

Topical Questions of

*Entropy and Gibbs free
energy*

Compiled By

Faizan Saleem

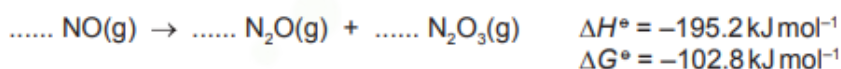
Topic: Entropy and Gibbs free Energy

Question No 1

- (a) State **one** natural and **one** man-made occurrence of oxides of nitrogen.

.....
..... [1]

- (b) Under conditions of high pressure and a catalyst, nitrogen monoxide, NO, forms two other oxides of nitrogen, dinitrogen monoxide, N₂O, and dinitrogen trioxide, N₂O₃.



- (i) Balance the equation above for the formation of N₂O and N₂O₃ from NO. [1]

- (ii) State how the oxidation number of nitrogen changes during this reaction.

NO → N₂O from to

NO → N₂O₃ from to

[1]

- (iii) Calculate the entropy change for the reaction at 298 K. Include the units in your answer.

$\Delta S^\circ =$

units = ...

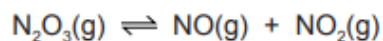
[2]

- (iv) State whether the sign of ΔS° calculated in (iii) agrees with that predicted from your balanced equation in (i). Explain your answer.

.....
..... [1]

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(c) At room temperature N_2O_3 dissociates.



(i) Write the expression for K_p for this equilibrium. Include the units in your answer.

$$K_p =$$

units = ...

[1]

A 1.00 dm^3 flask at 25°C is filled with pure $\text{N}_2\text{O}_3(\text{g})$ at an initial pressure of 0.60 atm .
At equilibrium, the partial pressure of $\text{NO}_2(\text{g})$ is 0.48 atm .

(ii) Calculate the partial pressures of $\text{NO}(\text{g})$ and $\text{N}_2\text{O}_3(\text{g})$ at equilibrium. Hence calculate the value of K_p at 25°C .

$$p(\text{NO}(\text{g})) = \dots\dots\dots$$

$$p(\text{N}_2\text{O}_3(\text{g})) = \dots\dots\dots$$

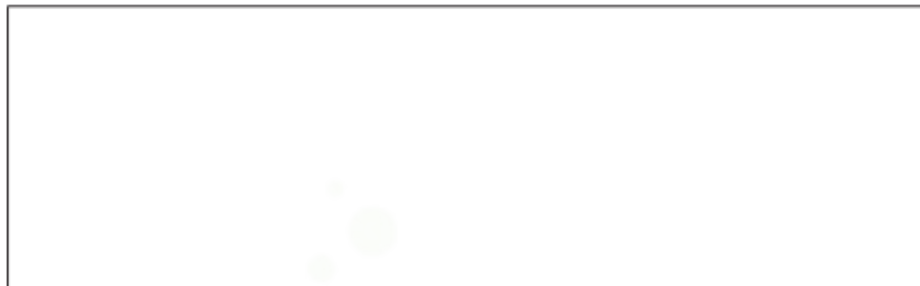
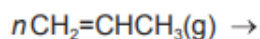
$$K_p = \dots\dots\dots$$

[2]

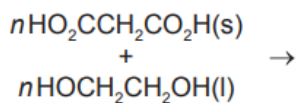
Question No 2

- (a) (i) Complete the equations to show the **two** types of polymerisation. Draw **one** repeat unit for each polymer. Include any other products.

- addition polymerisation



- condensation polymerisation



[3]

- (ii) Suggest the sign of the entropy changes, ΔS° , for each of the **two** types of polymerisation. Explain your answers.

- ΔS° for addition polymerisation

.....
.....
.....

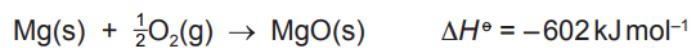
- ΔS° for condensation polymerisation

.....
.....
.....

[2]

Question No 3

(d) The equation for the formation of magnesium oxide from its elements is shown.



substance	$S^\ominus / \text{JK}^{-1} \text{mol}^{-1}$
Mg(s)	32.7
O ₂ (g)	205
MgO(s)	26.9

Use the equation and the data given in the table to calculate $\Delta G^\ominus$ for the reaction at 25 °C.

