23.4 GIBBS FREE ENERGY CHANGE

Introduction

1. An equation for the direct reduction of ZnS by carbon is shown.
   $2\ ZnS(s) + C(s) \rightarrow 2\ Zn(l) + CS_2(g)$ $\Delta H^{\ominus} = +733\ kJ\ mol^{-1}$
   $\Delta S^{\ominus} = +218\ J\ K^{-1}\ mol^{-1}$
   This reaction is **not** feasible at 800 °C.
   Calculate $\Delta G^{\ominus}$ for this reaction at 800 °C.

   $\Delta G^{\ominus}$ = ...(ans = 499.086) ..... kJ mol^{-1} [2]

2. The formation of CaF_2 at 298 K is shown.
   $Ca(s) + F_2(g) \rightarrow CaF_2(s)$ $\Delta H^{\ominus} = -1214\ kJ\ mol^{-1}$
   $\Delta G^{\ominus} = -1162\ kJ\ mol^{-1}$
   Calculate the entropy change, $\Delta S^{\ominus}$, in J K^{-1} mol^{-1}, for this reaction.

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

3. 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.

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

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

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

   $\Delta S^{\ominus}$ = ...(ans = -0.310)..... units = ...(ans = kJ mol^{-1}).....[2]

   (iii) 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]

4. Barium hydroxide reacts readily with ammonium chloride on mixing at room temperature.
$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4Cl\ (s) \rightarrow 2NH_3(g) + BaCl_2 \cdot 2H_2O(s) + 8H_2O(l)\ \Delta H_r = +133\ kJ\ mol^{-1}$
Some relevant standard entropies are given in Table 1.2.

substance	$Ba(OH)_2 \cdot 8H_2O(s)$	$NH_4Cl\ (s)$	$NH_3(g)$	$BaCl_2 \cdot 2H_2O(s)$	$H_2O(l)$
$S^{\ominus}$/ J K^{-1} mol^{-1}	427	95	192	203	70

Calculate the standard Gibbs free energy change, $\Delta G^{\ominus}$, for this reaction at 25 ° C.

$\Delta G^{\ominus}$ = ...(ans = -24.9) ..... kJ mol^{-1} [3]

5. The three physical states of H_2O have different standard entropies, $\Delta S^{\ominus}$, associated with them.
Table 2.1 shows these $\Delta S^{\ominus}$ values.

state of H_2O standard entropy,	$\Delta S^{\ominus}$ / J K^{-1} mol^{-1}
solid	+48.0
liquid	+70.1
gas	+188.7

The energy changes for $H_2O(s) \rightarrow H_2O(l)$ are shown.

$\Delta G^{\ominus} = 0.00\ kJ\ mol^{-1}$
$\Delta H = +6.03\ kJ\ mol^{-1}$

Use these data to show that the melting point of $H_2O(s)$ is 0 °C. [1]

6. Carbon disulfide, CS_2, is flammable and reacts readily with oxygen, as shown in reaction 1.
reaction 1 $CS_2(g) + 3O_2(g) \rightleftharpoons CO_2(g) + 2SO_2(g)$
Table 3.1 shows the standard enthalpy of formation, $\Delta H_f^{\ominus}$, and the standard entropy, $S^{\ominus}$, for some substances.
Table 3.1

	$CS_2(g)$	$O_2(g)$	$CO_2(g)$	$SO_2(g)$
$\Delta H_f^{\ominus}$ / kJ mol^{-1}	116.7	0.0	–393.5	–296.8
$S^{\ominus}$/ J K–1 mol^{-1}	237.8	205.2	213.8	248.2

Calculate the standard Gibbs free energy change, $\Delta G^{\ominus}$, in kJ mol^{-1}, for reaction 1 at 25 °C.

$\Delta G^{\ominus}$= ..... (ans = - 1061) ..... kJ mol^{-1} [3]

7. Hydrogen peroxide is a liquid at 298 K. It is moderately stable under room conditions but will decompose quickly if a catalyst is added.
reaction 1 $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$
Some standard entropies, $S^{\ominus}$, are shown in Table 3.2.

