanol molecules are suitable, so these are polar and were attracted towards the charged rod.

In alkanes, the bonds and the molecule shapes are not able to produce polar molecules, so they are not attracted towards the charged rod.
**Self Check**

**Q1.** In a covalent molecular substance
- A all molecules are joined to each other by covalent bonds
- B covalent bonds only exist within distinct molecules
- C London dispersion forces link atoms within each molecule
- D all the atoms in the structure form a giant covalent lattice

**Q2.** A solid covalent molecular compound has
- A covalent bonds and intermolecular forces
- B a network of covalent bonds
- C covalent bonds and ionic forces
- D only intermolecular forces throughout

**Q3.** Here is the outer electron picture for a molecule of hydrogen sulphide.
- The diagram shows that
  - A only sulphur has achieved a stable electron arrangement
  - B this is a diatomic molecule
  - C all the atoms have achieved stable electron arrangements
  - D this is an unstable molecule

**Q4.** The boiling point of C₆H₁₄ is greater than that of C₄H₁₀ because C₆H₁₄ has
- A more covalent bonds within the molecules
- B more London dispersion forces within the molecules
- C more covalent bonds between the molecules
- D more London dispersion forces between the molecules

**Q5.** CBr₄ has a higher boiling point than CCl₄ because
- A the larger bromine atoms result in weaker covalent bonds
- B bromine atoms are more stable than chlorine atoms
- C the larger bromine atoms result in stronger London dispersion forces
- D the C—Br bond is more reactive than the C—Cl bond

**Q6.** The diagrams show the structural formulae of two molecules.

\[
\text{X} \quad \text{Cl} - \text{Si} - \text{Si} - \text{Cl} \quad \text{Y} \quad \text{Cl} - \text{Si} - \text{Si} - \text{Si} - \text{Cl}
\]

Compared with molecule Y, molecule X will have
- A more intermolecular forces
- B a lower melting point
- C stronger covalent bonds
- D a higher boiling point

**Q7.** Hydrocarbon solids such as candle wax are soft because they have
- A a covalent lattice
- B many weak covalent bonds
- C only non-metallic atoms
- D a molecular structure

**Q8.** Carbon dioxide is a gas at room temperature because it has
- A strong covalent bonds
- B weak covalent bonds
- C strong intermolecular forces
- D weak intermolecular forces
## Self Check

**Q1.** Electrons in covalent bonds are not always shared equally because

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>some electrons are more stable than others</td>
</tr>
<tr>
<td>B</td>
<td>some bonds are stronger than others</td>
</tr>
<tr>
<td>C</td>
<td>different atoms have different stabilities</td>
</tr>
<tr>
<td>D</td>
<td>different atoms have different attractions for electrons</td>
</tr>
</tbody>
</table>

**Q2.** In a polar covalent bond a permanent dipole is created because the electrons are

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>slightly charged</td>
</tr>
<tr>
<td>B</td>
<td>unequally shared</td>
</tr>
<tr>
<td>C</td>
<td>completely uncharged</td>
</tr>
<tr>
<td>D</td>
<td>equally shared</td>
</tr>
</tbody>
</table>

**Q3.** If X has a stronger attraction for bonding electrons than Y, which diagram correctly depicts the X - Y bond?

<table>
<thead>
<tr>
<th>Diagram</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td><img src="image1" alt="Diagram A" /></td>
</tr>
<tr>
<td>B</td>
<td><img src="image2" alt="Diagram B" /></td>
</tr>
<tr>
<td>C</td>
<td><img src="image3" alt="Diagram C" /></td>
</tr>
<tr>
<td>D</td>
<td><img src="image4" alt="Diagram D" /></td>
</tr>
</tbody>
</table>

**Q4.** Electronegativity is a measure of

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>the energy required to add an electron to an atom.</td>
</tr>
<tr>
<td>B</td>
<td>the total number of electrons in an atom</td>
</tr>
<tr>
<td>C</td>
<td>the number of outer electrons in an atom</td>
</tr>
<tr>
<td>D</td>
<td>the attraction an atom has for the shared electrons in a covalent bond</td>
</tr>
</tbody>
</table>

**Q5.** In the periodic table the electronegativities of the elements

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>decrease down a group and decrease from left to right across a period</td>
</tr>
<tr>
<td>B</td>
<td>increase down a group and increase from left to right across a period</td>
</tr>
<tr>
<td>C</td>
<td>decrease down a group and increase from left to right across a period</td>
</tr>
<tr>
<td>D</td>
<td>increase down a group and decrease from left to right across a period</td>
</tr>
</tbody>
</table>

**Q6.** Use electronegativity values to calculate which of these bonds is the most polar.

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>H—Br</td>
</tr>
<tr>
<td>B</td>
<td>N—Cl</td>
</tr>
<tr>
<td>C</td>
<td>H—Br</td>
</tr>
<tr>
<td>D</td>
<td>Br—Cl</td>
</tr>
</tbody>
</table>

**Q7.** A polar molecule has

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>positive and negative ions</td>
</tr>
<tr>
<td>B</td>
<td>more than its fair share of electrons</td>
</tr>
<tr>
<td>C</td>
<td>only positive ions present</td>
</tr>
<tr>
<td>D</td>
<td>a slightly positive side and a slightly negative side</td>
</tr>
</tbody>
</table>

**Q8.** Polar-polar attractions can also be described in terms of dipoles as

<table>
<thead>
<tr>
<th>Option</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>temporary — induced</td>
</tr>
<tr>
<td>B</td>
<td>temporary — permanent</td>
</tr>
<tr>
<td>C</td>
<td>permanent — permanent</td>
</tr>
<tr>
<td>D</td>
<td>permanent — induced</td>
</tr>
</tbody>
</table>
**Home Practice 1.4A**

**Q1.** Silicon chloride is a liquid at room temperature.

\[ a) \] What is the formula of this compound?  
\[ b) \] Why can it be described as covalent molecular?  
\[ c) \] Draw a "dot-and-cross" outer electron picture for silicon tetrachloride.  
\[ d) \] What does each atom achieve, as shown in this diagram?  
\[ e) \] What type of bonding holds the molecules together in the liquid?  

**Q2.** Silicon tetrafluoride molecules have exactly the same shape as silicon tetrachloride molecules.

By considering atomic size and the intermolecular forces involved explain why silicon tetrafluoride is a gas when silicon tetrachloride is a liquid.

