

1 Describe how positive and negative ions are formed.

Positive ions – atoms lose electron(s)

Negative ions – atoms gain electron(s)

2 Give the formulae of the ions formed by the following elements: Li Ba Al Br O N

Li^+ Ba^{2+} Al^{3+} Br^- O^{2-} N^{3-}

3 State the formulae (with charges) of the following ions: nitrate sulfate hydroxide ammonium carbonate hydrogencarbonate phosphate(V)

Nitrate	Sulfate	Hydroxide	Ammonium
NO_3^-	SO_4^{2-}	OH^-	NH_4^+
Carbonate	Hydrogencarbonate	Phosphate	
CO_3^{2-}	HCO_3^-	PO_4^{3-}	

4 Work out the formulae of the following ionic compounds:

Lithium fluoride	LiF	Sodium hydrogencarbonate	NaHCO_3	Iron(II) phosphate	$\text{Fe}_3(\text{PO}_4)_2$
magnesium chloride	MgCl_2	Calcium nitrate	$\text{Ca}(\text{NO}_3)_2$	Sodium carbonate	Na_2CO_3
ammonium sulfate	$(\text{NH}_4)_2\text{SO}_4$	Barium hydroxide	$\text{Ba}(\text{OH})_2$	Copper(II) nitrate	$\text{Cu}(\text{NO}_3)_2$

5 Classify each of the following as having mainly ionic or mainly covalent bonding:

CO_2 CaS PCl_3 OF_2 MgO

mainly ionic	mainly covalent
CaS MgO	CO_2 PCl_3 OF_2

6 Explain what is meant by *ionic bonding*.

Strong electrostatic forces of attraction between oppositely-charged ions

7 Describe the structure of ionic compounds

Giant lattice structure of alternating positive and negative ions

8 Explain why ionic compounds

- have high melting point,
- are non-volatile
- are often soluble in water
- do not conduct electricity when solid
- conduct electricity when molten or in aqueous solution

high melting point	a lot of energy is required to overcome strong electrostatic forces of attraction in the giant lattice structure
Non-volatile	a lot of energy is required to overcome strong electrostatic forces of attraction in the giant lattice structure
often soluble in water	Water is a polar solvent – ion-dipole forces between water molecules and ions pay back the energy to break apart the lattice
do not conduct electricity when solid	Ions held tightly in place in giant lattice structure – not free to move
conduct electricity when molten or in aqueous solution	Ions free to move

9 Explain what is meant by a *covalent bond*

Electrostatic attraction between a shared pair of electrons and the nuclei of both atoms making up the bond.

10 Draw Lewis structures for:

O ₂	C ₂ H ₂	SO ₄ ²⁻
HCN	SO ₂	N ₂ H ₄
CO ₂	NH ₄ ⁺	ClF ₃
C ₂ H ₄	CO ₃ ²⁻	PCl ₅

	$\text{H}-\text{C}\equiv\text{C}-\text{H}$	
$\text{H}-\text{C}\equiv\text{N} $		

11 State and explain the differences in bond length and bond energy in

- C₂H₆, C₂H₄ and C₂H₂
- N₂, N₂H₄ and N₂H₂
-

C₂H₆ (C-C single bond – longest and weakest)

C₂H₄ (C=C double bond shorter than C-C and stronger than C-C)

C₂H₂ (C≡C triple bond – shortest and strongest – more shared electrons – greatest attraction between nuclei and shared electrons)

N₂H₄ (N-N single bond – longest and weakest)

N₂H₂ (N=N double bond shorter than N-N and stronger than N-N)

N₂ (N≡N triple bond – shortest and strongest)