Table 3.2

substance	$S^{\ominus}$ / J K^{-1} mol^{-1}
$H_2O_2(l)$	+102
$H_2O(l)$	+70

The enthalpy change and Gibbs free energy change for the following reaction are shown.
$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$ $\Delta H^{\ominus} = –196$ kJ mol^{-1}

$\Delta G^{\ominus} = –238$ kJ mol^{-1}

Use the data given to calculate the standard entropy of oxygen, $S^{\ominus}$, $O_2(g)$.

$S^{\ominus}$, $O_2(g)$ = ... (ans = 204.94) ..... J K^{-1} mol^{-1} [3]

Prediction of feasibility

8. Carbon disulfide reacts with chlorine to form tetrachloromethane, as shown in reaction 2.
reaction 2 $CS_2 + 3Cl_2 \rightleftharpoons CCl_4 + S_2Cl_2$ $\Delta H^{\ominus} = -261.6$ kJ mol^{-1}
$\Delta S^{\ominus} = -365.5$ J K^{-1} mol^{-1}

Calculate the maximum temperature, in K, for reaction 2 to be feasible.

temperature = ... (ans = 716) ..... K [2]

9. KI slowly oxidises in air, forming I_2.
reaction 1 $4KI(s) + 2CO_2(g) + O_2(g) \rightarrow 2K_2CO_3(s) + 2I_2(s)$ $\Delta H^{\ominus} = -203.4$ kJ mol^{-1}
Table 1.2 shows some data relevant to this question.

substance	standard entropy, $\Delta S^{\ominus}$ / J K^{-1} mol^{-1}
$CO_2(g)$	213.6
$I_2(s)$	116.1
$K_2CO_3(s)$	155.5
$KI(s)$	106.3
$O_2(g)$	205.2

(i) Calculate the standard entropy change, $\Delta S^{\ominus}$, of reaction 1.

$\Delta S^{\ominus}$ = ... (ans = - 514.4) ..... J K^{-1} mol^{-1} [2]

(ii) Use your answer to (i) to show that reaction 1 is spontaneous at 298 K. [2]

10. Potassium iodide, KI, is used as a reagent in both inorganic and organic chemistry.
KI forms an ionic lattice that is soluble in water.
KI(s) has a high solubility in water, although its enthalpy change of solution is endothermic.
Explain how this high solubility is possible.

.....

.....

..... [2]

Predicting the effect of temperature on the feasibility of a reaction

11. At 298 K, the Gibbs free energy change, ΔG, for the solution of compound T is +6.00 kJ mol^{-1}.
The enthalpy change of solution, ΔH_{sol}, of compound T is +30.0 kJ mol^{-1} at 298 K.
(i) Calculate the value of the entropy change, ΔS, for the solution of compound T at 298 K.

ΔS = ... J K^{-1} mol^{-1} [2]

(ii) Predict whether compound T becomes more or less soluble as the water is heated from 298 K to 360 K.
Explain your answer.

…….....

…….....[1]

12. Calcium fluoride, $CaF_2(s)$, can be synthesised directly from its elements.
The value of $\Delta H_f^{\ominus}$ ($CaF_2(s)$) is –1214 kJ mol^{-1}.

(i) Predict the sign of the entropy change, $\Delta S^{\ominus}$, for this synthesis. Explain your answer.

The sign of the entropy change is .....

explanation
.....

.....[1]

(ii) Use the value of $\Delta H_f^{\ominus}$ ($CaF_2(s)$) given and your answer to (i) to predict how the feasibility for this synthesis will change with increasing temperature.

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…….....[2]

13. Iron(II) chloride, $FeCl_2$, is oxidised by chlorine to form iron(III) chloride, $FeCl_3$, under standard conditions.
$2FeCl_2(s) + Cl_2(g) \rightarrow 2FeCl_3(s)$ $\Delta H^{\ominus} = -128$ kJ mol^{-1}

Table 3.2

species	$S^{\ominus}$ / J K^{-1} mol^{-1}
$Cl_2(g)$	223
$FeCl_2(s)$	120
$FeCl_3(s)$	142

(i) Use Table 3.2 and other data to calculate the Gibbs free energy change, $\Delta G^{\ominus}$, for this reaction.
Show your working.