**Q3.** Butane and pentane are both hydrocarbon molecules.

\[ a) \] What type of bonding force exists between hydrocarbon molecules?  
\[ b) \] Which of these hydrocarbons will have the higher boiling point and why?  
\[ c) \] Draw a diagram to illustrate the difference in intermolecular forces for these two molecules.

\[ \text{Total (10)} \]
**HOME PRACTICE 1.4B**

**Q1.** Copy each of the following and indicate the direction of polarity, if any, in each bond using an arrow on the bond and δ+ and δ- on the appropriate atoms.

   a) \( \text{H} \rightarrow \text{Br} \)  
   b) \( \text{H} \rightarrow \text{C} \)  
   c) \( \text{C} \rightarrow \text{C} \)

   

**Q2.** Use the following data to decide whether a stream of each of the following liquids would deflect if a charged rod was placed near it.

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Polarity</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) pentane</td>
<td>non-polar molecules</td>
</tr>
<tr>
<td>b) methanol</td>
<td>polar molecules</td>
</tr>
<tr>
<td>c) carbon tetrachloride</td>
<td>non-polar molecules</td>
</tr>
</tbody>
</table>

**Q3.** Use electronegativity values to decide which of each of the following pairs of bonds is more polar. Your working should be shown clearly.

   a) \( \text{C} \rightarrow \text{Br} \) or \( \text{H} \rightarrow \text{F} \)  
   b) \( \text{H} \rightarrow \text{I} \) or \( \text{N} \rightarrow \text{S} \)

   

\[ \text{Total} \ (10) \]
1.5 Polar Molecules

This lesson looks in more detail at how the shape of a molecule can affect its overall polarity, and how this, in turn, can affect the forces acting between molecules.

Shape & Polarity

This activity considers how the shape of a molecule can affect its overall polarity.

- The O—H bond is a **polar covalent bond**, with the oxygen slightly negative and the hydrogens slightly positive. Its shape results in a negative end and a positive end so it is a polar molecule.

- The C—H bond is a **polar covalent bond**, with the carbon slightly negative and the hydrogens slightly positive. Its shape doesn't produce a negative end and a positive end, it is not a polar molecule.

- The C = O bond is a **polar covalent bond**, with the oxygen slightly negative and the carbon slightly positive. Its shape doesn't produce a negative end and a positive end, it is not a polar molecule.

- The C—Cl bond is even more polar with the chlorine slightly negative and the carbon slightly positive. Overall this does produce a negative end and a positive end so it is a polar molecule.

Solubilities

This activity compares the intermolecular forces between polar and non-polar molecular compounds in terms of their ability to dissolve in different solvents.

**Polar→polar attractions** are intermolecular forces of attractions that exist between molecules which have a *permanent* overall dipole.

Polar→polar attractions are *perm* and, therefore, *stro* than the *temp* dipoles set up in non-polar molecules, London Dispersion forces.
### Chemical Changes & Structure

#### Topic 1 - Bonding & Periodicity

**CFE New Higher**

<table>
<thead>
<tr>
<th>Molecular Compound</th>
<th>Polarity</th>
<th>Solubility in heptane (non-polar)</th>
<th>Solubility in propanone (polar)</th>
</tr>
</thead>
<tbody>
<tr>
<td>hexane C\textsubscript{6}H\textsubscript{14}</td>
<td>non-polar</td>
<td></td>
<td></td>
</tr>
<tr>
<td>toluene CH\textsubscript{3}C\textsubscript{6}H\textsubscript{5}</td>
<td>nearly non-polar</td>
<td></td>
<td></td>
</tr>
<tr>
<td>trichloroethane CH\textsubscript{3}CCl\textsubscript{3}</td>
<td>polar</td>
<td></td>
<td></td>
</tr>
<tr>
<td>ethanol CH\textsubscript{3}CHO</td>
<td>polar</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

When the inter attr are fairly sim, then mol can mix easily be miscible - ‘like dissolves like’. If the molecules of one substance have much stro attr then they will tend to form sep layers, be immiscible.

### Comparing Boiling Points

This activity compares the boiling points of polar and non-polar molecular compounds.

It is only fair to compare molecules of

- similar size
- similar shape
- similar numbers of electrons

**iodine chloride**

- b.p. °C

Both molecules are of sim si, iden sha and have sim num of elec. Both molecules have sim London Dispersion forces.

**bromine**

- b.p. °C

Iodine chloride has a hi boiling point because of the ex perm po —po attractions that exist.

Both molecules are of sim si, sim sha and have methanal sim num of elec. Both molecules have b.p. °C

**methanal**

- b.p. °C

Methanal has a hi boiling point because of the ex perm po —po attractions that exist.

**ethane**

- b.p. °C

Both molecules are of sim si, sim sha and have sim num of elec. Both molecules have sim London Dispersion forces.

**ether**

- b.p. °C

Ether has a hi boiling point because of the ex perm po —po attractions that exist.

The stro forces of attr that exist between molecules, the hi the boiling point. The attr that exist between molecules with perm dipoles are stro than London Dispersion forces (temp →ind dipoles).
Self Check

Q1. Which diagram correctly shows the directions of the bond polarities in CFCl₃?

A. \[
\begin{array}{c}
\text{F} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]
B. \[
\begin{array}{c}
\text{F} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]
C. \[
\begin{array}{c}
\text{F} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]
D. \[
\begin{array}{c}
\text{F} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]

Q2. Which one of the following molecules is polar?

A. \[
\begin{array}{c}
\text{O} \\
\text{C} \\
\text{O}
\end{array}
\]
B. \[
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{Cl} \\
\text{Cl}
\end{array}
\]
C. \[
\begin{array}{c}
\text{Br} \\
\text{C} \\
\text{Br}
\end{array}
\]
D. \[
\begin{array}{c}
\text{Br} \\
\text{Br}
\end{array}
\]

Q3. Polar-polar attractions are the forces between
A. molecules which are polar
B. molecules which are non-polar
C. atoms of the same kind in a polar molecule
D. atoms at opposite ends of a polar covalent bond

Q4. The strength of polar-polar attractions is slightly
A. less than London dispersion forces
B. greater than London dispersion forces
C. greater than pure covalent bonding
D. less than pure covalent bonding

Questions 5 and 6 refer to these solvents.

