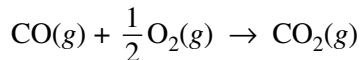


**AP<sup>®</sup> CHEMISTRY  
2006 SCORING GUIDELINES**

**Question 2**

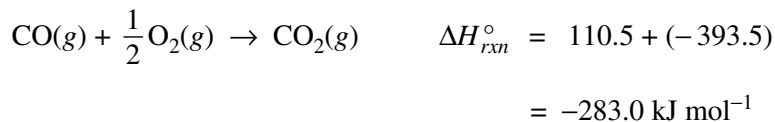


2. The combustion of carbon monoxide is represented by the equation above.

- (a) Determine the value of the standard enthalpy change,  $\Delta H_{rxn}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the following information.



Reverse the first equation and add it to the second equation to obtain the third equation.



**OR**

$$\Delta H_{rxn}^\circ = \Delta H_f^\circ \text{ of CO}_2(g) - \Delta H_f^\circ \text{ of CO}(g)$$

$$= -393.5 \text{ kJ mol}^{-1} - (-110.5 \text{ kJ mol}^{-1}) = -283.0 \text{ kJ mol}^{-1}$$

One point is earned for reversing the first equation.

One point is earned for the correct answer (with sign).

**OR**

Two points are earned for determining  $\Delta H_{rxn}^\circ$  from the enthalpies of formation.

(If sign is incorrect, only one point is earned.)

- (b) Determine the value of the standard entropy change,  $\Delta S_{rxn}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the information in the following table.

Substance	$S_{298}^\circ$ (J mol <sup>-1</sup> K <sup>-1</sup> )
CO(g)	197.7
CO <sub>2</sub> (g)	213.7
O <sub>2</sub> (g)	205.1

**AP<sup>®</sup> CHEMISTRY  
2006 SCORING GUIDELINES**

**Question 2 (continued)**

$\Delta S_{rxn}^{\circ} = 213.7 \text{ J mol}^{-1} \text{ K}^{-1} - (197.7 \text{ J mol}^{-1} \text{ K}^{-1} + \frac{1}{2}(205.1 \text{ J mol}^{-1} \text{ K}^{-1}))$ $= -86.5 \text{ J mol}^{-1} \text{ K}^{-1}$	<p>One point is earned for taking one-half of <math>S_{298}^{\circ}</math> for <math>\text{O}_2(\text{g})</math>.</p> <p>One point is earned for the answer (with sign).</p>
---	--

(c) Determine the standard free energy change,  $\Delta G_{rxn}^{\circ}$ , for the reaction at 298 K. Include units with your answer.

$\Delta G_{rxn}^{\circ} = \Delta H_{rxn}^{\circ} - T \Delta S_{rxn}^{\circ}$ $= -283.0 \text{ kJ mol}^{-1} - (298 \text{ K})(-0.0865 \text{ kJ mol}^{-1} \text{ K}^{-1})$ $\Delta G_{rxn}^{\circ} = -257.2 \text{ kJ mol}^{-1}$	<p>One point is earned for substituting the values from parts (a) and (b) into the equation.</p> <p>One point is earned for the answer (with sign and units).</p>
---	---

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

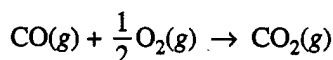
<p>Yes, the reaction is spontaneous because the value of <math>\Delta G_{rxn}^{\circ}</math> for the reaction is negative (<math>-257.2 \text{ kJ mol}^{-1}</math>).</p>	<p>One point is earned for an answer with justification (consistent with the answer in part (c)).</p>
--	---

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

$\Delta G_{rxn}^{\circ} = -RT \ln K_{eq} \Rightarrow \frac{\Delta G_{rxn}^{\circ}}{-RT} = \ln K_{eq}$ $\frac{-257,200 \text{ J mol}^{-1}}{-(8.31 \text{ J mol}^{-1} \text{ K}^{-1})(298 \text{ K})} = \ln K_{eq} \Rightarrow K_{eq} = 1.28 \times 10^{45}$	<p>One point is earned for correct substitution into the equation.</p> <p>One point is earned for the answer.</p>
--	---

Answer EITHER Question 2 below OR Question 3 printed on page 16. 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.

2A<sub>1</sub>



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

(a) Determine the value of the standard enthalpy change,  $\Delta H_{\text{rxn}}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the following information.



