AP® CHEMISTRY
# 2010 SCORING GUIDELINES

Question 2
(10 points)

A student performs an experiment to determine the molar enthalpy of solution of urea, H_2NCONH_2. The student places 91.95 g of water at 25°C into a coffee cup calorimeter and immerses a thermometer in the water. After 50 s, the student adds 5.13 g of solid urea, also at 25°C, to the water and measures the temperature of the solution as the urea dissolves. A plot of the temperature data is shown in the graph below.
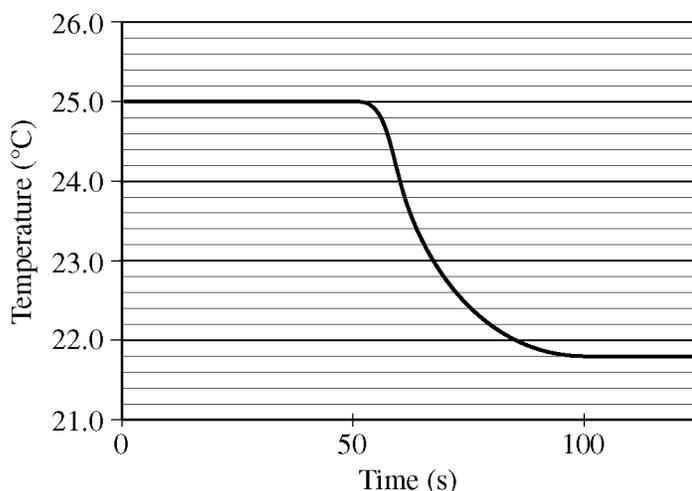
(a) Determine the change in temperature of the solution that results from the dissolution of the urea.

$\Delta T = 21.8 - 25.0 = -3.2$ Celsius degrees	One point is earned for the correct temperature change.

(b) According to the data, is the dissolution of urea in water an endothermic process or an exothermic process? Justify your answer.

The process is endothermic. The decrease in temperature indicates that the process for the dissolution of urea in water requires energy.	One point is earned for the correct choice with justification.

(c) Assume that the specific heat capacity of the calorimeter is negligible and that the specific heat capacity of the solution of urea and water is 4.2 J g^{-1} °C^{-1} throughout the experiment.

(i) Calculate the heat of dissolution of the urea in joules.

Assuming that no heat energy is lost from the calorimeter and given that the calorimeter has a negligible heat capacity, the sum of the heat of dissolution, q_{soln} and the change in heat energy of the urea-water mixture must equal zero. $$q_{soln} + mc\Delta T = 0 \quad \Rightarrow \quad q_{soln} = -mc\Delta T$$ $$m_{soln} = 5.13 \text{ g} + 91.95 \text{ g} = 97.08 \text{ g}$$ $$q_{soln} = -(97.08 \text{ g})(4.2 \text{ J g}^{-1}°C^{-1})(-3.2°C) = 1.3 \times 10^3 \text{ J}$$	One point is earned for the correct setup. One point is earned for the correct numerical result for the heat of dissolution.

(ii) Calculate the molar enthalpy of solution, ΔH_{soln}°, of urea in kJ mol^{-1}.

$$\Delta H_{soln}^\circ = \frac{q_{soln}}{\text{mol solute}}$$ molar mass of urea $= 4(1.0) + 2(14.0) + 12.0 + 16.0 = 60.0$ g mol^{-1} moles of urea $= 5.13$ g urea $\times \dfrac{1 \text{ mol urea}}{60.0 \text{ g urea}} = 0.0855$ mol $$\Delta H_{soln}^\circ = \frac{1.3 \times 10^3 \text{ J}}{0.0855 \text{ mol}} = 1.5 \times 10^4 \text{ J mol}^{-1} = 15 \text{ kJ mol}^{-1}$$	One point is earned for the calculation of moles of urea. One point is earned for the correct numerical result with correct algebraic sign.

(d) Using the information in the table below, calculate the value of the molar entropy of solution, ΔS_{soln}°, of urea at 298 K. Include units with your answer.

