ENTROPY

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Question 1

2 Silicon is the second most abundant element by mass in the Earth's crust.

(a) In industry, silicon is extracted from SiO_2 by reaction with carbon at over 2000 °C.

reaction 1 $SiO_2(s) + 2C(s) \rightarrow Si(l) + 2CO(g)$

(i) Explain why the entropy change, ΔS, of reaction 1 is positive.

....

.... [1]

(ii) Reaction 1 is highly endothermic.

Suggest the effect of an increase in temperature on the feasibility of this reaction.
Explain your answer.

....

....

.... [2]

(b) Silicon is purified by first heating it in a stream of $HCl(g)$ to form $SiHCl_3$. The $SiHCl_3$ formed is then distilled to remove other impurities.

reaction 2 $Si(s) + 3HCl(g) \rightarrow SiHCl_3(g) + H_2(g)$

(i) Table 2.1 shows some standard entropy data.

Table 2.1

compound	standard entropy, $S^{\ominus}$ / $J K^{-1} mol^{-1}$
$Si(s)$	19
$HCl(g)$	187
$SiHCl_3(g)$	314
$H_2(g)$	131

Use the data in Table 2.1 to calculate $\Delta S^{\ominus}$ for reaction 2.

$\Delta S^{\ominus}$ = ... $J K^{-1} mol^{-1}$ [2]

(ii) Reaction 3 is the reverse of reaction 2 and is used to obtain pure silicon.

reaction 3 $SiHCl_3(g) + H_2(g) \rightarrow Si(s) + 3HCl(g)$ $\Delta H = +219.3\,kJ\,mol^{-1}$

Use this information and your answer to **(b)(i)** to calculate the temperature, in K, at which reaction 3 becomes feasible.
Show your working.

[If you were unable to answer **(b)(i)**, you should use $\Delta S^{\circ} = -150\,J\,K^{-1}\,mol^{-1}$ for reaction 2. This is not the correct answer to **(b)(i)**.]

temperature = ... K [2]

9701/42/F/M/22 (Question 2)

Question	Answer	Marks
2(a)(i)	1 mol liquid and 2 mol gas formed from 3 mol solid **OR** two solid compounds converted to a liquid and a gas	**1**
2(a)(ii)	M1: (as T increases) $T\Delta S$ becomes greater (than ΔH) **OR** (as T increases) $T\Delta S$ becomes more positive M2: (as T increases) feasibility will increase as ΔG becomes more negative	**2**
2(b)(i)	M1: = 314 + 131 – (19 + 3 × 187) use of values and correct stoichiometry M2: **= –135** ($J\,K^{-1}\,mol^{-1}$)	**2**
2(b)(ii)	M1: $\Delta G = 0 \therefore T = \Delta H / \Delta S$ = +219.3 × 10^3 ÷ –**(b)(i)** M2: = **1624**(.4) (K)	**2**

Question 2

(c) (i) Explain what is meant by the term *entropy of a system*.

....

.... [1]

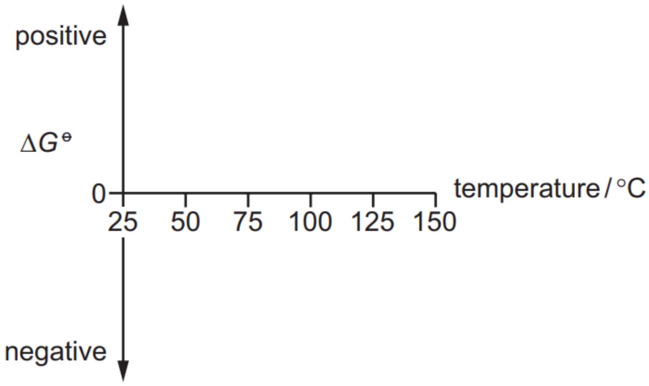
(ii) Place one tick (✓) in each row of the table to show the sign of each entropy change, ΔS.

process	ΔS is negative	ΔS is zero	ΔS is positive
NaCl dissolving in water			
water solidifying to ice			

[1]

(iii) The evaporation of one mole of water has a standard Gibbs free energy change, $\Delta G^{\ominus}$, of +8.6 kJ at 25 °C.

Sketch a graph on the axes to show how $\Delta G^{\ominus}$ changes for this process between 25 °C and 150 °C at 101 kPa.



[2]

(d) The reaction between A and B is feasible at low temperatures but is **not** feasible at high temperatures.

$$A + B \rightleftharpoons C + D$$

Deduce the signs of ΔH and ΔS for this reaction and explain why the feasibility changes with temperature.

sign of ΔH = ... sign of ΔS = ...

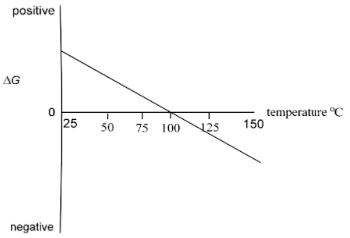
....

....

....

.... [2]

9701/42/M/J/21 (Question 4)

<table>
<tr><th>Question</th><th>Answer</th><th>Marks</th></tr>
<tr><td>4(c)(i)</td><td>measure / degree of (dis)order / randomness (of a system)
OR the number of possible arrangements of the particles and their energy (in a given system)</td><td>1</td></tr>
<tr><td>4(c)(ii)</td><td>
<table>
<tr><th></th><th>ΔS is negative</th><th>ΔS is zero</th><th>ΔS is positive</th></tr>
<tr><td>solid dissolving in water</td><td></td><td></td><td>✓</td></tr>
<tr><td>water solidifying to ice</td><td>✓</td><td></td><td></td></tr>
</table>
</td><td>1</td></tr>
<tr><td>4(c)(iii)</td><td>

two correct for 1 mark, three correct for two marks:
• starting at +8.6 kJ / in positive region close to the y-axis
• line passes through x-axis around 100°C
• negative gradient straight / curve line through the x-axis (no clear positive inflexions)</td><td>2</td></tr>
<tr><td>4(d)</td><td>M1: ΔH negative / – , ΔS negative / –

