Section A

Question	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Answer	B	C	B	B	D	C	B	C	C	D	A	A	D	A	D

MCQ worked solutions

1 (Ans: B)

From the mass numbers given in the table, the sample contains ^{10}B, ^{11}B and ^{23}Na (^{23}Na is the only impurity present). To compute the A_r of boron in the given sample, we need to consider only the isotopes, ^{10}B and ^{11}B, and their % abundance (i.e. ^{23}Na and its % abundance should not be included in the calculation). Note also that the sum of % abundance of ^{10}B and ^{11}B is not 100%.

A_r of boron in sample = $\dfrac{(10)(15.52)+(11)(74.48)}{15.52+74.48} = 10.83$

2 (Ans: C)

In the redox reaction, V_2O_5 is reduced and Fe^{2+} is oxidised to Fe^{3+}.

$n_{V_2O_5}$	:	n_{e^-}	:	$n_{Fe^{2+}}$
	:	1	:	1
1	:		:	2
1	:	2	:	2

Since
Oxidation: $Fe^{2+} \longrightarrow Fe^{3+} + e^-$
or $2Fe^{2+} \longrightarrow 2Fe^{3+} + 2e^-$
Reduction (unbalanced): $V_2O_5 + 2e^- \longrightarrow ?$
$\Rightarrow$ Each V atom accepted 1 e^-.
$\Rightarrow$ Oxidation no. of V decreases from +5 to +4.

Oxidation number of V — *0* — *+2* — *+4* — *+3*

A V **B** VO **C** VO_2 **D** V_2O_3

3 (Ans: B)

H^+, being positively charged, gives θ of +15° and $\left|\frac{q}{m}\right| = 1$

Since particle **Z** gives θ of –5°, particle **Z** is negatively charged and $\left|\frac{q}{m}\right| = \frac{1}{3}$

	proton	neutron	electron	q	m	$\left\|\frac{q}{m}\right\|$
1	1	0	4	–3	1	3
2	1	2	2	–1	3	$\frac{1}{3}$
3	3	3	1	+2	6	$\frac{1}{3}$
4	3	3	5	–2	6	$\frac{1}{3}$

Hence, only options 2 and 4 are correct.

4 (Ans: B)

There is a large jump in the 4th to 5th IE for one element, and in the 6th to 7th IE for another element. Hence these two elements belong to group 14 and 16 respectively, with minimally 8 electrons.

CS_2 is not possible because C has only 6 electrons to be removed.
Hence only SiO_2 is possible.

5 (Ans: D)

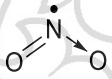
A (incorrect): **X** and **Y** do not have the same proton number and hence they are not isotopes.

B (incorrect): Electronegativity of an element increases across the period. Hence **Y** has a higher electronegativity than **X**.

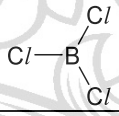
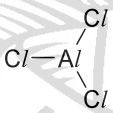
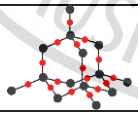
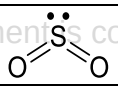
C (incorrect): Since **X** and **Y** have different proton number (and hence different no. of e^-), they will not have the same electronic configuration.

D (correct): As the ion of **Y** is isoelectronic with $\mathbf{X^+}$, the ion of **Y** has a +2 charge (i.e. $\mathbf{Y^{2+}}$). Hence, the ion of **Y** has a higher charge density (higher charge, smaller ionic radius) than the ion of **X**.

6 (Ans: C)

molecule	structure	no. of π bonds	no. of σ bonds
CO_2	O=C=O	2	2
HCN	H—C≡N	2	2
N_2	N≡N	2	1
NO_2		1	2 (Note: the dative covalent bond is also a σ bond)

7 (Ans: B)

		around underlined atom		
molecule	structure	no. of bond pairs	no. of lone pairs	shape
$\underline{B}Cl_3$		3	0	trigonal planar
$\underline{Al}Cl_3$		3	0	trigonal planar
$\underline{C}O_2$	O=C=O	2	0	linear
$\underline{Si}O_2$		4	0	tetrahedral
$H_2\underline{O}$		2	2	bent
$\underline{S}O_2$		2	1	bent

8 (Ans: C)

A (incorrect): Polar compounds may not always be soluble in water. If the energy released from the interactions formed between the compound and water is insufficient to compensate for the energy required to overcome the compound-compound and water-water interactions, then the compound may not be soluble. For example, chloroform, $CHCl_3$, is polar but it is insoluble in water.

B (incorrect): Compounds containing polar bonds can be non-polar if the dipole moments of the polar bonds cancel each other out, resulting in zero net dipole moment. CO_2 is one such example.

C (correct): When a charged rod is brought near a stream of a polar compound, the polar compound is attracted to the rod because of the dipoles present on the polar compound.

D (incorrect): All molecules experience id-id forces. Therefore, all polar compounds must experience both id-id and pd-pd forces. Also, some polar compounds are able to form intermolecular hydrogen bonding, in addition to id-id and pd-id interactions.

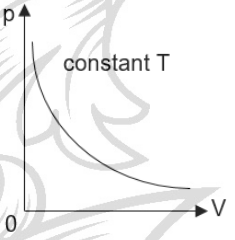
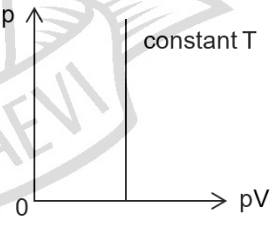
9 (Ans: C)

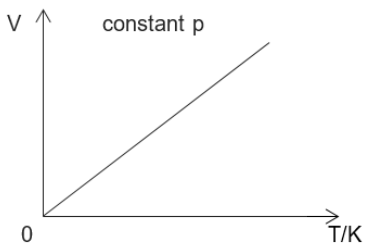
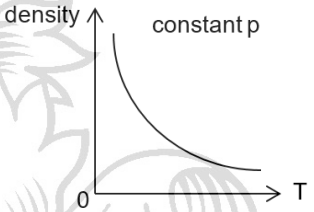
Statement 1 (incorrect): The compound cannot form hydrogen bonds with itself because it does not have any H atoms directly bonded to an F, O or N atom.