12 Explain in which of the following the central atom has an expanded octet:

BF₃

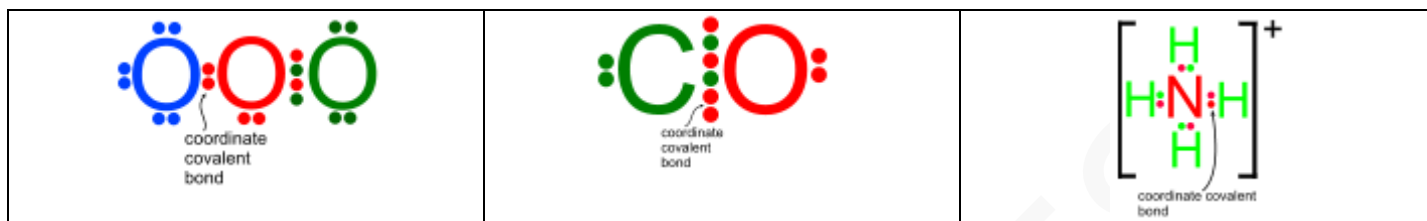
BrF₅ (12 e⁻ in outer shell of Br)

SO₃ (12 e⁻ in outer shell of S)

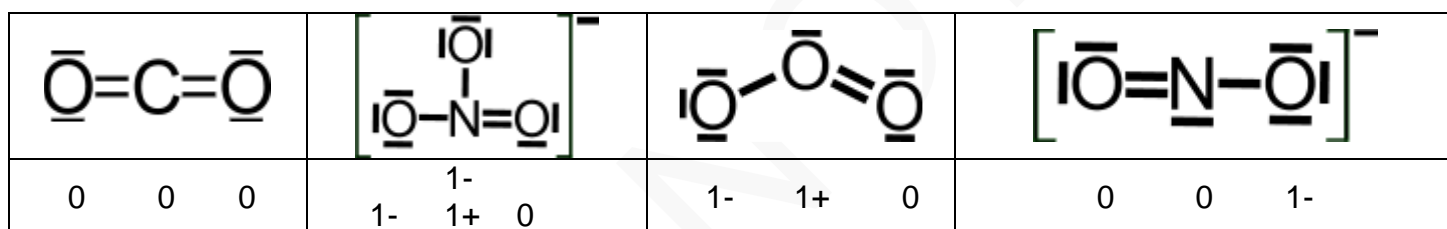
XeF₂ (10 e⁻ in outer shell of Xe)

SCl₂

- 13 Understand what is unusual about the Lewis structures for BCl_3 and BeCl_2
 B only has $6e^-$ in outer shell, Be only has $4e^-$ in outer shell – neither has a complete octet.
- 14 Explain what is meant by the term *coordinate covalent bond*
 Covalent bond where both shared electrons come from the same atom
- 15 Explain which of the following contain coordinate covalent bonds
 O_3 CO PCl_3 NH_4^+ NO_2^+

