Electrons, Bonding & Structure
AS & A Level

Model Answers 2

Level	A Level
Subject	Chemistry
Exam Board	OCR
Module	Foundations in Chemistry
Topic	Electrons, Bonding & Structure
Paper	AS & A Level
Booklet	Model Answers 2

Time allowed: **55 minutes**

Score: **/41**

Percentage: **/100**

Grade Boundaries:

A*	A	B	C	D	E
>85%	73%	60%	47%	34%	21%

Question 1

Solids exist as lattice structures.

(a) Giant metallic lattices conduct electricity. Giant ionic lattices do not. If a giant ionic lattice is melted, the molten ionic compound will conduct electricity.

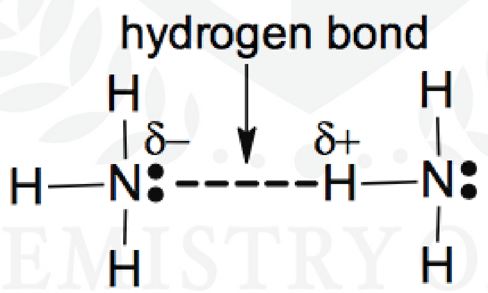
Explain these observations in terms of bonding, structure and particles present. [3]

- Metallic lattice has delocalised electrons, which can act as charge carriers.
- An ionic lattice is fixed when solid, so there are no mobile ions and no charge can flow.
- However, when molten, the ions are no longer in a fixed structure and hence mobile and able to carry charge.

(b) The solid lattice structure of ammonia, NH_3, contains hydrogen bonds.

(i) Draw a diagram to show hydrogen bonding between **two** molecules of NH_3 in a solid lattice.

Include relevant dipoles and lone pairs. [2]



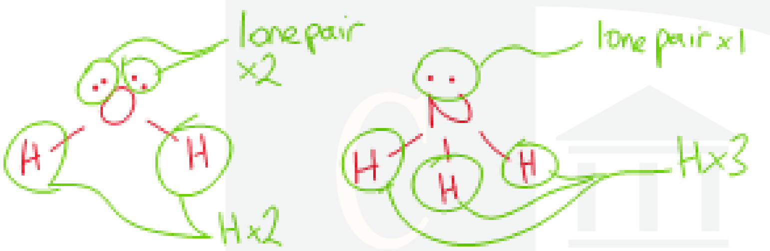
- Nitrogen is more electronegative so will attract electron density and therefore be δ- and the hydrogen will be δ+
- Hydrogen bonds are at 180°

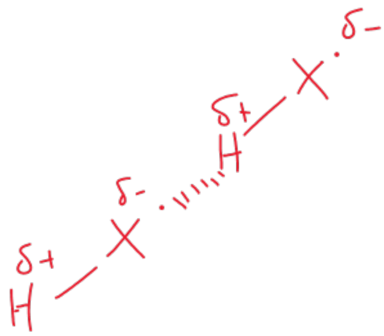
Two or more ammonia molecules with at least one $H^{\delta+}$ and at least one $N^{\delta-}$ ✓

H-bond between H in one ammonia and lone pair of N in another ammonia molecule ✓

(ii) Suggest why ice has a higher melting point than solid ammonia. [2]

- Ice has stronger hydrogen bonds than solid ammonia
- The oxygen in water has two lone pairs and the nitrogen in ammonia has only one





To form a hydrogen bond- two things must be present:

1. Hydrogen bonded to an electronegative element
2. A lone pair of electrons on the electronegative element

Since water has **two** hydrogens and **two** lone pairs of electrons on the oxygen, **two** hydrogen bonds can form.

Ammonia has **three** hydrogens bonded to nitrogen, but nitrogen only has **one** lone pair so is only able to form **one** hydrogen bond.

(c) Solid SiO_2 melts at 2230 °C. Solid $SiCl_4$ melts at –70 °C. Neither of the liquids formed conducts electricity.

Suggest the type of lattice structure in solid SiO_2 and in solid $SiCl_4$ and explain the difference in melting points in terms of **bonding** and **structure**.

In your answer you should use appropriate technical terms, spelled correctly. [5]

TIP: When a question asks you to explain something in terms of bonding and structure, always check to see that you have:

1) Described the bonding in all substances
2) Described the structure of all substances
3) Compare how they affect the physical property in the question (e.g. melting point)

SiO_2 must have a giant covalent lattice as the elements are both non-metals and the melting point is very high.

