

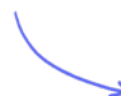
Structured Questions

Entropy & Spontaneity

Entropy (HL) / Calculating Standard Entropy Changes (HL) / Gibbs Free Energy (HL) / Spontaneous Reactions (HL) / Gibbs Free Energy & Equilibrium Constant (HL)

Easy (4 questions)	/38
Medium (5 questions)	/64
Hard (4 questions)	/41
Total Marks	/143

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Easy Questions

1 (a) State whether there is an increase or decrease in entropy for the following reactions.

i) $\text{H}_2\text{O} (\text{l}) \rightarrow \text{H}_2\text{O} (\text{g})$ [1]

ii) $\text{Cl}_2 (\text{g}) \rightarrow \text{Cl}_2 (\text{l})$ [1]

iii) $\text{CaCO}_3 (\text{s}) \rightarrow \text{CaO} (\text{s}) + \text{CO}_2 (\text{g})$ [1]

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(3 marks)

(b) State the equation to determine the entropy change, ΔS , of a reaction.

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(1 mark)

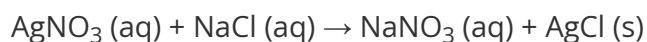
(c) Use your answer to part b) and section 13 of the data booklet to determine the standard entropy change, in $\text{J K}^{-1} \text{mol}^{-1}$, for water boiling.



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(1 mark)

(d) Explain why the entropy change of the following precipitation reaction is negative.



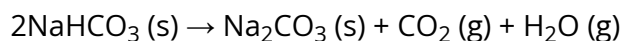
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(1 mark)

(e) State the point of a reversible reaction in which the Gibbs free energy is at its lowest.

(1 mark)

2 (a) Use section 13 of the data booklet and the information below to determine the following:



Compound	ΔH_f (kJ mol ⁻¹)
NaHCO ₃ (s)	-951
Na ₂ CO ₃ (s)	-1131

- i) The sum of the enthalpies of formation, ΔH_f , of the products in kJ mol⁻¹. [1]
- ii) The sum of the enthalpies of formation, ΔH_f , of the reactants in kJ mol⁻¹. [1]
- iii) Use your answers to part i) and ii) to determine the enthalpy change of reaction, ΔH_r , for the decomposition of sodium hydrogen carbonate, NaHCO₃ (s) in kJ mol⁻¹. [2]

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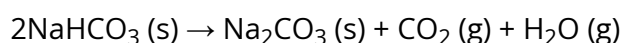
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(4 marks)

(b) Use section 13 in the data booklet and the information below to determine the following:



Compound	S (J K ⁻¹ mol ⁻¹)
NaHCO ₃ (s)	+102
Na ₂ CO ₃ (s)	+135

- i) The sum of the entropies, S , of the products in J K⁻¹ mol⁻¹

[1]

ii) The sum of the entropies, S , of the reactants in $\text{J K}^{-1} \text{mol}^{-1}$.

[1]

iii) Use your answers to part i) and part ii) to determine the standard entropy change for the decomposition of sodium hydrogen carbonate, $\text{NaHCO}_3(\text{s})$, in $\text{J K}^{-1} \text{mol}^{-1}$.

[2]

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(4 marks)

(c) Use your answers to part a) and b) as well as section 1 in the data booklet to determine the free energy change, in kJ mol^{-1} , for the decomposition of sodium hydrogen carbonate, $\text{NaHCO}_3(\text{s})$, at 500 K.

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(2 marks)

(d) Use your answer to part c) to state whether the decomposition of sodium hydrogen carbonate is spontaneous at 500K.

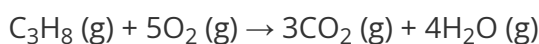
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(1 mark)

3 (a) State the equation which can be used to determine the standard free energy for a reaction change, ΔG , using ΔG formation data.

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(1 mark)

(b) Use section 13 of the data booklet and the equation below to determine the following.