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$\Delta G^\ominus = \dots\dots\dots$ units $\dots\dots\dots$ [4]

Question No 4

(b) Entropy is a measure of the disorder of a system.

Describe and explain what happens to the entropy of a gas when the temperature is increased.

.....
.....
..... [2]

(c) The table shows four reactions.

- (i) For each reaction, predict the sign of the entropy change, ΔS° . If you predict no entropy change, write 'no change' in the table below. The first one has been done for you.

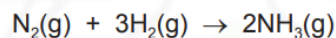
reaction	sign of ΔS°
$\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$	negative
$\text{Mg(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{MgO(s)}$	
$\text{CuSO}_4\text{(s)} + 5\text{H}_2\text{O(l)} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$	
$\text{NaHCO}_3\text{(s)} + \text{H}^+\text{(aq)} \rightarrow \text{Na}^+\text{(aq)} + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$	

[2]

- (ii) Explain why the entropy change for the first process is negative.

.....
 [1]

- (d) Calculate the standard entropy change, ΔS° , for this reaction.

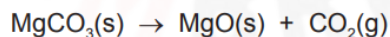


Standard entropies, S° , in $\text{JK}^{-1}\text{mol}^{-1}$ are given.

$\text{N}_2\text{(g)}$	$\text{H}_2\text{(g)}$	$\text{NH}_3\text{(g)}$
+192	+131	+193

ΔS° $\text{JK}^{-1}\text{mol}^{-1}$ [2]

- (e) Whether or not a chemical reaction is spontaneous (feasible) can be deduced by calculating the change in free energy, ΔG° , at a given temperature.



$$\begin{aligned}\Delta H^\circ &= +117 \text{ kJ mol}^{-1} \\ \Delta S^\circ &= +175 \text{ JK}^{-1}\text{mol}^{-1}\end{aligned}$$

- (i) Calculate the value of ΔG° at 298 K for the above reaction.

[2]

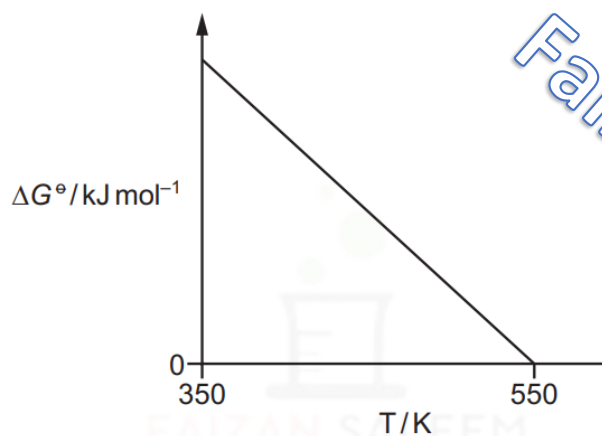
- (ii) Use your answer to (i) to explain whether or not this reaction is spontaneous at 298 K.

.....
 [1]

Question No 5

- (e) For a particular gas phase reaction the variation in standard Gibbs free energy change, ΔG° , with temperature is shown.

Assume standard enthalpy change, ΔH° , and standard entropy change, ΔS° , remain constant with temperature.



- (i) Write the equation that relates ΔG° to ΔH° and ΔS° .

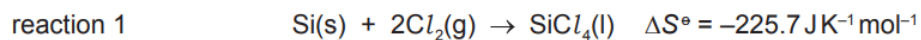
..... [1]

- (ii) Use this equation to explain why ΔG° becomes **less** positive as temperature increases in this reaction.

.....
..... [1]

Question No 6

- (c) (i) Silicon tetrachloride can be prepared according to reaction 1.



standard entropy of silicon, $S^\circ \text{ Si(s)}$	$18.7 \text{ J K}^{-1} \text{ mol}^{-1}$
standard entropy of silicon tetrachloride, $S^\circ \text{ SiCl}_4(\text{l})$	$239.0 \text{ J K}^{-1} \text{ mol}^{-1}$

Calculate the standard entropy of chlorine, $S^\circ \text{ Cl}_2(\text{g})$. Show all your working.