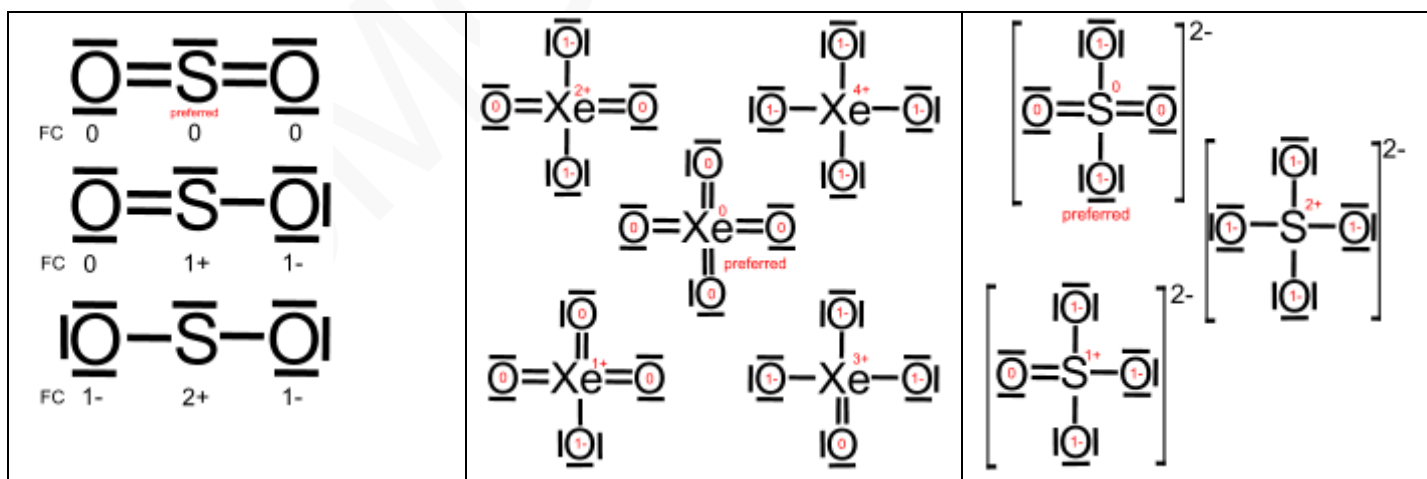


- 16 Explain what *formal charge* is.
 The charge an atom in a molecule/ion would have if all atoms had the same electronegativity – arises from formation of coordinate covalent bonds
- 17 Work out the formal charge on each atom in each of the following



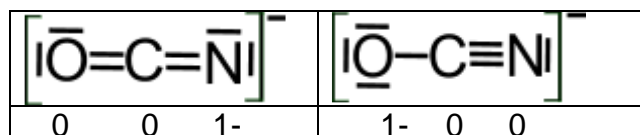
- 18 Draw at least 2 Lewis structures for each of the following and use formal charge to decide which is the preferred structure

(a) SO_2 (b) XeO_4 (c) SO_4^{2-}



- 19 Explain why formal charge cannot be used to deduce which of the Lewis structures shown for OCN^- is preferred.

Formal charges the same (only on different atoms)



20 Explain the valence shell electron pair repulsion theory for predicting shapes of molecules.

Pairs of electrons (electron domains) in the outer shell of the central atom

Repel each other

Take up positions in space to be as far away from each other as possible to minimise repulsions.

21 Predict the electron domain geometry, shape and bond angles of:

CH ₄	H ₂ O	C ₂ H ₄	SO ₂	H ₃ O ⁺	CO ₃ ²⁻
NH ₃	CO ₂	C ₂ H ₂	NH ₄ ⁺	BF ₃	PCl ₃

	Electron domain geometry	Shape	Suggested bond angle
CH ₄	Tetrahedral	Tetrahedral	109.5°
H ₂ O	Tetrahedral	Bent	104.5°
C ₂ H ₄	Trigonal planar (about each C)	Trigonal planar (about each C)	120°
SO ₂	Trigonal planar	Bent	118°
H ₃ O ⁺	Tetrahedral	Trigonal pyramidal	107°
CO ₃ ²⁻	Trigonal planar	Trigonal planar	120°
NH ₃	Tetrahedral	Trigonal pyramidal	107°
CO ₂	Linear	Linear	180°
C ₂ H ₂	Linear (about each C)	Linear (about each C)	180°
NH ₄ ⁺	Tetrahedral	Tetrahedral	109.5°
BF ₃	Trigonal planar	Trigonal planar	120°
PCl ₃	Tetrahedral	Trigonal pyramidal	107°

22 Predict the electron domain geometry, shape and bond angles of:

XeF ₄	XeOF ₄	SO ₄ ²⁻	ClF ₃	SF ₄ ²⁻
PCl ₅	BrF ₅	SF ₄	SF ₆	I ₃ ⁻

	Electron domain geometry	Shape	Suggested bond angle
XeF ₄	octahedral	Square planar	90°
XeOF ₄	octahedral	Square pyramidal	88°
SO ₄ ²⁻	tetrahedral	tetrahedral	109.5°
ClF ₃	Trigonal bipyramidal	T-shaped	86°
SF ₄ ²⁻	octahedral	Square planar	90°
PCl ₅	Trigonal bipyramidal	Trigonal bipyramidal	90°/120°
BrF ₅	octahedral	Square pyramidal	88°
SF ₄	Trigonal bipyramidal	See-saw	88°
SF ₆	octahedral	octahedral	90°
I ₃ ⁻	Trigonal bipyramidal	linear	180°

