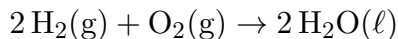


This print-out should have 72 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

For the reaction



find the value for the work done at 300 K.

1. 2.5 kJ
2. -7.5 kJ
3. 7.5 kJ **correct**
4. -2.5 kJ

Explanation:

$$T = 300 \text{ K}$$

$$\Delta n = -3 \text{ mol}$$

$$R = \frac{8.314 \text{ J}}{\text{mol} \cdot \text{K}}$$

$$\begin{aligned} w &= -P \Delta V \\ &= -\Delta n R T \\ &= -(-3 \text{ mol}) \left(\frac{8.314 \text{ J}}{\text{mol} \cdot \text{K}} \right) (300 \text{ K}) \\ &= 7482.6 \text{ J} \end{aligned}$$

002 10.0 points

The enthalpy of fusion of methanol (CH_3OH) is 3.16 kJ/mol. How much heat would be absorbed or released upon freezing 25.6 grams of methanol?

1. 2.52 kJ absorbed
2. 0.253 kJ absorbed
3. 2.52 kJ released **correct**
4. 3.95 kJ released
5. 3.95 kJ absorbed
6. 0.253 kJ released

Explanation:

$$\text{MW CH}_3\text{OH} = 32.042\text{g}$$

$$\Delta H(\text{fusion}) = -\Delta H(\text{freezing})$$

$$q = 25.6 \text{ g} \times \frac{1\text{mol}}{32.042\text{g}} \times \frac{-3.16 \text{ kJ}}{\text{mol}} = -2.525 \text{ kJ}$$

003 10.0 points

A 0.2 gram sample of a candy bar is combusted in a bomb calorimeter, increasing the temperature of the 2000 g of water from 25.00°C to 25.47°C. What is ΔU in kJ/g? Ignore any heat loss or gain by the calorimeter itself.

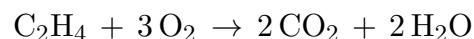
1. 19.6 kJ/g
2. -3.9 kJ/g
3. -0.08 kJ/g
4. -19.6 kJ/g **correct**
5. 3.9 kJ/g
6. 0.08 kJ/g

Explanation:

None

004 10.0 points

For the combustion reaction of ethylene (C_2H_4)



assume all reactants and products are gases, and calculate the ΔH_{rxn}^0 using bond energies.

1. 0 kJ/mol
2. 251 kJ/mol
3. 680 kJ/mol
4. -1300 kJ/mol **correct**
5. -251 kJ/mol
6. 1300 kJ/mol

7. -680 kJ/mol

Explanation:

$$\begin{aligned}\Delta H_{\text{rxn}}^0 &= \sum \text{BE}_{\text{reactants}} - \sum \text{BE}_{\text{products}} \\ &= [(C \equiv C) + 4(C - H) \\ &\quad + 3(O = O)] \\ &\quad - [4(C = O) + 4(H - O)] \\ &= \left[\left(602 \frac{\text{kJ}}{\text{mol}} \right) + 4 \left(413 \frac{\text{kJ}}{\text{mol}} \right) \right. \\ &\quad \left. + 3 \left(498 \frac{\text{kJ}}{\text{mol}} \right) \right] \\ &\quad - \left[4 \left(799 \frac{\text{kJ}}{\text{mol}} \right) + 4 \left(463 \frac{\text{kJ}}{\text{mol}} \right) \right] \\ &= -1300 \frac{\text{kJ}}{\text{mol}}\end{aligned}$$

005 10.0 points

Methyl tert-butyl ether or MTBE is an octane booster for gasoline. The combustion of 0.9211 grams of MTBE ($C_5H_{12}O(\ell)$, 88.15 g/mol) is carried out in a bomb calorimeter. The calorimeter's hardware has a heat capacity of 1.540 kJ/°C and is filled with exactly 2.022 L of water. The initial temperature was 26.336°C. After the combustion, the temperature was 29.849°C. Analyze this calorimeter data and determine the molar internal energy of combustion (ΔU) for this octane booster.

1. -3362 kJ/mol **correct**

2. -3120 kJ/mol

3. -3560 kJ/mol

4. -1957 kJ/mol

5. -2286 kJ/mol

6. -2748 kJ/mol

7. -4293 kJ/mol

Explanation:

$$\Delta T = 29.849 - 26.336 = 3.513^\circ$$

$$\begin{aligned}C_{\text{cal}} &= [2022(4.184) + 1540] / 1000 \\ &= 10.00 \text{ kJ}/^\circ\text{C}\end{aligned}$$

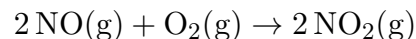
$$q_{\text{cal}} = 10(3.513) = 35.130 \text{ kJ}$$

$$\text{moles} = 0.9211 / 88.15 = 0.01045 \text{ mol}$$

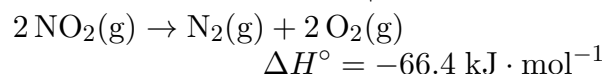
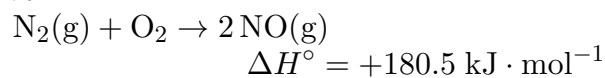
$$\begin{aligned}\Delta U &= -35.130 \text{ kJ} / 0.01045 \text{ mol} \\ &= -3362 \text{ kJ/mol}\end{aligned}$$

006 10.0 points

Calculate the standard reaction enthalpy for the oxidation of nitric oxide to nitrogen dioxide



given



1. -294.6 kJ · mol⁻¹

2. +114.1 kJ · mol⁻¹

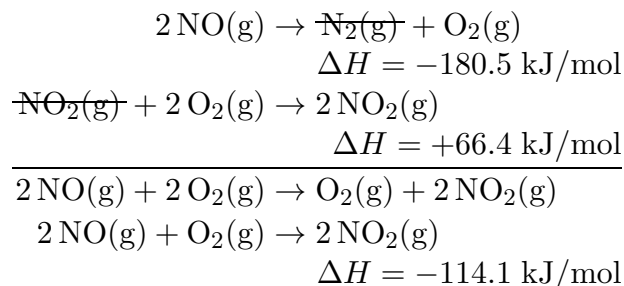
3. -114.1 kJ · mol⁻¹ **correct**

4. +246.9 kJ · mol⁻¹

5. -246.9 kJ · mol⁻¹

Explanation:

We need to reverse the reactions and add them:



007 10.0 points

You have a 12 oz. can (355 mL) of beer. You test the temperature and see that it reads 0°C. Now this isn't just any beer; this is Guinness and you've heard that Guinness is best at room temperature (20°C). If the specific heat

of Guinness is $4.186 \text{ J/g}\cdot^{\circ}\text{C}$, how much heat should you add in order to raise the temperature? The density of Guinness is 1.2 g/mL .

1. 33.6 kJ
2. 83 J
3. 33.6 J
4. 35.6 kJ **correct**

Explanation:

008 10.0 points

A student runs a reaction in a closed system. In the course of the reaction, 64.7 kJ of heat is released to the surroundings and 14.3 kJ of work is done on the system. What is the change in internal energy (ΔU) of the reaction?

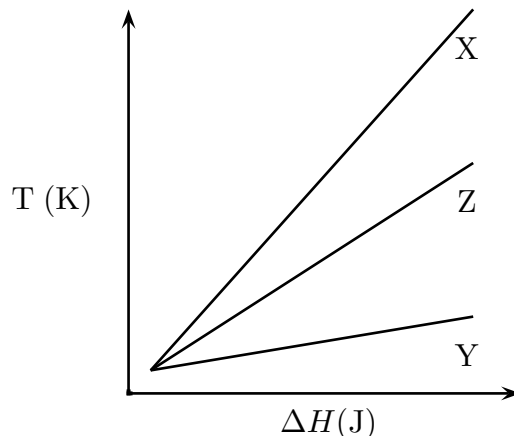
1. -79.0 kJ
2. 50.4 kJ
3. 79.0 kJ
4. -50.4 kJ **correct**
5. 90.4 kJ

Explanation:

The change in internal energy is given by the formula: $\Delta U = q + w$. In this reaction, q is -64.7 kJ and w is 14.3 kJ . The answer is -50.4 kJ .

009 10.0 points

Consider the plot below for three different samples of pure water.



Based on the plot, which answer choice below is a correct statement regarding the three samples of pure water?

1. All three samples have the same heat capacity.
2. Sample Z has the greatest heat capacity.
3. Sample X has the smallest mass. **correct**
4. All three samples have different specific heat capacities.
5. Sample Y would require the least heat to raise its temperature by 1 K .

