Answer EITHER Question 2 below OR Question 3 printed on page 8. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

$$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$$

2. The combustion of carbon monoxide is represented by the equation above.

(a) Determine the value of the standard enthalpy change, ΔH°_{rxn}, for the combustion of $CO(g)$ at 298 K using the following information.

$$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g) \qquad \Delta H^\circ_{298} = -110.5 \text{ kJ mol}^{-1}$$

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^\circ_{298} = -393.5 \text{ kJ mol}^{-1}$$

(b) Determine the value of the standard entropy change, ΔS°_{rxn}, for the combustion of $CO(g)$ at 298 K using the information in the following table.

Substance	S°_{298} ($\text{J mol}^{-1}\text{ K}^{-1}$)
$CO(g)$	197.7
$CO_2(g)$	213.7
$O_2(g)$	205.1

(c) Determine the standard free energy change, ΔG°_{rxn}, for the reaction at 298 K. Include units with your answer.

(d) Is the reaction spontaneous under standard conditions at 298 K ? Justify your answer.

(e) Calculate the value of the equilibrium constant, K_{eq}, for the reaction at 298 K.

Answer EITHER Question 2 OR Question 3 below. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.

2. Answer the following questions about voltaic cells.

   (a) A voltaic cell is set up using Al/Al^{3+} as one half-cell and Sn/Sn^{2+} as the other half-cell. The half-cells contain equal volumes of solutions and are at standard conditions.

      (i) Write the balanced net-ionic equation for the spontaneous cell reaction.

      (ii) Determine the value, in volts, of the standard potential, $E°$, for the spontaneous cell reaction.

      (iii) Calculate the value of the standard free-energy change, $\Delta G°$, for the spontaneous cell reaction. Include units with your answer.

      (iv) If the cell operates until $[Al^{3+}]$ is 1.08 M in the Al/Al^{3+} half-cell, what is $[Sn^{2+}]$ in the Sn/Sn^{2+} half-cell?

   (b) In another voltaic cell with Al/Al^{3+} and Sn/Sn^{2+} half-cells, $[Sn^{2+}]$ is 0.010 M and $[Al^{3+}]$ is 1.00 M. Calculate the value, in volts, of the cell potential, E_{cell}, at 25°C.

3. Answer the following questions about the thermodynamics of the reactions represented below.

Reaction X: $\frac{1}{2}I_2(s) + \frac{1}{2}Cl_2(g) \rightleftarrows ICl(g)$ $\quad \Delta H_f^\circ = 18 \text{ kJ mol}^{-1}$, $\Delta S_{298}^\circ = 78 \text{ J K}^{-1} \text{ mol}^{-1}$

Reaction Y: $\frac{1}{2}I_2(s) + \frac{1}{2}Br_2(l) \rightleftarrows IBr(g)$ $\quad \Delta H_f^\circ = 41 \text{ kJ mol}^{-1}$, $\Delta S_{298}^\circ = 124 \text{ J K}^{-1} \text{ mol}^{-1}$

   (a) Is reaction X, represented above, spontaneous under standard conditions? Justify your answer with a calculation.

   (b) Calculate the value of the equilibrium constant, K_{eq}, for reaction X at 25°C.

   (c) What effect will an increase in temperature have on the equilibrium constant for reaction X? Explain your answer.

   (d) Explain why the standard entropy change is greater for reaction Y than for reaction X.

   (e) Above what temperature will the value of the equilibrium constant for reaction Y be greater than 1.0? Justify your answer with calculations.

   (f) For the vaporization of solid iodine, $I_2(s) \rightarrow I_2(g)$, the value of ΔH_{298}° is 62 kJ mol^{-1}. Using this information, calculate the value of ΔH_{298}° for the reaction represented below.

$$I_2(g) + Cl_2(g) \rightleftarrows 2\,ICl(g)$$

STOP

If you finish before time is called, you may check your work on this part only.
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$$N_2(g) + 3\,F_2(g) \rightarrow 2\,NF_3(g) \qquad \Delta H^\circ_{298} = -264\ \text{kJ mol}^{-1};\ \Delta S^\circ_{298} = -278\ \text{J K}^{-1}\ \text{mol}^{-1}$$

2. The following questions relate to the synthesis reaction represented by the chemical equation in the box above.

(a) Calculate the value of the standard free energy change, ΔG°_{298}, for the reaction.

(b) Determine the temperature at which the equilibrium constant, K_{eq}, for the reaction is equal to 1.00. (Assume that ΔH° and ΔS° are independent of temperature.)