$\Delta G^{\ominus}$ = ...(ans = -74.7)..... kJ mol^{-1} [3]

(ii) Predict whether this reaction becomes more or less feasible at a higher temperature.
Explain your answer.

The reaction becomes ..... feasible.

explanation

.....

.....[1]

14. The feasibility of a chemical reaction depends on the standard Gibbs free energy change, $\Delta G^{\ominus}$.
This is dependent on the standard enthalpy and entropy changes, and the temperature.
Magnesium carbonate can be decomposed on heating.
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$ $\Delta H^{\ominus} = +117$ kJ mol^{-1}
Standard entropies are shown in Table 2.1.

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

(i) Calculate $\Delta G^{\ominus}$ for this reaction at 298 K.
Show your working.

$\Delta G^{\ominus}$ = ...(ans=64.79)..... kJ mol^{-1} [3]

(ii) Explain why this reaction is feasible only at high temperatures.

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15. The reaction of solid hydrated barium hydroxide, $Ba(OH)_2 \bullet 8H_2O$, with ammonium salts is endothermic. Calculate the minimum temperature at which the reaction of $Ba(OH)_2 \bullet 8H_2O$ with NH_4NO_3 becomes feasible. Show all your working.

$Ba(OH)_2 \bullet 8H_2O(s) + 2NH_4NO_3(s) \rightarrow 2NH_3(g) + Ba(NO_3)_2(s) + 10H_2O(l)$ $\Delta H_r = +132$ kJ mol^{-1}

$\Delta S^{\ominus} = +616$ J K^{-1} mol^{-1}

temperature = ... (ans = -58.7) ..... ° C [2]

16. (i) At 25 °C, the enthalpy change of solution of compound Z is +26 kJ mol^{-1}. The entropy change of solution of Z at the same temperature is +52 J K^{-1} mol^{-1}.
Calculate the value of the Gibbs free energy change, ΔG, for the solution of Z at 25 °C.

ΔG = ...(ans= 10.5)..... kJ mol^{-1} [2]

(ii) Use your answer to (i) to predict whether or not Z is soluble in water at 25 °C. Explain your answer.

…….....

……..... [1]

(iii) Predict whether Z becomes more or less soluble as the water is heated from 25 °C to 95 °C. Explain your answer.

…….....

……..... [1]

17. Silicon is the second most abundant element by mass in the Earth's crust.
In industry, silicon is extracted from SiO_2 by reaction with carbon at over 2000 ° C.
reaction 1 $SiO_2(s) + 2C(s) \rightarrow Si(l) + 2CO(g)$

Reaction 1 is highly endothermic.

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

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…….....

……..... [2]

18. The standard entropies, $S^{\ominus}$, of three species are given in the table.

species	$S^{\ominus}$ / J K^{-1} mol^{-1}
$H_2O(l)$	+70
$H_2(g)$	+131
$O_2(g)$	+205

(i) Calculate $\Delta S^{\ominus}$ for the reaction $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$.

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

(ii) $\Delta H^{\ominus}$ for the reaction $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ is +572 kJ mol^{-1}.
Calculate $\Delta G^{\ominus}$ for this reaction at 298 K.

$\Delta G^{\ominus}$ = ...(ans= 474.5)..... kJ mol^{-1} [2]

(iii) Predict the effect of increasing temperature on the spontaneity of this reaction.
Explain your answer.

…….....

..... [1]

19. The standard enthalpy change of solution, $\Delta H^{\ominus}_{sol}$, of $AgNO_3(s)$ in water is +22.6 kJ mol^{-1}.
Suggest how the feasibility of dissolving $AgNO_3(s)$ in water changes with temperature.
Explain your answer.

…….…….....

