



AP[®] Chemistry (Operational) 2004 Sample Student Responses

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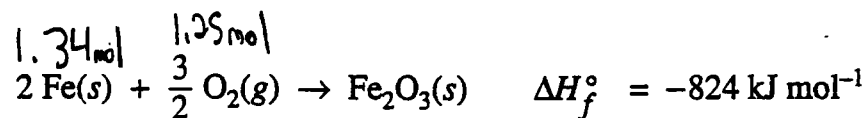
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2A,

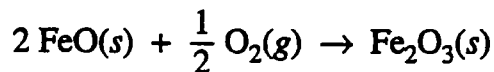
Answer EITHER Question 2 below OR Question 3 printed on page 16. Only one of these two questions will be graded. If you start both questions, be sure to cross out the question you do not want graded. The Section II score weighting for the question you choose is 20 percent.



2. Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 75.0 g sample of Fe(s) is mixed with 11.5 L of O₂(g) at 2.66 atm and 298 K.

- (a) Calculate the number of moles of each of the following before the reaction begins.
 - (i) Fe(s)
 - (ii) O₂(g)
- (b) Identify the limiting reactant when the mixture is heated to produce Fe₂O₃(s). Support your answer with calculations.
- (c) Calculate the number of moles of Fe₂O₃(s) produced when the reaction proceeds to completion.
- (d) The standard free energy of formation, ΔG_f[°], of Fe₂O₃(s) is -740. kJ mol⁻¹ at 298 K.
 - (i) Calculate the standard entropy of formation, ΔS_f[°], of Fe₂O₃(s) at 298 K. Include units with your answer.
 - (ii) Which is more responsible for the spontaneity of the formation reaction at 298 K, the standard enthalpy of formation, ΔH_f[°], or the standard entropy of formation, ΔS_f[°]? Justify your answer.

The reaction represented below also produces iron(III) oxide. The value of ΔH° for the reaction is -280. kJ per mole of Fe₂O₃(s) formed.



(e) Calculate the standard enthalpy of formation, ΔH_f[°], of FeO(s).

@ (i) $\frac{75.0 \text{ g Fe}}{55.85 \frac{\text{g}}{\text{mol}}} = 1.34 \text{ mol Fe}$

(ii) $PV = nRT$

$n = \frac{PV}{RT}$

$n = \frac{(2.66 \text{ atm})(11.5 \text{ L})}{(0.0821 \frac{\text{L atm}}{\text{K mol}})(298 \text{ K})}$

$n = 1.25 \text{ mol O}_2$

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$$\textcircled{b} \quad \frac{\frac{3}{2} \text{ mol } O_2}{2 \text{ mol Fe}} = \frac{1.25 \text{ mol } O_2}{x \text{ mol Fe}} \qquad \frac{\frac{3}{2} \text{ mol } O_2}{2 \text{ mol Fe}} = \frac{x \text{ mol } O_2}{1.34 \text{ mol Fe}}$$

$$x = 1.67 \text{ mol Fe} \qquad x = 1.01 \text{ mol } O_2$$

More oxygen than necessary \rightarrow only 1.01 mol O_2 needed.
Fe is the limiting reactant.

$$\textcircled{c} \quad \frac{2 \text{ mol Fe}}{1 \text{ mol } Fe_2O_3} = \frac{1.34 \text{ mol Fe}}{x \text{ mol } Fe_2O_3}$$

$$x = .670 \text{ mol } Fe_2O_3$$

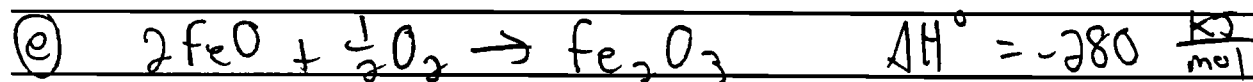
$$\textcircled{d} \quad \Delta G_f^\circ = \Delta H_f^\circ - T \Delta S_f^\circ$$

$$\text{(i) } \Delta S_f^\circ = \frac{\Delta H_f^\circ - \Delta G_f^\circ}{T}$$

$$= \frac{-824 \frac{\text{kJ}}{\text{mol}} - (-740 \frac{\text{kJ}}{\text{mol}})}{298 \text{ K}}$$

$$\Delta S_f^\circ = -.282 \frac{\text{kJ}}{\text{K} \cdot \text{mol}}$$

(ii) A reaction can be spontaneous if either its ΔH° is negative or its ΔS° is positive. In this reaction, the $\Delta H^\circ < 0$ + $\Delta S^\circ < 0$. Therefore the ΔH° contributes most to the spontaneity of the reaction ($\Delta G^\circ < 0$).

