Bonding, structure and trends (periodicity)

Bonding and structure revision

Objectives

- Draw and interpret dot and cross diagrams
- Use VSEPR to predict the shapes of molecules
- Describe covalent, ionic and metallic bonding
- Describe structures of substances
- Explain physical properties (MP, BP, conductivity etc based on structure and intermolecular forces).
- Describe trends in physical properties

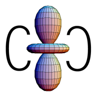
Definitions - LEARN

<u>Ionic bonding</u>

- Electrostatic interaction between oppositely charged ions

<u>Covalent bonding</u>

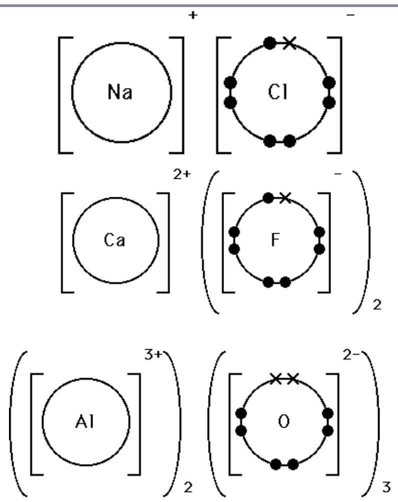
- A bond between two atoms involving a shared pair of electrons



Dot and cross diagrams - IONIC

Remember:

- Square brackets
- Charges
- Transferred electron(s)
- Balanced
- DO NOT show sharing



Past paper question

(b) Sodium reacts with oxygen to form sodium oxide, Na_2O, which is an ionic compound.

(i) Write the equation for the reaction of sodium with oxygen to form sodium oxide.

.... **[1]**

(ii) State what is meant by the term *ionic bond*.

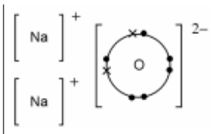
....

.... **[1]**

(iii) Draw a '*dot-and-cross*' diagram to show the bonding in Na_2O.

Show **outer** electrons only.

Mark scheme

(b)	(i)	$4\,Na + O_2 \longrightarrow 2\,Na_2O$ **OR** $2\,Na + \frac{1}{2}O_2 \longrightarrow Na_2O$ ✓	1	**ALLOW** correct multiples including fractions **IGNORE** state symbols
	(ii)	(electrostatic) attraction between oppositely charged ions✓	1	
	(iii)	 Na shown with either 8 or 0 electrons **AND** O shown with 8 electrons **with** 6 crosses and 2 dots (or vice versa) ✓ Correct charges on both ions ✓	2	**For 1st mark,** if 8 electrons shown around cation then 'extra' electron(s) around anion must match symbol chosen for electrons in cation Shell circles not required **IGNORE** inner shell electrons **ALLOW:** $2[Na^+]$ $2[Na]^+$ $[Na^+]_2$ (brackets not required) **DO NOT ALLOW** $[Na_2]^{2+}$ / $[Na_2]^+$ / $[2Na]^{2+}$ **DO NOT ALLOW:** $[Na_2]^{2+}$ $[Na_2]^+$ $[2Na]^{2+}$ $[Na]_2^+$

This question is about different models of bonding and molecular shapes.

(a) Magnesium sulfide shows ionic bonding.

(i) What is meant by the term *ionic bonding*?

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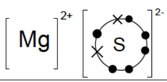
.... **[1]**

(ii) Draw a '*dot-and-cross*' diagram to show the bonding in magnesium sulfide.
Show outer electron shells only.

[2]

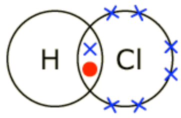
Mark scheme

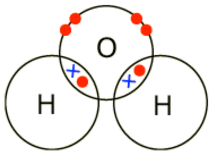
3	(a)	(i)	(Electrostatic) **attraction** between oppositely charged **ions**. ✓	1	**IGNORE** force **IGNORE** references to transfer of electrons **MUST** be ions, not particles
		(ii)	Mg shown with either 8 of 0 electrons **AND** S shown with 8 electrons **with** 2 crosses and 6 dots (or vice versa) ✓ Correct charges on both ions ✓ 	2	Mark charges on ions and electrons independently **For first mark**, if 8 electrons are shown around the Mg then 'extra electrons' around S must match the symbol chosen for electrons around Mg Shell circles not required **IGNORE** inner shell electrons Brackets are not required

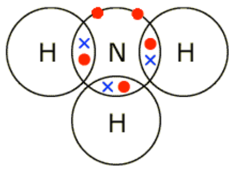
Dot and cross diagrams - COVALENT

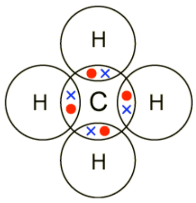
Remember

- Show sharing
- Overlap or touch circles
- Show outer electrons – easy to forget for halogens









Chemists have developed models for bonding and structure which are used to explain different properties.

(a) Ammonia, NH_3, is a covalent compound.

(i) Explain what is meant by a *covalent bond*.

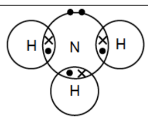
.... **[1]**

(ii) Draw a '*dot-and-cross*' diagram to show the bonding in NH_3.

Show **outer** electrons only.

[1]

Mark scheme

i	a shared pair of electrons ✓	1	**ALLOW** any response that communicates electron pair **ALLOW** shared pairs
ii	 ✓	1	Must be '*dot-and-cross*' circles for outer shells **NOT** needed **IGNORE** inner shells Non-bonding electrons of N do not need to be shown as a pair.

