

# Answers to Assessment Questions

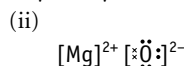
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- 1 (a) Electron [1], largest  $\frac{\text{charge}}{\text{mass}}$  ratio [1]
- (b) They have to be accelerated [1] and then deflected [1].
- (c) Suppose the fraction of peak 10 =  $x$  then fraction of peak 11 =  $1 - x$  and  $10x + 11(1 - x) = 10.8$  [1]  
 $\Rightarrow -x = -11 + 10.8$   
 $\therefore x = 0.2$  and  $y = 0.8$   
 $\therefore$  ratio of heights = 1 : 4 [1]
- (d) (i) Ratio of masses  
 B : H = 81.2 : 18.8 [1]  
 Ratio of moles  
 $B : H = \frac{81.2}{10.8} : \frac{18.8}{1}$  [1]  
 $= 7.52 : 18.8$   
 $= 1 : 2.5$   
 $= 2 : 5$   
 $\therefore$  empirical formula =  $B_2H_5$  [1]
- (ii)  $M_r(B_2H_5) = 26.6$   
 $\therefore$  molecular formula =  $B_4H_{10}$  [1]
- 2 (a)  $^{12}C$  is chosen as the standard [1] and given a relative atomic mass of exactly 12.
- (b) Mass of 1 mole of  $^1H = 1.0078$  g  
 Mass of 1 mole of  $^1H^+$   
 $= 1.0078 - (9.1091 \times 10^{-28})$   
 $\times 6.0225 \times 10^{23}$  [1]  
 $= 1.0078 - (5.4860 \times 10^{-4})$   
 $= 1.0078 - 0.0005$   
 $= 1.0073$  g [1]

- (c) (i) Sample is first vaporised by heating. [1] It is then ionised by bombardment with electrons. [1] This knocks out an electron to form a positive ion. [1] (any 2 for [2])
- (ii) To prevent formation of double positive ions ( $M^{2+}$ ) [1] rather than single positive ions (or to prevent fragmentation of molecular substances).
- (iii) Acceleration [1] by attraction through an electric field [1].
- (d) The mass of the electron lost from  $^6Li$  to form  $^6Li^+$  is negligible. [1]
- 3 (i) • Burn/heat  $[\frac{1}{2}]$  magnesium in air/oxygen  $[\frac{1}{2}]$   
 $2Mg + O_2 \rightarrow 2MgO$  [1]  
 • Heat  $[\frac{1}{2}]$  magnesium carbonate  $[\frac{1}{2}]$   
 $MgCO_3 \rightarrow MgO + CO_2$  [1]  
 • Heat  $[\frac{1}{2}]$  magnesium hydroxide  $[\frac{1}{2}]$   
 $Mg(OH)_2 \rightarrow MgO + H_2O$  [1]  
 • Heat magnesium  $[\frac{1}{2}]$  in steam  $[\frac{1}{2}]$   
 $Mg + H_2O \rightarrow MgO + H_2$  [1]  
 • Heat  $[\frac{1}{2}]$  magnesium nitrate  $[\frac{1}{2}]$   
 $Mg(NO_3)_2 \rightarrow$   
 $MgO + 2NO_2 + \frac{1}{2}O_2$  [1]

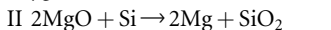
Any 2 ways  $\times$  [2]



(2 electrons transferred from Mg to O [1], charges on ions [1])

(iii) It is an oxide which acts as a base  $[\frac{1}{2}]$ , reacting with acids to form salts  $[\frac{1}{2}]$

(iv) I The silicon and the magnesium produced would react with oxygen in the air. [1]



2 moles MgO requires 1 mole of Si for reaction

$\Rightarrow 80.6$  g MgO requires 28.1 g of Si for reaction [1]

$\Rightarrow 500$  kg MgO requires

$\frac{28.1}{80.6} \times 500$  kg [1]

$= 174.3$  kg [1]

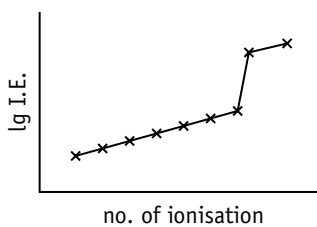
- 4 (a) (i)  $24 \text{ cm}^3 \text{ CO}_2 = \frac{24}{24\,000}$  moles  
 $= 0.001$  mole  $\text{CO}_2$  [1]
- (ii) Mass  $\text{CO}_2$  produced =  $0.001 \times 44$   
 $= 0.044$  g
- (b) (i) To ensure it has decomposed fully. [1]
- (ii) Total mass lost =  $0.240 - 0.160$   
 $= 0.080$  g [1]  
 $\therefore$  mass lost due to water  
 $= 0.080 - 0.044 = 0.036$  g [1]
- (iii) Number of moles of water  
 $= \frac{0.036}{18} = 0.002$
- (iv) Number of moles of copper  
 $= \frac{0.127}{63.5} = 0.0020$
- (c) (i) Mole ratio  $\text{CO}_2 : \text{H}_2\text{O} : \text{Cu}$   
 $= 0.001 : 0.002 : 0.002$   
 $= 1 : 2 : 2$
- (ii) The suggested formula should give a mole ratio of  $\text{CO}_2 : \text{Cu}$  which is 1 : 1 and not the calculated value of 1 : 2. [1]
- (d) The mole ratio of 1 : 2 for  $\text{CO}_2 : \text{Cu}$  suggests two copper compounds. The second compound must contain the same amount of Cu as  $\text{CuCO}_3$  and it must decompose giving off water to leave only copper(II) oxide. [1] The second compound is probably  $\text{Cu(OH)}_2$ . [1] Malachite cannot be simply  $\text{CuCO}_3 \cdot \text{Cu(OH)}_2$  because the mole ratio for  $\text{CO}_2 : \text{H}_2\text{O} : \text{Cu}$  for this would be 1 : 1 : 2. Thus, the probable formula of malachite is  $\text{CuCO}_3 \cdot \text{Cu(OH)}_2 \cdot \text{H}_2\text{O}$ .
- 5 (a) (i) The number of protons in an atom. [1]
- (ii) The number of protons plus neutrons in an atom. [1]
- (iii) The mass of an atom or isotope relative to the mass of an atom of  $^{12}\text{C}$  = 12 [1]