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


$$1(\Delta H_f^\circ Fe_2O_3) - 2(\Delta H_f^\circ FeO) = -280 \frac{\text{kJ}}{\text{mol}}$$

$$(-824 \frac{\text{kJ}}{\text{mol}}) - 2x = -280 \frac{\text{kJ}}{\text{mol}}$$

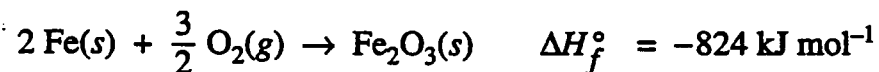
$$-2x = 544 \frac{\text{kJ}}{\text{mol}}$$

$$x = -272 \frac{\text{kJ}}{\text{mol}}$$

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2B,

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(a) Calculate the number of moles of each of the following before the reaction begins.

(i) Fe(s)

(ii) O₂(g)

(b) Identify the limiting reactant when the mixture is heated to produce Fe₂O₃(s). Support your answer with calculations.

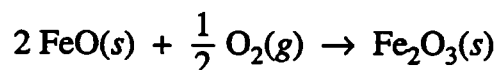
(c) Calculate the number of moles of Fe₂O₃(s) produced when the reaction proceeds to completion.

(d) The standard free energy of formation, ΔG_f° , of Fe₂O₃(s) is -740. kJ mol⁻¹ at 298 K.

(i) Calculate the standard entropy of formation, ΔS_f° , of Fe₂O₃(s) at 298 K. Include units with your answer.

(ii) Which is more responsible for the spontaneity of the formation reaction at 298 K, the standard enthalpy of formation, ΔH_f° , or the standard entropy of formation, ΔS_f° ? Justify your answer.

The reaction represented below also produces iron(III) oxide. The value of ΔH° for the reaction is -280. kJ per mole of Fe₂O₃(s) formed.



(e) Calculate the standard enthalpy of formation, ΔH_f° , of FeO(s).

$$a) i) \text{Fe}(s) \quad 75 \text{g Fe}(s) \cdot \frac{1 \text{ mol Fe}}{55.85 \text{g Fe}} = \boxed{1.34 \text{ mol Fe}}$$

$$ii) 11.5 \text{L O}_2 \cdot \frac{1 \text{ mol O}_2}{22.4 \text{L O}_2} = \boxed{0.513 \text{ mol O}_2}$$

$$b) \frac{\text{Fe}(s)}{\text{O}_2(g)} = \frac{2}{\frac{3}{2}} = 2 \cdot \frac{2}{3} = \frac{4}{3} = \frac{1.333 \text{ mol Fe}}{1 \text{ mol O}_2}$$

$$\frac{1.34 \text{ mol Fe}}{0.513 \text{ mol O}_2} = \frac{2.6157 \text{ mol Fe}}{1 \text{ mol O}_2}$$

Since $\frac{4}{3} < 2.6157$, O_2 is the limiting reactant
(There is more than enough Fe(s))

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2 B₂

ADDITIONAL PAGE FOR ANSWERING QUESTION 2.