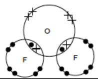
'*Dot-and-cross*' diagrams can be used to predict the shape of covalent molecules.

Fluorine has a covalent oxide called difluorine oxide, F_2O. The oxygen atom is covalently bonded to each fluorine atom.

(i) Draw a '*dot-and-cross*' diagram of a molecule of F_2O.
Show outer electron shells only.

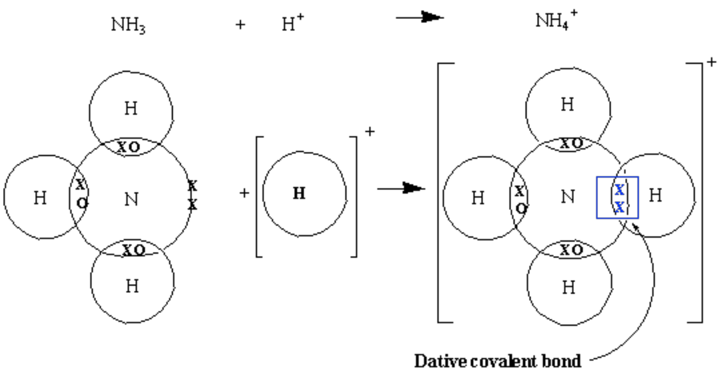
[2]

Mark scheme

(b)	**(i)**	Electron pairs in covalent bonds shown correctly using dots and crosses in a molecule of the F_2O ✓ Lone pairs correct on O and both F atoms ✓ 	**2**	Must be '*dot-and-cross*' circles for outer shells **NOT** needed **IGNORE** inner shells Non-bonding electrons of O do not need to be shown as pairs Non-bonding electrons of F do not need to be shown as pairs

Dative covalent bonds

- Both shared electrons come from one atom
- Show as two dots or two crosses together
- Key example is the ammonium ion



Ammonia reacts with hydrogen chloride, HC*l*, to form ammonium chloride, NH_4Cl.

NH_4Cl is an ionic compound containing NH_4^+ and Cl^- ions.

(i) Complete the electron configuration of the Cl^- ion.

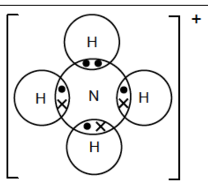
$1s^2$ .... **[1]**

(ii) Draw a '*dot-and-cross*' diagram to show the bonding in NH_4^+.

Show **outer** electrons only.

[1]

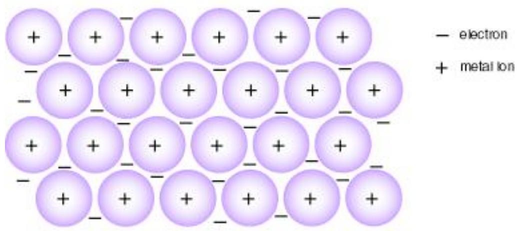
Mark scheme

i	$1s^2 2s^2 2p^6 3s^2 3p^6$ ✓	1	**ALLOW** subscripts
ii	 '*Dot-and-cross*' diagram to show four shared pairs of electrons one of which is a dative covalent bond (which must consist of the same symbols) ✓	1	**IGNORE** inner shells **IGNORE** '+' sign **BUT** a **DO NOT ALLOW** '–' sign. Brackets and circles not required

Metallic bonding

Definition: Electrostatic attraction between metal cations and delocalised electrons

Structure is giant metallic



This question compares the bonding, structure and properties of sodium and sodium oxide.

(a) Sodium, Na, is a metallic element.

Explain, with the aid of a labelled diagram, what is meant by the term *metallic bonding*.

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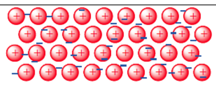
.... **[3]**

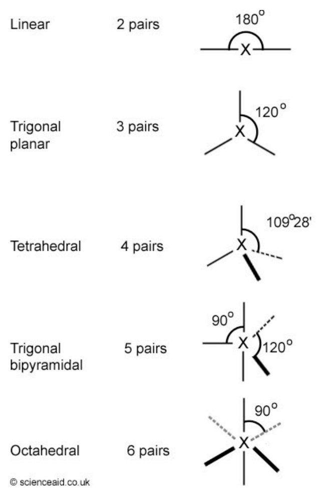
Mark scheme

2	(a)		 regular arrangement of **labelled** + ions with some attempt to show electrons ✓ scattering of labelled electrons **between** other species **OR** a statement anywhere of **delocalised** electrons (can be in text below) ✓ metallic bond as (electrostatic) **attraction** between the electrons and the positive ions ✓	3	Lattice must have at least 2 rows of positive ions If a metal ion is shown (e.g. Na^+), it must have the correct charge **ALLOW** for labels: + ions, positive ions, cations If '+' is unlabelled in diagram, award the label for '+' from a statement of 'positive ions' in text below **DO NOT ALLOW** as label or text positive atom **OR** protons **OR** nuclei **ALLOW** e^- **OR** e as label for electron **DO NOT ALLOW** '–' as label for electron

Shapes of molecules

- <u>Electron pairs repel</u> ensuring bonds are as far away from each other as possible.
- <u>Lone pairs repel more</u> then bonding pairs
- To determine shape/bond angle decide <u>how many bonding pairs and how many lone pairs</u> there are

