Anglo-Chinese School
(Independent)

YEAR 5 Final Examination 2022
INTERNATIONAL BACCALAUREATE DIPLOMA PROGRAMME
CHEMISTRY HIGHER LEVEL

PAPER 2

Monday **12th September 2022** **2 hours 15 minutes**

INSTRUCTIONS TO CANDIDATES

- Do not open this examination paper until instructed to do so.
- Write your **candidate session number in the box above**.
- A calculator is required for this paper.
- A copy of the Chemistry Data Booklet is required for this paper.
- Write your answers in the boxes provided.
- If you use additional sheets of paper for your answer, attach them to the booklet. Indicate the question number clearly on these sheets.
- All drawings must be in ink.

For examiner's use	
Qn 1	/10
Qn 2	/6
Qn 3	/11
Qn 4	/5
Qn 5	/13
Qn 6	/8
Qn 7	/17
Qn 8	/11
Qn 9	/9
Wrong s.f. /units	
Total	/90

This question paper consists of 26 printed pages including this cover page.

Answer **all** questions. Write your answers in the boxes provided.

1. Chile saltpetre is a mineral found in Chile and Peru, which mainly consists of sodium nitrate, $NaNO_3$. The mineral is purified to concentrate the $NaNO_3$ which is used as a fertilizer and in some fireworks.

To determine the purity of a sample of $NaNO_3$, a (1.64 ± 0.02) g impure sample was heated in NaOH (aq) with Devarda's alloy which contains aluminium. This reduces the $NaNO_3$ to ammonia which is boiled off and then dissolved in acid as shown in the following equation.

$$3NaNO_3\,(aq) + 8Al\,(s) + 5NaOH\,(aq) + 18H_2O\,(l) \rightarrow 3NH_3\,(g) + 8NaAl(OH)_4\,(aq)$$

The NH_3 gas produced is dissolved in (25.00 ± 0.06) cm^3 of 1.00 mol dm^{-3} H_2SO_4.

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

The resulting solution was titrated with 2.00 mol dm^{-3} NaOH to determine the amount of unreacted H_2SO_4. (16.20 ± 0.10) cm^3 of NaOH were required for complete neutralization.

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

(a) (i) Calculate the amount of H_2SO_4 present in 25.0 cm^3 of 1.00 mol dm^{-3} H_2SO_4. [1]

$n(H_2SO_4)$ = 25/1000 x 1.00 = 0.0250 mol

Marker's comments:
Well answered. 1m was deducted for incorrect unit (eg mols).

(ii) Calculate the amount of NaOH present in 16.2 cm^3 of 2.00 mol dm^{-3} NaOH. [1]

$n(H_2SO_4)$ = 16.2/1000 x 2.00 = 0.0324 mol

Marker's comments:
Well answered.

(b) (i) Use your answers in **(a)(i)** and **(a)(ii)**, determine the amount of $NaNO_3$ that is present in the sample. [2]

(If you do not have answers for **(a)(i)** and **(a)(ii)**, use n(H_2SO_4) = 0.0500 mol and n(NaOH) = 0.0250 mol, but these are not the correct answers).

n(H_2SO_4)$_{reacted}$ = 0.0250 - ½ (0.0324) = 0.00880 mol

n($NaNO_3$) = n(NH_3) = = 2 x 0.00880 = 0.0176 mol

Marker's comments:
Generally well answered. Some common mistakes include:

1. calculating no. mol of H_2SO_4 with incorrect value from part (a).
2. oversight on taking into account of back titration.
3. missing out the mole ratio in the chemical equations involved.

(ii) Calculate the percentage uncertainty for the amount $NaNO_3$ determined in **(b)(i)**. [2]

(If you do not have answers for **(a)(iii)**, use n($NaNO_3$) = 0.0100 mol but this is not the correct answer).

$$\frac{0.06}{25}(0.025) + \frac{0.10}{16.20}(\frac{1}{2} \times 0.0324) = 0.00016$$

$$\% \text{ unc for } n(NaNO_3) = \frac{0.00016}{0.00880} \times 100$$

$$= 1.818\% = \pm 1.8\% \text{ (2sf; <2\%)}$$

Marker's comments:
Poorly answered. Many candidates did not realise that the sum of absolute uncertainties is required as the corresponding calculation step in (b)(i) was subtraction.
Several candidates used the incorrect values for calculation and were not awarded full marks though a correct answer was obtained.
Candidates also overlooked the correct significant figure for the final calculated percentage uncertainty, which is 2 sf if it is <2%.

(iii) Use your answer in **(b)(i)** to determine the percentage composition of $NaNO_3$ present in the impure sample. [2]

(If you do not have answers for (c), use n($NaNO_3$) = 0.0100 mol but this is not the correct answer).

% composition ($NaNO_3$) = [0.0176 mol x M_r($NaNO_3$) / 1.64] x 100

= 91.21951 ≈ 91.2% (3 sf)

Marker's comments:
Generally well answered. However a few candidates were not familiar with the formula and performed incorrect calculation. Some candidates also calculated the percentage composition in moles instead of in mass, which should be discouraged.

