# How Far? - Entropy

## Question Paper

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|---|---|
| **Level** | International A Level |
| **Subject** | Chemistry |
| **Exam Board** | Edexcel |
| **Topic** | Rates, Equilibria & Further Organic Chemistry |
| **Sub Topic** | How Far? - Entropy |
| **Booklet** | Question Paper |

**Time Allowed:** **48 minutes**

**Score:** **/40**

**Percentage:** **/100**

**Grade Boundaries:**

| A* | A | B | C | D | E | U |
|---|---|---|---|---|---|---|
| >85% | '77.5% | 70% | 62.5% | 57.5% | 45% | <45% |

**1** In which of the following reactions is there a **decrease** in the entropy of the system?

☒ **A** $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

☒ **B** $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(l)$

☒ **C** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

☒ **D** $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

**(Total for Question 1 = 1 mark)**

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**2** Which of the following statements is always true for an exothermic reaction?

☒ **A** $\Delta S_{surroundings}$ doubles when the temperature in kelvin doubles.

☒ **B** $\Delta S_{surroundings}$ doubles when the natural log of the temperature in kelvin, $\ln T$, doubles.

☒ **C** The equilibrium constant, $K$, doubles when $\Delta S_{total}$ doubles.

☒ **D** The natural log of the equilibrium constant, $\ln K$, doubles when $\Delta S_{total}$ doubles.

**(Total for Question 2 = 1 mark)**

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**3** When one mole of magnesium chloride dissolves in water, the enthalpy change, $\Delta H_{solution}$, is more negative than the corresponding change for sodium chloride.

One explanation for this difference is that

☒ **A** the lattice enthalpy for magnesium chloride is more negative than the lattice enthalpy for sodium chloride.

☒ **B** the $\Delta H_{hydration}$ of magnesium ions is more negative than $\Delta H_{hydration}$ of sodium ions.

☒ **C** the $\Delta H_{formation}$ of magnesium chloride is more negative than $\Delta H_{formation}$ of sodium chloride.

☒ **D** magnesium chloride has more covalent character than sodium chloride.

**(Total for Question 3 = 1 mark)**

**4** The table shows some data about metal ions, non-metal ions and their compounds.

| Ion | Enthalpy change of hydration / kJ mol$^{-1}$ | Compound | Lattice energy / kJ mol$^{-1}$ |
|---|---|---|---|
| $Sr^{2+}(g)$ | –1443 | $SrF_2(s)$ | –2492 |
| $F^-(g)$ | –483 | | |
| $Rb^+(g)$ | –297 | RbCl(s) | –685 |
| $Cl^-(g)$ | –340 | | |

Use the data in the following calculations.

(a) What is the standard enthalpy change, in kJ mol$^{-1}$, for the following process?

$$Sr^{2+}(g) + 2F^-(g) \rightarrow Sr^{2+}(aq) + 2F^-(aq)$$

**(1)**

☒ **A** –477

☒ **B** –960

☒ **C** –1926

☒ **D** –2409

(b) What is the standard enthalpy change of solution, in kJ mol$^{-1}$, for rubidium chloride, RbCl?

**(1)**

☒ **A** –1322

☒ **B** –48

☒ **C** +48

☒ **D** +1322

**(Total for Question 4 = 2 marks)**

**5** Which of the following is the correct order of **decreasing** entropy?

| | Highest entropy | Middle entropy | Lowest entropy |
|---|---|---|---|
| ☒ **A** | ice at 0 °C | water at 0 °C | steam at 120 °C |
| ☒ **B** | ice at 0 °C | steam at 120 °C | water at 100 °C |
| ☒ **C** | steam at 100 °C | water at 0 °C | ice at –20 °C |
| ☒ **D** | steam at 100 °C | ice at –20 °C | water at 100 °C |

**(Total for Question 5 = 1 mark)**

**6** Ethanol burns in excess oxygen to produce carbon dioxide and water.

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

The standard molar entropies of the reactants and products at 298 K are given in the table below.