23 Explain what is meant by *electronegativity*

The attraction by an atom for the electrons in a covalent bond

24 State and explain which of the following bonds is most polar:

H-C (least polar – smallest difference in electronegativity)

H-Cl

H-F (most polar – biggest difference in electronegativity)

25 Explain which of the following molecules is/are polar:

N ₂	NH ₃	SO ₂	XeF ₄
Non	Polar	Polar	non
CO	CO ₂	CH ₂ Cl ₂	PCl ₅
Polar	Non	Polar	non

26 Explain the meaning of the term *hybridisation*.

Mixing of atomic orbitals

- forms same number of hybrid orbitals
- orbitals have s and p character (if $sp/sp^2/sp^3$ hybridisation)
- orbitals better arranged in space for bonding

27 Predict the type of hybridisation present on the central atom in:

CH ₄	BF ₃	HCN	H ₂ O	SO ₂	CO ₃ ²⁻	NH ₃	CO ₂	C ₂ H ₄
sp ³	sp ²	sp	sp ³	sp ²	sp ²	sp ³	sp	sp ²

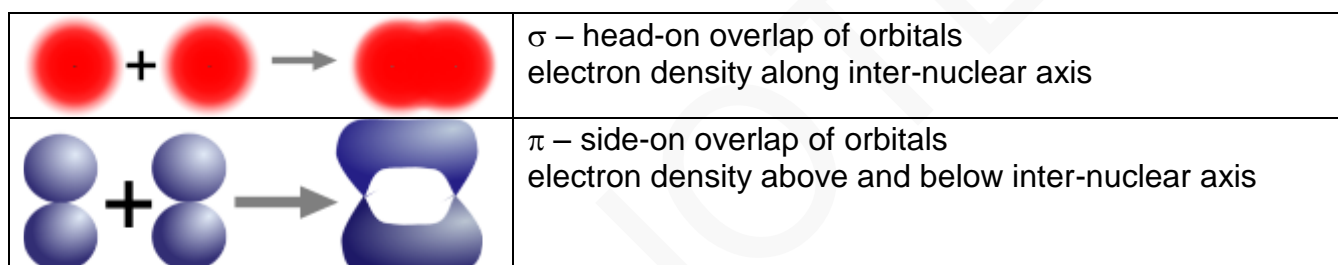
28 Predict the hybridisation of the O atoms in

CO ₂	SO ₂	CO
sp ²	sp ²	sp

29 Explain covalent bonding in terms of orbitals

Overlap of atomic/hybrid orbitals – usually both orbitals contain 1 e⁻ (but not in a coordinate bond)

30 Explain, using diagrams, what is meant by σ and π bonds.



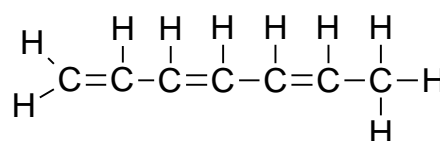
31 Work out the number of σ and π bonds in:

	C ₂ H ₄	HCN	C ₂ H ₂	N ₂ H ₄	CO
σ bonds	5	2	3	5	1
π bonds	1	2	2	0	2

32 In the organic molecule shown work out

(a) the number of σ and π bonds

16 σ , 3 π

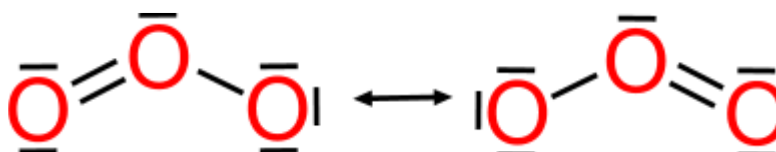


(b) the number of C atoms that are sp^2 hybridised 6

33 Explain what is meant by a resonance structure

A Lewis structure for a molecule/ion that cannot be described by one Lewis structure alone. The actual structure is best described as a hybrid of the resonance structures