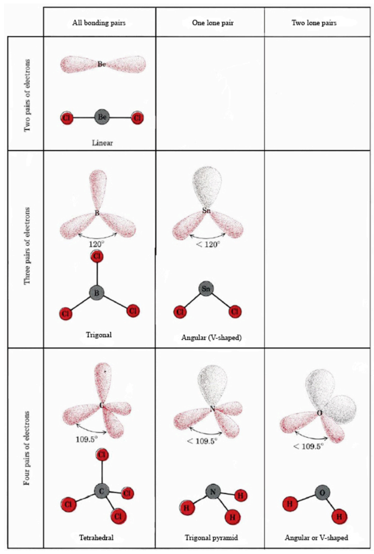
Shapes of molecules - all bonding pairs



Shapes of molecules with lone pairs

Common examples

- Water – <u>bent</u> 2LP
- SO_2 – <u>bent</u> 1LP
- Ammonia – <u>pyramidal</u> 1LP
- XeF_4 – <u>square planar</u> 2LP



Name the shape of the ammonia molecule.

Explain, using your '*dot-and-cross*' diagram, why ammonia has this shape and has a bond angle of 107°.

shape: ....

explanation: ....

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.... **[3]**

Mark scheme

Shape: pyramidal **OR** (trigonal) pyramid ✓ Explanation: There are 3 bonded pairs and 1 lone pair ✓ Lone pairs repel more than bonded pairs ✓	**3**	**ALLOW** 'bonds' for 'bonded pairs' **DO NOT ALLOW** 'atoms repel' **DO NOT ALLOW** electrons repel **ALLOW** LP for 'lone pair' **ALLOW** BP for bonded pair

Draw a '*dot-and-cross*' diagram to show the bonding in NH_4^+.

Show **outer** electrons only.

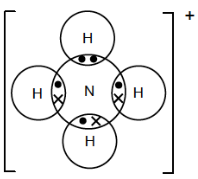
[1]

State the shape of, and bond angle in, an NH_4^+ ion.

shape: ....

bond angle: .... **[2]**

Mark scheme

ii	 '*Dot-and-cross*' diagram to show four shared pairs of electrons one of which is a dative covalent bond (which must consist of the same symbols) ✓	1	**IGNORE** inner shells **IGNORE** '+' sign **BUT** a **DO NOT ALLOW** '–' sign. Brackets and circles not required
iii	tetrahedral ✓ 109.5° ✓	2	**ALLOW** 109–110°

(c) Boron trifluoride, BF_3, ammonia, NH_3, and sulfur hexafluoride, SF_6, are all covalent compounds. The shapes of their molecules are different.

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

.... **[1]**

(ii) Using outer electron shells only, draw *'dot-and-cross'* diagrams for molecules of BF_3 and NH_3.

Use your diagrams to explain why a molecule of BF_3 has bond angles of 120° and NH_3 has bond angles of 107°.

BF_3	**NH_3**

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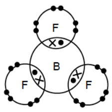
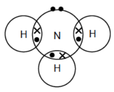
.... **[5]**

Mark scheme

(i)	octahedral **OR** octahedron ✓	1	
(ii)	  Diagram of BF_3 showing three '*dot-and-cross*' bonds between B and F and all F atoms with complete octet of electrons ✓ Diagram of NH_3 showing three '*dot-and-cross*' bonds between N and H and N atom has a lone pair ✓ ***Marking points 3, 4 and 5 may be awarded independently*** electron pairs repel ✓ NH_3 has **one lone** pair and **three bonding** pairs of electrons **AND** lone pair of electrons repels **more** than bonding pairs ✓ BF_3 has **three** (bonding) pairs of electrons (which repel equally) ✓	5	**Use annotations with ticks, crosses ECF etc. for this part** **ALLOW** diagrams without circles Must be '*dot-and-cross*' **IGNORE** 'electrons repel' **DO NOT ALLOW** 'atoms repel' **ALLOW** 'bonds repel' **ALLOW** 'bonds' for 'bonding pairs' **ALLOW** 'four pairs' in place of 'one lone pair and three bonding pairs' The third marking point can be gained from statements seen in fourth or fifth marking points

Structures

Ionic = giant ionic lattice

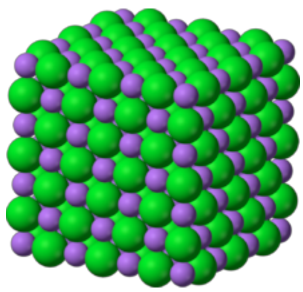
Metallic = giant metallic lattice

Covalent

1. Simple molecular
2. Simple molecular lattice
3. Giant covalent structure

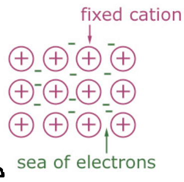
Giant ionic lattice

- High MP and BP due to strong ionic bonding
- Soluble in water as ions can become solvated
- Conducts when molten or dissolved as ions can move
- Examples NaCl, $MgCl_2$.



Giant metallic lattice

- High MP and BP due to strong metallic bonding - increases along a period (eg Na, Mg, Al) due to increasing cationic charge and more delocalised electrons
- Insoluble in water
- Conducts when solid or liquid due to delocalised electrons which can move



Past Paper Question

The table below shows the melting points and atomic radii of the elements in Period 3, Na to Cl.

element	Na	Mg	Al	Si	P	S	Cl
melting point/°C	98	639	660	1410	44	113	–101
atomic radius/pm	186	160	143	118	110	102	99

1 pm = 1×10^{-12} m

(a) (i) Explain the difference in melting point for the elements Na and Mg.