M2: as temperature increase, ΔG becomes (more) positive / less negative ora
OR at low(er) T, (ΔH more negative than TΔS) so ΔG is negative
OR at high(er) T, (ΔH less negative than TΔS) so ΔG is positive</td><td>2</td></tr>
</table>

Question 3

(f) NH_4IO_3 is an unstable compound that readily decomposes when warmed. The decomposition reaction is shown.

$$NH_4IO_3(s) \rightarrow \tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) + \tfrac{1}{2}I_2(g) + 2H_2O(l) \qquad \Delta H = -154.6\,kJ\,mol^{-1}$$

(i) Use the data in the table to calculate the entropy change of reaction, ΔS, of the decomposition of $NH_4IO_3(s)$.

compound	S / J K^{-1} mol^{-1}
$NH_4IO_3(s)$	42
$N_2(g)$	192
$O_2(g)$	205
$I_2(g)$	261
$H_2O(l)$	70

ΔS = ... J K^{-1} mol^{-1} [2]

(ii) This reaction is feasible at all temperatures.

Explain why, using the data in **(f)** and your answer to **(f)(i)**.

....

....

.... [1]

9701/42/F/M/21 (Question 3)

3(f)(i)	**M1:** ΔS = ½(192) + ½(205) + ½(261) + 2(70) – 42 **M2:** (+)427 (J K^{-1} mol^{-1}) ecf	**2**
3(f)(ii)	ΔG (always) negative because • $\Delta H < 0$ / negative **OR** exothermic **AND** • $\Delta S > 0$ / positive **OR** $-T\Delta S < 0$ for all T	**1**

Question 4

(b) Silver chloride, AgC*l*, is sparingly soluble in water. The equation for the enthalpy change of solution is shown.

$$AgCl(s) \rightarrow Ag^+(aq) + Cl^-(aq) \qquad \Delta H^{\ominus}_{sol} = +65.5\,kJ\,mol^{-1}$$

Standard entropies are shown in the table.

species	AgC*l*(s)	Ag^+(aq)	Cl^-(aq)
$S^{\ominus}$ / J K^{-1} mol^{-1}	+96.2	+72.7	+56.5

(i) Calculate the standard entropy change of solution, $\Delta S^{\ominus}$.

$\Delta S^{\ominus}$ = ... J K^{-1} mol^{-1} [1]

(ii) Explain, with the aid of a calculation, why AgC*l* is insoluble in water at 25 °C.

You should use data from this question and your answer to **(b)(i)**.

....

.... [3]

9701/43/M/J/20 (Question 7)

7(b)(i)	ΔS^{o} = 72.7 + 56.5 -96.2 = **+33.0** J K^{-1} mol^{-1}	**1**
7(b)(ii)	**M1** $\Delta G = \Delta H^{o} - T\Delta S^{o}$ **M2** ΔG = (65.5)- (298x0.033) = **+55.7** kJ mol^{-1} min 3sf **M3** ΔG = positive so not feasible/spontaneous	**3**

Question 5

3 **(a)** Explain what is meant by the term *entropy of a system*.

....

.... [1]

(b) State and explain whether the entropy change of each of the following processes is positive or negative. Do not consider the entropy change of the surroundings.

- liquid water at 80 °C is cooled to 60 °C

The entropy change is .... because ....

....

- solid calcium chloride is added to water and the mixture is stirred

The entropy change is .... because ....

....

- the change corresponding to the lattice energy of calcium chloride, ΔH_{latt} $CaCl_2$(s), takes place

The entropy change is .... because ....

....
[3]

(c) The reaction $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$ is not spontaneous at room temperature.

(i) Give the full name for the term $\Delta G^{\ominus}$.

.... [1]

(ii) Describe how the temperature at which the reaction becomes spontaneous can be calculated. Include an equation in your answer.

equation ....

....

....

....
[2]

[Total: 7]

9701/43/O/N/19 (Question 3)

3(a)	a **measure / degree** of **disorder / randomness** of a system	1
3(b)	**M1:** negative – molecules have less energy in the system **M2:** positive – solid being converted into an aqueous solution **M3:** negative – gaseous ions being converted into a solid	3
3(c)(i)	(standard) Gibbs free energy change	1
3(c)(ii)	**M1:** $(\Delta)G = \Delta H - T\Delta S$ **M2:** description of calculating the minimum value of T for which ΔG is zero / becomes negative **OR** $T = \Delta H / \Delta S$ [1]	2

Question 6

(f) Copper can form complexes with the ligands ammonia and *en*, $H_2NCH_2CH_2NH_2$, as shown.

$[Cu(H_2O)_6]^{2+}(aq) + en(aq) \rightleftharpoons [Cu(H_2O)_4(en)]^{2+}(aq) + 2H_2O(l)$ $K_{stab} = 3.98 \times 10^{10}$ equilibrium **4**

$[Cu(H_2O)_6]^{2+}(aq) + 2NH_3(aq) \rightleftharpoons [Cu(H_2O)_4(NH_3)_2]^{2+}(aq) + 2H_2O(l)$ $K_{stab} = 5.01 \times 10^{7}$ equilibrium **5**

(ii) The standard entropy change, $\Delta S^{\ominus}$, for equilibrium **4** is $+23\,J\,K^{-1}\,mol^{-1}$ and for equilibrium **5** is $-8.4\,J\,K^{-1}\,mol^{-1}$.

Suggest an explanation for this difference by reference to both equilibria.

....

....

