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

\[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\ell) \]

find the value for the work done at 300 K.

A. 2.5 kJ
B. -7.5 kJ
C. 7.5 kJ correct
D. -2.5 kJ

**Explanation:**

\[ w = -P \Delta V = -\Delta n RT \]

\[ = -(-3 \text{ mol}) \left( \frac{8.314 \text{ J}}{\text{mol} \cdot \text{K}} \right) (300 \text{ K}) \]

\[ = 7482.6 \text{ J} \]

002 10.0 points

The enthalpy of fusion of methanol (CH\text{3OH}) is 3.16 kJ/mol. How much heat would be absorbed or released upon freezing 25.6 grams of methanol?

A. 2.52 kJ absorbed
B. 0.253 kJ absorbed
C. 2.52 kJ released correct
D. 3.95 kJ released
E. 3.95 kJ absorbed
F. 0.253 kJ released

**Explanation:**

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

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 \(\Delta U\) in kJ/g? Ignore any heat loss or gain by the calorimeter itself.

A. 19.6 kJ/g
B. -3.9 kJ/g
C. -0.08 kJ/g
D. -19.6 kJ/g correct
E. 3.9 kJ/g
F. 0.08 kJ/g

**Explanation:**

None

004 10.0 points

For the combustion reaction of ethylene (C\text{2H}_4)

\[ \text{C}_2\text{H}_4 + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 2 \text{H}_2\text{O} \]

assume all reactants and products are gases, and calculate the \(\Delta H^0_{\text{rxn}}\) using bond energies.

A. 0 kJ/mol
B. 251 kJ/mol
C. 680 kJ/mol
D. -1300 kJ/mol correct
E. -251 kJ/mol
F. 1300 kJ/mol
7. $-680 \text{ kJ/mol}$

**Explanation:**

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

---

**006 10.0 points**

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

$$
2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)
$$
given

$$
\Delta H^\circ = +180.5 \text{ kJ/mol}^{-1}
$$

$$
2 \text{NO}_2(g) \rightarrow \text{N}_2(g) + 2 \text{O}_2(g)
$$

$$
\Delta H^\circ = -66.4 \text{ kJ/mol}^{-1}
$$

1. $-294.6 \text{ kJ/mol}^{-1}$
2. $+114.1 \text{ kJ/mol}^{-1}$
3. $-114.1 \text{ kJ/mol}^{-1}$ **correct**
4. $+246.9 \text{ kJ/mol}^{-1}$
5. $-246.9 \text{ kJ/mol}^{-1}$

**Explanation:**

We need to reverse the reactions and add them:

$$
2 \text{NO}(g) \rightarrow \text{N}_2(g) + \text{O}_2(g)
$$

$$
\Delta H = -180.5 \text{ kJ/mol}
$$

$$
\text{NO}_2(g) + 2 \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)
$$

$$
\Delta H = +66.4 \text{ kJ/mol}
$$

$$
2 \text{NO}(g) + 2 \text{O}_2(g) \rightarrow \text{O}_2(g) + 2 \text{NO}_2(g)
$$

$$
\Delta H = -114.1 \text{ kJ/mol}
$$

---

**007 10.0 points**

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

$$
C_{\text{cal}} = \frac{[2022(4.184) + 1540 ]/1000}{10.00 \text{ kJ/}^\circ \text{C}} = 0.9211/88.15 = 0.01045 \text{ mol}
$$

$$
\Delta U = -35.130 \text{ kJ} / 0.01045 \text{ mol} = -3362 \text{ kJ/mol}
$$

---

**Methyl tert-butyl ether or MTBE is an octane booster for gasoline. The combustion of 0.9211 grams of MTBE ($\text{C}_5\text{H}_{12}\text{O}(l)$, 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 ($\Delta U$) for this octane booster.**

1. $-3362 \text{ kJ/mol}$ **correct**
2. $-3120 \text{ kJ/mol}$
3. $-3560 \text{ kJ/mol}$
4. $-1957 \text{ kJ/mol}$
5. $-2286 \text{ kJ/mol}$
6. $-2748 \text{ kJ/mol}$
7. $-4293 \text{ kJ/mol}$

**Explanation:**

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

$$
q_{\text{cal}} = \frac{2022(4.184) + 1540}{1000} = 0.01045 \text{ mol}
$$

$$
\Delta U = -35.130 \text{ kJ} / 0.01045 \text{ mol} = -3362 \text{ kJ/mol}
$$
of Guinness is 4.186 J/g·°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:

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 (∆U) 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: ∆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:

\[
\text{Step 1: } 2.09 \text{ J} / \text{g} \cdot \text{°C} \cdot (12 \text{ g}) \cdot (0 - -40) \text{°C} = 1,003.2 \text{ J}
\]

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

Step 3: \[
\frac{4.184 \text{ J}}{\text{g} \cdot \text{°C}} \cdot (12 \text{ g}) \cdot (25 - 0) \text{°C} = 1,255.2 \text{ J}
\]

Total = 1,003.2 J + 4,008 J + 1,255.2 J = 6,266.4 J = 6.27 kJ

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:

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

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.

\[
\Delta U = q + w
\]

\[
= [-50 \text{ kJ} + (-250 \text{ kJ})] + (-500 \text{ kJ})
\]

\[
= -800 \text{ kJ}
\]
\[
(4.184 \text{ J/g} \cdot ^\circ \text{C}) (200 \text{ g})
\times (2.00^\circ \text{C})
= 1673 \text{ J}
\]

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:

\[
\Delta H \text{ of calorimeter} = (c_{\text{cal}})(\Delta T) = (150 \text{ J}/^\circ \text{C})(2.00^\circ \text{C}) = 300 \text{ J}
\]

\[
\Delta H \text{ of rxn} = \Delta H \text{ of water} + \Delta H \text{ of calorimeter}
= 1673 \text{ J} + 300 \text{ J} = 1973 \text{ J}
\]

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 \(\Delta H\) 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?

