2021 Y5 H2 Chemistry July Common Test – Suggested Solutions

Section A

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

MCQ worked solutions

1 Amount of CO_2 formed = 39.6 / 44 = 0.9 mol
Amount of H_2O formed = 16.2 / 18 = 0.9 mol

Since X and Y have the same empirical formula, the mole ratio of C : H in the products is the same as the mole ratio of C : H in X and Y.
$\Rightarrow$ mole ratio of C : H = 0.9 : 2(0.9) = 1 : 2
$\Rightarrow$ empirical formula of X and Y is CH_2.

Let the molecular formula of X be C_nH_{2n}. Then the molecular formulae of Y will be $C_{2n}H_{4n}$.
By considering the total amount of CO_2 formed from the combustion of 0.1 mol of C_nH_{2n} and 0.1 mol of $C_{2n}H_{4n}$,
$0.1 \times (n + 2n) = 0.9$
$n = 3$
$\Rightarrow$ Y has molecular formula C_6H_{12}.

2 Let the other species be X.

Oxidation number of Mn in MnO_4^{2-} = +6
Oxidation number of Mn in MnO_4^- = +7

oxidation: $MnO_4^{2-} \longrightarrow MnO_4^- + e^-$

1 mol of MnO_4^{2-} loses 1 mol of electron to form 1 mol of MnO_4^-.

Since this is a disproportionation reaction, MnO_4^{2-} undergoes oxidation to form MnO_4^- and reduction to form X.
$\Rightarrow$ Oxidation number of Mn in X should be lower than +6.
$\Rightarrow$ Option **D** is incorrect as oxidation number of Mn in Mn_2O_7 = +7

	X	Oxidation number of Mn	No. of electrons gained by 1 mol of MnO_4^{2-} to form 1 mol of X	Since number of electrons gained = number of electrons lost, ratio of X : MnO_4^-
A	$Mn(OH)_2$	+2	4 mol	1 : 4
B	MnO_2	+4	2 mol	1 : 2
C	Mn_2O_3	+3	3 mol	1 : 3

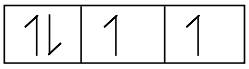
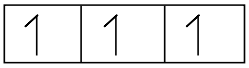
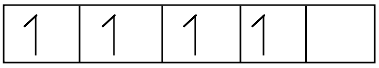
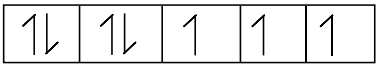
Only option **B** gives the correct ratio between the two manganese-containing products.

3 $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

Let the volume of each solution used be V dm^3.
Amount of HC*l* = V mol
Amount of NaOH = 1.5V mol
Since the mole ratio of HC*l* : NaOH = 1 : 1, HC*l* is limiting.

In the resulting solution:
Amount of NaOH unreacted = 0.5V mol
Amount of NaC*l* formed = V mol
Final total volume = 2V dm^3
[NaOH] = 0.5V / 2V = <u>0.25 mol dm^{-3}</u>
[NaC*l*] = V / 2V = <u>0.50 mol dm^{-3}</u>

4

Species	Electronic Configuration	Electrons-in-boxes diagram (for highest energy subshell)	Number of unpaired electrons
N^-	$1s^2 2s^2 2p^4$	2p 	2
F^{2+}	$1s^2 2s^2 2p^3$	2p 	3
Cr^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^4$	3d 	4
Co^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^7$	3d 	3

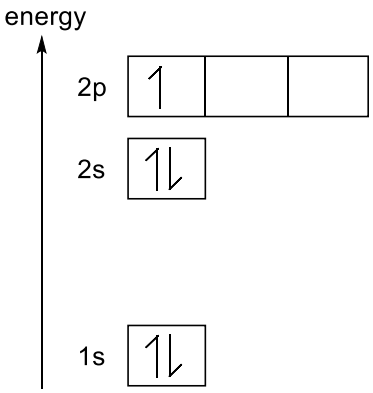
5 Option **A** is correct (as the statement is incorrect).
A neutron has no charge while an electron is negatively charged. Hence, they are <u>not oppositely charged</u>.

Option **B** incorrect (as the statement is correct).
A neutron is found in the nucleus at the centre of the atom while an electron is found around the nucleus. Hence, they are <u>found in different parts of the atom</u>.

Option **C** is incorrect (as the statement is correct).
A neutron has a relative mass of 1 while an electron has a relative mass of approximately 1/1840. Hence, the <u>neutron has a greater mass than the electron</u>.

Option **D** is incorrect (as the statement is correct).
The number of electrons in an atom must equal to the number of protons in an atom (for charge neutrality). The number of neutrons need not be the equal to the number of electrons in an atom. For example, H has 1 electron and 1 proton, but 0 neutron.

6 B: $1s^2\ 2s^2\ 2p^1$



The valence shell of B refers to the <u>electronic shell containing the 2s and 2p subshell</u>.

Option **A** is incorrect.
B only contains one half-filled orbital in the 2p subshell.

Option **B** is incorrect.
The 2s orbital is spherical in shape while the 2p orbitals are dumbbell in shape.

Option **C** is incorrect.
The 2p electron is at a higher energy level than the 2s electrons.

Option **D** is correct.
There is one orbital in the 2s subshell and three orbitals in the 2p subshell. Since each of the four orbitals can hold a maximum of 2 electrons, the valence shell of B can accommodate a maximum of eight electrons.

7 In the molecule, there are

- 15 single bonds $\Rightarrow$ 15 σ bonds
- 2 double bonds $\Rightarrow$ 2 σ bonds and 2 π bonds
- 1 triple bond $\Rightarrow$ 1 σ bond and 2 π bonds

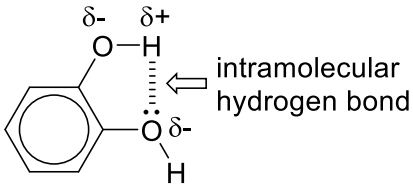
Hence, there are a total of <u>18 σ bonds and 4 π bonds</u>.