$SiCl_4$ must be simple molecular as the elements are both non-metals and the melting point is very low.

Neither are ionic as they do not conduct electricity when molten

Bonding

SiO_2 – covalent

$SiCl_4$ – covalent

Structure

SiO_2 – giant lattice

$SiCl_4$ – simple molecular

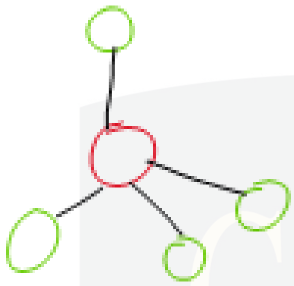
Melting points

SiO_2 – strong covalent bonds must be broken to separate particles (atoms)

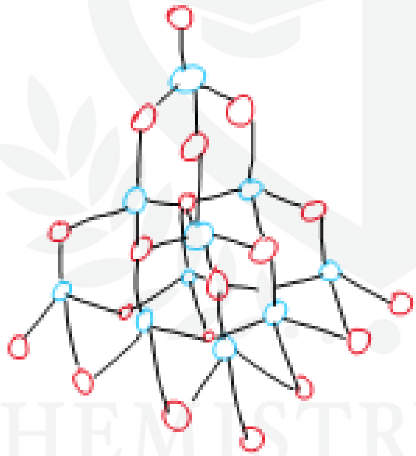
$SiCl_4$ – weak van der Waals' forces must be broken to separate particles (molecules)

More energy is needed to break forces in SiO_2 than $SiCl_4$

$SiCl_4$ molecular structure (similar to methane, CH_4)



SiO_2 giant covalent lattice structure



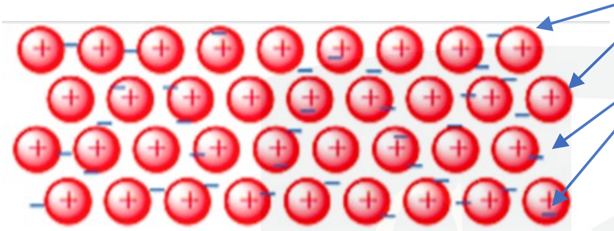
[Total: 12 Marks]

Question 2

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*. [3]



Regular arrangement of positive ions

Electrons are scattered between the positive ions. These are delocalised and can move freely throughout the whole structure

Metallic bonding = electrostatic **attraction** between the **electrons** and the **positive ions**

(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]

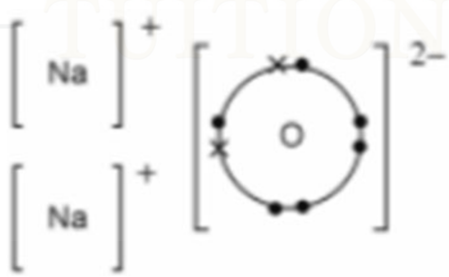
$4Na + O_2 \rightarrow 2Na_2O$

(ii) State what is meant by the term *ionic bond*. [1]

Electrostatic **attraction** between **oppositely charged ions**

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

Show **outer** electrons only. [2]



(c) Compare and explain the electrical conductivities of sodium and sodium oxide in the solid and liquid states. [5]

- **Na** conducts electricity in both solid and liquid states. This is because Na is a metal and it has delocalised electrons which can move freely throughout the structure and carry charge.
- **Na_2O** cannot conduct electricity when solid because the Na^+ and O^{2-} ions are fixed in place and cannot move.
- **Na_2O** conducts electricity when liquid because the ions are free to move and carry charge.

[Total: 12 Marks]

Question 3

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*? [2]