(b) Relative amounts  $^{20}\text{Ne} : ^{22}\text{Ne} = 90 : 10$

$$\text{So, } A_r(\text{Ne}) = \frac{90}{100} \times 20 + \frac{10}{100} \times 22 \quad [1]$$

$$= 20.2 \quad [1]$$

- (c) (i)  $\alpha$ :  $^4_2\text{He}^{2+}$ /He nucleus  $[\frac{1}{2}]$ ,  
poor/low penetration  $[\frac{1}{2}]$   
 $\beta$ :  $e^-$ /electron  $[\frac{1}{2}]$ , medium  
penetration  $[\frac{1}{2}]$   
 $\gamma$ : electromagnetic radiation  $[\frac{1}{2}]$ ,  
strong/high penetration  $[\frac{1}{2}]$
- (ii) Time taken for the amount of an  
isotope (or for its rate of decay)  
to fall to half its value. [1]
- (iii) A radioactive source is positioned  
on one side of the metal sheet  
and a radiation detector (GM  
tube) on the other. [1] The level  
of radiation penetrating the  
metal sheet can be used to  
monitor its thickness. [1]
- (iv) Because the level of radiation  
emitted must remain constant [1]  
if the thickness monitoring is to  
remain consistent [1].
- (d) (i) The energy change to convert 1  
mole of gaseous  $\text{F}^+$  ions [1] to 1  
mole of gaseous  $\text{F}^{2+}$  ions [1].



[2]

As electrons are removed, the  
I.E.s get more  
endothermic/larger. [1]  
Electrons at roughly same  
distance from nucleus/in same  
shell show relatively small  
increase from one to the next. [1]  
When a new shell starts, the jump  
in I.E. is large. [1]  
F has 7 electrons in a shell further  
from the nucleus (more easily  
removed) and 2 electrons in a  
shell close to the nucleus. [1]  
(max. [4])

- 6 (i)  $^{90}_{38}\text{Sr}$  [1]  
(ii)  $^{24}_{12}\text{Mg}$  and  $^{90}_{38}\text{Sr}$  [2]  
(iii)  $^{131}_{53}\text{I}$  [1]
- 7 (a) (i) A negatively charged particle [1]  
with a mass of approx. 1/2000th  
of a hydrogen-1 atom [1]

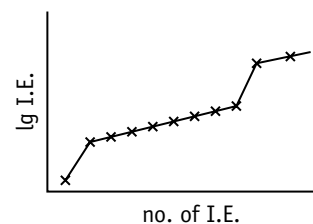
occupying the outer regions  
(parts) of all atoms [1].

- (ii) Electrons in atoms can only have  
certain (discrete/fixed) amounts  
of energy. [1] Groups of electrons  
can exist with roughly the same  
amount of energy. [1] These  
positions of roughly similar  
amounts of energy are called  
energy levels. [1] (max. [2])
- (b) X is  $\text{V}^{3+}$  [1]; Y is Na [1]; Z is  $\text{Cl}^-$  [1]
- (c) (i) Y [1]; (ii) X [1];  
(iii) Z [1]; (iv) Y [1]
- 8 (a) (i)  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$  [1]  
(ii) 20 protons [1]; 25 neutrons [1]  
(iii) Chemists are interested in  
chemical reactions which involve  
outer electrons/electron  
structures [1].
- (b) (i)  $^{45}_{21}\text{Sc} + {}^1_0\text{n} \rightarrow ^{45}_{20}\text{Ca} + {}^1_1\text{p}$  [1]  
(ii)  $6 \times 10^{-3} \text{ g} \rightarrow 7.5 \times 10^{-4} \text{ g}$   
involves 3 half-lives [1]  
 $\therefore$  half-life of  $^{45}_{20}\text{Ca}$   
= 180 days [1]  
(iii)  $^{45}_{20}\text{Ca} \rightarrow ^{45}_{21}\text{Sc} + {}^0_{-1}\text{e}$   
The radiation emitted would be  
electrons. [1] These would be  
attracted/deflected towards the  
positive terminal. [1]