8 Option **A** is incorrect.
SiO_2 has a <u>giant molecular structure</u> with <u>strong covalent bonds</u>. Boiling SiO_2 requires overcoming these strong covalent bonds.

Option **B** is correct.
Greater number of valence electrons increases the strength of metallic bonds.

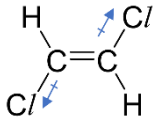
Option **C** is incorrect.
A*l*C*l*$_3$ has a <u>simple molecular structure</u> with <u>id-id</u> interactions.

Option **D** is incorrect.
In the first compound, the close proximity of the two -OH groups allows for intramolecular hydrogen bonding, which results in <u>less extensive intermolecular hydrogen bonding</u>. Since boiling involves overcoming the intermolecular hydrogen bonding, the first compound has a lower boiling point.

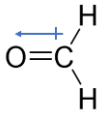


9 Given that the electronegativity values of C and H are similar, the C–H bond can be considered as non-polar. Hence, **A** is <u>non-polar</u>.

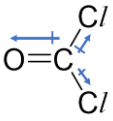
In **B**, the <u>C–C*l* bond dipoles cancel out</u> each other. Hence, **B** is also <u>non-polar</u>.



Both **C** and **D** are polar molecules. In **D**, the <u>dipole moment of C=O is partially offset by the dipole moments of the C–C*l* bonds</u>, resulting in a smaller overall dipole.

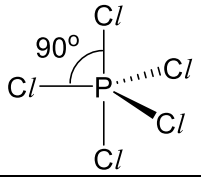
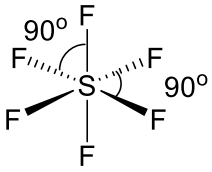
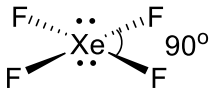


C



D

10

PCl_5	SF_6	XeF_4
		
trigonal bipyramidal (90° between axial and equatorial bonds)	octahedral (all bond angles are 90°)	square planar (all bond angles are 90°)

11

$$pV = nRT = \frac{mRT}{M}$$

$$M = \frac{mRT}{pV} = \frac{m}{V}\frac{RT}{p} = \text{density} \times \frac{RT}{p}$$

$$\text{Average } M_r = \text{density} \times \frac{RT}{p} = (5.00 \times 10^3) \times \frac{8.31 \times 298}{100 \times 10^3}$$

$$= 123.82$$

Note the conversion of units:
- density to g m^{-3}
- temperature to K
- pressures to Pa

Let the mole fraction of He and Xe be x and $(1 - x)$ respectively.
$123.82 = 4x + 131.3(1 - x)$
x = <u>0.059</u>

12 Using pV = nRT and making the y-axis the subject:

graph 1 (correct)	pV = nRT = constant at a particular T (since n, R and T are constants) ⇒ graph is similar to y = k (<u>horizontal straight line</u>) ⇒ <u>pV is larger at higher T</u>
graph 2 (correct)	$V = (\frac{nR}{p})(T) = (k)(T)$ at a particular p (since n, R and p are constants) ⇒ graph is similar to y = kx (<u>straight line that passes through origin</u>) ⇒ <u>gradient</u> ($= \frac{nR}{p}$) is <u>smaller at higher pressure</u>
graph 3 (incorrect)	$p = (nRT)(\frac{1}{V}) = (k)(\frac{1}{V})$ at a particular T (since n, R and T are constants) ⇒ graph is similar to y = kx (<u>straight line that passes through origin</u>) ⇒ <u>gradient</u> (= nRT) is <u>larger at higher T</u>

13 ΔH_{hyd} is directly proportional to the charge density (or charge ÷ radius) of the ion. The <u>higher the charge density</u> of the ion, the <u>greater the magnitude of its hydration energy</u>.

$$\left|\Delta H^{\ominus}_{hyd}[M^+(g)]\right| \propto \left|\frac{q_+}{r_+}\right|$$

ion	q_+	r_+	$\lvert q_+ / r_+ \rvert$
Cr^{2+}	2+	0.073	2 ÷ 0.073 = 27.4
Fe^{2+}	2+	0.061	2 ÷ 0.061 = 32.8
Ti^{3+}	3+	0.067	3 ÷ 0.067 = 44.8

14 Option **A** is correct.
ΔH_r = ΣBE of bonds in reactants – ΣBE of bonds in products
Since $\Delta H_r < 0$, energy released during bond forming is more than energy absorbed during bond breaking.

Option **B** is incorrect.
$\Delta H_r = \Sigma n \Delta H_c$ (reactants) – $\Sigma m \Delta H_c$ (products)
Hence, ΔH_c ($H_2(g)$) is also required.

Note: If formation enthalpies are given, $\Delta H_r = \Sigma n \Delta H_f$ (products) – $\Sigma m \Delta H_f$ (reactants), only ΔH_f ($C_2H_4(g)$) and ΔH_f ($C_2H_6(g)$) are required as ΔH_f ($H_2(g)$) = 0.

Option **C** is correct.
As 2 mol of gaseous reactants react to form only 1 mol of gaseous product, the disorder of the system decreases and hence $\Delta S < 0$.

Option **D** is correct.
$\Delta G = \Delta H - T\Delta S$
Since $\Delta H < 0$ and $\Delta S < 0$, $\Delta G < 0$ only at low T.

15 equation **1** is incorrect.
P = $\Delta H^{\ominus}_{atomisation}(C(s)) + \mathbf{2}\Delta H^{\ominus}_{atomisation}(O_2(g))$

equation **2** is correct.
By Hess' law, S = – P + Q + R
Since Q + R = $\Delta H^{\ominus}_{formation}(CO_2(g))$
Hence, S = – P + $\Delta H^{\ominus}_{formation}(CO_2(g))$

equation **3** is correct.
Q + R = $\Delta H^{\ominus}_{combustion}(C(s))$
Hence, Q = $\Delta H^{\ominus}_{combustion}(C(s))$ – R

Note: The equation for $\Delta H^{\ominus}_{combustion}(C(s))$ is the same as that for $\Delta H^{\ominus}_{formation}(CO_2(g))$

equation **4** is incorrect.
$\Delta H^{\ominus}_{combustion}(CO(g))$ is represented by the equation $CO(g) + \mathbf{½O_2}(g) \rightarrow CO_2(g)$

Section B

B1(a)(i) Let the percentage abundance of thallium-203 be x%. Then the percentage abundance of thallium-205 will be (100 – x)%.