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……..... [2]

20. The standard entropies, ΔS, of three species are given in the table.

species	$\Delta S^{\ominus}$ / J K^{-1} mol^{-1}
HCl (g)	+187
$H_2(g)$	+131
$Cl_2(g)$	+223

(i) Calculate $\Delta S^{\ominus}$ for the reaction $2HCl\ (g) \rightarrow H_2(g) + Cl_2(g)$.

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

(ii) $\Delta H^{\ominus}$ for the reaction $2\ HCl\ (g) \rightarrow H_2(g) + Cl_2(g)$ is +185 kJ mol^{-1}.
Calculate $\Delta G^{\ominus}$ for this reaction at 298 K.

$\Delta G^{\ominus}$ = ... kJ mol^{-1} [2]

(iii) Predict the effect of increasing temperature on the spontaneity of this reaction. Explain your answer.

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…… [1]

21. (i) Potassium chloride is very soluble in water at 20 ° C.
Explain the solubility of potassium chloride by reference to change in entropy, ΔS.

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…… [1]

(ii) Use the Gibbs equation and your answer to (i) to predict whether potassium chloride is more soluble in water at 20 ° C or at 80 ° C. Explain your answer.

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…… [1]

22. 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 .....

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……

……. [2]

23. The equation for the reduction of iron(III) oxide by carbon monoxide at 450 °C is shown.
$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ $\Delta G^{\ominus} = -36.2$ kJ mol^{-1}
Table 3.1 shows the enthalpy of formation, ΔH_f, and the entropy, S, for some substances.

Table 3.1

	$Fe_2O_3(s)$	$CO(g)$	$Fe(s)$	$CO_2(g)$
ΔH_f / kJ mol^{-1}	–824.2	–110.5	0.0	–393.5
S / J K^{-1} mol^{-1}	87.4	to be calculated	27.3	213.8

Use the data in Table 3.1 to calculate the entropy, S, of carbon monoxide at 450 °C.
Show your working.

S of CO(g) = ... J K^{-1} mol^{-1} [3]

(d) Iron(II) oxide can also be reduced to iron by carbon monoxide, as shown.

$FeO(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$ $\Delta H = -11.1$ kJ mol^{-1}
$\Delta S = -15.2$ J K^{-1} mol^{-1}

State the effect of increasing temperature on the feasibility of this reaction.
Explain your answer.

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……. [2]

24. Anhydrous barium chloride can be obtained from the hydrated salt, as shown in reaction 1.

reaction 1 $BaCl_2 \cdot 2H_2O(s) + 2SOCl_2(l) \rightarrow BaCl_2(s) + 2SO_2(g) + 4HCl\ (g)$

(i) Describe one observation when reaction 1 is carried out.

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..... [1]

(ii) The entropy change, ΔS^o, for reaction 1 at 25 °C is +768 J K^{-1} mol^{-1}.
Explain why ΔS^o has a large positive value.

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..... [1]

(iii) Table 2.1 shows the enthalpy changes of formation, $\Delta H_f^{\ominus}$, for the compounds in reaction 1.

Table 2.1

Compound	$\Delta H_f^{\ominus}$ / kJ mol^{-1}
$BaCl_2$ (s)	–859
$BaCl_2 \cdot 2H_2O(s)$	–1460
$SOCl_2(l)$	–246
$SO_2(g)$	–297
HCl (g)	–92

Calculate the standard Gibbs free energy change, $\Delta G^{\ominus}$, in kJ mol^{-1}, for reaction 1 at 25 °C.

$\Delta G^{\ominus}$ = ...(ans = 97.7)..... kJ mol^{-1} [3]

25. The exhaust systems of many diesel-fuelled cars contain an additional system to reduce vehicle emissions. This uses a liquid that is added to the exhaust system.
This liquid contains urea, $(NH_2)_2CO$, which decomposes on heating to isocyanic acid, HNCO, and ammonia.

reaction 1: $(NH_2)_2CO \rightarrow HNCO + NH_3$

Isocyanic acid reacts with water vapour to form ammonia and carbon dioxide.

reaction 2: $HNCO(g) + H_2O(g) \rightarrow NH_3(g) + CO_2(g)$

Some values for standard enthalpy changes of formation, $\Delta H_f^{\ominus}$, and standard entropies, S o, are given in Table.

