Pages 77–80 Exam practice questions

1 a) Bonding between atoms in a metal is the result of electrostatic attractions between positive metal ions in a lattice [1] and delocalised electrons in the outer shell of the metal atoms. [1] In sodium, this results from ions with a charge of 1+ and one delocalised electron in each atom. In magnesium, the bonding results from ions with a charge of 2+ and two delocalised electrons in each atom. The bonding between atoms is therefore weaker in Na [1] and its melting temperature is much lower.

b) P_4 and S_8 are relatively simple molecules with weak forces of attraction. However, S_8 is larger and heavier than P_4 [1] and it requires higher temperatures before its molecules have sufficient energy to move away from fixed positions [1] in the solid lattice.

c) i) Aluminium – giant metallic structure [1]; bonding involves the attraction of positive ions for delocalised electrons. [1]

ii) Silicon – giant covalent structure [1]; bonding involves the attraction of positive nuclei for shared electrons [1] in covalent bonds. [1]

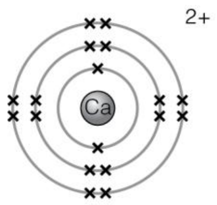
iii) Chlorine – simple molecular structure [1]; bonding involves the attraction of positive nuclei for shared electrons [1] in covalent bonds between atoms plus weak intermolecular forces. [1]

(*Any 6 points*)

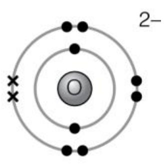
2 a) i) Bonding in calcium involves the electrostatic attraction of positive ions [1] in a giant lattice for the delocalised electrons [1] in the outermost shell of the calcium atoms. [1]

ii) Calcium contains delocalised electrons. [1] When calcium is connected to a battery, these delocalised electrons are attracted to the positive terminal [1] of the battery, forming an electric current.

b)



Ca^{2+} (2, 8, 8)



O^{2-} (2, 8)

[1] for both correct charges, 2 × [1] for each electron structure, [1] for two electrons of Ca transferred to O.

c) CaO conducts electricity when molten (liquid). [1]

Ca^{2+} [1] ions are attracted to the negative electrode from which they take electrons and form Ca metal. [1]

O^{2-} [1] ions are attracted to the positive electrode where they give up electrons and form O_2 gas. [1]

The giving of electrons to the positive electrode and taking of electrons from the negative electrode makes an electric current. [1]

3 a) Hydrogen chloride has covalent bonding. [1] The hydrogen and chlorine atoms share a pair of electrons [1] in a covalent bond. Dot-and-cross diagram. [1]

Sodium chloride has ionic bonding. [1] Sodium ions, Na^+ and chloride ions Cl^- are held in a giant lattice [1] by the attraction between oppositely charged ions. [1]

Copper has metallic bonding. [1] This involves the attraction of positive copper ions [1] in a giant lattice for delocalised electrons. [1]

b) Sodium chloride does not conduct when solid but will conduct when liquid (molten) or dissolved in water. [1]

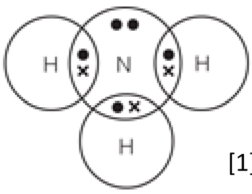
Under these conditions the ions are no longer held in a lattice and can move through the sodium chloride to the electrodes. [1]

In copper, delocalised electrons move through the metallic lattice. [1]

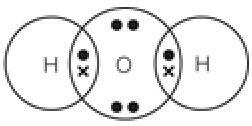
c) In copper, the metallic lattice structure means layers of identical atoms move over each other without repulsion. [1]

In sodium chloride, the ionic structure means that if one layer of ions is moved on impact the distance of only one ion diameter with respect to another layer [1], then like charges will be adjacent and will repel [1], causing the layers to separate.

