

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<sup>+</sup> Ba<sup>2+</sup> Al<sup>3+</sup> Br<sup>-</sup> O<sup>2-</sup> N<sup>3-</sup>

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

<b>Nitrate</b>	<b>Sulfate</b>	<b>Hydroxide</b>	<b>Ammonium</b>
NO <sub>3</sub> <sup>-</sup>	SO <sub>4</sub> <sup>2-</sup>	OH <sup>-</sup>	NH <sub>4</sub> <sup>+</sup>
<b>Carbonate</b>	<b>Hydrogencarbonate</b>	<b>Phosphate</b>	
CO <sub>3</sub> <sup>2-</sup>	HCO <sub>3</sub> <sup>-</sup>	PO <sub>4</sub> <sup>3-</sup>	

4 Work out the formulae of the following ionic compounds:

Lithium fluoride	LiF	Sodium hydrogencarbonate	NaHCO <sub>3</sub>	Iron(II) phosphate	Fe <sub>3</sub> (PO <sub>4</sub> ) <sub>2</sub>
magnesium chloride	MgCl <sub>2</sub>	Calcium nitrate	Ca(NO <sub>3</sub> ) <sub>2</sub>	Sodium carbonate	Na <sub>2</sub> CO <sub>3</sub>
ammonium sulfate	(NH <sub>4</sub> ) <sub>2</sub> SO <sub>4</sub>	Barium hydroxide	Ba(OH) <sub>2</sub>	Copper(II) nitrate	Cu(NO <sub>3</sub> ) <sub>2</sub>

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

CO<sub>2</sub> CaS PCl<sub>3</sub> OF<sub>2</sub> MgO

<b>mainly ionic</b>	<b>mainly covalent</b>
CaS	MgO

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 <sub>2</sub>	C <sub>2</sub> H <sub>2</sub>	SO <sub>4</sub> <sup>2-</sup>
HCN	SO <sub>2</sub>	N <sub>2</sub> H <sub>4</sub>
CO <sub>2</sub>	NH <sub>4</sub> <sup>+</sup>	ClF <sub>3</sub>
C <sub>2</sub> H <sub>4</sub>	CO <sub>3</sub> <sup>2-</sup>	PCl <sub>5</sub>

	H-C≡C-H	
H-C≡N		

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

- C<sub>2</sub>H<sub>6</sub>, C<sub>2</sub>H<sub>4</sub> and C<sub>2</sub>H<sub>2</sub>
- N<sub>2</sub>, N<sub>2</sub>H<sub>4</sub> and N<sub>2</sub>H<sub>2</sub>
- 

C<sub>2</sub>H<sub>6</sub> (C-C single bond – longest and weakest)

C<sub>2</sub>H<sub>4</sub> (C=C double bond shorter than C-C and stronger than C-C)

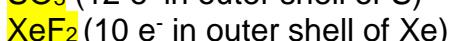
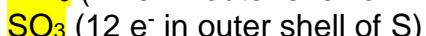
C<sub>2</sub>H<sub>2</sub> (C≡C triple bond – shortest and strongest – more shared electrons – greatest attraction between nuclei and shared electrons)

N<sub>2</sub>H<sub>4</sub> (N-N single bond – longest and weakest)

N<sub>2</sub>H<sub>2</sub> (N=N double bond shorter than N-N and stronger than N-N)

N<sub>2</sub>(N≡N triple bond – shortest and strongest)

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



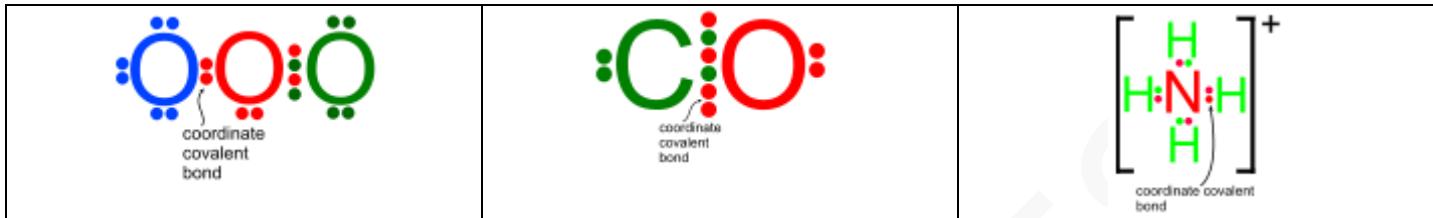
13 Understand what is unusual about the Lewis structures for  $\text{BCl}_3$  and  $\text{BeCl}_2$