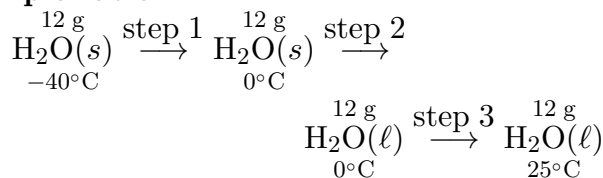
Explanation:

It is given that the samples are all pure water, which ensures that all three samples have the same specific heat capacity. Since the samples have different changes in T for a given input of energy, we can infer that all three samples have different heat capacities. Since sample X experiences the largest increase in T for a given amount of energy, it must be the least massive sample. Sample Z would actually have a heat capacity intermediate between that of X and Y. Sample Y requires the most heat to raise its temperature by any given amount.

010 10.0 points

What is the total heat flow when 12 grams of ice at -40°C are heated to become water at 25°C ?

1. 0.97 kJ
2. 2.26 kJ
3. 29.39 kJ
4. 4.01 kJ
5. 27.12 kJ
6. 6.27 kJ correct

Explanation:

$$\text{Step 1: } \frac{2.09 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \cdot (12 \text{ g}) \cdot (0 - (-40)^\circ\text{C}) = 1,003.2 \text{ J}$$

$$\text{Step 2: } \frac{334 \text{ J}}{\text{g}} \cdot (12 \text{ g}) = 4,008 \text{ J}$$

$$\text{Step 3: } \frac{4.184 \text{ J}}{\text{g} \cdot ^\circ\text{C}} \cdot (12 \text{ g}) \cdot (25 - 0)^\circ\text{C} = 1,255.2 \text{ J}$$

$$\text{Total} = 1,003.2 \text{ J} + 4,008 \text{ J} + 1,255.2 \text{ J} = 6,266.4 \text{ J} = 6.27 \text{ kJ}$$

011 10.0 points

A CD player and its battery together do 500 kJ of work, and the battery also releases 250 kJ of energy as heat and the CD player releases 50 kJ as heat due to friction from spinning. What is the change in internal energy of the system, with the system regarded as the battery and CD player together?

1. +200 kJ
2. -700 kJ
3. -750 kJ
4. -200 kJ
5. -800 kJ correct

Explanation:

Heat from the CD player is -50 kJ.

Heat from the battery is -500 kJ.

Work from both together on the surroundings is -250 kJ.

This question is testing your ability to see what the system is, and then look at ONLY the energy flow for the system. Here the system is the battery and the CD player together.

$$\begin{aligned} \Delta U &= q + w \\ &= [-50 \text{ kJ} + (-250 \text{ kJ})] + (-500 \text{ kJ}) \\ &= -800 \text{ kJ} \end{aligned}$$

012 10.0 points

3 g of a hydrocarbon fuel is burned in a bomb calorimeter that contains 200 grams of water initially at 25.00°C. After the combustion reaction, the temperature is 27.00°C. How much heat is evolved per gram of fuel burned? The heat capacity of the calorimeter (hardware only) is 150 J/°C.

1. 21220 J/g
2. 1673 J/g
3. 1973 J/g
4. 557 J/g
5. 7505 J/g
6. 7073 J/g
7. 300 J/g
8. 657 J/g correct

Explanation:

$$\begin{aligned} m_{\text{fuel}} &= 3 \text{ g} & m_{\text{water}} &= 200 \text{ g} \\ \Delta T &= 27.00^\circ\text{C} - 25.00^\circ\text{C} = 2.0^\circ\text{C} \end{aligned}$$

The amount of heat evolved by the reaction is equal to the amount of heat gained by the water plus the amount of heat gained by the calorimeter.

The specific heat of water is 4.184 J/g·°C, so we have to multiply by grams and temperature change in order to obtain Joules:

$$\Delta H \text{ of water} = (c_{\text{H}_2\text{O}})(m_{\text{water}})(\Delta T)$$

$$\begin{aligned}
 &= (4.184 \text{ J/g} \cdot ^\circ\text{C}) (200 \text{ g}) \\
 &\quad \times (2.00^\circ\text{C}) \\
 &= 1673 \text{ J}
 \end{aligned}$$

The heat capacity of the calorimeter is 150 J/°C. This is not per gram, so we just have to multiply by the temperature change to get Joules:

$$\begin{aligned}
 \Delta H \text{ of calorimeter} &= (c_{\text{cal}})(\Delta T) \\
 &= (150 \text{ J/}^\circ\text{C})(2.00^\circ\text{C}) \\
 &= 300 \text{ J}
 \end{aligned}$$

$$\begin{aligned}
 \Delta H \text{ of rxn} &= \Delta H \text{ of water} \\
 &\quad + \Delta H \text{ of calorimeter} \\
 &= 1673 \text{ J} + 300 \text{ J} = 1973 \text{ J}
 \end{aligned}$$

This is the total amount of heat evolved by the combustion of the 3 g of fuel. To get the amount per gram of fuel burned, we divide ΔH of the reaction by the amount of fuel burned (3 grams):

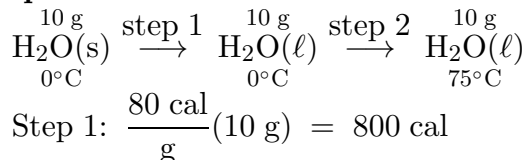
$$\frac{1973 \text{ J}}{3 \text{ g}} = 657 \text{ J/g}$$

013 10.0 points

The specific heat of water is 1.00 cal/g·°C, the heat of vaporization of water is 540 cal/g, and the heat of fusion of water is 80 cal/g. How much heat would be required to convert 10 grams of ice at 0°C to 10 grams of water at 75°C?

- 15.5 cal
- 6150 cal
- 155 cal
- 1.55 kcal **correct**
- 61.5 kcal

Explanation:



$$\text{Step 2: } \frac{1 \text{ cal}}{\text{g} \cdot ^\circ\text{C}}(10 \text{ g})(75 - 0)^\circ\text{C} = 750 \text{ cal}$$

$$\begin{aligned}
 \text{Total} &= 800 \text{ cal} + 750 \text{ cal} \\
 &= 1550 \text{ cal} \\
 &= 1.55 \text{ kcal}
 \end{aligned}$$

014 10.0 points

1 g of cake is combusted in a bomb calorimeter. The heat capacity of the calorimeter hardware is 12 calories · K⁻¹. The calorimeter contains 4 L of water; the specific heat capacity of water is 1 calorie · g⁻¹ · K⁻¹ and the density of water is 1 g · mL⁻¹. You detonate the cake and the temperature of the water increases by 1.2 K. Calculate the calories in the one-gram sample of cake, ΔU .

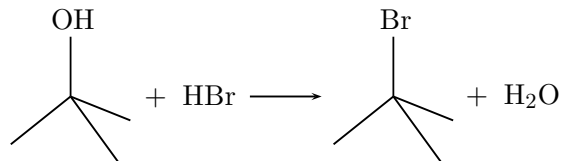
- 4814.4 calories **correct**
- 1150.7 calories
- 20083.2 calories
- 20143.4 calories
- 4800.0 calories
- 1147.2 calories

Explanation:

$$1 \text{ calorie} \cdot \text{g}^{-1} \cdot \text{K}^{-1} \times 4000 \text{ ml} \times 1 \text{ g} \cdot \text{mL}^{-1} \times 1.2 \text{ K}$$

015 10.0 points

Reaction of tertiary butyl alcohol with hydrobromic acid produces tertiary butyl bromide by the following reaction. Use bond energies (provided in preamble) to estimate the change in enthalpy, ΔH , for this reaction.



- +105 kJ/mol
- +186 kJ/mol
- +24 kJ/mol

4. -105 kJ/mol 5. -24 kJ/mol **correct**6. -186 kJ/mol **Explanation:**

In this reaction, you break two bonds in the reactants: a C-O bond and a H-Br bond. You form new bonds in the products: a C-Br bond and a H-O bond.