$(\frac{x}{100} \times 203) + (\frac{100-x}{100} \times 205) = 204.4$

x = 30.0

Percentage abundance of thallium-203 = <u>30.0%</u>
Percentage abundance of thallium-205 = 100 – 30.0 = <u>70.0%</u>

Comments:
- Generally well done. Unless otherwise stated, students should present their final answer to 3 s.f.
- Some students misread the question and incorrectly assumed the total abundance to be 92.5%.

B1(a)(ii) Relative atomic mass refers to the relative <u>average mass</u> of <u>1 atom</u> of an element compared to <u>1/12 the mass of a carbon-12 atom</u>.

OR

$$\text{Relative atomic mass} = \frac{\text{(weighted) average mass of 1 atom of the element}}{\frac{1}{12} \text{ x mass of 1 atom of carbon-12}}$$

Comments:
- Students need to revise the definition and be consistent in whether they are comparing the ratio in terms of 1 atom or 1 mole of atoms. Students also need to specify the isotope carbon-12 in the definition.
- It is an average mass to take into consideration the different isotopes of the element and their relative proportions.
- Many students forgot to include '$\frac{1}{12}$' in the definition.

B1(b)(i) Percentage thallium in thallium(I) bromide = $\frac{204.4}{204.4 + 79.9} \times 100\% = \underline{71.9\%}$

Comments:
- Generally well done.

B1(b)(ii) Mass of thallium in final precipitate = 71.9% × 7.45 x 10^{-4} = 5.36 × 10^{-4} g
Mass of thallium in 100 cm^3 solution = 5.36 x 10^{-4} × (100/25) = 2.14 × 10^{-3} g
Mass of thallium in bone = 2.14 x 10^{-3} × (100/92.5) = 2.31 × 10^{-3} g

Comments:
- Several students did not find the mass of thallium in 100 cm^3 solution because they did not account for the 92.5% recovery and/or the fact that only ¼ of the 100 cm^3 solution was analysed.

B1(b)(iii) $\frac{2.31 \times 10^{-3}}{50} \times 10^6 = \underline{46.2 \text{ ppm}}$

Comments:
- Some students failed to read the question at the start of (b) and did not realise that the bone sample is 50 g.

B1(c)(i) $SO_2 + Br_2 + 2H_2O \longrightarrow SO_4^{2-} + 2Br^- + 4H^+$

Comments:
- Generally well done.

B1(c)(ii) For the same amount of T*l* present in 50 g of the bone, since $\underline{M_r \text{ of } TlBr_3 \text{ is greater than } TlBr}$, $\underline{\text{having some } TlBr_3 \text{ in the precipitate would lead to the mass of precipitate in step 7 to be higher than expected}}$. The $\underline{\text{calculated concentration in (b)(iii) would increase/be higher.}}$

Comments:
- Students need to be clear whether they are explaining in terms of T*l*Br in ppt, or T*l*Br and T*l*Br_3 in ppt.
- Students must also substantiate their answers when stating facts like "the real percentage of T*l* in the bone will become lower". This is because the real percentage of T*l* in the bone could have become higher if the "impurity" were not T*l*Br_3 but T*l*.

B1(d) Any Pb^{2+} ions present in the bone would be precipitated out when H_2SO_4 is added as $PbSO_4$, an insoluble salt, is formed.

Comments:
- Students must recognise that Pb^{2+} is removed from solution because it formed $PbSO_4$ precipitate that was filtered away.

B1(e) Pipette / burette

Comments:
- Students need to realise that the 25 cm^3 solution measured is to be used for titration and the volume measured needs to be very precise and thus only pipette / burette can be used.

B2(a) (Any two of the following)
- An ideal gas consists of particles of negligible volume.
- The gas particles exert negligible attractive forces on one another.
- The gas particles are in constant random motion.
- Collisions between the gas particles are perfectly elastic.
- The average kinetic energy of particles in a gas is constant at constant temperature. The average kinetic energy is proportional to the absolute temperature.

Comments:
- Generally well done. A common mistake is to state that the volume of an ideal gas is negligible. This is incorrect as volume of a gas is the volume of the container it fills and hence is not negligible. It is the volume of gas **particles** (which is the volume actually occupied by the particles) which is negligible.

B2(b)(i) $n = \frac{pV}{RT} = \frac{(18 - 1) \times 101325 \times \frac{15}{1000}}{8.31 \times 293} = 10.6 \text{ mol}$

OR

Since n and T are constant, pV = constant

Volume of gas at released under room temperature and pressure = $\frac{17 \times 15}{1} = 255 \text{ dm}^3$

Amount of gas released = $\frac{255}{24} = 10.6$ mol

Comments:
- Students should note that R has the value of 8.31 J mol^{-1} K^{-1} only if p is in Pa, V in m^3, n in mol and T in K. Hence, to calculate n, students need to convert
  - p in atm to Pa ($\times$ 101325) and
  - V in dm^3 to m^3 ($\times 10^{-3}$).
- Many students did not take into consideration that the residual pressure is 1 atm and used p of 18 atm (instead of 17 atm) to calculate n.
- Students should note that amount of a substance refers to the number of moles of that substance, and not its volume or mass.