<table>
<thead>
<tr>
<th></th>
<th></th>
<th>heptane</th>
<th>trichloroethane</th>
<th>ethanol</th>
</tr>
</thead>
<tbody>
<tr>
<td>X</td>
<td>Y</td>
<td>CH₃C₃H₁₄</td>
<td>CH₃CCl₃</td>
<td>CH₃CHO</td>
</tr>
<tr>
<td>C₇H₁₄</td>
<td>CH₃</td>
<td>non-polar</td>
<td>polar</td>
<td>polar</td>
</tr>
</tbody>
</table>

Q5. Chloroform, CHCl₃, is a polar liquid. In which of these solvents will it be soluble?
A. only X
B. only Y
C. only Y and Z
D. all three

Q6. Cyclohexane, C₆H₁₂, is a non-polar liquid. In which of these solvents will it be soluble?
A. only X
B. only Y
C. only Y and Z
D. all three

Q7. Substances P and Q have different boiling points.

The difference in boiling point is most likely due to a difference in
A. their London dispersion forces
B. the size of the molecules
C. their polar-polar attractions
D. the number of covalent bonds

Q8. Why can the effect of polarity on boiling points not be compared for molecules of different size?
A. size effects polar-polar attractions
B. Larger molecules lessen the intermolecular forces.
C. polar covalent bonds cannot form between large molecules.
D. London dispersion forces would also differ
HOME PRACTICE

Q1. Carbon dioxide has molecules which are linear in shape.
   
   a) Draw the carbon dioxide molecule showing the direction of the dipoles for the polar bonds
   b) Why is carbon dioxide a non-polar molecule?
   c) What bonding forces exist between its molecules?
   d) How does this explain why carbon dioxide is a gas at room temperature?

Q2. Ethanal has molecules of similar size to those of carbon dioxide but the shape is angular.
   
   a) Draw the ethanal molecule showing the direction of the dipole on the polar bond.
   b) Why is methanal described as a polar molecule?
   c) What bonding forces exist between its molecules?
   d) How does this explain why ethanal is a liquid at room temperature?

Q3. The bonding around carbon atoms is usually tetrahedral.
   
   a) Explain why CF₄ is non-polar
   b) Explain why HCF₃ is polar

Total (10)
1.6 Hydrogen Bonding & Water

This lesson looks at hydrogen bonding, how it arises, and its effect on properties such as melting point, boiling point, and viscosity.

This activity is about the extra bonding strength of the hydrogen bond and the situations in which it can arise.

Molecules such as \textit{ammonia} (NH\textsubscript{3}), \textit{water} (H\textsubscript{2}O) and \textit{hydrogen fluoride} (HF) have stronger attractions than London Dispersion forces or polar→polar attractions. These stronger attractions are called hydrogen bonds.

\textit{Hydrogen bonds} are only set up where a strongly \textit{polar} covalent bond already exists between a \textit{hydrogen} atom and a small, strongly \textit{electron attracting} atom. There are only 3 really suitable bonds:

\begin{align*}
\text{N}^{\delta-} &- \text{H}^{\delta+} \\
\text{O}^{\delta-} &- \text{H}^{\delta+} \\
\text{F}^{\delta-} &- \text{H}^{\delta+}
\end{align*}

These bonds must also \textit{line up} so that the \textit{hydrogen} is in a \textit{straight line} between two \textit{slightly negative} atoms: this leads to much stronger than normal attractions called \textit{hydrogen bonds}.

<table>
<thead>
<tr>
<th>Weak Bonding Forces</th>
<th>Intermediate strength</th>
<th>Very Strong Forces</th>
</tr>
</thead>
<tbody>
<tr>
<td>London Dispersion forces</td>
<td>Hydrogen bonds</td>
<td>Covalent bonds</td>
</tr>
<tr>
<td>polar→polar attractions</td>
<td></td>
<td>Ionic bonds</td>
</tr>
</tbody>
</table>

Boiling Points

This activity examines the boiling points of the hydrides of group 7 elements in relation to the intermolecular forces present.

\begin{center}
\begin{tabular}{c|c|c|c|c}
 & row 2 & row 3 & row 4 & row 5 \\
\hline
Boiling point (°C) & & & & \\
0 & & & & \\
-20 & & & & \\
-40 & & & & \\
-60 & & & & \\
-80 & & & & \\
-100 & & & & \\
\end{tabular}
\end{center}
As molecular size decreases, we expect the boiling point to decrease, especially as London Dispersion forces decrease.

We would expect, therefore that the boiling point for hydrogen fluoride would be about $-108^\circ$C. It is in fact, $19^\circ$C, about $120^\circ$ higher than expected. The extra energy is needed to overcome the much stronger hydrogen bonds that exist between molecules of hydrogen fluoride.

Viscosity & Bonding

This activity examines the effect of hydrogen bonding on the viscosity of a liquid by timing how long two liquids take to drain from a burette.

Two molecules of similar size, similar shape and number of electrons are used to ensure that London Dispersion forces would be very similar.

<table>
<thead>
<tr>
<th>Molecule X</th>
<th>Molecule Y</th>
</tr>
</thead>
<tbody>
<tr>
<td>(propanone)</td>
<td>(propan-2-ol)</td>
</tr>
<tr>
<td>polar hydrogen bonding is present</td>
<td>polar hydrogen bonding is present</td>
</tr>
<tr>
<td>not present</td>
<td>seconds to drain from 0 to 50</td>
</tr>
<tr>
<td></td>
<td>seconds to drain from 0 to 50</td>
</tr>
</tbody>
</table>

Propan-2-ol is more viscous than propanone because the H—O bond is very polar and the slightly positive hydrogen atom (H$^{\delta+}$) can end up in a straight line between two slightly negative oxygen atoms (O$^{\delta-}$).

The extra attractions (hydrogen bonds) set up between the molecules of propan-2-ol make the molecules tend to stick together and make it a more viscous liquid.

Boiling Point of Water

This activity considers the boiling point of water in relation to the boiling points of the other group 6 elements.
As molecular size decreases, we expect the boiling point to decrease, especially London dispersion forces.

We would expect, therefore that the boiling point for water would be about \(-70^\circ\text{C}\). It is in fact, \(100^\circ\text{C}\), about \(170^\circ\text{C}\) higher than expected.

The extra energy is needed to overcome the much stronger hydrogen bonds that exist between molecules of water.

**Structure of Ice**

This activity looks at the structure and bonding in ice and how this explains its unusual density.

Normally **mol** solids are denser than their own **liquids**. This is because the increasing **attractions** (London dispersion) pull the molecules closer together as the solid forms.

For example, a piece of **can** wax will sink when dropped into liquid wax.

Ice floats on water because it is **less dense**. The strongest attractions are the hydrogen bonds.