$S^\circ \text{ Cl}_2(\text{g}) = \dots\dots\dots \text{ J K}^{-1} \text{ mol}^{-1}$ [2]

- (ii) Explain why the entropy change for reaction 1 is negative.

.....
..... [1]

- (d) The standard enthalpy change of formation of silicon tetrachloride, $\Delta H_f^\circ \text{ SiCl}_4(\text{l})$, is -640 kJ mol^{-1} .

Reaction 1 is spontaneous at lower temperatures, but it is not spontaneous at very high temperatures.

Calculate the temperature above which reaction 1 is **not** spontaneous.

temperature = ... K [2]

Question No 7

The table lists the standard enthalpy changes of formation, $\Delta H_f^\ominus$, for some compounds and aqueous ions.

species	$\Delta H_f^\ominus / \text{kJ mol}^{-1}$
$\text{Ba}^{2+}(\text{aq})$	-538
$\text{OH}^-(\text{aq})$	-230
$\text{CO}_2(\text{g})$	-394
$\text{BaCO}_3(\text{s})$	-1216
$\text{H}_2\text{O}(\text{l})$	-286

(b) (i) Reaction 1 occurs when $\text{CO}_2(\text{g})$ is bubbled through an aqueous solution of $\text{Ba}(\text{OH})_2$.

Use the data in the table to calculate the standard enthalpy change for reaction 1, $\Delta H_{r1}^\ominus$.



$\Delta H_{r1}^\ominus = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

If $\text{CO}_2(\text{g})$ is bubbled through an aqueous solution of $\text{Ba}(\text{OH})_2$ for a long time, the precipitated $\text{BaCO}_3(\text{s})$ dissolves, as shown in reaction 2.



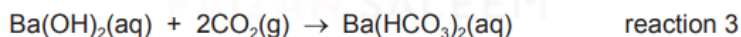
The standard enthalpy change for reaction 2, $\Delta H_{\text{r}2}^\ominus = -26 \text{ kJ mol}^{-1}$.

- (ii) Use this information and the data in the table to calculate the standard enthalpy change of formation of the $\text{HCO}_3^-(\text{aq})$ ion.

$$\Delta H_{\text{f}}^\ominus \text{HCO}_3^-(\text{aq}) = \dots\dots\dots \text{ kJ mol}^{-1} \quad [2]$$

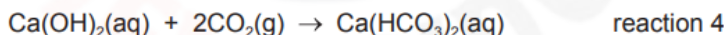
- (iii) The overall process is shown by reaction 3.

Use your answer to (ii), and the data given in the table, to calculate the standard enthalpy change for reaction 3, $\Delta H_{\text{r}3}^\ominus$.



$$\Delta H_{\text{r}3}^\ominus = \dots\dots\dots \text{ kJ mol}^{-1} \quad [1]$$

- (iv) How would the value of $\Delta H_{\text{r}3}^\ominus$ compare with the value of $\Delta H_{\text{r}4}^\ominus$ for the similar reaction with $\text{Ca}(\text{OH})_2(\text{aq})$ as shown in reaction 4?
Explain your answer.



.....

 [2]

- (c) The standard entropy change for reaction 1 is $\Delta S_{\text{r}1}^\ominus$.

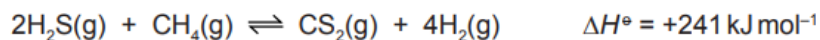
Suggest, with a reason, how the standard entropy change for reaction 3 might compare with $\Delta S_{\text{r}1}^\ominus$.

.....

 [2]

Question No 8

- (c) (i) Predict the sign of ΔS° for this reaction. Explain your answer.



.....
..... [1]

The free energy change, ΔG° , for this reaction at 1000 K is +51 kJ mol⁻¹.

- (ii) Calculate the value of ΔS° for this reaction, stating its units.

$\Delta S^\circ =$ units [2]

- (d) How would the value of ΔG° , and hence the spontaneity (feasibility) of this reaction change as the temperature increases? Explain your answer.