(c) Calculate the standard enthalpy change, ΔH°, that occurs when a 0.256 mol sample of $NF_3(g)$ is formed from $N_2(g)$ and $F_2(g)$ at 1.00 atm and 298 K.

The enthalpy change in a chemical reaction is the difference between energy absorbed in breaking bonds in the reactants and energy released by bond formation in the products.

(d) How many bonds are formed when two molecules of NF_3 are produced according to the equation in the box above?

(e) Use both the information in the box above and the table of average bond enthalpies below to calculate the average enthalpy of the F – F bond.

Bond	Average Bond Enthalpy (kJ mol^{-1})
N≡N	946
N–F	272
F–F	?

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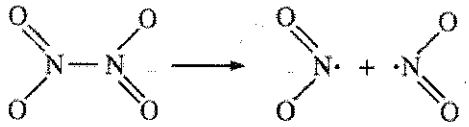
6. Use principles of thermodynamics to answer the following questions.

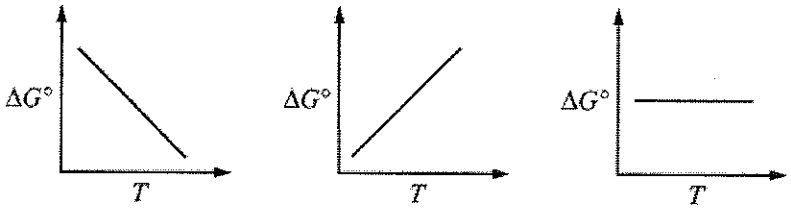
(a) The gas N_2O_4 decomposes to form the gas NO_2 according to the equation below.



(i) Predict the sign of ΔH° for the reaction. Justify your answer.

(ii) Predict the sign of ΔS° for the reaction. Justify your answer.

(b) One of the diagrams below best represents the relationship between ΔG° and temperature for the reaction given in part (a). Assume that ΔH° and ΔS° are independent of temperature.



Draw a circle around the correct graph. Explain why you chose that graph in terms of the relationship $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$.

(c) A reaction mixture of N_2O_4 and NO_2 is at equilibrium. Heat is added to the mixture while the mixture is maintained at constant pressure.

(i) Explain why the concentration of N_2O_4 decreases.

(ii) The value of K_{eq} at 25°C is 5.0×10^{-3}. Will the value of K_{eq} at 100°C be greater than, less than, or equal to this value?

(d) Using the value of K_{eq} at 25°C given in part (c)(ii), predict whether the value of ΔH° is expected to be greater than, less than, or equal to the value of $T\Delta S^\circ$. Explain.

STOP

END OF EXAM

Answer Question 5 and Question 6. The Section II score weighting for these questions is 15 percent each.

Your responses to these questions will be graded on the basis of the accuracy and relevance of the information cited. Explanations should be clear and well organized. Examples and equations may be included in your responses where appropriate. Specific answers are preferable to broad, diffuse responses.

Reaction	Equation	ΔH°_{298}	ΔS°_{298}	ΔG°_{298}
X	$C(s) + H_2O(g) \rightleftarrows CO(g) + H_2(g)$	+131 kJ mol^{-1}	+134 J mol^{-1} K^{-1}	+91 kJ mol^{-1}
Y	$CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$	+41 kJ mol^{-1}	+42 J mol^{-1} K^{-1}	+29 kJ mol^{-1}
Z	$2\,CO(g) \rightleftarrows C(s) + CO_2(g)$	?	?	?

5. Answer the following questions using the information related to reactions X, Y, and Z in the table above.

   (a) For reaction X, write the expression for the equilibrium constant, K_p.

   (b) For reaction X, will the equilibrium constant, K_p, increase, decrease, or remain the same if the temperature rises above 298 K? Justify your answer.

   (c) For reaction Y at 298 K, is the value of K_p greater than 1, less than 1, or equal to 1? Justify your answer.

   (d) For reaction Y at 298 K, which is larger: the total bond energy of the reactants or the total bond energy of the products? Explain.

   (e) Is the following statement true or false? Justify your answer.

   "On the basis of the data in the table, it can be predicted that reaction Y will occur more rapidly than reaction X will occur."

   (f) Consider reaction Z at 298 K.

      (i) Is ΔS° for the reaction positive, negative, or zero? Justify your answer.

      (ii) Determine the value of ΔH° for the reaction.

      (iii) A sealed glass reaction vessel contains only $CO(g)$ and a small amount of $C(s)$. If a reaction occurs and the temperature is held constant at 298 K, will the pressure in the reaction vessel increase, decrease, or remain the same over time? Explain.