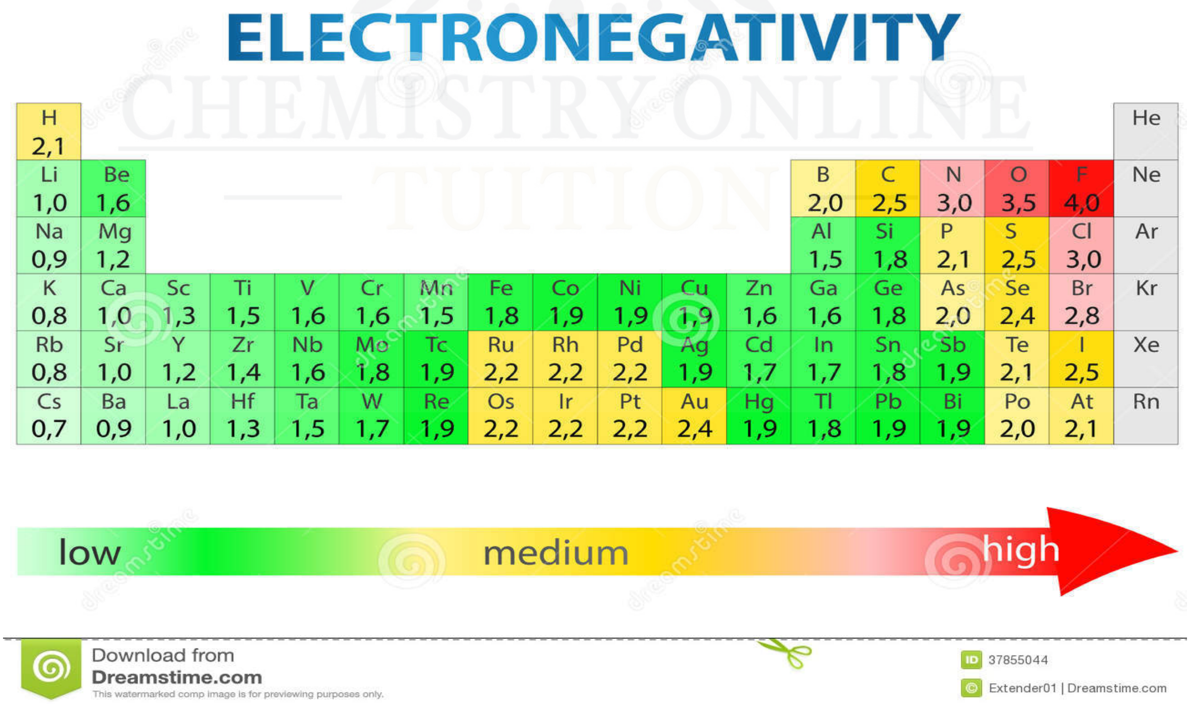
The ability of an atom to attract electrons ✓

...in a covalent bond ✓

(b) Show, using δ+ and δ– symbols, the permanent dipoles on each of the following bonds. [1]

N—F N—Br

Electronegativity values



H 2,1																	He
Li 1,0	Be 1,6											B 2,0	C 2,5	N 3,0	O 3,5	F 4,0	Ne
Na 0,9	Mg 1,2											Al 1,5	Si 1,8	P 2,1	S 2,5	Cl 3,0	Ar
K 0,8	Ca 1,0	Sc 1,3	Ti 1,5	V 1,6	Cr 1,6	Mn 1,5	Fe 1,8	Co 1,9	Ni 1,9	Cu 1,9	Zn 1,6	Ga 1,6	Ge 1,8	As 2,0	Se 2,4	Br 2,8	Kr
Rb 0,8	Sr 1,0	Y 1,2	Zr 1,4	Nb 1,6	Mo 1,8	Tc 1,9	Ru 2,2	Rh 2,2	Pd 2,2	Ag 1,9	Cd 1,7	In 1,7	Sn 1,8	Sb 1,9	Te 2,1	I 2,5	Xe
Cs 0,7	Ba 0,9	La 1,0	Hf 1,3	Ta 1,5	W 1,7	Re 1,9	Os 2,2	Ir 2,2	Pt 2,2	Au 2,4	Hg 1,9	Tl 1,8	Pb 1,9	Bi 1,9	Po 2,0	At 2,1	Rn

Electronegativity increases from left to right and from bottom to top of the periodic table.

Fluorine is the most electronegative element, followed by oxygen and then nitrogen and chlorine.

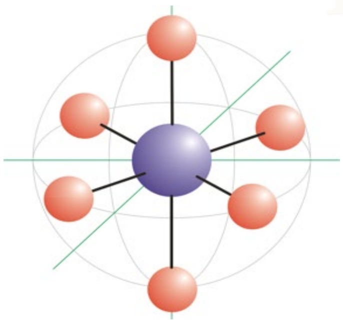
The more electronegative atom 'pulls' the electron density towards itself, creating a partially negative (δ^-) end to the molecule.