- 9 (a) (i)  $\text{O}(\text{g}) \rightarrow \text{O}^+(\text{g}) + e^-$  [1]  
(ii) Helium has electrons in only the  
first shell and a larger nuclear  
charge than H. Therefore, He has  
the smallest atomic  
radius/electrons closest to the  
nucleus [1] and no shielding of  
these electrons from the nuclear  
charge [1].  
(iii) O has the electron structure  
 $1s^2 2s^2 2p^4$  and N has the electron  
structure  $1s^2 2s^2 2p^3$ . [1] In N the  
2p sub-shell is half-filled and the  
symmetrical distribution of charge  
confers extra stability [1] making  
it more difficult to remove an  
electron from N than O.  
(iv) As atomic number increases, each  
Group 1 metal has an additional  
shell of electrons, so the  
outermost electron is further  
from the nucleus [1] and there is  
more shielding of the outermost  
electron from the nuclear  
charge [1].
- (b) (i)  $\text{O}(\text{g}) + e^- \rightarrow \text{O}^-(\text{g})$  [1]  
 $\text{O}^-(\text{g}) + e^- \rightarrow \text{O}^{2-}(\text{g})$   
(ii) The first electron is captured by  
 $\text{O}(\text{g})$  atoms readily due to the

overall nuclear attraction. [1] The  
second electron is added only  
endothermically because the  
incoming electron is repelled by  
 $\text{O}^-(\text{g})$ . [1]

- (c) Although the formation of both  
 $\text{Mg}^{2+}(\text{g})$  and  $\text{O}^{2-}(\text{g})$  are endothermic  
processes, the  $\text{Mg}^{2+}(\text{g})$  and  $\text{O}^{2-}(\text{g})$   
ions then combine to form  
 $\text{Mg}^{2+}\text{O}^{2-}(\text{s})$ . [1] This process (lattice  
energy) is very exothermic [1] and the  
overall reaction is exothermic.
- (d) Successive ionisation energies of an  
element can give evidence for shells of  
electrons. When the I.E.s or  $\lg(\text{I.E.s})$   
are plotted against the number of the  
ionisation, points at similar energy  
levels indicate shells of electrons with  
similar energies. [1] E.g. for Na:



[1]

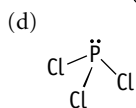
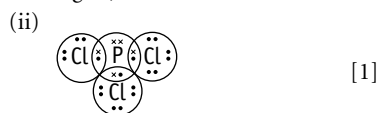
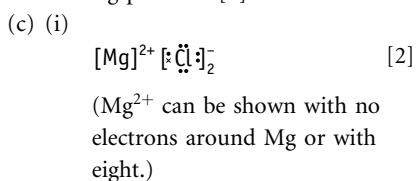
The plot for Na indicates a shell  
structure of 2, 8, 1 [1] (max. [2])

- 10 (a) The energy required to remove one  
electron from each atom [1] of 1 mole  
of gaseous atoms of the element [1]  
under standard conditions.  
(b) Neon has a full shell of electrons  
symmetrically disposed which make  
the atom extra stable. [1] It has a  
greater nuclear charge than F [1] and  
so holds its outer electrons in the same  
shell more strongly [1]. (any 2 for [2])  
(c) N has the electron structure  $1s^2 2s^2 2p^3$ ,  
whereas O has the structure  $1s^2 2s^2 2p^4$ .  
The structure in N with a half-filled 2p  
sub-shell is symmetrical and more  
stable than that in O. [1] Two of the  
outer electrons in the 2p sub-shell of O  
are paired. [1] The repulsion between  
this pair makes it easier to remove one  
of them. [1] (any 2 for [2])  
(d) The outermost electron in Na is in a  
new 3rd shell. [1] This is further from  
the nucleus [1] and is also more  
shielded [1] so is easier to remove.  
(any 2 for [2])  
(e) The electron structure in C is  
 $1s^2 2s^2 2p^2$ . The electron structure in N  
is  $1s^2 2s^2 2p^3$ . [1] The outer 2p electrons

will be attracted by 7 protons in N, but 6 protons in C. (fewer protons in C) [1] Shielding by inner electrons will be similar for C and N. [1] Predict 1st I.E. for C =  $\frac{6}{7} \times$  1st I.E. for N = 1200 kJ mol<sup>-1</sup> [1] (any 3 for [3])

- 11 (a) (i) From neutral to acidic (pH decreases) [1]  
 (ii)  $\text{SiCl}_4 + 4\text{H}_2\text{O} \rightarrow \text{Si(OH)}_4 + 4\text{HCl}$   
 or  $\text{SiCl}_4 + 2\text{H}_2\text{O} \rightarrow \text{SiO}_2 + 4\text{HCl}$   
 ([1] for correct products, [1] for balanced equation)

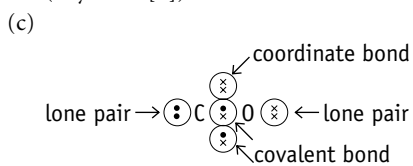
(b) Alternately sodium ions and chloride ions throughout the structure with no missing positions [2]



(pyramidal w.r.t. atoms [1], bond angles 100–108° [1])

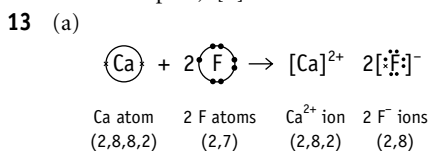
- (e) (i) tetrahedral [1]  
 (ii) octahedral [1]