(iv) Calculate the absolute uncertainty of the percentage composition of $NaNO_3$. [2]

$[0.91\% + \frac{0.02}{1.64} \times 100\%)]$ **x 91.21951**

= 1.942 ≈ ±2% (1sf)

Deduct 1m for incorrect sf (overall)

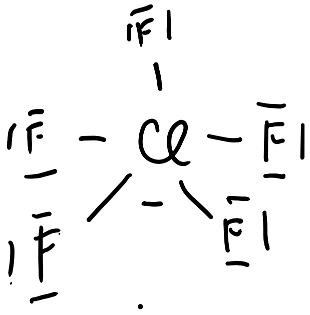
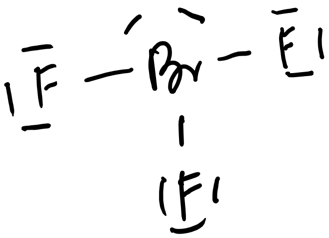
Marker's comments:
Fairly well answered. Some candidates were unable to obtain full marks as they were supposed to determine the absolute uncertainty by multiplying the percentage uncertainty with the percentage composition value. Many also did not leave the final absolute uncertainty in 1 sf.

2. (a) Fluorine reacts with other elements in Group 17 to form interhalogen compounds. The formulae of two such compounds are as shown.

(i) Draw the Lewis structure. [1]

(ii) State the correct bond angle. [1]

(iii) State the molecular geometry for each molecule. [2]

	ClF_5	BrF_3
Lewis structure		
Bond angle of F-X-F	**81° to 89°**	**81° to 89°** **Accept 110° to 119° or slightly >180°**
Molecular Geometry	**Square pyramidal**	**T-shape**

Marker's comments:
Poorly attempted. Common mistakes include:

1. missing lone pairs on the central and terminal atoms
2. not stating a specific bond angle (the term <90º is not allowed for question command term "state")
3. unfamiliar with the assigned shape given the no. bond pairs and lone pairs.

(b) The boiling point of the molecules are given in the table. [2]

compound	ClF_5	BrF_3
boiling point / °C	-13.1	125.7

Suggest why the boiling points of the molecules are significantly different.

Both molecules are polar and thus there is presence of permanent dipole-dipole forces between the molecules. However, the permanent dipole-dipole force between BrF_3 molecules is stronger than that in $C\mathit{l}F_5$ due to greater dipole as a result of larger electronegativity difference between Br-F than Cl-F. Hence more energy is required to overcome the dipole-dipole forces between BrF_3 molecules.

Marker's comments:
Poorly attempted. Most candidates were unable to identify the correct reasons behind the variation in boiling point by stating irrelevant justifications including polar vs non polar molecules, or varying in London dispersion force due to different Mr.

3. (a) (i) Explain why copper is considered a transition metal while zinc is not. [2]

Copper forms Cu^{2+} ions while Zn forms Zn^{2+} ions.

For Cu^{2+}, it has a partially filled 3d orbitals while for Zn^{2+}, the 3d orbitals are fully filled.

Marker's comments:
Very poorly attempted. Most of the students did not know the definition of transition elements and were unable to answer this question.

(ii) Explain why complexes of Zn^{2+} (aq) are colourless whereas complexes of Cu^{2+} (aq) are coloured. [4]

For Zn^{2+} complexes, the 3d orbitals are fully filled and no d-d transition can take place.

For Cu^{2+} complexes, the presence of ligands split the 3d orbitals into 2 sets of different energies✓. Cu^{2+} has partially filled d-orbitals.

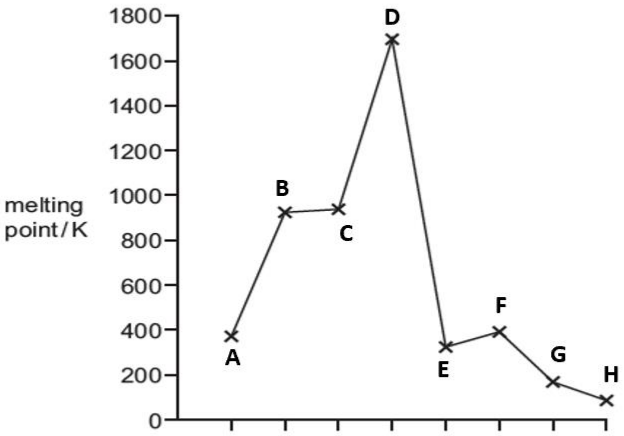
The 3d electron from the lower energy d orbital can absorb light of a particular wavelength and move up to a higher energy d orbital. ✓

Energy gap corresponds to light in the visible region of the spectrum. ✓

The colour seen is the complement of the colour absorbed. ✓

Marker's comments:
In general, students did well to explain why Cu^{2+} complexes are coloured. However, there are misconceptions regarding why Zn^{2+} complexes are not coloured such as Zn^{2+} ion unable to accommodate the incoming ligands; Zn^{2+} complexes do not undergo any d-d splitting.

A plot of the melting points of the period 3 elements **A – H** is shown below.



(b) (i) Element **C** forms an amphoteric oxide. Construct two chemical equations to indicate the amphoteric nature of the oxide of **C**. [2]

$C_2O_3 + 6HCl \rightarrow 2CCl_3 + 3H_2O$

$C_2O_3 + 3H_2O + 2NaOH \rightarrow 2NaC(OH)_4$

(accept A*l* as C)

Marker's comments:
In general, students can write the equation to show how the oxide reacts with the acid but unable to write the equation to show how the oxide reacts with the base.

(ii) Explain the significant difference in the melting point between elements **D** and **E**. [3]

Element D has a giant molecular structure while element E has a simple molecular structure. The atoms in element D are held by strong covalent network while there is presence of weak LDF (accept VDW) between molecules of element E. Thus more energy is required to overcome the strong covalent network than weak LDF.

Marker's comments:
Generally well answered.