The change in enthalpy is the energy in for breaking the bonds combined with the energy out gained from forming the new bonds.

$$\Delta H_{\text{rxn}} = \text{B.E.}_{\text{C-O}} + \text{B.E.}_{\text{H-Br}} - \text{B.E.}_{\text{C-Br}} - \text{B.E.}_{\text{H-O}}$$

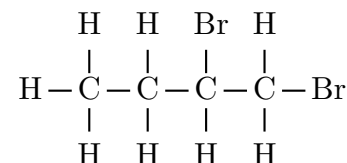
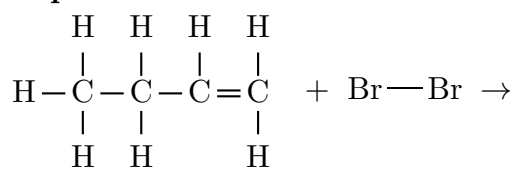
$$\Delta H_{\text{rxn}} = 358 + 366 - 285 - 463$$

$$\Delta H_{\text{rxn}} = -24 \text{ kJ mol}^{-1}$$

016 10.0 points

Estimate the heat released when 1-butene ($\text{CH}_3\text{CH}_2\text{CH}=\text{CH}_2$) reacts with bromine to give $\text{CH}_3\text{CH}_2\text{CHBrCH}_2\text{Br}$. Bond enthalpies are

C—H : 412 kJ/mol ; C—C : 348 kJ/mol ;
 C=C : 612 kJ/mol ; C—Br : 276 kJ/mol ;
 Br—Br : 193 kJ/mol .

1. 317 kJ/mol 2. 288 kJ/mol 3. 181 kJ/mol 4. 507 kJ/mol 5. 95 kJ/mol **correct****Explanation:**

$$\begin{aligned} \Delta H &= \sum E_{\text{break}} - \sum E_{\text{make}} \\ &= \left[(\text{C}=\text{C}) + (\text{Br}-\text{Br}) \right] \\ &\quad - \left[2(\text{C}-\text{Br}) + (\text{C}-\text{C}) \right] \\ &= 612 \text{ kJ/mol} + 193 \text{ kJ/mol} \\ &\quad - \left[2(276 \text{ kJ/mol}) + 348 \text{ kJ/mol} \right] \\ &= -95 \text{ kJ/mol}, \end{aligned}$$

which means 95 kJ/mol of heat was released.

017 10.0 points

Which of the following is/are a reason that water is a desirable heat sink for use in calorimeters?

- I) Water's heat specific capacity is very precisely known.
- II) Water is readily available.
- III) Water has an unusually large specific heat capacity.

1. I only

2. II and III

3. I and II

4. I, II and III **correct**

5. II only

6. I and III

7. III only

Explanation:

Water is a good heat sink for all of the reasons listed above. Moreover, its large heat capacity, liquid state and ready availability enable us to easily set up a calorimeter such that ΔT is large enough that it can be easily measured and small enough that phase transition temperatures are not reached.

018 10.0 points

Consider a thermodynamic system that is simultaneously releasing heat and doing work. The internal energy of this system will:

1. Decrease **correct**
2. Increase, decrease, or stay the same depending on the magnitudes of heat and work
3. Stay exactly the same.
4. Increase

Explanation:

The change in internal energy is equal to the sum of the heat absorbed by the system and work done on the system based on the equation: $\Delta U = q + w$. In this case, q and w are both negative. Therefore the internal energy will be decreasing regardless of the magnitudes of heat and work.

019 10.0 points

Which of the following statements is/are true?

- I) For a given process, ΔH must be zero when external pressure is zero.
- II) For a given process, ΔU and ΔH must have different values.
- III) For a given process, ΔU_{sys} and ΔU_{surr} must have the same magnitude.

1. I, II
2. I, II, III
3. I only
4. III only **correct**
5. II, III
6. II only
7. I, III

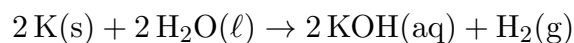
Explanation:

ΔH is heat transferred at constant pressure. For a given process ΔU and ΔH do

not always have different values; for example, they can both be zero for a given process. Because ΔU_{univ} must be zero according to the 1st law, ΔU of a system and its surroundings not only can but must have the same magnitude and opposite signs.

020 10.0 points

If you drop a piece of potassium metal into water you get the following exothermic reaction:



What are the values of q and w for this reaction, at constant temperature and pressure?

1. Both are positive.
2. q is negative and w is positive.
3. q is positive and w is negative.
4. Both are negative. **correct**

Explanation:

021 10.0 points

The formation of chemical bonds from separated atoms

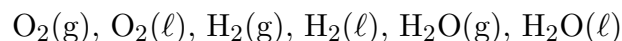
1. is never spontaneous.
2. increases entropy.
3. may be either endothermic or exothermic.
4. is always exothermic. **correct**
5. is always endothermic.

Explanation:

The energies to break bonds are positive (require energy input, positive ΔH) and energies to form bonds are negative (release energy, negative ΔH).

022 10.0 points

Which of



have a standard enthalpy of formation equal to zero?

1. $\text{O}_2(\text{g})$, $\text{O}_2(\ell)$, $\text{H}_2(\text{g})$, $\text{H}_2(\ell)$, $\text{H}_2\text{O}(\text{g})$, $\text{H}_2\text{O}(\ell)$
2. $\text{O}_2(\text{g})$, $\text{H}_2(\text{g})$, $\text{H}_2\text{O}(\text{g})$
3. $\text{O}_2(\text{g})$, $\text{O}_2(\ell)$, $\text{H}_2(\text{g})$, $\text{H}_2(\ell)$
4. $\text{O}_2(\text{g})$, $\text{H}_2(\text{g})$ **correct**
5. All of them, but only at absolute zero

Explanation:

Molecules in their native state at STP have a standard enthalpy of formation of zero.

023 10.0 points

When 1 mol of methane is burned at constant pressure, -890 kJ/mol of energy is released as heat. If a 3.64 g sample of methane is burned at constant pressure, what will be the value of ΔH ? (Hint: Convert the grams of methane to moles. Also make sure your answer has the correct sign for an exothermic process.)

1. -61.1875
2. -202.475
3. -176.888
4. -268.669
5. -257.544
6. -233.625
7. -132.387
8. -140.731
9. -264.219
10. -115.144

Correct answer: -202.475 kJ .

Explanation:

CH_4 : $\Delta H_c = -890 \text{ kJ/mol}$ $m = 3.64 \text{ g}$
 $\Delta H = ?$

$$n_{\text{CH}_4} = (3.64 \text{ g CH}_4) \left(\frac{\text{mol CH}_4}{16 \text{ g CH}_4} \right)$$

$$= 0.2275 \text{ mol CH}_4$$

$$\Delta H = (0.2275 \text{ mol})(-890 \text{ kJ/mol}) = -202.475 \text{ kJ}$$

024 10.0 points

Calculate the quantity of energy required to change 3.00 mol of liquid water at 100°C to steam at 100°C . The molar heat of vaporization of water is 40.6 kJ/mol .

1. 300 kJ
2. 122 kJ **correct**
3. None of these
4. 40.6 kJ
5. 13.5 kJ

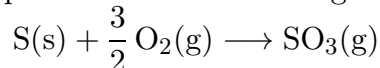
Explanation:

$n = 3.0 \text{ mol}$

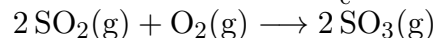
$$\frac{40.6 \text{ kJ}}{\text{mol}} \cdot 3 \text{ mol} = 121.8 \text{ kJ}$$

025 10.0 points

Calculate the heat of formation for 2.6 mol of sulfur dioxide (SO_2) from its elements, sulfur and oxygen. Use the balanced chemical equation and the following information.



$$\Delta H_c^0 = -395.2 \text{ kJ/mol}$$



$$\Delta H^0 = -198.2 \text{ kJ/mol}$$

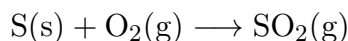
1. -414.54
2. -562.59
3. -384.93
4. -769.86
5. -592.2
6. -503.37
7. -651.42
8. -710.64
9. -621.81
10. -532.98

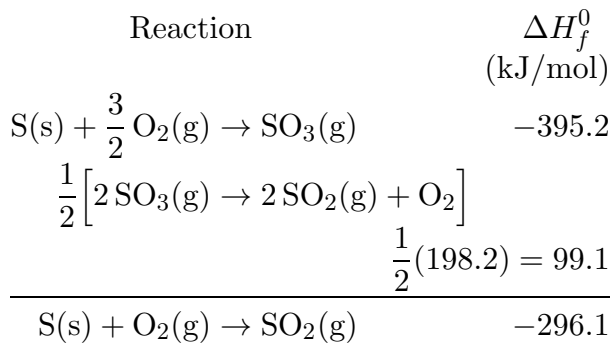
Correct answer: -769.86 kJ .

Explanation:

ΔH_f of SO_2 from S and O = ?

The balanced equation is





For 2.6 mol sulfur dioxide, the energy required is

$$(-296.1 \text{ kJ/mol})(2.6 \text{ mol}) = -769.86 \text{ kJ}$$

026 10.0 points

? heat capacity is the amount of heat required to raise the temperature of one ? of an object by 1°C. It is an ? property.

