

This print-out should have 6 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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**LDE Bomb Calorimeter 006**

**001** 10.0 points

Assuming the apparatus itself absorbs no heat, what will be the change in temperature if 1000 mL of water are used as a heat sink for a reaction that releases 8 kJ of heat?

1. 1.91 K **correct**

2. 8.00 K

3. -271.09 °C

4. -265 °C

5. -1.91 K

**Explanation:**

$$q = m \cdot c \cdot \Delta T$$

$$8,000 \text{ J} = (1000 \text{ g}) \cdot \left(4.184 \frac{\text{J}}{\text{g} \cdot \text{K}}\right) \cdot \Delta T$$

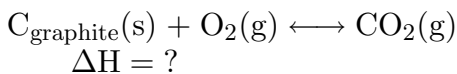
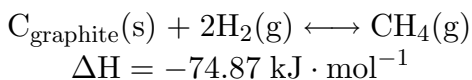
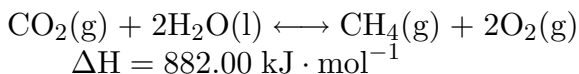
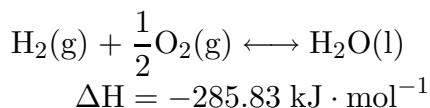
$$\Delta T = 1.91 \text{ K}$$

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**LDE Hess' Law 006**

**002** 10.0 points

Calculate the change in enthalpy for the combustion of graphite using the data below.



1. -230.04 kJ · mol<sup>-1</sup>

2. 521.30 kJ · mol<sup>-1</sup>

3. -515.87 kJ · mol<sup>-1</sup>

4. -386.41 kJ · mol<sup>-1</sup> **correct**

5. 235.47 kJ · mol<sup>-1</sup>

**Explanation:**

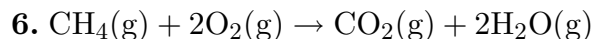
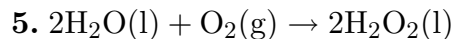
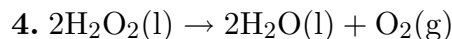
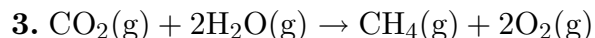
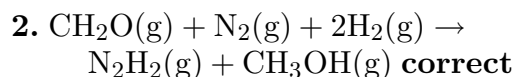
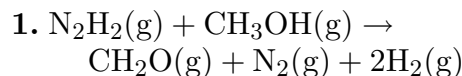
In order for the three provided reactions to cancel to result in the unknown reaction (the combustion of graphite), the first reaction needs to be reversed and doubled, the second needs to be reversed, and the third will remain unchanged. Consequently, the overall change in enthalpy for the reaction is:  $\Delta H_{\text{rxn}} = -2 \times -285.83 + -1 \times 882.00 + -74.87 = -386.41 \text{ kJ} \cdot \text{mol}^{-1}$

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**LDE Thermodynamic Work 0034**

**003** 10.0 points

For which of the following reactions at room temperature (25 °C) would there be 5.0 kJ of work done on the system?



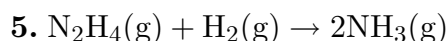
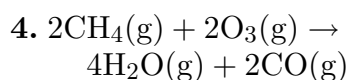
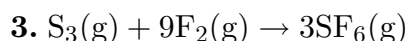
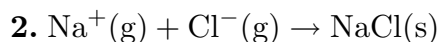
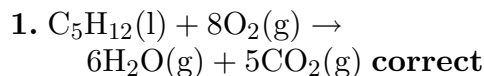
**Explanation:**

At room temperature (298 K), the product of the gas constant ( $R = 8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ ) and T is very close to  $2.5 \text{ kJ} \cdot \text{mol}^{-1}$ . Based on the equation  $5.0 \text{ kJ} = -\Delta n_{\text{gas}} \cdot 2.5 \text{ kJ} \cdot \text{mol}^{-1}$ , the reaction for which  $\Delta n_{\text{gas}}$  is -2 will be the correct answer.

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**LDE Entropy 002****004** 10.0 points

Which of the reactions below will likely have the largest increase in entropy ( $\Delta S_{rxn}$ )?

**Explanation:**

The reaction with the greatest positive value for  $\Delta n_{gas}$  will have the greatest value of  $\Delta S_{rxn}$ .

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**LDE Sign Conventions 004****005** 10.0 points

The formation of ammonia from hydrogen and nitrogen gases becomes less and less spontaneous as temperature is increased, eventually becoming non-spontaneous at sufficiently high temperatures. Which of the following statements must be true?

1. The change in entropy is large.
2. The reaction is endothermic.
3. The reaction is exothermic. **correct**
4. The change in entropy is small.

**Explanation:**

Since the reaction becomes more and more spontaneous as the temperature is lowered, it must be spontaneous at  $T = 0$  K. Since  $\Delta G = \Delta H$  at  $T = 0$  K,  $\Delta H$  must be negative and the reaction is exothermic.

How much energy would be associated with the non-vibrational motion of 2 moles of  $N_2$ ?

1.  $\frac{5}{2}kT$

2.  $3RT$

3.  $5RT$  **correct**

4.  $5kT$

5.  $6kT$

6.  $\frac{5}{2}RT$

**Explanation:**

Total number of motional modes is equal to 3 times the number of atoms in the molecule. In the case of  $N_2$ , this is 6. It has 3 translational and 2 rotational modes (rather than 3, because it is linear). The sixth and final mode is vibrational; for a total of 5 non-vibrational modes. This would amount to  $5RT$ , 2 moles times  $1/2 RT$  per mode times 5 modes.

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**LDE Stat Thermo Energy 002****006** 10.0 points