Hydrogen bonds are **stronger**, making ice a **strong structure** (think of the Titanic!).

Hydrogen bonding also makes the water molecules line up so that the hydrogen atoms (H\(^{\delta^+}\)) are in a **straight line** between two oxygen atoms (O\(^{\delta^-}\)).

The water molecules end up further apart in **solid ice** than in **liquid water** - so ice is **less dense** than water.
### Dissolving in water

This activity looks at the type of substances which dissolve readily in water and explains it in terms of intermolecular forces.

Water is made up of small **molecules** with hydrogen **bonds** the main **molecular** forces, even in the liquid state.

For another substance to **dissolve** in water it has to able to produce **in molecular forces** of **similar strength**.

<table>
<thead>
<tr>
<th>Molecular Compound</th>
<th>Polarity</th>
<th>Hydrogen bonding present</th>
<th>Ability to form hydrogen bonds with water</th>
<th>Solubility in water</th>
</tr>
</thead>
<tbody>
<tr>
<td>hexane C₆H₁₄</td>
<td>non-polar</td>
<td>✗</td>
<td>✗</td>
<td>✗</td>
</tr>
<tr>
<td>toluene CH₃C₆H₅</td>
<td>nearly non-polar</td>
<td>✗</td>
<td>✗</td>
<td>✗</td>
</tr>
<tr>
<td>trichloroethane CH₃CCl₃</td>
<td>polar</td>
<td>✗</td>
<td>✗</td>
<td>✗</td>
</tr>
<tr>
<td>ethanal CH₃CHO</td>
<td>polar</td>
<td>✓</td>
<td>✓</td>
<td>✓</td>
</tr>
<tr>
<td>propan-2-ol CH₃CHOHCH₃</td>
<td>polar</td>
<td>✓</td>
<td>✓</td>
<td>✓</td>
</tr>
</tbody>
</table>

Hydrogen bonding in water must be **significantly strong** as many **polar** molecules, as well as all **non-polar** molecules, are unable to **dissolve** **(mix)** in water.

- **non-polar** — **non-polar**
- **polar** — **polar**
- **polar** — **polar**

London dispersion

- **temporary** — **induced**
- **permanent** — **permanent**

Hydrogen bonding

- **permanent** — **permanent**

Miscible liquids, those that mix and **dissolve** in water, must have **hydrogen bonds** already or the **ability to form hydrogen bonds** with water molecules.
Most of the time, however, you will need to identify the ‘best solvent’, not by thinking too deeply about the various functional groups/polarity/intermolecular forces, but by simply looking for a solvent that is from the same area as the target molecule - look for a hydrocarbon to dissolve another hydrocarbon, an alcohol to dissolve an alcohol and so on.

Erythrose can be used in the production of a chewing gum that helps prevent tooth decay.

\[
\text{OH} \quad \text{OH} \\
\text{HO} \quad \text{CH}_2 \quad \text{CH} \quad \text{CH} \quad \text{C} \\
\quad \text{H}
\]

Caryophyllene is a natural product which can be extracted from clove oil using a solvent.

\[
\begin{align*}
\text{H}_3\text{C} & \quad \text{CH} \\
\text{H}_3\text{C} & \quad \text{CH} \\
\text{H}_2\text{C} & \quad \text{CH} \\
\text{CH}_2 & \quad \text{CH}_2 \\
\text{C} & \quad \text{CH}_3 \\
\text{C} & \quad \text{CH}_3 \\
\text{HC} & \quad \text{CH}_2 \\
\text{CH} & \quad \text{CH}_2 \\
\end{align*}
\]
**Self Check**

Q1. Hydrogen bonds occur in compounds where an atom bonded to hydrogen has

A a strong attraction for electrons
B a weak attraction for electrons.
C a strong attraction for protons
D a weak attraction for protons

Q2. Which of the following diagrams represents two molecules held together by hydrogen bonding?

A

\[
\begin{array}{c}
\text{H-F} \\
\text{\textvdash} \\
\text{H-F}
\end{array}
\]

B

\[
\begin{array}{c}
\text{H-F} \\
\text{\textdash} \\
\text{H-F}
\end{array}
\]

C

\[
\begin{array}{c}
\text{F-H} \\
\text{\textdash} \\
\text{F-H}
\end{array}
\]

D

\[
\begin{array}{c}
\text{F-H} \\
\text{\textvdash} \\
\text{F-H}
\end{array}
\]

Q3. Which of the following would exhibit hydrogen bonding between its molecules?

A hydrogen phosphide, PH\(_3\)
B hydrogen sulphide, H\(_2\)S
C hydrogen nitride, NH\(_3\)
D hydrogen iodide, HI

Q4. Hydrogen bonding is stronger than

A London dispersion forces and polar-polar attractions
B pure covalent and polar covalent bonds
C polar-polar attractions and ionic bonding
D London dispersion forces and covalent bonds

Q5. Atoms of which three elements can be involved with hydrogen in producing hydrogen bonding?

A fluorine, chlorine, iodine
B sodium, nitrogen, oxygen
C carbon, silicon, sulphur
D nitrogen, oxygen, fluorine

Q6. Hydrogen bonding increases the boiling point of a substance because

A the bonding within the molecules is weaker
B the bonding within the molecules is stronger
C the attractions between the molecules are greater
D the attractions between the molecules are less

Q7. You could test the viscosity of a liquid by running the liquid out of a burette and

A neutralising it with acid
B timing how long it took to deliver a fixed volume
C testing its deflection toward a charged rod
D finding its mass

Q8. Hydrogen bonding increases the viscosity of a substance because

A the bonding within the molecules is weaker
B the bonding within the molecules is stronger
C the attractions between the molecules are greater
D the attractions between the molecules are less
**Self Check**

**Q1.** The unusual boiling point of water is due to

A polar covalent bonds between molecules  
B greater London dispersion forces  
C hydrogen bonding between molecules  
D strong polar-polar attractions

**Q2.** Which diagram shows the way water molecules are attracted to each other in ice?

A  
B  
C  
D

**Q3.** The unusual density of ice is caused by

A covalent bonds holding the molecules in a closely packed structure  
B covalent bonds holding the molecules in an open network structure  
C hydrogen bonds holding the molecules in a closely packed structure  
D hydrogen bonds holding the molecules in an open network structure

**Q4.** Ice is unusually strong for a molecular compound due to

A the strength of ionic attractions  
B a network of hydrogen bonding  
C the strength of polar covalent bonds  
D a network of London dispersion forces

**Q5.** Ice floating in water is unusual because

A it melts away very easily  
B solids usually sink in their own liquids  
C liquids and solids do not usually exist together  
D its molecules change shape on melting

**Q6.** To be miscible in water a covalent molecular compound must

A be strongly polar  
B have hydrogen bonding  
C be non-polar  
D be able to form hydrogen bonds

**Q7.** The boiling point of H₂S is lower than that of H₂O because

A the London dispersion forces are less  
B it has weaker covalent bonds  
C it has smaller molecules than H₂O  
D hydrogen bonding is not present

**Q8.** Coniceine is a deadly poison extracted from the plant hemlock.

Which of the following would be the best solvent for coniceine?