34 Draw two resonance structures for O₃ and explain why the structure of ozone cannot be described by just one of them.



The O-O bond lengths in ozone are equal which one structure alone would not predict
The actual structure is a hybrid of these 2 structures

- 35 Explain what is meant by delocalisation of electrons and explain why a delocalised bonding scheme is necessary to describe the structure of ozone.

Electrons are shared between 3 or more atoms

The O-O bond lengths in ozone are equal and between O=O and O-O

Delocalisation of π electrons in O_3

- 36 Explain why UV radiation of shorter wavelength is required to dissociate O_2 than O_3

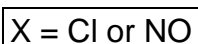
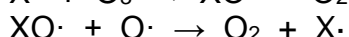
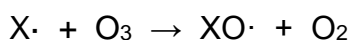
Bond order in O_2 is 2

Bond order in O_3 is 1.5

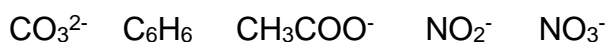
Stronger bond in O_2

Therefore shorter wavelength (higher energy) UV radiation required to dissociate O_2

- 37 Write equations to show the mechanism of ozone depletion as catalysed by CFCs and NO_x .



- 38 Draw resonance structures for each of the following



CO_3^{2-}	
C_6H_6	
CH_3COO^-	
NO_2^-	
NO_3^-	

- 39 Explain why a delocalised bonding scheme is necessary to describe the structure of molecules/ions in 38.

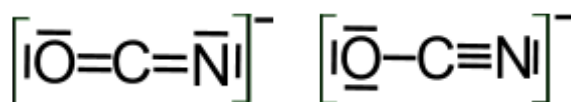
CO_3^{2-}	All C-O bond length are equal and between C-O and C=O – Lewis structure would predict two different bond lengths.
CH_3COO^-	
C_6H_6	All C-C bond length are equal and between C-C and C=C – Lewis structure would predict two different bond lengths.
NO_2^-	Both N-O bond length are equal and between N-O and N=O – Lewis structure would predict two different bond lengths.
NO_3^-	

40 Explain whether the C-O bond length in CO_3^{2-} or CH_3COO^- is longer

CO_3^{2-} longer bond – bond order 1.33

CH_3COO^- shorter bond – bond order 1.5

41 Use the resonance structures shown here to predict a C-O bond length for OCN^-



Bond	Bond length / pm
C-O	143
C=O	122
C≡C	113

Between 122 pm and 143 pm

42 Explain which of the following contain delocalised electrons

H_2SO_4 PO_4^{3-} CH_3COOH H_2O_2 NO_2^+ SO_3^{2-} H_2CO_3 HCO_3^-

Multiple Lewis structures can be written for each one which are equivalent except for where double/single bonds and associated lone pairs are – cannot be described by just 1 Lewis structures as X-O bond lengths are equal.

43 Explain the origin of London (dispersion) forces.

Unsymmetrical distribution of electrons

Instantaneous dipole –induced dipole attraction.

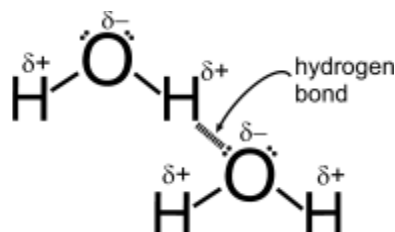
44 Explain what van der Waals forces are.

Intermolecular forces – London forces and permanent dipole-dipole interactions – **not** hydrogen bonds or ion-dipole interactions.