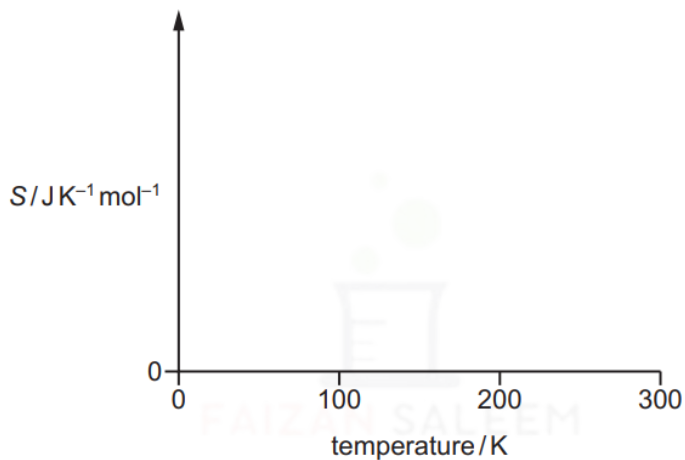
.....
.....
..... [2]

Question No 9

Entropy is a measure of the disorder of a system.

(a) Assume the entropy, S , for H_2O is zero at 0 K.

Sketch a graph on the axes to show how the entropy changes for H_2O between 0 K and 300 K.



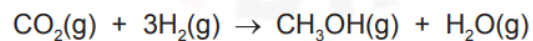
[2]

(b) Place **one tick (✓)** in **each row** of the table to show the sign of the entropy changes, ΔS .

	ΔS is negative	ΔS is positive
solid dissolving in water		
water boiling to steam		

[1]

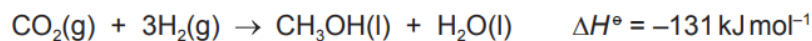
(c) The equation for a reaction that produces methanol is shown.



Use relevant bond energies from the *Data Booklet* to calculate the enthalpy change, ΔH , for this gas phase reaction.

$\Delta H = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

(d) At 298 K, both products of this reaction are liquid.



Standard entropies are shown in the table.

substance	$\text{CO}_2(\text{g})$	$\text{H}_2(\text{g})$	$\text{CH}_3\text{OH}(\text{l})$	$\text{H}_2\text{O}(\text{l})$
$S^\circ/\text{JK}^{-1}\text{mol}^{-1}$	+214	+131	+127	+70

(i) Calculate the standard entropy change, ΔS° , for this reaction.

$$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1}\text{mol}^{-1} \quad [2]$$

(ii) Calculate the standard Gibbs free energy change, ΔG° , for this reaction at 298 K.

$$\Delta G^\circ = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

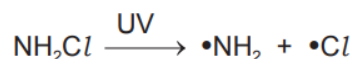
(iii) Predict the effect of increasing the temperature on the feasibility of this reaction.

.....
..... [1]

Question No 10

- (a) Chloramine, NH_2Cl , can be used in the treatment of drinking water to kill bacteria. Excess chloramine in water is destroyed using UV light. The mechanism for this involves free radicals.

The initiation step in this process is shown.



- (i) What is meant by the term *free radical*?

.....
..... [1]

The equation for a possible propagation step in the process is shown.



- (ii) Suggest an equation for a possible termination step in this process.

..... [1]

- (b) (i) Draw the 'dot-and-cross' diagram of NH_2Cl . Show outer electrons only.

[1]

- (ii) State the hybridisation of the nitrogen atom and suggest the H-N-Cl bond angle in the NH_2Cl molecule.

hybridisation of N

H-N-Cl bond angle

[1]

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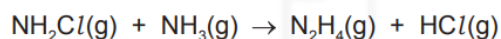
- (c) Some values for standard enthalpy changes of formation, ΔH_f° , and standard entropies, S° , are given in the table.

	$\Delta H_f^\circ/\text{kJ mol}^{-1}$	$S^\circ/\text{JK}^{-1}\text{mol}^{-1}$
$\text{NH}_2\text{Cl(g)}$	+80.1	+241
$\text{NH}_3\text{(g)}$	-45.9	+198
$\text{N}_2\text{H}_4\text{(g)}$	+95.4	+237
HCl(g)	-92.3	+187

- (i) Define the meaning of the term *entropy*.