1. Specific; gram; extensive
2. Molar; gram; intensive
3. Specific; mole; extensive
4. Molar; gram; extensive
5. Specific; gram; intensive **correct**
6. Molar; mole; extensive

Explanation:

None

027 (part 1 of 3) 10.0 points

Consider the following chemical and physical changes:

- A. $H_2O(l) \rightarrow H_2O(g)$
- B. $H_2O(l) \rightarrow H_2O(s)$
- C. $\frac{1}{2} H_2(g) + \frac{1}{2} F_2(g) \rightarrow HF(g)$
 $\Delta H = -271.1 \text{ kJ/mol}$
- D. $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
- E. $N_2(g) + O_2(g) \rightarrow 2NO(g)$
 $\Delta H = 180.50 \text{ kJ/mol}$

Which change(s) are endothermic?

1. C and D only
2. B and E only
3. C only
4. A, D, and E only
5. A, C, and D only
6. A and E only **correct**

Explanation:

None

028 (part 2 of 3) 10.0 points

For which change(s) would $\Delta H = \Delta U$?

1. B and E only
2. C and D only
3. A and B only
4. A and D only
5. B, C, and E only **correct**

Explanation:

None

029 (part 3 of 3) 10.0 points

For which change(s) would $\Delta H_{rxn} = \Delta H_f$ of the product?

1. A, B, and C only
2. C only **correct**
3. A and C only
4. A, B, C, and E only
5. C and E only

Explanation:

None

030 10.0 points

The standard enthalpy of formation of $\text{Br}_2(\ell)$ is

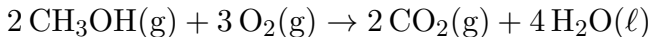
1. negative.
2. zero. **correct**
3. positive.

Explanation:

Since bromine is a liquid at 25°C and 1 atm pressure, it is already in its standard state, so the standard enthalpy of formation is zero.

031 10.0 points

Consider the combustion reaction below.



If this reaction took place in a closed, rigid container, work would be (positive/negative/zero) and heat would be (positive/negative/zero).

1. positive, zero
2. positive, negative
3. zero, positive
4. negative, positive
5. negative, zero
6. zero, negative **correct**

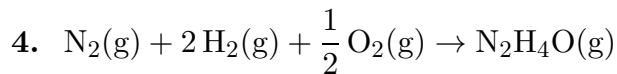
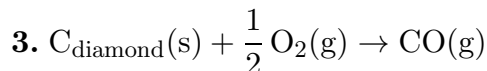
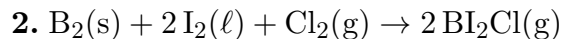
Explanation:

Despite the fact that the reaction results in a net decrease in moles of gas, which would normally result in positive work (work done on the system), the closed, rigid container precludes a change in volume and results in zero work. Like other combustion reactions, heat is released, and so heat would be negative.

032 10.0 points

Which of the reactions below is a formation reaction?

1. $2 \text{Fe}(\text{s}) + 3 \text{O}(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s})$



correct

Explanation:

A formation reaction produces exactly one mole of one product from elements in their standard states.

033 10.0 points

Energy in the amount of 455 J is added to a 67.0 g sample of water at a temperature of 7.00°C . What will be the final temperature of the water?

1. 26.7039
2. 8.62465
3. 15.1616
4. 27.1092
5. 17.8501
6. 30.404
7. 29.7016
8. 13.5327
9. 15.6545
10. 3.29054

Correct answer: 8.62465°C .

Explanation:

$$q = 455 \text{ J}$$

$$m = 67 \text{ g}$$

$$SH_{\text{H}_2\text{O}} = 4.18 \text{ J/g} \cdot \text{K}$$

$$T_i = 7^\circ\text{C}$$

$$T_f = ?$$

$$\begin{aligned} SH_{\text{H}_2\text{O}} &= \frac{q}{m \Delta T} \\ \Delta T &= \frac{q}{m SH_{\text{H}_2\text{O}}} \\ &= \frac{455 \text{ J}}{(67 \text{ g})(4.18 \text{ J/g} \cdot \text{K})} \\ &= 1.62465 \text{ K} \\ &= 1.62465^\circ\text{C} \end{aligned}$$

$$\Delta T = T_f - T_i$$

$$T_f = \Delta T + T_i$$

$$= 1.62465^\circ\text{C} + 7^\circ\text{C}$$

$$= 8.62465^\circ\text{C}$$

034 10.0 points

A system did 150 kJ of work and its internal energy increased by 60 kJ. How much energy did the system gain or lose as heat?

1. The system gained 60 kJ of energy as heat.
2. The system gained 90 kJ of energy as heat.
3. The system lost 210 kJ of energy as heat.
4. The system lost 90 kJ of energy as heat.
5. The system gained 210 kJ of energy as heat. **correct**

Explanation:

035 10.0 points

An important reaction that takes place in the atmosphere is



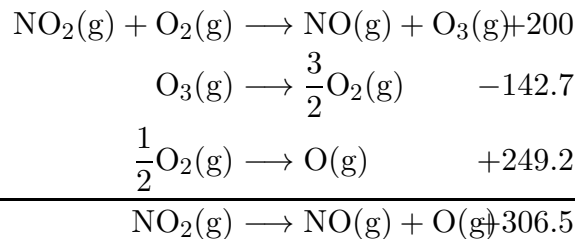
which is brought about by sunlight. Calculate the standard enthalpy of the reaction from the following information

reaction	ΔH° (kJ)
$\text{O}_2(\text{g}) \rightarrow 2 \text{O}(\text{g})$	+498.4
$\text{NO}(\text{g}) + \text{O}_3(\text{g}) \longrightarrow \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$	-200.0
$\frac{3}{2} \text{O}_2(\text{g}) \longrightarrow \text{O}_3(\text{g})$	+142.7

1. 306.5 kJ **correct**
2. 820.5 kJ
3. 320.2 kJ
4. 555.7 kJ
5. 963.8 kJ
6. 106.5 kJ
7. 449.2 kJ

Explanation:

Using Hess' Law we add the reverse (flip) of reaction 2; the reverse (flip) of reaction 3; and one half of reaction 1:

**036 10.0 points**

What is the value of work when an external pressure of 2 atm compresses a piston from an initial volume of 11.2 liters to a final volume of 2 liters.

1. -18.4 kJ
2. 18.4 kJ
3. 1.86 kJ **correct**
4. -1.86 kJ

Explanation:

For expansion against a constant external pressure,

$$\begin{aligned} w &= -P_{\text{ext}} \Delta V \\ &= (-2 \text{ atm})(2 \text{ L} - 11.2 \text{ L}) \\ &\quad \times (101.325 \text{ J} \cdot \text{L}^{-1} \cdot \text{atm}^{-1}) \\ &= 1.86 \text{ kJ}. \end{aligned}$$

037 10.0 points

A bomb calorimeter with a heat capacity of 30 J/C contains 1000 g of water with an initial temperature of 25°C. A 0.5 g sample of a candy bar is placed in a bomb calorimeter and ignited, resulting in a new water temperature of 30°C. What is ΔE for this reaction?

1. -42 kJ/g **correct**
2. 0 kJ/g

3. +21 kJ/g

4. -300 kJ/g

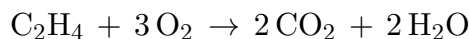
5. +300 kJ/g

6. -21 kJ/g

7. +42 kJ/g

Explanation:**038 10.0 points**

For the combustion reaction of ethylene (C₂H₄)



assume all reactants and products are gases, and calculate the ΔH_{rxn}^0 using bond energies from the list below.

C—H : 413 kJ/mol; H—O : 463 kJ/mol;

O—O : 146 kJ/mol; O=O : 498 kJ/mol.

C—C : 346 kJ/mol; C=C : 602 kJ/mol.

C—O : 358 kJ/mol; C=O : 799 kJ/mol.

1. 0 kJ/mol

2. -251 kJ/mol

3. -680 kJ/mol

4. -1300 kJ/mol **correct**

5. 1300 kJ/mol

6. 680 kJ/mol

7. 251 kJ/mol

Explanation:

$$\begin{aligned}\Delta H_{\text{rxn}}^0 &= \sum \text{BE}_{\text{reactants}} - \sum \text{BE}_{\text{products}} \\ &= [(C=C) + 4(C-H) \\ &\quad + 3(O=O)] \\ &\quad - [4(C=O) + 4(H-O)] \\ &= \left[\left(602 \frac{\text{kJ}}{\text{mol}} \right) + 4 \left(413 \frac{\text{kJ}}{\text{mol}} \right) \right.\end{aligned}$$

$$\begin{aligned}&\quad \left. + 3 \left(498 \frac{\text{kJ}}{\text{mol}} \right) \right] \\ &\quad - \left[4 \left(799 \frac{\text{kJ}}{\text{mol}} \right) + 4 \left(463 \frac{\text{kJ}}{\text{mol}} \right) \right] \\ &= -1300 \frac{\text{kJ}}{\text{mol}}\end{aligned}$$

039 10.0 points

2.26 g of liquid water at 23.5 °C was completely converted to ice at 0 °C. How much heat was (absorbed/released) by the system during this process?