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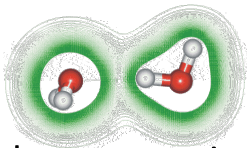
.... [3]

Mark scheme

Magnesium ions have a greater charge✓ Magnesium has more (delocalised **OR** outer) **electrons**✓ Magnesium has greater attraction between **ions** and **electrons OR** has stronger **metallic** bonds ✓	3	***USE annotations with ticks, crosses, ecf, etc for this part.*** **ALLOW REVERSE ARGUMENT** e.g. sodium ions have a smaller charge **ALLOW** Mg^{2+} / Mg ion / Na ion / Na^{+} ion **ALLOW** 'charge density' as alternative to 'charge' **ALLOW REVERSE ARGUMENT** e.g. sodium has fewer electrons **ALLOW REVERSE ARGUMENT** e.g. sodium has less attractions between **ions** and **electrons** **OR** has weaker **metallic** bonds ✓

Simple molecular

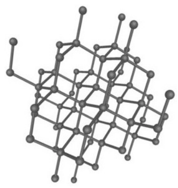
- Low MP and BP as need to break weak intermolecular forces such as instantaneous dipole-induced dipole, permanent dipole-permanent dipole and hydrogen bonding
- Examples:
- H_2O, CO_2, N_2, F_2, SF_6



- No mobile charge carriers (electrons or ions) so do not conduct

Giant Covalent Lattice

- High MP and BP due to strong covalent bonds
- Generally do not conduct (except graphite*)
- Insoluble



Examples:
Si, C (diamond, graphite)

*Graphite conducts electricity because of delocalised electrons between it's layer structure. There are weak intermolecular forces between these layers which is why graphite is soft and diamond is hard.

(b) One form of naturally occurring carbon is graphite.

The table below lists some properties of graphite.

electrical conductivity	good conductor
hardness	soft
melting point	very high

- Describe the bonding and structure in graphite.
- Explain, in terms of bonding and structure, the properties of graphite shown above.

In your answer, you should use appropriate technical terms, spelt correctly.

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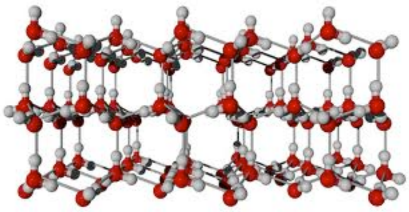
.... [5]

Mark scheme

	5	Use annotations with ticks, crosses etc. for this part.
		All five marking points are independent
giant covalent (lattice) ✓		**ALLOW** giant atomic **OR** giant molecular **OR** macromolecular
layers ✓		**ALLOW** planes **OR** sheets Allow diagram showing at least two layers
Each of the three properties below must be linked to explanation *good conductor* - because it has mobile electrons **OR** delocalised electrons **OR** electrons can move ✓		Electron(s) must be spelt correctly ONCE
high melting / boiling point - because strong **OR** covalent bonds have to be broken ✓		**DO NOT ALLOW** 'strong ionic bonds' **OR** strong metallic bonds.
soft - because there are van der Waals' forces **OR**		

Simple molecular lattice

- Low MP and BP as need to break weak intermolecular forces such as instantaneous dipole-induced dipole, permanent dipole-permanent dipole and hydrogen bonding
- Examples: I_2, H_2O (ice), P_4
- Do not conduct



Intermolecular forces in simple covalent structures

These are the forces holding molecules together in simple covalent structures / molecules / monatomic gases.

- Permanent dipole – permanent dipole
- Van der Waals
- Hydrogen Bonding

Dipole – dipole interactions

- Molecules such as HCl have a permanent dipole. δ+ H-Cl δ-
- This means that the molecules will attract each other
- This is called dipole-dipole

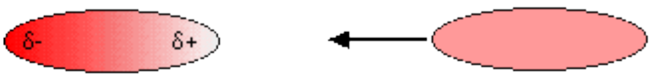
δ+ δ− δ+ δ−
H—Cl- - - - -H—Cl

Van der Waals Forces

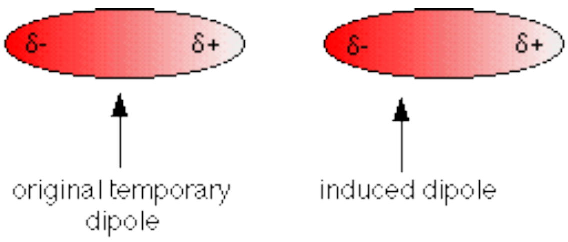
All atoms and molecules contain fast moving electrons.

At some point in time more electrons may be on one side of the molecule than the other causing instantaneous dipoles in the molecule

Imagine a molecule which has a INSTANTANEOUS polarity being approached by one which happens to be entirely non-polar just at that moment



This sets up an ***induced dipole*** in the approaching molecule, which is orientated in such a way that the d+ end of one is attracted to the d - end of the other.



INSTANTANEOUS DIPOLE INDUCED DIPOLE

Chlorine, bromine and iodine are halogens commonly used in school and college experiments.

(a) Halogens have van der Waals' forces between their molecules.

(i) Describe how van der Waals' forces arise.