B2(b)(ii)

$n_{\text{each balloon}} = \frac{pV}{RT} = \frac{1.2 \times 101325 \times 5.50 \times 10^{-3}}{8.31 \times 293} = 0.2747$ mol

Number of balloons that can be filled = 10.61 / 0.2747 = 38.6 = <u>38</u> (rounded down)

OR

p_1V_1 (in gas tank) = p_2V_2 (in balloons)

$V_2 = \frac{p_1V_1}{p_2} = \frac{17 \times 15}{1.2} = 212.5$ dm^3

Number of balloons that can be filled = 212.5 ÷ 5.50 = 38.6 = <u>38</u> (rounded down)

Comments:

- As the question requires students to calculate the maximum number of balloons that can be inflated to 5.50 dm^3, the expected answer is a whole number and should be rounded down.

B2(c)

$n_{He} = \frac{pV}{RT} = 0.8 \text{ x } \frac{(18) \times 101325 \times \frac{15}{1000}}{8.31 \times 293} = 8.989$ mol

At –200 °C, $p = \frac{nRT}{V} = \frac{8.989 \times 8.31 \times 73}{\frac{15}{1000}} = \underline{3.64 \times 10^5 \text{ Pa}}$

OR

$\frac{p_1}{T_1} = \frac{p_2}{T_2}$

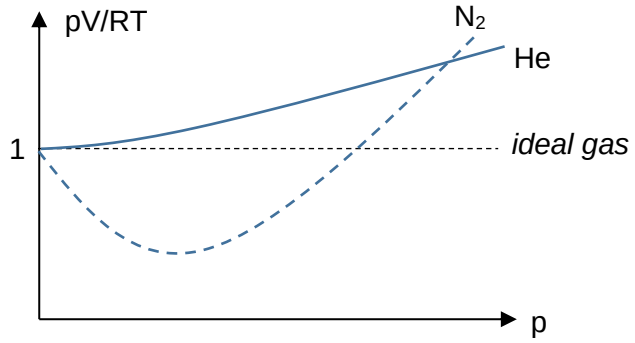
$\frac{0.8 \times 18}{293} = \frac{p_2}{73}$

$p_2 = \underline{3.59 \text{ atm}}$

Comments:

- As the gas mixture in the tank contains 80% He,
  $n_{He} = 0.8 \times n_{gas}$ in tank or
  p_{He} at 293 K = 0.8 × pressure of tank
- At −200 °C, N_2 is liquefied and the only gas remaining is He. Therefore,
  new internal pressure of the tank = p_{He} at −200 °C (or 73 K)

B2(d)



At moderately high pressure, negative deviation is observed for N_2. Both N_2 and He form intermolecular instantaneous dipole-induced dipole interactions. N_2 has more electrons and a larger electron cloud which is more easily polarised, thus N_2 molecules form stronger instantaneous dipole-induced dipole interactions. N_2 deviates more from ideality than He.

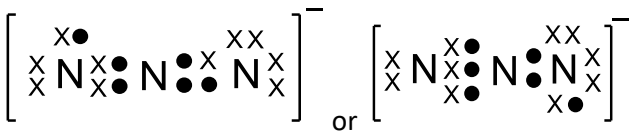
OR

At very high pressure, positive deviation is observed. The intermolecular repulsive force between N_2 molecules is more significant than that of He as the molecular size / size of electron cloud of N_2 molecules is larger than that of He atoms. Hence, N_2 deviates more from ideality than He.

Comments:

- Some students gave the wrong formula for nitrogen (N instead of N_2) and/or for helium (H_2 instead of He).
- As both N_2 and He have simple molecular structures, the intermolecular forces of instantaneous dipole−induced dipole (id−id) interactions should be named.
- Since the strength of id−id interactions is dependent on ease of distortion of electron cloud, the explanation of why N_2 has stronger id−id interactions than He should include N_2 having greater number of electrons (NOT N_2 having greater M_r or N_2 is a diatomic molecule while He is monoatomic).
- As He has 2 electrons, there are still id−id interactions between He atoms even though the id−id interactions are very weak.

B3(a)



Comments:

- Unless otherwise stated, students should use only dot (•) and cross (×) to represent the electrons in a dot-and-cross diagram.
- As N is a Period 2 element with no low-lying vacant d orbitals, N cannot expand octet. Hence, in both structures, there is a dative bond from the central N atom to one of the side N atoms.
- Several students missed out indicating the negative charge of the azide ion.
- Extra electrons are usually placed on the (more electronegative) side atoms. Students are advised to use a different symbol, i.e. a symbol different from that representing the electrons of an atom, to represent any extra electrons on that atom.

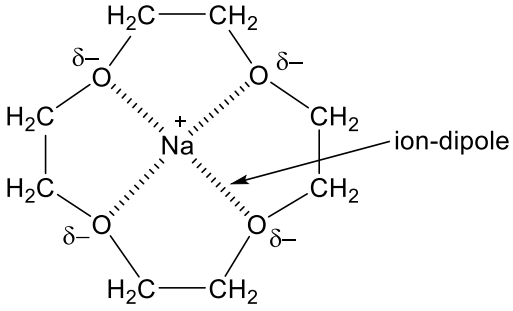
B3(b)(i) Since there are 2 lone pairs and 2 bond pairs about each O atom, the shape about each O atom is bent.

As lone pairs exert greater repulsion than bond pairs, the bond angle about each O atom is 105^o.

Comments:

- Some students showed confusion between electron pair geometry and molecular geometry/shape. In this question, the oxygen atoms have a tetrahedral electron pair geometry (due to four regions of electron density). However, as two of these electron pairs are lone pairs, the shape/molecular geometry about the oxygen atoms is bent.
- To account for why the bond angle is smaller than 109.5^{o} (bond angle in a tetrahedral), students need to compare the strength of repulsion exerted by lone pairs vs bond pairs, i.e. lone pair-lone pair > lone pair-bond pair > bond pair-bond pair repulsion.
- A few students mistakenly thought that there is only one lone pair around each oxygen atom.

B3(b)(ii)



Comments:

- Reading of question is important. This question requires students to draw and <u>label</u> the interaction between the Na^+ (only <u>one Na^+</u>) and crown ether. The answer will need to include:
  - '+' on sodium ion
  - 'δ–' on O
  - labelling of ion-dipole interaction
- Note that the Na^+ is a positively charged ion and hence does not possess partial δ+ charge.
- Ion-dipole interaction should be represented using a dotted line. Some students misrepresented it by using a line (which is used for covalent bond).

