

## 1. 9701\_w19\_qp\_43 Q3

- 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,  $\Delta H_{\text{latt}}$  CaCl<sub>2</sub>(s), takes place

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

.....

[3]

- (c) The reaction ZnCO<sub>3</sub>(s) → ZnO(s) + CO<sub>2</sub>(g) is not spontaneous at room temperature.

- (i) Give the full name for the term  $\Delta G^\circ$ .

..... [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]

## 2. 9701\_s19\_qp\_42 Q5

- 5 (a) Complete the table by placing **one** tick ( $\checkmark$ ) in each row to indicate the sign of each type of energy change under standard conditions.

energy change	always positive	always negative	either negative or positive
lattice energy			
enthalpy change of neutralisation			

[1]

- (b) Define, in words, the term *enthalpy change of solution*.

.....  
.....

[1]

- (c) The following enthalpy changes are given.

enthalpy change	value / kJ mol <sup>-1</sup>
standard enthalpy change of formation, $\Delta H_f^\ominus$ , for $K_3PO_4(s)$	-2035
standard enthalpy change, $\Delta H^\ominus$ , for $P(s) + 2O_2(g) + 3e^- \rightarrow PO_4^{3-}(aq)$	-1284
standard enthalpy change, $\Delta H^\ominus$ , for $K(s) \rightarrow K^+(aq) + e^-$	-251

Determine the standard enthalpy change of solution of potassium phosphate,  $K_3PO_4(s)$ . It may be helpful to draw a labelled energy cycle.

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

- (d) Some lattice energy values are shown in the table.

compound	lattice energy value / $\text{kJ mol}^{-1}$
$\text{CaBr}_2(\text{s})$	-2176
$\text{KBr}(\text{s})$	-679

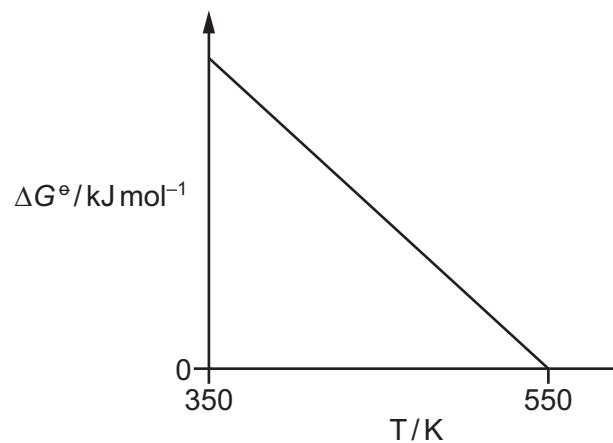
Suggest an explanation for why  $\Delta H_{\text{latt}}^\circ \text{ CaBr}_2$  is **more** exothermic than  $\Delta H_{\text{latt}}^\circ \text{ KBr}$ .

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

[2]

- (e) For a particular gas phase reaction the variation in standard Gibbs free energy change,  $\Delta G^\circ$ , with temperature is shown.

Assume standard enthalpy change,  $\Delta H^\circ$ , and standard entropy change,  $\Delta S^\circ$ , remain constant with temperature.



- (i) Write the equation that relates  $\Delta G^\circ$  to  $\Delta H^\circ$  and  $\Delta S^\circ$ .

..... [1]

- (ii) Use this equation to explain why  $\Delta G^\circ$  becomes **less** positive as temperature increases in this reaction.

.....  
.....

[1]

[Total: 9]

**3. 9701\_s18\_qp\_42 Q1**

- 1 Silicon tetrachloride,  $\text{SiCl}_4$ , is formed when silicon reacts with chlorine under suitable conditions. It is a colourless liquid with a low boiling point.

(a) Explain why  $\text{SiCl}_4$  has a low boiling point.

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

[2]

(b)  $\text{SiCl}_4$  reacts with water to produce an acidic solution.

(i) Write an equation for this reaction.

.....

[1]

(ii) Describe **two** visual observations when silicon tetrachloride is added drop by drop to a small amount of water.

1 .....

2 .....

[2]

(iii) A sample of 0.8505 g of  $\text{SiCl}_4$  is added to 800 cm<sup>3</sup> of water. All of the soluble acidic product is dissolved in the water.

Calculate the pH of the solution obtained.

pH = ..... [3]

- (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.

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

- (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  $-640 \text{ kJ mol}^{-1}$ .

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.

$$\text{temperature} = \dots \text{ K} \quad [2]$$

[Total: 13]

**4. 9701\_s17\_qp\_43 Q1**

- 1 (a) Describe and explain the variation in the solubilities of the hydroxides of the Group 2 elements.

.....  
 .....  
 .....  
 .....  
 ..... [4]

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,  $\Delta H_{r1}^\ominus$ .