4. A novel idea to extract oxygen gas can be carried out by using a moon rock and heating it to very high temperatures, by concentrating the Sun's rays onto the sample. Most moon rocks are composed of metal oxides and silicon dioxides. At these temperatures, all the bonds in the sample are broken and the sample atomises. Some of the atoms may be ionised into positively charged ions. The ions may then be separated using magnetic or electric fields. Positively charged ions can be deflected in an electric field and the relationship is given by:

$$\text{Angle of deflection} \propto \frac{charge}{mass}$$

For a monoatomic, gaseous element at one atmosphere and at a temperature T (in K), the fraction, α, of the atoms that are ionised may be approximated by the equation:

$$\alpha = \sqrt{\frac{Ce^{-E_i/RT}}{T^{-\frac{5}{2}} + Ce^{-E_i/RT}}}$$

where E_i is the first ionization energy of the element in J mol^{-1}
R is the molar gas constant = 8.314 J K^{-1} mol^{-1}
T is temperature in kelvins
and C is a constant with a numerical value of 6.58×10^{-7}

(a) Rock sample 1 is composed of the elements O, Na, Mg, Al and Si. By using their relative atomic mass, calculate the $\frac{charge}{mass}$ ratio for the following ions and state which ion will deflect the most in an electric field. [2]

Positive Ions	**Charge/mass ratio**
O^+	**0.0625**
Na^+	**0.0434**
Mg^{2+}	**0.0833**

Marker's comments:
Generally well answered. A number of students left the charge/mass ratio in fraction form and not in decimal form.

(b) Rock sample 2 contains the elements Ti, Al, Mg, Si and O. Given the first ionization energy of oxygen is 1314 kJ mol^{-1}, use the equation given to calculate the percentage of the oxygen atoms that would be ionized at this temperature of 10000 K. [3]

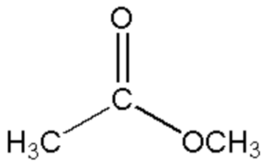
From the equation, first ionisation energy of O = 1.314×10^6 J mol^{-1}

Substitute values into α and calculate α = 0.02999

Percentage ionised = 2.99%

Marker's comments:
Generally well answered. Students should be able to get the correct value if the ionization energy is converted to J mol^{-1}.

5. (a) Simple esters such as methyl ethanoate and ethyl ethanoate are flammable liquids. The structure of methyl ethanoate, $C_3H_6O_2$, is shown below.



(i) Define the standard enthalpy change of formation of methyl ethanoate. [1]

Enthalpy change when one mole of methyl ethanoate is formed from its constituent elements at standard states / 1 bar / conditions

OR $3C(s) + 3H_2(g) + O_2(g) \rightarrow C_3H_6O_2(l)$

Marker's comments:
Very well done, most students can give the key words in the definition. But a small handful of students seem to think that elements are in the atomic form.

(ii) Calculate the standard enthalpy change of formation of methyl ethanoate. [2]

	$\Delta H^{\ominus}_{c}$ / kJ mol^{-1}
carbon	−393.5
hydrogen	−285.8
methyl ethanoate	−1592.1

$\Delta_f H^{\ominus} = 3 \times (-393.5) + 3\,(-285.8) - (-1592.1) = -445.8$ kJ mol^{-1}

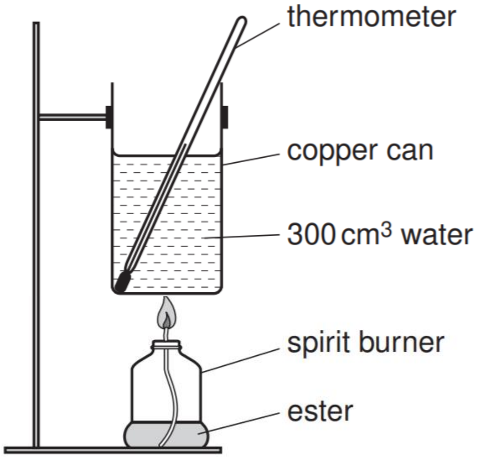
OR manipulation by equations
$3C(s) + 3H_2(g) + O_2(g) \rightarrow C_3H_6O_2(l)$

$3C(s) + 3O_2(g) \rightarrow 3CO_2(g)$	**$3 \times (-393.5)$**
$3H_2(s) + 1\tfrac{1}{2}O_2(g) \rightarrow 3H_2O(l)$	**$3 \times (-285.8)$**
$3CO_2(g) + 3H_2O(l) \rightarrow C_3H_6O_2(l) + 3\tfrac{1}{2}O_2(g)$	**$-(-1592.1)$**

Marker's comments:
Not well answered. Common mistakes included

- use the wrong coefficient of the elements in forming the ester
- apply the formula $\Delta H^\theta_{reaction} = \Sigma\ \Delta H^\theta_f$ (products) – $\Sigma\ \Delta H^\theta_f$ (reactants) without realizing that $\Delta H^\theta_{combustion}$ data are given in this context
- a small handful of students even added in bond enthalpy of oxygen in the consideration

(b) A student used the experimental set-up shown to carry out experiments to determine the standard enthalpy change of combustion for methyl ethanoate.



The mass of copper can is 250 g and the volume of water is 300 cm^3.

(i) An experiment was carried out using methyl ethanoate. The ester was combusted in a spirit burner underneath a copper can so that the flame from the burner heated 300 cm^3 of water in the can. It was found that 0.980 g of ester was required to raise the temperature of the water in the can by 9.0 °C. [2]

Calculate the total thermal energy in kJ gained by the water and the copper can in this initial experiment.