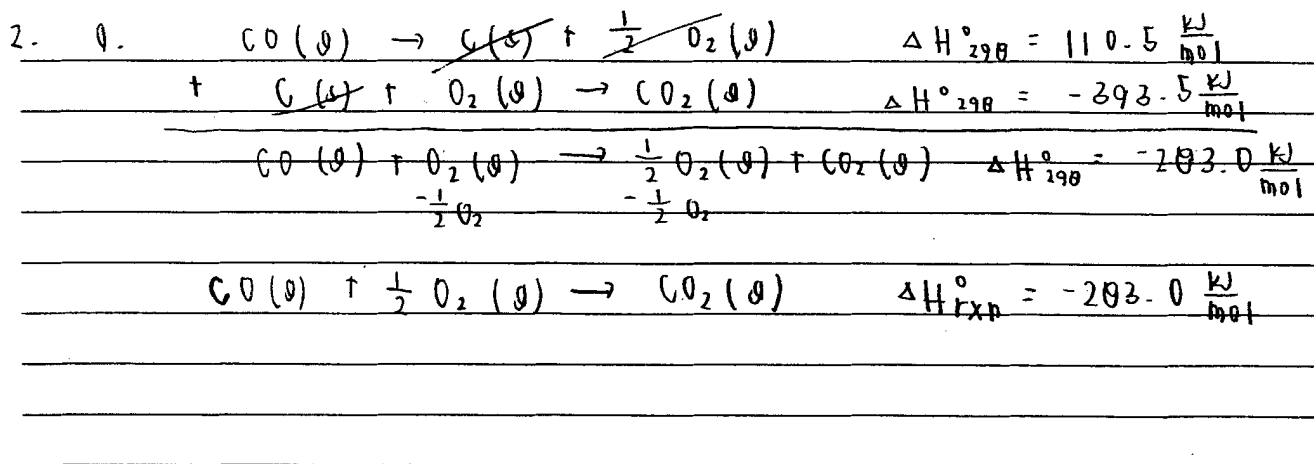
(b) Determine the value of the standard entropy change,  $\Delta S_{\text{rxn}}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the information in the following table.

Substance	$S_{298}^\circ$ ( $\text{J mol}^{-1} \text{K}^{-1}$ )
$\text{CO}(g)$	197.7
$\text{CO}_2(g)$	213.7
$\text{O}_2(g)$	205.1

(c) Determine the standard free energy change,  $\Delta G_{\text{rxn}}^\circ$ , 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,  $K_{\text{eq}}$ , for the reaction at 298 K.



GO ON TO THE NEXT PAGE.

$$\begin{aligned}
 b. \quad \Delta S_{\text{rxn}}^{\circ} &= (213.7 \frac{\text{J}}{\text{mol}\cdot\text{K}}) - \left[ (197.7 \frac{\text{J}}{\text{mol}\cdot\text{K}}) + \frac{1}{2} (205.1 \frac{\text{J}}{\text{mol}\cdot\text{K}}) \right] \\
 &= 213.7 - (197.7 + 102.55) \\
 &= 213.7 \frac{\text{J}}{\text{mol}\cdot\text{K}} - 300.25 \frac{\text{J}}{\text{mol}\cdot\text{K}} \\
 &= -86.6 \frac{\text{J}}{\text{mol}\cdot\text{K}}
 \end{aligned}$$

$$\begin{aligned}
 c. \quad \Delta G^{\circ} &= \Delta H^{\circ} - T \Delta S^{\circ} \\
 &= -203.0 \frac{\text{kJ}}{\text{mol}} - 290 \text{K} \left( \frac{-86.6 \text{J}}{\text{mol}\cdot\text{K}} \right) \left( \frac{1 \text{kJ}}{1000 \text{J}} \right) \\
 &= -203.0 \frac{\text{kJ}}{\text{mol}} - -25.8 \frac{\text{kJ}}{\text{mol}} \\
 &= -257.2 \frac{\text{kJ}}{\text{mol}}
 \end{aligned}$$

d. Yes the reaction is spontaneous at 290 K.