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

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,  $\Delta H_{r2}^\ominus$ , =  $-26 \text{ kJ mol}^{-1}$ .

- (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.

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

- (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,  $\Delta H_{r3}^\ominus$ .



$$\Delta H_{r3}^\ominus = \dots \text{ kJ mol}^{-1} \quad [1]$$

- (iv) How would the value of  $\Delta H_{r3}^\ominus$  compare with the value of  $\Delta H_{r4}^\ominus$  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  $\Delta S_{r1}^\ominus$ .

Suggest, with a reason, how the standard entropy change for reaction 3 might compare with  $\Delta S_{r1}^\ominus$ .

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

[Total: 13]

## 5. 9701\_w19\_qp\_42 Q4

- 4 The table shows some standard entropy data.

substance	standard entropy, $S^\circ$ $/ \text{JK}^{-1} \text{mol}^{-1}$
PbO <sub>2</sub> (s)	77
PbO(s)	69
O <sub>2</sub> (g)	205

Lead(IV) oxide, PbO<sub>2</sub>, decomposes to lead(II) oxide, PbO, and oxygen when heated.



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

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

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

$$T = \dots \text{ K} \quad [3]$$

- (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.

.....  
.....

[Total: 6]

## 6. 9701\_w19\_qp\_41 Q3

- 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,  $\Delta H_{\text{latt}}$  CaCl<sub>2</sub>(s), takes place

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

.....

[3]

- (c) The reaction ZnCO<sub>3</sub>(s) → ZnO(s) + CO<sub>2</sub>(g) is not spontaneous at room temperature.

- (i) Give the full name for the term  $\Delta G^\circ$ .

..... [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]

**7. 9701\_s17\_qp\_41 Q1**

- 1 (a)** Describe and explain the variation in the solubilities of the hydroxides of the Group 2 elements.

.....  
 .....  
 .....  
 .....  
 ..... [4]

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,  $\Delta H_{r1}^\ominus$ .



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

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,  $\Delta H_{r2}^\ominus$ , =  $-26 \text{ kJ mol}^{-1}$ .

- (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.

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

- (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,  $\Delta H_{r3}^\ominus$ .



$$\Delta H_{r3}^\ominus = \dots \text{ kJ mol}^{-1} \quad [1]$$

- (iv) How would the value of  $\Delta H_{r3}^\ominus$  compare with the value of  $\Delta H_{r4}^\ominus$  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  $\Delta S_{r1}^\ominus$ .

Suggest, with a reason, how the standard entropy change for reaction 3 might compare with  $\Delta S_{r1}^\ominus$ .

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

[2]

[Total: 13]

## 8. 9701\_s20\_qp\_42 Q2

- 2 (a) Describe and explain how the solubility of the Group 2 sulfates varies down the group.

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.....  
.....  
.....  
..... [4]

- (b) The trend in the decomposition temperatures of Group 2 peroxides,  $\text{MO}_2$ , is similar to that of Group 2 carbonates.

Suggest which of barium peroxide,  $\text{BaO}_2$ , and calcium peroxide,  $\text{CaO}_2$ , will decompose at the **lower** temperature. Explain your answer.

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

- (c) Magnesium iodate(V),  $\text{Mg}(\text{IO}_3)_2$ , decomposes when heated to form magnesium oxide, oxygen and iodine.

Construct an equation for this reaction.

..... [1]

- (d) Calcium iodate(V),  $\text{Ca}(\text{IO}_3)_2$ , is sparingly soluble in water.

The concentration of its saturated solution is  $5.6 \times 10^{-3} \text{ mol dm}^{-3}$  at 298 K.

- (i) Write an expression for the solubility product,  $K_{\text{sp}}$ , of  $\text{Ca}(\text{IO}_3)_2$ , and state its units.

$$K_{\text{sp}} =$$

$$\text{units} = \dots \quad [2]$$

- (ii) Calculate the numerical value for  $K_{\text{sp}} \text{ Ca}(\text{IO}_3)_2$  at 298 K.

$$K_{\text{sp}} = \dots \quad [1]$$

- (iii) When a few cm<sup>3</sup> of concentrated Ca(NO<sub>3</sub>)<sub>2</sub>(aq) is added to a saturated solution of Ca(IO<sub>3</sub>)<sub>2</sub> a white precipitate forms.

Identify the white precipitate and give an explanation for this observation.

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.....  
.....

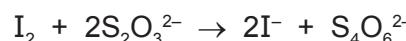
[2]

- (e) Iodised salt is sodium chloride mixed with a small amount of sodium iodate(V), NaIO<sub>3</sub>.