Compound	ΔH_f^o / kJ mol^{-1}	S o / J K^{-1} mol^{-1}
HNCO(g)	–101.7	+238.2
$H_2O(g)$	–241.8	+188.8
$NH_3(g)$	–45.9	+192.8
$CO_2(g)$	–393.5	+213.8

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

.....

..... [1]

(ii) Use the data in Table 5.1 to calculate ΔG^o for reaction 2 at 25 ° C.
Show your working.

$\Delta G^{\ominus}$ = ...(ans= 89.9)..... kJ mol^{-1} [4]

26. NH_4IO_3 is an unstable compound that readily decomposes when warmed. The decomposition reaction is shown.
$NH_4IO_3(s) \rightarrow 1/2\ N_2(g) + 1/2O_2(g) + 1/2I_2(g) + 2H_2O(l)$ $\Delta H = -154.6$ 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 and your answer to (i).

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..... [1]

27. 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: $Si(s) + 3\,HCl\,(g) \rightarrow SiHCl_3(g) + H_2(g)$

(i) Table 2.1 shows some standard entropy data.

Table 2.1

compound	standard entropy, S° / 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 ΔS° for reaction 2.

ΔS° = ...(-135)..... 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}

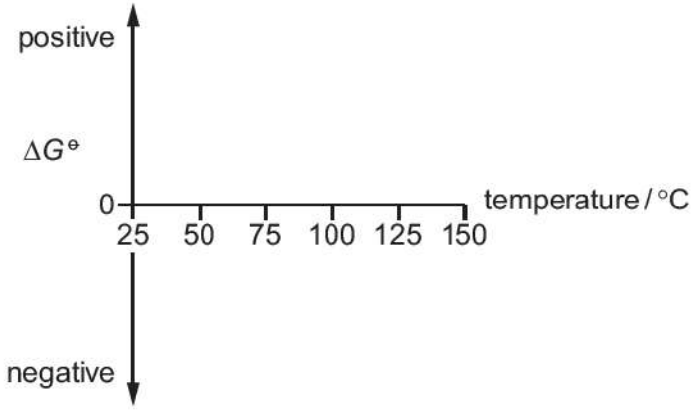
Use this information and your answer to (i) to calculate the temperature, in K, at which reaction 3 becomes feasible.

Show your working.

temperature = ...(ans = 1624)..... K [2]

28. 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.



29. The reaction between A and B is feasible at low temperatures but is not feasible at high temperatures.
$A + B \rightarrow C + D$
Deduce the signs of ΔH and ΔS for this reaction and explain why the feasibility changes with temperature.

sign of ΔH = ... sign of ΔS = ...

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..... [2]

Thermal stabilities of nitrates and carbonates

30. Write an equation for the decomposition of $Ca(NO_3)_2$. [1]

31. Lead(II) nitrate, $Pb(NO_3)_2$, decomposes on heating in a similar manner to the Group 2 nitrates.
Write an equation for the decomposition of lead(II) nitrate.

..... [1]

Copper(II) nitrate decomposes in a similar manner to Group 2 nitrates.
Write an equation for the decomposition of $Cu(NO_3)_2$. [1]

Calcium nitrate, $Cu(NO_3)_2$, decomposes on heating.
Write an equation for the decomposition of calcium nitrate. [1]

32. Write an equation for the thermal decomposition of $CaCO_3$.

33. Copper(II) carbonate decomposes on heating in a similar way to the carbonates of Group 2.
Write an equation for the decomposition of copper(II) carbonate.

..... [1]

34. Group 2 nitrates decompose when heated.
Describe how the thermal stability of Group 2 nitrates changes with increasing proton number.
Explain your answer.

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..... [3]

Explain why the thermal stability of the Group 2 nitrates increases down the group. [2]

Describe the trend in the decomposition temperature of the Group 2 nitrates.
Explain your answer. [3]

35. Lithium nitrate, $LiNO_3$, decomposes on heating in a similar way to Group 2 nitrates to give the metal oxide, a brown gas, and oxygen.
(i) Write an equation for the decomposition of $LiNO_3$.

