# 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.
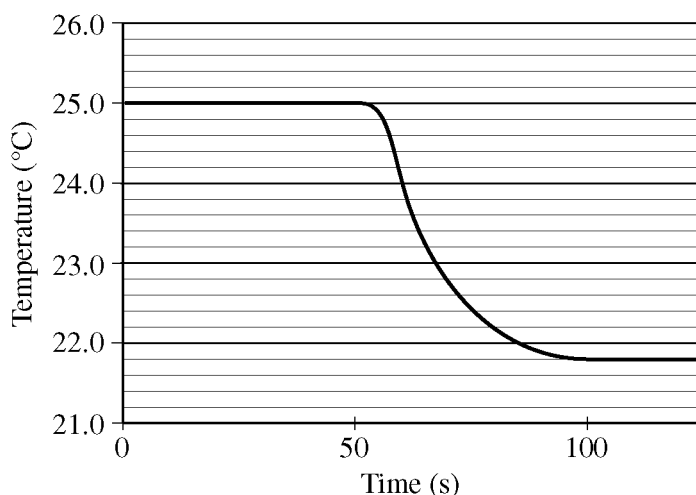
(a) Determine the change in temperature of the solution that results from the dissolution of the urea.

$\Delta T = 21.8 - 25.0 = -3.2$ 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. $q_{soln} + mc\Delta T = 0 \Rightarrow q_{soln} = -mc\Delta T$ $m_{soln} = 5.13\text{ g} + 91.95\text{ g} = 97.08\text{ g}$ $q_{soln} = -(97.08\text{ g})(4.2\text{ J g}^{-1}°C^{-1})(-3.2°C) = 1.3 \times 10^3\text{ J}$	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}.

$\Delta H^\circ_{soln} = \dfrac{q_{soln}}{\text{mol solute}}$ molar mass of urea $= 4(1.0) + 2(14.0) + 12.0 + 16.0 = 60.0\text{ g mol}^{-1}$ moles of urea $= 5.13\text{ g urea} \times \dfrac{1\text{ mol urea}}{60.0\text{ g urea}} = 0.0855\text{ mol}$ $\Delta H^\circ_{soln} = \dfrac{1.3 \times 10^3\text{ J}}{0.0855\text{ mol}} = 1.5 \times 10^4\text{ J mol}^{-1} = 15\text{ kJ mol}^{-1}$	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}

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ $-6.9\text{ kJ mol}^{-1} = 14.0\text{ kJ mol}^{-1} - (298\text{ K})(\Delta S^\circ)$ $\Delta S^\circ_{soln} = 0.0701\text{ kJ mol}^{-1}\text{ K}^{-1} = 70.1\text{ J mol}^{-1}\text{ K}^{-1}$	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.