Enthalpy, Entropy & Free Energy
A Level only
Model Answers 1

Level	A Level
Subject	Chemistry
Exam Board	OCR
Module	Physical Chemistry & Transition Elements
Topic	Enthalpy, Entropy & Free Energy
Paper	A Level only
Booklet	Model Answers 1

Time allowed: **32 minutes**

Score: **/24**

Percentage: **/100**

Grade Boundaries:

A*	A	B	C	D	E
>85%	73%	60%	47%	34%	21%

Question 1

The table below shows standard entropies, $S^{\ominus}$.

Substance	CO(g)	$H_2(g)$	$CH_3OH(l)$
$S^{\ominus}$ / J mol^{-1} K^{-1}	197.6	130.6	239.7

What is the entropy change, $\Delta S^{\ominus}$, in J mol^{-1} K^{-1}, for the following reaction?

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(l)$$

A –219.1

B –88.5

C +88.5

D +219.1

[1]

Step	**Working out**
1. Use the equation $\Delta S^{\ominus}_{system} = \Delta S^{\ominus}_{products} - \Delta S^{\ominus}_{reactants}$ The $\Delta S^{\ominus}$ value for $H_2(g)$ is multiplied by 2 as per the reaction stoichiometry.	$\Delta S^{\ominus}$ = (239.7) – ((197.6 + (2 x 130.6) $\Delta S^{\ominus}$ = (131) - (458.8)
2. Calculate $\Delta S^{\ominus}$ of the reaction.	$\Delta S^{\ominus}$ = -219.9 J K^{-1} mol^{-1}

Question 2

This question looks at different aspects of entropy.

(a) Three processes are given below.

For each process, state and explain whether the change would be accompanied by an increase or decrease in entropy.

(i) The freezing of water.

increase or decrease **.** [1]

(entropy) decreases

AND

(solid/ice has) less disorder/ more order/ fewer ways of arranging energy/ less freedom/ less random molecules

(ii) The reaction of calcium carbonate with hydrochloric acid.

increase or decrease [1]

(entropy) increases

AND

(CO_2) **gas** is **formed**

Could be from equation with $CO_2(g)$

(iii) The formation of $O_3(g)$ from $O_2(g)$.

increase or decrease [1]

entropy decreases

AND

3 mol O_2 form 2 mol O_3

OR $3O_2 \rightarrow 2O_3$

OR 3 mol gas form 2 mol gas

Increase / decrease and the reason must be stated in each process.

(b) The enthalpy and entropy changes of a reaction both have a negative sign.

Discuss how the feasibility of this reaction will change as the temperature increases. [2]

Feasibility AND ΔG

Reaction becomes/is less feasible/not feasible

AND

ΔG increases

OR ΔG becomes/is less negative/more positive

OR $\Delta G > 0$ **OR** $\Delta H - T\Delta S > 0$

OR ΔH – TΔS becomes/is less negative/more positive

OR $\Delta H > T\Delta S$

OR $T\Delta S$ becomes/is more negative than ΔH

Effect on $T\Delta S$

$T\Delta S$ becomes more negative **OR** $T\Delta S$ decreases

OR $-T\Delta S$ becomes more positive **OR** $-T\Delta S$ increases

OR magnitude of $T\Delta S$ increases

OR $|\ T\Delta S\ |$ increases

Since both enthalpy and entropy changes are negative, then the reaction is spontaneous at low temperatures. Considering the equation $\Delta G = \Delta H - T\Delta S$, increasing temperature increases the value of $T\Delta S$, and since ΔS is negative, then the effect is to increase the value for ΔG , making the reaction less feasible.

(c) The metal tungsten is obtained on a large scale from its main ore, wolframite.
Wolframite contains tungsten(VI) oxide, WO_3.

Tungsten is extracted from wolframite by reduction with hydrogen:

$$WO_3(s) + 3H_2(g) \longrightarrow W(s) + 3H_2O(g) \qquad \Delta H = +115\,kJ\,mol^{-1}$$

Standard entropies are given in the table below.

Substance	$WO_3(s)$	$H_2(g)$	W(s)	$H_2O(g)$
$S^{\ominus}$ **J K^{-1} mol^{-1}**	76	131	33	189

(i) Calculate the free energy change, ΔG, in kJ mol^{-1}, for this reaction at 25 °C.