$$c) \text{ Hence, } \frac{.513 \text{ mol O}_2 \cdot 1 \text{ mol Fe}_2\text{O}_3}{\frac{3}{2} \text{ mol O}_2} = \boxed{.342 \text{ mol Fe}_2\text{O}_3}$$

$$d) \Delta G_f^\circ \text{ Fe}_2\text{O}_3 = -740 \text{ kJ/mol}$$

$$a) \Delta G_f^\circ = \Delta H - T\Delta S^\circ$$

$$-740 \text{ kJ/mol} = -824 \text{ kJ/mol} - 298 \text{ K} (\Delta S_f^\circ)$$

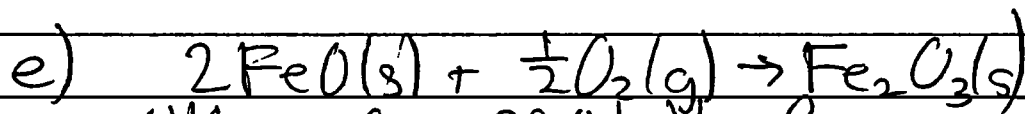
$$84 \text{ kJ/mol} = -298 \text{ K} (\Delta S_f^\circ)$$

$$\Delta S_f^\circ = \boxed{-282 \text{ kJ/mol}\cdot\text{K}}$$

ii) If you examine both values, it is clear that one dominates the other.

$$\Delta H_f^\circ = -824 \text{ kJ/mol} \quad -T\Delta S = 84 \text{ kJ/mol}$$

Clearly, ΔH_f° is the larger value, and thus the standard enthalpy of formation is responsible for the spontaneity of the reaction. ΔS_f° is far too small ($-282 \text{ kJ/mol}\cdot\text{K}$) to be a significant factor.



$$\Delta H_f^\circ \text{ Fe}_2\text{O}_3 = -824 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ Fe}_2\text{O}_3 = -824 \text{ kJ/mol}$$

$$\Delta H_f^\circ \text{ FeO} = ?$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$-280 \text{ kJ/mol} = -824 \text{ kJ/mol} - [2(\Delta H_f^\circ \text{ FeO}) + \frac{1}{2}(\Delta H_f^\circ \text{ O}_2)]$$

$$\text{Since } \Delta H_f^\circ \text{ O}_2 = 0,$$

$$-280 \text{ kJ/mol} = -824 \text{ kJ/mol} - 2(\Delta H_f^\circ \text{ FeO})$$

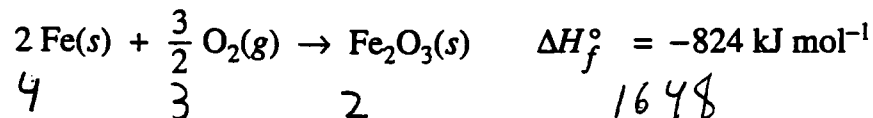
$$544 \text{ kJ/mol} = -2(\Delta H_f^\circ \text{ FeO})$$

$$\boxed{-272 \text{ kJ/mol} = \Delta H_f^\circ \text{ FeO}}$$

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2C1

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.



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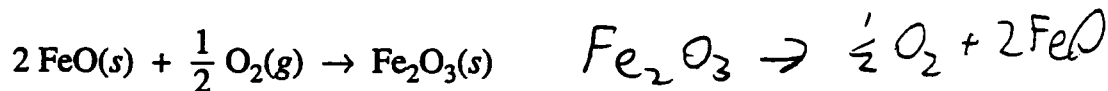
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(e) Calculate the standard enthalpy of formation, ΔH_f[∘], of FeO(s).

a. i. Fe 1.34 moles

ii. O₂ 2.66 (11.5) = n(0.0821)(298)

n = 1.25 moles

b. 1.34 × (3/2) = 1.005 moles O₂ needed Fe is limiting

1.25 × (1/2) = 1.66 moles Fe needed

c. 1.34 × (1/2) = .67 moles Fe₂O₃

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$$\begin{aligned}d. i \quad \Delta G &= \Delta H - T\Delta S \\ &= -1480000 = -824000 - (298)\Delta S \\ &\quad -656000 = -298\Delta S \\ \Delta S &= 2201 \text{ J/mol}\cdot\text{K}\end{aligned}$$

i. Standard entropy of formation is more responsible because it is a measure of the order of the reaction. It tells whether or not the reaction is orderly or chaotic.

$$e. \quad \frac{-(-280)}{2} = 140 \text{ kJ/mol}$$

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