1. 1478 J; absorbed

2. 755 J; absorbed

3. 1478 J; released

4. 977 J; absorbed

5. 977 J; released **correct**

6. 755 J; released

Explanation:

for 1 gram (cooling + freezing):

$$23.5(4.184) + 334 = 432.324 \text{ J/g}$$

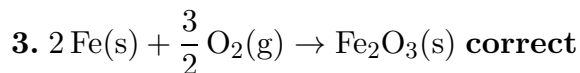
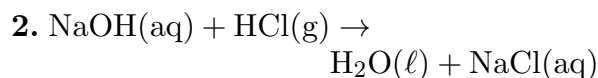
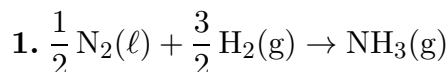
scale up to 2.26 g :

$$432.324(2.26) = 977.052 \text{ J}$$

= 977 J released

040 10.0 points

Which of the following reactions is an enthalpy of formation reaction?

**Explanation:**

Formation reactions describe production of exactly one mole of one product from stoichiometric quantities of elements in their standard states.

041 10.0 points

Consider a system where 2.50 L of ideal gas expands to 6.25 L against a constant external pressure of 330 torr. Calculate the work (w) for this system.

1. -1238 J
2. +1238 J
3. -1.63 J
4. +165 J
5. +1.63 J
6. -165 J **correct**

Explanation:

Convert torr to atm, and then convert answer in L·atm to joules. The answer will be negative due to expansion of the gas.

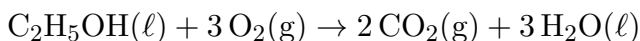
$$w = -P\Delta V = -(330/760)(3.75L)$$

$$w = -1.628 \text{ L atm}$$

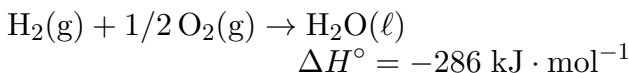
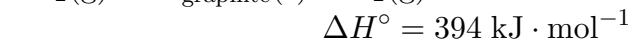
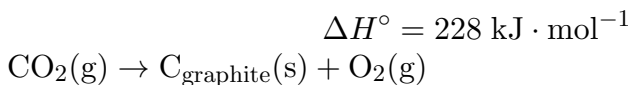
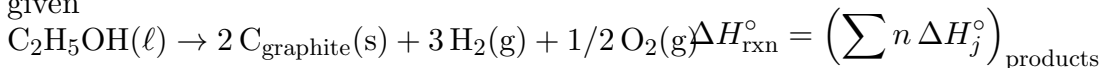
$$\times 101.325 \text{ J/(L atm)} = -165 \text{ J}$$

042 10.0 points

Calculate the standard reaction enthalpy for the reaction



given



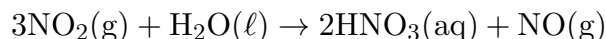
1. 730 kJ · mol⁻¹
2. -846 kJ · mol⁻¹
3. -1,418 kJ · mol⁻¹ **correct**
4. 336 kJ · mol⁻¹
5. -452 kJ · mol⁻¹

Explanation:

We need to reverse and double the second reaction as well as triple the third reaction and then add all three reactions.

043 10.0 points

Calculate the standard reaction enthalpy ($\Delta H_{\text{rxn}}^\circ$) for the final stage in the production of nitric acid, when nitrogen dioxide dissolves in and reacts with water:



1. -370 kJ
2. +70 kJ
3. -104 kJ
4. +136 kJ
5. -304 kJ
6. -137 kJ **correct**

Explanation:

Values for ΔH_f° from external table are in order (from reaction) +33, -286, -207, and +90

$$\begin{aligned} \Delta H_{\text{rxn}}^\circ &= \left(\sum n \Delta H_j^\circ \right)_{\text{products}} \\ &\quad - \left(\sum n \Delta H_j^\circ \right)_{\text{reactants}} \\ &= \left[2 \Delta H_{\text{f, HNO}_3(\text{aq})}^\circ + \Delta H_{\text{f, O}(\text{g})}^\circ \right] \\ &\quad - \left[3 \Delta H_{\text{f, NO}_2(\text{g})}^\circ + \Delta H_{\text{f, H}_2\text{O}(\ell)}^\circ \right] \\ &= \left[2(-207) + 90 \right] - \left[3(33) + (-286) \right] \\ &= -137 \text{ kJ} \end{aligned}$$

044 10.0 points

The molar heat capacity of $\text{C}_6\text{H}_6(\ell)$ is 136 J/mol · °C and of $\text{C}_6\text{H}_6(\text{g})$ is 81.6 J/mol · °C. The molar heat of fusion for benzene is 9.92 kJ/mol and its molar heat of vaporization is

30.8 kJ/mol. The melting point of benzene is 5.5°C, its boiling point is 80.1°C, and its molecular weight 78.0 g/mol. How much heat would be required to convert 234 g of solid benzene (C₆H₆(s)) at 5.5°C into benzene vapor (C₆H₆(g)) at 100.0°C?

1. 97.2715 kJ
2. 157.468 kJ **correct**
3. 4931.72 kJ
4. 60.1968 kJ
5. 152.597 kJ

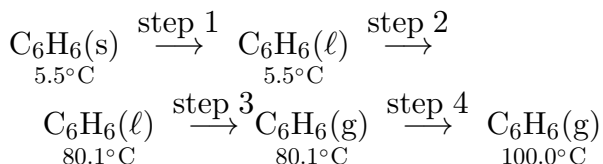
Explanation:

$$m_{\text{benzene}} = 234 \text{ g}$$

$$T_2 = 100.0^\circ\text{C}$$

$$T_1 = 5.5^\circ\text{C}$$

$$234 \text{ g} \times \frac{\text{mol}}{78.0 \text{ g}} = 3 \text{ mol}$$



$$\text{Step 1 : } \frac{9.92 \text{ kJ}}{\text{mol}} \times 3 \text{ mol} = 29.76 \text{ kJ}$$

$$\text{Step 2 : } \frac{136 \text{ J}}{\text{mol} \cdot ^\circ\text{C}} \times (3 \text{ mol}) \times (80.1 - 5.5)^\circ\text{C}$$

$$= 30436.8 \text{ J} = 30.4368 \text{ kJ}$$

$$\text{Step 3 : } \frac{30.8 \text{ kJ}}{\text{mol}} \times (3 \text{ mol}) = 92.4 \text{ kJ}$$

$$\text{Step 4 : } \frac{81.6 \text{ J}}{\text{mol} \cdot ^\circ\text{C}} \times (3 \text{ mol})$$

$$\times (100.0 - 80.1)^\circ\text{C}$$

$$= 4871.52 \text{ J} = 4.87152 \text{ kJ}$$

$$\text{Total} = 29.76 \text{ kJ} + 30.4368 \text{ kJ}$$

$$+ 92.4 \text{ kJ} + 4.87152 \text{ kJ}$$

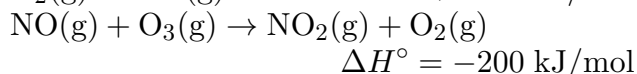
$$= 157.468 \text{ kJ}$$

045 10.0 points

Calculate the standard reaction enthalpy for the reaction



given +142.7 kJ/mol for the standard enthalpy of formation of ozone and



Remember the definition of the standard enthalpy of formation of a substance.