B only has  $6\text{e}^-$  in outer shell, Be only has  $4\text{e}^-$  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



16 Explain what *formal charge* is.

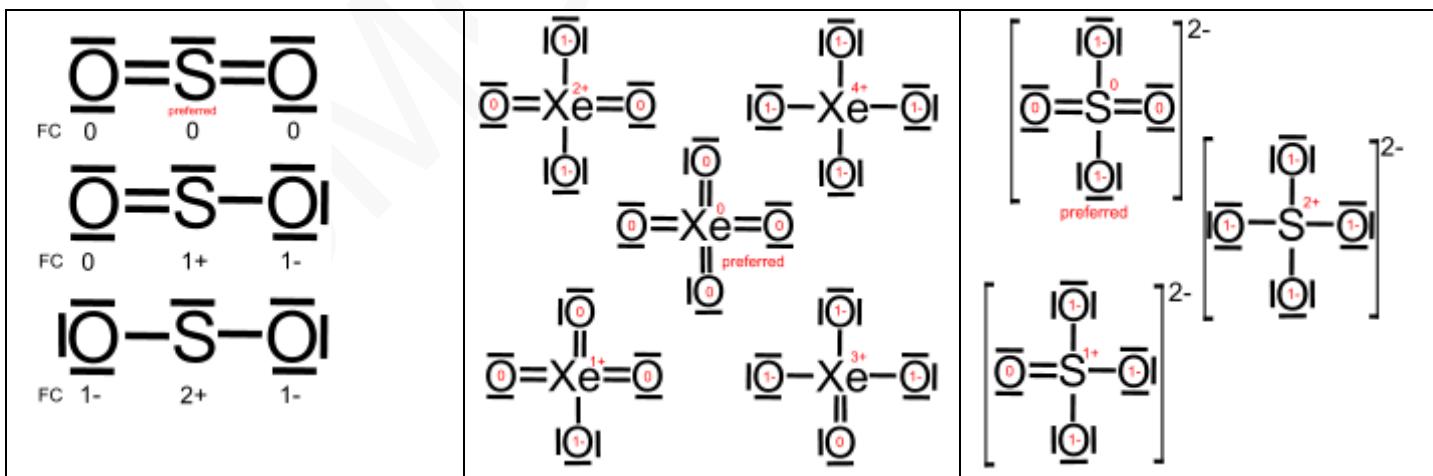
The charge an atom in a molecule/ion would have if all atoms had the same electronegativity – arises from formation of coordinate covalent bonds

17 Work out the formal charge on each atom in each of the following

$\bar{\text{O}}=\text{C}=\bar{\text{O}}$	$[\bar{\text{O}}-\text{N}=\bar{\text{O}}]^-$	$\bar{\text{O}}-\bar{\text{O}}-\bar{\text{O}}=\bar{\text{O}}$	$[\bar{\text{O}}=\text{N}-\bar{\text{O}}]^-$
0    0    0	1-    1+    0	1-    1+    0	0    0    1-

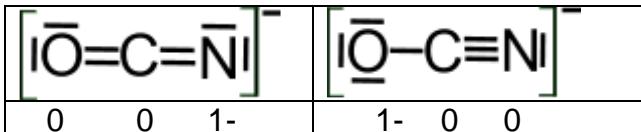
18 Draw at least 2 Lewis structures for each of the following and use formal charge to decide which is the preferred structure

- (a)  $\text{SO}_2$       (b)  $\text{XeO}_4$       (c)  $\text{SO}_4^{2-}$



19 Explain why formal charge cannot be used to deduce which of the Lewis structures shown for  $\text{OCN}^-$  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:

$\text{CH}_4$	$\text{H}_2\text{O}$	$\text{C}_2\text{H}_4$	$\text{SO}_2$	$\text{H}_3\text{O}^+$	$\text{CO}_3^{2-}$
$\text{NH}_3$	$\text{CO}_2$	$\text{C}_2\text{H}_2$	$\text{NH}_4^+$	$\text{BF}_3$	$\text{PCl}_3$