- 12 (a) <sup>14</sup>N has 7p, 7n and 7e. [1] <sup>15</sup>N has 7p, 8n and 7e. [1]  
 (b) Volatile (low m.pt., low b.pt.), non-conductor (poor conductor) of electricity, soluble in non-polar solvents (insoluble in polar solvents). (any 2 for [2])

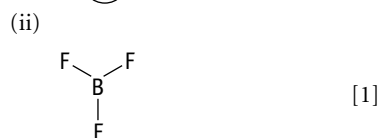


([1] for each of 3 correct labels)

- (d) (i) Same number of electrons/same electronic structure [1]  
 (ii)  $\text{H} \text{ : } \text{C} \text{ : } \text{N} \text{ :}$  [2]  
 (iii) CO and HCN are polar (have a dipole). N<sub>2</sub> is non-polar (has no dipole). [1]



Each Ca atom contains 2 electrons in its outermost shell, electron structure 2,8,8,2. [1] These 2 outer electrons are transferred to two F atoms leaving a Ca<sup>2+</sup> ion (2,8,8) [1]. The two fluorine atoms each have seven electrons in their outer shell, electron structure 2,7 [1]. Each F atom gains one electron to form a F<sup>-</sup> ion (2,8). [1] (Marks can be obtained for a diagram or for text)



The three pairs of bonded electrons repel each other equally [1] and move as far apart as possible.

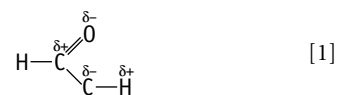
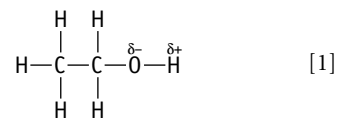
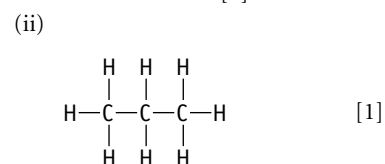
- (iii) Dipoles/polar bonds [1] cancel out [1] their effects *or* centres of + and - charge [1] in the molecule coincide [1].  
 (iv) Co-ordinate or dative covalent. [1]

- 14 (a) (i) A bond formed by the sharing of a pair of electrons [1], one electron being contributed by each atom to the shared pair [1].  
 (ii) By Van der Waals forces (induced dipole attractions). [1]  
 (b) It has a higher M<sub>r</sub>. [1] It also has more electrons and can therefore form induced dipoles more easily. [1] This results in stronger Van der Waals forces. [1] (any 2 for [2])

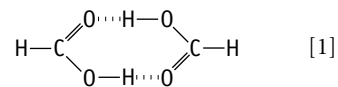
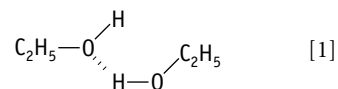
- (c) (i) The power of an atom in a covalent bond [1] to attract electrons [1].  
 (ii) HCl contains two atoms with an electronegativity difference of 0.9. [1] It therefore has a significant permanent dipole. [1] CH<sub>4</sub> has four H atoms distributed symmetrically around a C atom. It has no permanent dipole. [1] The intermolecular (dipole-dipole) attractions in HCl are much stronger than the induced dipole attractions in CH<sub>4</sub> (Van der Waals forces). [1] Hence, the b.pt. of HCl is higher.

- 15 (a) (i) A bond in which the centres of positive and negative charges do not coincide/a bond in which there is an unequal distribution of charge. [1]

The atoms of the bond have different electronegativities and their nuclei attract electrons to a different extent. [1]



- (b) (i) Induced (temporary) dipole attractions/Van der Waals forces. [1] These forces are very weak. [1]  
 (ii) Hydrogen bonds (H bonds) [1]



(or with just one H-bond)

Methanoic acid is more polar/possibility of two H bonds between 2 molecules/overall stronger H bonding. (any 1 for [1])

- 16 (a) The power of an atom/element in a covalent bond [1] to attract electrons [1].  
 (b) (i) B-H, C-H, N-H, F-H [1]  
 (ii) The greater the difference in electronegativity of the atoms in the bond [1], the greater the ionic character of the bond [1].  
 (c) (i) H-bonding is shown by compounds containing either N-H, O-H or F-H bonds. [1] These three bonds have hydrogen attached to the three most electronegative

elements. [1] This creates a bond with high polarity. [1]



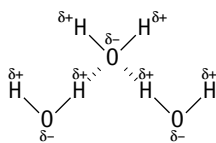
Electrons in the bonds are attracted towards the more electronegative oxygen, leaving an H atom without any surrounding electrons (effectively a bare proton). [1] This creates a more polar bond (greater dipole) than any other and the extra-strong dipole attractions are called hydrogen bonds. [1] (max. [4])

(ii)  $\text{NH}_3$ ,  $\text{CH}_3\text{OH}$  and  $\text{D}_2\text{O}$  ([2] for all correct, [1] for 2 correct)

17 (a) The power of an atom in a covalent bond [1] to attract the electrons in the bond [1].