B3(b)(iii) A molecule of crown ether forms <u>more extensive ion-dipole</u> interactions with the sodium ion than a molecule of dimethyl ether. Hence, a molecule of crown ether is able to stabilise the sodium ion more than a molecule of dimethyl ether.

Comments:

- For students who recognised that the O atoms in crown ether form interactions with Na^+ in (b)(ii), most were able to state that crown ether forms <u>four</u> ion-dipole interactions with Na^+ compared to only <u>one</u> in dimethyl ether as crown ether has four δ– O atoms while dimethyl ether has only one δ– O atom.
- For students who recognised that dimethyl ether is too small to encapsulate Na^+, many did not go on to explain that the more exposed Na^+ is unable to form favourable interactions with the non-polar solvent and is hence less stabilised.
- Do note the use of correct terminologies for this question, e.g. oxygen atoms (NOT oxygen molecules)

B3(b)(iv) The <u>non-polar hydrocarbon chain / CH_2CH_2 groups</u> of crown ether molecules can form favourable <u>instantaneous dipole-induced dipole interactions</u> with the solvent molecules.

Comments:

- Reading of question is important. For this question, students are to refer to the <u>structure</u> of crown ether to explain why the <u>positively charged</u> crown ether with encapsulated Na^+ is still able to dissolve in the non-polar organic solvent. Hence, it is necessary to specify
  - <u>which part</u> of crown ether can form favourable interaction with the non-polar solvent
  - and the <u>type of interaction</u> formed.
- Several students discussed about the dipole moments in crown ether cancelling out resulting in a non-polar crown ether, without realising that the question already stated that the crown ether with encapsulated Na^+ is positively charged.

B3(b)(v) The <u>positively charged</u> crown ether with the encapsulated sodium ion can form <u>electrostatic forces of attraction</u> with the <u>negatively charged</u> azide ion, hence allowing the azide ion to enter the non-polar organic solvent.

OR

With the stabilisation of the sodium ion by the crown ether, there will be a <u>build-up of positive charge</u> in the non-polar organic solvent. In order to <u>maintain charge neutrality</u>, the negatively charged azide ions will enter the non-polar solvent.

Comments:

- Few students were able to answer this question.
- Reading of the context of question is important. In this question, <u>aqueous</u> NaN_3 is added to CH_3CH_2Br dissolved in a <u>non-polar organic solvent</u>. As the two layers are immiscible, crown ethers can be added to increase the solubility of NaN_3 in the non-polar organic solvent for reaction (shown by equation 1) to occur.
  From this information, students should note that NaN_3 is not in the solid state and hence should not consider enthalpy changes involving breaking of ionic bonds in solid NaN_3.
- From the information given, only Na^+ is encapsulated in the crown ether.
- Note that azide ion, N_3^-, is negatively charged and cannot form favourable interactions with the non-polar organic solvent.

B3(c) All three compounds exist as discrete molecules with similar electron cloud size.

Dimethyl ether has the lowest boiling point because of the weaker intermolecular permanent dipole–permanent dipole (pd-pd) interactions. Hence the least amount of energy is required to overcome the weaker pd-pd interactions.

Both compounds **A** and **B** have intermolecular hydrogen bonding, which are stronger than pd-pd interactions.

B is able to form more extensive intermolecular hydrogen bonds compared to **A**. Therefore **B** has a higher boiling point than **A**.

Comments:

- Some students mistakenly thought that dimethyl ether is a non-polar compound. Similar to water, the bond dipoles in dimethyl ether do not cancel out and hence dimethyl ether is a polar compound.
- Most students were able to recognise that only A and B have intermolecular (NOT intramolecular) hydrogen bonds and that the hydrogen bonds in B are more extensive. B has an average of two hydrogen bonds per molecule while A has an average of one hydrogen bond per molecule.
- Some students fail to see that the –OH group in A allows it to form intermolecular hydrogen bonding.
- Note that all three compounds are straight chain compounds with no branching.
- When comparing differences in strength of id-id interactions due to branching, this is usually applied in the context of isomers (compounds with same molecular formula but different arrangement of the atoms in their molecules).

Section C

C1(a) Combining atomic nuclei requires large amount of energy to overcome the repulsion from the like charges present in the positively charged nucleus.

Comments:

- Students need to understand that the atomic nucleus consists of neutrons and protons (not electrons). Protons are positively charged and hence, the nucleus is positively charged.
- Students also need to understand that the combination of like charges would result in repulsion.

C1(b) **X** = $^{23}_{11}$Na or sodium-23

Y = $^{1}_{0}$n or neutron

Comments:

- Students need to take into consideration the nucleon number, proton number and neutron number. Students also need to understand what each number written beside the element denotes.

$$^{12}_{6}C + {}^{12}_{6}C \rightarrow \mathbf{X}(^{23}_{11}Na) + {}^{1}_{1}H$$

- It is important to include the nucleon number (23) of Na as different isotopes of Na have different masses.
- A common mistake for **X** is ^{23}Mg. Note that Mg has 12 protons instead of 11 protons. Another common mistake for **X** is V. Note that V has 23 protons instead of having a nucleon number of 23.

$$^{12}_{6}C + {}^{12}_{6}C \rightarrow {}^{23}_{12}Mg + \mathbf{Y}(^{1}_{0}n)$$

- A common mistake for **Y** is ^{1}H. Note that H has 1 proton while **Y** has no proton and 1 neutron.