$\Delta G^{\circ} = -257.2 \frac{\text{kJ}}{\text{mol}}$ ; a negative number meaning the reaction is spontaneous.

$$e. \quad \Delta G = \Delta G^{\circ} + RT \ln Q$$

$$\Delta G = 0 \text{ at equilibrium}$$

$$0 = \Delta G^{\circ} + RT \ln K$$

$$\Delta G^{\circ} = -RT \ln K$$

$$-257.2 \frac{\text{kJ}}{\text{mol}} = - \left( 8.31 \frac{\text{J}}{\text{mol}\cdot\text{K}} \right) \left( \frac{1 \text{kJ}}{1000 \text{J}} \right) (290 \text{K}) \ln K$$

$$-257.2 \text{kJ} = -2.47638 \frac{\text{kJ}}{\text{mol}} \ln K$$

$$\ln K = 103.8612014$$

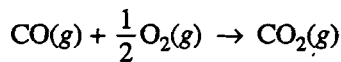
$$0 \quad 0$$

$$K_{\text{eq}} = 1.20 \times 10^{45}$$

GO ON TO THE NEXT PAGE.

Answer EITHER Question 2 below OR Question 3 printed on page 16. 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.

2B,



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

- (a) Determine the value of the standard enthalpy change,  $\Delta H_{\text{rxn}}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the following information.  $\Delta H_{\text{rxn}}^\circ = -283.0 \text{ kJ/mol}$



- (b) Determine the value of the standard entropy change,  $\Delta S_{\text{rxn}}^\circ$ , for the combustion of  $\text{CO}(g)$  at 298 K using the information in the following table.  $\Delta S_{\text{rxn}}^\circ = -86.6 \text{ J/mol K}$

Substance	$S_{298}^\circ$ ( $\text{J mol}^{-1} \text{K}^{-1}$ )
$\text{CO}(g)$	197.7
$\text{CO}_2(g)$	213.7
$\text{O}_2(g)$	205.1

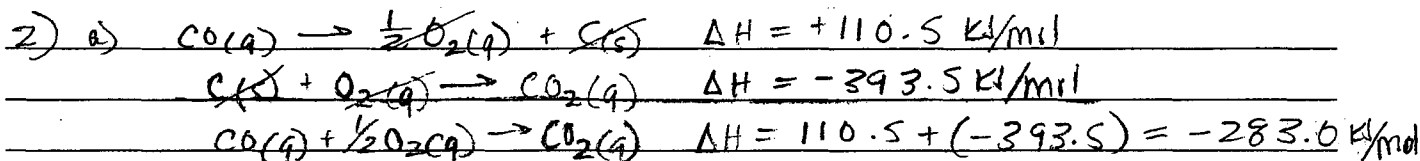
- (c) Determine the standard free energy change,  $\Delta G_{\text{rxn}}^\circ$ , for the reaction at 298 K. Include units with your answer.  $\Delta G_{\text{rxn}}^\circ = -308.8 \text{ kJ/mol}$

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

No because  $\Delta G$  must be positive value for the reaction to be spontaneous.

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

$$K_{\text{eq}} = 1.43 \times 10^{54}$$



b)  $\Delta S = \sum \Delta S_{\text{prod}} - \sum \Delta S_{\text{react}}$

$$\Delta S = (213.7) - (\frac{1}{2}(205.1) + (197.7))$$

$$\Delta S = -86.6 \text{ J/mol K}$$

$$-86.6 \text{ J} \times \frac{\text{kJ}}{1000 \text{ J}} =$$

c)  $\Delta G = \Delta H - T\Delta S$

$$\Delta G = (-283.0 \frac{\text{kJ}}{\text{mol}}) - (298 \text{ K})(-0.0866 \frac{\text{kJ}}{\text{mol K}})$$

$$\Delta G_{\text{rxn}}^\circ = -308.8 \text{ kJ/mol}$$

GO ON TO THE NEXT PAGE.

d) The reaction is not spontaneous because  $\Delta G$  must be a positive value for the reaction to be spontaneous. The  $\Delta G$  for this reaction is a negative value and therefore the reaction is nonspontaneous.  $\Delta G$  calculated is the amount of energy you need to make the reaction occur.

e)  $\Delta G = -RT \ln K_{eq}$ .

$$-308.9 \frac{\text{kJ}}{\text{mol}} = -(0.00831)(298)(\ln K_{eq})$$

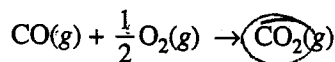
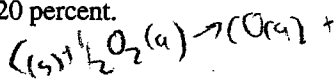
$$\ln K_{eq} = 124.698$$

$$K_{eq} = 1.43 \times 10^{54}$$

GO ON TO THE NEXT PAGE.

Answer EITHER Question 2 below OR Question 3 printed on page 16. 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.