Statement 2 (correct): The compound is ionic so it can form ion-dipole interactions with water.

Statement 3 (correct): $|LE| \propto \left|\frac{q_+ q_-}{r_+ + r_-}\right|$. Since the cation has a small charge and a very large size, the magnitude of the lattice energy is small.

10 (Ans: D)

	manipulation of ideal gas equation	correct graph
A	pV = nRT since n, R and T are constants, pV = constant $p \propto \frac{1}{V}$ hence graph is similar to $y = \frac{k}{x}$	
B	pV = nRT Since n, R and T constants, pV = constant at a given T. Hence graph is similar to x = c (vertical straight line)	

C	pV = nRT Since n, R and p are constants, $V \propto T$ Hence, graph is similar to y = *k*x (straight line that passes through origin)	
D	$pV = nRT \Rightarrow pV = \frac{m}{M}RT$ $\Rightarrow p = \frac{m}{V} x \frac{RT}{M}$ = density x $\frac{RT}{M}$ $\Rightarrow$ density = $\frac{pM}{RT}$ since p, M and R are constants, density $\propto \frac{1}{T}$ hence graph is similar to y = $\frac{k}{x}$	

11 (Ans: A)

Assumption of an ideal gas:
1. the gas particles have <u>negligible volume</u> and
2. the gas particles exert <u>negligible attractive forces</u> on one another.

Neon behaves most ideally at rtp as the
- volume of neon gas particles is the smallest and
- intermolecular forces of attraction (id-id) in neon is the weakest
  (Intermolecular hydrogen bonding in NH_3 is stronger than the id-id interaction in Ne, Cl_2 and CO_2. Amongst Ne, Cl_2 and CO_2, Ne has the weakest id-id interaction as it has the smallest and least polarisable electron cloud.)

12 (Ans: A)

Statement 1 (incorrect): Standard enthalpy change of atomisation of $Br_2(l)$ refers to the energy absorbed to <u>form **one mole** of gaseous Br atoms</u> from $Br_2(l)$ at 298 K and 1 bar. In the equation, <u>two moles of Br(g) are formed</u>.

Statement 2 (correct): Standard enthalpy change of combustion of C(s) refers to the energy released when <u>one mole of C(s)</u> is burnt in excess oxygen at 298 K and 1 bar. In the equation, there is <u>one mole of C(s)</u> being burnt to give $CO_2(g)$. The formation of $CO_2(g)$ only indicates complete combustion. (Note: In incomplete combustion, CO(g) is formed as well).

Statement 3 (incorrect): Standard enthalpy change of formation of $H_2O(l)$ refers to the energy change when **<u>one mole</u>** <u>of $H_2O(l)$ is formed</u> from its constituent elements at 298 K and 1 bar. In the equation, there are <u>two moles of $H_2O(l)$ formed</u>.

13 (Ans: D)

Thermal decomposition of $MgCO_3$ is an endothermic process as heat needs to be supplied in order for the reaction to proceed. Hence, ΔH is positive.

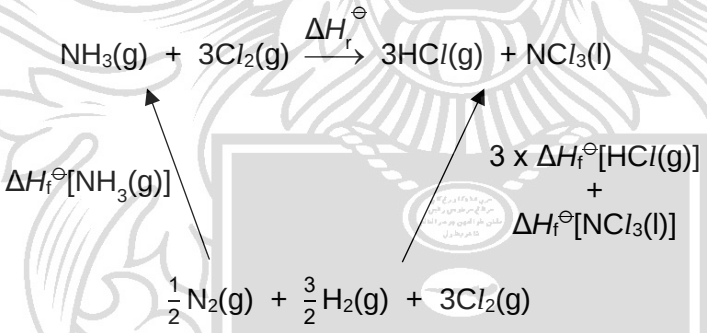
Since the decomposition happens only at high temperatures, it means that $\Delta G < 0$ only at high temperatures.

$\Delta G = \Delta H - T\Delta S$

Since $\Delta H > 0$, ΔS must be positive so that at high temperatures, $T\Delta S > \Delta H$, which will lead to $\Delta G < 0$.

Alternative explanation:
$MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$
Thermal decomposition of $MgCO_3$ is an endothermic process as heat needs to be supplied in order for the reaction to proceed. Hence, ΔH is positive.
Since there is an increase in the number of gas particles after reaction, entropy has increased and the system becomes more disordered. Hence, ΔS is positive.

14 (Ans: A)



By Hess' Law:
$\Delta H_r^\ominus = -\Delta H_f^\ominus[NH_3(g)] + 3 \times \Delta H_f^\ominus[HCl(g)] + \Delta H_f^\ominus[NCl_3(l)]$
$= -(-45.9) + 3(-92.3) + (+230.0)$
$= -1.0 \text{ kJ mol}^{-1}$

(Note: Take note of coefficients. Recall that standard enthalpy change of formation is for the formation of one mole of the compound.)

15 (Ans: D)

$n(CO_2) = 1.76 / (12.0 + 2 \times 16.0) = 0.04$ mol
$n(H_2O) = 0.72 / (2 \times 1.0 + 16.0) = 0.04$ mol
ratio of C : H in compound **P** = 0.04 : 2(0.04) = 1 : 2 $\Rightarrow$ **A** and **B** are incorrect

$$n(\text{O atoms}) \text{ in 1.20 g compound } \mathbf{P} = \frac{1.20 - (12.0)(0.04) - (1.0)(0.08)}{16.0}$$
$= 0.04$ mol $\Rightarrow$ **C** is incorrect

