1 Chemists use three energy terms, enthalpy, entropy and free energy, to help them make predictions about whether reactions may take place.

(a) The table below shows five processes. Each process has either an increase in entropy or a decrease in entropy.

For each process, tick (✓) the appropriate box.

process		**increase in entropy**	**decrease in entropy**
A	$C_2H_5OH(l) \rightarrow C_2H_5OH(g)$		
B	$C_2H_2(g) + 2H_2(g) \rightarrow C_2H_6(g)$		
C	$NH_4Cl(s) + aq \rightarrow NH_4Cl(aq)$		
D	$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$		
E	$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$		

[2]

(b) At 1 atm (101 kPa) pressure, ice melts into water at 0 °C.

Complete the table below using the symbols '+', '–' or '0' to show the sign of ΔH and ΔS for the melting of ice at 0 °C and 1 atm.

For each sign, explain your reasoning.

energy change	**sign +, – or 0**	**reasoning**
ΔH		
ΔS		

[2]

(c) Much of the hydrogen required by industry is produced by reacting natural gas with steam:

$$CH_4(g) + H_2O(g) \longrightarrow 3H_2(g) + CO(g)$$

Standard entropies are given in the table below.

substance	$CH_4(g)$	$H_2O(g)$	$H_2(g)$	$CO(g)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	186	189	131	198

(i) Calculate the standard entropy change, in J K^{-1} mol^{-1}, for this reaction of natural gas with steam.

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

(ii) State **two** large-scale uses for the hydrogen produced.

1. ....

2. .... **[1]**

(d) Ammonium chloride, NH_4Cl, can dissociate to form ammonia, NH_3, and hydrogen chloride, HCl.

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

At 298 K, ΔH = +176 kJ mol^{-1} and ΔG = +91.2 kJ mol^{-1}.

- Calculate ΔG for this reaction at 1000 K.
- Hence show whether this reaction takes place spontaneously at 1000 K.

Show **all** your working.

ΔG = ... kJ mol^{-1} **[4]**

[Total: 11]

2 Free energy changes can be used to predict the feasibility of processes.

(a) Write down the equation that links the free energy change with the enthalpy change and temperature.

.... **[1]**

(b) You are provided with equations for five processes.

For each process, predict the sign of ΔS.

process	**sign of ΔS**
$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$	
$NaCl(s) + (aq) \rightarrow NaCl(aq)$	
$H_2O(l) \rightarrow H_2O(s)$	
$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$	
$CuSO_4(s) + 5H_2O(l) \rightarrow CuSO_4{\bullet}5H_2O(s)$	

[2]

(c) Ammonia can be oxidised as shown in the equation below.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

Standard entropies are given in the table below.

substance	$NH_3(g)$	$O_2(g)$	$NO(g)$	$H_2O(g)$
$S^{\ominus}$ / J K^{-1} mol^{-1}	192	205	211	189

Calculate the standard entropy change, in J K^{-1} mol^{-1}, for this oxidation of ammonia.

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

(d) The exothermic reaction below occurs spontaneously at low temperatures but does **not** occur at very high temperatures.

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

Explain why.

....

....

....

....

....

.... **[2]**

(e) An ore of iron contains iron(III) oxide, Fe_2O_3.
Iron is extracted from this ore by heating with carbon.
The equation below shows one of the reactions which takes place.

$$Fe_2O_3(s) + 3C(s) \rightarrow 2Fe(s) + 3CO(g)$$

$\Delta S = +543\,J\,K^{-1}\,mol^{-1}$ and $\Delta H = +493\,kJ\,mol^{-1}$

Calculate the minimum temperature at which this reaction becomes feasible.

Show **all** your working.

minimum temperature = ... **[3]**

[Total: 10]

3 Entropy changes are an important factor in determining the feasibility of reactions.

(a) You are provided with equations for four processes.

A $2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$
B $H_2O(l) \longrightarrow H_2O(g)$
C $H_2(g) + ½O_2(g) \longrightarrow H_2O(l)$
D $2C(s) + O_2(g) \longrightarrow 2CO(g)$

For each process, explain why ΔS has the sign shown below.

A: sign of ΔS: negative

reason for sign: ....

....

B: sign of ΔS: positive

reason for sign: ....

....

C: sign of ΔS: negative

reason for sign: ....

....

D: sign of ΔS: positive

reason for sign: ....

.... **[4]**

(b) Calcium oxide, CaO, is used to make cement. Calcium oxide is manufactured by the thermal decomposition of calcium carbonate.

$$CaCO_3(s) \longrightarrow CaO(s) + CO_2(g) \qquad \Delta H = +178\,kJ\,mol^{-1}$$

Standard entropies of $CaCO_3(s)$, $CaO(s)$ and $CO_2(g)$ are given in the table below.

substance	$CaCO_3(s)$	$CaO(s)$	$CO_2(g)$
$\boldsymbol{S}$ **/ J K^{-1}mol^{-1}**	89	40	214

- Using the information in the table, show that the entropy change, ΔS, for the decomposition of calcium carbonate is $0.165\,kJ\,K^{-1}\,mol^{-1}$.
- Show that calcium carbonate is stable at room temperature (25 °C).
- Calculate the minimum temperature needed to decompose calcium carbonate.

Show all your working.

[7]

[Total: 11]