	Electron domain geometry	Shape	Suggested bond angle
$\text{CH}_4$	Tetrahedral	Tetrahedral	$109.5^\circ$
$\text{H}_2\text{O}$	Tetrahedral	Bent	$104.5^\circ$
$\text{C}_2\text{H}_4$	Trigonal planar (about each C)	Trigonal planar (about each C)	$120^\circ$
$\text{SO}_2$	Trigonal planar	Bent	$118^\circ$
$\text{H}_3\text{O}^+$	Tetrahedral	Trigonal pyramidal	$107^\circ$
$\text{CO}_3^{2-}$	Trigonal planar	Trigonal planar	$120^\circ$
$\text{NH}_3$	Tetrahedral	Trigonal pyramidal	$107^\circ$
$\text{CO}_2$	Linear	Linear	$180^\circ$
$\text{C}_2\text{H}_2$	Linear (about each C)	Linear (about each C)	$180^\circ$
$\text{NH}_4^+$	Tetrahedral	Tetrahedral	$109.5^\circ$
$\text{BF}_3$	Trigonal planar	Trigonal planar	$120^\circ$
$\text{PCl}_3$	Tetrahedral	Trigonal pyramidal	$107^\circ$

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

$\text{XeF}_4$	$\text{XeOF}_4$	$\text{SO}_4^{2-}$	$\text{ClF}_3$	$\text{SF}_4^{2-}$
$\text{PCl}_5$	$\text{BrF}_5$	$\text{SF}_4$	$\text{SF}_6$	$\text{I}_3^-$

	Electron domain geometry	Shape	Suggested bond angle
$\text{XeF}_4$	octahedral	Square planar	$90^\circ$
$\text{XeOF}_4$	octahedral	Square pyramidal	$88^\circ$
$\text{SO}_4^{2-}$	tetrahedral	tetrahedral	$109.5^\circ$
$\text{ClF}_3$	Trigonal bipyramidal	T-shaped	$86^\circ$
$\text{SF}_4^{2-}$	octahedral	Square planar	$90^\circ$
$\text{PCl}_5$	Trigonal bipyramidal	Trigonal bipyramidal	$90^\circ/120^\circ$
$\text{BrF}_5$	octahedral	Square pyramidal	$88^\circ$
$\text{SF}_4$	Trigonal bipyramidal	See-saw	$88^\circ$
$\text{SF}_6$	octahedral	octahedral	$90^\circ$
$\text{I}_3^-$	Trigonal bipyramidal	linear	$180^\circ$

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:

$\text{N}_2$	$\text{NH}_3$	$\text{SO}_2$	$\text{XeF}_4$
Non	Polar	Polar	non
$\text{CO}$	$\text{CO}_2$	$\text{CH}_2\text{Cl}_2$	$\text{PCl}_5$
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:

$\text{CH}_4$	$\text{BF}_3$	$\text{HCN}$	$\text{H}_2\text{O}$	$\text{SO}_2$	$\text{CO}_3^{2-}$	$\text{NH}_3$	$\text{CO}_2$	$\text{C}_2\text{H}_4$
$\text{sp}^3$	$\text{sp}^2$	$\text{sp}$	$\text{sp}^3$	$\text{sp}^2$	$\text{sp}^2$	$\text{sp}^3$	$\text{sp}$	$\text{sp}^2$

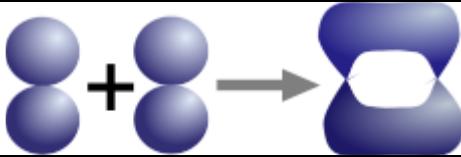
28 Predict the hybridisation of the O atoms in

$\text{CO}_2$	$\text{SO}_2$	$\text{CO}$
$\text{sp}^2$	$\text{sp}^2$	$\text{sp}$

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  $\sigma$  and  $\pi$  bonds.