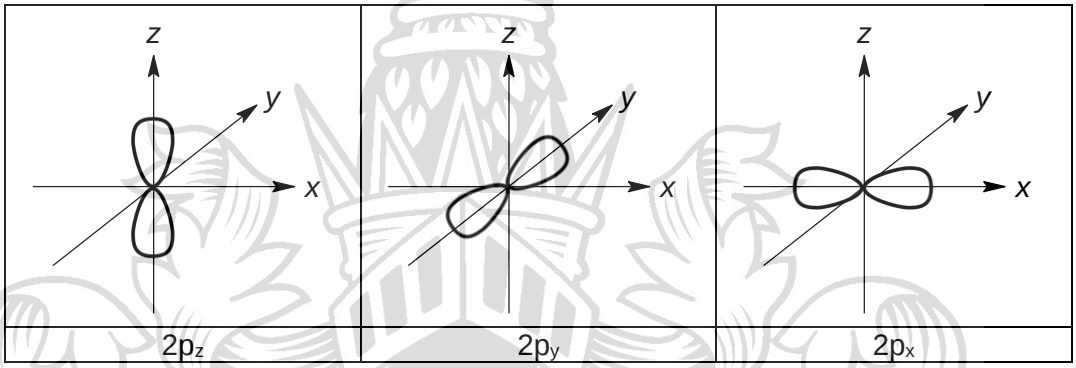
ratio of C : H : O in compound **P** = 0.04 : 2(0.04) : 0.04 = 1 : 2 : 1
$\therefore$ Empirical formula of **P** is CH_2O $\Rightarrow$ **D** is correct

Section B

B1(a)(i) Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

Comments:
- Generally well done. Do note that the number of electrons in each subshell should be written as a "superscript".

B1(a)(ii)



$2p_z$ $2p_y$ $2p_x$

Comments:
- Students should be specific and label the orbital as 2p (or $2p_x$, $2p_y$, $2p_z$) orbital, and not just as p orbital.
- Note that p orbitals are always aligned along (and not between) either the x or y or z axis.

B1(a)(iii) $Na^+(g) \rightarrow Na^{2+}(g) + e^-$

It is the energy required/absorbed/needed/supplied to remove 1 mole of electrons from 1 mole of gaseous Na^+ ions to form 1 mole of gaseous Na^{2+} ions.

Comments:
- Students need to show understanding that ionisation energies are endothermic. Hence, answers stating 'energy change' or 'energy evolved' or 'energy released' are not accepted.
- A small number of students did not understand that ionisation energies are computed based on a per mole basis. Hence, answers stating that it is the energy required to remove one electron from gaseous Na^+ ions to form gaseous Na^{2+} ions is not accepted.
- State symbols are compulsory for equation representing ionisation energies.

B1(a)(iv) Successive ionisation energies of sodium increase as each successive (negatively charged) electron is removed from an ion of increasing positive charge which attracts the electrons more strongly.

Comments:
- Students need to show understanding that it is increasingly harder to remove a negatively charged electron from an ion of increasing positive charge.

B1(b)(i) Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$
Ne: $1s^2\ 2s^2\ 2p^6$

Na has a lower first ionisation energy as it has 1 more principal quantum shell than Ne, hence the distance between the nucleus and valence electron is greater. Both nuclear charge and shielding effect also increases. Electrostatic attraction between the nucleus and valence electrons decreases, resulting in decrease in energy required to remove the valence electron from Na than in Ne.

Comments:
- Students need to show understanding that the difference in the first ionisation energies is primarily due to the distance of the valence electron to be removed from the nucleus. This is further elaborated by the fact that the valence electron from sodium is found in the n=3 principal quantum shell and the valence electron from neon is found in the n=2 principal quantum shell.
- Students should not explain in terms of achieving a stable octet configuration.

B1(b)(ii) Despite sodium having a lower first IE than neon, it is a solid at room temperature. Hence energy must be absorbed to vapourise/atomise sodium (ie. Na(s) $\rightarrow$ Na(g)) before ionisation occurs.

The combined energy of vapourisation and ionisation of sodium exceeds the ionisation of neon. Hence, when the lamp was first turned on, ionisation of neon takes place first before the ionisation of sodium.

Comments:
- A number of students failed to recognise that sodium needs to be in the vapour state before ionisation can take place and hence the orange glow of sodium is observed later.
- Purely stating that sodium is in the solid state and neon in the gaseous state is not sufficient as this is already given in the question.

B2(a)(i) $IO_3^- + 5I^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$

Comments:
Solve using half-equations (under acidic medium):
oxidation: $2I^- \longrightarrow I_2 + 2e^-$
reduction: $2IO_3^- + 12H^+ + 10e^- \longrightarrow I_2 + 6H_2O$
overall: $2IO_3^- + 10I^- + 12H^+ \longrightarrow 6I_2 + 6H_2O$
$IO_3^- + 5I^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$ (in simplest ratio)
(Please approach your tutor for the alternative method to work out the overall equation)
- Always ensure that there is both mass and charge balance in the overall equation.
- No electrons (e^-) should appear in the overall equation.
- Students are encouraged to give the overall equation in the simplest ratio.
- Although state symbols are not required for this question, I_2 is in the aqueous (aq) state (not s, l or g).

B2(a)(ii) $n(IO_3^-) = 19.80 / 1000 \times 0.120 = 0.002376$ mol

mole ratio of IO_3^- : I_2 : $C_6H_8O_6$ = 1 : 3 : 3

$n(C_6H_8O_6) = 0.002376 \times 3 = 0.007128$ mol

$[C_6H_8O_6] = 0.007128 / 25 \times 1000$
$= \underline{0.285 \text{ mol dm}^{-3}}$

Comments:

$IO_3^- + 5I^- + 6H^+ \longrightarrow 3I_2 + 3H_2O$ ----- (1)

$C_6H_8O_6 + I_2 \longrightarrow C_6H_6O_6 + 2I^- + 2H^+$ -----(2)

- I_2, which is produced in reaction (1), is reacted with $C_6H_8O_6$ in reaction (2). Hence, the mole ratio of IO_3^- : I_2 : $C_6H_8O_6$ should be used instead of the mole ratio of IO_3^- : I^- : $C_6H_8O_6$.
- Units for molar concentration is mol dm^{-3} (not g dm^{-3})

B2(a)(iii) Unlike potassium iodate(V), iodine is volatile and its concentration will vary during the course of titration, hence causing the titration results to be unreliable.