C1(c)(i) Fe: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6\,4s^2$
Fe^{2+}: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^6$

Comments:

- Students need to be familiar with the correct representation of electronic configuration. Common mistakes include: 1s2,2s2,2p6; $1s_2 2s_2 2p_6$; 1s22s22p6.
- Students should know that the 4s orbitals are filled up before 3d. However, in writing, 3d needs to be written before 4s. Hence, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2 3d^6$ will not be accepted.
- Students should also know that electrons are removed from 4s orbital first before 3d orbitals during the formation of cations.
- Students should work out the electronic configuration of cation from the electronic configuration of the parent atom.
- When the question asks for you to state the electronic configuration, you are not allowed to write in the following manner: [Ar] $3d^6\,4s^2$. This representation is only permissible if you provide the electronic configuration to aid the explanation of concepts like ionisation energy.

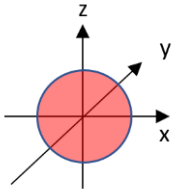
C1(c)(ii) The radius of Fe^{2+} is <u>smaller</u> than Fe.

Both Fe and Fe^{2+} have the same number of protons and hence have the same nuclear charge. However, Fe^{2+} has <u>one less electronic shell</u> compared to Fe. Thus, the electrostatic attraction between the nucleus and the valence electrons increases, resulting in a smaller radius for Fe^{2+}.

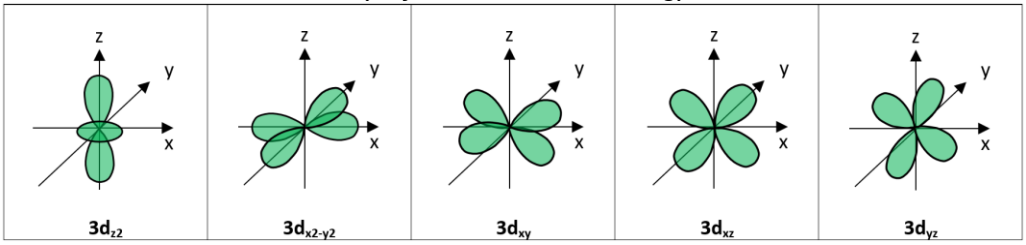
Comments:

- Students need to state that the radius of Fe^{2+} is smaller than Fe. Also, students should know that the main reason is due to Fe^{2+} having one less electronic shell despite having the same nuclear charge as Fe.
- Explanations involving the Fe^{2+} having more protons than electrons are not accepted.

C1(c)(iii) Second electron removed from 4s.



Third electron removed from: (any one of the following)



Comments:
- Most students are able to draw and label the 4s orbital. However, students need to be more familiar with the labelling and representation of the 3d orbitals.
- The 3d orbital drawn should correspond with the name of the orbital (i.e. $3d_{xz}$, $3d_{xy}$).

C1(c)(iv) The 4s orbital is non-directional while the 3d orbital is directional.

OR

The 4s orbital is larger / more diffuse than the 3d orbital.

Comments:
- Students need to make specific reference to the orbitals in their answers.
- Some students incorrectly stated that 4s orbital contains 2 electrons while 3d orbital contains 10 electrons. According to Hund's rule, the maximum number of electrons found in an orbital is 2.

C1(c)(v) For D^-, $\frac{q}{m} = -0.5$ and $\theta = -18^o$.

Since $\theta = k(\frac{q}{m})$, $k = 36$

For $^{56}Fe^{2+}$, $\frac{q}{m} = \frac{2}{56}$.

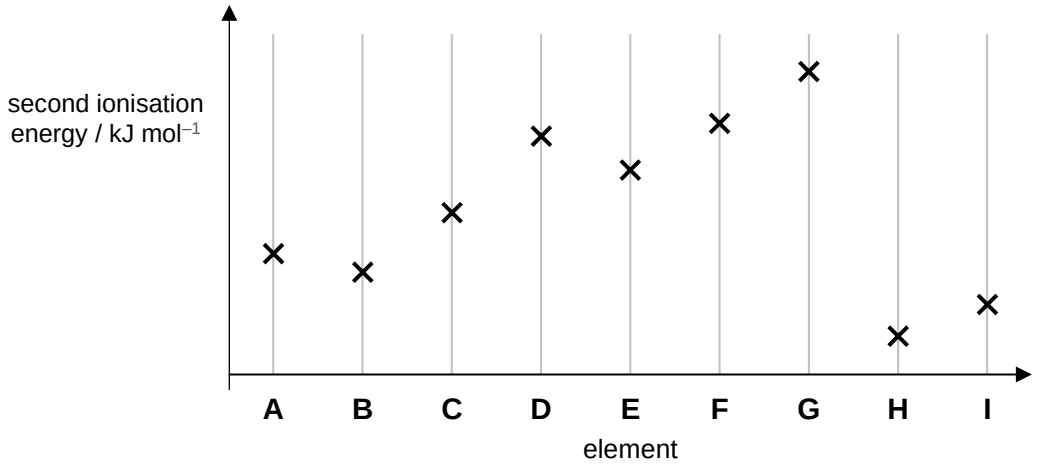
Hence, $\theta = 36 \times \frac{2}{56} = \underline{+1.29^o}$

Comments:
- Students need to indicate the sign in front of the angle of deflection as it suggests the direction of the deflection.

C1(d)(i) $A^+(g) \rightarrow A^{2+}(g) + e^-$

Comments:
- Students are required to answer this question in terms of A.
- Students need to indicate that ionisation involves the formation of gaseous ions from gaseous ions/atoms.
- A small number of students wrongly suggested that electrons are in the gaseous state, i.e. $e^-(g)$.

C1(d)(ii)



Comments:
- From the large drop in IE, students should be able to determine that H^+ has an additional electronic shell compared to G^+. Hence, C^+ has the electronic configuration ending with ns^2np^2 and E^+ has the electronic configuration ending with ns^2np^4.

C1(d)(iii) B^+: ns^2np^1
C^+: ns^2np^2
C^+ has one more proton and hence has a higher nuclear charge than B^+. Though C^+ also has one more electron than B^+, the increase in shielding effect is minimal since this additional electron occupies the outermost shell. The valence electrons in C^+ experience a higher effective nuclear charge and are more strongly attracted by the nucleus. Thus, second ionisation energy of C is higher than B.