.....
 [1]

Hydrazine, N_2H_4 , can be produced from chloramine and ammonia as shown.



- (ii) Calculate the standard entropy change, ΔS° , for this reaction.

$$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1}\text{mol}^{-1} \quad [1]$$

- (iii) Calculate the standard enthalpy change, ΔH° , for this reaction.

$$\Delta H^\circ = \dots\dots\dots \text{kJ mol}^{-1} \quad [1]$$

- (iv) Calculate the standard Gibbs free energy change, ΔG° , for this reaction at 298 K.

$$\Delta G^\circ = \dots\dots\dots \text{kJ mol}^{-1} \quad [2]$$

- (v) Explain, with reference to ΔG° , why this reaction becomes **less** feasible at higher temperatures.

.....

 [1]

Question No 11

The spontaneity (feasibility) of a chemical reaction depends on the standard Gibbs free energy change, ΔG° . This is related to the standard enthalpy and entropy changes by the equation shown.

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

- (a) State and explain whether the following processes will lead to an increase or decrease in entropy.

- (i) the reaction of magnesium with hydrochloric acid

entropy change

explanation

[1]

- (ii) solid potassium chloride dissolving in water

entropy change

explanation

[1]

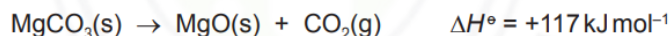
- (iii) steam condensing to water

entropy change

explanation

[1]

- (b) Magnesium carbonate can be decomposed.



Standard entropies are shown in the table.

substance	$\text{MgCO}_3(\text{s})$	$\text{MgO}(\text{s})$	$\text{CO}_2(\text{g})$
$S^\circ / \text{J mol}^{-1} \text{K}^{-1}$	+65.7	+26.9	+214

- (i) Calculate ΔG° for this reaction at 298K.
Include a relevant sign and give your answer to **three** significant figures.

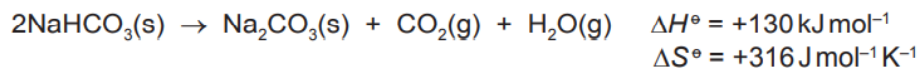
$$\Delta G^\circ = \dots\dots\dots \text{kJ mol}^{-1} \quad [3]$$

- (ii) Explain, with reference to ΔG° , why this reaction becomes more feasible at higher temperatures.

.....

..... [1]

(c) On heating, sodium hydrogencarbonate decomposes into sodium carbonate as shown.



Calculate the **minimum** temperature at which this reaction becomes spontaneous (feasible).
Show your working.

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temperature = ... K [2]

(d) The solubility of Group 2 sulfates decreases down the Group.

Explain this trend.

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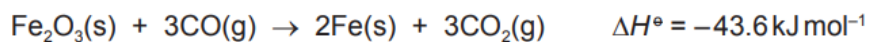
.....

.....

..... [2]

Question No 12

(c) Iron(III) oxide can be reduced to iron metal using carbon monoxide at a temperature of 1000 °C.



Some relevant standard entropies are given in the table.

substance	$\text{Fe}_2\text{O}_3(\text{s})$	$\text{CO}(\text{g})$	$\text{Fe}(\text{s})$	$\text{CO}_2(\text{g})$
$S^\circ / \text{JK}^{-1} \text{mol}^{-1}$	+90	+198	+27	+214

(i) What is meant by the term *entropy*?

.....
..... [1]

(ii) Calculate the standard entropy change, ΔS° , for this reaction.

$\Delta S^\circ = \dots\dots\dots \text{JK}^{-1} \text{mol}^{-1}$ [2]

(iii) Calculate the standard Gibbs free energy change, ΔG° , for this reaction at 25 °C.

$\Delta G^\circ = \dots\dots\dots \text{kJ mol}^{-1}$ [2]

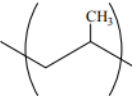
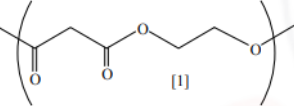
(iv) Suggest why a temperature of 1000 °C is usually used for this reaction, even though the reaction is spontaneous (feasible) at 25 °C. Explain your answer.