(b) (i) The molecules increase in  $M_r$  from  $\text{H}_2\text{S}$  to  $\text{H}_2\text{Te}$ . [1] They, therefore, have increasing numbers of electrons [1] and induced dipoles form more easily [1]. These induced dipoles constitute the Van der Waals forces between molecules and the stronger Van der Waals forces require higher temperatures to separate (boil) the molecules. [1] (any 3 for [3])

(ii) O is much more electronegative than S, Se and Te. [1] Thus, the H—O bonds in water are much more polar than in other Group VI hydrides. [1] This greater polarity results in extra-strong dipole–dipole attractions in  $\text{H}_2\text{O}$  called hydrogen bonds. [1] These extra-strong intermolecular forces cause the anomalously high boiling point for  $\text{H}_2\text{O}$ . [1]



Correct showing of H bonds between  $\text{H}_2\text{O}$ . [1] Indication that  $\text{H}_2\text{O}$  molecules can have two H bonds per molecule. [1] (max. [5])

(c) 3D structures of proteins are held in place largely by H bonds [1] between H of N—H groups [1] and the O of C=O groups [1]. In addition, there

are bridges (bonds) between two sulphur atoms (disulphide bridges) [1] and links (H bonds or ionic attractions) between some of the side-chain R groups [1]. (max. [3])

18 (a) (i) Type of bond – ionic compounds tend to dissolve, covalent compounds tend not to. [1]

Strength of bond – strong bonds between ions or between atoms will tend to reduce solubility. [1]

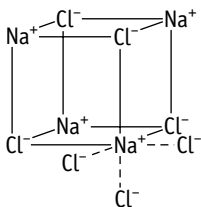
Polarity of bonds – the more polar the bonding, the more soluble the substance. [1]

Compounds with hydrogen bonding tend to be very soluble in water which also has hydrogen bonds. [1]

(any 2 for [2])

(ii) Ethanol has polar O—H bonds which can form H bonds with water molecules. [1] Hexane is a non-polar substance with weak Van der Waals bonds only. [1] It cannot form bonds with water molecules similar in strength to the H bonds between water molecules themselves. [1]

(b) (i)



(cubic structure  $[\frac{1}{2}]$ , ions alternating  $[\frac{1}{2}]$ )

Co-ordination number of a sodium ion = 6 [1]

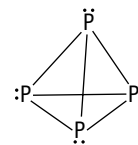
(ii)  $\text{Cs}^+$  is much larger than  $\text{Na}^+$  and can fit more  $\text{Cl}^-$  ions around it. [1]

(iii) Sodium chloride has strong ionic bonds between oppositely charged ions. [1] Iodine has only weak intermolecular/Van der Waals forces between  $\text{I}_2$  molecules. [1] The stronger forces in NaCl require a much higher temperature to separate the ions than the weak forces between  $\text{I}_2$  molecules. [1]

19 (a) (i) P has only one isotope. [1]  $\text{P}_4^+$  (allow  $\text{P}_4$ ) [1]

(ii) White phosphorus contains  $\text{P}_4$  molecules. [1] Four P atoms are

joined together by covalent bonds. [1]



[1]

The covalent bonds within the  $\text{P}_4$  molecules are strong but there are only weak (Van der Waals) forces between the separate  $\text{P}_4$  molecules. [1]

(iii) Silicon has a giant molecular structure. [1] Each Si atom is joined by covalent bonds [1] to four other Si atoms [1] in a giant molecule. A large amount of energy is required to break the covalent bonds. [1]

(b) (i) Ratio of masses  
P : O = 56.4 : 43.6 [1]

Ratio of moles

$$\text{P : O} = \frac{56.4}{31} : \frac{43.6}{16} [1]$$

$$= 1.82 : 2.73$$

$$= 2 : 3$$

$\therefore$  empirical formula of A  
=  $\text{P}_2\text{O}_3$  [1]

(ii)  $M_r$  [1]

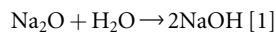
(iii)  $\text{P}_4$  (or  $4\text{P}$ ) +  $5\text{O}_2 \rightarrow \text{P}_4\text{O}_{10}$  [1]  
124 g P  $\rightarrow$  284 g  $\text{P}_4\text{O}_{10}$

$$\therefore 3.2 \text{ g P} \rightarrow \frac{284}{124} \times 3.2 \text{ g } \text{P}_4\text{O}_{10}$$

$$= 7.33 \text{ g [1]}$$

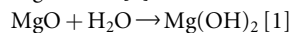
(iv)  $\text{P}_4\text{O}_{10} + 6\text{H}_2\text{O} \rightarrow 4\text{H}_3\text{PO}_4$  [1]  
 $\text{H}_3\text{PO}_4$  (phosphoric acid) is a (strong) acid. [1]

(c)  $\text{Na}_2\text{O}$  ionic [1]



high pH (11–15) [1]

$\text{MgO}$  ionic [1]



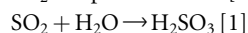
high pH (8–11) [1]

$\text{SiO}_2$  giant molecular [1]

insoluble [1]

no effect on pH (7) [1]

$\text{SO}_2$  simple molecular [1]



low pH (3–5) [1]

20 (a) Enthalpy change when the molar quantities as shown in the neutralisation equation [1] react under standard conditions [1].