Comments:

- In the direct titration of vitamin C syrup with aqueous iodine, iodine can be placed in the burette or conical flask. Starch indicator is often used to make the end-point colour change more distinct to allow for more accurate determination of the end-point.
- In the case of titration of iodine (in conical flask) against vitamin C syrup (in burette), starch indicator is added when the concentration of unreacted iodine is low (i.e. solution in conical flask is pale yellow in colour) to prevent iodine molecules from being trapped in the starch molecules.
  Note that after the dark blue starch-iodine complex is formed, the iodine can still be released from the complex to react with the ascorbic acid in the titration.
- The use of correct scientific terminologies is important. Description of iodine 'escaped' or 'liberated' is not accepted. Also, *aqueous* iodine is used in the titration and it is incorrect to state that iodine sublimes (a phase transition from *solid* to *gaseous* state).
- A number of students incorrectly suggested that iodine is oxidised to iodide ions by atmospheric oxygen. The conversion of iodine to iodide ions is a reduction reaction.
- Some students suggested that the iodide ions formed in the titration can be oxidised back to iodine by atmosphere oxygen. Although this reaction is possible under acidic medium, the process takes time and has minimal/insignificant effect on the titration results.

B2(b)(i)

1. Fill a burette with 0.005 mol dm^{-3} potassium manganate(VII), $KMnO_4$.
2. Pipette 25.0 cm^3 of solution **A** into a 250 cm^3 conical flask.

3. Using a 50 cm^3 measuring cylinder, add 15 cm^3 of 0.050 mol dm^{-3} sulfuric acid into the conical flask.

Calculations (not required by question):
Assuming 1.0 g of iron(II) sulfate in 250 cm^3 of solution **A**,
$n(FeSO_4)$ in 250 cm^3 solution **A** = 1.0 / 151.9
= 0.006583 mol
= $n(Fe^{2+})$ in 250 cm^3 solution **A**

$n(Fe^{2+})$ in 25.0 cm^3 solution **A** = 0.006583 / 250 x 25.0 = 0.0006583 mol
$n(H^+)$ required = 0.0006583 / 5 x 8 = 0.001053 mol
$n(H_2SO_4)$ required = 0.001053 / 2 = 0.0005267 mol
Volume of H_2SO_4 required = 0.0005267 / 0.050 x 1000 = 10.53 cm^3

4. Titrate solution **A** against $KMnO_4$.

5. Stop the titration when the end-point is reached, i.e. when one drop of $KMnO_4$ added changes the colour of the solution in the conical flask from yellow to pale orange/pale pink.

6. Record the titration results using an appropriate table.

7. Repeat the titration until at least two consistent results are obtained (i.e. the titre volumes are within 0.10 cm^3 of each other).

Key points to include in the procedure:
- Apparatus (50 cm^3 burette, 25.0 cm^3 pipette, measuring cylinder, conical flask)
- Appropriate volume of H_2SO_4 added (10.55 to 100 cm^3)
- Colour change at end-point
- Consistent results

Comments:
- For planning experiments, the procedure should be presented in steps.
- As the redox titration involves coloured species, the titration is self-indicating. Use of indicators such as methyl orange, thymolphthalein etc is not required. Such indicators are used in acid-base titrations which do not involve coloured species.
- Students should read the question carefully.
  - The experiment involves 'titrating solution A (*in conical flask*) against $KMnO_4$ (*in burette*) under acidic conditions'.
  - Solution A is already provided. Hence there is no need to include preparation of solution A in the procedure.
- As the titration is done under acidic conditions, students should perform a brief calculation on the volume of H_2SO_4 (a dibasic acid) required for each titration. Hence an appropriate volume of H_2SO_4 used should ensure that
  - H^+ is present *in excess*
  - total volume of solution in the conical flask should not exceed half its maximum capacity
- To acidify the reaction mixture, it is recommended to add H_2SO_4 to the conical flask containing 25.0 cm^3 of solution A, instead of preparing a burette containing acidified $KMnO_4$.
- Some students prepared a mixture of solution A and H_2SO_4, followed by pipetting 25.0 cm^3 of this mixture for titration. Note that for this method,
  $n(FeSO_4)$ in 25.0 cm^3 of mixture $\neq$ $n(FeSO_4)$ in 25.0 cm^3 of solution A

B2(b)(ii) $n(KMnO_4)$ used = V / 1000 x 0.005 = $(5.00 \times 10^{-6})V$ mol
$n(FeSO_4)$ in 25.0 cm^3 of solution **A** = 5 x $(5.00 \times 10^{-6})V$ = $(2.50 \times 10^{-5})V$ mol
$n(FeSO_4)$ in 250 cm^3 of solution **A** = $(2.50 \times 10^{-5})V$ / 25.0 x 250
= $\underline{(2.50 \times 10^{-4})V \text{ mol}}$

Molar mass of $FeSO_4$ =151.9 g mol^{-1}
mass of $FeSO_4$ in **five** iron tablets = $(2.50 \times 10^{-4})V$ x 151.9 = 0.03798V g
mass of $FeSO_4$ in **one** iron tablet = 0.03798V / 5 = $\underline{0.00760V \text{ g}}$

Comments:
- A handful of students forgot that the *five* iron tablets are dissolved in *250 cm^3* (not 25 cm^3 which is used in the titration) of solution A.
- Students should use the M_r of $FeSO_4$ given in the question.

B2(b)(iii) Black ppt is $\underline{MnO_2}$.
The student $\underline{\text{missed the step on addition of } H_2SO_4}$.

Comments:
- Manganese oxide is not accepted. It should be manganese (IV) oxide or manganese dioxide.
- Most student correctly identified the missing step although the identity of the black ppt may be incorrect.
- Recall that
  - In *acidic* medium (presence of excess H^+), MnO_4^- ion is reduced to Mn^{2+}.