4 a)



[1]



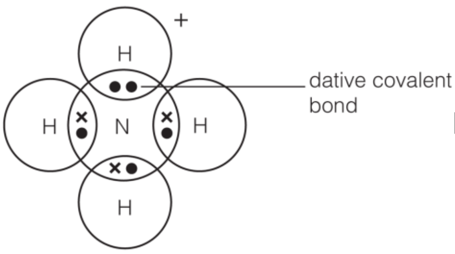
[1]

b) i) $NH_3(g) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

[1] for products, [1] for state symbols

ii) A dative covalent bond is a bond in which two atoms share a pair of electrons [1], both of the electrons being contributed by one atom. [1]

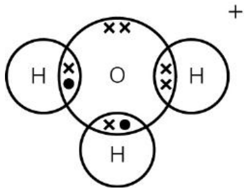
iii)



[1]

[1]

5 a)

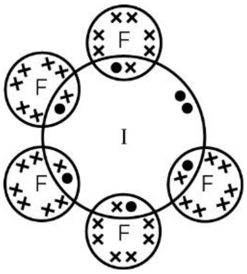


[2]

(trigonal) pyramidal [1] 107° [1]

(based on tetrahedral electron arrangement)

b)

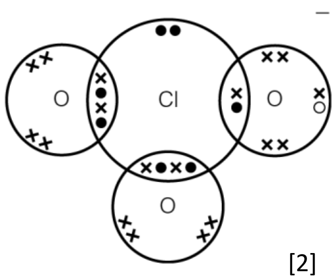


[2]

(square) pyramidal [1] 90° [1]

(based on octahedral electron arrangement)

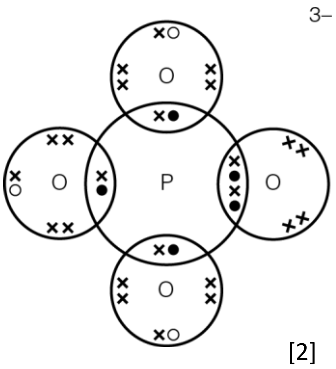
c)



[2]

pyramidal [1] 107° [1]

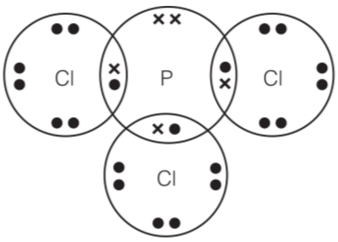
d)



[2]

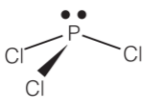
tetrahedral [1] 109.5° [1]

6 a)



[2]

b)



[1]

Pyramidal [1]; bond angle 107° [1]

c) There are four pairs of electrons around the phosphorus atom [1] (one lone pair and three shared pairs). These four pairs repel each other as far as possible, taking up tetrahedral positions around the phosphorus atom [1] and leading to a pyramidal shape for atoms in PCl_3. The angle is reduced from 109.5° to 107° by the extra repulsion of the bonding pairs by the lone pair. [1]

d) The lone pair of electrons on the phosphorus atom in PCl_3 [1] donates this pair in a dative covalent bond [1] to the boron atom in BCl_3, giving the boron atom a full second shell of electrons. [1]

e) Both P and B now have 4 bonding pairs in their outer shell [1] so both have tetrahedral arrangement [1] leading to a bond angle of 109.5°. [1]

7 a) i) London forces. [1]

ii) London forces, hydrogen bonding. [1]

b) Propane and ethanol are molecules of about the same size and molar mass so the contribution from London forces is about the same for both molecules. [1]

Hydrogen bonds are a much stronger type of intermolecular force than London forces. [1]

c) Glycerol, with three −OH groups, has much more extensive hydrogen bonding between its molecules than ethanol. All three −OH groups can take part in hydrogen bonding. [1]

This extra hydrogen bonding makes it difficult for the molecules to move over each other in the liquid, which is very viscous. [1] It is also harder to separate the glycerol molecules from each other so the boiling temperature of glycerol is higher than that of ethanol. [1]

8 CF_4 is tetrahedral [1] and therefore non-polar. [1]

SF_4 is distorted tetrahedral (often called ‘see-saw’) [1] with four bonding pairs and a lone pair, so is polar. [1]

XeF_4 is square planar [1] and therefore non-polar. [1]

