This print-out should have 8 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

LDE Bomb Calorimeter 001
001   5.0 points

For this problem assume that the calorimeter itself absorbs no heat. If we set up a bomb calorimetry experiment to determine the molar enthalpy of combustion of octane (C_8H_{18}) using 1 L of water as our heat sink, 1.142 g of octane, and measure an initial and final temperature of 25.01 °C and 37.99 °C, respectively, what will be the experimentally determined molar enthalpy of combustion of octane? Assume the density of water is $1.00 \text{ g} \cdot \text{mL}^{-1}$. The specific heat capacity of water is $4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$.

1. $-543 \text{ kJ} \cdot \text{mol}^{-1}$

2. $-54.3 \text{ kJ} \cdot \text{mol}^{-1}$

3. $-5,430 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

4. $-54,300 \text{ kJ} \cdot \text{mol}^{-1}$

5. $-5.43 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:
$\Delta T = T_f - T_i = 37.99\,°C - 25.01\,°C = 12.98\,°C = 12.98$ K

$$m = 1 \text{ L} \cdot \frac{1000 \text{ mL}}{\text{L}} \cdot \frac{1.00 \text{ g}}{\text{mL}} = 1000 \text{ g}$$

$$n = 1.142 \text{ g octane} \cdot \frac{114.26 \text{ g}}{\text{mol}} = 0.01 \text{ mol}$$
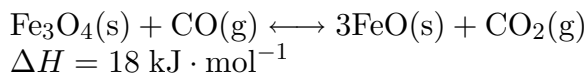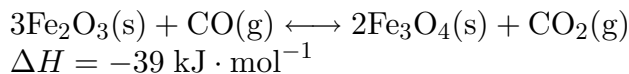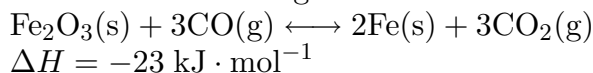
$$\Delta H_{rxn} = -\Delta H_{cal} = -mc\Delta T$$

$$\frac{-mc\Delta T}{n} = -\frac{1000 \text{ g} \cdot \frac{4.184 \text{ J}}{\text{g} \cdot \text{K}} \cdot 12.98 \text{ K}}{0.01 \text{ mol}}$$

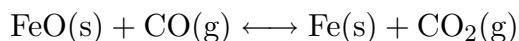$$= -5,430 \text{ kJ} \cdot \text{mol}^{-1}$$

LDE Hess' Law 001
002   5.0 points

Given the following data:
$$Fe_2O_3(s) + 3CO(g) \longleftrightarrow 2Fe(s) + 3CO_2(g)$$
$$\Delta H = -23 \text{ kJ} \cdot \text{mol}^{-1}$$

$$3Fe_2O_3(s) + CO(g) \longleftrightarrow 2Fe_3O_4(s) + CO_2(g)$$
$$\Delta H = -39 \text{ kJ} \cdot \text{mol}^{-1}$$

$$Fe_3O_4(s) + CO(g) \longleftrightarrow 3FeO(s) + CO_2(g)$$
$$\Delta H = 18 \text{ kJ} \cdot \text{mol}^{-1}$$

calculate ΔH for the reaction

$$FeO(s) + CO(g) \longleftrightarrow Fe(s) + CO_2(g)$$

1. $-11 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

2. $-66 \text{ kJ} \cdot \text{mol}^{-1}$

3. $-2 \text{ kJ} \cdot \text{mol}^{-1}$

4. $-18.5 \text{ kJ} \cdot \text{mol}^{-1}$

5. $-13.5 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:
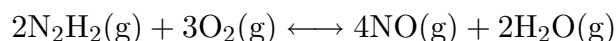$$\Delta H_{rxn} = \frac{1}{2}(-23 \text{ kJ} \cdot \text{mol}^{-1})$$
$$-\frac{1}{6}(-39 \text{ kJ} \cdot \text{mol}^{-1}) - \frac{1}{3}(18 \text{ kJ} \cdot \text{mol}^{-1})$$
$$= -11 \text{ kJ} \cdot \text{mol}^{-1}$$