(b) Diagram should include:  
burette or pipette to add NaOH;  
polystyrene calorimeter or lagged

calorimeter; lid; accurate thermometer. (all 4 items [2], 2 or 3 items [1])

(c) Points plotted correctly [1], two intersecting straight lines [1].

(d) (i)  $(37 \pm 1) \text{ cm}^3$  [1]

(ii)  $6.6^\circ\text{C}$  [1]

(e) (i) Enthalpy change

$$= m \times c \times \Delta T [1]$$

$$= (62 \pm 1) \times 4.2 \times 6.6$$

$$= (1719 \pm 28) \text{ J} [1]$$

(ii) Number of moles of  $\text{H}_2\text{SO}_4$

$$= \frac{25}{1000} \times 0.6 = 0.015 [1]$$

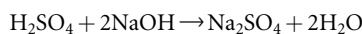
$$0.015 \text{ moles } \text{H}_2\text{SO}_4 \rightarrow 1720 \text{ J}$$

$\therefore$

$$1 \text{ mole } \text{H}_2\text{SO}_4 \rightarrow 1720 \times \frac{1}{0.015}$$

$$= 115\,000 \text{ J} [1]$$

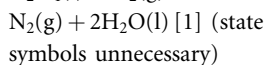
$\therefore$  for the reaction



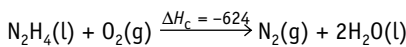
$$\Delta H_{\text{neutr}} = -115 \text{ kJ (sign [1], value [1])}$$

21 (a) Bonds between molecules in ethane involve only Van der Waals (induced dipole) forces. [1] Those between molecules in hydrazine involve much stronger hydrogen bonds. [1]

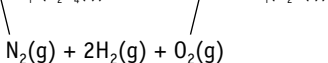
(b) (i)  $\text{N}_2\text{H}_4(\text{l}) + \text{O}_2(\text{g}) \rightarrow$



(ii)



$$\Delta H_f(\text{N}_2\text{H}_4(\text{l}))$$



$$\Delta H_f(\text{N}_2\text{H}_4(\text{l}))$$

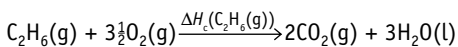
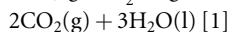
$$= 2 \times \Delta H_f(\text{H}_2\text{O}(\text{l})) - \Delta H_c$$

$$= 2(-286) - (-624)$$

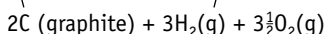
$$= +52 \text{ kJ mol}^{-1} \text{ (correct}$$

numerical answer [1], sign and units [1], evidence of correct calculation [1])

(c) (i)  $\text{C}_2\text{H}_6(\text{g}) + 3\frac{1}{2}\text{O}_2(\text{g}) \rightarrow$



$$\Delta H_f(\text{C}_2\text{H}_6(\text{g}))$$



$$2\Delta H_f(\text{CO}_2(\text{g})) + 3\Delta H_f(\text{H}_2\text{O}(\text{l}))$$

$$\Delta H_c(\text{C}_2\text{H}_6(\text{g}))$$

$$= -\Delta H_f(\text{C}_2\text{H}_6(\text{g}))$$

$$+ 2\Delta H_f(\text{CO}_2(\text{g}))$$

$$+ 3\Delta H_f(\text{H}_2\text{O}(\text{l}))$$

$$= -(-85) + 2(-394) + 3(-286)$$

$$= -1561 \text{ kJ mol}^{-1} \text{ (correct}$$

numerical value [1], sign and units [1], evidence of correct calculation [1])

(d) It can be stored more easily as a

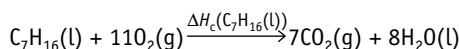
liquid/it requires less oxygen per unit mass. (any 1 for [1])

22 (a) (i) Enthalpy change when 1 mole of a substance [1] is completely burnt (combusted) in air or  $\text{O}_2$  [1] under standard conditions.

(ii) 298 K ( $25^\circ\text{C}$ ) and 1 atm

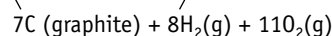
(101 kPa/100 kPa) [1]

(b)



$$\Delta H_f(\text{C}_7\text{H}_{16}(\text{l}))$$

$$7\Delta H_f(\text{CO}_2(\text{g})) + 8\Delta H_f(\text{H}_2\text{O}(\text{l}))$$



$$\Delta H_c(\text{C}_7\text{H}_{16}(\text{l}))$$

$$= -(-224.4) + 7(-393.5)$$

$$+ 8(-285.9)$$

$$= +224.4 - 2754.5 - 2287.2$$

$$= -4817.3 \text{ kJ mol}^{-1} \text{ (correct}$$

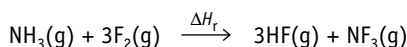
numerical answer [1], sign and units [1], evidence of correct calculation [1])

(c)  $\text{C}_7\text{H}_{16} \rightarrow \text{C}_5\text{H}_{12} + \text{C}_2\text{H}_4$  [2] (similar other equations are possible)

23 (a) Enthalpy change when 1 mole of a substance is formed from its elements [1] in their normal physical states [1] under standard conditions [1].