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$
purple → faint pink/colourless in dilute solutions

  - In *neutral or alkaline* medium, MnO_4^- ion is reduced to solid MnO_2.

$MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$
purple → brown/black solid

Since MnO_4^- : Fe^{2+} becomes 1 : 3, more MnO_4^- is required, causing the titre volume to be much higher than expected.

B2(b)(iv) The $\underline{\text{ascorbic acid present can also be oxidised}}$ by acidified potassium manganate(VII), causing the titration results to be inaccurate/much higher than expected.

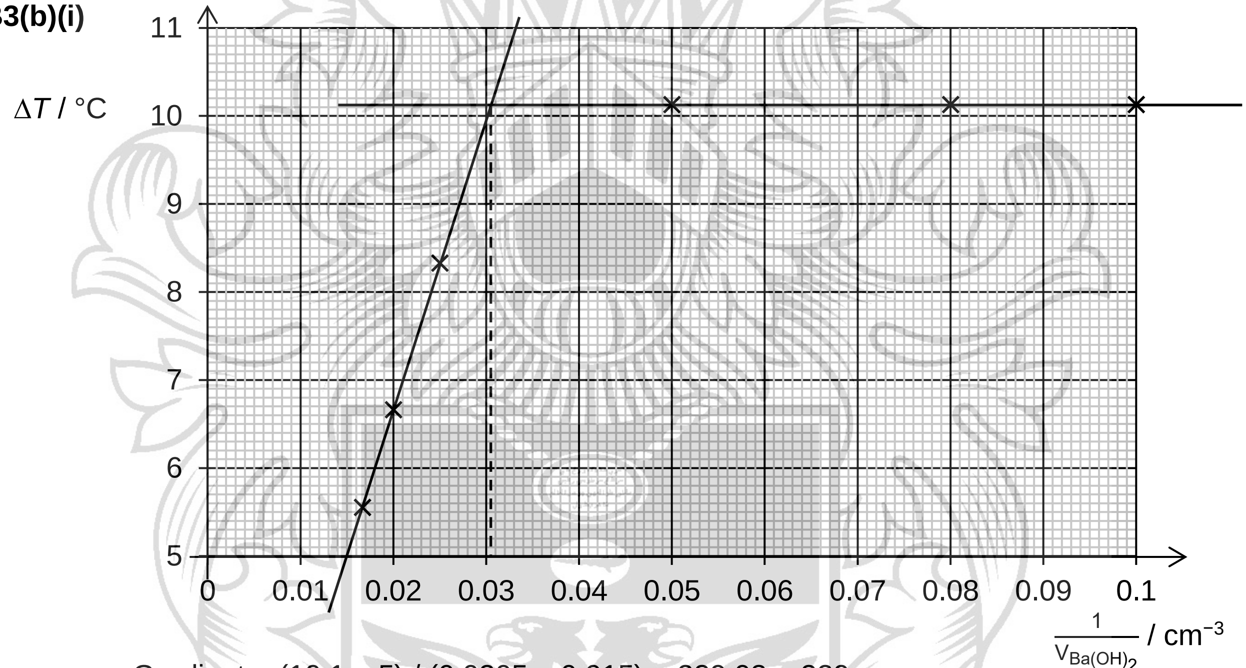
Comments:
- From parts (a) and (b) of the question, both ascorbic acid AND iron(II) sulfate can be oxidised (i.e. both are reducing agents). Hence, when Ferrograd® is titrated against $KMnO_4$, both ascorbic acid AND iron(II) sulfate can be oxidised by $KMnO_4$.
- A redox reaction involves both oxidation and reduction. Since both ascorbic acid and iron(II) sulfate are reducing agents, ascorbic acid cannot oxidised Fe^{2+} to Fe^{3+} ions.

B3(a) Definition: Standard enthalpy change of neutralisation ($\Delta H_{neut}^{\ominus}$) is the <u>energy change</u> when an acid and a base react to form <u>one mole of water</u> <u>at 298 K and 1 bar</u>.

Ionic equation: $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

<u>Comments:</u>
- Some students mistook "standard conditions" for "s.t.p."
- A common error was to leave out <u>"one mole"</u> of water.
- Note that an <u>ionic equation</u> (with no spectator ions) is required by the question.
- State symbols were sometimes incorrect or omitted.

B3(b)(i)



Gradient = (10.1 – 5) / (0.0305 – 0.015) = 329.03 = <u>329</u>

<u>Comments:</u>
- Some students drew best fit lines that were not up to standard. If the question had been read carefully, students should draw a horizontal line for the last 3 plotted points and a best fit straight line for the first 3 plotted points.
- The gradient should be derived from 2 points obtained from the best fit line. A common mistake was to take 1 point and use it to calculate the gradient based on the equation:

$$\Delta T = \frac{q}{c} \times \frac{1}{V_{Ba(OH)_2}}$$

  This is not accepted because experimental data is subjected to error and the line of best fit may not fit the given equation exactly. Also, students should infer from the question that the line of best fit was needed for the calculation of the gradient.
- Some students did not read off the graph properly and calculated the gradient incorrectly.

B3(b)(ii) Since gradient = $\frac{q}{c}$

heat change, q = 329.03 x 4.18 = <u>1375.3 J</u>

$2HX + Ba(OH)_2 \rightarrow BaX_2 + 2H_2O$

$HX : H_2O = 1:1$

So $\Delta H = -\frac{q}{n_{H_2O}} = -\frac{q}{0.025} = -55013 \text{ J mol}^{-1} = \underline{-55.0 \text{ kJ mol}^{-1}}$

<u>Comments:</u>

- Many students who attempted this question tried to apply their knowledge of q = mcΔT and calculated q by substituting m = 60. However, this shows a lack of understanding of the question as q is simply calculated by multiplying the gradient by c, which is 4.18.
- Students need to recognise that ΔT remains the same whether its units is in °C or K. Therefore, the 4.18 J K^{-1} cm^{-3} can be applied to the gradient (units of °C cm^3) directly without having to do any conversion.
- Based on the definition for standard enthalpy of neutralisation, students must divide q by the <u>amount of water</u> produced (which happened to be the same as the amount of limiting HX for this question).
- In Energetics, units are crucial (e.g. 1000 J is very different from 1000 kJ). All evaluated numerical figures should therefore be accompanied by the appropriate units.