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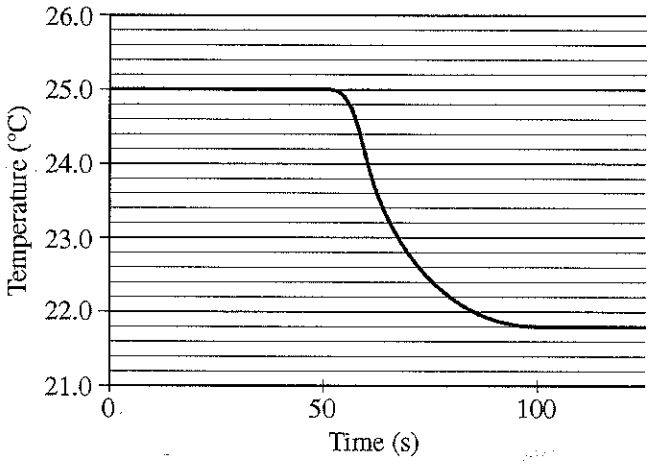
2. A student performs an experiment to determine the molar enthalpy of solution of urea, H_2NCONH_2. The student places 91.95 g of water at 25°C into a coffee-cup calorimeter and immerses a thermometer in the water. After 50 s, the student adds 5.13 g of solid urea, also at 25°C, to the water and measures 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.

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

(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.

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

(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}

(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.

(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.

3. Hydrogen gas burns in air according to the equation below.

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

(a) Calculate the standard enthalpy change, ΔH°_{298}, for the reaction represented by the equation above. (The molar enthalpy of formation, ΔH°_f, for $H_2O(l)$ is -285.8 kJ mol^{-1} at 298 K.)

(b) Calculate the amount of heat, in kJ, that is released when 10.0 g of $H_2(g)$ is burned in air.

(c) Given that the molar enthalpy of vaporization, ΔH°_{vap}, for $H_2O(l)$ is 44.0 kJ mol^{-1} at 298 K, what is the standard enthalpy change, ΔH°_{298}, for the reaction $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$?

A fuel cell is an electrochemical cell that converts the chemical energy stored in a fuel into electrical energy. A cell that uses H_2 as the fuel can be constructed based on the following half-reactions.

Half-reaction	E° (298 K)
$2\ H_2O(l) + O_2(g) + 4\ e^- \rightarrow 4\ OH^-(aq)$	0.40 V
$2\ H_2O(l) + 2\ e^- \rightarrow H_2(g) + 2\ OH^-(aq)$	−0.83 V

(d) Write the equation for the overall cell reaction.

(e) Calculate the standard potential for the cell at 298 K.

(f) Assume that 0.93 mol of $H_2(g)$ is consumed as the cell operates for 600. seconds.

(i) Calculate the number of moles of electrons that pass through the cell.

(ii) Calculate the average current, in amperes, that passes through the cell.

(g) Some fuel cells use butane gas, C_4H_{10}, rather than hydrogen gas. The overall reaction that occurs in a butane fuel cell is $2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(l)$. What is one environmental advantage of using fuel cells that are based on hydrogen rather than on hydrocarbons such as butane?

STOP

If you finish before time is called, you may check your work on this part only.
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3. Answer the following questions about glucose, $C_6H_{12}O_6$, an important biochemical energy source.

(a) Write the empirical formula of glucose.

In many organisms, glucose is oxidized to carbon dioxide and water, as represented by the following equation.

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

A 2.50 g sample of glucose and an excess of $O_2(g)$ were placed in a calorimeter. After the reaction was initiated and proceeded to completion, the total heat released by the reaction was calculated to be 39.0 kJ.

(b) Calculate the value of $\Delta H°$, in kJ mol^{-1}, for the combustion of glucose.

(c) When oxygen is not available, glucose can be oxidized by fermentation. In that process, ethanol and carbon dioxide are produced, as represented by the following equation.

$$C_6H_{12}O_6(s) \rightarrow 2\ C_2H_5OH(l) + 2\ CO_2(g) \qquad \Delta H° = -68.0 \text{ kJ mol}^{-1} \text{ at } 298 \text{ K}$$

The value of the equilibrium constant, K_p, for the reaction at 298 K is 8.9×10^{39}.

(i) Calculate the value of the standard free-energy change, $\Delta G°$, for the reaction at 298 K. Include units with your answer.

(ii) Calculate the value of the standard entropy change, $\Delta S°$, in J K^{-1} mol^{-1}, for the reaction at 298 K.

(iii) Indicate whether the equilibrium constant for the fermentation reaction increases, decreases, or remains the same if the temperature is increased. Justify your answer.

(d) Using your answer for part (b) and the information provided in part (c), calculate the value of $\Delta H°$ for the following reaction.

