AP® Chemistry
2001 Sample Student Responses

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2A

Answer EITHER Question 2 below OR Question 3 printed on page 12. 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	?

a. $\frac{114.1\ kJ}{2\ mol\ NO} \times \frac{1\ mol\ NO}{30.007\ g\ NO} \times 73.1\ g\ NO =$ **139 kJ released**

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ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

2bi $\Delta G^\circ = -RT \ln K$

$\ln K = \frac{-\Delta G^\circ}{RT} = \frac{+70,400\ J}{(8.31\ J/mol\cdot K)(298\ K)} = 28.4$

$K = e^{28.4} = \boxed{2.22 \times 10^{12}}$

ii $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

As T increases, ΔS° becomes more significant.

Because $-T\Delta S^\circ$ is positive and increasing,

ΔG° will increase or become less negative

c. $\Delta S^\circ_{rxn} = 2S^\circ_{NO_2} - 2S^\circ_{NO} - S^\circ_{O_2}$

$-146.5\ J/K = 2mol(240.1\ J/K\cdot mol) - 2mol(210.8\ J/K\cdot mol) - S^\circ_{O_2}$

$S^\circ_{O_2} = 480.2\ J/K - 421.6\ J/K + 146.5\ J/K$

$= \boxed{205.1\ J/K\cdot mol}$

d. $\Delta H^\circ = -4$(bond energy in NO_2) $+ 2$(bond energy in NO) $+$ (bond energy in O_2)

$-114.1\ kJ = -4x + 2(607\ kJ/mol) + 495\ kJ/mol$

x = bond energy of N-O bond in NO_2

$= \frac{1}{4}(1214\ kJ + 495\ kJ + 114.1\ kJ)$

$= \boxed{455.8\ kJ/mol}$

2B

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

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

a) $2NO + O_2 \rightarrow 2NO_2$ $\quad \Delta H^\circ = -114.1\text{ KJ}$

a) 73.1 g NO | 1 mol / 30 g | 2 mol NO_2 / 2 mol NO | −114.1 KJ / 2 mol NO_2 = **−139 KJ / mol**

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ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

b) $\Delta G° = -70.4$ KJ

i) $\Delta G° = -RT\ln K \quad \therefore \ln K = \frac{-\Delta G}{RT}$

$$\ln K = \frac{-(-70.4)\,\text{kJ}}{\text{mol}} \left|\begin{array}{l} \frac{1000\,\text{J}}{1\,\text{kJ}} \\ \end{array}\right. \left|\begin{array}{l} \frac{\text{mol}\cdot\text{K}}{8.31\,\text{J}} \\ \end{array}\right. \left|\begin{array}{l} \\ 298\,\text{K} \end{array}\right. = 28.4$$

$\ln K = 28.4$

$e^{28.4} = \boxed{2.16 \times 10^{12}} = K_{eq}$

ii) Because this is an exothermic reaction (ΔH is negative), energy is a "product" of the reaction. Therefore, according to Lechatlier's principle if the temperature is increased (thus adding more product), the equilibrium will shift to favor the reactants. Therefore, the fwd. rxn. will become less spontaneous and ΔG will become LESS NEGATIVE.

c) $\Delta S° = \Sigma S_p - \Sigma S_R$

$-146.5 = [2(240.1)] - [2(210.8) + x]$

$(-146.5 - 480.2 + 421.6) = -x$

$+205.1 = +x$

$S°$ for O_2 =
$205.1\ \text{J mol}^{-1}\,\text{K}^{-1}$

d) $2(N-O) + (O-O) \rightarrow 2(O-N-O)$

$\Delta H = \Sigma H_{\text{bonds broken}} - \Sigma H_{\text{bonds made}}$

$-114.1 = [2(607) + 1(495)] - 2x$

$-114.1 = 1709 - 2x$

-1709

$\frac{-1823.1}{2} = \frac{-2x}{2} \qquad +x = +912 \qquad x = \boxed{912\ \text{KJ/mol}}$

2C

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73.1 g

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

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

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(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.

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(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	?

(a) $q = mc\Delta t$

$q = (73.1 \times 3.0 \times 10^{8}\ ms^{-1})$

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ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

(b)(i) $K_{eq} = \frac{[NO_2]^2}{[O_2][NO]^2}$

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

$= (-) - (\uparrow T)(-)$ ex

$= (-) - (25)(-) \Rightarrow (-) + 25 = -300 + 25 = -275$

↑ temp = $(-) - (50)(-) \Rightarrow - + 50 = -300 + 50 = -250$

therefore ΔG will become less negative

(c) $\Delta S = \Sigma S^\circ$ products $- \Sigma S^\circ$ reactants

$\Delta S = 2(NO_2) - [O_2 + 2(NO)]$

$-146.5 = 2(240.1) - [O_2 + 2(210.8)]$

$-146.5 = 480.2 - [O_2 + 421.6]$

$-626.7 = -O_2 - 421.6$

$-205.1 = -O_2$

$O_2 = 205.1$ J/Kmol

(d) $\Delta H = 2(NO_2) - [O_2 + 2(NO)]$

$-114.1 = 2(NO_2) - [495 + 2(607)]$

$-114.1 = 2(NO_2) - 1709$

$\frac{1594.9}{2} = \frac{2(NO_2)}{2}$

$NO_2 = 797.45 \Rightarrow$ 798 kJ/mol