B3(b)(iii) There would be <u>no change</u> to the enthalpy change of neutralisation. This is because even though <u>twice as much heat is being produced</u>, <u>twice as much water is also produced</u>.

<u>Comments:</u>

- Many students simply re-stated the definition of the enthalpy change of neutralisation as being "for one mole of water" and therefore concluded that the enthalpy change remained the same. However, this argument is insufficient as the amount of water produced would have doubled. Therefore, the key is to show explicit understanding that <u>both heat released</u> and <u>amount of water produced are doubled, so they cancel out</u>.

B3(c)(i) The point of intersection is where HX and $Ba(OH)_2$ are in the correct stoichiometric ratios of 2:1.

$$\frac{1}{V_{Ba(OH)_2}} = 0.0305$$

$$V_{Ba(OH)_2} = 32.787$$

$$= \underline{32.8 \text{ cm}^3}$$

<u>Comments:</u>

- Generally well done. Students managed to identify that the point of intersection is where HX is neither limiting nor in excess (i.e. in stoichiometric amount with respect to $Ba(OH)_2$).

B3(c)(ii) Amount of $Ba(OH)_2$ = 0.5 x 0.025 = 0.0125 mol

$[Ba(OH)_2]$ = 0.0125 / (32.787/1000) = 0.38125 = <u>0.381 mol dm^{-3}</u>

<u>Comments:</u>
- A common mistake was to forget the ratio of HX : $Ba(OH)_2$ and therefore using 0.025 mol in the answer.

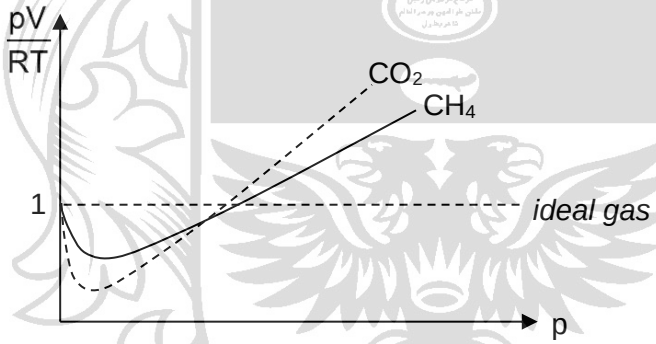
Section C

C1(a)(i) An ideal gas consists of <u>particles of negligible volume</u>. In other words, the size of the gas particles is negligible compared to the volume of the container.

The gas particles <u>exert negligible attractive forces</u> on one another.

<u>Comments:</u>
- The two assumptions given above are the two main assumptions. Students are encouraged to give these two assumptions rather than the other three assumptions on constant random motion, collisions being elastic and average energy being directly proportional to the absolute temperature.
- Volume of gas refers to the volume occupied by the gas in the container (which is equal to the volume of container as the gas fills up the container) ⇒ Hence, volume of gas is <u>not</u> negligible. Instead, it is the <u>volume of gas **particles**</u> that is negligible.

C1(a)(ii)



CO_2 deviates more from ideal behaviour at moderate pressure since it has <u>stronger instantaneous dipole-induced dipole interactions</u> due to its <u>larger and more polarisable electron cloud</u>.

<u>Comments:</u>
- The type of intermolecular forces must be specified.
- Some students mistakenly thought that CO_2 is polar. CO_2 is non-polar as it has a linear shape so the bond dipole moments cancel out each other.
- Some students mistakenly thought that CH_4 is polar. C is just slightly more electronegative than H and C−H bond is considered essentially non-polar. Also, CH_4 is tetrahedral so the bond dipole moments, if any, will cancel out.

- The strength of id-id interactions is due to the larger electron cloud (which is more polarisable). A larger M_r or larger size of molecules should not be used to explain the stronger id-id interactions.
- Some students mistakenly stated that CO_2 and CH_4 have 44 and 16 electrons respectively.
- A handful of students mistakenly related the intramolecular C=O bonds to stronger intermolecular id-id interactions.

C1(b) For ideal gas under same conditions of temperature and pressure, volume ratio is equivalent to mole ratio.

$$M_r = \frac{58(28.0)+21(2.0)+9(44.0)+7(16.0)+4(32.0)+34.1}{100} = \underline{23.4}$$

Comments:
- Students are reminded to give final numerical answers to 3 s.f.
- A handful of students showed calculation and rounding off errors.

C1(c)

$pV = nRT \Rightarrow \frac{pV}{T} = nR = \text{constant}$

$\therefore \frac{p(250)}{(37+273)} = \frac{(1)(300)}{(20+273)} \Rightarrow p = 1.27 \text{ atm}$

Alternatively,

At rtp condition, $n = \frac{300}{24000} = 0.0125$ mol

In the rectum, $pV = nRT$

$p \times 250 \times 10^{-6} = 0.0125 \times 8.31 \times (37 + 273)$

$\Rightarrow p = 1.29 \times 10^5$ Pa

Comments:
- Many students incorrectly used the formula $p_1V_1 = p_2V_2$, which is valid only under constant temperature. The appropriate formula is $\frac{p_1V_1}{T_1} = \frac{p_2V_2}{T_2}$, where the units of p and V must be consistent and T must be in Kelvin.
- The conditions for r.t.p. can be found in the *Data Booklet* (293 K and 101325 Pa).
- Many students calculated the number of moles of flatus using pV = nRT and mistakenly converted V to dm^3 instead of m^3, though the error was nullified when applying pV = nRT the second time.

C1(d)(i) 10 cm^3

Comments:
- Generally well done. Students are expected to know that the only acidic gases present are CO_2 and H_2S.
- Although H_2 or CH_4 contain H atoms, they are not acids as they do not produce H^+.