.... [1]

9701/42/M/J/19 (Question 3)

3(f)(ii)	equilibrium **4** has a (net) increase in moles of product / 2 moles goes to 3 moles whereas equilibrium **5** has same number of moles of reactants and products / 3 moles vs 3 moles [1]	1

Question 7

(b) Under conditions of high pressure and a catalyst, nitrogen monoxide, NO, forms two other oxides of nitrogen, dinitrogen monoxide, N_2O, and dinitrogen trioxide, N_2O_3.

$$\text{......}\ NO(g) \rightarrow \text{......}\ N_2O(g) + \text{......}\ N_2O_3(g) \qquad \Delta H^\ominus = -195.2\,kJ\,mol^{-1}$$
$$\Delta G^\ominus = -102.8\,kJ\,mol^{-1}$$

(i) Balance the equation above for the formation of N_2O and N_2O_3 from NO. [1]

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

$NO \rightarrow N_2O$ from .... to ....

$NO \rightarrow N_2O_3$ from .... to ....

[1]

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

$\Delta S^\ominus$ = ...

units = ...

[2]

(iv) State whether the sign of $\Delta S^\ominus$ calculated in **(iii)** agrees with that predicted from your balanced equation in **(i)**. Explain your answer.

....

.... [1]

9701/42/F/M/19 (Question 1)

1(b)(i)	$4NO \longrightarrow N_2O + N_2O_3$	**1**
1(b)(ii)	+2 to +1 **AND** +2 to +3	**1**
1(b)(iii)	$\Delta S = (\Delta H - \Delta G) / T$ $= (-195.2+102.8) / 298$ $= -0.310\ kJ\ mol^{-1}\ K^{-1}$ **M1** numerical answer **M2** units	**2**

Question 8

4 The table shows some standard entropy data.

substance	standard entropy, $S^\ominus$ / $J\,K^{-1}\,mol^{-1}$
$PbO_2(s)$	77
$PbO(s)$	69
$O_2(g)$	205

Lead(IV) oxide, PbO_2, decomposes to lead(II) oxide, PbO, and oxygen when heated.

$$2PbO_2(s) \rightarrow 2PbO(s) + O_2(g) \qquad \Delta H^\ominus = +118\,kJ\,mol^{-1}$$

(a) Use the data to calculate the value of $\Delta S^\ominus$ for this reaction.

$\Delta S^\ominus$ = ... $J\,K^{-1}\,mol^{-1}$ [2]

(b) Use the value of $\Delta H^\ominus$ and your answer to **(a)** to calculate the temperature at which this reaction becomes feasible.

T = ... K [3]

(c) Solid lead(II) oxide can be made by heating lead metal in air.

Predict the **sign** of the standard entropy change of this reaction. Explain your answer.

....

.... [1]

[Total: 6]

9701/42/O/N/19

4(a)	**M1:** correct use of stoichiometry **M2:** answer + 189	**2**
4(b)	**M1:** States or uses correct form of Gibbs equation $\Delta G = \Delta H - T\Delta S$ **M2:** appreciates / includes ΔG = 0 at temperature required **M3:** uses 1000 correctly and answer +624(.339) Award 3 marks for correct answer	**3**
4(c)	negative **and** decrease in number / amount of gas molecules	**1**

Question 9

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

reaction 1 $Si(s) + 2Cl_2(g) \rightarrow SiCl_4(l)$ $\Delta S^{\ominus} = -225.7\,J\,K^{-1}\,mol^{-1}$

standard entropy of silicon, $S^{\ominus}$ Si(s)	$18.7\,J\,K^{-1}\,mol^{-1}$
standard entropy of silicon tetrachloride, $S^{\ominus}$ $SiCl_4$(l)	$239.0\,J\,K^{-1}\,mol^{-1}$

Calculate the standard entropy of chlorine, $S^{\ominus}$ Cl_2(g). Show all your working.

$S^{\ominus}$ Cl_2(g) = ... $J\,K^{-1}\,mol^{-1}$ [2]

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

....

.... [1]

(d) The standard enthalpy change of formation of silicon tetrachloride, $\Delta H_f^{\ominus}\ SiCl_4(l)$, 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]

9701/42/M/J/18

Question 10

7 (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

$n\,CH_2{=}CHCH_3(g) \rightarrow$

- condensation polymerisation

$n\,HO_2CCH_2CO_2H(s) + n\,HOCH_2CH_2OH(l) \rightarrow$

(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]

(b) An amide bond forms when a carboxylic acid reacts with an amine.

(i) Complete the equation by writing the products in the box.

$R\text{–}CO_2H + H_2N\text{–}R' \rightarrow$ [............ + ...]

[1]

(ii) Use your answer to **(i)** to work out the bonds that are broken and the bonds that are formed during the reaction between a carboxylic acid and an amine.

- bonds that are broken

....

- bonds that are formed

....

[2]

(iii) Use bond energy values from the *Data Booklet* to calculate the enthalpy change, $\Delta H^\ominus$, when **one** mole of amide bonds is formed in the reaction in **(i)**.

$\Delta H^\ominus$ = ... kJ [2]

(c) Amide bonds can also be formed by reacting acyl chlorides with amines.

The enthalpy change for this process, $\Delta H^\ominus$, is $-6.00\,kJ\,mol^{-1}$.

Calculate the minimum entropy change, $\Delta S^\ominus$, for this reaction to be spontaneous (feasible) at 298 K.

$\Delta S^\ominus$ = ... $J\,K^{-1}\,mol^{-1}$ [2]

9701/42/F/M/18

Question 11

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

$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta H^\ominus = -602\,kJ\,mol^{-1}$$

substance	$S^\ominus / J\,K^{-1}\,mol^{-1}$
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.

$\Delta G^\ominus$ = ... units ....

[4]

9701/42/F/M/17

Question 12

(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, $\Delta S^{\ominus}$. If you predict no entropy change, write 'no change' in the table below. The first one has been done for you.

reaction	sign of $\Delta S^{\ominus}$
$CO(g) + O_2(g) \rightarrow CO_2(g)$	negative
$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$	
$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4.5H_5O(s)$	
$NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + CO_2(g) + H_2O(l)$	

[2]

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

....