	$\sigma$ – head-on overlap of orbitals electron density along inter-nuclear axis
	$\pi$ – side-on overlap of orbitals electron density above and below inter-nuclear axis

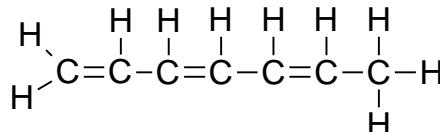
31 Work out the number of  $\sigma$  and  $\pi$  bonds in:

	$\text{C}_2\text{H}_4$	$\text{HCN}$	$\text{C}_2\text{H}_2$	$\text{N}_2\text{H}_4$	$\text{CO}$
$\sigma$ bonds	5	2	3	5	1
$\pi$ bonds	1	2	2	0	2

32 In the organic molecule shown work out

(a) the number of  $\sigma$  and  $\pi$  bonds

16  $\sigma$ , 3  $\pi$

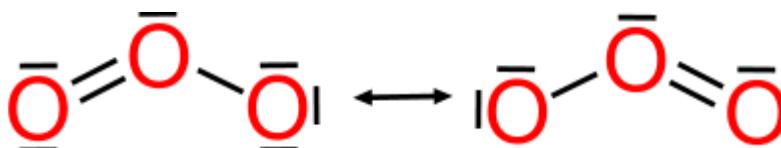


(b) the number of C atoms that are  $sp^2$  hybridised 6

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  $\text{O}_3$  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  $\pi$  electrons in  $O_3$

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

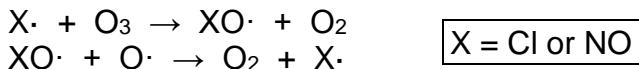
Bond order in  $O_2$  is 2

Bond order in  $O_3$  is 1.5

Stronger bond in  $O_2$

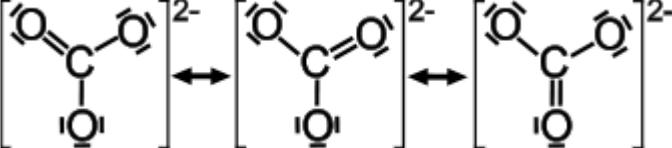
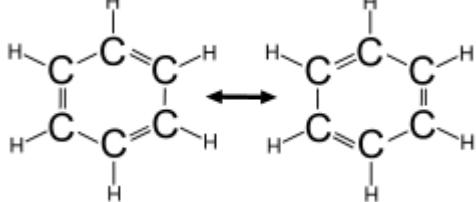
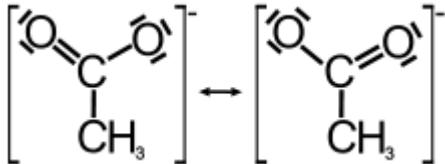
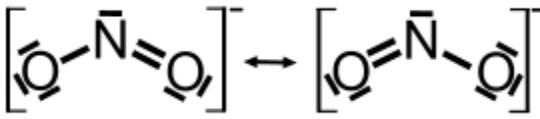
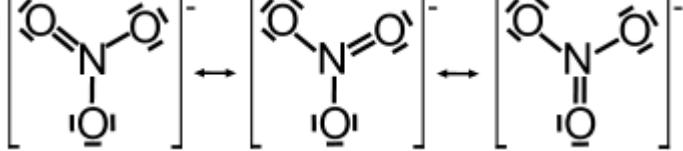
Therefore shorter wavelength (higher energy) UV radiation required to dissociate  $O_2$

- 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.
$NO_2^-$	
$NO_3^-$	Both N-O bond length are equal and between N-O and N=O – Lewis structure would predict two different bond lengths.

- 40 Explain whether the C-O bond length in  $\text{CO}_3^{2-}$  or  $\text{CH}_3\text{COO}^-$  is longer

$\text{CO}_3^{2-}$  longer bond – bond order 1.33

$\text{CH}_3\text{COO}^-$  shorter bond – bond order 1.5

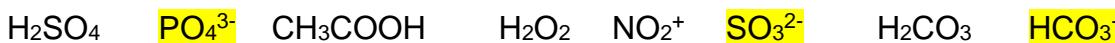
- 41 Use the resonance structures shown here to predict a C-O bond length for  $\text{OCN}^-$



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



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.

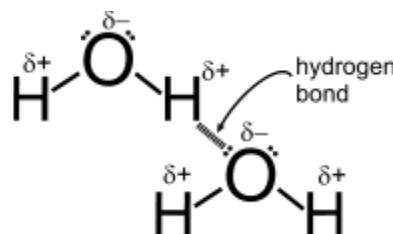
- 44 Explain what van der Waals forces are.