C1(d)(ii) In the 10 cm^3 of gas removed, 9 cm^3 was CO_2 and 1 cm^3 was H_2S.

$n(NaOH) = 2\ n(CO_2) = 2\ (\frac{9}{24000}) = 7.50 \times 10^{-4}$ mol

$n(NaOH) = 2\ n(H_2S) = 2\ (\frac{1}{24000}) = 8.33 \times 10^{-5}$ mol

Total $n(NaOH)$ required = $7.50 \times 10^{-4} + 8.33 \times 10^{-5}$ = $\underline{8.33 \times 10^{-4}\ mol}$

Or

$n(NaOH) = 2\ n(\text{acidic gases}) = 2\ (\frac{10}{24000}) = \underline{8.33 \times 10^{-4}\ mol}$

Minimum $[NaOH] = \frac{8.33 \times 10^{-4}}{0.050} = \underline{0.0167\ mol\ dm^{-3}}$

Comments:

- At r.t.p., amount of gas = $\frac{V}{\text{molar volume}} = \frac{V\ (\text{in } cm^3)}{24000}$ or $\frac{V\ (\text{in } dm^3)}{24}$
- 1 mol of dibasic acid reacts with 2 mol of OH^-. Hence, to find the amount of NaOH, students need to multiply the amount of acidic gases by 2, not 4.

C1(e)(i) $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$

Comments:

- When asked to write equation to represent an enthalpy change, state symbols MUST be given.
- Many students gave the equation $2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$ to represent $\Delta H_c(H_2(g))$. This is incorrect as $\Delta H_c(H_2(g))$ is the energy change when 1 mol of H_2 is completely burnt in excess O_2.
- At r.t.p. (as specified in the question), H_2O is in the liquid state.

C1(e)(ii) $n(CH_4) = \frac{7}{24000} = \underline{2.92 \times 10^{-4}}$ mol

$n(H_2) = \frac{21}{24000} = \underline{8.75 \times 10^{-4}\ mol}$

Heat evolved = $(2.92 \times 10^{-4})(880) + (8.75 \times 10^{-4})(280)$
= $\underline{502\ J}$

Comments:

- Heat evolved = $n(H_2) \times \Delta H_c(H_2) + n(CH_4) \times \Delta H_c(CH_4)$.
- Some students were unable to calculate amount of H_2 and CH_4. Students are reminded that under r.t.p. conditions, $n_{gas} = \frac{V_{gas}\ (\text{in } cm^3)}{24000}$ or $\frac{V_{gas}\ (\text{in } dm^3)}{24}$.

C1(e)(iii) From (e)(i), combustion of H_2 result in a decrease in entropy ($\Delta S < 0$).
Hence $T\Delta S$ is a negative term and the reaction is only spontaneous at temperature, T, where $|T\Delta S| < |\Delta H|$ or low temperatures such that $\Delta G < 0$.

Comments:
- Generally well done.
- Many students were able to deduce that $\Delta S < 0$, but fewer students concluded that $\Delta G < 0$ (for spontaneous reaction) could occur only for low temperatures specifically.

C2(a)(i) The electronegativity of an atom in a molecule is a relative measure of its ability to attract bonding electrons.

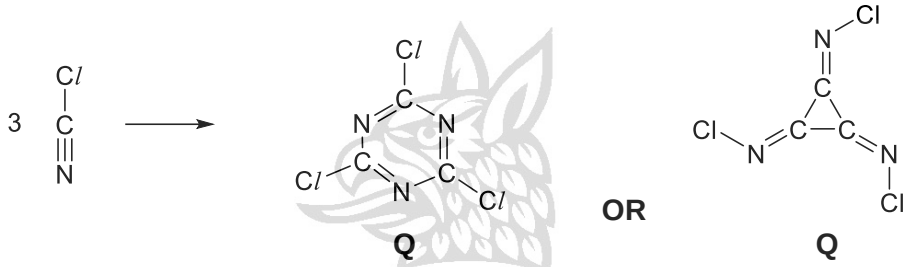
Comments:
- Many students tried to use their own words to explain electronegativity, which failed to convey the same idea. Students should follow the definition given in the lecture notes.
- Many students did not mention 'bonding electrons' or that the shared electrons belongs to a covalent bond.
- Some students confused electronegativity with electron affinity.

C2(a)(ii) Both the carbon-nitrogen and carbon-chlorine have the same electronegativity difference. The overall dipole moment of the molecule will be zero / the molecule is non-polar.

C*l*CN has instantaneous dipole-induced dipole (id-id) interactions between its molecules. Electrons are constantly moving and at any given moment, the electron density / electron cloud of a molecule can be unsymmetrical, resulting in an instantaneous dipole, which induces a short-lived dipole in a neighbouring molecule.

Comments:
- Many students used the abbreviations 'id-id' and 'pd-pd' interactions without stating what these abbreviations stand for.
- id-id and pd-pd are interactions, not bonds.
- Students are expected to explain how an instantaneous dipole arises, which later induces dipoles in neighbouring molecules.
- An instantaneous dipole arises, not due to the constant motion of electrons but due to a momentary distortion of a molecule's electron cloud.
- A large number of students were unable to recognise that C*l*CN is a non-polar molecule. Instead, most students explained that the molecule is polar because the bonds are polar or that there was an electronegativity difference.
- Based on information given by the question, even though the C-C*l* and C≡N bonds are polar, their dipole moments cancel out due to the linear geometry of the molecule. Hence, pd-pd interactions are not present.
- Many students have the wrong idea that only non-polar molecules have intermolecular id-id interactions. All molecules have them.

C2(a)(iii)



Comments:

- Students should be able to produce a molecular structure that has a symmetrical shape that enables the dipole moments of all the bonds to cancel each other.
- Many students misinterpreted the question and stated that **Q** has a "simple molecular structure". The question did not ask for the structure and bonding of **Q**. Moreover, stating that **Q** has a simple molecular structure will not explain why it has an overall dipole moment.
- Some students also misinterpreted the question by stating the molecular shape or geometry.