A propanoic acid  
B propan-1-ol  
C heptane  
D water
**HOME PRACTICE 1.6A**

**Q1.** Apart from hydrogen, an atom of one of three elements must be present for hydrogen bonding to be possible

\[ a) \text{ Name these } \textbf{three} \text{ elements.} \]

\[ b) \text{ What } \textbf{two} \text{ things are special about atoms of these elements?} \]

**Q2.** Hydrogen fluoride is a compound in which hydrogen bonding occurs.

\[ a) \text{ Draw a diagram to illustrate hydrogen bonding between two HF molecules} \]

\[ b) \text{ How does the hydrogen bond compare in strength with the polar covalent bond in the HF molecule?} \]

\[ c) \text{ How does the hydrogen bond compare in strength with London dispersion forces?} \]

\[ d) \text{ Hydrogen chloride (b.p. -85 °C) has larger molecules than hydrogen fluoride (b.p. 19 °C) and yet the boiling point is much lower. Explain why this should be.} \]

**Q3.** Two polar liquids are compared for the time it takes 50 ml to drain from a burette.

\[ \text{Liquid A} \]

\[ \text{Liquid B} \]

\[ a) \text{ Which liquid will drain faster and why?} \]

\[ b) \text{ Why can the differences in viscosity not be attributed to London dispersion forces in these examples?} \]

*Total (10)*
Q1. At -10 °C water is solid whereas hydrogen sulphide, H₂S, is a gas.

a) Which of these two compounds shows an unusual melting point in relation to the size of its molecules? 1
b) What are the two main types of bond in ice? 2
c) What kind of shape do the rings of water molecules in ice produce? 1
d) What gives ice its strength? 1
e) What makes ice less dense than liquid water? 1

Q2. The following table provides information about some molecular compounds.

<table>
<thead>
<tr>
<th>Name</th>
<th>Structure</th>
<th>Polarity</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) methanal</td>
<td>C = O</td>
<td>polar</td>
</tr>
<tr>
<td>b) chloroform</td>
<td>Cl</td>
<td>polar</td>
</tr>
<tr>
<td>c) methane</td>
<td>C - H</td>
<td>non-polar</td>
</tr>
<tr>
<td>d) ethanol</td>
<td>C - O - H</td>
<td>polar</td>
</tr>
</tbody>
</table>

4

Total (10)
1.7 Ionic & Covalent Networks

This lesson topic looks at compounds which have a covalent or ionic network structures.

Atomic Ions

This activity is about the formation of atomic ions from neutral atoms.

<table>
<thead>
<tr>
<th>Element</th>
<th>Covalent Radius (pm)</th>
<th>Atom Electron Arrangement</th>
<th>Ion Electron Arrangement</th>
<th>Ion Symbol</th>
<th>Ionic Radius (pm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium</td>
<td>154</td>
<td>2,8,1</td>
<td>2,8</td>
<td>Na⁺</td>
<td>102</td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>aluminium</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>phosphorus</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>sulphur</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>chlorine</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>calcium</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>chromium (II)</td>
<td></td>
<td>2,8,13,1</td>
<td>2,8,12</td>
<td></td>
<td></td>
</tr>
<tr>
<td>iron (II)</td>
<td></td>
<td>2,8,14,2</td>
<td>2,8,14</td>
<td></td>
<td></td>
</tr>
<tr>
<td>iron (III)</td>
<td></td>
<td>2,8,14,2</td>
<td>2,8,13</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Me atoms tend to form pos ions because they have 1, 2, or 3 electrons in their out shell. They will try to lo these electrons to reach the nea noble gas electron arrangement. Their pos charges (in the nuc ) outnumbers their elec so they become, overall, pos.

Non-me atoms tend to form neg ions because they have 5, 6, or 7 electrons in their out shell. They will try to ga extra electrons to reach the nea noble gas electron arrangement. Their pos charges (in the nuc ) are outnumbered by their elec so they become, overall, neg.

For an atom to lo electrons, another atom must be available to ga the electrons. A me atom cannot lo electrons if there is not a n metal atom available. Simple, two element, io compounds must always contain a met and a n metallic element.

Me ions are always sm than the ori atom either because:

1. they have lost the outer shell of electrons,

   or 2. rem electrons experience a larger effective charge from nucleus as proton:electron ratio has incr (due to loss of elec ).
Non-metal ions are always **bigger** than the original atom because:
removal electrons experience a weaker effective charge from nucleus as proton:electron ratio has **decreased** (due to **gain of electron**).

There will also be increased repulsion between the electrons which can also explain why the ionic radius increases.

**Simple Ionic Structures**

This activity looks more closely at the structure and bonding in a simple ionic compound such as sodium chloride.

The **ionic bond** is the **electrostatic attraction** between a **positive** and a **negative** ion.

An ion is able to attract, and be attracted by, several ions in all directions.

In sodium chloride, each Na⁺ ion attracts Cl⁻ ions.

Each Cl⁻ is also able to attract Na⁺. Overall, there are exactly equal numbers of the two ions, hence the formula NaCl.

The ionic bonding extends throughout the structure (net) and a regular repeating pattern also exists (lattice).

The strength of an ionic bond is affected by two factors:-

1. the amount of positive/negative charge
2. the size of the ions

A bigger ion has a weaker attraction than a smaller ion with the same charge: the charge is spread over a larger volume, weakening its effect.

This is why, for example, sodium fluoride (1000°C) has a higher MPt than sodium iodide (660°C):- the ionic bonding in NaF is stronger than in NaI.

This is why, for example, magnesium oxide (2825°C) has a higher MPt than calcium oxide (2054°C):- the ionic bonding in MgO is stronger than in CaO.

