

## THERMODYNAMICS FRQS #1

1. Liquid heptane,  $C_7H_{16}(l)$ , is completely combusted to produce  $CO_2(g)$  and  $H_2O(l)$ . The heat of combustion,  $\Delta H^\circ_{\text{comb}}$ , for heptane is  $-4.85 \times 10^3$  kJ.

A. Write a balanced equation for the reaction.

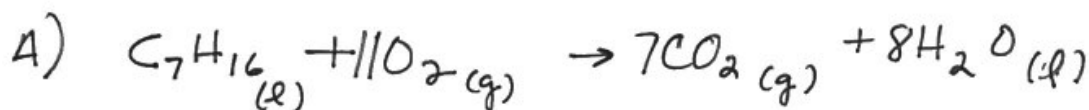
B. Using the information in the table below, calculate the value of  $\Delta H^\circ_f$  for heptane in kJ/mole.

Compound	$\Delta H^\circ_f$ (kJ/mole)
$CO_2(g)$	-393.5
$H_2O(l)$	-285.8

C. 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/ $^\circ\text{C}$ , calculate the temperature change of the calorimeter.



$$B) \quad -4.85 \times 10^3 \text{ kJ} = [7(-393.5) + 8(-285.8)] - [H_f^\circ + 11(0)]$$

$$-4.85 \times 10^3 \text{ kJ} = -2754.5 - 2286.4 - H_f^\circ$$

$$H_f^\circ = -190.9 \text{ kJ/mol}$$

$$C) \quad i) \quad \frac{0.0108 \text{ mol} \times -4.85 \times 10^3 \text{ kJ}}{1 \text{ mol}} = 52.4 \text{ kJ released}$$

$$ii) \quad \frac{52.4 \text{ kJ}}{9.273 \text{ kJ/}^\circ\text{C}} = 5.65^\circ\text{C}$$

$$\Delta T$$

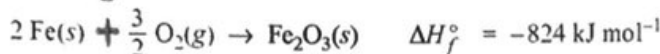
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#2

Question 2

75.0g



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 O<sub>2</sub>(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)

$n_{\text{Fe}} = 75.0 \text{ g Fe} \times \left( \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \right) = 1.34 \text{ mol Fe}$	1 point for number of moles of Fe(s)
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(ii) O<sub>2</sub>(g)

$PV = nRT$ $n_{\text{O}_2} = \frac{PV}{RT} = \frac{2.66 \text{ atm} \times 11.5 \text{ L}}{0.0821 \frac{\text{L atm}}{\text{mol K}} \times 298 \text{ K}}$ $n_{\text{O}_2} = 1.25 \text{ mol O}_2$	1 point for number of moles of O <sub>2</sub> (g)
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(b) Identify the limiting reactant when the mixture is heated to produce Fe<sub>2</sub>O<sub>3</sub>(s). Support your answer with calculations.

$n_{\text{O}_2 \text{ reacting}} = 1.34 \text{ mol Fe} \times \left( \frac{1.5 \text{ mol O}_2}{2 \text{ mol Fe}} \right)$ $= 1.01 \text{ mol O}_2$ <p>There is 1.25 mol O<sub>2</sub> initially, so there is an excess of O<sub>2</sub>, and Fe is the limiting reactant.</p> <p style="text-align: center;"><u>OR</u></p> $n_{\text{Fe reacting}} = 1.25 \text{ mol O}_2 \times \left( \frac{2 \text{ mol Fe}}{1.5 \text{ mol O}_2} \right)$ $= 1.67 \text{ mol Fe}$ <p>There is 1.34 mol Fe initially, so there is not enough Fe, and Fe is the limiting reactant.</p>	<p>1 point for identifying limiting reactant</p> <p>1 point for supporting calculation</p>
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$$\frac{1.34 \text{ moles Fe}}{1} \times \frac{1 \text{ mole Fe}_2\text{O}_3}{2 \text{ moles Fe}} = .67 \text{ mole Fe}_2\text{O}_3$$

$$\frac{1.25 \text{ mole O}_2}{1} \times \frac{1 \text{ mole Fe}_2\text{O}_3}{1.5 \text{ mole O}_2} = .83 \text{ mole Fe}_2\text{O}_3$$

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**Question 2 (cont'd.)**

- (c) Calculate the number of moles of  $\text{Fe}_2\text{O}_3(s)$  produced when the reaction proceeds to completion.

$n_{\text{Fe}_2\text{O}_3} = 1.34 \text{ mol Fe} \times \left( \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \right) = 0.670 \text{ mol Fe}_2\text{O}_3$	1 point for number of moles of $\text{Fe}_2\text{O}_3$ produced
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- (d) The standard free energy of formation,  $\Delta G_f^\circ$ , of  $\text{Fe}_2\text{O}_3(s)$  is  $-740. \text{ kJ mol}^{-1}$  at 298 K.