| Substance | $S^{\ominus}$/J mol$^{-1}$ K$^{-1}$ |
|---|---|
| $C_2H_5OH(l)$ | 161 |
| $O_2(g)$ | 205 |
| $CO_2(g)$ | 214 |
| $H_2O(l)$ | 70 |

The value of $\Delta S^{\ominus}_{system}$ for this reaction, in J mol$^{-1}$ K$^{-1}$, is

☒ **A** –138

☒ **B** –82

☒ **C** +82

☒ **D** +138

**(Total for Question 6 = 1 mark)**

**7** Nitrogen reacts with hydrogen to form ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \quad \Delta H = -92.2 \text{ kJ mol}^{-1}$$

What is the value of $\Delta S^{\ominus}_{\text{surroundings}}$, in J mol$^{-1}$ K$^{-1}$, for this reaction at 25 °C?

☒ **A** –3688

☒ **B** –309.4

☒ **C** +309.4

☒ **D** +3688

**(Total for Question 7 = 1 mark)**

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**8** The signs of $\Delta H^{\ominus}_{\text{reaction}}$ and $\Delta S^{\ominus}_{\text{system}}$ for four different gaseous reactions are shown in the table. Which reaction must be thermodynamically feasible at **all** temperatures?

| | $\Delta H^{\ominus}_{\text{reaction}}$ | $\Delta S^{\ominus}_{\text{system}}$ |
|---|---|---|
| ☒ **A** | negative | negative |
| ☒ **B** | negative | positive |
| ☒ **C** | positive | negative |
| ☒ **D** | positive | positive |

**(Total for Question 8 = 1 mark)**

**9** Use the data in the table to calculate the enthalpy change of solution of calcium chloride, $CaCl_2$.

| | |
|---|---|
| Lattice energy of calcium chloride | –2258 kJ $mol^{-1}$ |
| Hydration enthalpy of $Ca^{2+}$ | –1650 kJ $mol^{-1}$ |
| Hydration enthalpy of $Cl^-$ | –364 kJ $mol^{-1}$ |

The enthalpy change of solution of calcium chloride, in kJ $mol^{-1}$, is

☒ **A** –244

☒ **B** –120

☒ **C** +120

☒ **D** +244

**(Total for Question 9 = 1 mark)**

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**10** The expression that relates $\Delta S_{system}$ to the equilibrium constant, $K$, for a reaction is

☒ **A** $\Delta S_{system} = R\ln K + \Delta H/T$

☒ **B** $\Delta S_{system} = R\ln K - \Delta H/T$

☒ **C** $\Delta S_{system} = T\ln K + \Delta H/R$

☒ **D** $\Delta S_{system} = T\ln K - \Delta H/R$

**(Total for Question 10 = 1 mark)**

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**11** Energy is given out when one mole of gaseous strontium ions is hydrated.

$$Sr^{2+}(g) + aq \rightarrow Sr^{2+}(aq)$$

This reaction is less exothermic than the corresponding reaction for magnesium ions, $Mg^{2+}(g)$, because

☒ **A** the sum of the first two ionization energies of magnesium is more than that of strontium.

☒ **B** the lattice energies of magnesium compounds are more exothermic than the lattice energies of corresponding strontium compounds.

☒ **C** the solubility of magnesium hydroxide is less than the solubility of strontium hydroxide.

☒ **D** the ionic radius of $Mg^{2+}$ is less than the ionic radius of $Sr^{2+}$.

**(Total for Question 11 = 1 mark)**

**12** Use the data below to calculate the enthalpy change of solution of magnesium chloride.

| | |
|---|---|
| Lattice energy of magnesium chloride | –2526 kJ $mol^{-1}$ |
| Enthalpy of hydration of $Mg^{2+}(g)$ | –2003 kJ $mol^{-1}$ |
| Enthalpy of hydration of $Cl^{-}(g)$ | –340 kJ $mol^{-1}$ |

The enthalpy change of solution of magnesium chloride, in kJ $mol^{-1}$, is

☒ **A** +183

☒ **B** +157

☒ **C** –157

☒ **D** –183

**(Total for Question 12 = 1 mark)**

**13** Nitrogen monoxide, NO, reacts with oxygen as shown below.

$$NO(g) + \tfrac{1}{2}O_2(g) \rightarrow NO_2(g) \qquad \Delta H^{\ominus} = -57 \text{ kJ mol}^{-1}$$

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

The standard molar entropy of $\tfrac{1}{2}O_2(g)$ is 102.5 J $mol^{-1}$ $K^{-1}$.