Predictions across a period are complicated by the fact that several factors can be changing at the same time. Namely:-

1. **size** of the ions (radius) - Na ( ) Mg ( ) Al ( )
2. **charge** on the ions - Na⁺ Mg²⁺ Al
3. **ratio** of the ions (formula) - MgO
4. **structure** of the lattice - some are 6:6 others are 8:8
Some ionic substances can also dissolve in water.

The polar nature of the water molecule allows it to attract both positive and negative ions.

The attractions between water molecules and these ions are similar in strength to the hydrogen bonds that exist between water molecules.

Water is reasonably unique in this ability to dissolve both ionic and covalent substances.

**Silicon Carbide**

This activity looks at the bonding and structure of the compound silicon carbide (also known as carborundum).

Silicon carbide is a covalent network compound:
- two different atoms are bonded together
- the covalent bonding extends throughout structure
- both are non-metals and will overlap their valence shells and share electrons

The silicon and carbon atoms are bonded together by the forces of attraction that exist between the shared pairs of electrons and the nuclei of the sharing atoms.

By sharing electrons, both the silicon and the carbon atoms achieve the extra stability of a full outer shell.

Like all network compounds, the structure of silicon carbide goes on ‘forever’. However, there is a simple repeating pattern of Si-C-Si-C-Si-C etc, so it can also be described as a covalent lattice and its form is simply SiC.
This activity looks at the structure and bonding in silicon dioxide (also known as silica)

Silicon dioxide is a covalent network compound:
- two different atoms are bonded together
- the covalent bonding extends throughout the structure
- both are non-metals and will overlap their electron shells and share electrons

The silicon and oxygen atoms are bonded together by the forces of attraction that exist between the shared pairs of electrons and the nuclei of the sharing atoms.

By sharing electrons, both the silicon and the oxygen atoms achieve the extra stability of a full outer shell.

Like all networks, the structure of silicon dioxide goes on ‘forever’. However, there is a simple repeating pattern so it can also be described as a covalent lattice. Each silicon atom is bonded to 4 oxygens while each oxygen is bonded to only 2 silicon atoms. Overall, there are twice as many oxygen atoms as silicon, and its formula is simplified to $\text{SiO}_2$.

Silicon carbide (carborundum) and silicon dioxide (silica) have the typical properties of covalent network substances.

**Very High Melting points** - a large number of very strong covalent bonds require to be broken before covalent networks can melt.

**Very Hard** - a large number of very strong covalent bonds require to be broken before fragments of covalent networks can be broken off.

**Very Unreactive** - a large number of very strong covalent bonds require to be broken before covalent networks can react with other chemicals.
Cov net are vulnerable to sudden stress (are bri) as the directional nature of the bonding (typically tetra) makes it difficult or impossible for bonds to re-establish if any of the atoms change pos. Diamonds, for example, are extr ha but can still be cut and shaped by the application of sudden impact.

Io net are similarly vulnerable to sudden stress (are bri). Though io bonding can act in all dir, shifting the pos of ions can result in ions of the same cha being forced too close. The resulting rep can break the network apart.

Met networks, however, are mall because the met bond can also act in all dir but can be easily re-established, even when atoms are moved quite considerably.
**Self Check**

**Q1.** Magnesium gives away two electrons to form the Mg\(^{2+}\) ion because this

A. gives a stable electron arrangement  
B. balances the nuclear charge  
C. gives an even number of energy levels  
D. balances the number of neutrons

**Q2.** Ionic bonding is the electrostatic attraction between

A. shared electrons and a pair of ions  
B. oppositely charged ions  
C. ions and free-moving electrons  
D. ions of the same charge

**Q3.** The structure of sodium chloride can be described as an ionic network because it has

A. covalent bonds linking ions together  
B. a regular repeating pattern  
C. more than one type of bonding  
D. ionic bonding throughout all of it

**Q4.** The structure of sodium chloride can be described as an ionic lattice because it has

A. covalent bonds linking ions together  
B. a regular repeating pattern  
C. more than one type of bonding  
D. ionic bonding throughout all of it

**Q5.** The formula of sodium chloride is NaCl because

A. one sodium ion is joined to one chloride ion  
B. the structure contains distinct NaCl molecules  
C. the overall structure gives an ionic lattice  
D. the number of sodium ions is equal to the number of chloride ions

**Q6.** The caesium ion is larger than the sodium ion. When comparing caesium chloride with sodium chloride you would expect caesium chloride to have

A. a lower melting point  
B. much stronger bonding  
C. a higher boiling point  
D. the same strength of bonding

**Q7.** Which of the following statements is true?

A. the potassium ion is larger than the potassium atom  
B. the chloride ion is smaller than the chlorine atom  
C. the sodium atom is larger than the sodium ion  
D. the oxygen atom is larger than the oxide ion

**Q8.** A positively charged particle with electron arrangement 2, 8 could be

A. a neon atom  
B. a fluoride ion  
C. a sodium atom  
D. an aluminium ion
**SELF CHECK**

**Q1.** Silicon carbide is called a covalent network compound because it has
A. a regular repeating pattern
B. atoms which share electrons
C. a network of covalent molecules
D. covalent bonds throughout the entire structure

**Q2.** Which structure represents the bonding arrangement in silicon carbide?

A. ![Structure A]
B. ![Structure B]
C. ![Structure C]
D. ![Structure D]

**Q3.** Covalent network compounds are usually
A. very hard and have low melting points
B. very flexible and have high melting points
C. very hard and have high melting points
D. very flexible and have low melting points

**Q4.** Silicon dioxide is a hard substance because it has
A. covalent bonds throughout
B. both ionic and covalent bonds
C. layers bonded by London forces
D. a molecular network structure

**Q5.** Silicon dioxide is used
A. for flexible garden hoses
B. as a high temperature lubricant
C. for adding fizz to lemonade
D. as the abrasive on sandpaper

**Q6.** Crucibles used for melting metals are often made of silicon carbide. This is because silicon carbide
A. is a very light-weight material
B. withstands very high temperatures
C. is a good conductor of heat
D. heats up very easily

**Q7.** The formula of silicon carbide is SiC because
A. the structure contains SiC molecules
B. each silicon atom is joined to one carbon atom.
C. the ratio of silicon atoms to carbon atoms is one to one
D. silicon atoms form covalent bonds with carbon atoms

**Q8.** The formula of silicon dioxide is SiO₂ because
A. the valency of silicon is two
B. each silicon atom is joined to two oxygen atoms
C. the valency of oxygen is two
D. there are twice as many oxygen atoms as silicon atoms
HOME PRACTICE

Q1. Work out the ion symbol and the electron arrangement for the ion that would be formed from each of the following atoms.