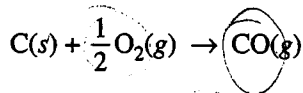
20,



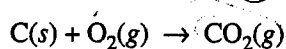
2. The combustion of carbon monoxide is represented by the equation above.

(a) Determine the value of the standard enthalpy change,  $\Delta H_{rxn}^\circ$ , for the combustion of  $CO(g)$  at 298 K using the following information.

$\Delta H_{prod} - \Delta H_{react}$



$$\Delta H_{298}^\circ = -110.5 \text{ kJ mol}^{-1}$$



$$\Delta H_{298}^\circ = -393.5 \text{ kJ mol}^{-1} - 393.5 \text{ kJ mol}^{-1}$$

(b) Determine the value of the standard entropy change,  $\Delta S_{rxn}^\circ$ , for the combustion of  $CO(g)$  at 298 K using the information in the following table.

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

(c) Determine the standard free energy change,  $\Delta G_{rxn}^\circ$ , 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.

$$201. \Delta H^\circ = \sum \Delta H_{prod}^\circ - \sum \Delta H_{react}^\circ \quad \Delta H_f^\circ \frac{1}{2} O_2(g) = 0$$

$$\Delta H_{rxn}^\circ = \Delta H_f^\circ CO_2(g) - \Delta H_f^\circ CO(g) = -393.5 \text{ kJ mol}^{-1} - (-110.5 \text{ kJ mol}^{-1}) = -283 \text{ kJ mol}^{-1}$$

$$b) \Delta S_{rxn}^\circ = \sum S_{products}^\circ - \sum S_{reactants}^\circ = 213.7 \text{ J mol}^{-1} \text{ K}^{-1} - (197.7 + 205.1) = -189.1 \text{ J mol}^{-1} \text{ K}^{-1}$$

GO ON TO THE NEXT PAGE.

2C<sub>2</sub>

$$c) \Delta G_{rxn}^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$$

$\frac{kJ}{mol} - (K \cdot \frac{J}{mol \cdot K})$

$$= -283 \text{ kJ mol}^{-1} - (298 \text{ K} \cdot -189.1 \text{ J mol}^{-1} \text{ K}^{-1})$$

$$= -283 \text{ kJ mol}^{-1} - (298 \text{ K} \cdot -0.1891 \text{ kJ mol}^{-1} \text{ K}^{-1})$$

$$= -226.6482 \text{ kJ mol}^{-1}$$

d) ~~Yes this reaction is spontaneous because the change in free energy is negative, and it is exothermic,  $\Delta H_{rxn}$  is negative, so excess energy is given off.~~

d) Yes This reaction is spontaneous because  $\Delta H_{rxn}$  is <sup>neg. entis</sup> exothermic and  $\Delta G$  is negative (change in free energy), and because  $\Delta S$  is negative (change in entropy).

GO ON TO THE NEXT PAGE.



**AP<sup>®</sup> CHEMISTRY**  
**2006 SCORING COMMENTARY**

**Question 2**

**Overview**

The intent of this question was to test students' knowledge of basic thermodynamic relationships, including enthalpy, entropy, and free energy changes, and the equilibrium constant associated with a chemical reaction. The first task was to calculate the values for enthalpy, entropy, and free energy changes from the information provided. Students were asked what the calculated thermodynamic quantities implied for the spontaneity of the reaction given,  $\text{CO} + \frac{1}{2} \text{O}_2 \rightarrow \text{CO}_2$ , and had to calculate the thermodynamic equilibrium constant.

**Sample: 2A**

**Score: 9**

This response earned all 9 points: 2 points for part (a), 2 points for part (b), 2 points for part (c), 1 point for part (d), and 2 points for part (e). Note that the student has an unusual way of writing the numeral "8," which is consistently used throughout the response.

**Sample: 2B**

**Score: 7**

Only 1 point was earned in part (c) because the negative sign in  $(-.0866 \text{ kJ/mol K})$  is not carried through correctly, constituting a math error. The point was not earned in part (d) because of the incorrect conclusion that a negative Gibbs free energy is an indication of nonspontaneity.

**Sample: 2C**

**Score: 5**

Only 1 point was earned in part (b) because the entropy of oxygen is not multiplied by  $\frac{1}{2}$ , its coefficient in the combustion equation. This incorrect value is used correctly in part (c), but only 1 out of 2 points were earned in this part because the number of significant figures in the answer is off by more than one. In part (d) 1 point was earned, but part (e) is not attempted.