……..... [1]

(ii) The other Group 1 nitrates, MNO_3, decompose on heating to form the metal nitrite, MNO_2, and oxygen. The thermal stability of these nitrates increases down the group.
Suggest why the thermal stability of MNO_3 increases down the group.

……..... [2]

36. $Zn(NO_3)_2$ undergoes thermal decomposition when heated. The reaction is similar to the thermal decomposition of Group 2 nitrates.
(i) Construct an equation for the thermal decomposition of $Zn(NO_3)_2$. [1]

(ii) The radii of some Group 2 cations and Zn^{2+} are shown in Table 1.2.

cation	Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}	Zn^{2+}
radius/ pm	65	99	113	135	74

State and explain the trend in thermal stability of the Group 2 nitrates down the group. [2]

(iii) Use Table 1.2 to suggest which Group 2 nitrates are less thermally stable than zinc nitrate.

..... [1]

37. Calcium nitrate, $Ca(NO_3)_2$, is a white crystalline solid. When heated, it starts to decompose at approximately 500°C.

(i) Suggest temperatures at which $Mg(NO_3)_2$ and $Ba(NO_3)_2$ start to decompose.
Explain your answer.
temperature at which $Mg(NO_3)_2$ starts to decompose ..... °C

temperature at which $Ba(NO_3)_2$ starts to decompose .....°C

explanation .....

……..... [3]

38. Copper(II) nitrate, $Cu(NO_3)_2$, and barium nitrate, $Ba(NO_3)_2$, both decompose when heated.
Copper(II) nitrate decomposes at a lower temperature than barium nitrate.
Suggest a reason for this difference. Explain your answer.

..... [2]

39. Describe the trend in the thermal stabilities of the carbonates of the Group 2 elements.
Explain your answer.

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..... [3]

40. Both calcium carbonate, $CaCO_3$, and barium carbonate, $BaCO_3$, decompose when heated to form the metal oxide and a gas.
State which of $CaCO_3$ and $BaCO_3$ decomposes at a lower temperature.
Explain your answer.

The compound that decomposes at a lower temperature is ..... .

explanation .....

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..... [2]

41. The Group 1 carbonates are much more thermally stable than the Group 2 carbonates.
State and explain the trend in the thermal stability of the Group 2 carbonates. [2]

42. Group 2 metals form stable carbonates and sulfates.
State and explain the trend in the thermal stability of the Group 2 carbonates down the group. [3]

43. The trend in the decomposition temperatures of Group 2 peroxides, MO_2, is similar to that of Group 2 carbonates.
Suggest which of barium peroxide, BaO_2, and calcium peroxide, CaO_2, will decompose at the lower temperature. Explain your answer.

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..... [2]

44. Magnesium iodate(V), $Mg(IO_3)_2$, decomposes when heated to form magnesium oxide, oxygen and iodine.
Construct an equation for this reaction.

..... [1]

45. Group 2 hydroxides decompose on heating to give the corresponding metal oxide and water vapour.
Suggest which of $Mg(OH)_2$ and $Sr(OH)_2$ will decompose at a lower temperature.
Explain your answer.

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..... [2]

46. Group 1 hydrogencarbonates, $MHCO_3$, decompose on gentle heating to give the corresponding metal carbonate, carbon dioxide and water vapour.
(i) Write an ionic equation for the decomposition of the hydrogencarbonate ion.

..... [1]

(ii) The thermal stability of Group 1 hydrogencarbonates increases down the group.
Suggest an explanation for the trend in thermal stability of the Group 1 hydrogencarbonates.