- (i) Calculate the standard entropy of formation,  $\Delta S_f^\circ$ , of  $\text{Fe}_2\text{O}_3(s)$  at 298 K. Include units with your answer.

$\Delta G_f^\circ = \Delta H_f^\circ - T \Delta S_f^\circ$ $-740. \text{ kJ mol}^{-1} = -824 \text{ kJ mol}^{-1} - (298 \text{ K}) \Delta S_f^\circ$ $+84 \text{ kJ mol}^{-1} = -(298 \text{ K}) \Delta S_f^\circ$ $\Delta S_f^\circ = \frac{+84 \text{ kJ mol}^{-1}}{-298 \text{ K}} = -0.28 \text{ kJ mol}^{-1} \text{ K}^{-1}$ <p align="center"><i>-280 J mol<sup>-1</sup> K<sup>-1</sup></i></p>	<p>1 point for calculation of <math>\Delta S_f^\circ</math></p> <p>1 point for correct units</p>
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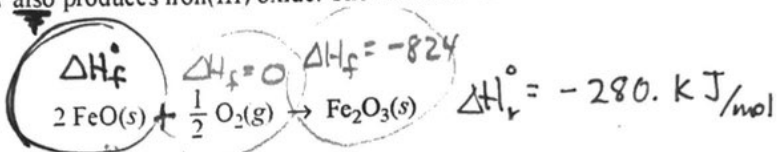
- (ii) Which is more responsible for the spontaneity of the formation reaction at 298 K, the standard enthalpy of formation,  $\Delta H_f^\circ$ , or the standard entropy of formation,  $\Delta S_f^\circ$ ? Justify your answer.

<p><math>\Delta H_f^\circ</math> is the more important factor. The reaction is exothermic, which favors spontaneity. <math>\Delta S_f^\circ</math> is negative, which means the system becomes more ordered as the reaction proceeds. Greater order will not increase the spontaneity of the reaction.</p>	<p>1 point for indicating that <math>\Delta H_f^\circ</math> is responsible and for an explanation that addresses the signs of <math>\Delta H^\circ</math> and <math>\Delta S^\circ</math></p>
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Question 2 (cont'd.)

The reaction represented below also produces iron(III) oxide. The value of  $\Delta H^\circ$  for the reaction is  $-280.$  kJ per mole of  $\text{Fe}_2\text{O}_3(\text{s})$  formed.



(e) Calculate the standard enthalpy of formation,  $\Delta H_f^\circ$ , of  $\text{FeO}(\text{s})$ .

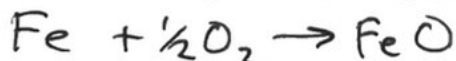
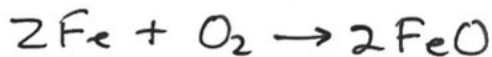
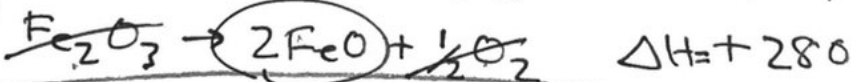
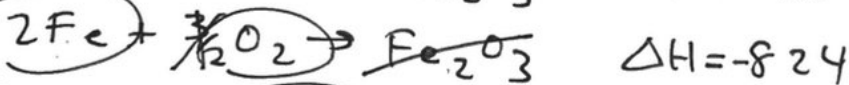
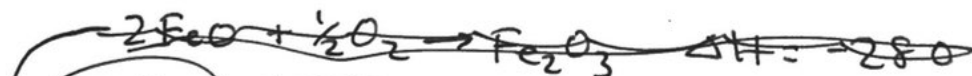
$\Delta H_{rxn}^\circ = \sum \Delta H_f^\circ (\text{products}) - \sum \Delta H_f^\circ (\text{reactants})$ $\Delta H_{rxn}^\circ = \Delta H_f^\circ \text{Fe}_2\text{O}_3(\text{s}) - [2\Delta H_f^\circ \text{ of FeO}(\text{s}) + \frac{1}{2} \Delta H_f^\circ \text{O}_2(\text{g})]$ $-280. \text{ kJ mol}^{-1} = [-824 \text{ kJ mol}^{-1}] - [2(\Delta H_f^\circ \text{FeO}(\text{s})) + \frac{1}{2}(0)]$ $+544 \text{ kJ mol}^{-1} = -2\Delta H_f^\circ \text{FeO}(\text{s})$ $-272 \text{ kJ mol}^{-1} = \Delta H_f^\circ \text{FeO}(\text{s})$	<p>1 point for correct stoichiometry</p> <p>1 point for correct calculation</p>
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$$+1 \quad -544 \text{ kJ} \quad -280 = -824 - 2x$$

$$+544 = -2x$$

$$\frac{-2}{-2} \quad \frac{-2}{-2}$$

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





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**Question 2 (continued)**

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