D^+: ns^2np^3
E^+: ns^2np^4
The p electron to be removed from E^+ is a paired electron while that to be removed from D^+ is an unpaired electron. Due to the inter-electronic repulsion between paired electrons in the same orbital, less energy is required to remove the paired p electron from E^+. Hence, the second ionisation energy of E is lower than D.

Comments:
- Students need to be aware that across the period, the shielding effect remains approximately constant (NOT the same or constant).
- Many students were able to identify the presence of paired electron in E^+ and mentioned about inter-electronic repulsion accordingly.

C2(a)(i) enthalpy change of combustion: always negative
enthalpy change of formation of compound: either positive or negative

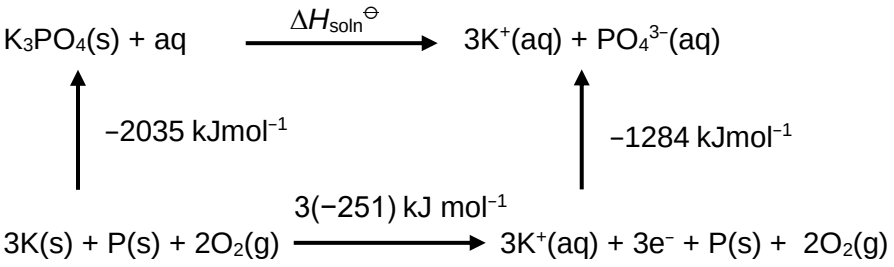
Comments:
- Generally well done.

C2(a)(ii) Standard enthalpy change of solution of potassium phosphate, $K_3PO_4(s)$, is the energy change when one mole of $K_3PO_4(s)$ is completely dissolved in a solvent to form an infinitely dilute solution at 298 K and 1 bar.

$K_3PO_4(s) + aq \longrightarrow 3K^+(aq) + PO_4^{3-}(aq)$ $\Delta H_{soln}^{\ominus}$

Comments:
- Students need to revise the definitions to be precise in recalling key terms of the definition, such as "energy change" as opposed to "energy released/ required".
- Many students forgot to state the standard conditions; some incorrectly stated standard conditions as 293 K and 1 atm.

C2(a)(iii)



$\Delta H_{soln}^{\ominus} = -(-2035) + 3(-251) + (-1284) = \underline{-2.00 \text{ kJ mol}^{-1}}$

Comments:
- Students need to label every arrow drawn in the energy cycle with either the ΔH representation or value.
- Students need to check to ensure that every equation in the energy cycle is balanced, in terms of mass and charge.
- Chemical equations should be balanced by addition of chemical species instead of deduction of chemical species.

C2(a)(iv) $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$
$-283 = (-251) - 298\ (\Delta S^{\ominus})$
$\Delta S^{\ominus} = +0.107 \text{ kJ mol}^{-1} \text{ K}^{-1} = \underline{+107 \text{ J mol}^{-1} \text{ K}^{-1}}$

Since the reaction results in a change from a highly ordered solid state to a disordered aqueous state, there is an increase in the number of ways of arrangement of particles so $\Delta S^{\ominus}$ is positive.

Comments:
- Students need to read the question carefully to avoid substituting the wrong value for $\Delta H^{\ominus}$ in the calculation of $\Delta G^{\ominus}$.
- Many students gave the wrong units for $\Delta S^{\ominus}$ as J mol^{-1} instead of J mol^{-1} K^{-1}.
- Some students made the mistake of stating that K changes from solid to liquid state, instead of solid to aqueous state.
- Many students did not link the increase in entropy to the increase in the number of ways of arrangement of particles.

C2(b)(i) $HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$ $\Delta H^{\ominus} = -53.4$ kJmol^{-1}

Amount of $HCl = \frac{30.0}{1000} \times 1.20 = 0.036$ mol

Amount of $NH_3 = \frac{40.0}{1000} \times 0.85 = 0.034$ mol

Since the mole ratio of HCl : NH_3 is 1 : 1, the limiting reagent is $NH_3(aq)$.

$q = -n(\Delta H) = -0.034(-53.4 \times 10^3) = +1815.6$ J
$q = C\Delta T + mc\Delta T$
$+1815.6 = (15.0 \times \Delta T) + (70.0 \times 4.18 \times \Delta T)$
$\Delta T = \underline{+5.90\ °C}$

Comments:
- Many students did not account for the calorimeter, which would also gain heat from the reaction.

C2(b)(ii) HCl, HNO_3 and H_2SO_4 are strong acids, while CH_3COOH is a weak acid.

$HCl + NaOH \rightarrow NaCl + H_2O$
$HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$
For neutralisation reactions using 1 mol of strong monobasic acid (HCl and HNO_3) 1 mol of water is formed and the enthalpy change of reaction for 1 mol of each acid is −58 kJ mol^{-1}.