The specific heat capacities of water and copper are 4.18 and 0.384 $J\ g^{-1}\ K^{-1}$, respectively. Assume that the water and copper are in thermal equilibrium with each other.

q = 4.18 x 300 x 9 + 0.384 x 250 x 9
= 13500 J = 12.15 kJ

Marker's comments:
Generally well done. Common mistakes included

- did not consider the thermal energy taken up by the copper can
- took 0.980 as the mass of water and copper can
- convert ΔT of 9 °C = (273 + 9) K, note that ΔT in °C = ΔT in K as it is the change in temperature, not absolute temperature
- multiplying the total mass and total specific heat capacity of water and copper can

(ii) Given that the heat transfer to the entire experimental set-up is 65 % efficient and using your answer in (b)(i), calculate the standard enthalpy change of combustion for methyl ethanoate. [2]

(If you do not have answer for (b)(i), use 10.0 kJ, but this is not the correct answer.)

q from reaction = $12.15 \Big/ \frac{65}{100}$ = 18.69 kJ

$\Delta H^{\ominus}_c = -18.69 \Big/ \frac{0.980}{74.09} = -1413 = -1410\ \text{kJ mol}^{-1}$

Alternative answer:

q from reaction = $10.0 \Big/ \frac{65}{100}$ = 15.385 kJ; $\Delta H^{\ominus} = -15.385 \Big/ \frac{0.980}{74.09} = -1163 = -1163\ \text{kJ mol}^{-1}$

Marker's comments:
Generally well done although a handful of students either did not consider the heat transfer efficiency despite given in the question or did not seem to understand the idea of efficiency. Some students forgot to include the negative sign for exothermic combustion reaction. A small number of students use the amount of water as the reactant in this combustion reaction.

(iii) With reference to standard enthalpy change of combustion given in (a)(ii), calculate the percentage error in the experimentally determined value of $\Delta_c H^{\ominus}$ for methyl ethanoate in (b)(ii). [1]

(If you do not have answer for (b)(ii), use the magnitude of 1250 kJ mol^{-1}, but this is not the correct answer.)

Percentage error = $\left|\frac{1413 - 1592.1}{1592.1}\right|$ x 100 = 11.2 %

Alternative answer: Percentage error = $\left|\frac{1250 - 1592.1}{1592.1}\right|$ x 100 = 21.5 %

Marker's comments:
A mixture of answers obtained for this question, showing the lack in understanding of percentage error. Many students use the ΔH^{θ}_f data calculated in (a)(ii) as the literature value when question is asking for ΔH^{θ}_c. Some students were obviously confused between percentage error and percentage uncertainty.

(iv) Suggest **one** source of systematic error in the experiment that would lead to the deviation from the literature value. [1]

- **Incomplete combustion of ester**
- **Evaporation of ester / water during heating**
- **Other sensible suggestion for the lower-than-expected value**

Mark not awarded for:

- **heat loss to surrounding (accounted in heat transfer efficiency)**
- **heat gained by apparatus (copper can) (accounted in the calculation)**

Marker's comments:
Very poorly done as majority of students use heat loss as the source of error despite heat transfer efficiency and heat capacity of the copper cup being considered in the calculation. Students should be aware of other experimental limitations other than heat loss in calorimetric experiments. Students are also discouraged from suggesting errors due to experimenter's poor technique or equipment fault, unless suggested in the question.

(v) Suggest **one** improvement to the experimental set-up that could increase the accuracy of the raw data recorded in these experiments. [1]

- **Put a lid on the calorimeter**
- **Add insulation around the side and/or top of the calorimeter**
- **Stir the water in the copper pot**
- **Draw hot vapour from the flame through a calorimeter using suction**
- **Put a cap on the spirit burner when it isn't lit to avoid evaporative losses**
- **draft excluders**
- **Do repeats and take an average of consistent results to eliminate outliers**
- **Other sensible refinement**

Mark not awarded for:

- **improving the thermometer / use a more precise thermometer**
- **comments about height of the can above the burner**
- **use of a different burner or different material for the can**

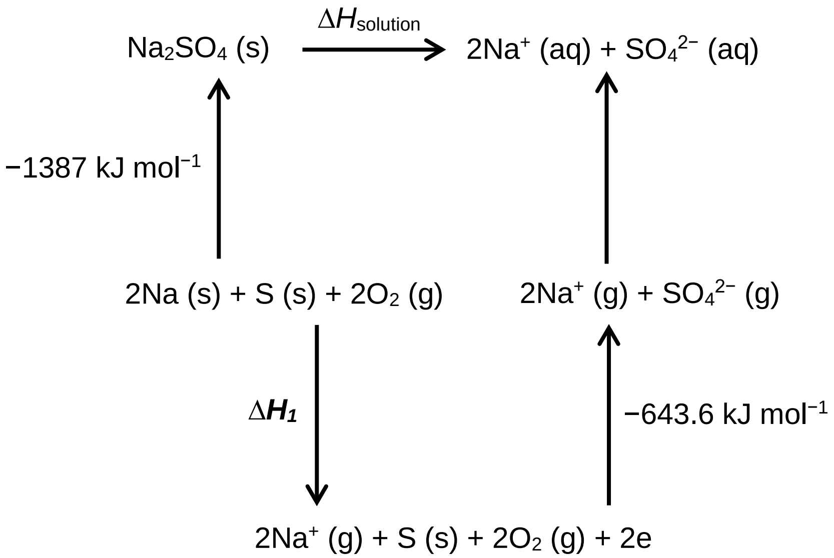
Marker's comments:
Generally well done as marking is very lenient.

Note that for practical purpose, the use of insulation in this context should only be on the side and top of the can, as heat transfer is through the bottom.

Likewise, the change of material for the copper can is not practical, as a container with higher heat capacity will take longer to reach a measurable ΔT, allowing for more heat loss while one with lower heat capacity will experience faster heat loss.