9 a) F is more electronegative than Cl so F−Cl is polar [1] but in a fluorine molecule both atoms are the same so the molecule is not polar. [1]

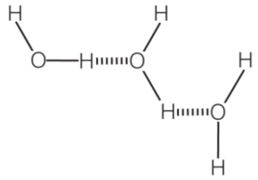
b) There is a lone pair of electrons on the S atom in SO_2 so the molecule is not linear. Hence it is polar because the S=O bonds are polar. [1]

CO_2, with just two double bonds, is linear and the polarities of the two bonds cancel. [1]

c) There is a lone pair of electrons on the N in NCl_3 so the molecule is pyramidal. With three polar N−Cl bonds the molecule is overall polar. [1]

With no lone pairs on BCl_3 the molecule is planar and symmetrical. Hence it is not polar despite the polarity of the B−Cl bonds. [1]

10 a) Hydrogen bonding [1]



A lone pair on the oxygen atom [1] is attracted to the very electron-deficient hydrogen atom [1] on another molecule.

b) i) Water is the only one of the compounds affected by hydrogen bonding. Its boiling temperature is much higher than expected. [1]
The trend in boiling temperatures for the other hydrides in the group is as expected for a series of compounds with similar structures but increasing numbers of electrons per molecule. The molecules become more polarisable down the group and so London forces increase. [1]

ii) Molecules in ice are held together by hydrogen bonding. The molecules form a giant lattice structure in which each oxygen atom is bonded to two hydrogen atoms by covalent bonds and two others by hydrogen bonds. This gives rise to an open lattice. [1]
As ice melts the hydrogen-bonded structure collapses bringing the water molecules closer together. So the melt water is denser than the ice. [1]

iii) Non-polar pentane molecules [1] cannot break into the hydrogen-bonded network of molecules in water. [1]

iv) The intermolecular forces in methoxymethane are London forces and attractions between permanent dipoles. [1] The hydrogen bonding between ethanol molecules is much stronger. [1]

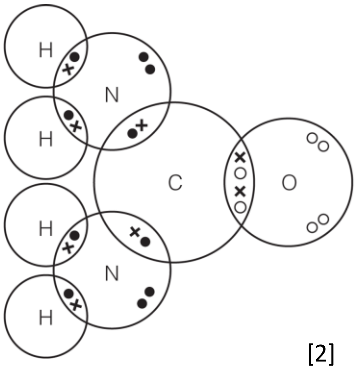
11 a) Allotropes are different forms of the same element [1] in the same state. [1]

b) Yes, because they are another solid form of carbon. [1]

c) i) Any two from: graphite fibres will conduct electricity; graphite fibres act as lubricants maintaining smooth contact; graphite fibres are flexible and will not break as the contacts open and close. [2]

ii) Diamonds do not conduct electricity. [1] Diamonds would be too expensive. [1]
Diamonds would scratch any surface with which they made contact. [1]

12 a)



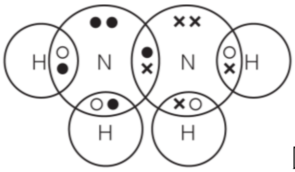
[2]

b) i) Trigonal planar [1]

ii) Pyramidal [1]

c) Any two of: urea provides higher percentage of nitrogen; urea would not result in nitrate contamination of the soil; urea would not cause eutrophication like nitrate; urea is not as soluble as NH_4NO_3, so it is not washed out of the soil so readily. [2]

13 a)



[2]

b) Allow 106°–108°. [1] The lone pair will repel more strongly than the shared pairs of electrons [1] around the N atoms. So, the H–N–H bond angle will be slightly less than 109.5° [1], similar to that in ammonia.

c) $N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$

Formulae of reactants and products [1]; correctly balanced with state symbols [1]

(Allow $2H_2O(l)$. Allow equations in which the products are H_2O and N_2O or NO or NO_2.)