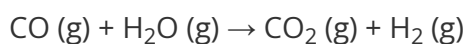
i) The sum of the Gibbs free enthalpy of formation, ΔG_f , of the products in kJ mol^{-1} .
[1]

ii) The sum of the Gibbs free enthalpy of formation, ΔG_f , of the reactants, kJ mol^{-1} .
[1]

iii) The Gibbs free energy change, ΔG , for the combustion of propane, kJ mol^{-1} .
[1]

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(3 marks)

(c) The reaction of carbon monoxide with water is as follows:



The entropy change, ΔS , for this reaction is $-135 \text{ J K}^{-1} \text{ mol}^{-1}$ and the enthalpy change, ΔH , for the reaction is $-41.4 \text{ kJ mol}^{-1}$.

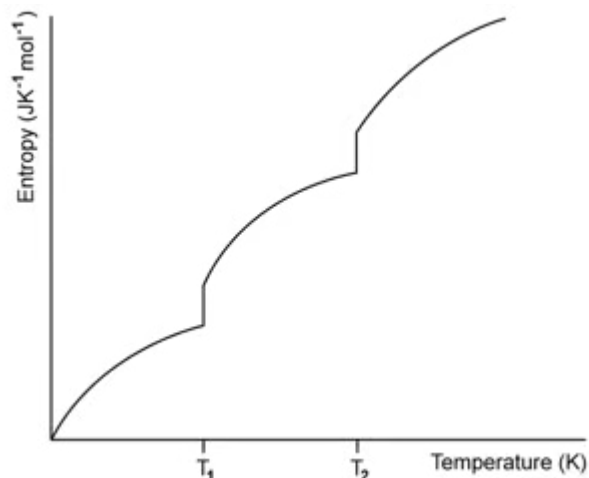
Determine the free energy change, ΔG , for the reaction at 700 K in kJ mol^{-1} .

(2 marks)

- (d)** Use your answer to part b) state whether the reaction of carbon monoxide and water is spontaneous at 700 K.

(1 mark)

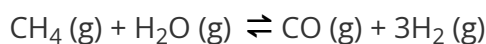
4 (a) The image below shows how the entropy of compound Y varies with temperature



State the changes occurring at T₁ and T₂.

(2 marks)

(b) Study the reaction of methane, CH₄, with water shown.



Use sections 13 and 14 of the data booklet to determine the following.

- i) The sum of the enthalpies of formation, ΔH_f , of the products in kJ mol⁻¹. [1]
- ii) The sum of the enthalpies of formation, ΔH_f , of the reactants in kJ mol⁻¹. [1]
- iii) Use your answers to part i) and ii) to determine the enthalpy change of reaction, ΔH_r , for the reaction kJ mol⁻¹. [2]

(4 marks)

(c) The standard entropy of hydrogen gas is $131.0 \text{ J K}^{-1} \text{ mol}^{-1}$.

Use section 13 in the data booklet to determine the following for the reaction in part a).

i) The sum of the entropies, S , of the products in $\text{J K}^{-1} \text{ mol}^{-1}$ [1]

ii) The sum of the entropies, S , of the reactants in $\text{J K}^{-1} \text{ mol}^{-1}$. [1]

iii) Use your answers to part i) and part ii) to determine the reaction given in part a) in $\text{J K}^{-1} \text{ mol}^{-1}$. [2]

(4 marks)

(d) We can use the equation $\Delta G = \Delta H - T\Delta S$ to determine the temperature, in K, at which a reaction becomes feasible.

Use your answers to part b) and c) as well as section 1 of the data booklet to determine the following.