Show your working. [2]

ΔS = (33 + 3 x 189) – (76 + 3 x 131)

ΔS = (+)131 (J K^{-1} mol^{-1})

Use the equation $\Delta S^{\ominus}_{system} = \Delta S^{\ominus}_{products} - \Delta S^{\ominus}_{reactants}$ to calculate the total entropy of the system.

The $\Delta S^{\ominus}$ values for $H_2O(g)$ and $H_2(g)$ are multiplied by 3 in accordance with the reaction stoichiometry. Standard entropy values are per mole so these values must be multiplied by the coefficients from the balanced equation in order for the calculations to be correct.

ΔG = 115 – (298 x 0.131)

= (+) 75.962 **OR** 75.96 **OR** 76.0 **OR** 76 (kJ K^{-1} mol^{-1})

Use $\Delta G = \Delta H - T\Delta S$ to then calculate free energy. 131 is divided by 1000 to convert to kJ mol^{-1} as the value for ΔH given in the question is in kJ mol^{-1}.

(ii) Calculate the minimum temperature, in K, at which this reaction becomes feasible.

Show your working. [2]

(Minimum temperature when) $\Delta G = 0$ **OR** $\Delta H - T\Delta S = 0$

OR

(For feasibility) $\Delta G = 0$ **OR** $\Delta G < 0$ **OR** $\Delta H - T\Delta S < 0$

OR $T = \dfrac{\Delta H}{\Delta S}$

$T = \dfrac{115}{0.131} = 878$ K

The minimum temperature for the reaction to be feasible is when ΔG is zero or just below zero. To find this temperature, substitute a value of 0 into the equation for Gibbs free energy and solve for T.

Step	**Working out**
1.Substitute 0 for ΔG *and rearrange for* $\Delta H - T\Delta S$	$\Delta H - T\Delta S = 0$
2. Substitute the values for entropy and enthalpy and rearrange for T	$115 - T(0.131) = 0$ $-T(0.131) = -115$ $T(0.131) = 115$
3. Solve for T	So $T = \dfrac{115}{0.131} = 878$ K

[Total: 9 Marks]

Question 3

The equation for the reaction of CO_2 and H_2O to produce glucose, $C_6H_{12}O_6$, and O_2 is shown below.

$6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$ $\Delta H = +2879$ kJ mol^{-1}; $\Delta S = -256$ J K^{-1} mol^{-1}

Standard entropies are given in the table below.

Substance	$CO_2(g)$	$H_2O(l)$	$O_2(g)$
$S^{\ominus}$ / **J K^{-1} mol^{-1}**	214	70	205

(a) (i) Calculate the standard entropy of glucose. [2]

$-256 = (6 \times 205) + S(C_6H_{12}O_6) - (6 \times 214 + 6 \times 70)$

OR $S(C_6H_{12}O_6) = -256 - (6 \times 205) + (6 \times 214 + 6 \times 70)$

OR $-256 + 474 = 218$ (J K^{-1} mol^{-1})

Use to the equation $\Delta S^{\ominus}system = \Sigma S^{\ominus}products - \Sigma S^{\ominus}reactant$

$-256 = (1230 + X) - (1284 + 420)$

$-256 = 1230 + X - 1704$

$X = -256 - 1230 + 1704$

$X = 218$

(ii) Calculate ΔG, in kJ mol^{-1}, at 25 °C.

Show all your working. [2]

$\Delta G = +2879 - 298 \times -0.256$

$= (+)2955$ (kJ mol^{-1})

Use $\Delta G = \Delta H - T\Delta S$ to calculate the Gibbs free energy.

(iii) Explain why this reaction is **not** feasible at **any** temperature. [1]

ΔH is positive **OR** $\Delta H > 0$

AND

ΔS is negative **OR** $T\Delta S$ is negative **OR** $\Delta S < 0$ **OR** $T\Delta S < 0$

AND ΔG will always be positive **OR** $\Delta G > 0$

From the equation for Gibbs energy, ΔH is +ve, $T\Delta S$ is -ve, therefore $(\Delta H - T\Delta S)$ will always produce a positive number, so the reaction will never be feasible.

(b) Although the reaction between CO_2 and H_2O to form $C_6H_{12}O_6$ and O_2 appears not to be feasible, plants are able to make the reaction take place spontaneously by photosynthesis.

Each year, 3.4×10^{18} kJ of solar energy is taken in by all the plants on the Earth to make photosynthesis take place.

Calculate the mass of carbon dioxide that is removed each year from the atmosphere by photosynthesis on Earth.