	Accepted Value
ΔH_{soln}° of urea	14.0 kJ mol^{-1}
ΔG_{soln}° of urea	−6.9 kJ mol^{-1}

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$ $$-6.9 \text{ kJ mol}^{-1} = 14.0 \text{ kJ mol}^{-1} - (298 \text{ K})(\Delta S^\circ)$$ $$\Delta S_{soln}^\circ = 0.0701 \text{ kJ mol}^{-1} \text{ K}^{-1} = 70.1 \text{ J mol}^{-1} \text{ K}^{-1}$$	One point is earned for the correct setup. One point is earned for the correct numerical result with correct units.

(e) The student repeats the experiment and this time obtains a result for ΔH°_{soln} of urea that is 11 percent below the accepted value. Calculate the value of ΔH°_{soln} that the student obtained in this second trial.

Error $= (0.11)(14.0 \text{ kJ mol}^{-1}) = 1.54 \text{ kJ mol}^{-1}$ $14.0 \text{ kJ mol}^{-1} - 1.54 \text{ kJ mol}^{-1} = 12.5 \text{ kJ mol}^{-1}$	One point is earned for the correct numerical result.

(f) The student performs a third trial of the experiment but this time adds urea that has been taken directly from a refrigerator at 5°C. What effect, if any, would using the cold urea instead of urea at 25°C have on the experimentally obtained value of ΔH°_{soln} ? Justify your answer.

There would be an increase in the obtained value for ΔH°_{soln} because the colder urea would have caused a larger negative temperature change.	One point is earned for the correct prediction with justification.

2. a) $\Delta T = 21.8 - 25.0$ -25.0

       $= -3.2°C$ $\frac{+21.8}{-3.2}$

b) If the temperature decreased, that means that the reactants absorbed energy from their surroundings, so the process is endothermic.

c) i) $q = mc\Delta T$

$= (91.95 + 5.13\ g)\left(\frac{4.2\ J}{g\cdot°C}\right)(3.2°C)$

$= 1300\ J$

ii) $\Delta H°_{soln} = \dfrac{q}{mol\ urea}$ $mol\ urea = \dfrac{1\ mol}{60.06\ g} \times 5.13\ g$

$= \dfrac{1300\ J}{0.0854\ mol}$ $= 0.0854\ mol$

$= 15000\ J/mol \rightarrow 15\ kJ/mol$

d) $\Delta G° = \Delta H° - T\Delta S°$

$-6.9\ kJ\ mol^{-1} = 14.0\ kJ\ mol^{-1} - 298K\cdot\Delta S°$

$298\Delta S° = 14.0 + 6.9$

$298\Delta S° = 20.9\ kJ\ mol^{-1}$

$\Delta S° = 0.0701\ kJ\ mol^{-1}K^{-1}$

e) $100 - 11 = 89.0\%$

$\Delta H°_{soln}(trial) = 0.890 \times 14.0\ kJ/mol$

$= 12.5\ kJ/mol$

f) Using cold urea instead of urea at 25°C would increase the experimentally obtained value of $\Delta H°_{soln}$, because the change in temperature for urea is also greater. Because $q = mc\Delta T$, if ΔT incr. so does q, as well as $\Delta H°_{soln}$.

2 B

a) $\Delta T = T_i - T_f$
$\Delta T = 25.0°C - 21.8°C$
$\Delta T = 3.2°C$

b) Endothermic, b/c as the reaction progressed the T dropped, meaning that the reaction absorbed heat from its surroundings

c) i. $\Delta H = (m)(\Delta T)(\text{specific heat})$
$\Delta H = (91.95)(3.2°C)(4.2 J g^{-1} °C^{-1})$
$\Delta H = 1200 J$

ii. $\Delta H°_{soln} = 1.2 kJ / (.0854 mol)$ $M = 60.062$
$= 14. kJ mol^{-1}$ $5.13g / 60.062g = .0854 mol$

d) $-\Delta G = \Delta H - \Delta S(T)$
$6.9 kJ mol^{-1} = 14.0 kJ mol^{-1} - \Delta S (298)$
$\Delta S = .024 kJ mol^{-1} K^{-1}$

e) $\dfrac{14.0 kJ mol^{-1} - x}{14.0 kJ} = .11$
$x = 12.5 kJ mol^{-1}$

f) The cold area would increase the experimentally obtained value of $\Delta H°_{soln}$ b/c ΔT would appear larger as is cooled the sol'n further.