1. +306 kJ/mol **correct**

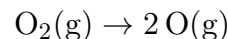
2. +355 kJ/mol

3. +192 kJ/mol

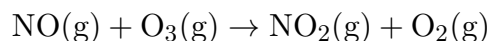
4. +592 kJ/mol

5. +555 kJ/mol

Explanation:

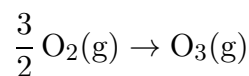


$$\Delta H^\circ = +498.4 \text{ kJ/mol}$$



$$\Delta H^\circ = -200 \text{ kJ/mol}$$

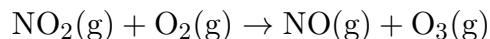
The standard formation of ozone is



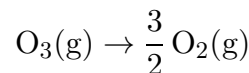
$$\Delta H = +142.7 \text{ kJ/mol}$$

We calculate the ΔH_{rxn} using Hess' Law:

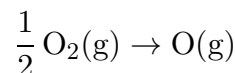
To combine the reactions and get the desired reaction, reverse the second and third equations and add half of the first one:



$$\Delta H = +200 \text{ kJ/mol}$$



$$\Delta H = -142.7 \text{ kJ/mol}$$



$$\Delta H = \frac{1}{2}(498.4 \text{ kJ/mol})$$



$$\Delta H_{\text{rxn}} = 306.5 \text{ kJ/mol}$$

046 10.0 points

A coffee cup calorimeter measures the heat at constant ? whereas a bomb calorimeter measures the heat at constant ?

1. pressure ($q_p = \Delta H$); volume ($q_v = \Delta U$)
correct

2. pressure ($q_p = \Delta U$); volume ($q_v = \Delta H$)

3. volume ($q_v = \Delta H$); pressure ($q_p = \Delta U$)

4. volume ($q_v = \Delta U$); pressure ($q_p = \Delta H$)

Explanation:

None

047 10.0 points

You have two liquids of identical mass, and both with initial temperatures of 15°C . One is ethanol, $\text{C}_2\text{H}_5\text{OH}$, with a specific heat of $2.46 \text{ J/g}^\circ\text{C}$ and the other is benzene, C_6H_6 , with a specific heat of $1.74 \text{ J/g}^\circ\text{C}$. If both liquids absorb the same amount of heat, which one will have the highest final temperature? Assume that neither liquid reaches its boiling point.

1. Cannot tell without more information given.

2. ethanol

3. Both liquids will have the same final temperature.

4. benzene **correct**

Explanation:

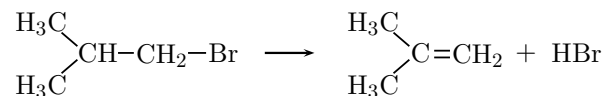
Temperature rise (ΔT) is inversely proportional to the heat capacity.

$$\Delta T = \frac{q}{mC_s}$$

Therefore, because benzene has a smaller heat capacity, C_s , it will have the larger temperature rise.

048 10.0 points

1-bromo-isobutane will undergo an elimination reaction to yield isobutene and hydrogen bromide as shown in the reaction below. Use bond energies (provided in preamble) to estimate the change in enthalpy, ΔH , for this gas phase reaction.



1. -270 kJ/mol

2. $+270 \text{ kJ/mol}$

3. -76 kJ/mol

4. $+337 \text{ kJ/mol}$

5. $+76 \text{ kJ/mol}$ **correct**

6. -337 kJ/mol

Explanation:

In this reaction, you break three bonds in the reactants: a C-C, a C-H, and a C-Br bond. You form new bonds in the products: a H-Br bond and a C=C bond.

The change in enthalpy is the energy in for breaking the bonds combined with the energy out gained from forming the new bonds.

$$\Delta H_{\text{rxn}} = 346 + 413 + 285 - 366 - 602$$

$$\Delta H_{\text{rxn}} = +76 \text{ kJ mol}^{-1}$$

049 10.0 points

Which is true, considering the first law of thermodynamics?

1. $\Delta U = q - w$, where heat and work can both be positive for the same process

2. $\Delta U = q + w$, where heat and work can never both be positive for the same process

3. $\Delta U = q - w$, where heat and work can never both be positive for the same process

4. $\Delta U = q + w$, where heat and work can both be positive for the same process. **cor-**

rect

Explanation:

none

050 10.0 points

For which of the following chemical equations would $\Delta H_{\text{rxn}}^{\circ} = \Delta H_{\text{f}}^{\circ}$?

- $\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\ell)$ **correct**
- $\text{C}(\text{s, graphite}) + \frac{3}{2}\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
- $\text{CO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$
- $\text{N}_2(\ell) + 3 \text{F}_2(\text{g}) \rightarrow 2 \text{NF}_3(\ell)$

Explanation:

For $\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\ell)$, $\Delta H_{\text{f}}^{\circ}$ of $\text{O}_2(\text{g})$ and $\text{H}_2(\text{g})$ are 0. Therefore, $\Delta H_{\text{rxn}}^{\circ} = \Delta H_{\text{f}}^{\circ}(\text{H}_2\text{O}_2(\ell))$

051 10.0 points

The combustion of methane gas (CH_4) forms $\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$. Calculate the heat produced by burning 1.98 mol of the methane gas. Use these $\Delta H_{\text{f}}^{\circ}$ data to help:

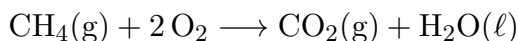
$\text{CH}_4(\text{g}) = -74.9 \text{ kJ/mol}$
 $\text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$
 $\text{H}_2\text{O}(\ell) = -285.8 \text{ kJ/mol}$.

- 1566.75
- 1513.34
- 1290.79
- 1459.93
- 1726.99
- 1175.06
- 1424.32
- 1201.77
- 1121.65
- 1762.6

Correct answer: 1762.6 kJ.

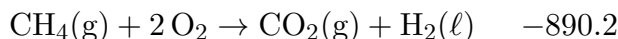
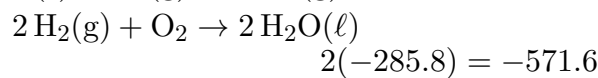
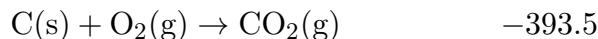
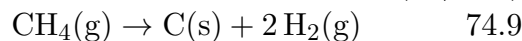
Explanation:

$n_{\text{CH}_4} = 1.98 \text{ mol}$



ΔH° for combustion of $\text{CH}_4 = ?$

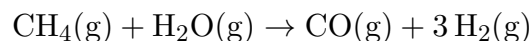
Reaction

 $\Delta H_{\text{f}}^{\circ}$
(kJ/mol)

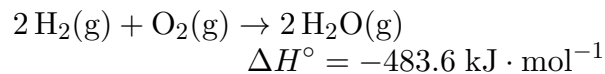
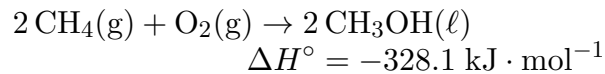
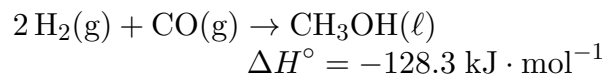
$$\Delta H^{\circ} = (890.2 \text{ kJ/mol}) (1.98 \text{ mol}) = 1762.6 \text{ kJ}$$

052 10.0 points

Calculate the standard reaction enthalpy for the reaction.



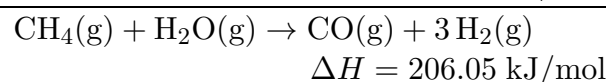
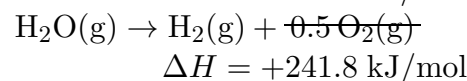
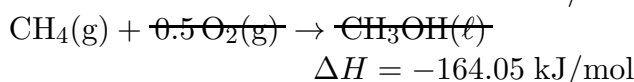
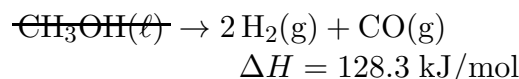
given



- $+155.5 \text{ kJ} \cdot \text{mol}^{-1}$
- $+206.1 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**
- $+216 \text{ kJ} \cdot \text{mol}^{-1}$
- $+412.1 \text{ kJ} \cdot \text{mol}^{-1}$
- $+42.0 \text{ kJ} \cdot \text{mol}^{-1}$

Explanation:

We need to reverse the first reaction, halve the second, halve and reverse the third and add the results:



053 10.0 points

A system absorbs 237 J of heat while it performs 435 J of work. What is the change in the internal energy of the system?

1. 672 J
2. 198 J
3. -198 J **correct**
4. -672 J

Explanation:

$$\Delta U = q + w = 237 \text{ J} +^{-} 435 \text{ J} = -198 \text{ J}$$

054 10.0 points

Calculate the enthalpy change that occurs when 1.00 kg of acetone condenses at its boiling point (329.4 K). The standard enthalpy of vaporization of acetone is $29.1 \text{ kJ} \cdot \text{mol}^{-1}$.