.... [1]

(d) Calculate the standard entropy change, $\Delta S^{\ominus}$, for this reaction.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Standard entropies, $S^{\ominus}$, in $J\,K^{-1}\,mol^{-1}$ are given.

$N_2(g)$	$H_2(g)$	$NH_3(g)$
+192	+131	+193

$\Delta S^{\ominus}$ .... $J\,K^{-1}\,mol^{-1}$ [2]

(e) Whether or not a chemical reaction is spontaneous (feasible) can be deduced by calculating the change in free energy, $\Delta G^{\ominus}$, at a given temperature.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$$

$\Delta H^{\ominus} = +117\,kJ\,mol^{-1}$
$\Delta S^{\ominus} = +175\,J\,K^{-1}\,mol^{-1}$

(i) Calculate the value of $\Delta G^{\ominus}$ 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 13

(c) **(i)** Predict the sign of $\Delta S^{\ominus}$ for this reaction. Explain your answer.

$$2H_2S(g) + CH_4(g) \rightleftharpoons CS_2(g) + 4H_2(g) \qquad \Delta H^{\ominus} = +241\,kJ\,mol^{-1}$$

....

.... [1]

The free energy change, $\Delta G^{\ominus}$, for this reaction at 1000 K is +51 kJ mol^{-1}.

(ii) Calculate the value of $\Delta S^{\ominus}$ for this reaction, stating its units.

$\Delta S^{\ominus}$ = ... units .... [2]

(d) How would the value of $\Delta G^{\ominus}$, and hence the spontaneity (feasibility) of this reaction change as the temperature increases? Explain your answer.

....

....

.... [2]

s/16/qp41

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

Question 14

5 Cadmium ions form complexes with primary amines and with 1,2-diaminoethane.

$$Cd^{2+}(aq) + 4CH_3NH_2(aq) \rightleftharpoons [Cd(CH_3NH_2)_4]^{2+}(aq) \qquad K_{stab} = 3.6 \times 10^{6} \quad \text{equilibrium I}$$

$$Cd^{2+}(aq) + 2H_2NCH_2CH_2NH_2(aq) \rightleftharpoons [Cd(H_2NCH_2CH_2NH_2)_2]^{2+}(aq) \qquad K_{stab} = 4.2 \times 10^{10} \quad \text{equilibrium II}$$

(b) Values for ΔH° and ΔG° for equilibria **I** and **II**, and the value of ΔS° for equilibrium **I**, are given in the table below. All values are at a temperature of 298 K.

equilibrium	ΔH° / kJ mol^{-1}	ΔG° / kJ mol^{-1}	ΔS° / J K^{-1} mol^{-1}
I	–57.3	–37.4	–66.8
II	–56.5	–60.7	to be calculated

(i) Suggest a reason why the ΔH° values for the two equilibria are very similar.

....

.... [1]

(ii) Calculate ΔS° for equilibrium **II**.

ΔS° = ... J K^{-1} mol^{-1} [1]

(iii) Suggest a reason for the difference between the ΔS° you have calculated for equilibrium **II** and that for equilibrium **I** given in the table.

....

....

.... [1]

(iv) Which of the two complexes is the more stable? Give a reason for your answer.

....

.... [1]

S/16/qp42

(b) (i)	(each complex is formed by) making (4 ×)N-Cd bonds and breaking (6 ×) O-Cd bonds ***or*** same types of / similar bonds forming / breaking ***or*** same number of bonds forming / breaking	1
(ii)	$\Delta S = (\Delta H - \Delta G) / T$ = (60.7 – 56.5) × 1000 / 298 = (+)**14** / (+)**14.1**	1
(iii)	fewer moles (of solutes) are forming (one mole of) the complex (so less loss of disorder) ***or*** one *en* displaces two H_2O whereas one CH_3NH_2 only displaces one H_2O	1
(iv)	The $[Cd(H_2NCH_2CH_2NH_2)_2]^{2+}$ / equilibrium 2 complex (is more stable) because: *either* K_{stab} is greater or ΔG° is more negative.	1

Question 15

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

$$2H_2S(g) + CH_4(g) \rightleftharpoons CS_2(g) + 4H_2(g) \qquad \Delta H^\circ = +241\,kJ\,mol^{-1}$$

....

.... [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.

ΔS° = ... 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]

S/16/qp43

(c) (i)	ΔS° will be positive, because more gas moles on the RHS / products	[1]
(ii)	$\Delta S^\circ = (\Delta H^\circ - \Delta G^\circ) / T = (241 - 51)/1000 = 0.19$ 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 16

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

$$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta H^{\ominus} = -602\,kJ\,mol^{-1}$$

substance	$S^{\ominus}$ / $J\,K^{-1}\,mol^{-1}$
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.

$\Delta G^{\ominus}$ = ... units ....

[4]

m/17/qp42

M1 correct use of ΔG = ΔH – TΔS	1
M2 $\Delta S = 26.9 - (32.7 + 102.5) = -108.3\,J\,K^{-1}\,mol^{-1}$ **OR** $-0.1083\,kJ\,K^{-1}\,mol^{-1}$	1
M3 $\Delta G = -602 - (298 \times (-0.1083)) = -570$	1
M4 units: $kJ\,mol^{-1}$	1

Question 17

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

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g) \qquad \Delta H^\circ = -43.6\,kJ\,mol^{-1}$$

Some relevant standard entropies are given in the table.

substance	$Fe_2O_3(s)$	$CO(g)$	$Fe(s)$	$CO_2(g)$
$S^\circ / J\,K^{-1}\,mol^{-1}$	+90	+198	+27	+214

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

....