The equation required to determine the temperature at which a reaction becomes spontaneous

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The temperature at which the reaction of methane, CH_4 , and water becomes spontaneous

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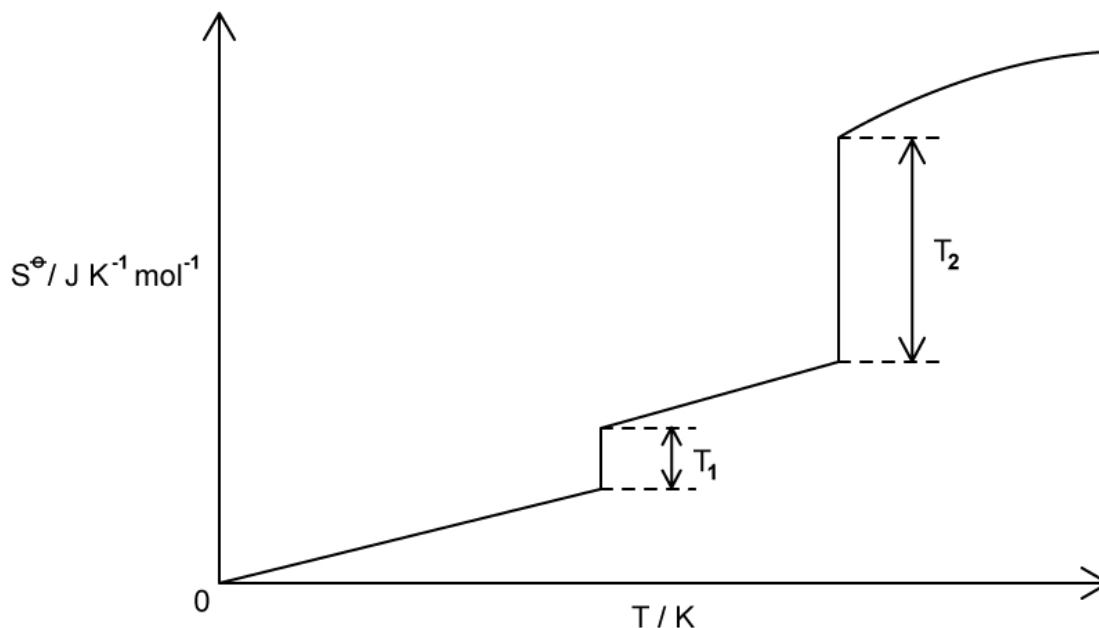
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(3 marks)

Medium Questions

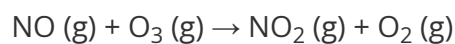
1 (a) This question looks at how the entropy change of water varies with temperature.



- i) The entropy of water is zero when the temperature is zero Kelvin. Explain why, with reference to the water molecules in your answer. [2]
- ii) Explain why the entropy change, ΔS , is larger at temperature T_2 than at temperature T_1 . [2]
- iii) On the figure, draw the boiling point (T_b) of water on the appropriate axis. [1]

(5 marks)

- (b) Standard entropies can be used to calculate the entropy change of a reaction, ΔS . For example, for the formation of nitrogen monoxide from nitrogen and oxygen.



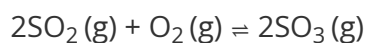
Substance	Entropy value ($\text{J K}^{-1} \text{ mol}^{-1}$)
NO (g)	210.8
O ₂ (g)	205.2
NO ₂ (g)	240.0
O ₃ (g)	238.9

Use the data given to calculate the entropy change of the reaction between nitric oxide and ozone at 298K.

(2 marks)

- (c) The contact process is a method used industrially to form sulfur trioxide, by reacting sulfur dioxide and oxygen together over a vanadium(V) oxide catalyst.

The equation for this reaction is shown below:



Substance	Formation enthalpy values (kJ mol^{-1})
SO ₂ (g)	-297
SO ₃ (g)	-395

i)

Calculate the standard enthalpy change of the contact process reaction using the data provided.

[2]

- ii) The standard entropy change of this reaction is $-189 \text{ J K}^{-1} \text{ mol}^{-1}$. Use this value and your enthalpy value from part (i) to calculate a value for the free energy change for this reaction at 298K.