1. -29.1 kJ
2. +29.1 kJ
3. +501 kJ
4. -501 kJ **correct**
5. $-2.91 \times 10^4 \text{ kJ}$

Explanation:

Acetone is CH_3COCH_3
MM = 58.0798 g/mol

$$n = \frac{(1000 \text{ g})(1 \text{ mol acetone})}{58.0798 \text{ g/mol}} = 17.2177 \text{ mol}$$

Condensation is the opposite of vaporization:

$$\begin{aligned} \Delta H &= q = n \Delta H_{\text{cond}} \\ &= n(-\Delta H_{\text{vap}}) \\ &= (17.2177 \text{ mol})(-29.1 \text{ kJ/mol}) \\ &= -501.035 \text{ kJ} \end{aligned}$$

055 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?

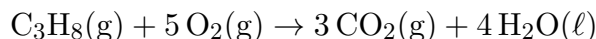
1. $\text{N}_2\text{H}_2(\text{g}) + \text{CH}_3\text{OH}(\text{g}) \rightarrow \text{CH}_2\text{O}(\text{g}) + \text{N}_2(\text{g}) + 2 \text{H}_2(\text{g})$
2. $\text{CH}_2\text{O}(\text{g}) + \text{N}_2(\text{g}) + 2 \text{H}_2(\text{g}) \rightarrow \text{N}_2\text{H}_2(\text{g}) + \text{CH}_3\text{OH}(\text{g})$ **correct**
3. $2 \text{H}_2\text{O}(\ell) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}_2(\ell)$
4. $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$
5. $2 \text{H}_2\text{O}_2(\ell) \rightarrow 2 \text{H}_2\text{O}(\ell) + \text{O}_2(\text{g})$
6. $\text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g}) \rightarrow \text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g})$

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 $5.0 \text{ kJ} = -\Delta n_{\text{gas}} (2.5 \text{ kJ} \cdot \text{mol}^{-1})$, the reaction for which $\Delta n_{\text{gas}} = -2$ will be the correct answer.

056 10.0 points

The value of ΔH for the reaction



is -2220 kJ/mol rxn . How much heat is given off when 33.0 g of propane gas (C_3H_8) is burned at constant pressure?

1. 1665 kJ **correct**
2. 22420 kJ
3. 2220 kJ
4. 25.96 kJ
5. 555 kJ
6. 50.5 kJ

7. 6660 kJ

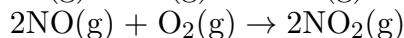
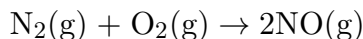
Explanation:

$$\Delta H = -2220 \text{ kJ/mol} \quad m_{\text{C}_3\text{H}_8} = 33.0 \text{ g}$$

$$q = \left(\frac{-2220 \text{ kJ}}{\text{mol rxn}} \right) \left(\frac{1 \text{ mol rxn}}{1 \text{ mol C}_3\text{H}_8} \right) \\ \times \left(\frac{1 \text{ mol C}_3\text{H}_8}{44 \text{ g C}_3\text{H}_8} \right) (33 \text{ g C}_3\text{H}_8) \\ = -1665 \text{ kJ} \quad \text{or} \quad 1665 \text{ kJ released}$$

057 10.0 points

The two reactions shown below are both endothermic. For which reaction is $\Delta H < \Delta U$?



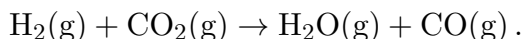
1. $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$ **correct**
2. Neither reaction has $\Delta H < \Delta U$.
3. Both reactions have $\Delta H < \Delta U$.
4. $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g})$

Explanation:

.

058 10.0 points

Consider the following reaction



ΔH_f for $\text{CO}_2(\text{g})$ is -22.5 kJ/mol ;

ΔH_f for $\text{CO}(\text{g})$ is -6.3 kJ/mol ;

ΔH_f for $\text{H}_2\text{O}(\text{g})$ is -13.8 kJ/mol .

1. ΔH of the reaction is negative.
2. ΔH of the reaction is zero.
3. ΔH of the reaction is positive. **correct**

Explanation:

Reactants:

$$\Delta H_f \text{ CO}_2(\text{g}) = -22.5 \text{ kJ/mol}$$

Products:

$$\Delta H_f \text{ CO}(\text{g}) = -6.3 \text{ kJ/mol}$$

$$\Delta H_f \text{ H}_2\text{O}(\text{g}) = -13.8 \text{ kJ/mol}$$

$$\Delta H_{\text{rxn}}^0 = \sum n H_{\text{f prod}}^0 - \sum n H_{\text{f rct}}^0 \\ = \left(-13.8 \frac{\text{kJ}}{\text{mol}} - 6.3 \frac{\text{kJ}}{\text{mol}} \right) \\ - \left(-22.5 \frac{\text{kJ}}{\text{mol}} \right) \\ = 2.4 \frac{\text{kJ}}{\text{mol}}$$

ΔH is positive.

059 10.0 points

Consider the following specific heats: copper, $0.384 \text{ J/g}\cdot^\circ\text{C}$; lead, $0.159 \text{ J/g}\cdot^\circ\text{C}$; water, $4.18 \text{ J/g}\cdot^\circ\text{C}$; glass, $0.502 \text{ J/g}\cdot^\circ\text{C}$. Which substance, once warmed, would be more likely to maintain its heat and keep you warm through a long football game on a cold night?

1. water **correct**
2. glass
3. copper
4. lead

Explanation:

Water has the highest specific heat of the substances listed, so it has the capacity to emit the largest quantity of heat with minimal temperature loss; the emitted heat keeps you warm. The substance continues to warm you until its temperature is at or below your body temperature.

060 10.0 points

A block of aluminum at 25°C and 1 atm is heated until it is a liquid at 700°C . It is then cooled back down until it is back in the initial state of being a solid at 25°C and 1 atm. For this entire process (heating and cooling) ΔH is...

1. positive
2. less than ΔU
3. zero **correct**

4. greater than ΔU

5. negative

Explanation:

Since the initial and final states of the system are identical, and since the value of a state function such as ΔH is dependent only on the state of the system, we can conclude that ΔH is zero.

061 10.0 points

Which statement about internal energy is true?

1. The internal energy of a system is equal to w at constant volume.

2. The internal energy of a system is constant at constant volume.

3. The internal energy of a system is equal to w at constant pressure.

4. The internal energy of a system is equal to q at constant volume. **correct**

5. The internal energy of a system is equal to q at constant pressure.

6. The internal energy of a system is constant at constant pressure.

Explanation:

062 10.0 points

When 0.100 g of graphite is burned completely in a bomb calorimeter (heat capacity = 3.344 kJ/°C), containing 3000 g of water, a temperature rise of 0.21°C is observed. What is ΔE for the combustion of graphite? The specific heat of liquid water is 4.184 J/g·°C.

1. $\Delta E = +3.34$ kJ/mol

2. $\Delta E = -40.1$ kJ/mol

3. $\Delta E = -285.$ kJ/mol

4. $\Delta E = -3.34$ kJ/mol

5. $\Delta E = -401.0$ kJ/mol **correct**

Explanation:

$$\begin{aligned} m_{\text{graphite}} &= 0.100 \text{ g} & m_{\text{water}} &= 3000 \text{ g} \\ \Delta T &= 0.21^\circ\text{C} & SH_{\text{water}} &= 4.184 \text{ J/g}\cdot^\circ\text{C} \\ HC &= 3.344 \text{ kJ/}^\circ\text{C} \end{aligned}$$

The amount of heat responsible for the temperature increase for 3000 g of water is

$$\begin{aligned} q &= \left(\frac{4.184 \text{ J}}{\text{g}\cdot^\circ\text{C}}\right)(3000 \text{ g})(0.21^\circ\text{C})\left(\frac{1 \text{ kJ}}{1000 \text{ J}}\right) \\ &= 2.6359 \text{ kJ} \end{aligned}$$

The amount of heat responsible for the warming of the calorimeter is

$$q = \left(3.344 \frac{\text{kJ}}{^\circ\text{C}}\right)(0.21^\circ\text{C}) = 0.7022 \text{ kJ}$$

The amount of heat released for this reaction is

$$2.6359 \text{ kJ} + 0.7022 \text{ kJ} = 3.3381 \text{ kJ}$$

The reaction was exothermic and there were 0.1 g of graphite, so

$$\left(\frac{-3.3381 \text{ kJ}}{0.1 \text{ g}}\right)\left(12 \frac{\text{g}}{\text{mol}}\right) = -401 \frac{\text{kJ}}{\text{mol}}$$

063 10.0 points

When a given reaction was run at a constant pressure of 1 atm, the system absorbed 5 kJ of heat and the gases were consumed, causing the volume to decrease from 3.5 L to 1.5 L. What are ΔH and ΔU , respectively?