LDE Bond Enthalpy 001
003   5.0 points

Using the provided bond enthalpy data, calculate the change in enthalpy for the following reaction:

$$2N_2H_2(g) + 3O_2(g) \longleftrightarrow 4NO(g) + 2H_2O(g)$$

1. $-514 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

2. $1,531 \text{ kJ} \cdot \text{mol}^{-1}$

3. $514 \text{ kJ} \cdot \text{mol}^{-1}$

4. $-1,166 \text{ kJ} \cdot \text{mol}^{-1}$
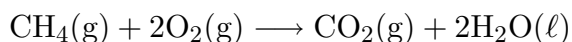
5. $-1,531 \text{ kJ} \cdot \text{mol}^{-1}$

6. $1,166 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:
$\Delta H_{rxn} = \Sigma BE_{\text{reactants}} - \Sigma BE_{\text{products}}$

$= 2 \cdot 409 \text{ kJ} \cdot \text{mol}^{-1} + 4 \cdot 388 \text{ kJ} \cdot \text{mol}^{-1}$
$+ 3 \cdot 496 \text{ kJ} \cdot \text{mol}^{-1} - 4 \cdot 630 \text{ kJ} \cdot \text{mol}^{-1}$
$- 4 \cdot 463 \text{ kJ} \cdot \text{mol}^{-1}$

$= -514 \text{ kJ} \cdot \text{mol}^{-1}$

LDE Thermodynamic Work 001
004　5.0 points

Approximate the work for the following reaction at 300 K.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$$

1. -5 kJ

2. 0 kJ

3. 5 kJ **correct**

4. -2.5 kJ

5. 2.5 kJ

Explanation:
At room temperature (300 K), the product of the RT term in $W = -\Delta n_{\text{gas}} RT$ is approximately 2.5 kJ. Since Δn_{gas} is -2 for this reaction, the work function is approximately 5 kJ.

LDE Sign Conventions 001
005　5.0 points
Work done (by/on) the system is negative, and heat (gained/lost) by the system is positive.

1. by, gained **correct**

2. on, lost

3. by, lost

4. on, gained

Explanation:
The sign conventions for process functions (like work and heat) are such that a positive sign is attributed to the system if it gains energy (in the form of heat or work).

LDE Entropy 001
006　5.0 points

If both Δn_{gas} and Δn for a particular reaction are zero, which statement below would be the most accurate prediction we could make about the change in entropy?

1. It is definitely negative.

2. It is definitely positive.

3. It may be either positive or negative and is probably very large.

4. It may be either positive or negative and is probably very small. **correct**

Explanation:
Because the best way to predict the sign and magnitude of the change in entropy is based on the change in the moles of reactants and products (particularly of gaseous reactants or products), the sign would be impossible to predict, but the change would likely be very small.

LDE Reaction Spontaneity 001
007　5.0 points
Whether a reaction is spontaneous or non-spontaneous will be temperature independent if:
  I) ΔH is positive and ΔS is negative.
 II) ΔH is positive and ΔS is positive.
III) ΔH is negative and ΔS is negative.
IV) ΔH is negative and ΔS is positive.

1. I only

2. II only

3. I, IV **correct**

4. II, III

5. III only

6. IV only

Explanation:

If ΔH is positive and ΔS is negative, the reaction is non-spontaneous at all temperatures, and thus temperature independent. If ΔH is negative and ΔS is positive, the reaction is spontaneous at all temperatures, and thus temperature independent.

LDE Thermodynamic Theory 002
008   5.0 points

Which of the following statements about the first and second laws of themodynamics is/are true?

I) The energy of the universe is always conserved.

II) The energy of a system is always conserved.

III) The energy of the system always increases.

IV) The entropy of the universe always increases.

V) The entropy of the universe always conserved.

VI) The entropy of the system is always conserved.

1. I and V

2. II and V

3. II and VI

4. III and VI

5. III and IV

6. I and IV **correct**

Explanation: