Metallic lattice structures

1. [2015 / P1 / Q4]

The melting point of calcium, 839 °C, is much higher than the melting point of sodium, 98 °C. Which statement is most relevant in explaining this difference?

- **A** The calcium atom is larger than the sodium atom.
- **B** The calcium atom is heavier than the sodium atom.

C The calcium ion, Ca²⁺, has a higher charge than the sodium ion, Na⁺.

D The calcium ion, Ca^{2+} , contains more electrons than the sodium ion, Na^{+} .

The best answer is C. Option D should have the word "valence" electrons in order to be correct.

Examiners' comments:

This is one of the questions which candidates found challenging. Many candidates felt the number of electrons in the ions was the deciding factor rather than the charge on the ions.

AfL: Get students to suggest what is wrong with the option D.	· - 1
Hinge Question: What affects the strength of metallic bonding?	- į
 Ans: 1)Number of <u>valence electrons</u> contributed to the sea of delocalised electrons 2) Cationic radius 	_

Dot cross diagram, VSEPR theory, Shapes of molecules

2. N2016/3/5 (modified)

(a) Outline the principles of the Valence Shell Electron Pair Repulsion (VSEPR) Theory.

Central Atom's Valence Shell Electron Pairs (both bond pairs and lone pairs) arrange themselves in space to minimise repulsion and maximise separation.

Lone Pair-Lone Pair repulsion > Lone Pair-Bond Pair repulsion > Bond Pair-Bond Pair repulsion

(b) Although they have similar formulae, the following three compounds have differently shaped molecules.

(i) Predict and explain their shapes, using diagrams where appropriate.

	BF ₃	NF ₃	ClF ₃
Bond Pairs	3	3	3
Lone Pairs	0	1	2
Steric Number	3	4	5
Shape	Trigonal Planar	Trigonal Pyramidal	T-shaped
Diagram	F B F	F F	F . F

(ii) Label the partial charges, δ + and δ -, present on one B-F bond. State the cause of these partial charges. Hence, state, with a reason, whether or not BF₃ is a polar molecule.

δ+ -δ-

B – F

<u>F is more electronegative</u> than B and hence pull the shared electrons closer towards itself, giving rise to the dipole. BF_3 has no net dipole moment. Hence, there is no dipole moment.

<mark>AfL: Hinge</mark> Whi	e ques ch dot	tion (tutors may do this question before going through the terms of ter	Q3):
[Dot-cross diagram	Molecule
	1.	н <u>*</u> О <u>*</u> О <u>*</u> Ч	H2O2
	2.	· 0 × (e * 0: **	ClO3.
	3.	$\begin{bmatrix} Na \end{bmatrix}^{\dagger} \begin{bmatrix} x & x \\ x & y \\ x & x \end{bmatrix}^{-1}$	NaCl
	4.	Br š Še š Br	SeBr ₂
	5.	[× če × č:] -	ClO-

Reason for my answer is because (you may choose more than 1 answer):

2

- 1. Diagram 1 should have alternate dot cross on O.
- 2. Diagram 2 should have no dative bonds.

I

I

- 3. Diagram 3 have missing valence electrons on Na⁺.
- 4. Diagram 4 have missing valence electrons.
- 5. Diagram 5 should have O gaining the additional electron.

Ans: Incorrect diagram:1, 2,4, 5 Correct reason: 1, 2, 4, 5

3. <u>N2010/3/4c</u>

(a) Draw dot–and–cross diagrams to show the bonding in the molecules of NO₂ and O₃.
 Each molecule contains a dative covalent bond. In the NO₂ molecule, the central atom is nitrogen.

In each case you should distinguish carefully between electrons originating from the central atom and those from the two other most atoms. Include all lone pairs in your diagram.

- (b) Suggest a value for the bond angle in each of the above two molecules, giving reasons for your choice.
 - In O₃, there are <u>2 bond pairs and 1 lone pair</u> around the central O atom. Since <u>lone-pair bond pair repulsion is greater than bond pair bond pair repulsion</u>, the bond angle is about <u>112</u>° (ie. 109° < bond angle < 120°)
 - In NO₂, there is a <u>lone–electron and 2 bond pairs</u> around the central N atom. Thus the <u>lone–electron will repel the bond pair to a smaller extent than a lone pair</u>. Thus the bond angle in NO₂ is <u>134</u>°. (ie. 120° < bond angle < 180°)
 - Note: Given values are actual values for NO₂ and O₃. NO₂ (should have greater bond angle than O₃. Allow any value between that give for O₃ and 170°); O₃ (accept any value between 110–120°).

AfL: Hinge question According to VSEPR, Lp- lp repulsion > lp - bp repulsion > bp- bp repulsion How will the lone electron-bond pair repulsion compare with the rest? Ans: Lp- lp repulsion > lp - bp repulsion > bp- bp repulsion > lone e – bp repulsion

Examiners' comments:

• Estimating the bond angle in these two molecules is more difficult. As NO2 has an unpaired electron it is not so readily covered by the Valence Shell Electron Pair Repulsion theory, and therefore the logical approach is to begin with ozone. The central oxygen is surrounded by two bonds and a lone electron pair: the basic angle is therefore 120°, but allowing for the fact that a lone pair is closer to the oxygen than a bond pair, or is said to 'occupy a greater volume', this angle will be diminished to something less than 120°; a value between 110° and 120° was allowed. Moving to the NO₂ molecule, a single unpaired

electron will be less effective in diminishing the basic angle, and so the bond angle for NO_2 should be greaterthan that for O_3 . Examiners allowed any value between that given for O_3 and 170°

(c) The compound FO₂ does not exist but C/O₂ does. By considering the possible types of bonding in the two compounds, suggest reasons for this difference. (Assume that the halogen atom occupies a central position in each of these molecules.) [6]

A possible structure for XO_2 (where X = F or CI) is

$$O = X = O$$
 or $O = X \rightarrow O$,

in which there are more than 8 electrons around the central X atom in both structures. As <u>*Cl*</u> is in period 3 and has <u>energetically accessible d-orbitals</u> to <u>expand its octet structure</u>, C/O_2 exists. The above structures do not exist for FO₂ since fluorine lacks energetically accessible d-orbitals to expand its octet structure.

Another possible structure for FO_2 is $O \leftarrow F \rightarrow O$. However, while there are only <u>7 electrons</u> around the central F atom, fluorine is too electronegative to form dative bonds. Thus, the above structure does not exist too.

Examiners' comments:

Candidates were asked to suggest a reason why the compound CIO_2 exists, but FO₂ does not. Very many candidates pointed out that the CIO_2 molecule would involve chlorine expanding its octet of electrons, which it is able to do through the 3d orbital which is accessible for electron occupation. Very few candidates went further with the discussion. If fluorine cannot expand its octet, then the other alternative would be for the fluorine to provide two dative bonds: it does not do this because it is too electronegative for dative bonding.



(i) Draw clear diagrams of these three molecular arrangements, each showing the five pairs of electrons. State which arrangement, if any, would result in a molecule with no dipole moment.
 [2]



Sigma and pi bonds

4. Which one of the compounds is ionic and contains both sigma and pi bonds?

A $Fe(OH)_3$ **B** HC/O **C** H_2S **D** NaCN

A and D is ionic. However, only $C \equiv N^{-}$ contains both sigma and pi bonds.

Dative Bonding

5. N2016/31/1 (modified)

(a) Carbon monoxide reacts with boron hydride. BH₃, at high pressure to give the compound H₃BCO, in which carbon is bonded to both boron and oxygen. Draw a dot-and-cross diagram to show the bonding in H₃BCO, clearly indicating any dative bonds it contains.



(b) Explain why these two molecules form a product when they react in the molar ratio 1:1.

BH₃

B in BH_3 contains <u>one empty orbital</u> and <u>C</u> in CO contains 1 lone pair of electrons. Hence, the C atom donates its lone pair electrons to B to form <u>dative/coordinate bonding</u>.

(c) By considering the numbers of bonding and non-bonding electron pairs, draw diagrams to show the likely shapes of carbon monoxide, boron hydride and the product. In your diagrams, clearly state the values of the bond angles.

СО

Product



(d) Suggest the structural formulae of possible similar products from

(i) beryllium difluoride and carbon monoxide



Physical Properties relating to structure: Melting and Boiling Points

6. The following table lists the boiling points of some organic compounds.

Compound	Formula	M _r	Boiling point / °C
A	CH ₃ CH ₂ CH ₂ CH ₂ OH	74.0	118
В	$CH_3CH_2CH_2CH_2CH_3$	72.0	36
С	CH ₃ -C-N C ₂ H ₅	191	130

With reference to the type and extent of relevant intermolecular forces, explain the differences in boiling points

- (i) between **A** and **B**, and
- (ii) between **A** and **C**.

(i) Between A & B:

A is a polar simple covalent molecule held by hydrogen bonding.

B is a non-polar **simple covalent molecule** held by **weak instantaneous dipole-induced dipole** (id-id) attractions.

Less energy is needed to overcome the weaker id-id attractions in B.

Hence, **B** has lower boiling point.

(ii) Between A & C:

A is a polar simple covalent molecule held by hydrogen bonding.

C is a polar **simple covalent molecule** held by **permanent dipole- permanent dipole (pd-pd)** attractions.

However, **C** has <u>more electrons</u> and forms <u>stronger id–id attractions</u> than **A**.

<u>More energy</u> is needed to overcome the stronger id-id attractions of attraction in **C**. Therefore **C** has higher boiling point.

7. Which statement describes a phenomenon which cannot be explained by hydrogen bonding?

- A Ice has a lower density than water at 0°C.
- B 2-nitrophenol is more volatile than 4-nitrophenol.



- C Ethanoic acid, CH_3COOH , has a higher melting point than expected.
- D The boiling point of pentanol, $CH_3CH_2CH_2CH_2CH_2OH$, is higher than 2-methylbutanol, $CH_3CH(CH_3)CH_2CH_2OH$.
- (A) In ice, each water molecule is surrounded by 4 neighbouring molecules in a tetrahedral cluster through hydrogen bonding. As water freezes to ice, the water molecules have less kinetic energy to overcome the intermolecular hydrogen bonds. This creates a regular lattice structure of a tetrahedral arrangement of the water molecules which results in a fairly open structure on ice. This accounts for why ice has a lower density than water.
- (B) Due to the proximity of the 2 functional groups (–OH and –NO₂) in 2-nitrophenol, it forms intramolecular hydrogen bonding. Hence, it has less extensive intermolecular hydrogen bonding and as a result more volatile than 4-nitrophenol.



(C) Ethanoic acids forms a dimer, as a result it is bonded by two hydrogen bonds, which results in a higher melting point.

(D) Both pentanol and 2-methylbutanol are polar simple covalent molecules held together by hydrogen bonds. However, 2-methylbutanol is a branched molecule hence its instantaneous dipole-induced dipole attractions is weaker than pentanol. Therefore having a lower boiling point than pentanol.



- Teacher can ask if students understand the term volatility and what happens when a covalent compound boils (what interactions/bonding is broken).
(Volatility $\alpha \frac{1}{Boiling point}$ and stress that the covalent bonds are not broken during melting and
boiling; only the intermolecular interactions are broken)
Teacher can go on to ask what happens when covalent bonds in a molecule are broken
Ans: molecule undergoes decomposition e.g 2 $H_2O_2 \rightarrow 2 H_2O + O_2$
Teacher can get students to draw out the intramolecular hydrogen bonding in B and the intermolecular hydrogen bonding in C:

8. Explain the following observations.

(a) Name the structure of solid iodine and solid sodium iodide. Explain why solid iodine turns into a vapour when it is heated gently, but solid sodium iodide does not.

Iodine – Simple covalent molecule

Sodium iodide – Giant ionic lattice

The <u>weak instantaneous dipole – induced dipole interactions</u> between iodine molecules require <u>little</u> <u>amount of energy</u> to overcome as compared to the <u>strong electrostatic forces of attraction between</u> <u>oppositely charged ions</u> in sodium iodide. Hence, solid iodine vapourises when heated gently while solid soldium iodide does not.

(b) Explain why $BeCl_2$ is a simple covalent molecule.

Be²⁺ has a high charge density and the <u>CI⁻ anion is large.</u> Hence, CI⁻ anion is easily polarised by

the Be²⁺ cation. Hence BeC l_2 is a simple covalent molecule.

	AfL: Hinge question Explain why $AICI_3$ is a covalent compound and not ionic.	
l	Ans: Al^{3+} has a <u>high charge density</u> and the <u>Cl⁻ anion is large</u> . Hence, Cl^- anion is <u>easily</u> <u>polarised</u> by the Al^{3+} cation. Hence $AlCl_3$ is a covalent compound.	j

(c) State the type of bonding and describe the lattice structure of the crystalline solids sodium chloride, iodine, graphite and copper. You may include labelled diagrams in your answer.

NaCI – <u>Ionic Bonding & Giant ionic lattice structure</u> with <u>strong electrostatic forces of attraction</u> <u>between oppositely charged ions</u>. lodine – <u>Covalent bonding</u> & <u>Simple covalent molecule</u> with <u>instantaneous dipole – induced dipole</u> <u>interactions</u> of attraction between molecules.

Graphite – <u>Covalent bonding</u> & <u>Giant molecular structure</u> with <u>extensive network of covalent bonds</u> <u>between atoms in a layer</u> <u>and</u> weak <u>instantaneous dipole-induced dipole attraction between layers</u>.

Copper – <u>Metallic bonding & Giant metallic lattice structure</u> with <u>regular arrangement of copper ions</u> <u>surrounded by delocalized electrons</u>.

AfL: Teacher can do a comparison of table between diamond and graphite

	Diamond	Graphite
Shape around each C atom (shape)	tetrahedral	Trigonal planar
Type of bonding	Covalent bonds	Covalent bonds within each layer; weak id-id interactions between between the layers

11





O C atom

9. [N2016/P3/4]

Explain the following comparison in terms of the types of intermolecular forces and how they arise. Use diagrams when appropriate to illustrate your answer.

- The boiling point of ethane, C_2H_6 , (-89 °C) is higher than that of methane, CH_4 , (-164 °C).
- The boiling point of methanol, CH₃OH, (65 °C) is higher than that of ethane.

The boiling point of ethane, C_2H_6 , (-89 °C) is higher than that of methane, CH_4 , (-164 °C).

Both C₂H₆ and CH₄ molecules are held together by <u>instantaneous dipole-induced dipole</u> <u>attractions (id-id)</u>. <u>Temporary fluctuations in the electron cloud distribution of non-polar</u> <u>molecules give rise to an instantaneous dipole which induces neighbouring electron clouds</u> to be another dipole. Hence, id-id occurs due to <u>electrostatic attraction between the δ+ end of</u>

one dipole to the **o**- of the next.



 C_2H_6 molecules have <u>stronger intermolecular id-id</u> because it has <u>more electrons</u>, hence a more <u>polarisable electron cloud</u>. <u>More energy</u> is required to break these id-id, thus C_2H_6 has a higher b.p.

The boiling point of methanol, CH₃OH, (65 °C) is higher than that of ethane.

CH₃OH molecules are **predominantly held together by hydrogen bonds**, **which arises due to H bonded to the highly-electronegative atom O**. This makes the **H very electropositive**, δ +,

resulting in hydrogen bonds when the 5+ H is attracted to the lone pair of electrons found on

another O of methanol.

The hydrogen bonds in CH_3OH are <u>much stronger</u> than id-id in ethane. Hence, it requires <u>more</u> <u>energy</u> to break them. Thus CH_3OH has a higher b.p. than C_2H_6 .



Criteria of success: hydrogen bond:	•••
δ+ on H, and δ- on	:
electronegative atoms	
Lone pair of e ⁻ on F, O or N	:
Dotted line and label to	:
represent hydrogen bond	:

Physical Properties relating to structure: Solubility

10. Explain the following statements. Use diagrams when appropriate to illustrate your answer.

- Ethanol, CH_3CH_2OH , is much more soluble in water than in benzene, C_6H_6 .
- Cyclohexane, C_6H_{12} , is more soluble in benzene than in water.

(a) **Energy released from forming hydrogen bonds** between ethanol and water is **sufficient** to overcome the **hydrogen bonds** between ethanol molecules and **hydrogen bonds** between water molecules.

However, **<u>energy released from forming instantaneous dipole-induced dipole (id-id)</u> <u>attractions</u> between ethanol and benzene is <u>insufficient</u> to overcome the <u>hydrogen bonds</u> between ethanol molecules and the <u>id-id attractions</u> between benzene molecules.**

(b) **Energy released from forming id-id attractions** between cyclohexane and benzene is **sufficient** to overcome the **id-id attractions** between cyclohexane molecules and **id-id attractions** between benzene molecules.

However, **<u>energy released from forming id-id attractions</u>** between cyclohexane and water is **<u>insufficient</u>** to overcome the <u>**hydrogen bonds**</u> between water molecules and <u>**id-id attractions**</u> between cyclohexane molecules.

11. When a sample of solid sodium chlorite, NaC/O₂, is dissolved in water, the cations and anions are each surrounded by water molecules. Explain with the aid of suitable labelled diagrams why sodium chlorite is soluble in water.



Energy released from forming ion-dipole attractions between the ions and water is **sufficient** to overcome the **ionic bonds** between the ions and **hydrogen bonds** between water molecules.

Physical Properties relating to structure: Liquefaction

12. The gases butane, 2-methylpropane and propane are kept in a sealed container at atmosphere pressure and room temperature.

Which row represents the order in which these gases will liquefy as the pressure in the container is gradually increased?

		Increasing pressure	•
A	Butane	2-methylpropane	Propane
В	2-methylpropane	Butane	Propane
С	Propane	Butane	2-methylpropane
D	Propane	2-methylpropane	Butane

Butane and 2-methylpropane has stronger instantaneous dipole-induced dipole (id-id) attractions than propane due to the larger number of electrons. Butane has more extensive id-id attractions than 2-methylpropane due to butane having a larger surface area than 2-methylpropane. **The stronger the attractions, the more the compound will liquefy.**



Giant Covalent Molecule

13. Which statement about the properties of graphite and diamond is incorrect?

- 1. Graphite can be used as a lubricant because of the weak forces of attraction between layers of atoms.
- 2. Graphite is a good conductor of electricity in the direction parallel to each layer.

- 3. In graphite, the carbon to carbon distance between different layers is the same as that within each plane.
- 4. All covalent bonds in diamond are of the same length but those in graphite are not.
- 5. The C-C-C bond angle is larger in diamond than in graphite.

(A) 1, 2 only (B) 3, 4, 5 only (C) 4, 5 only (D) 2, 4, 5 only

- (1) Correct. Graphite has weak instantaneous dipole-induced dipole attractions between the layers.
- (2) Correct. Each C in graphite has an unhybridised p orbital which contains a single electron that will overlap sideways with the p-orbitals of its neighbouring carbon atoms which results in an extended π -electron cloud above and below the plane of carbon. Hence, these π -electrons are delocalised within each layer and graphite can conduct electricity in a direction parallel to the layers.
- (3) Incorrect. In graphite, the carbon to carbon distance between different layers is longer than that within each plane.
- (4) Incorrect. All covalent bonds in diamond and graphite are of the same length.
- (5) Incorrect. The C-C-C bond angle in diamond is 109.5° but the C-C-C bond angle in graphite is 120°.