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

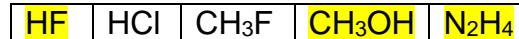
- 45 For each of the following pairs, explain in terms of intermolecular forces, which has the **higher boiling point**.

$\text{Cl}_2$ or $\text{Br}_2$	Higher relative molecular mass – stronger London forces
$\text{CH}_3\text{Cl}$ or $\text{CH}_3\text{Br}$	
$\text{CH}_3\text{CH}_2\text{CH}_3$ or $\text{CH}_3\text{OCH}_3$	Similar relative molecular masses – London forces similar. $\text{CH}_3\text{OCH}_3$ is polar so also has permanent dipole – permanent dipole forces between molecules – stronger intermolecular forces
$\text{CH}_3\text{CH}_2\text{OH}$ or $\text{CH}_3\text{OCH}_3$	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.
$\text{NH}_3$ or $\text{PH}_3$	$\text{NH}_3$ has hydrogen bonding between molecules, $\text{PH}_3$ does not. $\text{NH}_3$ 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



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    MgCl<sub>2</sub>    **SiCl<sub>4</sub>**

NaCl and MgCl<sub>2</sub> are ionic so have higher boiling points than SiCl<sub>4</sub> which is covalent molecular.

NaCl/MgCl<sub>2</sub> – strong electrostatic forces of attractions between oppositely-charged ions need a lot of energy to overcome.

SiCl<sub>4</sub> – 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**    SiCl<sub>4</sub>    SCl<sub>2</sub>

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

SiCl<sub>4</sub>/SCl<sub>2</sub> molecular in the liquid state – no ions.

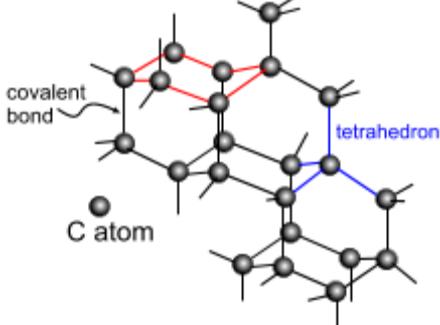
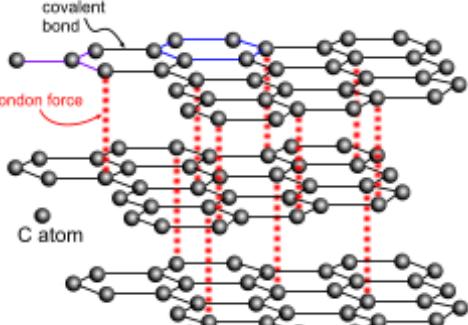
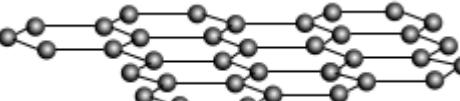
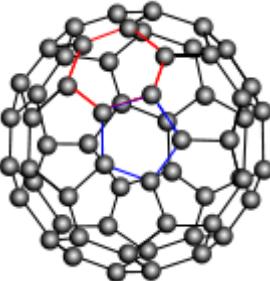
## 51 Explain the relative solubilities of ionic and covalent substances in polar and non-polar solvents.

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.

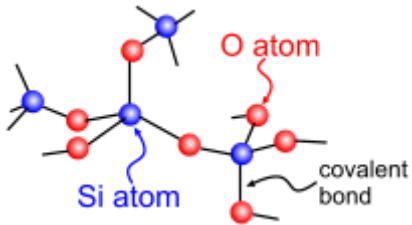
Non-polar solvents – Ionic substances tend to be insoluble: forces between ions and non-polar molecules are weak – not enough energy released to break apart the lattice. Non-polar covalent substances tend to be soluble: similar strength of intermolecular forces in solute, solvent and solution – energy 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
- C<sub>60</sub> buckminsterfullerene

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

## 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 C<sub>60</sub> buckminsterfullerene

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

C<sub>60</sub> - 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 C<sub>60</sub> 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

C<sub>60</sub> – 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+

Mg<sup>2+</sup> smaller than Na<sup>+</sup>

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.