.....
.....
..... [1]

Question No 1

1(a)	natural: lightning, bacterial decomposition, volcanic emissions man-made: exhaust fumes, power stations, jet / car/ vehicle engines	1
1(b)(i)	$4\text{NO} \longrightarrow \text{N}_2\text{O} + \text{N}_2\text{O}_3$	1
1(b)(ii)	+2 to +1 AND +2 to +3	1
1(b)(iii)	$\Delta S = (\Delta H - \Delta G) / T$ $= (-195.2 + 102.8) / 298$ $= -0.310 \text{ kJ mol}^{-1} \text{ K}^{-1}$ M1 numerical answer M2 units	2
1(b)(iv)	yes as there is a decrease in no. of moles of gas OR yes as moles of (gaseous) reactants is greater than moles of (gaseous) products	1
1(c)(i)	$K_p = p(\text{NO})p(\text{NO}_2) / p(\text{N}_2\text{O}_3)$ AND units: atm OR Pa	1
1(c)(ii)	M1 $p(\text{NO}) = p(\text{NO}_2) = 0.48 \text{ atm}$ $p(\text{N}_2\text{O}_3)_{\text{eqm}} = p(\text{N}_2\text{O}_3)_0 - 0.48 = 0.12 \text{ atm}$ M2 $K_p = 0.48^2 / 0.12 = 1.92 \text{ (atm)}$	2

Question No 2

Question	Answer	Marks
7(a)(i)	 [1]  [1] + H ₂ O [1]	3
7(a)(ii)	- for addition polymerisation: ΔS will be negative, as many gas molecules are combining to form one (large) molecule - for condensation polymerisation: ΔS likely to be positive, (as each pair of monomer molecules join to chain, two molecules of) <u>water</u> forms	2

Question No 3

(d)	M1 correct use of $\Delta G = \Delta H - T\Delta S$	1
	M2 $\Delta S = 26.9 - (32.7 + 102.5) = -108.3 \text{ J K}^{-1} \text{ mol}^{-1}$ OR $-0.1083 \text{ kJ K}^{-1} \text{ mol}^{-1}$	1
	M3 $\Delta G = -602 - (298 \times (-0.1083)) = -570$	1
	M4 units: kJ mol^{-1}	1

Question No 4

(b)	(higher temperature means that) particles have more energy; entropy (of the gas/system) increases because of an increase in the amount of disorder/randomness;	2										
(c) (i)	<table><tr><th>reaction</th><th>sign of ΔS°</th></tr><tr><td>$\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$</td><td>negative</td></tr><tr><td>$\text{Mg(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{MgO(s)}$</td><td>negative</td></tr><tr><td>$\text{CuSO}_4\text{(s)} + 5\text{H}_2\text{O(l)} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$</td><td>negative</td></tr><tr><td>$\text{NaHCO}_3\text{(s)} + \text{H}^+\text{(aq)} \rightarrow \text{Na}^+\text{(aq)} + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$</td><td>positive</td></tr></table>	reaction	sign of ΔS°	$\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$	negative	$\text{Mg(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{MgO(s)}$	negative	$\text{CuSO}_4\text{(s)} + 5\text{H}_2\text{O(l)} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$	negative	$\text{NaHCO}_3\text{(s)} + \text{H}^+\text{(aq)} \rightarrow \text{Na}^+\text{(aq)} + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$	positive	2
reaction	sign of ΔS°											
$\text{CO(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)}$	negative											
$\text{Mg(s)} + \frac{1}{2}\text{O}_2\text{(g)} \rightarrow \text{MgO(s)}$	negative											
$\text{CuSO}_4\text{(s)} + 5\text{H}_2\text{O(l)} \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O(s)}$	negative											
$\text{NaHCO}_3\text{(s)} + \text{H}^+\text{(aq)} \rightarrow \text{Na}^+\text{(aq)} + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$	positive											
(ii)	there is a reduction in the overall number of <u>gaseous</u> molecules	1										
(d)	$\Delta S_f^\circ = 386 - (192 + (3 \times 131))$ $= -199 \text{ (JK}^{-1} \text{mol}^{-1}\text{)}$	2										
(e) (i)	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ $= 117 - ((298 \times 175) / 1000)$ $= (+) 64.85 \text{ (kJ mol}^{-1}\text{)}$	2										
(ii)	<u>ΔG° is positive</u> and so the reaction is <u>not spontaneous</u> (at 298 K)	1										