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ADDITIONAL PAGE FOR ANSWERING QUESTION 2

(a) $\Delta T = T_F - T_i$

$= 21.8 - 25$

$= -3.2°C$

(b) The dissolution of urea in water is an exothermic process because ΔT is negative and energy is therefore released.

(c) i. $4.2 \frac{J}{g \cdot c} \times 5.13g \times -3.2°C = -68.9 J$

ii. $\Delta H^\circ_{soln} = 4.2 \frac{J}{g \cdot c} \times 60.08 g \ldots$

(d) $\Delta G = \Delta H - T\Delta S$

$-6.9 \frac{kJ}{mol} = 14.0 \frac{kJ}{mol} - 298(K)\Delta S$

$\Delta S = 0.0701 \ kJ = 70.1 \ J/mol \cdot K$

(e) $\Delta H^\circ_{soln} = 14.0 \ kJ$ (accepted value) $\times .11 = 1.54$

$14.0 - 1.54 = 12.46 \ kJ/mol$

(f) The ΔH_{soln} would decrease. Because the urea would be at a lower temperature, some heat would be required to raise the heat of urea so less heat would be released.

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Question 2

Overview

This question assessed students' knowledge and skills pertaining to a laboratory experience that involved preparing a solution in a coffee-cup calorimeter and measuring the results. Parts of the question were mathematical and parts were conceptual in nature. In part (a) students were required to examine a temperature–time graph and use graph-reading skills to determine the temperature change that occurred as a solution of urea and water formed. Part (b) required classification of the process as either endothermic or exothermic based on the change in temperature. In parts (c)(i) and (c)(ii) students had to calculate the heat of dissolution of urea with the equation $q = mc\Delta T$ and then convert that calculated energy into a molar enthalpy. Part (d) required students to calculate a molar entropy of solution with the equation $\Delta G = \Delta H - T\Delta S$ given a table of accepted values for the $\Delta H°_{soln}$ and $\Delta G°_{soln}$ of the solute (urea). Part (e) supplied a percent error from a student experiment and required calculation of the experimental value obtained from that experiment. Part (f) involved an error analysis whereby students needed to predict the effect on $\Delta H°_{soln}$ of using a sample of solute at a colder temperature and then justify their prediction.

Sample: 2A
Score: 10

This response earned all 10 points: 1 point for part (a), 1 point for part (b), 2 points for part (c)(i), 2 points for part (c)(ii), 2 points for part (d), 1 point for part (e), and 1 point for part (f).

Sample: 2B
Score: 8

In part (c)(i) this response earned 1 point for the setup by substituting a mass, a specific heat, and the ΔT determined in part (a) into the equation $q = mc\Delta T$. The second point was not earned because the substituted mass represents the mass of the solvent and does not include the mass of the solute, thereby resulting in an incorrect number of joules. In part (d) the response earned 1 point for the setup by substituting the correct values from the table into the equation $\Delta G = \Delta H - T\Delta S$, but the second point was not earned because the answer calculated is incorrect.

Sample: 2C
Score: 5

The point was not earned in part (b) of this response because the process is incorrectly identified as exothermic. In part (c)(i) the response earned 1 point for the setup by substituting a mass, a specific heat, and the ΔT determined in part (a) into the equation $q = mc\Delta T$. The second point was not earned because the substituted mass represents the mass of the solute and does not include the mass of the solvent, thereby resulting in an incorrect number of joules. The points were not earned in part (c)(ii) because the student does not calculate the number of moles of urea and divide the energy calculated in part (c)(i) by this amount. The point was not earned in part (f) because the prediction of the direction of change in the value of $\Delta H°_{soln}$ is incorrect.