[3]

- iii) Use your answer to part (ii) to explain whether the reaction is feasible at 298 K.

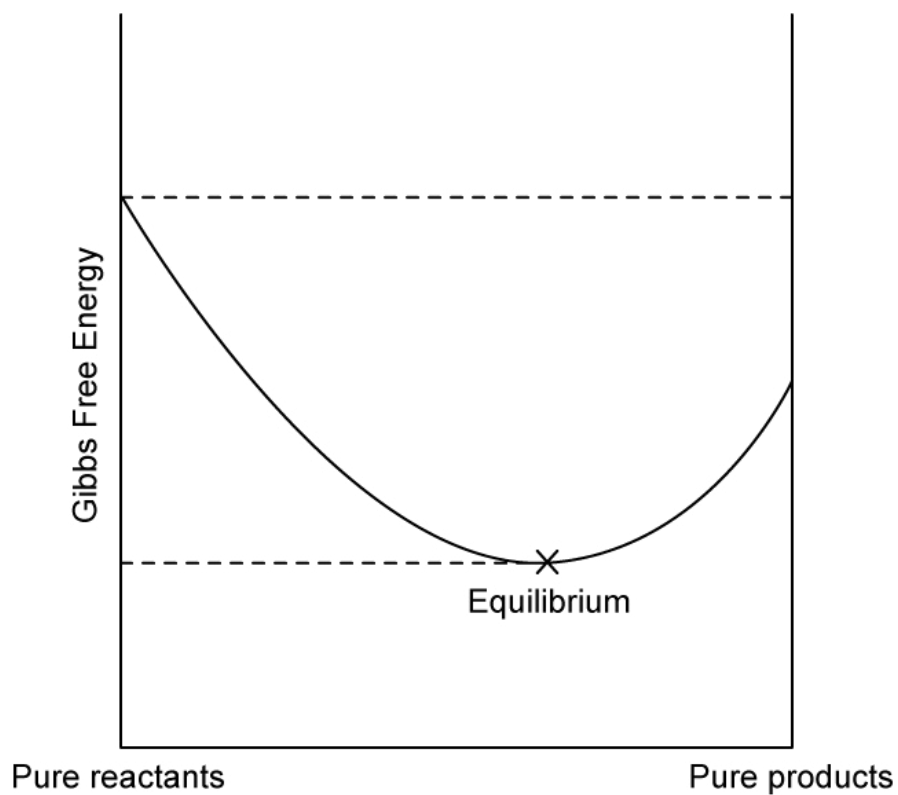
[1]

(6 marks)

- (d) The value for the free energy change is an indication whether the forward or backward reaction is favoured.

The curve that we would expect to see for the reaction between sulfur dioxide and oxygen is shown below.

Explain why the curve for this reaction is shifted to the right hand side.



(2 marks)

- 2 (a) The enthalpy of solution of sodium chloride is $+4 \text{ kJ mol}^{-1}$. Explain why the free energy change for dissolving sodium chloride in water is negative, despite the enthalpy change being a positive value.

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(3 marks)

- (b) Calcium carbonate thermally decomposes to form calcium oxide and carbon dioxide, as shown below:



The enthalpy change of the above reaction is $\Delta H^\ominus = +178 \text{ kJ mol}^{-1}$ and the entropy change is $\Delta S^\ominus = +161 \text{ J K}^{-1} \text{ mol}^{-1}$

Calculate the temperature at which the free-energy change, ΔG^\ominus , for this process is zero.

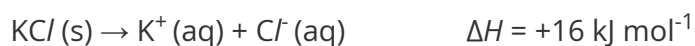
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(3 marks)

- (c) Some ionic compounds such as potassium chloride, KCl , will dissolve in water at room temperature in an endothermic process.