Question No 5

(e)(i)	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ [1]	1
(e)(ii)	$T\Delta S$ is more positive OR $-T\Delta S$ becomes more negative [1]	1

Question No 6

(c)(i)	$-225.7 = 239.0 - (18.7 + 2x)$	1
	$x = +223$	1
(c)(ii)	decrease in number of moles of gas /more moles of gas on left / reactants (ora)	1
1(d)	use of $\Delta G = \Delta H - T\Delta S$ with $\Delta G = 0$ / $\Delta G > 0$ or $T = \Delta H / \Delta S$ or $T = (640\,000 / 225.7)$	1
	2836 / 2840 (2835.6)	1

Question No 7

(b)(i)	$\Delta H_{r1} - (538 + 2 \times 230 + 394) = -(1216 + 286)$	1
	$\Delta H_{r1} - 1392 = -1502$	
	$\Delta H_{r1} = \mathbf{-110}$	1
(b)(ii)	let $\Delta H_f(\text{HCO}_3^-(\text{aq})) = y$	1
	$2y - 538 = -1216 - 394 - 286 - 26$	
	$y = \mathbf{-692}$	1
(b)(iii)	$\Delta H_{r3} - 538 - 2(230 + 394) = -538 - 2(692)$	1
	$\Delta H_{r3} = \mathbf{-136}$	
(b)(iv)	ΔH_{r3} will be identical to ΔH_{r4} , / unchanged	1
	as the reaction is the same, or:	1
	$2\text{OH}^-(\text{aq}) + 2\text{CO}_2(\text{g}) \longrightarrow 2\text{HCO}_3^-(\text{aq})$ or	
	metal ions stay in solution/metal ions are unchanged / are spectators	

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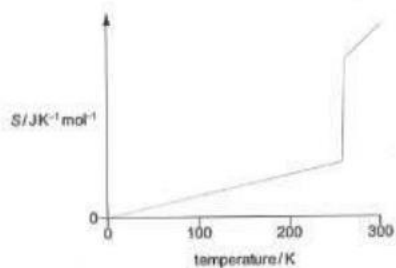
Question	Answer	Marks
1(c)	more gaseous moles are being consumed (in reaction 3) or more CO₂ moles are being consumed (in reaction 3)	1
	ΔS is therefore expected to be more negative/less positive for reaction 3.	1

Question No 8

(c) (i)	ΔS° will be positive, because more gas moles on the RHS/products	[1]
(ii)	$\Delta S^\circ = (\Delta H^\circ - T\Delta G^\circ) / T = (241 - 51) / 1000 = 0.19$ OR 190 kJ mol ⁻¹ K ⁻¹ OR J mol ⁻¹ K ⁻¹	[1] [1]
(d)	ΔG° will become less positive/more negative as T increases, ...because ΔS° is positive (or $-T\Delta S^\circ$ is more negative) ...therefore the reaction becomes more feasible/spontaneous as T increases	[2]

Question No 9

l(a)



2

M1 continuous increase in S from 0–300 K (excluding m.p.) [1]

M2 steep vertical increase in S ONLY at the m.p. **AND** continuous increase in S after m.p. [1]

l(b)

[1] for each correct tick

1

	negative ΔS°	positive ΔS°
solid dissolving in water		✓
water boiling to steam		✓

l(c)

$\Delta H^\circ = (2 \times \text{C=O}) + (3 \times \text{H-H}) - (3 \times \text{C-H}) - (\text{C-O}) - (3 \times \text{O-H})$
 $\Delta H^\circ = (2 \times 805) + (3 \times 436) - (3 \times 410) - (1 \times 360) - (3 \times 460)$ [1]
 $\Delta H^\circ = 1610 + 1308 - 1230 - 360 - 1380 = -52 \text{ (kJ mol}^{-1}\text{)}$ [1] ecf correct answer scores [2]

