IB · HL · Chemistry

60 mins 3 questions

Practice Paper 1B

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Total Marks **/35**

1 (a) Every winter, road salt is spread on icy surfaces to lower the freezing point of water. A student investigates how the mass of calcium chloride ($CaCl_2$) added to water affects the freezing point of the solution.

In each trial, a measured mass of solid $CaCl_2$ is added to 100 g of distilled water. The solution is stirred and cooled, and the freezing point is recorded.

Write an ionic equation to show how $CaCl_2$ dissociates in water.

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(1 mark)

(b) Explain how the dissociation of $CaCl_2$ helps to lower the freezing point of water.

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(1 mark)

(c) The table below shows the freezing points recorded during the investigation.

Mass of $CaCl_2$ added / g	**Freezing point / ^{o}C**
0.0	0.0
5.0	–2.6
10.0	–5.1
15.0	–7.5
20.0	–9.8
25.0	–11.6

(i) Describe the trend shown in the data.

[1]

(ii) Calculate the average decrease in freezing point per gram of $CaCl_2$ added.

[2]

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..... **(3 marks)**

(d) Estimate the freezing point if 30.0 g of $CaCl_2$ were added.

..... **(1 mark)**

(e) Explain one reason why your estimate in (d) may not be completely accurate.

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..... **(2 marks)**

(f) The student later discovers the $CaCl_2$ used was slightly damp and not pure.

Explain how this would affect the results of the experiment.

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..... **(2 marks)**

(g) A second student suggests:

"The bigger the molar mass of the salt, the more it lowers the freezing point."

Evaluate this suggestion.

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..... **(2 marks)**

2 (a) A student investigates the effectiveness of four commercial antacids in reducing stomach acidity. The active ingredients of the antacids are shown below:

Brand	Active ingredients
A	$Mg(OH)_2$, $Al(OH)_3$
B	$NaHCO_3$, $CaCO_3$
C	$CaCO_3$
D	$Mg(OH)_2$, Al_2O_3

Each tablet is crushed and added to 25.0 cm^3 of 1.00 mol dm^{-3} hydrochloric acid (HCl). After 5 minutes, the final pH is recorded.

(i) Write a balanced chemical equation for the reaction between calcium carbonate and hydrochloric acid.

[1]

(ii) State the formula of one ion produced in the reaction that causes the pH to increase.

[1]

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(2 marks)

(b) Suggest two experimental variables, other than time, that should be controlled to ensure a fair comparison between the antacid tablets.

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(2 marks)

(c) The results of the experiment are shown below.

Antacid	Dose used	Final pH
A	1 tablet (0.80 g)	3.42
B	1 tablet (1.25 g)	5.01
C	0.5 tablet (0.75 g)	3.48
D	1 tablet (0.95 g)	2.12

(i) Assuming that the initial pH of the acid was 1.00, calculate the change in pH for antacid A.

[1]

(ii) Calculate the uncertainty in the pH change for antacid A, using ±0.02 for each pH value.

[1]

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..... **(2 marks)**

(d) Explain one reason why antacid B may appear more effective than C, even though both contain calcium carbonate.

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..... **(2 marks)**

(e) The student concludes that "Antacid B is the most effective."

Use the data to evaluate this conclusion.

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..... **(3 marks)**

(f) The student later discovers the antacid B tablet was slightly damp.

Explain how this might affect the result.

..... **(1 mark)**

(g) Suggest one environmental concern with using excess antacid tablets that contain carbonate or hydroxide compounds.

..... **(1 mark)**

3 (a) A student investigates how temperature affects the position of equilibrium in the following reversible reaction:

$$CO\ (g) + H_2O\ (g) \rightleftharpoons CO_2\ (g) + H_2\ (g)$$

$$\Delta H^{\ominus} = -41.2\ kJ\ mol^{-1}$$

The student collects data at different temperatures and uses the equilibrium concentrations to calculate the equilibrium constant, *K*, for each trial.

The table below shows her results:

Temperature / K	**ln *K***
500	1.24
550	0.85
600	0.51
650	0.23
700	–0.05

The student notes that the value of ln K decreases as temperature increases.

Using the data, describe how the position of equilibrium changes with increasing temperature. Support your answer with reference to the values of *K*.

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(1 mark)

(b) Explain this effect using Le Châtelier's principle and the sign of $\Delta H^{\ominus}$.

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(2 marks)

(c) Use the data to calculate the standard Gibbs free energy change ($\Delta G^{\ominus}$) at 600 K.

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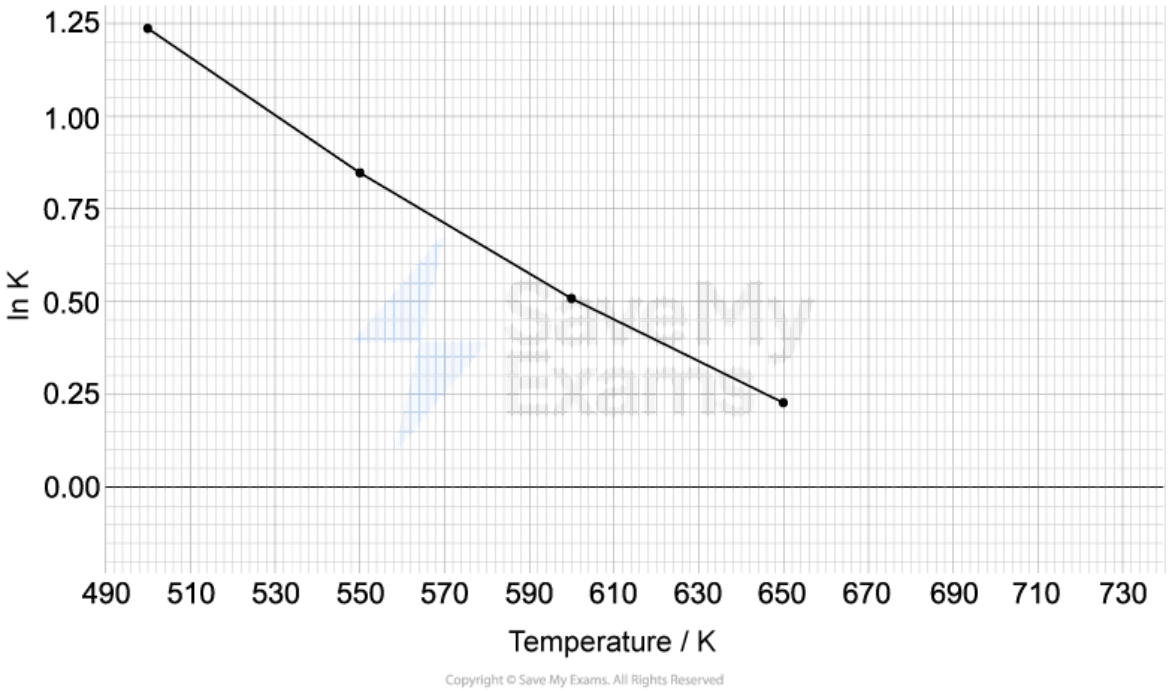
(2 marks)

(d) The student then plots a graph of ln *K* against temperature (T) using the data.



Use the graph to estimate the temperature at which $\Delta G^{\ominus}$ for the reaction is zero. Justify your answer.

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(2 marks)

(e) Evaluate whether this reaction would be thermodynamically feasible at 750 K.

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(3 marks)