C2(b)(i)

Comments:

- For dot-and-cross diagrams, all the valence electrons should be shown.
- For the dative bond from O to C, the two electrons should be drawn using the same symbol as the valence electrons on O.
- The arrow '⟶' to show the dative bond from O to C should not be drawn here because the question did not ask for the structure of the molecule.
- Finally, electrons should not be drawn on overlapping circles representing electron shells anymore. Electrons occupy orbitals, which do not always have spherical shapes.

C2(b)(ii) $F_3B \leftarrow CO$

Comments:

- Many students drew the dot-and-cross diagram for the compound. This was not accepted because the question asked for the *structure*. All bonds should be represented by lines (or arrows, in the case of the dative bond), not dots and crosses.
- The dative bond needs to be drawn from C (the electron pair donor) to B (the acceptor atom).

C2(b)(iii)

	BF_3	F_3BCO
no of bond pairs	3	4
no of lone pairs	0	0
shape around B atom	trigonal planar	tetrahedral

Comments:
- Many students did not read the question which requires them to "deduce" the change in shape. Hence, students need to state the change in the number of bond pairs and lone pairs to conclude the change in shape around the B atom.

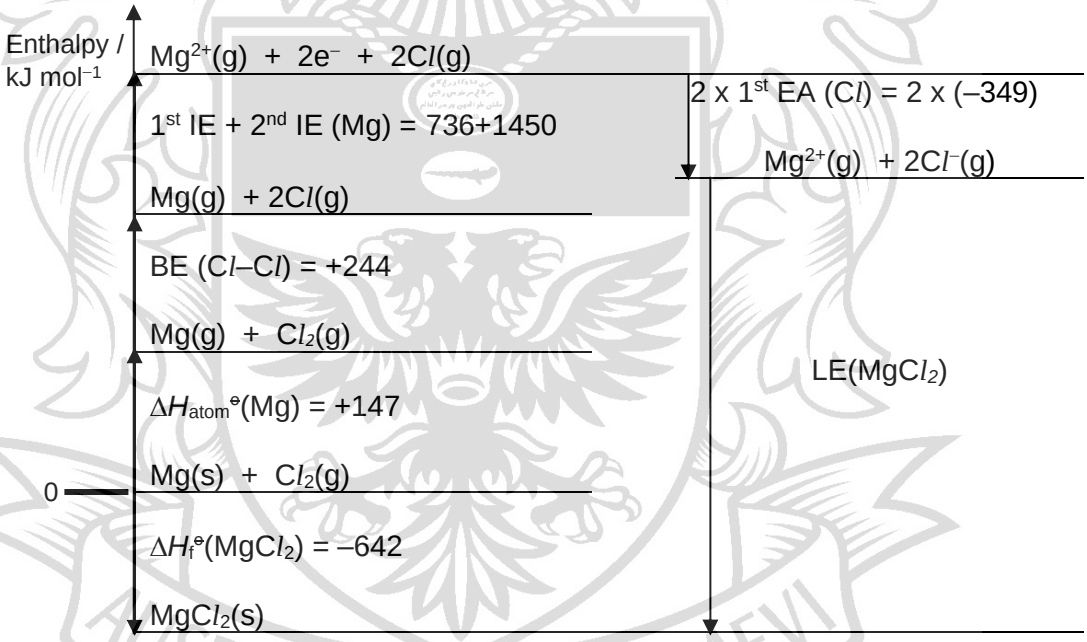
C2(c) MgF_2 has a giant ionic structure while PF_3 is a polar molecule and has a simple molecular structure.

More energy is required to overcome the stronger electrostatic forces of attraction between oppositely charged ions / strong ionic bonds in MgF_2 compared to the weak intermolecular forces of attractions (permanent dipole–permanent dipole interactions) between PF_3 molecules.

Hence the melting point of MgF_2 is higher than that of PF_3.

Comments:
- Many students did not state the structures of PF_3 and MgF_2.

C2(d)(i)



By Hess' law:

$\Delta H_f^{\ominus}(MgCl_2)$
$= \Delta H_{atom}^{\ominus}(Na) + BE\ (Cl–Cl) + 1^{st}\ IE + 2^{nd}\ IE\ (Mg) + 2 \times 1^{st}\ EA\ (Cl) + LE(MgCl_2)$
$–642 = +147 + 244 + 736 + 1450 + 2(–349) + LE(MgCl_2)$

$LE(MgCl_2) = –2521\ kJ\ mol^{–1} = \underline{–2520\ kJ\ mol^{–1}}$

Comments:

- Many students left out the BE(C*l*–C*l*) or forgot to multiply the EA of chlorine by 2.
- The axis need to be labelled and elements in their standard states should start at '0'.

C2(d)(ii) $\left| \text{lattice energy} \right| \propto \left| \frac{q_+ \times q_-}{r_+ + r_-} \right|$

As Mg^{2+} has a higher charge and is smaller than Na^+, the inter-ionic distance between Mg^{2+} and Cl^- ions will be shorter than that between Na^+ and Cl^- ions. Thus, the lattice energy of $MgCl_2$ will be more exothermic/negative than that of NaC*l*.

Comments:

- Both the charge and size of Mg^{2+} must be included in the explanation. The size of Mg^{2+} affects the inter-ionic distance, not the charge density.

C2(d)(iii) $\frac{\text{cation radius}}{\text{anion radius}} = \frac{0.065}{0.181} = 0.359$

coordination geometry of Mg^{2+} is tetrahedral and coordination no. = 4

Comments:

- Generally well done.

C2(d)(iv) Ba^{2+} has a larger coordination number as it has a larger ionic radius than Mg^{2+}, thus more Cl^- ions could be packed around the larger Ba^{2+}ion.

OR

$\frac{0.135}{0.181} = 0.746 > 0.732$

Thus Ba^{2+} has a larger coordination number which is greater than 6.

Comments:

- Students do not need to explain why Ba^{2+} is larger than Mg^{2+}.