Criteria of Success Checklist
Tick if you have
Dot-and-cross diagrams
Used only dot and cross shapes to represent electrons.
\square For ionic compound, ensure that the charges on the ions are balanced
For covalent molecules or polyatomic ions:
1. Alternate dot and cross is used when more than 2 types of atoms are
present or when there is more than 1 centre atoms present.
2. Central atom has no more than 8 electrons if it is from Period 2

3. Dative bonds are used when atoms cannot	expand octet or	when atoms
have full valance shell but needs to be bor	ided to another	atom
4. Indicate the charge of the species (if any)	
Shapes and bond angles		
Number of bond pairs and lone pairs of electrons is determined	d before assigning s	shape based on
VSEPR theory		
Lewis structure or shape of molecules/charged species are dro	awn reflecting the s	hape as set out
by VSEPR theory		
Bond angles are labelled between bond pairs of electrons		
Explaining dative bond formed between 2 species		
Acceptor atom is identified and number of empty orbitals/ number orbitals/ number of empty orbitals/ number orbitals/ number of empty orbitals/ number of empty orbitals/ number orbitals/ number orbita	nber of electrons i [.]	t is short of is
stated		
Donor atom with lone pair of electron is identified		
Number of dative bond formed is determined and is used to ex	xplain the molar rat	io in which the
species reacted		
Explaining physical properties such as m.p and b.p		
SIA (structure, interactions, amount of energy) format is used	1	
interaction	action and the solut	e-solvent
Compared the strengths of interaction		
Concluded solubility by considering energy terms		
Diagram to illustrate hydrogen bonding		
$\Box~\delta$ + on H, and δ - on electronegative atoms		
Lone pair on N, O or F		
Dotted line to represent hydrogen bond and label the hydroge	n bond	
Assessment Objectives	Self-check Qn	Discussion Qn