<table>
<thead>
<tr>
<th>Atom</th>
<th>Electron arrangement</th>
</tr>
</thead>
<tbody>
<tr>
<td>a) Na</td>
<td>2, 8, 1</td>
</tr>
<tr>
<td>b) P</td>
<td>2, 8, 5</td>
</tr>
<tr>
<td>c) Se</td>
<td>2, 8, 18, 6</td>
</tr>
</tbody>
</table>

Q2. The diagram represents the structure of calcium oxide.

a) What is the main type of bonding in this structure?

b) Within the structure (i.e. away from the edges) how many oxide ions are there around each calcium ion?

c) Why is the formula given simply as CaO?

d) Calcium comes immediately below magnesium in the periodic table and CaO (m.p. 2580 °C) has the same structure as MgO (m.p. 2800 °C). Explain in detail why the melting point for CaO is lower than that of MgO.

Q3. The enthalpy of lattice breaking is the energy required to completely separate the ions from one mole of an ionic solid. The table shows the enthalpies of lattice breaking, in kJ mol\(^{-1}\), for some alkali metal halides.

<table>
<thead>
<tr>
<th>Ions</th>
<th>F(^-)</th>
<th>Cl(^-)</th>
<th>Br(^-)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Li(^+)</td>
<td>1030</td>
<td>834</td>
<td>788</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>910</td>
<td>769</td>
<td>732</td>
</tr>
<tr>
<td>K(^+)</td>
<td>808</td>
<td>701</td>
<td>671</td>
</tr>
</tbody>
</table>

a) Write a general statement linking the enthalpy of lattice breaking to ion size

b) Predict the enthalpy of lattice breaking of lithium iodide.

Total (10)
HOME PRACTICE

Q1. Silicon carbide is used as an abrasive.

   a) What is the formula of silicon carbide?  1
   b) What type of bonding holds the atoms together in the structure?  1
   c) Why can the structure be referred to as a network?  1
   d) Why can the word lattice also be applied to this structure?  1
   e) What makes silicon carbide an extremely strong material?  1
   f) Why do structures like silicon carbide have very high melting points?  1

Q2. Silicon dioxide is also used as an abrasive, although it is not quite as strong as silicon carbide.

   a) How many oxygen atoms are bonded to each silicon atom in silicon dioxide?  1
   b) Why is the formula given as SiO$_2$?  1
   c) Quartz crystals are made of silicon dioxide. Which of the following temperatures would you expect for the melting point of quartz? 16 $^\circ$C 161 $^\circ$C 1610 $^\circ$C  1
   d) What is an abrasive?  1

Total (10)
Q1.  

a) Complete the table below by adding the name of an element for each of the types of bonding and structure described.

<table>
<thead>
<tr>
<th>Bonding and structure at room temperature and pressure</th>
<th>Name of element</th>
</tr>
</thead>
<tbody>
<tr>
<td>metallic solid</td>
<td>sodium</td>
</tr>
<tr>
<td>monatomic gas</td>
<td></td>
</tr>
<tr>
<td>covalent network solid</td>
<td></td>
</tr>
<tr>
<td>discrete covalent molecular gas</td>
<td></td>
</tr>
<tr>
<td>discrete covalent molecular solid</td>
<td></td>
</tr>
</tbody>
</table>

b) Many patterns in the physical and chemical properties of elements are observed. Why does the electronegativity of elements increase across the second row of the Periodic Table from lithium to fluorine?

c) The Periodic Table groups together elements with similar properties. In most Periodic Tables hydrogen is placed at the top of Group 1, but on some it is placed at the top of Group 7.

Using your knowledge of Chemistry, comment on why hydrogen can be placed in both Group 1 and Group 7.
CONSOLIDATION

Q1. The table below contains information about some diatomic molecules.

<table>
<thead>
<tr>
<th></th>
<th>H–H</th>
<th>H–Cl</th>
<th>Cl–Cl</th>
<th>I–Cl</th>
<th>Br–Br</th>
</tr>
</thead>
<tbody>
<tr>
<td>Boiling point / °C</td>
<td>-253</td>
<td>-85</td>
<td>-35</td>
<td>97</td>
<td>59</td>
</tr>
<tr>
<td>Bond enthalpy / kJ mol⁻¹</td>
<td>432</td>
<td>428</td>
<td>243</td>
<td>211</td>
<td>194</td>
</tr>
</tbody>
</table>

a) Boiling points can be used to show the effect of intermolecular forces. Explain why a comparison of the boiling points of ICl and Br₂ provides good evidence about the strength of permanent dipole-permanent dipole interactions.

b) In the table above, which of the diatomic molecules listed has the strongest covalent bond?

Q2. The Periodic Table allows chemists to make predictions about the properties of elements.

a) The elements lithium to neon make up the second period of the Periodic Table.
   i) Name an element from the second period that exists as a covalent network.
   ii) Why do the atoms decrease in size from lithium to neon?

b) On descending Group 1 from lithium to caesium, the electronegativity of the elements decreases. Explain clearly why the electronegativity of elements decreases as you go down the group.
Q1. Electronegativity values can be used to predict the type of bonding present in substances.

The type of bonding between two elements can be predicted using the diagram below.

\[ \text{Average electronegativity} \]

\[ \text{Difference in electronegativities} \]

\( a) \) Using the information in the diagram, state the highest average electronegativity found in ionic compounds.

\( b) \) The diagram can be used to predict the bonding in tin iodide.

Electronegativity of tin = 1.8 \hspace{1cm} \text{Average electronegativity} = 2.2

Electronegativity of iodine = 2.6 \hspace{1cm} \text{Difference in electronegativity} = 0.8

Predict the type of bonding in tin iodide.

\( c) \) The electronegativities of arsenic and chlorine are shown below.

Electronegativity of arsenic = 2.2

Electronegativity of chlorine = 3.0

Place a small cross on the diagram to show the position of arsenic chloride. *Show calculations clearly.*
Consolidation

Q1. Hydrogen cyanide is highly toxic. Molecules of hydrogen cyanide and molecules of nitrogen gas both have 14 electrons.

<table>
<thead>
<tr>
<th>Structure</th>
<th>Boiling point (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen cyanide</td>
<td>H–C≡N</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N≡N</td>
</tr>
</tbody>
</table>

Explain why the boiling point of hydrogen cyanide is much higher than the boiling point of nitrogen.