.... [1]

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

ΔS° = ... $J\,K^{-1}\,mol^{-1}$ [2]

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

ΔG° = ... $kJ\,mol^{-1}$ [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]

3(c)(i)	(entropy is a measure / degree of the) disorder of a **system / substance**	1	1
3(c)(ii)	$\Delta S^{\circ} = (2\times27) + (3\times214) - (90) - (3\times198)$ **OR** 696 – 684 $\Delta S^{\circ} = (+) 12$ (J K^{-1} mol^{-1})	1 1	2
3(c)(iii)	$\Delta G^{\circ} = -43.6 - (298 \times 12 / 1000)$ $\Delta G^{\circ} = -47.2$ (kJ mol^{-1})	1 1	2
3(c)(iv)	high E_a **and** to speed up the rate	1	1

Question 18

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]

Question 19

(b) Magnesium carbonate can be decomposed.

$$MgCO_3(s) \rightarrow MgO(s) + CO_2(g) \qquad \Delta H^\ominus = +117\,kJ\,mol^{-1}$$

Standard entropies are shown in the table.

substance	$MgCO_3(s)$	$MgO(s)$	$CO_2(g)$
$S^\ominus / J\,mol^{-1}\,K^{-1}$	+65.7	+26.9	+214

(i) Calculate $\Delta G^\ominus$ for this reaction at 298 K.
Include a relevant sign and give your answer to **three** significant figures.

$\Delta G^\ominus$ = ... $kJ\,mol^{-1}$ [3]

(ii) Explain, with reference to $\Delta G^\ominus$, why this reaction becomes more feasible at higher temperatures.

....

.... [1]

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

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + CO_2(g) + H_2O(g) \qquad \Delta H^\ominus = +130\,kJ\,mol^{-1}$$
$$\Delta S^\ominus = +316\,J\,mol^{-1}\,K^{-1}$$

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

temperature = ... K [2]

w/16/qp43

3(a)(i)	(entropy) increases / is positive **and** H_2 / gas is formed	1	**1**
3(a)(ii)	(entropy) increases / is positive **and** (KC*l* (aq)) solution has (free) moving / mobile ions / aqueous ions	1	**1**
3(a)(iii)	(entropy) decreases / is negative **and** decrease in **gas**	1	**1**
3(b)(i)	$\Delta S^{\circ} = 26.9 + 214 - 65.7 = (+)\,175.2\ (\text{J K}^{-1}\ \text{mol}^{-1})$ $\Delta G^{\circ} = 117 - (298 \times 175.2 / 1000)$ **OR** $\Delta G^{\circ} = 117\,000 - (298 \times 175.2)$ $\Delta G^{\circ} = +64.8\ (\text{kJ mol}^{-1})$	1 1 1	**3**
3(b)(ii)	TΔS is more positive than ΔH / TΔS increases / –TΔS 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** (K)	1 1	**2**

2016 Specimen Paper Question on Entropy

7 **(a)** The table lists the equations for five processes.

For each process, predict the sign of ΔS.

process	**sign of ΔS**
$NaBr(s) + (aq) \rightarrow NaBr(aq)$	
$H_2O(l) \rightarrow H_2O(g)$	
$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$	
$CoCl_2(s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O(s)$	

[2]

(b) Ethanol can be combusted as shown in the equation.

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

Standard entropies are shown in the table.

substance	$CH_3CH_2OH(l)$	$O_2(g)$	$CO_2(g)$	$H_2O(l)$
$S^{\ominus}$, J K^{-1} mol^{-1}	161	205	214	70

Calculate the standard entropy change, $\Delta S^{\ominus}$, for this reaction.

$\Delta S^{\ominus}$ = ... J K^{-1} mol^{-1} [2]

(c) The combustion of ethanol is an exothermic reaction.

This reaction occurs spontaneously at low temperatures but does **not** occur at very high temperatures. Explain why.

....

....

....

.... [2]

(d) The decomposition of calcium carbonate is an endothermic reaction.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

$\Delta H = +178$ kJ mol^{-1} and $\Delta S = +159$ J K^{-1} mol^{-1}

Calculate the **minimum** temperature at which this reaction becomes feasible.
Show all your working.

[3]

[Total: 9]

Marking Scheme

(a)

process	sign of ΔS
$NaBr(s) + (aq) \rightarrow NaBr(aq)$	+
$H_2O(l) \rightarrow H_2O(g)$	+
$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$	–
$CoCl_2(s) + 6H_2O(l) \rightarrow CoCl_2.6H_2O(s)$	–

2 correct, (1) mark
4 correct, (2) marks [2]

(b) $\Delta S^{\ominus} = (214 \times 2) + (70 \times 3) - (161 \times 1) - (205 \times 3)$
$= -138$ J K^{-1} mol^{-1} [2]

(c) As temperature increases $T\Delta S$ is more negative or $-T\Delta S$ increases (1)
At high temperature $T\Delta S$ is more negative than ΔH (so ΔG is positive) (1) [2]

(d) the reaction is feasible, ΔG is negative so
T > *$\Delta H / T\Delta S$ or use of $T = \Delta H / T\Delta S$* (1)

T = 178000/159 (1)
T = 1119.5 K units required or T>1120 K (1) [3]

IB QUESTIONS on ENTROPY

1. Which change leads to an increase in entropy?

A. $CO_2(g) \rightarrow CO_2(s)$

B. $SF_6(g) \rightarrow SF_6(l)$

C. $H_2O(l) \rightarrow H_2O(s)$

D. $NaCl(s) \rightarrow NaCl(aq)$

(Total 1 mark)

2. The reaction between but-1-ene and water vapour produces butan-1-ol.

$$C_4H_8(g) + H_2O(g) \rightarrow C_4H_9OH(l)$$

The standard entropy values ($S^{\ominus}$) for but-1-ene, water vapour and butan-1-ol are 310, 189 and 228 $J\ K^{-1}\ mol^{-1}$ respectively. What is the standard entropy change for this reaction in $J\ K^{-1}\ mol^{-1}$?

A. –271

B. +271

C. –107

D. +107

(Total 1 mark)

3. What are the signs of $\Delta H^{\ominus}$ and $\Delta S^{\ominus}$ for a reaction that is non-spontaneous at low temperature but spontaneous at high temperature?