[2]

amount of CO_2 removed

= $3.4 \times 10^{18} \times 6/2879$ **OR** 7.09×10^{15} (mol)

mass of $CO_2 = 44.0 \times 7.09 \times 10^{15} =$ **3.12×10^{17}g**

Step	**Working out**
1.Write down the moles of CO_2 produced (from the balanced equation given in question)	$6CO_2 + 6H_2O$ + light = $C_6H_{12}O_6 + 6O_2$ So 6 moles CO_2
2.From the question, 2879 kJ removes 6 moles of CO_2, so use this to calculate the total moles of CO_2 removed using the total amount of solar energy	Total moles $CO_2 = \frac{3.4\ x10^{18}}{2879} \times 6 = 7.09 \times 10^{15}$
3.Convert moles to mass	Mass $CO_2 = 7.09 \times 10^{15} \times 44$ $= 3.12 \times 10^{17}$g

[Total 7 Marks]

Question 4

A student is asked to calculate ΔG at 25 ºC for the combustion of butan-1-ol. The teacher provides two pieces of information.

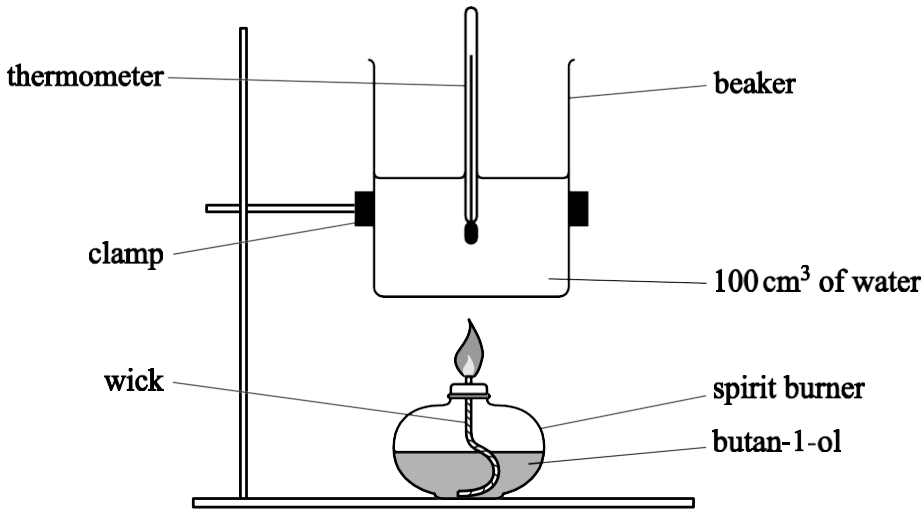
- The equation for the combustion of butan-1-ol.

$$CH_3(CH_2)_3OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l) \qquad \textbf{Equation 2}$$

- Standard entropies of butan-1-ol, oxygen, carbon dioxide and water.

	$CH_3(CH_2)_3OH(l)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	228	205	214	70

The student carries out an experiment using the apparatus below and obtains the following results. The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.



Mass of burner and butan-1-ol before burning / g	98.997
Mass of burner and butan-1-ol after burning / g	98.738
Initial temperature / ºC	18.5
Maximum temperature reached / ºC	39.0

Use the information on the previous page to calculate ΔG, in kJ mol^{-1}, for the combustion of But an-1-ol according to **Equation 2** at 25 °C. **[7]**

Step	Working out
1. Use Q = m x c x ΔT to calculate Q.	Q = 100 x 4.18 x 20.5 Q = -8569 J = -8.569 kJ (negative as the reaction is exothermic).
2. Use the result from step 1 to calculate the molar enthalpy value.	Mass of fuel used = 0.259 g n(Butanol) = $\frac{0.259}{74}$ = 3.5 x 10^{-3} moles ΔH = $\frac{-8.569}{3.5\ x\ 10^{-3}}$ = -2448 kJ mol^{-1}
3. Calculate ΔS using the equation: ΔS = ∑(products) - ∑(reactants).	ΔS = ∑(4 x 214) + (5 x 70) - ∑(228 - (6 x 205) ΔS = -252 JK^{-1} mol^{-1} OR -0.252 kJK^{-1} mol^{-1}
4. Use ΔG = ΔH - TΔS to calculate ΔG.	ΔG = -2448 - (298 x -0.252) **ΔG = -2373 (kJ mol^{-1})**