In your answer you should mention the intermolecular forces involved and how they arise.

Q2. Sodium is the first element in the third period of the periodic table.

<table>
<thead>
<tr>
<th>Na</th>
<th>Mg</th>
<th>Al</th>
<th>Si</th>
<th>P</th>
<th>S</th>
<th>Cl</th>
<th>Ar</th>
</tr>
</thead>
</table>

a) Describe the trend in covalent radii values across this period from Na to Cl.

b) Explain the trend in strength of metallic bonding values across this period from Na to Al.

c) The spike graph shows the first four ionisation energies for aluminium.

Explain why the fourth ionisation energy of aluminium is much higher than the third ionisation energy.
### KA 2.1

**The first 20 elements in the Periodic Table are categorised according to bonding and structure.**

Periodic trends and underlying patterns and principles.

<table>
<thead>
<tr>
<th>Elements are arranged in the Periodic Table in order of increasing atomic number.</th>
</tr>
</thead>
<tbody>
<tr>
<td>The Periodic Table allows chemists to make accurate predictions of physical properties and chemical behaviour for any element based on its position.</td>
</tr>
<tr>
<td>The first 20 elements in the Periodic Table can be categorised according to bonding and structure:</td>
</tr>
<tr>
<td>• metallic (Li, Be, Na, Mg, Al, K, Ca)</td>
</tr>
<tr>
<td>• covalent molecular ( \text{H}_2, \text{N}_2, \text{O}_2, \text{F}_2, \text{Cl}_2, \text{P}_4, \text{S}<em>8 ) and fullerenes (eg ( \text{C}</em>{60} ))</td>
</tr>
<tr>
<td>• covalent network (B, C (diamond, graphite), Si)</td>
</tr>
<tr>
<td>• monatomic (noble gases)</td>
</tr>
</tbody>
</table>

### KA 2.2

**Covalent radius, ionisation energy, electronegativity and trends in groups and periods, related to atomic structure.**

<table>
<thead>
<tr>
<th>The covalent radius is a measure of the size of an atom.</th>
</tr>
</thead>
<tbody>
<tr>
<td>The trends in covalent radius across periods and down groups can be explained in terms of the number of occupied shells, and the nuclear charge.</td>
</tr>
<tr>
<td>The trends in ionisation energies across periods and down groups can be explained in terms of the atomic size, nuclear charge and the screening effect due to inner shell electrons.</td>
</tr>
<tr>
<td>Atoms of different elements have different attractions for bonding electrons</td>
</tr>
<tr>
<td>Electronegativity is a measure of the attraction an atom involved in a bond has for the electrons of the bond.</td>
</tr>
<tr>
<td>Electronegativity values increase across a period and decrease down a group.</td>
</tr>
<tr>
<td>Electronegativity trends can be rationalised in terms of nuclear charge, covalent radius and the presence of ‘screening’ inner electrons.</td>
</tr>
</tbody>
</table>

### KA 3.1

**Bonding continuum**

Polar covalent bonds and their position on the bonding continuum, dipole formation and notation. \( \delta^+ \) \( \delta^- \),

\[ \text{eg} \quad \text{H}^{\delta^+} \rightarrow \text{Cl}^{\delta^-} \]

<table>
<thead>
<tr>
<th>In a covalent bond, atoms share pairs of electrons.</th>
</tr>
</thead>
<tbody>
<tr>
<td>The covalent bond is a result of two positive nuclei being held together by their common attraction for the shared pair of electrons.</td>
</tr>
<tr>
<td>Polar covalent bonds are formed when the attraction of the atoms for the pair of bonding electrons is different.</td>
</tr>
<tr>
<td>Delta positive ( (\delta^+) ) and delta negative ( (\delta^-) ) notation can be used to indicate the partial charges on atoms, which give rise to a dipole.</td>
</tr>
<tr>
<td>Pure covalent bonding and ionic bonding can be considered as being at opposite ends of a bonding continuum with polar covalent bonding lying between these two extremes.</td>
</tr>
<tr>
<td>The larger the difference in electronegativities between bonded atoms, the more polar the bond will be. If the difference is large then the movement of bonding electrons from the element of lower electronegativity to the element of higher electronegativity is complete resulting in the formation of ions.</td>
</tr>
<tr>
<td>Compounds formed between metals and non-metals are often, but not always ionic.</td>
</tr>
</tbody>
</table>
All molecular elements and compounds and monatomic elements condense and freeze at sufficiently low temperatures. For this to occur, some attractive forces must exist between the molecules or discrete atoms.

Any ‘intermolecular’ forces acting between molecules are known as van der Waals’ forces.

There are several different types of van der Waals’ forces such as London dispersion forces and permanent dipole: permanent dipole interactions, which include hydrogen bonding.

London dispersion forces are forces of attraction that can operate between all atoms and molecules. These forces are much weaker than all other types of bonding. They are formed as a result of electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules.

The strength of London dispersion forces is related to the number of electrons within an atom or molecule.

A molecule is described as polar if it has a permanent dipole.

The spatial arrangement of polar covalent bonds can result in a molecule being polar.

Permanent dipole-permanent dipole interactions are additional electrostatic forces of attraction between polar molecules.

Permanent dipole-permanent dipole interactions are stronger than London dispersion forces for molecules with similar numbers of electrons.

Bonds consisting of a hydrogen atom bonded to an atom of a strongly electronegative element such as fluorine, oxygen or nitrogen are highly polar.

Hydrogen bonds are electrostatic forces of attraction between molecules which contain these highly polar bonds.

A hydrogen bond is stronger than other forms of permanent dipole-permanent dipole interaction, but weaker than a covalent bond.

Melting points, boiling points and viscosity can all be rationalised in terms of the nature and strength of the intermolecular forces which exist between molecules.

By considering the polarity and number of electrons present in molecules, it is possible to make qualitative predictions of the strength of the intermolecular forces.

The melting and boiling points of polar substances are higher than the melting and boiling points of non-polar substances with similar numbers of electrons.

The anomalous boiling points of ammonia, water and hydrogen fluoride are a result of hydrogen bonding. Boiling points, melting points, viscosity and solubility/miscibility in water are properties of substances which are affected by hydrogen bonding.

Hydrogen bonding between molecules in ice results in an expanded structure which causes the density of ice to be less than that of water at low temperatures.

Ionic compounds and polar compounds tend to be soluble in polar solvents such as water and insoluble in non-polar solvents.

Non-polar substances tend to be soluble in non-polar solvents and insoluble in polar solvents.