	$\Delta H^{\ominus}$	$\Delta S^{\ominus}$
A.	–	–
B.	+	–
C.	–	+
D.	+	+

(Total 1 mark)

4. Consider the following reaction:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

(i) Suggest why this reaction is important for humanity.

(1)

(ii) Using the average bond enthalpy values in Table 10 of the Data Booklet, calculate the standard enthalpy change for this reaction.

(4)

(iii) The absolute entropy values, S, at 238 K for $N_2(g)$, $H_2(g)$ and $NH_3(g)$ are 192, 131 and 193 J K^{-1} mol^{-1} respectively. Calculate $\Delta S^\ominus$ for the reaction and explain the sign of $\Delta S^\ominus$.

(2)

(iv) Calculate $\Delta G^\ominus$ for the reaction at 238 K. State and explain whether the reaction is spontaneous.

(3)

(v) If ammonia was produced as a liquid and not as a gas, state and explain the effect this would have on the value of $\Delta H^\ominus$ for the reaction.

(2)
(Total 12 marks)

5. Which reaction causes a decrease in the entropy of the system?

A. $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

B. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

C. $2C(s) + O_2(g) \rightarrow 2CO(g)$

D. $2SO_3(g) \rightarrow 2SO_2(g) + O_2(g)$

(Total 1 mark)

6. Consider the following reaction.

$$2CH_3OH(g) + H_2(g) \rightarrow C_2H_6(g) + 2H_2O(g)$$ □□

(a) The standard enthalpy change of formation for $CH_3OH(g)$ at 298 K is –201 kJ mol^{-1} and for $H_2O(g)$ is –242 kJ mol^{-1}. Using information from Table 11 of the Data Booklet, determine the enthalpy change for this reaction. (Check the Marking Scheme for Missing Data)

....

....

....

(2)

(b) The standard entropy for $CH_3OH(g)$ at 298 K is 238 J K^{-1} mol^{-1}, for $H_2(g)$ is 131 J K^{-1} mol^{-1} and for $H_2O(g)$ is 189 J K^{-1} mol^{-1}. Using information from Table 11 (Check Marking Scheme) of the Data Booklet, determine the entropy change for this reaction.

....

....

....

(2)

(c) Calculate the standard change in free energy, at 298 K, for the reaction and deduce whether the reaction is spontaneous or non-spontaneous.

....

....

....

....

....

....

(3)
(Total 7 marks)

7. What is the standard entropy change, $\Delta S^{\ominus}$, for the following reaction?

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

	CO(g)	**O_2(g)**	**CO_2(g)**
$S^{\ominus}$/J K^{-1} mol^{-1}	198	205	214

A. –189

B. –173

C. +173

D. +189

(Total 1 mark)

8. A reaction has a standard enthalpy change, $\Delta H^{\ominus}$, of +10.00 kJ mol^{-1} at 298 K. The standard entropy change, $\Delta S^{\ominus}$, for the same reaction is +10.00 J K^{-1} mol^{-1}. What is the value of $\Delta G^{\ominus}$ for the reaction in kJ mol^{-1}?

A. +9.75

B. +7.02

C. –240

D. –2970

(Total 1 mark)

9. Which reaction has the greatest increase in entropy?

A. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

B. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

C. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

D. $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$

(Total 1 mark)

10. Which reaction has the largest increase in entropy?

A. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

B. $Al(OH)_3(s) + NaOH(aq) \rightarrow Al(OH)_4^-(aq) + Na^+(aq)$

C. $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

D. $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

(Total 1 mark)

11. When hydrogen peroxide decomposes, the temperature of the reaction mixture increases.

$$2H_2O_2(aq) \rightarrow O_2(g) + 2H_2O(l)$$

What are the signs of ΔH, ΔS and ΔG for this reaction?

	ΔH	ΔS	ΔG
A.	–	–	–
B.	–	+	–
C.	+	+	–
D.	–	+	+

(Total 1 mark)

12. Which reaction has the greatest increase in entropy?

A. $SO_2(g) + 2H_2S(g) \rightarrow 2H_2O(l) + 3S(s)$

B. $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$

C. $CaC_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$

D. $N_2(g) + O_2(g) \rightarrow 2NO(g)$

(Total 1 mark)

13. $\Delta G^\ominus$ calculations predict that a reaction is always spontaneous for which of the following combinations of $\Delta H^\ominus$ and $\Delta S^\ominus$?

A. $+\Delta H^\ominus$ and $+\Delta S^\ominus$

B. $+\Delta H^\ominus$ and $-\Delta S^\ominus$

C. $-\Delta H^\ominus$ and $-\Delta S^\ominus$

D. $-\Delta H^\ominus$ and $+\Delta S^\ominus$

(Total 1 mark)

14. Some words used in chemistry can have a specific meaning which is different to their meaning in everyday English.

State what the term *spontaneous* means when used in a chemistry context.

....

....

(Total 1 mark)

15. Propene can be hydrogenated in the presence of a nickel catalyst to form propane.
Use the data below to answer the questions that follow.

Compound	**Formula**	$\Delta H^{\ominus}_f$ **/ kJ mol^{-1}**	$S^{\ominus}$**/ J K^{-1} mol^{-1}**
hydrogen	$H_2(g)$	0	+ 131
propane	$C_3H_8(g)$	– 104	+ 270
propene	$C_3H_6(g)$	+ 20.4	+ 267

(i) Outline why the value for the standard enthalpy change of formation of hydrogen is zero.

....

....

(1)

(ii) Calculate the standard enthalpy change for the hydrogenation of propene.

....

....

....

....

(2)

(iii) Calculate the standard entropy change for the hydrogenation of propene.

....

....

....

....

(2)

(iv) Determine the value of $\Delta G^{\ominus}$ for the hydrogenation of propene at 298 K.

....

....

....

....