Suggestions should reduce systematic errors to improve the accuracy of the results.

(c) The energy cycle of sodium sulfate is shown.



(i) Given that the $\Delta H_{atomisation}$ of sodium is +107 kJ mol^{-1}, calculate a value for ΔH_1 using section 8 of the data booklet. [1]

ΔH_1 = 2 (107 + 496) = (+) 1206 kJ mol^{-1}

Marker's comments:
Generally well done. Students either forgot to multiply by the coefficient of 2 or use first electron affinity as first ionization energy.

(ii) Determine the enthalpy change of solution of Na_2SO_4, in kJ mol^{-1}, using your answer to (c)(i) and section 20 of data booklet. [2]

(If you do not have answer for (c)(i), use the magnitude of 1000 kJ mol^{-1}, but this is not the correct answer.)

$\Delta H_{solution}$ = 1387 + 1206 + −643.6 + 2(−424) + −1099
= +2.4 kJ mol^{-1}

Alternative answer:

$\Delta H_{solution}$ = 1387 + 1000 + −643.6 + 2(−424) + −1099 = −203.6 kJ mol^{-1}

Marker's comments:
Generally well done. Students who are able to obtain the standard enthalpy change of hydration of 2 moles of Na^+ (g) and SO_4^{2-} (g) are able to get the correct answer.

6. The elements of Group 14 can form monoxides and dioxides. The stabilities of the monoxides, with respect to disproportionation into the element and the dioxide, vary. The equations for some of the disproportionation reactions are given below together with some thermodynamic data for the reactions.

disproportionation equation	$\Delta H^\ominus$ / kJ mol^{-1}	$\Delta S^\ominus$ / J K^{-1} mol^{-1}	$\Delta G^\ominus$ / kJ mol^{-1}
$2CO(g) \rightarrow C(s) + CO_2(g)$	−172.5	−175.9	−120.1
$2SiO(g) \rightarrow Si(s) + SiO_2(s)$	−711.5	−362.9	−603.4
$2GeO(s) \rightarrow Ge(s) + GeO_2(s)$	−126.8		
$2PbO(s) \rightarrow P(s) + PbO_2(s)$	+157.2	−4.000	+158.4

(a) Explain why the magnitude of the entropy change for the disproportionation of

(i) SiO is much bigger than for CO, [2]

For CO change is from 2 moles of gas to 1 mole of gas and 1 mole of solid whereas for SiO change is from 2 moles of gas to 2 moles of solid.

Decrease in disorder / randomness is greater for SiO.

Marker's comments:
Mostly well answered. Students who did not get the full mark simply compared the final state of both systems rather than looking at the change over the two systems. Likewise, numerous students did not elaborate on how the change in the number of moles of gas particles affected the disorderliness of the system, hence the entropy.

(ii) PbO is close to zero. [1]

2 moles of solid produce 2 moles of solid; negligible change in disorderliness.

Marker's comments:
Very well done.

(b) The standard molar entropies for germanium and its oxides are provided.

Species	$S^{\ominus}$ / J K^{-1} mol^{-1}
germanium, Ge(s)	31.1
germanium monoxide, GeO(s)	50.0
germanium dioxide, GeO_2(s)	55.3

(i) Calculate the standard entropy change, $\Delta S^{\ominus}$, for the disproportionation of germanium monoxide. [1]

$\Delta S^{\ominus} = 55.3 + 31.1 - 2 \times 50.0 = -13.6$ J K^{-1} mol^{-1}

Marker's comments:
Surprisingly, this was not too well answered. Majority of the students either did not consider the coefficient of the reactant or mixed up the entropy value of GeO and GeO_2.

(ii) Using your answer to **(b)(i)**, calculate the standard free energy change, $\Delta G^{\ominus}$, for the disproportionation of germanium oxide and determine its spontaneity at 25 °C. [2]

(If you do not have answer for **(b)(i)**, use −30 J K^{-1} mol^{-1}, but this is not the correct answer.)

$\Delta G^{\ominus} = -126.8 - (298 \times -13.6 / 1000) = -123$ kJ mol^{-1}

Reaction is spontaneous since $\Delta G^{\ominus}$ is negative.

Alternative answer:

$\Delta G^{\ominus} = -126.8 - (298 \times -30 / 1000) = -118$ kJ mol^{-1}

Marker's comments:
Common mistakes like not considering the units for ΔS and dividing by 1000, and not answering the question fully. Other than that, well done.

(iii) Predict the effect of increasing the temperature on the spontaneity of the disproportionation of germanium monoxide. [2]

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$

Both enthalpy change and entropy change are negative, increasing the temperature would make $-T\Delta S^{\ominus}$ more positive and hence $\Delta G^{\ominus}$ to be more positive.

Reaction becomes non-spontaneous / less spontaneous at higher temperature.

Marker's comments:
Most students are able to state the effect of temperature on the spontaneity of the reaction correctly, but the explanation is lacking in terms of considering the signs of ΔH and ΔS.

7. (a) In a Cannizzaro Reaction between an aldehyde, RCHO, in presence of a strong base such as sodium hydroxide, its respective carboxylate salt $RCOO^-$ and alcohol, RCH_2OH, were formed.

$$2RCHO + OH^- \rightarrow RCOO^- + RCH_2OH$$

The kinetic study of the reaction could be followed by measuring the rate of disappearance of hydroxide ions by varying initial concentrations of the aldehyde, RCHO and hydroxide ions, OH^- at a constant temperature, T_1.