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.... **[3]**

Mark scheme

Creating the dipole mark uneven distribution of electrons ✓	3	Use annotations with ticks, crosses ECF etc. for this part **ALLOW** movement of electrons **ALLOW** changing electron density
Type of dipole mark creates an instantaneous dipole **OR** temporary dipole ✓		**ALLOW** 'transient', 'oscillating', 'momentary', 'changing'
Induction of a second dipole mark causes induced dipole(s) in neighbouring molecules ✓		**ALLOW** 'induces a dipole in neighbouring molecules' **ALLOW** 'causes a resultant dipole in neighbouring molecules' **ALLOW** 'atoms' for 'molecules'

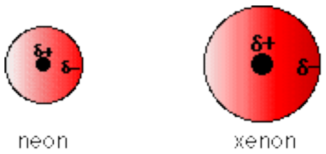
Noble gases

helium	-269°C
neon	-246°C
argon	-186°C
krypton	-152°C
xenon	-108°C
radon	-62°C

As you go down the group is that the number of electrons increases, and so also does the radius of the atom.

This makes the atom more polarisable and so the Van der Waals forces increase.

This increases the boiling point.



The boiling points of some Group 7 elements are shown below.

Group 7 element	boiling point/°C
chlorine	–35
bromine	59
iodine	184

Explain why the halogens show this trend in boiling points.

In your answer, you should use appropriate technical terms, spelt correctly.

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.... [3]

Mark scheme

van der Waals' forces **OR** induced dipole interactions ✓ number of electrons increases ✓ **Down the group**, intermolecular forces / van der Waals' forces increase **OR** **Down the group**, more energy needed to break intermolecular / van der Waals' forces ✓	3	**electron(s) must be seen and spelt correctly ONCE** **ALLOW** number of electron shells increases **ALLOW** iodine has most electrons **ALLOW** chlorine has the least electrons For **'Down the group'** **ALLOW** 'Increase in boiling points' or 'Molecules get bigger'

(ii) State **and** explain the trend in the boiling points of chlorine, bromine and iodine.

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.... **[3]**

(b) The halogen astatine does **not** exist in large enough quantities to observe any of its reactions.

Why would astatine be expected to react similarly to other halogens?

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.... **[1]**

Mark scheme

	(ii)	boiling points increase down the group ✓ greater number of electrons **OR** stronger intermolecular forces **OR** stronger van der Waals' forces ✓ more energy needed to break intermolecular **OR** van der Waals' forces ✓	3	Use annotations with ticks, crosses ECF etc. for this part **ALLOW** Bpt of iodine is highest **OR** Bpt of chlorine is lowest **ALLOW** Cl for chlorine etc. For 'down the group' **ALLOW** 'as molecules get bigger' **ALLOW** number of **electron** shells increases **IGNORE** 'more shells' (if no reference to electrons) **ALLOW** 'more' for 'stronger' **ALLOW** iodine has most electrons **ALLOW** chlorine has fewest electrons **DO NOT ALLOW** any implication that the attraction is between atoms not molecules for third mark
(b)		Same number of **outer(most)** electrons **OR** same **outer(most)** electron structure ✓	1	**ALLOW** same number of electrons in outer shell **ALLOW** It has seven outer electrons **IGNORE** same group **DO NOT ALLOW** 'same number of electrons'

Molecular mass and shape

- Heavier molecules contain more electrons and generally have stronger Van der Waals forces.
- Straight molecules can approach more closely than branched molecules. This means the Van der Waals forces are stronger.

butane	CH_3—CH_2—CH_2—CH_3	B.Pt: -0.5°C
2-methylpropane	CH_3—CH(CH_3)—CH_3	B.Pt: -11.7°C

Past paper question

Sulfur exists as S_8 molecules and chlorine as Cl_2 molecules. Use this information to explain the difference in their melting points.

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.... [2]

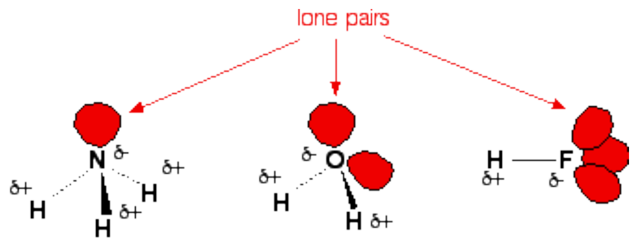
Mark scheme

Cl_2 **OR** S_8 has intermolecular **OR** van der Waals' forces ✓ S_8 has stronger intermolecular forces **OR** van der Waals' forces than Cl_2 **OR** S_8 has more electrons ✓	2	**ALLOW REVERSE ARGUMENT** ie Cl_2 has weaker intermolecular forces **OR** van der Waals' forces **DO NOT ALLOW** comparison involving covalent bonds **ALLOW REVERSE ARGUMENT** Cl_2 has fewer electrons

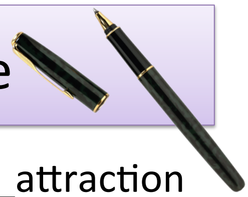
The origin of hydrogen bonding

To have a hydrogen bond you need:

1. Hydrogen; bonded to
2. Nitrogen, oxygen or fluorine (the most electronegative elements)
3. At least 1 lone pair on the N, O or F.