(c)instruct and describe: (i)instruct and describe: (ii)instruct and describe: (iii)instruct and describe: (iii) (iii)instruct and describe: (iii) (iii)instruct and describe: (iii) (iii)instruct and describe: (iii) (iii) (co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the A_LC_L moleculeinstruct and (iii) (iiii) (iii) (iiii) (iii) (iii) (ii
i) ionic bond as the electrostatic attraction between oppositely charged ions; ii) covalent bond as the electrostatic attraction between a shared pair of electrons and positively charged nuclei, (iii) metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons 4, 3 (b) describe, including the use of 'dot-and-cross' diagrams: 5 4, 3 (i) ionic bonding as in sodium chloride and magnesium oxide 5 4, 3 (ii) covalent bonding as in hydrogen; oxygen; nitrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene 5 4, 3 (iii) co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the AJ_CL_ molecule 4 5 (c) describe covalent bonding in terms of orbital overlap (limited to 4, 5 5 6, 7 4, 3, 2 (d) explain the shapes of, and bond angles in, molecules such as BF ₃ (trigonal planar); CO ₂ (linear); CH ₄ (terahedral); NH ₅ (trigonal planar); CO ₂ (linear); CH ₂ (terahedral) by using the Valence Shell Electron Pair Repulsion theory 5 4, 3 (e) predict the shapes of, and bond angles in, molecules such as analogous to those specified in (d) 6 7 2 (g) deduce the polarity on theory (e) predict the shapes of, and bond angles in, molecules 5 4, 3 2 (g) deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d));
(i)Data the original production of the constraint of the c
Opposition of a software electrostatic attraction between a shared pair of electrons and positively charged nuclei, (iii) metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons4, 3(b) describe, including the use of 'dot-and-cross' diagrams: (i) ionic bonding as in sodium chloride and magnesium oxide (ii) covalent bonding as in hydrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene54, 3(c) describe covalent bonding in terms of orbital overlap (limited to the ammonium ion and in the AL/CL_molecule4,5(c) describe covalent bonding in terms of orbital overlap (limited to entheme (limited to Organic Chemistry)5, 6, 74, 3, 2(d) explain the shapes of, and bond angles in, molecules such as BF ₅ (trigonal planar); CO ₂ (linear); CH ₄ (tetrahedral); NH ₃ (trigonal planar); CO ₂ (linear); CH ₄ (tetrahedral); NH (f) explain and deduce bond polarity using the concept of electronegativity [quantitative treatment of electro-negativity is not required]6, 72(g) deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d));28, 7,9(h) describe the following forces of attraction (electrostatic in nature); (i) intermolecular forces, based on permanent and induced dipoles, as in liquid and gaseous CHCl ₃ , Br ₂ and the noble gases (ii) hydrogen bonding, using ammonia and water as examples of molecules containing –NH and –OH groups12(i) outline the importance of intermolecular forces to the liquefaction of gases when subjected to high pressure and / or low temperature.12
(ii)Evaluation of the electrostatic statusStatus(iii)metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons(b)describe, including the use of 'dot-and-cross' diagrams: (i)54, 3(i)ionic bonding as in sodium chloride and magnesium oxide (iii)54, 3(ii)covalent bonding as in hydrogen; oxygen; nitrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene54, 3(iii)co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the AkCh, molecule45(c)describe covalent bonding in terms of orbital overlap (limited to s and p orbitals only), giving σ and $\pi \pi$ bonds (see also topic on Introduction to Organic Chemistry)5, 6, 74, 3, 2(d)explain the shapes of, and bond angles in, molecules such as BF3, (trigonal planar); CO_ (linear); CH4, (tetrahedral); NH3 (trigonal pyramidal); H_2O (bent); SF6 (octahedral) by using the Valence Shell Electron Pair Repulsion theory54, 3(e)predict the shapes of, and bond angles in, molecules analogous to those specified in (d)54, 3(f)explain and deduce bond polarity using the concept of electronegativity [quantitative treatment of electro-negativity is not required]2(g)deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d));8, 7,9(h)describe the following forces of attraction (electrostatic in nature);8, 7,9(i)intermolecular forces, based on permanent and induced dipoles, as
Strate pair of electrons and delocalised electrons(iii) metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons5(b) describe, including the use of 'dot-and-cross' diagrams: (i) ionic bonding as in sodium chloride and magnesium oxide (ii) covalent bonding as in hydrogen; oxygen; nitrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene (iii) co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the Al_5Cl_{h} molecule4, 3(c) describe covalent bonding in terms of orbital overlap (limited to s and p orbitals only), giving σ and $\pi \pi$ bonds (see also topic on Introduction to Organic Chemistry)5, 6, 74, 3, 2(d) explain the shapes of, and bond angles in, molecules such as BF ₃ (trigonal planar); CO ₂ (linear); CH ₄ (tetrahedral); NH ₃ (trigonal planar); CO ₂ (linear); CH ₄ (tetrahedral) by using the Valence Shell Electron Pair Repulsion theory54, 3(e) predict the shapes of, and bond angles in, molecules analogous to those specified in (d)6, 72(f) explain and deduce bond polarity using the concept of electronegativity [quantitative treatment of electro-negativity is not required]8, 7,9(g) deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d));8, 7,9(h) describe the following forces of attraction (electrostatic in nature): (i) intermolecular forces, based on permanent and induced dipoles, as in liquid and gaseous CHCl ₃ , Br ₂ and the noble gases8, 7,9(ii) hydrogen bonding, using ammonia and water as examples of molecules containing –NH and –OH groups12(i) outlin
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temperature.
(j) outline the importance of hydrogen bonding to the physical 7
properties of substances, including ice and water
(k) explain the terms bond energy and bond length for covalent
bonds
(I) compare the reactivities of covalent bonds in terms of bond
energy, bond length and bond polarity
(m) describe, in simple terms, the lattice structure of a crystalline 1, 2 13, 8
solid which is:
(i) ionic, as in sodium chloride, magnesium oxide
(ii) simple molecular, as in iodine
(iii) giant molecular, as in graphite: diamond
(iv) hydrogen-bonded, as in ice
(v) metallic, as in copper
[the concept of the 'unit cell' is not required]
(n) describe interpret and/or predict the effect of different types of 1.3 1.11.10.6
structure and bonding on the physical properties of substances
(o) suggest the type of structure and bonding present in a 8.9.10 8
substance from given information