Experiment number	[RCHO] / mol dm^{-3}	[OH^-] / mol dm^{-3}	Initial rate of disappearance of OH^- / mol dm^{-3} s^{-1}
1	0.050	0.100	3.3×10^{-6}
2	0.100	0.050	6.3×10^{-6}
3	0.100	0.100	1.3×10^{-5}
4	0.200	0.150	

(i) Deduce with reasoning, the order of reaction with respect to RCHO and OH^-. Hence, deduce the overall order of the reaction. [3]

Comparing experiments 1 and 3, by doubling the concentrations of RCHO while keeping the concentration of OH^- constant, the rate increases by 4 times. AND The order of reaction with respect to RCHO is 2;

Comparing experiments 2 and 3, by doubling the concentrations of OH^- RCHO while keeping the concentration of RCHO constant, the rate increases by 2 times. AND The order of reaction with respect to OH^- is 1;

Overall order of reaction = 3;

Marker's comments:
Generally well answered, but there were a number of students who did not give overall order of reaction despite finding the individual order of reactions.

(ii) State the rate expression. [1]

Rate = k[RCHO]^2[OH$^-$]

Marker's comments:
Generally well answered.

(iii) Calculate the value of *k* and state its units clearly. [2]

Using expt 1: $3.3 \times 10^{-6} = k[0.050]^2[0.100]$, k= <u>0.013 mol^{-2} dm^6 s^{-1}</u> ***or***

Using expt 2: $6.3 \times 10^{-6} = k[0.100]^2[0.050]$, k= <u>0.013 mol^{-2} dm^6 s^{-1}</u> ***or***

Using expt 3: $1.3 \times 10^{-5} = k[0.100]^2[0.100]$, k= <u>0.013 mol^{-2} dm^6 s^{-1}</u> ***or***

Marker's comments:
Most students were able to do it and have substituted the correct values, however the students' calculated values were different which indicate some misuse of calculator.

(iv) Calculate the rate of disappearance of OH$^-$ for experiment 4. [1]

Using Experiments 2 and 4, Rate of disappearance = $6.3 \times 10^{-6} \times 3 \times 2^2$ = <u>7.6×10^{-5} mol dm^{-3} s^{-1}</u>
or
Using Experiments 1 and 4, Rate of disappearance = $3.3 \times 10^{-6} \times 1.5 \times 4^2$ = <u>7.9×10^{-5} mol dm^{-3} s^{-1}</u>
or
Using Experiments 3 and 4, Rate of disappearance = $1.3 \times 10^{-5} \times 1.5 \times 2^2$ = <u>7.8×10^{-5} mol dm^{-3} s^{-1}</u>

Marker's comments:
Well answered. Some candidates lost mark over careless mistakes, eg incorrect values / operation used for calculation or missing units/ significant figure.

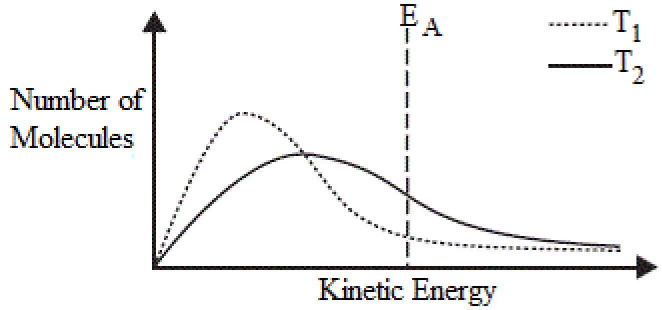
(v) Given that the reaction proceeds in a two-step mechanism, forming $[RCHO(OH)]^-$ as an intermediate, propose a possible mechanism. [3]

Step 1: $RCHO + OH^- \rightarrow [RCHO(OH)]^-$ (fast step)
Step 2: $[RCHO(OH)]^- + RCHO \rightarrow RCOO^- + RCH_2OH$ (slow step)

Marker's comments:
Many candidates did not specify the slow/fast step in the mechanism. Some candidates overlooked that the coefficient of the elementary steps must equal to the overall equation.

(b) (i) An additional experiment was carried out at a higher temperature T_2. [1]

On the same axes, sketch the Maxwell-Boltzmann energy distribution curve at the two temperatures T_1 and T_2, where $T_2 > T_1$. Label your graphs and axes clearly.



(ii) Draw a dotted line to represent the activation energy on your graphs in **(b)(i)**. Explain in terms of collision theory, the effect on rate of reaction when temperature is increased. [3]

At a higher temperature T_2, more particles / higher fraction of particles has energy with $K_E \geq E_a$, this increases in the frequency of effective collisions, leading to in an increase in the rate of reaction.

Marker's comments:
A number of candidates could not draw the Boltzmann distribution curve properly or indicated incorrect units on the graph.
Some candidates also lost marks due to the missing key phrases in the explanation.

(c) Describe how the activation energy of a chemical reaction could be determined experimentally. [3]

Alternative 1 (graphical method)
Carry out the reaction at several temperatures (if a number is indicated, it has to be more than 2) and measure the rate of reaction;
Plot a graph of In *k* against 1/T (K^{-1}) (units of T must be indicated to be in Kelvin);
Accept rate constant in place of *k*.
Determine Ea by using Ea= -gradient x R.