1. +5 kJ, +0.2 kJ

2. -5 kJ, -4.8 kJ

3. +5 kJ, +5.2 kJ **correct**

4. +5.2 kJ, +5 kJ

5. +5 kJ, +4.8 kJ

6. -4.8 kJ, +0.2 kJ

7. -5 kJ, -5.2 kJ

8. +5 kJ, +5 kJ

9. -5 kJ, -5 kJ

Explanation:

none

064 10.0 points

Juan freezes a bottle of water to ice (500.mL) in preparation for a road trip. How much heat can be absorbed by that ice before it is fully melted?

1. 2090 kJ

2. 167 kJ **correct**

3. 0 kJ

4. 1130 kJ

5. 0.500 kJ

6. 1.50 kJ

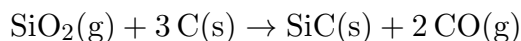
7. 6.02 kJ

Explanation:

Density of water is 1 g/mL. So 1 g = 1 mL for water. $q = (500\text{g}) \times (334 \text{ J/g}) = 167000 \text{ J} = 167 \text{ kJ}$.

065 10.0 points

How much heat is absorbed in the complete reaction of 3.00 grams of SiO_2 with excess carbon in the reaction below?



ΔH for the reaction is +624.7 kJ/mol rxn.

1. 31.2 kJ **correct**

2. 1.33×10^4 kJ

3. 5.06 kJ

4. 366 kJ

5. 1.13×10^5 kJ

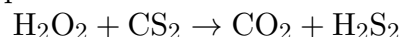
Explanation:

$$m_{\text{SiO}_2} = 3.00 \text{ g} \quad \Delta H = +624.7 \text{ kJ/mol}$$

$$\begin{aligned} q &= \left(\frac{624.7 \text{ kJ}}{\text{mol rxn}} \right) \left(\frac{1 \text{ mol rxn}}{1 \text{ mol SiO}_2} \right) \\ &\quad \times \left(\frac{1 \text{ mol SiO}_2}{60 \text{ g SiO}_2} \right) (3 \text{ g SiO}_2) \\ &= 31.2 \text{ kJ} \end{aligned}$$

066 10.0 points

Using bond energies, estimate the enthalpy change for the reaction between hydrogen peroxide (H_2O_2) and carbon disulfide (CS_2) to produce carbon dioxide (CO_2) and hydrogen disulfide (H_2S_2) according to the balanced equation:



1. -577 kJ/mol

2. -106 kJ/mol

3. 292 kJ/mol

4. 106 kJ/mol

5. -292 kJ/mol **correct**

6. 577 kJ/mol

Explanation:

$$\Delta H_{\text{rxn}} = \Sigma \text{BE}_{\text{broken}} - \Sigma \text{BE}_{\text{formed}}$$

Total bonds broken are 2 hydrogen oxygen single bonds (463 kJ/mol), 1 oxygen oxygen single bond (146 kJ/mol) and 2 carbon sulfur double bonds (577 kJ/mol). Total bonds formed are 2 hydrogen sulfur single bonds (347 kJ/mol), 1 sulfur sulfur single bond (226 kJ/mol) and 2 carbon oxygen double bonds (799 kJ/mol).

$$\Delta H_{\text{rxn}} = -292 \text{ kJ/mol}$$

067 10.0 points

The following reaction occurs during the production of metallic iron:



Calculate ΔH for this reaction at 25°C and 1 atm.

ΔH_f for $\text{CO}_2(\text{g}) = -393.51 \text{ kJ/mol}$, and

ΔH_f for $\text{Fe}_2\text{O}_3(\text{s}) = -824.2 \text{ kJ/mol}$.

1. There is insufficient information to answer this question.

2. +467.9 kJ **correct**

3. -430.7 kJ

4. +430.7 kJ

5. -467.9 kJ

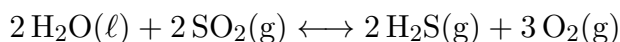
Explanation:

$$\begin{aligned} \Delta H_{\text{rxn}} &= \Delta H_{\text{products}} - \Delta H_{\text{reactants}} \\ &= (-393.51)(3) - (-824.2)(2) \\ &= 467.9 \text{ kJ} \end{aligned}$$

Note: The enthalpy of a pure element is always zero.

068 10.0 points

Based on thermodynamic table data calculate ΔH_{rxn} for



1. $560 \text{ kJ} \cdot \text{mol}^{-1}$

2. $-560 \text{ kJ} \cdot \text{mol}^{-1}$

3. $1120 \text{ kJ} \cdot \text{mol}^{-1}$ **correct**

4. $-1120 \text{ kJ} \cdot \text{mol}^{-1}$

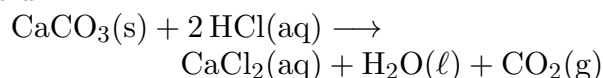
Explanation:

$$\begin{aligned} \Delta H_{\text{rxn}} &= \sum \Delta H_{f, \text{ products}} - \sum \Delta H_{f, \text{ reactants}} \\ &= [2(-21 \text{ kJ} \cdot \text{mol}^{-1}) + 3(0)] \end{aligned}$$

$$\begin{aligned} &- [2(-286 \text{ kJ} \cdot \text{mol}^{-1}) \\ &\quad + 2(-297 \text{ kJ} \cdot \text{mol}^{-1})] \\ &= 1,120 \text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

069 10.0 points

Calculate the standard reaction enthalpy for the reaction of calcite with hydrochloric acid



The standard enthalpies of formation are:

for $\text{CaCl}_2(\text{aq}) : -877.1 \text{ kJ/mol}$;

for $\text{H}_2\text{O}(\ell) : -285.83 \text{ kJ/mol}$;

for $\text{CO}_2(\text{g}) : -393.51 \text{ kJ/mol}$;

for $\text{CaCO}_3(\text{s}) : -1206.9 \text{ kJ/mol}$;

and for $\text{HCl}(\text{aq}) : -167.16 \text{ kJ/mol}$.

1. -38.2 kJ/mol

2. -98.8 kJ/mol

3. -116 kJ/mol

4. -15.2 kJ/mol **correct**

5. -72.7 kJ/mol

6. -165 kJ/mol

7. -215 kJ/mol

Explanation:

We use Hess' Law:

$$\begin{aligned} \Delta H^\circ &= \sum n \Delta H_{j, \text{prod}}^\circ - \sum n \Delta H_{j, \text{reac}}^\circ \\ &= \Delta H_{f, \text{CaCl}_2(\text{aq})}^\circ + \Delta H_{f, \text{H}_2\text{O}(\ell)}^\circ \\ &\quad + \Delta H_{f, \text{CO}_2(\text{g})}^\circ - \left[\Delta H_{f, \text{CaCO}_3(\text{s})}^\circ \right. \\ &\quad \left. + 2 \left(\Delta H_{f, \text{HCl}(\text{aq})}^\circ \right) \right] \\ &= -877.1 \text{ kJ/mol} + (-285.83 \text{ kJ/mol}) \\ &\quad + (-393.51 \text{ kJ/mol}) \\ &\quad - \left[-1206.9 \text{ kJ/mol} \right. \\ &\quad \left. + 2(-167.16 \text{ kJ/mol}) \right] \\ &= -15.22 \text{ kJ/mol}. \end{aligned}$$

070 10.0 points

For an exothermic reaction, the sum of bond energies for the reactants are (greater/lesser) than those of the products.

1. lesser **correct**

2. greater

Explanation:

For an exothermic reaction, the change in enthalpy is negative.

4. -828 kJ/mol **correct**

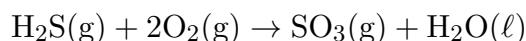
5. -168.2 kJ/mol

Explanation:

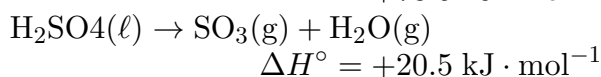
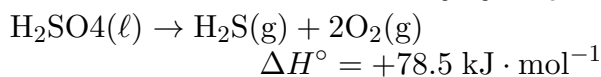
None

071 10.0 points

Calculate the standard reaction enthalpy for the reaction



given



1. +88.0 kJ

2. -69.0 kJ **correct**

3. +110.0 kJ

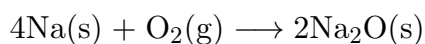
4. -47.0 kJ

Explanation:

The overall reaction is sum of reaction 1, reaction 3 and the reverse of reaction 2.

072 10.0 points

When 17.8 g sodium is treated with excess oxygen, 160.2 kJ of heat is produced. What is the ΔH_{rxn} for the below reaction?



1. -1682 kJ/mol

2. -15.2 kJ/mol

3. -152 kJ/mol