Substance	Entropy value $\text{J K}^{-1} \text{ mol}^{-1}$
$\text{KCl} (\text{s})$	+83
$\text{K}^+ (\text{aq})$	+103
$\text{Cl}^- (\text{aq})$	+57

- i) Using the data provided, prove that this process is feasible at 298 K. [4]
- ii) Use your knowledge of structure and bonding to explain why ΔH^\ominus is positive for this process. [2]

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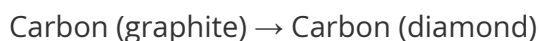
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(6 marks)

(d) Diamond and graphite are both allotropes of carbon.

The conversion of graphite into diamond is represented as follows



Use this data below to calculate values for ΔH and ΔS for the reaction. Use these values to explain why this reaction is **not** feasible under standard pressure at any temperature.

	C (graphite)	C (diamond)
ΔH (kJ mol ⁻¹)	0	+1.9
ΔS (J K ⁻¹ mol ⁻¹)	+5.7	+2.4

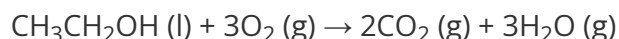
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(3 marks)

3 (a) Ethanol is used in large quantities in the production of alcoholic beverages and as a fuel.

The combustion of ethanol is represented by the equation:



The standard entropy, S^\ominus , of $\text{O}_2 \text{ (g)}$ is $205 \text{ J K}^{-1} \text{ mol}^{-1}$.

Using the data given and Section 13 in the Data Booklet, determine the entropy change, ΔS^\ominus , for the combustion of ethanol at 298K.

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(3 marks)

(b) Using the enthalpy of combustion for ethanol from Section 14 in the Data Booklet and the ΔS^\ominus determined in part (a), calculate the standard free energy for the combustion of ethanol.

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(3 marks)

(c) Explain whether changing the temperature for the combustion of ethanol will alter the spontaneity of the reaction.

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(3 marks)

(d) Using Section 13 of the Data Booklet, explain the difference in the standard entropy values between methanol, CH_3OH and ethanol, $\text{CH}_3\text{CH}_2\text{OH}$.

(1 mark)

4 (a) Ammonia, NH_3 , is produced by the Haber process and is an important chemical in the manufacture of fertilisers and clearing products.

Ammonia gas can react with oxygen to produce nitrogen monoxide and steam, and is the first step in the Ostwald process which produces nitric acid.

- i) Write an equation for the reaction of ammonia with oxygen to produce nitrogen monoxide and steam.
- ii) Using the given values determine the entropy for change for this reaction at 298 K.

Substance	Entropy values ($\text{J K}^{-1} \text{mol}^{-1}$)
$\text{NH}_3 (\text{g})$	192.8
$\text{O}_2 (\text{g})$	205.2
$\text{H}_2\text{O} (\text{g})$	188.8
$\text{NO} (\text{g})$	210.8

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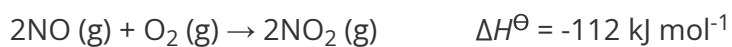
(5 marks)

(b) Explain why the standard entropy change for the reaction is positive.

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(1 mark)

- (c) The second step in the Ostwald process produces nitrogen dioxide as shown in the equation



The standard entropy for $\text{NO}_2 \text{ (g)}$ is $240.0 \text{ J K}^{-1} \text{ mol}^{-1}$

Determine the value for the free energy change for this reaction at 298 K using the information given

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(6 marks)

- (d) Explain whether changing the temperature for the production of nitrogen dioxide will alter the spontaneity of the reaction

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(2 marks)

- 5 (a) The boiling point of a liquid is the temperature at which its solid and liquid phases are in equilibrium as shown in the equation for the vaporisation of water.



Use Section 13 of the Data Booklet to determine values for the enthalpy change, ΔH^\ominus , and entropy change, ΔS^\ominus , for the reaction at 298 K.

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(4 marks)

- (b) Use your answer to part (a) to estimate a temperature, in K, that the reaction becomes feasible.

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(3 marks)

- (c) Explain how your answer to part (b) could be made more accurate.