2

d)(i)

$\Delta S^\circ = 127 + 70 - (214 + 3 \times 131)$ [1]
 $= -410 \text{ (J K}^{-1} \text{ mol}^{-1}\text{)}$ [1] ecf correct answer scores [2]

2

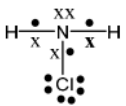
d)(ii)

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ [1]
 $\Delta G^\circ = -131 - (298 \times -0.41) = -8.8(2) \text{ (kJ mol}^{-1}\text{)}$ [1] correct answer scores [2]

2

Question	Answer	Marks
d)(iii)	(as temperature increases) feasibility decreases	1

Question No 10

(a)(i)	species with an unpaired electron	1
(a)(ii)	$\text{NH}_2 + \text{Cl} \rightarrow \text{NH}_2\text{Cl}$	1
(b)(i)		1
(b)(ii)	sp^3 AND $100-107^\circ$	1
(c)(i)	(entropy) is a measure of the disorder/randomness of a system	1
(c)(ii)	$\Delta S^\circ = 237 + 187 - (241 + 198) = -15.0 \text{ (J K}^{-1} \text{ mol}^{-1}\text{)}$	1
(c)(iii)	$\Delta H^\circ = 95.4 - 92.3 - (80.1 - 45.9) = -31.1 \text{ (kJ mol}^{-1}\text{)}$	1
(c)(iv)	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ [1] $\Delta G^\circ = -31.1 - (298 \times -0.015) = -26.6 \text{ (kJ mol}^{-1}\text{)}$ [1]	2
(c)(v)	(at higher temperatures) $T\Delta S^\circ$ becomes more negative so ΔG° becomes more positive OR (at high temperatures) $-T\Delta S^\circ$ is becomes more positive so ΔG° becomes more positive	1

Question No 11

(a)(i)	(entropy) increases/is positive and H_2 /gas is formed	1	1
(a)(ii)	(entropy) increases/is positive and (KCl (aq)) solution has (free) moving/mobile ions/aqueous ions	1	1
(a)(iii)	(entropy) decreases/is negative and decrease in gas	1	1
(b)(i)	$\Delta S^\circ = 26.9 + 214 - 65.7 = (+) 175.2 \text{ (J K}^{-1} \text{ mol}^{-1}\text{)}$ $\Delta G^\circ = 117 - (298 \times 175.2/1000)$ OR $\Delta G^\circ = 117\,000 - (298 \times 175.2)$ $\Delta G^\circ = +64.8 \text{ (kJ mol}^{-1}\text{)}$	1 1 1	3
(b)(ii)	$T\Delta S$ is more positive than ΔH / $T\Delta S$ increases/ $-T\Delta S$ more negative and ΔG is negative/decrease/less positive	1	1
3(c)	use of $\Delta G = 0$ or $\frac{T\Delta S}{\Delta H} = 1$ $T = 130 / (316/1000) = 410/411/412/411.4 \text{ (K)}$	1 1	2
(d)	hydration enthalpy and lattice energy both more endothermic/more positive/less exothermic/less negative (down the group) ΔH_{hyd} decreases more/faster and ΔH_{sol} becomes (more) endothermic/(more) positive/less exothermic/less negative	1 1	2

Question No 12

c)(i)	(entropy is a measure /degree of the) disorder of a system/substance	1	1
c)(ii)	$\Delta S^\circ = (2 \times 27) + (3 \times 214) - (90) - (3 \times 198)$ OR $696 - 684$ $\Delta S^\circ = (+) 12 \text{ (J K}^{-1} \text{ mol}^{-1}\text{)}$	1 1	2
c)(iii)	$\Delta G^\circ = -43.6 - (298 \times 12 / 1000)$ $\Delta G^\circ = -47.2 \text{ (kJ mol}^{-1}\text{)}$	1 1	2
c)(iv)	high E_a and to speed up the rate	1	1