45 For each of the following pairs, explain in terms of intermolecular forces, which has the **higher boiling point**.

Cl_2 or Br_2	Higher relative molecular mass – stronger London forces
CH_3Cl or CH_3Br	
$\text{CH}_3\text{CH}_2\text{CH}_3$ or CH_3OCH_3	Similar relative molecular masses – London forces similar. CH_3OCH_3 is polar so also has permanent dipole – permanent dipole forces between molecules – stronger intermolecular forces
$\text{CH}_3\text{CH}_2\text{OH}$ or CH_3OCH_3	Same relative molecular masses – London forces similar. Both polar. $\text{CH}_3\text{CH}_2\text{OH}$ has hydrogen bonding between molecules – stronger intermolecular force.
NH_3 or PH_3	NH_3 has hydrogen bonding between molecules, PH_3 does not. NH_3 has stronger intermolecular forces.

46 Draw a diagram to show the hydrogen bonding between molecules of water.



47 Explain which of the following will have hydrogen bonding between molecules

HF	HCl	CH_3F	CH_3OH	N_2H_4
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Have H covalently bonded to N, O or F and a lone pair of electrons on N/O/F

- 48 Explain why covalent molecular compounds:
- have low melting and boiling points
 - are volatile
 - do not conduct electricity when liquid
 - tend to be soluble in organic solvents

low melting and boiling points	Weak intermolecular forces require little energy to overcome
volatile	Weak intermolecular forces require little energy to overcome
do not conduct electricity when liquid	No ions and electrons held tightly in covalent bonds – no charged particles which are free to move
soluble in organic solvents	The intermolecular forces in organic solvents are more similar to those in covalent compounds

- 49 Explain which of the following chlorides will have the lowest boiling point

NaCl MgCl₂ SiCl₄

NaCl and MgCl₂ are ionic so have higher boiling points than SiCl₄ which is covalent molecular.

NaCl/MgCl₂ – strong electrostatic forces of attractions between oppositely-charged ions need a lot of energy to overcome.

SiCl₄ – London forces between molecules require little energy to overcome.

- 50 Explain which of the following chlorides will be the best conductor of electricity in the liquid state

NaCl SiCl₄ SCl₂

NaCl is ionic so mobile ions present in the liquid state.

SiCl₄/SCl₂ molecular in the liquid state – no ions.

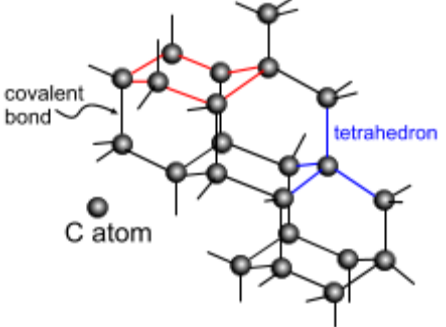
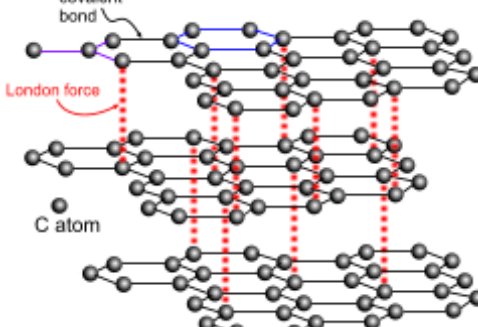
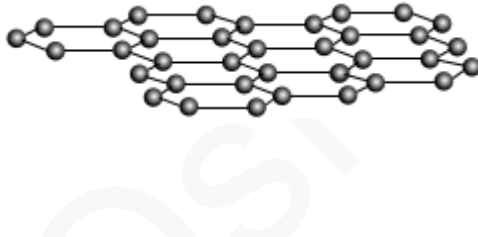
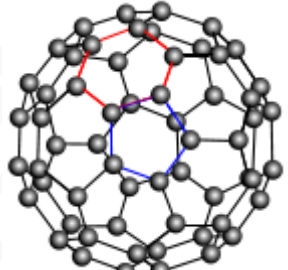
- 51 Explain the relative solubilities of ionic and covalent substances in polar and non-polar solvents.