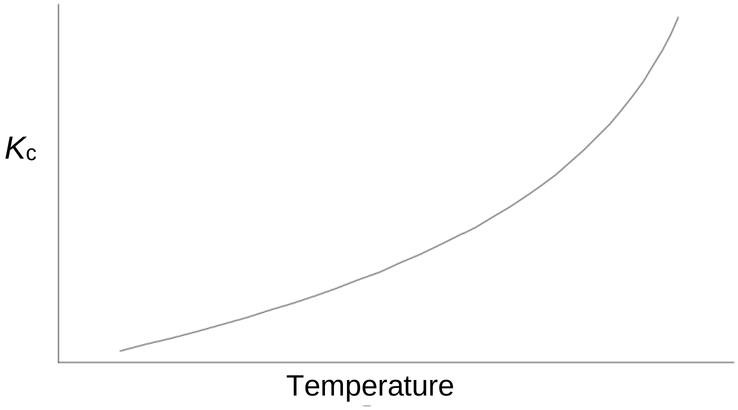
Alternative 2 (2-point method)
Carry out reaction at two temperatures;
to determine two rate constant values; Or rate of reaction at both temperatures;
Determine Ea by using formula $In\frac{k_1}{k_2} = \frac{E_a}{R}(\frac{1}{T_1} - \frac{1}{T_2})$ **or** $\frac{rate_1}{rate_2} = e^{-\frac{E_a}{R}\left(\frac{1}{T_1}-\frac{1}{T_2}\right)}$

Marker's comments:
Most students are able to obtain 2 marks as they are familiar with the use of Arrhenius equation to get E_a. But many did not state how to get the values for plotting experimentally. A small handful of students are confused between rate constant and equilibrium constant.

8. Nitrogen monoxide, NO is formed when nitrosyl chloride, NOCl, dissociates according to the following equation.

$$2NOCl\ (g) \rightleftharpoons 2NO\ (g) + Cl_2\ (g)$$

(a) (i) The sketch graph below shows how the value of K_c for this reaction changes with temperature. [2]



State and explain whether the reaction is exothermic or endothermic.

Reaction is **endothermic**
As **temperature increases**, K_c increases implying **position of equilibrium has shifted right** to favour endothermic reaction.

Marker's comments:
Generally well done.

(ii) State and explain how increasing the pressure of the reaction mixture at constant temperature will affect the position of equilibrium and the equilibrium constant, K_c. [2]

Position of equilibrium shifts **left** to **decrease the number of gaseous particles**.
K_c remains **unchanged**.

Marker's comments:
Students generally know where the position of equilibrium has shifted but a couple of students did not mention the reason why. In addition, some students mentioned that K_c decreases due to the POE shift to the left which is incorrect as K_c is only affected by temperature. Again, a small handful of students did not address whether there is a change in the K_c value.

(ii) At 230 °C, the equilibrium constant, K_c, is 0.0045. [2]

At a given time, 0.400 mol NOCl, 0.300 mol NO and 0.100 mol Cl_2 were mixed in a 2.00 dm^3 flask.

Calculate the reaction quotient, Q, and deduce, showing your reasoning, if the forward or the reverse reaction is favoured at this time.

$Q=\frac{(0.150)^2(0.0500)}{0.200^2}=\underline{\mathbf{0.0281}}$

Since **$\underline{Q > K_c}$**, the **reverse reaction is favoured**.

Marker's comments:
There's an alarmingly high number of students who did not convert amount to concentration and hence were not able to obtain the correct numerical value of Q. A small number of students compared K_c <1 instead of to Q to determine which reaction is favored. A small number of students also mistakenly think that Q> K_c will favor the forward reaction.

(iii) As a first step in the manufacture of nitric acid it has been suggested that nitrogen monoxide, NO, can be formed from nitrogen and oxygen in a reversible reaction. [2]

The value of K_c for this reaction is 1×10^{-5} at 1500 K.

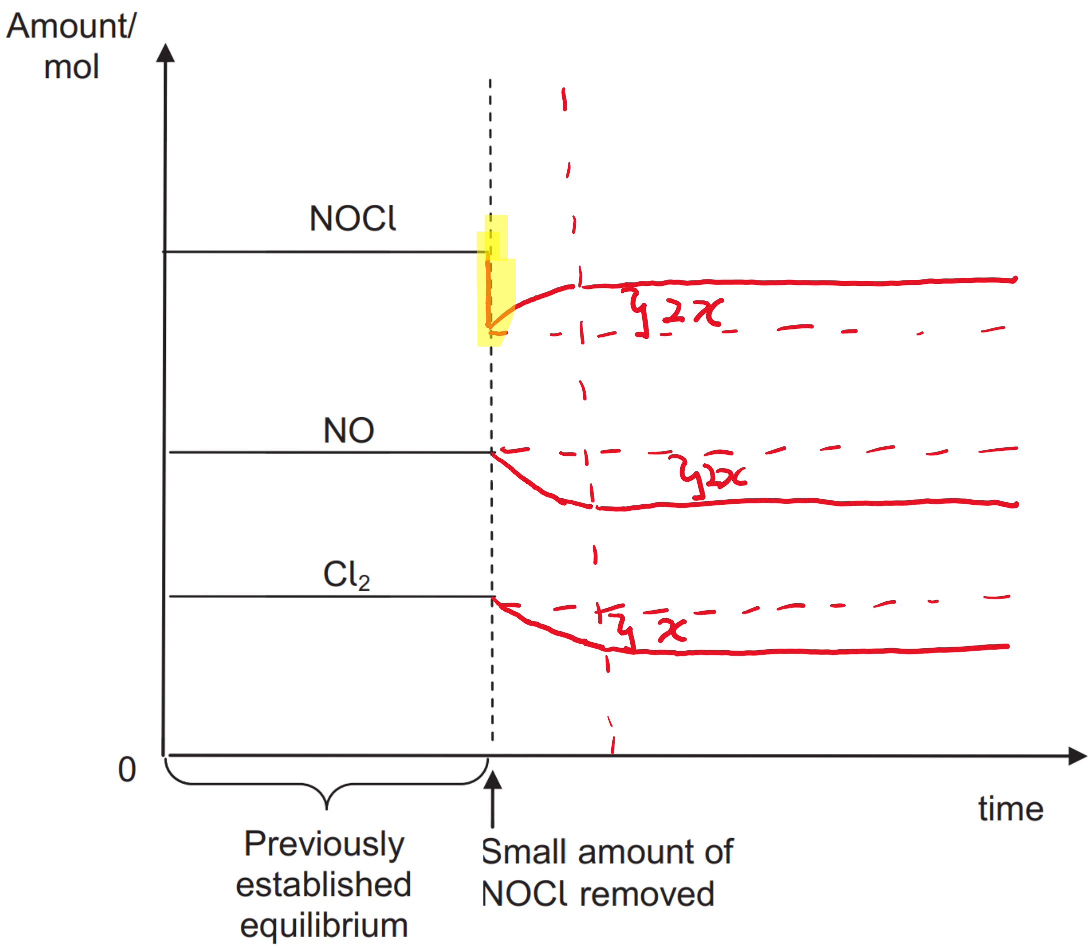
Explain the significance of this value for an industrial chemist interested in manufacturing nitrogen monoxide by the direct combination of the elements.