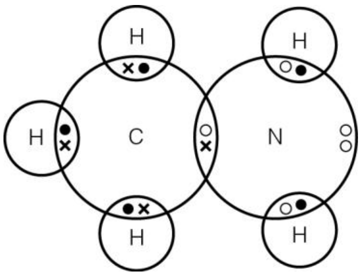
d) The molar mass of hydrazine = 32.0 g mol^{-1}

∴ When 1 mol N_2H_4 burns completely, the enthalpy change of combustion of hydrazine

= –(18.3 kJ g^{-1} × 32.0 g mol^{-1}) = –585.6 kJ mol^{-1} = –586 kJ mol^{-1} (3 s.f.)

[1] for 586, [1] for sign and units.

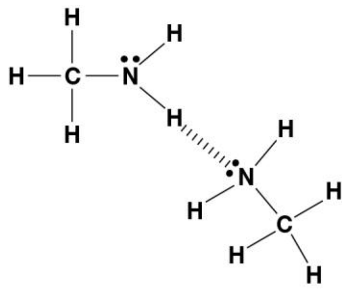
14 a)



[1]

b) H–C–H is 109.5° [1]; tetrahedral electron arrangement all bonding pairs. [1]

C–N–H is 107° [1]; tetrahedral electron arrangement three bonding pairs and one lone pair. [1] Lone pair repels more strongly than bonding pairs so reduces the bond angle. [1]

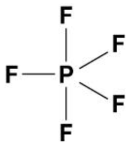
c)  [1]

Lone pair and dotted hydrogen bond [1]; 180° [1]

d) i) $CH_3NH_3^+Cl^-$ [1]

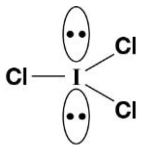
ii) 109.5° [1]; tetrahedral electron arrangement all bonding pairs. [1]

15 ICl_3 has five electron pairs so its structure is based on the trigonal bipyramidal electron arrangement [1] like PF_5.



But there are three bonding pairs and two lone pairs, so there are three possible arrangements:

A Both lone pairs vertical



[1]

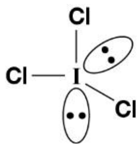
Repulsions

lp–lp 1 × 180°

lp–bp 6 × 90°

bp–bp 3 × 120° [1]

B One lone pair vertical one horizontal



[1]

Repulsions

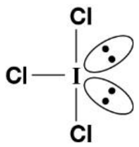
lp–lp 1 × 90°

lp–bp 2 × 120° & 3 × 90° & 1 × 180°

bp–bp 1 × 120° & 2 × 90°

[1]

C Both lone pairs horizontal



[1]

Repulsions

lp–lp 1 × 120°

lp–bp 4 × 90° & 2 × 120°

bp–bp 2 × 90° & 1 × 180°

[1]

The shape with the minimum overall repulsions is (C) where the atoms form a 'T-shaped' molecule. [1]

Although the lone pairs are not as far apart as in shape (A), the planar molecule, the lone-pair bonding pair repulsions are less in (C) than (A) and this is favoured. [1]

16 This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning. Assess the quality of the answers taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make, and illustrate, in an answer:

- Group 7 elements are non-polar molecules.
- Down the group the number of electrons in the molecule increases.
- Therefore the London forces between molecules increase.
- Group 1 elements are metals.
- Down the group, the number of delocalised electrons per atom remains the same but the size of the atoms increases.

17 a) Polar molecules are attracted to the charged rod. The tiny dipoles orientate themselves so that their opposing charge is pointing towards the charged rod. [1]

Non-polar molecules have no dipoles and are not attracted by the charged rod. [1]

b) Polar molecules that are deflected towards the charged rod: water and bromoethane.