Copy and complete

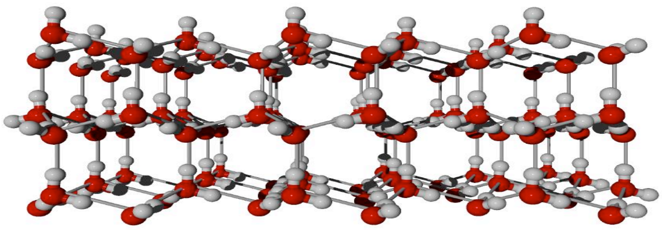


A __________ is a strong __________attraction between an electron deficient __________ atom on one molecule and a __________of electrons on a highly __________ atom on a different molecule.

[dipole-dipole, electronegative, hydrogen bond, hydrogen, lone pair]

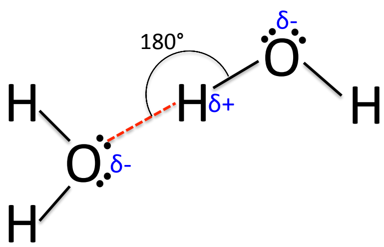
The anomalous properties of water

- Ice has a lower density than water because the hydrogen bonds hold the structure apart in an open lattice.
- Water has a relatively high freezing point and boiling point for its Mr because the hydrogen bonds require a lot of energy in order to break them.

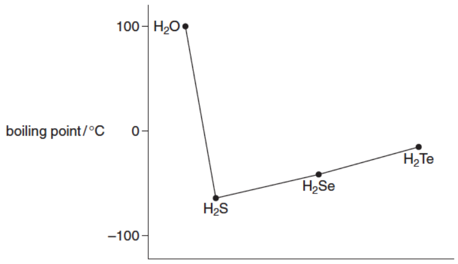


Drawing hydrogen bonds

- For exam questions it is really important to draw hydrogen bonds a certain way.
- Up to 4 marks could be available!



(b) The figure below shows the boiling points of four hydrides of Group 6 elements.



(i) Explain, with the aid of a diagram, the intermolecular forces in H_2O that lead to the relatively high boiling point of H_2O.

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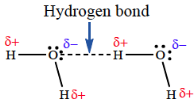
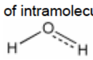
.... **[3]**

Mark scheme

		Answer	Marks	Guidance
(b)	**(i)**	 Shape of water with at least one H with δ+ and at least one O with δ− ✓ H-bond between H in one water molecule and a lone pair of an O in another water molecule✓ hydrogen bond labelled **OR** H_2O has hydrogen bonding ✓	**3**	all marks can be awarded from a labelled diagram If HO_2 shown then **DO NOT ALLOW** 1st mark Dipole could be described in words so it does **not** need to be part of diagram. At least one hydrogen bond **must** clearly hit a lone pair Lone pair interaction could be described in words so it does **not** need to be part of diagram. **DO NOT ALLOW** hydrogen bonding if described in context of intramolecular bonding, *ie* 

Liquid ammonia, NH_3, and water, H_2O, both show hydrogen bonding.

(i) Draw a labelled diagram to show hydrogen bonding between two molecules of liquid **ammonia**.

[3]

(ii) Water has several anomalous properties as a result of its hydrogen bonding.

Describe and explain **one** anomalous property of water which results from hydrogen bonding.

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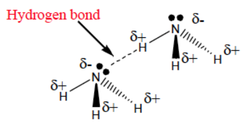
.... **[2]**

(i)	(At least) two **NH_3** molecules with correct dipole shown with at least one H with δ^+ and one N with δ^- ✓ (Only) one hydrogen bond from N atom on one molecule to a H atom on another molecule ✓ Lone pair shown on the N atom and hydrogen bond must hit the lone pair ✓ 	3	**DO NOT ALLOW** first mark for ammonia molecules with incorrect lone pairs **DO NOT ALLOW** first mark if H_2O, NH_2 or NH is shown **ALLOW** hydrogen bond need not be labelled as long as it clear the bond type is different from the covalent N–H bond **ALLOW** a line (i.e. looks like a covalent bond) as long as it is labelled 'hydrogen bond) **ALLOW** 2-D diagrams **ALLOW** two marks if water molecules are used. One awarded for a correct hydrogen bond and one for the involvement of lone pair
(ii)	Liquid H_2O is denser than solid ✓ In solid state H_2O molecules are held apart by hydrogen bonds **OR** ice has an open lattice ✓ **OR** H_2O has a relatively high boiling point **OR** melting point ✓ (relatively strong) hydrogen bonds need to be broken **OR** a lot of energy is needed to overcome hydrogen bonds **OR** hydrogen bonds are strong ✓	2	ORA **ALLOW** ice floats for first mark **ALLOW** higher melting **OR** boiling point than expected **DO NOT ALLOW** H_2O has a high melting / boiling point **ALLOW** other properties caused by hydrogen bonding not mentioned within the specification E.g. high surface tension – strong hydrogen bonds on the surface

Properties - Conductivity

- Conductivity is dependent on **mobile charge carriers**
- These are delocalised electrons in metals and graphite
- In ionic solids and liquids and in solution the charge carriers are ions

Past paper question

Compare and explain the electrical conductivities of sodium and sodium oxide in the solid and liquid states.