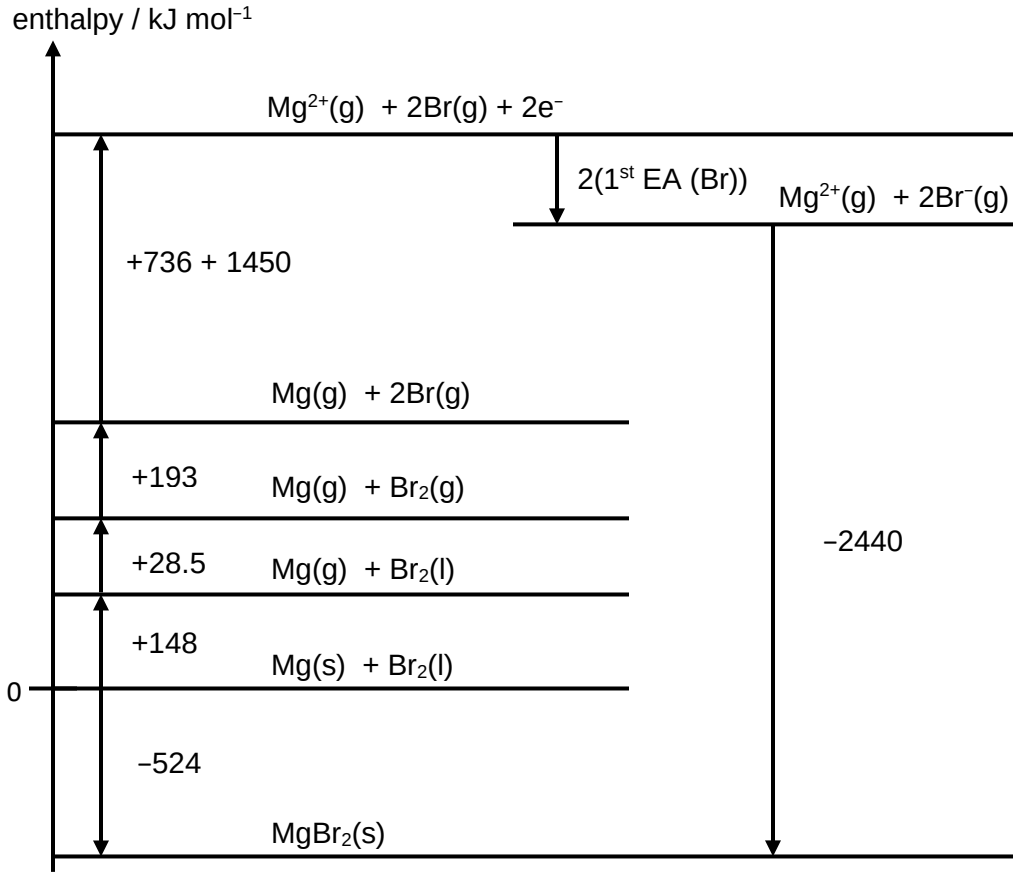
$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
For neutralisation reaction with 1 mol of H_2SO_4, a strong dibasic acid, 2 mol of water is formed. Hence, q and ΔH will be approximately double that of a strong monobasic acid.

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
For neutralisation reaction with 1 mol of CH_3COOH to form 1 mol of water, additional energy is required for the ionisation of CH_3COOH. Hence the magnitude of q and ΔH will be smaller that of a strong monobasic acid.

Comments:

- To compare between strong monobasic and strong dibasic acids, reference should first be made to them as strong acids, and then to how one is monobasic while the other is dibasic. To compare between the strong monobasic and weak monobasic acids, reference should be made to the similarity of being monobasic and then the difference where the weak acid requires energy for ionisation.
- Some misconceptions included thinking that sulfuric acid is the 'strongest' acid. A strong acid completely ionises without the need for energy to do so. Sulfuric acid, nitric acid and hydrochloric acids are all strong acids.
- Some also mistakenly thought that the number of H atoms represented the number of H^+ ions that would be formed in solution. CH_3COOH is a monobasic acid. Only the -COOH group ionises to form $-COO^-$ and H^+.
- Some mentioned that hydrochloric acid is basic, whereas the correct terminology is monobasic.

C2(c)(i)



−524 = +148 + 28.5 + 193 + 736 + 1450 + 2(1st EA(Br)) −2440
1st EA (Br)= <u>−320 kJ mol−1</u>

Comments:

- The standard state of Br_2 is liquid, not gaseous. Hence, $Br_2(l)$ has to be vaporised to $Br_2(g)$, before applying the BE(Br-Br) to break the gaseous Br-Br bond to form 2Br(g) atoms. Recall that the definition of bond energy involves the breaking of 1 mol of bonds in the <u>gaseous state</u>.

- Some students mistook enthalpy change of vaporisation of $Br_2(l)$ as BE(Br-Br) or enthalpy change of atomisation of Br.
- Students should label '0' for the energy level containing only elements in their standard states.
- Note that the cation (Mg^{2+}) should be formed before the anion (Br^-), i.e. ionisation energy should precede electron affinity.
- Some students did not realise that the first electron affinity is exothermic and the arrow should be pointing downwards.
- Arrows need to be drawn from one energy level to another, and should not be floating in mid-air.
- There should not be a combination of multiple, unrelated steps in the energy level diagram.
- Some other common mistakes include:
    - missing out the $2e^-$ after the ionisation of Mg(g)
    - forgetting to multiply 1st EA (Br) by 2 (since 2 mol of Br(g) atoms are involved)
    - missing out the sign for the final answer

C2(c)(ii) The magnitude of the lattice energy of magnesium bromide is <u>smaller</u> than that of magnesium nitride as the ionic charge of Br^- is smaller than that of N^{3-}, and ionic radius of Br^- is larger than that of N^{3-}.

$$|LE| \propto \left| \frac{q_+ q_-}{r_+ + r_-} \right|$$

Comments:
- Although the question clearly asked for <u>magnitude</u> of LE, many students referred to LE, omitting the word 'magnitude', making it unclear whether they were referring to the magnitude or the value on the number line. Answers need to state clearly that the magnitude of LE of $MgBr_2$ is smaller or the LE of $MgBr_2$ is less exothermic.
- The nitride ion is N^{3-} (NOT N^- or NO_3^-).
- Many students referred to the Br^- ion as Br or bromine. The correct terminology should be used: bromide ion or Br^- ion (NOT bromine ion or atom).
- Some students used charge density in their answer, which is incorrect.

C2(c)(iii) As Br^- ion has a larger and more polarisable electron cloud than F^- ion, $MgBr_2$ has greater covalent character than MgF_2. Hence, $MgBr_2$ has a greater discrepancy between experimental LE and that predicted by the ionic model.

Comments:
- Note the use of correct terminology: bromide ion or Br^- ion (NOT bromine ion).
- The terms "polarising power" is used for cations and "polarisability" for anions. It is the cations that polarises the electron cloud of the anions, and not the other way round.
- The question requires a discussion of the greater discrepancy in $MgBr_2$. However, some misread the question and simply mentioned that there was covalent character.