Use other standard molar entropy values from your Data Booklet.

Include a sign and units in your answer.

**(2)**

(ii) Calculate the entropy change of the surroundings, $\Delta S^{\ominus}_{surroundings}$ at 298 K and hence the total entropy change, $\Delta S^{\ominus}_{total}$ at this temperature.

Include a sign and units in your answers.

**(2)**

(iii) Calculate the temperature at which the reaction ceases to be spontaneous.

**(2)**

(b) Nitrogen monoxide and carbon monoxide are formed in car engines. To prevent these pollutant gases being released into the atmosphere, car exhausts are fitted with a catalyst and the reaction below occurs.

$$CO(g) + NO(g) \rightarrow CO_2(g) + \tfrac{1}{2}N_2(g)$$

At the temperature of the car exhaust, $\Delta S_{total}$ for this reaction is positive.

Suggest why this reaction needs a catalyst.

**(1)**

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**(Total for Question 13 = 7 marks)**

**14** This question is about sucrose, the chemical commonly known as sugar. Some thermochemical data for sucrose and oxygen are given in the table below.

| | |
|---|---|
| Standard entropy of sucrose, $S^{\ominus}$ [$C_{12}H_{22}O_{11}(s)$] | +392.4 J $mol^{-1}$ $K^{-1}$ |
| Standard enthalpy change of combustion of sucrose, $\Delta H_c^{\ominus}$ | –5639.7 kJ $mol^{-1}$ |
| Standard entropy of oxygen, $S^{\ominus}$ [$½O_2(g)$] | +102.5 J $mol^{-1}$ $K^{-1}$ |

The equation for the complete combustion of sucrose, $C_{12}H_{22}O_{11}$, is

$$C_{12}H_{22}O_{11}(s) + 12O_2(g) \rightarrow 12CO_2(g) + 11H_2O(l)$$

(a) (i) Calculate the standard entropy change of the system, $\Delta S^{\ominus}_{system}$, for this combustion, using the data given in the table and your Data Booklet. Include a sign and units in your answer.

**(3)**

(ii) Calculate the standard entropy change of the surroundings, $\Delta S^{\ominus}_{surroundings}$, for this combustion at 298 K. Include a sign and units in your answer.

**(2)**

(iii) Calculate the total standard entropy change for the combustion, $\Delta S^{\ominus}_{total}$, at 298 K.

State the significance of your answer.

**(2)**

(iv) State and explain the effect, if any, of increasing the temperature on $\Delta S^{\ominus}_{surroundings}$, $\Delta S^{\ominus}_{total}$ and the extent of the reaction.

**(3)**

(v) Icing sugar can be hazardous when it is being finely powdered in a factory.

Explain why sucrose is stable at room temperature, in spite of your answer to part (iii), but its manufacture is hazardous.

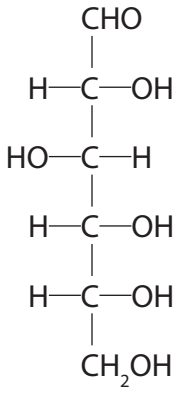
**(2)**

(vi) Suggest **two** risks associated with high levels of sucrose in the diet.

**(2)**

(b) Sucrose can be hydrolysed by warming with dilute hydrochloric acid to form glucose and fructose.

In aqueous solution, a structure of glucose is




(i) Circle or mark with an asterisk (*) all the chiral centres on the structure of glucose.

**(2)**

(ii) State the physical property associated with molecules which have chiral centres.

**(1)**

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(iii) State what change you would expect to see when glucose is boiled with Benedict's or Fehling's solutions.

Explain the chemistry involved in this reaction.

**(3)**

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**(Total for Question 14 = 20 marks)**