Hexane and tetrachloromethane are non-polar molecules overall so they are not deflected

4 × (answer + prediction) for [1] each; total [4]

c) Polar molecules always orientate themselves so that the region of the molecules with the opposite charge to the charged rod is closer to the rod. [1] This means that attractive forces are always larger than repulsive forces between the rod and the molecules. [1]

18

Substance	a) **Bonding**	b) **Identity**
A	Covalent molecular [1]	1-Bromobutane [1]
B	Metallic lattice [1]	Manganese [1]
C	Ionic lattice [1]	Sodium bromide [1]
D	Giant covalent [1]	Silicon dioxide [1]
E	Covalent molecular [1]	Hydrogen bromide [1]
F	Ionic lattice [1]	Aluminium oxide [1]
G	Mobile metal atoms in liquid [1]	Mercury [1]

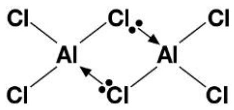
19 a) Both metals have ions with a charge of 1+ and one delocalised electron in each atom. [1] But sodium is a smaller ion than potassium [1] so the attraction for the delocalised electrons is stronger [1] so more energy is required to separate the ions. [1]

b) Oxide ion has greater charge than chloride [1], so force of attraction between ions is greater [1] so more energy is required to separate the ions. [1]

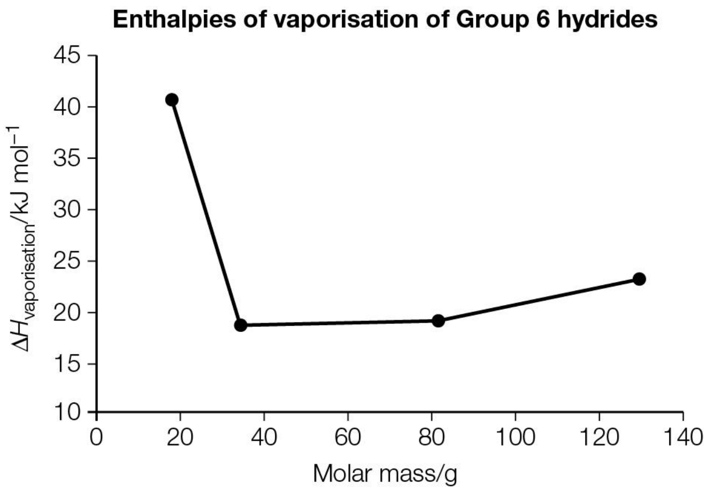
c) Weak forces between chlorine molecules are easily overcome [1] so little energy is needed to boil chlorine. [1]

The covalent bonds between chlorine atoms are strong [1] so a high temperature (or UV light) is required to provide the required energy. [1]

d) The bonding in pure aluminium chloride is largely covalent [1] because Al^{3+} polarises chloride ions. [1] An $AlCl_3$ molecule contains an Al atom with only 6 electrons in its outer energy level. [1] This vacant orbital can accept a pair of electrons from a chlorine atom forming a dative covalent bond. [1]



20 a)



Calculation of molar mass. [1] Axes [1]; points correctly plotted. [1] Good use of scale. [1]

b) The main intermolecular forces between the molecules of H_2S and H_2Se are permanent dipole–dipole attractions. [1]

c) Extrapolating the graph gives a value of about 18.5 kJ mol^{-1} for the $\Delta H_{vaporisation}$ of water from intermolecular forces other than hydrogen bonding. [1]

d) So the contribution of hydrogen bonding is about 22 kJ mol^{-1}. [1]

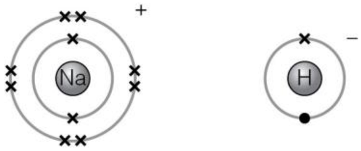
21 This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning. Assess the quality of the answers taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make, and illustrate, in an answer:

- the C–Cl bonds are polar
- the C–C single bond in 1,2-dichloroethane can rotate so the C–Cl bonds will not necessarily be opposite

- therefore the molecule has an overall dipole
- the C=C bond in 1,2-dichloroethene cannot rotate
- therefore the C–Cl bonds in the *E*-isomer will be opposite so there is no overall dipole
- and the C–Cl bonds in the *Z*-isomer will not be opposite so there is an overall dipole

22 a) i) $2Na(s) + H_2(g) \rightarrow 2NaH(s)$ [1] for equation; [1] for symbols

ii) 

[1] for each ion.

b) Hydrogen is discharged at the positive electrode. [1]