(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]

Octahedral

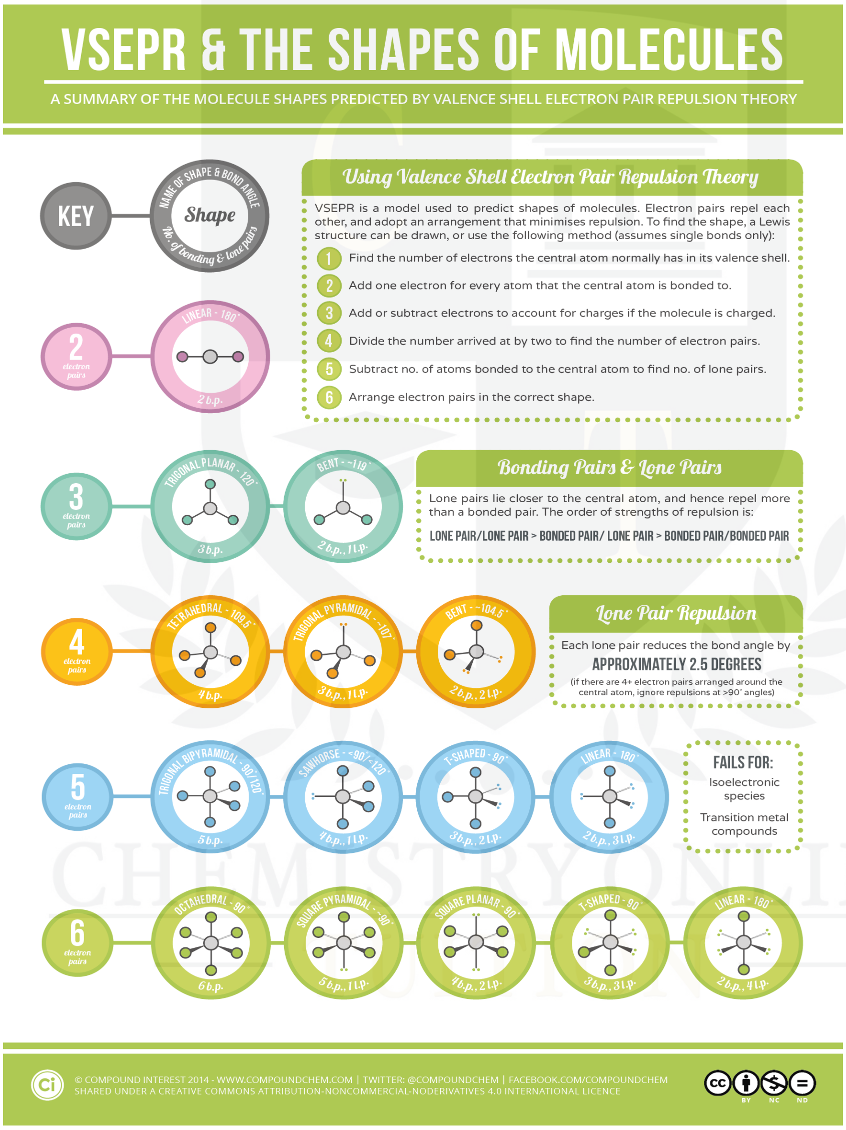
It has six bonding pairs (and no lone pairs) arranged around the sulfur. To minimise repulsion, this results in octahedral geometry:



TIP: Learn the different combinations of bonding pairs to lone pairs, the resulting shape of the molecules produced, and the bond angles involved.

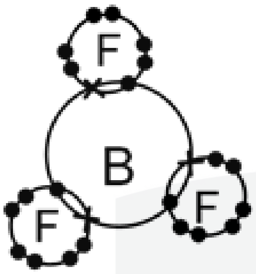
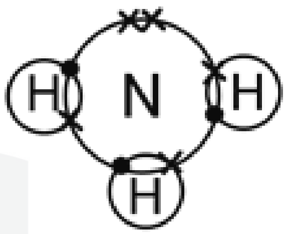
A poster, such as the one below, will help you do this.

http://www.compoundchem.com/2014/11/13/vsepr/



(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°. [5]

	
BF_3	**NH_3**

- Boron has 3 electrons, all of which bond covalently: shown by three 'dot and cross bonds' between B and F atoms.
- Fluorine atoms have 7 electrons, and can use one of these to form a single covalent bond, leaving 6 non-bonding electrons: shown by 6 of the same symbol (dot or cross) arranged in 3 lone pairs ✓
- N has 5 outer electrons and uses 3 of these to form covalent bonds (reaching a complete octet of electrons): shown by three 'dot and cross bonds' between N and H atoms and one lone pair of electrons on the N ✓
- Each H atom shares its single electron: shown by having no lone pairs of electrons

Explanation

Electron pairs repel one another and minimize repulsion from each other by their geometric arrangement.