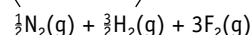
(b) If reactants A and B are converted to products C and D via different intermediates, the enthalpy change is the same whichever route (path) is taken. [1]

(c) (i)



$$\Delta H_f(\text{NH}_3(\text{g}))$$

$$3 \times \Delta H_f(\text{HF}(\text{g})) + \Delta H_f(\text{NF}_3(\text{g}))$$



$\Delta H_r$  (enthalpy change for reaction)

$$= -\Delta H_f(\text{NH}_3(\text{g}))$$

$$+ 3\Delta H_f(\text{HF}(\text{g}))$$

$$+ \Delta H_f(\text{NF}_3(\text{g})) [2]$$

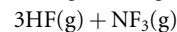
$$= -(-46) + 3(-269) + (-114)$$

$$= 46 - 807 - 114$$

$$= -875 \text{ kJ mol}^{-1} \text{ (correct}$$

numerical value [1], sign and units [1])

(ii)  $\text{NH}_3(\text{g}) + 3\text{F}_2(\text{g}) \rightarrow$



*Bonds broken*

$$3 \times \text{N—H} = +3 \times 388$$

$$3 \times \text{F—F} = +3 \times 158$$

*Bonds formed*

$$3 \times \text{H—F} = -3 \times 562$$

$$3 \times \text{N—F} = -3 \times 272$$

$$\therefore \Delta H_r = +3 \times 388 + 3 \times 158$$

$$- 3 \times 562 - 3 \times 272 [2]$$

$$= 1164 + 474 - 1686 - 816$$

$$= -864 \text{ kJ mol}^{-1} \text{ (correct}$$

numerical value [1], sign and units [1], max. [3])

(d) The bond enthalpies in (c) (ii) are average values [1] in the case of N—H and N—F and are not specific to  $\text{NH}_3$  or  $\text{NF}_3$  [1]. The  $\Delta H_f$  values, on the other hand, relate to the specific substances concerned. [1]

24 (a) (i)



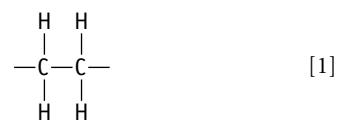
or  $\text{CH}_3\text{CH}_2^+$  [1]

(ii) The lone (non-bonded) pairs of electrons on the water molecule. [1]

(iii) The rate of production of ethanol will increase [1] but the amount of ethanol at equilibrium will decrease [1].

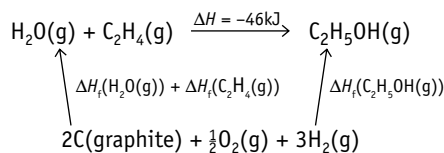
(iv) By Le Chatelier's principle, an increase in pressure will produce more ethanol at equilibrium [1] as 2 moles of gas form 1 mole of product gas in the equation [1].

(v) poly(ethene)/polythene [1]



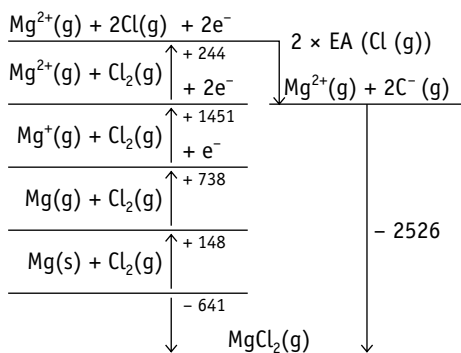
(b) The higher the pressure, the more costly the equipment to withstand the pressure. [1]

(c)



$$\begin{aligned} \Delta H_f(\text{C}_2\text{H}_5\text{OH}(\text{g})) &= \Delta H_f(\text{H}_2\text{O}(\text{g})) + \Delta H_f(\text{C}_2\text{H}_4(\text{g})) \\ &\quad + \Delta H [1] \\ &= -242 + 52.3 - 46 \\ &= -235.7 \text{ (-236) kJ [1]} \end{aligned}$$

25 (a) (i)



(each stage of cycle correct with state symbols =  $\frac{1}{2}$ ];  $6 \times \frac{1}{2}$ ] = [3]  
E.A. =  $\frac{1}{2}(-244 - 1451 - 738 - 148 - 641 + 2526)$  [1]  
=  $\frac{1}{2} \times -696$   
=  $-348 \text{ kJ mol}^{-1}$  [1]

(ii) The theoretical value assumes a perfect ionic structure [1] in which the ions are discrete spherical particles [1] with no polarisation causing partial covalent bonding [1]. In fact, the  $\text{Mg}^{2+}$  ion will cause polarisation of the  $\text{Cl}^-$  ion. [1] (max. [3])

(b) (i) As the cations get larger,  $\text{Ca}^{2+} \rightarrow \text{Sr}^{2+} \rightarrow \text{Ba}^{2+}$ , their charge density decreases. [1] This results in weaker attraction for the polar water molecules [1] and thus hydration enthalpies which are less exothermic.