- 50.00 g of iodised salt is dissolved in distilled water and the solution made up to 250 cm<sup>3</sup> in a volumetric flask with distilled water.
- 50.0 cm<sup>3</sup> of this solution is pipetted into an excess of aqueous acidified potassium iodide.



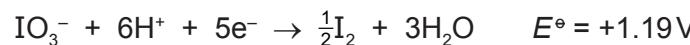
- The iodine produced requires 12.40 cm<sup>3</sup> of 0.00200 mol dm<sup>-3</sup> aqueous sodium thiosulfate solution for complete reaction.



Calculate the mass of sodium iodate(V) present in 50.00 g of iodised salt.

$$\text{mass of NaIO}_3 = \dots \text{g} \quad [3]$$

- (f) The half-equation for the reduction of iodate(V) ions is shown.

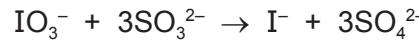


Use data from the *Data Booklet* to predict whether a reaction is feasible when aqueous solutions of acidified iodate(V) ions and bromide ions are mixed. Explain your answer.

.....  
.....

[1]

- (g) Iodate(V) ions react with sulfite ions in acidic solution at pH 5.00 as shown.



The initial rate of reaction was found to be first order with respect to  $\text{IO}_3^-$ , first order with respect to  $\text{SO}_3^{2-}$  and first order with respect to  $\text{H}^+$ .

- (i) Write the rate equation for this reaction, stating the units of the rate constant,  $k$ .

rate = ..... mol dm<sup>-3</sup> s<sup>-1</sup>

units of  $k$  = ..... [2]

- (ii) The rate of reaction depends on the pH of the solution. Assume all other concentrations remain the same.

Use the expression  $x = \frac{\text{rate at pH 5.00}}{\text{rate at pH 4.00}}$  to calculate the value of  $x$ .

$x = \dots$  [1]

[Total: 19]

## 9. 9701\_s22\_qp\_41 Q1

- 1 (a) The solubility of the Group 2 sulfates decreases down the group.

Explain this trend.

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

[3]

- (b) Describe what is observed when magnesium and barium are reacted separately with an excess of dilute sulfuric acid.

magnesium .....

barium .....

[1]

- (c) The solubility product,  $K_{\text{sp}}$ , of  $\text{BaSO}_4$  is  $1.08 \times 10^{-10} \text{ mol}^2 \text{ dm}^{-6}$  at 298 K.

Calculate the solubility of  $\text{BaSO}_4$  in g per  $100 \text{ cm}^3$  of solution.

solubility of  $\text{BaSO}_4$  = ..... g per  $100 \text{ cm}^3$  of solution

[2]

(d) (i) The equation for the formation of a gaseous sulfate ion is shown.



Calculate the standard enthalpy change of formation,  $\Delta H_f^\ominus$ , of  $SO_4^{2-}(g)$ . It may be helpful to draw a labelled energy cycle. Use relevant data from Table 1.1 in your calculations.

**Table 1.1**

energy change	value / kJ mol <sup>-1</sup>
lattice energy of barium sulfate, $BaSO_4(s)$	-2469
standard enthalpy change of formation of barium sulfate	-1473
standard enthalpy change of atomisation of barium	+180
first ionisation energy of barium	+503
second ionisation energy of barium	+965
standard enthalpy change of atomisation of sulfur	+279
standard enthalpy change for $S(g) \rightarrow S^{2-}(g)$	+440
standard enthalpy change for $O(g) \rightarrow O^{2-}(g)$	+657
$O=O$ bond energy	+496

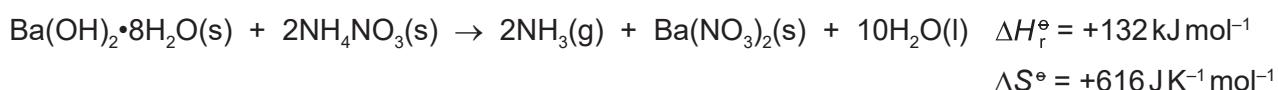
$$\Delta H_f^\ominus \text{ of } SO_4^{2-}(g) = \dots \text{ kJ mol}^{-1} \quad [3]$$

- (ii) Suggest how the lattice energy of  $\text{BaSO}_4(\text{s})$  differs from the lattice energy of  $\text{Cs}_2\text{SO}_4(\text{s})$ . Explain your answer.

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

- (e) The reaction of solid hydrated barium hydroxide,  $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$ , with ammonium salts is endothermic.

- (i) Calculate the **minimum** temperature at which the reaction of  $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$  with  $\text{NH}_4\text{NO}_3$  becomes feasible. Show all your working.



$$\text{temperature} = \dots \text{ }^\circ\text{C} \quad [2]$$

- (ii) Barium hydroxide reacts readily with ammonium chloride on mixing at room temperature.



Some relevant standard entropies are given in Table 1.2.