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(1 mark)

- (d) Explain why the reaction is spontaneous above the boiling point of water.

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(2 marks)

Hard Questions

1 (a) The equations for two separate reversible reactions are as follows:

- Reaction **A** $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$
- Reaction **B** $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$

Use the information in the following table to determine the enthalpy change, ΔH_r , for reaction **A**.

	SO₂	O₂	SO₃	H₂
$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	-296.8	0	-395.7	0
$S^\ominus / \text{J K}^{-1} \text{mol}^{-1}$	248.2	205.1	256.8	131

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(2 marks)

Using the information from the table in part a), calculate the standard entropy change, ΔS , of reaction **A**.

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(b) (2 marks)

(c) Use your answer to parts a) and b) to determine the temperature at which reaction **A** becomes feasible.

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(2 marks)

(d) By using the data from part a) and section 13 of the data booklet, deduce if reaction **B** is feasible at a low temperature.

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(4 marks)

2 (a) Magnesium carbonate decomposes at a relatively high temperature.

	MgCO ₃	MgO
$\Delta H_f^\ominus / \text{kJ mol}^{-1}$	-1095.8	-601.7
$\Delta S^\ominus / \text{J mol}^{-1}$	65.7	26.9

i) Write an equation for the decomposition and use section 13 of the data booklet and the information given to determine the enthalpy change for the decomposition of magnesium carbonate.

[3]

ii) Use section 13 of the data booklet and the information given to determine the standard entropy change for the decomposition of magnesium carbonate.

[2]

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(5 marks)

(b) Using your answer to part a) to determine if the decomposition of magnesium carbonate is feasible at 280 °C.

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(1 mark)

(c) Using your answer to part b) to determine the temperature, in °C, at which the decomposition of magnesium carbonate becomes feasible.

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(3 marks)

3 (a) Used in the production of polymers, methanol is manufactured in large quantities.

i) State the equation for the combustion of methanol. [1]

ii) Comment on the entropy change of the reaction. [2]

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(3 marks)

(b) Use section 13 of the data booklet to determine the enthalpy change, ΔH_r , for the production of methanol from carbon monoxide and hydrogen.

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(3 marks)

Use your answer to part b) and section 13 of the data booklet to calculate the free energy change, ΔG , for the formation of methanol from carbon monoxide and hydrogen at 300 K.

$$\Delta S^\ominus \text{ hydrogen} = + 131 \text{ J K}^{-1} \text{ mol}^{-1}$$

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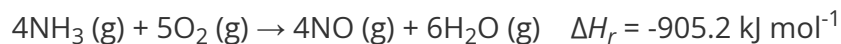
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(c) (3 marks)

(d) A student states that if the temperature is lowered for the formation of methanol, the reaction will become non-spontaneous. Comment on the student's statement.

(4 marks)

- 4 (a) The Ostwald process to produce nitric acid involves the oxidation of ammonia. The equation is shown below:



	NH₃ (g)	O₂ (g)
S[⊖] (J K⁻¹ mol⁻¹)	192.5	205

The free energy change, ΔG^\ominus , for the oxidation of ammonia at 298 K is -959 kJ mol^{-1} . Use section 13 of the data booklet, to calculate the entropy change for the oxidation of ammonia in $\text{J K}^{-1} \text{ mol}^{-1}$.

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(3 marks)

- (b) Use your answer to part a) and section 13 of the data booklet to determine the standard entropy of nitric oxide gas.

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(3 marks)

- (c) A 1.00 mol sample of NOCl was placed in reactor and heated to 227°C until the system reached equilibrium. The value for the equilibrium constant at this temperature is K_c is $4.5 \times 10^{-4} \text{ mol dm}^{-3}$.



Write an expression for K_c .

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(1 mark)

- (d)** Using section 1 and 2 in the data booklet determine the value for the free energy change, ΔG^\ominus , in kJ mol^{-1} for the reaction in part c).

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(2 marks)