(ii) Solubility decreases from  $\text{CaSO}_4$  to  $\text{BaSO}_4$ . [1]  
Solubility is related to  $\Delta H_{\text{soln}} = -\text{lattice enthalpy} + \Delta H_{\text{hyd}}$  [1]  
 $\Delta H_{\text{soln}}$  cannot be calculated fully because  $\Delta H_{\text{hyd}}(\text{SO}_4^{2-})$  is not provided [1] but:  $\Delta H_{\text{soln}}$  is

becoming more endothermic ( $830 + \Delta H_{\text{hyd}}(\text{SO}_4^{2-})$ ;  $1004 + \Delta H_{\text{hyd}}(\text{SO}_4^{2-})$ ;  $1014 + \Delta H_{\text{hyd}}(\text{SO}_4^{2-})$ ) [1] and as a result solubility is decreasing [1]. (max. [4])

(iii)  $\text{BaSO}_4$  is very insoluble. [1]

(iv)  $\text{Ba}(\text{s}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{Ba}(\text{OH})_2(\text{aq}) + \text{H}_2(\text{g})$  (correct products [1], balanced equation [1])

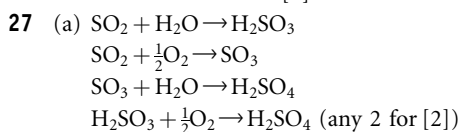
(v) Add dilute  $\text{H}_2\text{SO}_4$  or a soluble sulphate to the solution. [1] Filter off the  $\text{BaSO}_4$  ppt. [2] Wash with water and dry. [1]

26 (a)  $\Delta S^\ominus = S^\ominus(\text{products}) - S^\ominus(\text{reactants})$  [1]  
=  $3 \times 134 + 83 - 4 \times 112$   
=  $+37 \text{ J K}^{-1} \text{ mol}^{-1}$  [1]

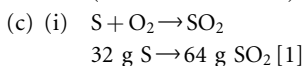
(b)  $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus$  [1]  
=  $16.8 - 298 \times \frac{37}{1000}$   
=  $16.8 - 11.0$   
=  $+5.8 \text{ kJ mol}^{-1}$  [1] for numerical value, [1] for correct sign and units.

(c) (i)  $\Delta G$  is negative ( $\Delta G \leq 0$ ) [1]

(ii)  $\Delta G = \Delta H - T\Delta S$   
The reaction is just feasible when  $\Delta G = 0$   
 $\therefore 0 = \Delta H - T\Delta S$  [1]  
 $\Rightarrow T = \frac{\Delta H}{\Delta S}$   
=  $\frac{16.8 \times 1000}{37}$  [1]  
=  $454 \text{ K}$  [1]



(b) (i) Damages leaves/lowers pH of soil/damages roots (any 2 for [2])  
(ii) Chalk will react with the acids (neutralise the acids). [1]



$$\Rightarrow 10^6 \times \frac{1.92}{100} \text{ g S} \rightarrow 10^6 \times \frac{1.92}{100} \times \frac{64}{32} \text{ g SO}_2$$

$$= 38\,400 \text{ g (38.4 kg) SO}_2$$

(ii) Number of moles of  $\text{SO}_2 = \frac{38\,400}{64} = 600$  [1]  
Volume of  $\text{SO}_2 = 600 \times 24 = 14\,400 \text{ dm}^3$  [1]

(iii) Reduce the use of fossil fuels (use alternative power supplies).  
Remove S or sulphides from fuel before use (flotation to remove pyrites from coal).

Trap  $\text{SO}_2$  using basic materials (flue gas desulphurisation). (any 2 for [2])

28 (a) (i) Rate of reaction is the change in concentration or amount of a given reactant or product per unit time. [1]

(ii) Overall order of reaction is the sum of the powers [1] of the reactants' concentrations [1] in the rate equation.

For the rate equation shown, overall order =  $m + n$ . [1]

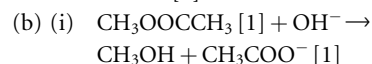
(b) (i) Order with respect to A = 1. [1]  
Order with respect to B = 2. [1]

(ii) Rate =  $k[\text{A}]^1[\text{B}]^2$   
Using data from the first experiment,  
 $1.23 \times 10^{-3} = k(0.1)(0.01)^2$   
 $\therefore k = 1.23 \times 10^2$  [1]  
( $\text{mol dm}^{-3}$ ) $^{-2} \text{ s}^{-1}$  or  $\text{mol}^{-2} \text{ dm}^6 \text{ s}^{-1}$  [1]

(c) (i) Increase in temperature causes the reaction rate to increase [1] because molecules/particles will move faster at higher temperatures [1], giving more collisions per second [1] and also giving a greater fraction of successful collisions as the particles collide with greater energy [1].

(ii) A (positive) catalyst will increase the rate of reaction [1] because it provides an alternative pathway/route for the reaction mechanism [1] with a lower activation energy [1] which gives more molecules/collisions sufficient energy for reaction [1].

29 (a) Order with respect to C = 2 [1]  
Order with respect to D = 0 [1]  
Initial rate in experiment 4 =  $1.6 \times 10^{-5}$  [1]  $\text{mol dm}^{-3} \text{ s}^{-1}$



(ii) Rate =  $k[\text{CH}_3\text{OOCCH}_3][\text{OH}^-]$  [1]  
 $8.8 \times 10^{-5} = k \times 0.02 \times 0.02$

$$k = \frac{8.8 \times 10^{-5}}{4 \times 10^{-4}} [1]$$

$$= 2.2 \times 10^{-1} [1] \text{ mol}^{-1} \text{ dm}^3 \text{ s}^{-1} [1]$$