Polar solvents – ionic substances tend to be soluble: ion-dipole forces between polar molecules and ions pay back the energy to break apart the lattice. Non-polar covalent substances: less soluble in polar solvents – weak (London) forces formed between solute and solvent molecules - does not pay back the energy to break dipole-dipole forces between solvent molecules.

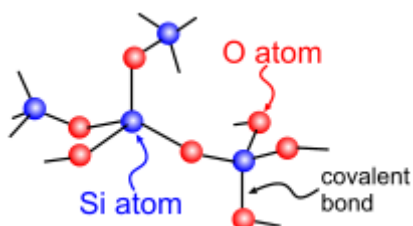
Non-polar solvents – Ionic substances tend to be insoluble: forces between ions and non-polar molecules are weak – not enough energy released to break apart the lattice. Non-polar covalent substances tend to be soluble: similar strength of intermolecular forces in solute, solvent and solution – energy released when solute-solvent interactions formed (London forces) enough to overcome London forces in pure solute/solvent.

52 Describe the structure and bonding of

- Diamond
- Graphite
- Graphene
- C₆₀ buckminsterfullerene

diamond		<p>Giant covalent</p> <p>Each C joined to 4 others by covalent bonds</p> <p>Tetrahedral arrangement about each C</p> <p>Carbon atoms in puckered hexagons</p>
graphite		<p>Giant covalent</p> <p>Each C joined to 3 others by covalent bonds</p> <p>Trigonal planar arrangement about each C</p> <p>Carbon atoms in planar hexagons layers with London forces between them</p>
graphene		<p>Giant covalent</p> <p>Single layer of graphite</p> <p>Each C joined to 3 others by covalent bonds</p> <p>Trigonal planar arrangement about each C</p>
C ₆₀ buckminsterfullerene		<p>Covalent molecular</p> <p>Each C joined to 3 others by covalent bonds</p> <p>hexagons and pentagons of C atoms</p> <p>Bond angles less than 120°</p> <p>London forces between molecules</p>

53 Describe the structure and bonding of silicon dioxide



Giant covalent.

Each Si covalently bonded to 4 O atoms

Each O covalently bonded to 2 Si atoms

 54 Explain why diamond has a much higher melting/boiling point than C₆₀ buckminsterfullerene

Diamond – giant covalent – lots of energy required to break strong covalent bonds

C₆₀ - simple molecular covalent – not much energy required to break weak London forces between molecules.

- 55 Explain why graphite and graphene are better conductors of electricity than diamond and C₆₀ fullerene.

Graphite/graphene – each C forms 3 bonds – 1 electron on each C atom delocalised over the layer – delocalised electrons free to move

Diamond – all electrons held tightly in covalent bonds – not free to move

C₆₀ – molecular structure – electrons delocalised within molecules but cannot easily move between them.

- 56 Explain what is meant by *metallic bonding*.

giant lattice of positive ions and sea of delocalised electrons

electrostatic attraction between positive ions and delocalised electrons

- 57 Explain why metals conduct electricity and are malleable.

Conduct electricity – delocalised electrons free to move

Malleable – layers of positive ions can slide over each other without changing the bonding – bonding non-directional.

- 58 Explain why magnesium has a higher melting point than sodium

Mg – stronger electrostatic attraction between positive ions and delocalised electrons

Mg forms 2+ ion but Na forms 1+

Mg²⁺ smaller than Na⁺

Mg more delocalised electrons than Na

- 59 Explain what an *alloy* is.

Homogeneous mixtures of two or more metals, or of a metal with a non-metal

- 60 Explain why alloys tend to be stronger than pure metals.

A differently-sized atom in the structure prevents layers of ions sliding over each other as easily.