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.... [5]

Mark scheme

sodium is a (good) conductor because it has mobile electrons **OR** delocalised electrons **OR** electrons can move ✓ sodium oxide does not conduct as a solid ✓ sodium oxide conducts when it is a liquid ✓ ions cannot move in a solid ✓ **ions** can move **OR** are mobile when liquid ✓	5	**Throughout this question, 'conducts' and 'carries charge' are treated as equivalent terms.** **DO NOT ALLOW** 'free electrons' for mobile electrons **ALLOW** poor conductor **OR** bad conductor 'Sodium oxide only conducts when liquid' is insufficient to award 'solid conductivity' mark **ALLOW** ions are fixed in place **IGNORE** electrons **IGNORE** charge carriers **IGNORE** 'delocalised ions' or 'free ions' for mobile ions Any mention of electrons moving is a **CON**

Past paper question

Fluorine reacts with lithium at room temperature to form a white crystalline solid, lithium fluoride. Lithium fluoride is a good conductor of electricity when molten but not when solid.

(i) Draw a '*dot-and-cross*' diagram to show the bonding in lithium fluoride.
Show the outer electrons only.

[2]

(ii) Explain why lithium fluoride conducts electricity when molten but **not** when solid.

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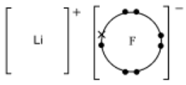
.... **[2]**

Mark scheme

5	c	i	 Li shown with either 2 or 0 electrons **and** F shown with 8 electrons with 7 crosses and one dot (or *vice versa*) ✓ correct charges on both ions ✓	2	**For first mark**, if 2 electrons are shown in the cation then the 'extra' electron in the anion must match symbol chosen for electrons in the cation **IGNORE** inner shell electrons **ALLOW** 'Fl' for fluorine Circles not essential **DO NOT ALLOW** Li^+ with 8 electrons Second mark is independent
		ii	Ions cannot move in a solid ✓ **Ions** can move **OR** are mobile when molten ✓	2	**ALLOW** ions are fixed in place **IGNORE** electrons **IGNORE** 'charge carriers' or 'charged particles' **DO NOT ALLOW** ions can move when in solution **IGNORE** charge carriers **IGNORE** 'delocalised ions' or 'free ions' **ALLOW** '**Ions** can only move when molten' for one mark Any mention of electrons moving when molten is a **CON**

Properties – high MP or BP

- High MP or BP usually associated with breaking covalent, ionic or metallic bonds
- ie to melt a giant covalent structure, giant ionic lattice or giant metallic lattice
- This requires a large amount of energy

Past paper question

This question is about elements in Period 2 of the Periodic Table.

(a) Lithium has a giant metallic structure and a boiling point of 1342 °C.

Describe, with the aid of a labelled diagram, the structure and bonding in lithium and explain why lithium has a high boiling point.

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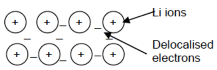
.... **[3]**

Mark scheme

5	a	 Diagram showing a regular arrangement of **labelled** '**Li^+**' or **'+ ions'** with some attempt to show electrons ✓ Scattering of **labelled** electrons **between** other species **OR** a statement anywhere of **delocalised** electrons (can be in text or in diagram) ✓ The attraction between + ions and e^- is strong **OR** metallic bonding is strong ✓	3	Lattice diagram must have at least two rows of correctly charged ions and a minimum of 2 ions per row **ALLOW** as label: + ions, positive ions, cations If '+' is unlabelled in diagram, award label from a correct statement within the text below **DO NOT ALLOW** 2+, 3+ etc ions **DO NOT ALLOW** for label or in text: nuclei **OR** positive atom **OR** protons **ALLOW** e^- **OR** e as label for electron **ALLOW** a lot of energy is needed to break the (metallic) bond **DO NOT ALLOW** incorrect particles or incorrect attraction e.g. 'intermolecular attraction' or 'nuclear attraction'

Properties – high MP or BP

- Sometimes this is contrasting Van der Waals with hydrogen bonding
- In water for example the boiling point is higher than would be expected for Van der Waals alone.

Properties – low MP or BP

- This is usually associated with simple molecular covalent substances or simple molecular lattices or monatomic gases
- Eg water, SF_6, N_2, I_2, H_2, Ne, Ar
- It requires breaking only weak intermolecular forces such as Van der Waals or permanent dipole - permanent dipole or hydrogen bonds

The third period of the Periodic Table features the elements magnesium and chlorine. The table below shows the melting points of these elements.

element	melting point / °C
magnesium	650
chlorine	–101

Describe the structure and bonding shown by these elements. Use your answer to explain the difference in melting points.

In your answer, you should use appropriate technical terms spelt correctly.

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.... [6]

Mg has a **giant** structure ✓	6	
Mg has **metallic** bonding OR description of metallic bonding as positive ions and **delocalised** electrons ✓		**Metallic OR delocalised** seen spelt correctly at least **ONCE**
(There is electrostatic attraction between) positive ions and electrons ✓		**DO NOT ALLOW** as label nuclei **OR** protons for positive ions **ALLOW labelled** diagram of metallic bonding for second and third marks  Lattice must have at least two rows of positive ions. If a Mg ion is shown it must correct charge **ALLOW** for labels:+ ions, positive ions, cations **DO NOT ALLOW** as label nuclei **OR** protons for positive ions **ALLOW** e^- or e as label for electron **DO NOT ALLOW** '–' without label for electron
Cl has a simple molecular **OR** simple covalent (lattice) ✓		**Covalent OR molecule OR molecular** seen spelt correctly at least **ONCE** **ALLOW** Cl is a (covalent) **molecule**
Cl has van der Waals' forces (between molecules) **OR** Cl has instantaneous dipole–induced dipoles **OR** temporary dipole–temporary dipole ✓		**IGNORE** Cl has intermolecular bonding
van der Waals' forces are weak **and** metallic bonds are strong **OR** van der Waals' forces are weaker than metallic bonds **OR** Less energy is needed to overcome van der Waals' than metallic bonds ✓		**ALLOW** ECF from incorrect descriptions of giant structure with strong bonds; e.g. Mg has giant ionic structure **ALLOW** ECF from any incorrect intermolecular forces e.g. permanent dipole –dipole from marking point 5 **ALLOW** vdW easier to break ORA

Electronegativity

Electronegativity is a measure of the attraction of a bonded atom for the pair of electrons in a covalent bond.