$2H^- \rightarrow H_2 + 2e^-$ [1]

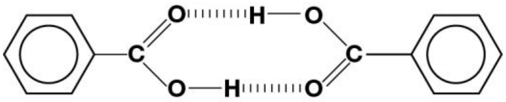
c) $H^- + H_2O \rightarrow H_2 + OH^-$ [1]

The hydride ion is acting as a proton acceptor or a base. [1]

d) The smaller size and larger charge makes Mg^{2+} more polarising than Na^+ [1], so MgH_2 has even more covalent character than NaH. [1] Heating will decompose MgH_2 into its elements [1] or treatment with water:

$MgH_2 + 2H_2O \rightarrow Mg(OH)_2 + 2H_2$ [1]

23 a) i) Despite the polar COOH group, the large non-polar benzene ring [1] limits the possible interactions with water molecules and the ability of the molecule to break into the hydrogen bonded network of water molecules. [1]

ii) 

[1]

The COOH groups cannot hydrogen bond to the non-polar solvent, [1] but they can form hydrogen bonds to another COOH. [1] The dimer then disperses throughout the non-polar solvent. [1]

b) i) $C_6H_5COOH + NaOH \rightarrow C_6H_5COONa + H_2O$ [1]

ii) M_r = 122.0 g mol^{-1} [1]

$$2.90 \text{ g acid} = \frac{2.90 \text{ g}}{122.0 \text{ g mol}^{-1}} = 0.02377 \text{ mol [1]}$$

$$\text{Volume of } 0.500 \text{ mol dm}^{-3} \text{ NaOH required} = \frac{0.02377 \text{ mol}}{0.500 \text{ mol dm}^{-3}}$$

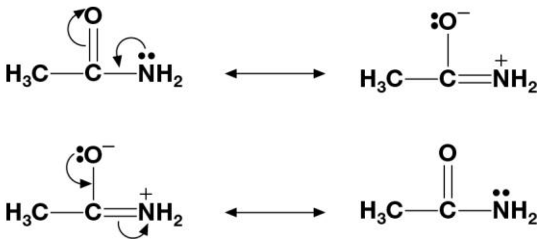
$$= 0.0475 \text{ dm}^3 \text{ or } 47.5 \text{ cm}^3 \text{ [1]}$$

iii) The salt formed is ionic [1] Hydration of the ions provides sufficient energy to overcome the lattice energy of the ionic compound. [1]

24 a) Ammonia and water contain hydrogen bonding [1], but the electronegativity of oxygen is greater than nitrogen [1] so the hydrogen bonding in water is stronger than that in ammonia. [1]

Methanol contains only one O–H bond per molecule and water contains two O–H bonds. [1] Although there are two lone pairs on each oxygen [1], fewer hydrogen bonds can be formed. [1]

b) Hydrogen bonding alone would suggest that the acid should have the higher boiling temperature, so extra interactions must be present. [1]



The lone pair on the N can be delocalised [1] (towards the more electronegative oxygen). This leads to an ionic structure [1] which will exert stronger electrostatic forces on similar species. [1]

25 Tin fluoride:

Sn	F
$\frac{61.0}{118.7}$ = 0.514	$\frac{39.0}{19.0}$ = 2.05
Empirical formula = SnF_4 [1]	

Tin iodide:

Sn	I
$\frac{19.0}{118.7}$ = 0.160	$\frac{81.0}{127}$ = 0.638
Empirical formula = SnI_4 [1]	

Electronegativity values Sn = 1.8 F = 4.0 difference = 2.2 probably ionic [1]

Electronegativity values Sn = 1.8 I = 2.5 difference = 0.7 probably covalent [1]

Melting temperature of SnF_4 is quite high so likely to be ionic. [1]

Melting temperature of SnI_4 is lower so likely to be molecular (tetrahedral shape so non-polar molecule). [1]

Sn^{4+} is very polarising. [1]

F^- is small and not easily polarised [1] so SnF_4 is mainly ionic.

I^- is large and is easily polarised to give a covalent compound SnI_4. [1]