(2)

(v) At 298 K the hydrogenation of propene is a spontaneous process. Determine the temperature above which propane will spontaneously decompose into propene and hydrogen.

....

....

....

....

(2)

(Total 9 marks)

16. (Not in Syllabus) What is the standard free energy change, $\Delta G^{\ominus}$, in kJ, for the following reaction?

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$$

Compound	**$\Delta G_f^{\ominus}$ / kJ mol^{-1}**
$C_2H_5OH(l)$	–175
$CO_2(g)$	–394
$H_2O(g)$	–229
$O_2(g)$	0

A. –1650

B. –1300

C. – 448

D. +1300

(Total 1 mark)

17. Which reaction has the most negative change in entropy?

A. $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

B. $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$

C. $PbCl_2(s) \rightarrow Pb^{2+}(aq) + 2Cl^{-}(aq)$

D. $C(s) + O_2(g) \rightarrow CO_2(g)$

(Total 1 mark)

MARKING SCHEME

1. D

[1]

2. A

[1]

3. D

[1]

4. (i) fertilizers / increasing crop yields;
production of explosives for mining; 1 max

(ii) ΔH = (sum of energies of bonds broken) – (sum of energies of bonds formed);

Can be implied by working.

correct substitution of values and numbers of bonds broken;
correct substitution of values and numbers of bonds made;
(ΔH = (N≡N) + 3(H–H) – 6(N–H) = 944 + 3(436) – 6(388) =) –76.0 (kJ); 4

Allow ECF.
Do not penalize for sig. fig. or units.
*Award **[4]** for correct final answer.*

(iii) ($\Delta S^{\ominus}$[2×193] – [192 + 3×131]) = –199 (J K^{-1} mol^{-1}); 2

Allow ECF.
four <u>gaseous</u> molecules generating two <u>gaseous</u> molecules / fewer molecules of <u>gas</u>;

(iv) ($\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$ = –76.0 – 298(–0.199)) = –16.7 (kJ);
Spontaneous;
ΔG is negative; 3

Do not penalize for SF.

(v) heat released when gas → liquid;
$\Delta H^{\ominus}$ becomes more negative; 2

[12]

5. B

[1]

6. (a) $\Delta H^{\ominus}_{reaction} = \Sigma\Delta H_f^{\ominus}(products) - \Sigma\Delta H_f^{\ominus}(reactants)$
= [(1)(–85) + (2)(–242)] – [(2)(–201)];
= –167 (kJ/kJ mol^{-1});
*Award **[1]** for (+) 167.* 2

(b) $\Delta S^{\ominus}_{reaction} = \Sigma S^{\ominus}(products) - \Sigma S^{\ominus}(reactants)$

= [(1)(230) + (2)(189)] – [(2)(238) + (1)(131)];
= 1 (J K^{-1}/J K^{-1} mol^{-1}); 2

(c) $\Delta G^{\Theta}_{reaction} = (\Delta H^{\Theta} - T\Delta S^{\Theta}) = (-167) - (298)(0.001)$;
*Award **[1]** for correct substitution of values.*

= –167 kJ/–167000 J;
Units needed for mark in (c) only.
Accept –167 kJ mol^{-1} or –167000 J mol^{-1}.

spontaneous;

Award marks for final correct answers throughout in each of (a), (b) and (c). 3

[7]

7. B

[1]

8. B

[1]

9. A

[1]

10. C

[1]

11. B

[1]

12. C

[1]

13. D

[1]

14. the reaction gives out (Gibbs Free) energy that can do work;
ΔG for the reaction has a negative value;
a reaction that occurs without adding energy (beyond that required to overcome energy barrier); 1 max

[1]

15. (i) by definition $\Delta H_h^{\ominus}$ of elements (in their standard states) is zero / no reaction involved / *OWTTE*; 1

(ii) $\Delta H = -104 - (+20.4)$;
$= -124.4$ (kJ mol^{-1}); 2
*Award **[1 max]** for 124.4 (kJ mol^{-1}).*
*Award **[2]** for correct final answer.*

(iii) $\Delta S = 270 - (267 + 131)$;
$= -128$ (J K mol^{-1}); 2
*Award **[1 max]** for +128 (J K^{-1} mol^{-1}).*
*Award **[2]** for correct final answer.*

(iv) $\Delta G = \Delta H - T\Delta S = -124.4 - \frac{(-128 \times 298)}{1000}$;
$= -86.3$ kJ mol^{-1}; 2
Units needed for the mark.
*Award **[2]** for correct final answer.*
Allow ECF if only one error in first marking point.

(v) $\Delta G = \Delta H - T\Delta S = 0$ / $\Delta H = T\Delta S$;
$T = \frac{-124.4}{-128/1000} = 972$ K / 699 °C; 2
Only penalize incorrect units for T and inconsistent ΔS value once in (iv) and (v).

[9]

16. B

[1]

17. A

[1]

SHOEIFFAT QUESTIONS

Basic Questions

BQ 1 Predict the sign of the entropy change, ΔS°, for the following processes:

a. $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$

1 mole of a gas is producing 2 moles of gases ⇒ disorder increases.
ΔS° is positive.

b. $2H(g) \rightarrow H_2(g)$

2 moles of gases are producing 1 mole of gas ⇒ disorder decreases.
ΔS° is negative.

(c) Fluorine gas, an oxidizing agent, is bubbled into a solution of potassium bromide at 25°C. The balanced equation for the reaction occurring is: $F_2(g)+2Br^-(aq) \rightarrow 2F^-(aq)+Br_2(aq)$

Predict the sign of ΔS° for the reaction at 25°C. Justify your prediction.

The sign of ΔS° is negative.
One of the reactants, F_2, is a gas at 25°C, but there are no gaseous products. Gases have high entropies, so the entropy of the reactants is greater than the entropy of the products, making ΔS° negative.