The product **yield** is very **small**.
Yield does not justify **cost of producing it at high temperature**.

Marker's comments:
Since the K_c value is given, there is a need to comment on the value which quite a number did not. In addition, since the K_c at other temperatures are not given, one could not comment on whether reducing temperature will indeed result in a reasonable yield as a method of production. Hence, students can only conclude that it is not very efficient at the current temperature given the low K_c that it may not be a suitable production method.

(c) A small amount of NOCl is suddenly removed from the equilibrium mixture containing NOCl, NO and Cl_2, keeping the temperature constant. The mixture is then allowed to reach equilibrium again. [3]

Complete the sketch below to show how the amount of each of the three gases would change until a new equilibrium is established.



Marker's comments:
Most students were able to get the sharp dip in [NOCl] due to the removal, and also the change in concentrations of the reactants and the products due to the POE shift. However, there is no indication/labelling of when the equilibrium has re-established.

9. The key stage in the manufacture of sulfuric acid is the reaction between sulfur dioxide and oxygen.

$$2SO_2\ (g) + O_2\ (g) \rightleftharpoons 2SO_3\ (g)$$

(a) (i) State the equilibrium constant expression, K_c of this reaction. [1]

$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$

Marker's comments:
Generally well answered

(ii) When a 5.2 mol of SO_2 and 2.6 mol of O_2 were allowed to reach equilibrium in a 2 dm^3 vessel at 500 °C, the amount of SO_3 at equilibrium is 4.8 mol. Calculate the value of K_c at 500 °C. [2]

	$2SO_2$(g)	+	O_2(g)	=	2SO(g)
Initial amount/ mol	5.20		2.60		0
Change	-4.80		-2.4		+4.80
Eqm amount/ **mol**	0.40		0.20		4.80
Concentration	**0.20**		**0.10**		**2.40**

$K_c = \frac{(2.4)^2}{(0.20)^2(0.10)} = \underline{1440}$

Marker's comments:
A high number of students has mistaken the amount of SO_2 and O_2 given as the equilibrium amount and hence were not able to get the correct K_C value. In addition, a high number of students were able to construct the ice table but did not convert mol to concentration.

(b) (i) At exactly 727 ^{o}C, the value of the equilibrium constant is 280. Calculate the standard Gibbs free energy change, $\Delta G^{\ominus}$, for the reaction, in kJ, using sections 1 and 2 of the data booklet. State your answer to three significant figures. [2]

$\Delta G^{\ominus}$=-RTln K_c = -8.31x 1000 x ln(280) =**-46800 J**

= **-46.8 kJ**

Marker's comments:
Some students did not convert from J to kJ despite the question asking for answers to be in kJ. In addition, some students are not aware that the units of $\Delta G^{\ominus}$ is J instead of kJ when using the formula.

(ii) Comment on the value of ΔG when the reaction quotient equals the equilibrium constant, $Q = K$ [2]

ΔG = 0

reaction at equilibrium
OR rate of forward and reverse reaction is the same
OR constant macroscopic properties

Marker's comments:
Some students state that **ΔG** is at its minimum which is not true as it could be negative. The total Gibbs free energy is at its minimum instead.

(c) The reaction could be speed up by using V_2O_5 catalyst. State and explain how the use of the catalyst could affect the yield of SO_3 produced in this reaction. [2]

No change in yield.
Catalyst has **no effect on position of equilibrium** as it **increases the rate of the forward and backward reaction by the same extent**.

Marker's comments:
A handful of students did not address about the yield of SO_3 and a few students gave the definition of the catalyst which was not needed. In addition, some students pointed out that it increases the rate of the forward and backward reaction but did not commit that the rate increase Is equal. This is important as the POE will still shift if the increase is not the same.

2022 Year 5 HL Chemistry Promos Paper 1 – Answers

1	**B**	11	**A**	21	**B**	31	**B**
2	**C**	12	**D**	22	**B**	32	**B**
3	**A**	13	**D**	23	**D**	33	**B**
4	**B**	14	**D**	24	**A**	34	**D**
5	**C**	15	**A**	25	**C**	35	**D**
6	**A**	16	**C**	26	**D**	36	**D**
7	**D**	17	**B**	27	**D**	37	**D**
8	**B**	18	**A**	28	**A**	38	**C**
9	**B**	19	**C**	29	**B**	39	**C**
10	**A**	20	**C/D**	30	**D**	40	**B**