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

STOP

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3. A sample of $CH_3CH_2NH_2$ is placed in an insulated container, where it decomposes into ethene and ammonia according to the reaction represented above.

Substance	Absolute Entropy, $S°$, in J/(mol·K) at 298 K
$CH_3CH_2NH_2(g)$	284.9
$CH_2CH_2(g)$	219.3
$NH_3(g)$	192.8

(a) Using the data in the table above, calculate the value, in $J/(mol_{rxn} \cdot K)$, of the standard entropy change, $\Delta S°$, for the reaction at 298 K.

(b) Using the data in the table below, calculate the value, in kJ/mol_{rxn}, of the standard enthalpy change, $\Delta H°$, for the reaction at 298 K.

Bond	C–C	C = C	C–H	C–N	N–H
Average Bond Enthalpy (kJ/mol)	348	614	413	293	391

(c) Based on your answer to part (b), predict whether the temperature of the contents of the insulated container will increase, decrease, or remain the same as the reaction proceeds. Justify your prediction.

An experiment is carried out to measure the rate of the reaction, which is first order. A 4.70×10^{-3} mol sample of $CH_3CH_2NH_2$ is placed in a previously evacuated 2.00 L container at 773 K. After 20.0 minutes, the concentration of the $CH_3CH_2NH_2$ is found to be 3.60×10^{-4} mol/L.

(d) Calculate the rate constant for the reaction at 773 K. Include units with your answer.

(e) Calculate the initial rate, in $M\ min^{-1}$, of the reaction at 773 K.

(f) If $\frac{1}{[CH_3CH_2NH_2]}$ is plotted versus time for this reaction, would the plot result in a straight line or would it result in a curve? Explain your reasoning.

STOP

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Answer Question 5 and Question 6. The Section II score weighting for these questions is 15 percent each.

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Process	$\Delta H°$ (kJ/mol$_{rxn}$)
$Br_2(l) \rightarrow Br_2(g)$	30.91
$I_2(s) \rightarrow I_2(g)$	62.44

5. At 298 K and 1 atm, the standard state of Br_2 is a liquid, whereas the standard state of I_2 is a solid. The enthalpy changes for the formation of $Br_2(g)$ and $I_2(g)$ from these elemental forms at 298 K and 1 atm are given in the table above.

   (a) Explain why $\Delta H°$ for the formation of $I_2(g)$ from $I_2(s)$ is larger than $\Delta H°$ for the formation of $Br_2(g)$ from $Br_2(l)$. In your explanation identify the type of particle interactions involved and a reason for the difference in magnitude of those interactions.

   (b) Predict which of the two processes shown in the table has the greater change in entropy. Justify your prediction.

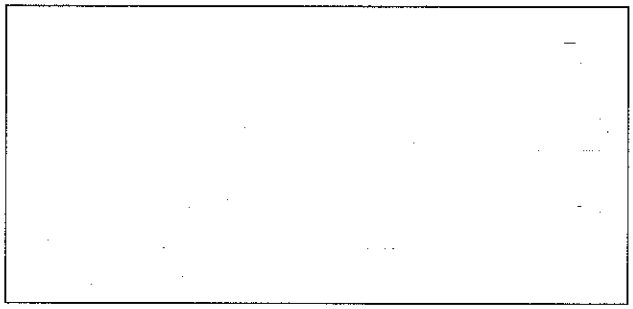
   (c) $I_2(s)$ and $Br_2(l)$ can react to form the compound $IBr(l)$. Predict which would have the greater molar enthalpy of vaporization, $IBr(l)$ or $Br_2(l)$. Justify your prediction.

   An experiment is performed to compare the solubilities of $I_2(s)$ in different solvents, water and hexane (C_6H_{14}). A student adds 2 mL of H_2O and 2 mL of C_6H_{14} to a test tube. Because H_2O and C_6H_{14} are immiscible, two layers are observed in the test tube. The student drops a small, purple crystal of $I_2(s)$ into the test tube, which is then corked and inverted several times. The C_6H_{14} layer becomes light purple, while the H_2O layer remains virtually colorless.

   (d) Explain why the hexane layer is light purple while the water layer is virtually colorless. Your explanation should reference the relative strengths of interactions between molecules of I_2 and the solvents H_2O and C_6H_{14}, and the reasons for the differences.

(e) The student then adds a small crystal of KI(*s*) to the test tube. The test tube is corked and inverted several times. The I^- ion reacts with I_2 to form the I_3^- ion, a linear species.

(i) In the box below, draw the complete Lewis electron-dot diagram for the I_3^- ion.



(ii) In which layer, water or hexane, would the concentration of I_3^- be higher? Explain.