BQ 2 Consider the following process: $Zn(s) + \frac{1}{2}O_2(g) \rightarrow ZnO(s)$

a. Calculate ΔS° at 298K, given the following S° values, in $J.K^{-1}\,mol^{-1}$:
ZnO: 44; Zn: 42; O_2: 205.

$$\Delta S° = S^\circ_{ZnO(s)} - S^\circ_{Zn(s)} - \frac{1}{2}S^\circ_{O_2(g)} = (44) - (42) - (\frac{205}{2}) = -100.5 J/mol\,K$$

b. Is the sign of ΔS° expected?
yes, because 1 mole of solid and 0.5 mole of gas are becoming 1 mole of solid.

BQ 4 Consider the following process: $H_2O(s) \rightarrow H_2O(l)$ ΔH° = 6.03 kJ
Given ΔS° = 22.1 J/K.mol, show that the melting of ice becomes spontaneous at 0°C.

Given ΔS° = 22.1 J/K.mol, ΔH° = 6.03 kJ, T = 0°C + 273 = 273K
RTF: rxn is spontaneous

$$\Delta G° = \Delta H° - T\,\Delta S° = (6.03) - (273)(22.1 \times 10^{-3}) = -3.3 \times 10^{-3} kJ$$

Since ΔG° is negative then the process is spontaneous.

BQ 5 Consider the following process: $C_6H_6(l) \rightarrow C_6H_6(g)$
Given for the above process ΔH°vap = 30.8 kJ/mol, and ΔS°vap = 87.2 J/K.mol.
Calculate the boiling point of benzene.

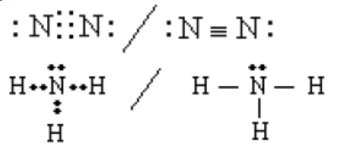
Given: ΔH°vap = 30.8 kJ/mol, ΔS°vap = 87.2 J/K.mol RTF: boiling point of benzene

A phase change at equilibrium is isothermal, ie occurs at same t, therefore ΔG° = 0kJ

$$\Delta G° = \Delta H°vap - T\Delta S°vap = (30.8) - T(87.2 \times 10^{-3}) = 0kJ \Rightarrow T = \frac{30.8}{87.2 \times 10^{-3}} = 353K \Rightarrow t = 80°C$$

BQ 7 Answer the following questions about nitrogen, hydrogen, and ammonia.

a. Draw the complete Lewis electron-dot diagrams for N_2 and NH_3.



(b) Calculate the standard free-energy change, ΔG, that occurs when 24.0 g of $H_2(g)$ react with excess $N_2(g)$ at 298 K according to the reaction represented below.

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$ $\Delta G^\circ_{298} = -34$ kJ mol^{-1}

Given: m of H_2 = 24.0 g, $\Delta G^\circ_{298} = -34$ kJ mol^{-1}

RTF: ΔG

$$n_{H_2} = \frac{24.0}{2} = 12.0\text{moles}$$

$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)$ $\Delta G^\circ_{298} = -34$ kJ mol^{-1}

3moles -34kJ

12.0moles ΔG_{298} $\Delta G_{298} = \frac{12.0\,(-34)}{3} = -136$ **kJ**

Multiple Choice

1. Which of the following is true of a reaction that is spontaneous at <u>higher</u> temperatures?

[-A-] ΔS° and ΔH° are both negative
[-B-] ΔS° and ΔH° are both positive
[-C-] ΔS° is negative and ΔH° is positive
[-D-] ΔS° is positive and ΔH° is negative
[-E-] ΔS° and ΔH° are both equal to zero

2. When solid NH_4SCN is mixed with solid $Ba(OH)_2$ in a closed container, the temperature drops and a gas is produced. Which of the following indicates the correct signs for ΔG, ΔH and ΔS for the process?

	ΔG	ΔH	ΔS
[-A-]	-	-	-
[-B-]	-	+	-
[-C-]	**-**	**+**	**+**
[-D-]	+	-	+
[-E-]	+	-	-

A gas is produced from solids ⇒ entropy increased ⇒ ΔS is positive
Temperature drops ⇒ reaction is endothermic ⇒ ΔH is positive

3. When a solid sample of $NaNO_3$ is added to a cup of water, the temperature of the resulting solution decreases. Which of the following must be true?

	ΔG	ΔH	ΔS
[-A-]	-	-	-
[-B-]	-	+	-
[-C-]	**-**	**+**	**+**
[-D-]	+	-	+
[-E-]	+	-	-

Sodium nitrate dissolves ⇒ reaction is spontaneous ⇒ ΔG negative
Temperature decreases ⇒ reaction is endothermic ⇒ enthalpy must be positive
Products are aqueous ⇒ entropy increased ⇒ΔS is positive

5. $X(s) \leftrightarrow X(l)$
Which of the following is true for any substance undergoing the process represented above at its normal melting point?

[-A-] $\Delta S < 0$
[-B-] $\Delta H = 0$
[-C-] $\Delta H = T\,\Delta G$
[-D-] $T\Delta S = 0$
[-E-] $\Delta H = T\Delta S$

6. For a certain reaction, the standard free energy is -70.0kJ at 100K and -40.0Kj at 200 K. For this reaction
[-A-] $\Delta H > 0$; $\Delta S < 0$
[-B-] $\Delta H < 0$; $\Delta S < 0$
[-C-] $\Delta H > 0$; $\Delta S > 0$
[-D-] $\Delta H < 0$; $\Delta S < 0$

Sample Questions

SQ1 Define a spontaneous process?
A process is said to be spontaneous if it occurs without external or outside interference.
SQ2 What are the driving forces for spontaneous processes?
minimizing energy and maximizing randomness
SQ Define entropy. **It is a thermodynamic function that measures randomness or disorder.**
SQ4 State the second law of thermodynamics.
The second law of thermodynamics states that in any spontaneous process there is always an increase in the entropy of the universe.
SQ5 When will water have higher entropy, at 25°C or at 80°C?
At 80°C, water will have greater entropy. Entropy increases with an increase in temperature.