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..... [2]

47. Group 2 carbonates decompose when heated to form the metal oxide and carbon dioxide.
Suggest a mechanism for the decomposition of the carbonate ion by adding **two** curly arrows in Fig. 1.1. [1]

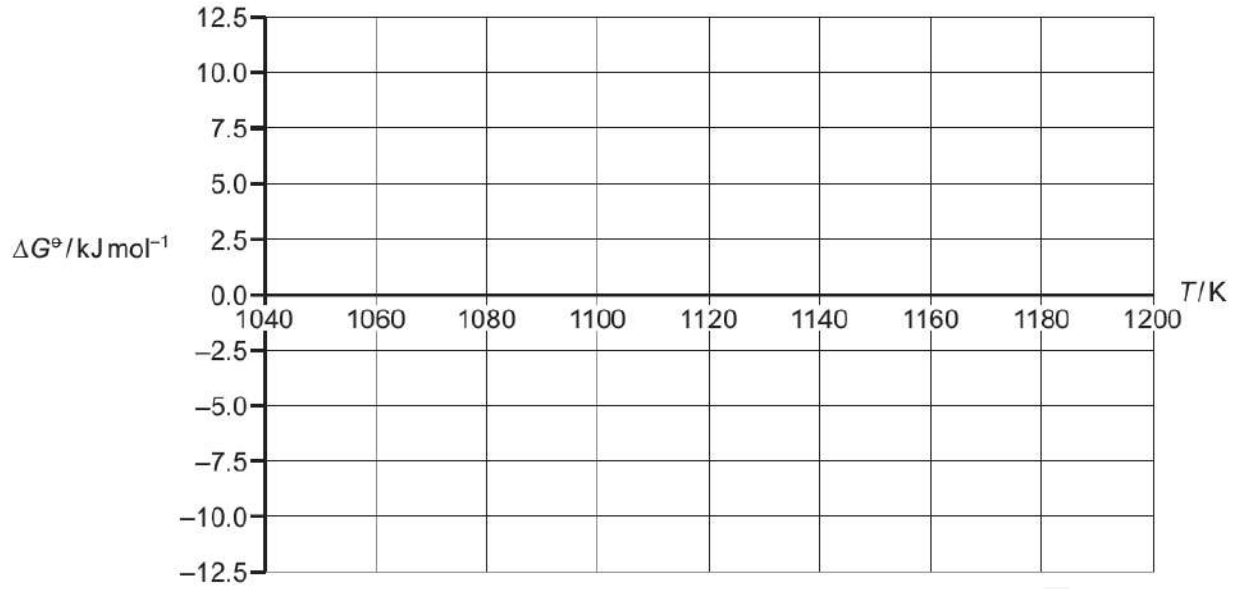
Fig. 1.1

48. Calcium carbonate decomposes on heating.
$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
Table 4.1 shows the values of the Gibbs free energy change, $\Delta G^{\ominus}$, for this reaction at various temperatures.

Table 4.1

T / K	$\Delta G^{\ominus}$ / kJ mol–1
1050	9.9
1085	4.3
1120	–1.3
1148	–5.8
1176	–10.3

Assume the standard enthalpy change, ΔH o, and the standard entropy change, ΔS o, for this reaction remain constant over this temperature range.
Use the data in Table 4.1 to plot a graph of $\Delta G^{\ominus}$ against T on the grid. [2]



(ii) Calculate the gradient of your graph. Determine the ΔS^{o} in $J\ K^{-1}\ mol^{-1}$ for this reaction.
Show all working.

ΔS^{o} = ... $J\ K^{-1}\ mol^{-1}$ [2]

49. Iron(II) carbonate, $FeCO_3$, and nickel(II) carbonate, $NiCO_3$, both decompose when heated.
$FeCO_3$ decomposes at a lower temperature than $NiCO_3$.
Suggest a possible reason for this difference. Explain your answer. **9701_w22_qp_43**

Marking Scheme
Fe^{2+} is smaller / has a smaller radius
OR Fe^{2+} greater charge density [1]
polarises/distorts the anion / $CO_3^{(2)-}$ more [1] [2]

I believe this is a wrong question and wrong answer accordingly. As the Fe^{2+} ion is bigger than the Ni^{2+} ion.

Decomposition of Oxalates

50. Barium ethanedioate, BaC_2O_4, decomposes on heating to produce barium oxide and a mixture of two different gases.
Construct an equation for the decomposition of barium ethanedioate.