**Table 1.2**

substance	$\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}(\text{s})$	$\text{NH}_4\text{Cl}(\text{s})$	$\text{NH}_3(\text{g})$	$\text{BaCl}_2 \cdot 2\text{H}_2\text{O}(\text{s})$	$\text{H}_2\text{O}(\text{l})$
$S^\circ / \text{JK}^{-1} \text{ mol}^{-1}$	427	95	192	203	70

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

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

[Total: 16]

## 10. 9701\_s22\_qp\_42 Q4

- 4 (a) Calcium carbonate decomposes on heating.



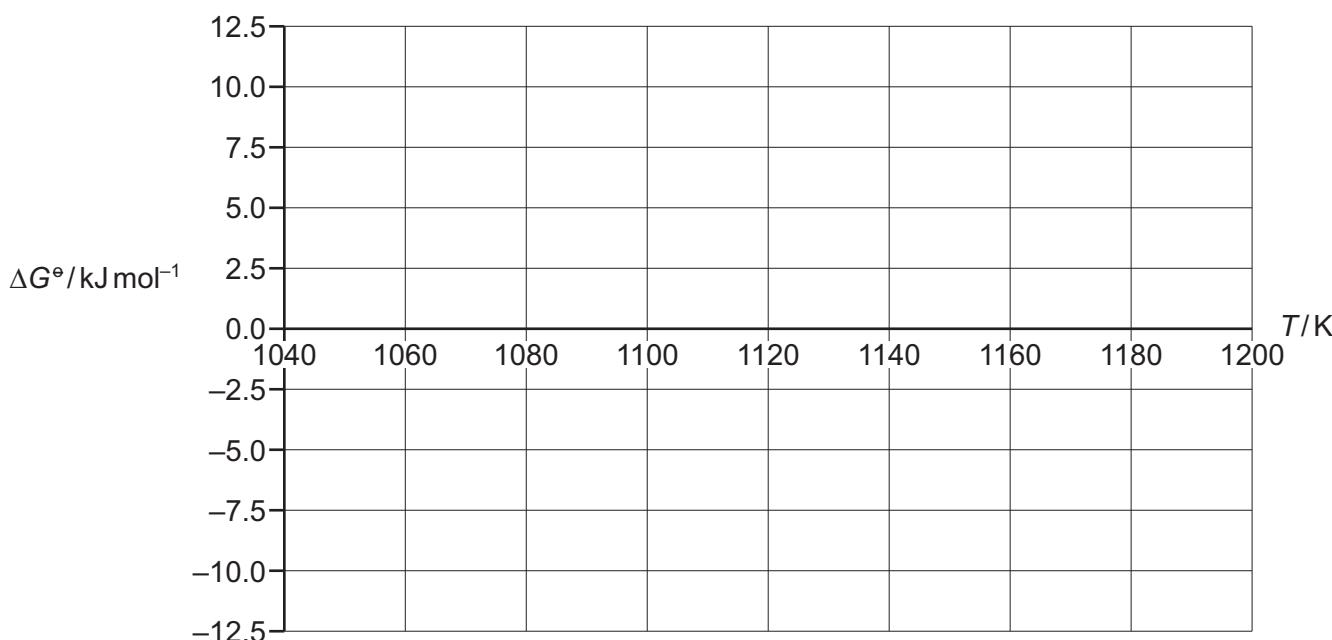
Table 4.1 shows the values of the Gibbs free energy change,  $\Delta G^\circ$ , for this reaction at various temperatures.

**Table 4.1**

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

Assume the standard enthalpy change,  $\Delta H^\circ$ , and the standard entropy change,  $\Delta S^\circ$ , for this reaction remain constant over this temperature range.

- (i) Use the data in Table 4.1 to plot a graph of  $\Delta G^\circ$  against  $T$  on the grid.



[2]

- (ii) Calculate the gradient of your graph. Determine the  $\Delta S^\circ$  in  $\text{JK}^{-1}\text{mol}^{-1}$  for this reaction. Show all working.

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

- (b) Group 1 hydrogencarbonates,  $\text{MHCO}_3$ , 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.

.....

.....

.....

- (c) The buffer system in seawater contains a mixture of  $\text{HCO}_3^-$  and  $\text{H}_2\text{CO}_3$ .



- (i) Define a buffer solution.

.....

.....

- (ii) Construct **two** equations to show how equilibrium 5 acts as a buffer solution.

.....

.....

- (iii) The  $[\text{HCO}_3^-]/[\text{H}_2\text{CO}_3]$  ratio in a sample of seawater is 14.1.

Calculate the pH of this sample.

[ $pK_a$ :  $\text{H}_2\text{CO}_3$ , 6.35]

pH = ..... [3]

[Total: 14]