Electronegativity is measure on the Pauling scale

Bond polarity

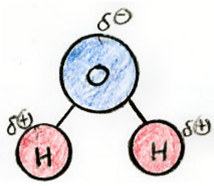
When two atoms of differing electronegativity are bonded together the electrons are drawn towards the atom with the higher electronegativity.

A permanent dipole is a small charge difference across a bond that results from a difference in the electronegativities of the bonded atoms.

A polar covalent bond has a permanent dipole

Predict the polarity - AfL

- Draw the dipole(s) for the following molecules.
- Eg water



- Ammonia
- Hydrogen chloride
- Carbon tetrafluoride
- Di-chloro, di-fluoro methane

Is the molecule polar?

- For molecules that are symmetrical, the dipoles of any bonds within the molecule may <u>cancel out</u>.
- For example CCl_4 or CBr_2CBr_2.
- We can say that the centre of positive charge is in the same place as the centre of negative charge.
- Where the dipoles do not cancel out the molecule has an area of partially positive and an area of partially negative charge.

Past paper question

Linus Pauling was a Nobel prize winning chemist who devised a scale of electronegativity.

Some Pauling electronegativity values are shown in the table.

element	electronegativity
B	2.0
Br	2.8
N	3.0
F	4.0

(a) What is meant by the term *electronegativity*?

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....

....

.... **[2]**

(b) Show, using δ+ and δ– symbols, the permanent dipoles on each of the following bonds.

N—F　　　　　N—Br

[1]

Mark scheme

The ability of an **atom** to attract electrons ✓ in a covalent bond ✓	2	**ALLOW** 'attraction of an **atom** for electrons' **ALLOW** 'pull' for 'attract' **DO NOT ALLOW** 'element' for 'atom' **ALLOW** 'shared pair' or 'bond(ing) pair' for 'covalent bond'
δ^{+}N–Fδ^{-} **AND** δ^{-}N–Brδ^{+} ✓	1	**ALLOW** d+ / d- **DO NOT ALLOW** + / –

Past paper question

Molecules of BF_3 contain polar bonds, but the molecules are non-polar.

Suggest an explanation for this difference.

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.... **[2]**

Mark scheme

BF_3 is **symmetrical** ✓ The **dipoles** cancel out ✓	2	**IGNORE** 'polar bonds cancel' **IGNORE** 'charges cancel'

Who discovered the periodic table?

Mendeleev was the first to publish (in 1869) a version of the Table that we would recognise today …

…but does he deserve all the credit?

Other key figures include:

John Newlands (1 March 1865), Lothar Mayer (1868)
Alexandre-Emile Béguyer de Chancourtois (1862)

Structure of the periodic table

- In the modern periodic table elements are arranged in order of their atomic number.
- Each horizontal row is called a period.
- Each vertical column is called a group.

Periods and groups

<u>Periods (horizontal)</u>
Elements show gradual changes in properties across a period. These trends are repeated for each period. This is called PERIODICITY

<u>Groups (vertical)</u>
Elements in a group have similar chemical properties.

Past paper question

Fluorine will react violently with gallium to produce gallium fluoride.

Mendeleev originally called gallium 'eka-aluminium' as he predicted that gallium would have similar properties to aluminium.

(i) Complete the electron structure of the gallium atom.

$1s^2$ .... [1]

(ii) Use Mendeleev's prediction to suggest the empirical formula of gallium fluoride.

.... [1]

Mark scheme

(i)	$(1s^2)\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^1$ ✓	1	**IGNORE** $1s^2$ twice **ALLOW** $4s^2$ before $3d^{10}$ **ALLOW** '3D'
(ii)	GaF_3 ✓	1	

Past paper question

Chemists use the Periodic Table to predict the behaviour of elements.

(a) Early attempts at developing a Periodic Table arranged elements in order of increasing atomic mass.

(i) State which two elements from the **first twenty** elements of the modern Periodic Table are not arranged in order of increasing atomic mass.

.... **[1]**

(ii) Why does the modern Periodic Table **not** arrange some elements, such as those in **a(i)**, in order of increasing atomic mass?

....

....

.... **[1]**

Mark scheme

(a)	**(i)**	Potassium **AND** argon ✓	**1**	**ALLOW** K and Ar
	(ii)	They are arranged in increasing atomic number **OR** Neither would show properties **OR** trends of rest of group **OR** Neither would show properties **OR** trends of rest of period **OR** They are arranged by electron configuration ✓	**1**	**ALLOW** any correct property difference e.g. This would place a reactive metal in the same group as noble gases **ALLOW** they do not fit in with the rest of the group