In BF_3 there are three electron pairs, all involved in bonding. These repel each other equally to give a bond angle of 120°.

In NH_3 there are three bonding pairs and an additional lone pair which repels more than bonding pairs.

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

Suggest an explanation for this difference. [2]

BF_3 is symmetrical, therefore, the three dipoles cancel each other out in all directions to leave no net dipole.

[Total: 11 Marks]

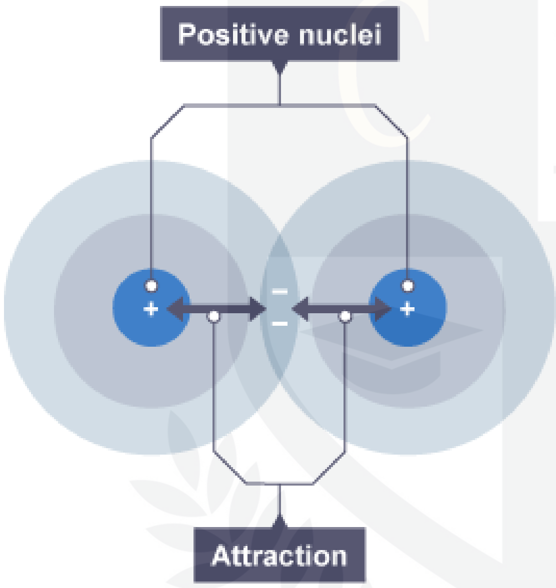
Question 4

Simple molecules are covalently bonded.

(a) State what is meant by the term *covalent bond*. [1]

A **covalent bond** is a chemical **bond** that involves **a shared pair of electrons** between the nuclei of two atoms.

A diagram to show the attraction involved in a covalent bond:

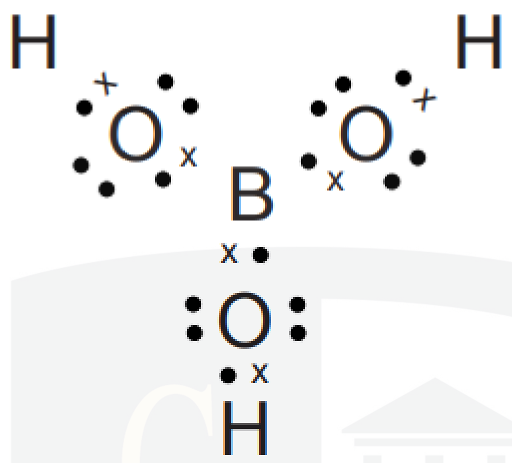


(b) Chemists are able to predict the shape of a simple covalent molecule from the number of electron pairs surrounding the central atom.

(i) Explain how this enables chemists to predict the shape. [2]

Pairs of electrons surrounding a central atom repel. The shape is determined by the number of bond pairs and number of lone pairs of electrons. Lone pair-lone pair repulsion > lone pair-bond pair repulsion > bond pair-bond pair repulsion.

(ii) The '*dot-and-cross*' diagram of the simple covalent molecule, H_3BO_3, is shown below. [2]



Predict the O–B–O and B–O–H bond angles in a molecule of H_3BO_3.

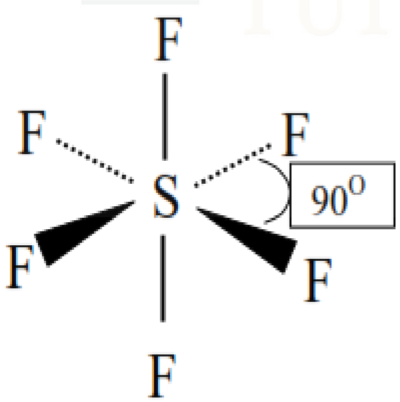
B in H_3BO_3 has 3 bonding electron pairs, and 0 lone pair, so the shape of the molecule is trigonal planar. Therefore the bond angles of O-B-O is **120º**

O in H_3BO_3 has 2 bonding electron pairs, and 2 lone pairs, so the shape of the molecule is bent. Therefore the bond angles of O-B-O is **104.5º**

(c) Give an example of a simple covalent molecule which has all bond angles equal to 90°. [1]

SF_6 (or any molecule that has 6 bonding pairs and no lone pairs therefore forming an octahedral shape)

Diagram to show the shape and the bond angles in SF_6:



[Total 6 Marks]