..... [1]

(ii) The trend in the decomposition temperatures of the Group 2 ethanedioates is similar to that of the Group 2 nitrates.
Suggest which of CaC_2O_4 and BaC_2O_4 will decompose at the lower temperature. Explain your answer.

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..... [2]

51. Some of the ionic compounds of Group 2 elements undergo thermal decomposition.
Thermal decomposition of solid anhydrous magnesium ethanedioate, MgC_2O_4, occurs above 650 °C. The products are magnesium oxide and a mixture of two different gases, one of which gives a white precipitate with saturated calcium hydroxide solution.
(a) Complete the equation for the thermal decomposition of MgC_2O_4.

MgC_2O_4 [1]

(b) Suggest which of MgC_2O_4 or CaC_2O_4 undergoes thermal decomposition at a lower temperature. Explain your answer. [2]

52. When a sample of hydrated lithium ethanedioate, $Li_2C_2O_4•H_2O$, is gently heated, two gaseous products are formed, and a white solid residue remains.
The residue is added to $HNO_3(aq)$. A gas is produced that turns limewater milky.
Complete the equation for the decomposition of $Li_2C_2O_4•H_2O$.

$Li_2C_2O_4•H_2O$ ..... + ... + ... [1]

..... [1]

Stoichiometric Calculations

53. Separate samples of 0.02 mol of calcium carbonate and 0.02 mol of barium carbonate are heated until completely decomposed to the metal oxide and carbon dioxide.
(a) State which of these two Group 2 carbonates requires the **higher** temperature before it begins to decompose. Explain your answer.

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..... [2]
(b) After decomposition is complete, the 0.02 mol sample of calcium oxide is taken and added to 2.00 dm^3 of water. A solution is formed with no solid present. Dilute sulfuric acid is then added dropwise until a precipitate is seen.
The same procedure is repeated with the 0.02 mol sample of barium oxide, using the same concentration solution of dilute sulfuric acid.
Identify the sample to which the **most** sulfuric acid must be added to cause a precipitate to appear.
Explain your answer. You should refer to the solubilities of the precipitates and relevant energy terms in your answer.

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..... [3]
(c) (i) Calculate the mass, in g, of CO_2 produced by the decomposition of 0.020 moles of calcium carbonate.

mass of CO_2 = ... g [1]

(ii) Calculate the minimum mass, in g, of propane that would, on complete combustion, produce the same mass of CO_2 calculated in **(c)(i)**.
Give your answer to three significant figures.

mass of propane = ... g [2]

54. Separate samples of 0.01 mol of magnesium nitrate and 0.01 mol of strontium nitrate are heated until completely decomposed to the metal oxide, nitrogen dioxide and oxygen.
(a) State which of these two Group 2 nitrates requires the **higher** temperature before it begins to decompose. Explain your answer.

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..... [2]

(b) After decomposition is complete the 0.01 mol sample of magnesium oxide is taken and increasing amounts of water are added to it, with stirring, until no solid remains.
This procedure is repeated with the 0.01 mol sample of strontium oxide.
Identify the sample to which most water must be added to cause all the solid to dissolve.
Explain your answer by reference to the solubilities of the products formed when water is added to the oxides. You should refer to relevant energy terms in your answer.

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..... [3]

(c) The nitrogen dioxide given off by the decomposition of 0.0100 mol of strontium nitrate is dissolved in water. The oxidising agent $H_2O_2(aq)$ is then added to give 150.0 cm^3 of a solution in which nitric acid, HNO_3, is the only nitrogen-containing product.
(i) Calculate the concentration, in mol dm^{-3}, of HNO_3 in the 150.0 cm^3 of solution.

concentration = ... mol dm^{-3} [1]

(i) The HNO_3 present in 25.0 cm^3 of this solution is neutralised using 0.125 mol dm^{-3} NaOH(aq).
Calculate the minimum volume, in cm^3, of NaOH(aq) needed. Give your answer to three significant figures.

volume = ... cm^3 [1]

***... ***