$$O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$$

6. Consider the reaction represented above.

(a) Referring to the data in the table below, calculate the standard enthalpy change, $\Delta H°$, for the reaction at 25°C. Be sure to show your work.

	$O_3(g)$	$NO(g)$	$NO_2(g)$
Standard enthalpy of formation, $\Delta H_f°$, at 25°C (kJ mol^{-1})	143	90.	33

(b) Make a qualitative prediction about the magnitude of the standard entropy change, $\Delta S°$, for the reaction at 25°C. Justify your answer.

(c) On the basis of your answers to parts (a) and (b), predict the sign of the standard free-energy change, $\Delta G°$, for the reaction at 25°C. Explain your reasoning.

(d) Use the information in the table below to write the rate-law expression for the reaction, and explain how you obtained your answer.

Experiment Number	Initial [O_3] (mol L^{-1})	Initial [NO] (mol L^{-1})	Initial Rate of Formation of NO_2 (mol L^{-1} s^{-1})
1	0.0010	0.0010	x
2	0.0010	0.0020	$2x$
3	0.0020	0.0010	$2x$
4	0.0020	0.0020	$4x$

(e) The following three-step mechanism is proposed for the reaction. Identify the step that must be the slowest in order for this mechanism to be consistent with the rate-law expression derived in part (d). Explain.

Step I: $O_3 + NO \rightarrow O + NO_3$

Step II: $O + O_3 \rightarrow 2\,O_2$

Step III: $NO_3 + NO \rightarrow 2\,NO_2$

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.

$$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g) \qquad \Delta H^\circ = -114.1\ kJ,\ \Delta S^\circ = -146.5\ J\ K^{-1}$$

2. The reaction represented above is one that contributes significantly to the formation of photochemical smog.

(a) Calculate the quantity of heat released when 73.1 g of $NO(g)$ is converted to $NO_2(g)$.

(b) For the reaction at 25°C, the value of the standard free-energy change, ΔG°, is −70.4 kJ.

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

(ii) Indicate whether the value of ΔG° would become more negative, less negative, or remain unchanged as the temperature is increased. Justify your answer.

(c) Use the data in the table below to calculate the value of the standard molar entropy, S°, for $O_2(g)$ at 25°C.

	Standard Molar Entropy, S° ($J\ K^{-1}\ mol^{-1}$)
$NO(g)$	210.8
$NO_2(g)$	240.1

(d) Use the data in the table below to calculate the bond energy, in $kJ\ mol^{-1}$, of the nitrogen-oxygen bond in NO_2. Assume that the bonds in the NO_2 molecule are equivalent (i.e., they have the same energy).

	Bond Energy ($kJ\ mol^{-1}$)
Nitrogen-oxygen bond in NO	607
Oxygen-oxygen bond in O_2	495
Nitrogen-oxygen bond in NO_2	?

GO ON TO THE NEXT PAGE.

3. In an experiment, a sample of an unknown, pure gaseous hydrocarbon was analyzed. Results showed that the sample contained 6.000 g of carbon and 1.344 g of hydrogen.

(a) Determine the empirical formula of the hydrocarbon.

(b) The density of the hydrocarbon at 25°C and 1.09 atm is 1.96 g L^{-1}.

(i) Calculate the molar mass of the hydrocarbon.

(ii) Determine the molecular formula of the hydrocarbon.

In another experiment, liquid heptane, $C_7H_{16}(l)$, is completely combusted to produce $CO_2(g)$ and $H_2O(l)$, as represented by the following equation.

$$C_7H_{16}(l) + 11\ O_2(g) \rightarrow 7\ CO_2(g) + 8\ H_2O(l)$$

The heat of combustion, ΔH°_{comb}, for one mole of $C_7H_{16}(l)$ is -4.85×10^3 kJ.

(c) Using the information in the table below, calculate the value of ΔH°_f for $C_7H_{16}(l)$ in kJ mol^{-1}.

Compound	ΔH°_f (kJ mol^{-1})
$CO_2(g)$	–393.5
$H_2O(l)$	–285.8

(d) A 0.0108 mol sample of $C_7H_{16}(l)$ is combusted in a bomb calorimeter.

(i) Calculate the amount of heat released to the calorimeter.

(ii) Given that the total heat capacity of the calorimeter is 9.273 kJ $°C^{-1}$, calculate the temperature change of the calorimeter.

Answer EITHER Question 7 below OR Question 8 printed on page 14. 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 15 percent.

7. Answer the following questions about thermodynamics.

Substance	Combustion Reaction	Enthalpy of Combustion, ΔH°_{comb}, at 298 K (kJ mol^{-1})
$H_2(g)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	–290
$C(s)$	$C(s) + O_2(g) \rightarrow CO_2(g)$	–390
$CH_3OH(l)$		–730

(a) In the empty box in the table above, write a balanced chemical equation for the complete combustion of one mole of $CH_3OH(l)$. Assume products are in their standard states at 298 K. Coefficients do not need to be whole numbers.

(b) On the basis of your answer to part (a) and the information in the table, determine the enthalpy change for the reaction $C(s) + H_2(g) + H_2O(l) \rightarrow CH_3OH(l)$.

(c) Write the balanced chemical equation that shows the reaction that is used to determine the enthalpy of formation for one mole of $CH_3OH(l)$.

(d) Predict the sign of ΔS° for the combustion of $H_2(g)$. Explain your reasoning.

(e) On the basis of bond energies, explain why the combustion of $H_2(g)$ is exothermic.

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} (J mol^{-1} 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.

Based on answer in part (c) will K be less than or greater than 1? Explain

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.
Do not turn to the other part of the test until you are told to do so.

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

If you finish before time is called, you may check your work on this part only.
Do not turn to the other part of the test until you are told to do so.

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

If you finish before time is called, you may check your work on this part only.
Do not turn to the other part of the test until you are told to do so.

7. Aluminum metal can be recycled from scrap metal by melting the metal to evaporate impurities.

☆ (a) Calculate the amount of heat needed to purify 1.00 mole of Al originally at 298 K by melting it. The melting point of Al is 933 K. The molar heat capacity of Al is 24 J/(mol·K), and the heat of fusion of Al is 10.7 kJ/mol.

☆ (b) The equation for the overall process of extracting Al from Al_2O_3 is shown below. Which requires less energy, recycling existing Al or extracting Al from Al_2O_3 ? Justify your answer with a calculation.

$$Al_2O_3(s) \rightarrow 2\,Al(s) + \frac{3}{2}O_2(g) \qquad \Delta H° = 1675\ \text{kJ/mol}_{rxn}$$

STOP

END OF EXAM