1. 15.5 cal
2. 6150 cal
3. 155 cal
4. 1.55 kcal correct
5. 61.5 kcal

**Explanation:**

\[
\frac{1 \text{ calorie} \cdot ^\circ \text{C}^{-1} \cdot \text{g}^{-1} \times 4000 \text{ ml} \times 1 \text{ g} \cdot \text{ml}^{-1} \times 1.2 \text{K}}{20083.2 \text{ calories}}
\]

**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, \(\Delta U\).

1. 4814.4 calories correct
2. 1150.7 calories
3. 20083.2 calories
4. 20143.4 calories
5. 4800.0 calories
6. 1147.2 calories

**Explanation:**

\[
\frac{1 \text{ calorie} \cdot ^\circ \text{C}^{-1} \cdot \text{g}^{-1} \times 4000 \text{ ml} \times 1 \text{ g} \cdot \text{mL}^{-1} \times 1.2 \text{K}}{20143.4 \text{ calories}}
\]

**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, \(\Delta H\), for this reaction.

\[
\begin{align*}
\text{OH} & \quad + \quad \text{HBr} \quad \rightarrow \quad \text{Br} & \quad + \quad \text{H}_2\text{O} \\
10^\circ \text{C} & \quad \rightarrow & \quad 75^\circ \text{C}
\end{align*}
\]

1. +105 kJ/mol
2. +186 kJ/mol
3. +24 kJ/mol
4. \(-105 \text{ kJ/mol}\)

5. \(-24 \text{ kJ/mol correct}\)

6. \(-186 \text{ kJ/mol}\)

**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.

\[\Delta H_{\text{rxn}} = B.E_{\text{C-O}} + B.E_{\text{H-Br}} - B.E_{\text{C-Br}} - B.E_{\text{H-O}}\]

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

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

016 10.0 points

Estimate the heat released when 1-butene (CH₃CH₂CH=CH₂) reacts with bromine to give CH₃CH₂CHBrCH₂Br. 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:**

\[\begin{array}{cccc}
\text{H} & \text{H} & \text{Br} & \text{H} \\
\text{H–C–C–C} & \text{C} & \text{Br–Br} \rightarrow \\
\text{H} & \text{H} & \text{H} & \text{H}
\end{array}\]

\[\Delta H = \sum B.E_{\text{break}} - \sum B.E_{\text{make}}\]

\[= [(\text{C=O}) + (\text{Br-Br})] - 2(\text{C-Br}) + (\text{C=C})\]

\[= 612 \text{ kJ/mol} + 193 \text{ kJ/mol} - 2(276 \text{ kJ/mol}) + 348 \text{ kJ/mol}\]

\[= -95 \text{ kJ/mol},\]

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 \(\Delta T\) 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, $\Delta H$ must be zero when external pressure is zero.
II) For a given process, $\Delta U$ and $\Delta H$ must have different values.
III) For a given process, $\Delta U_{\text{sys}}$ and $\Delta U_{\text{surr}}$ must have the same magnitude.

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

**Explanation:**
$\Delta H$ is heat transferred at constant pressure. For a given process $\Delta U$ and $\Delta H$ do not always have different values; for example, they can both be zero for a given process. Because $\Delta U_{\text{univ}}$ must be zero according to the 1st law, $\Delta U$ 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:

$$2 \text{K(s)} + 2 \text{H}_2\text{O(ℓ)} \rightarrow 2 \text{KOH(aq)} + \text{H}_2(\text{g})$$

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 $\Delta H$) and energies to form bonds are negative (release energy, negative $\Delta H$).

022  10.0 points
Which of $\text{O}_2(\text{g})$, $\text{O}_2(\text{ℓ})$, $\text{H}_2(\text{g})$, $\text{H}_2(\text{ℓ})$, $\text{H}_2\text{O(ℓ)}$, $\text{H}_2\text{O(ℓ)}$.
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, \(-890 \text{ kJ/mol}\) 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 \(\Delta H\) ? (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: \(-202.475 \text{ kJ}\).

Explanation:
\( \text{CH}_4: \Delta H_c = -890 \text{ kJ/mol} \)
\( m = 3.64 \text{ g} \)

\[ \Delta H = ? \]

\[ n_{\text{CH}_4} = \frac{(3.64 \text{ g CH}_4)}{16 \text{ g CH}_4} \]

\[ = 0.2275 \text{ mol CH}_4 \]

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

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:
\[ n = 3.0 \text{ mol} \]

\[ \frac{40.6 \text{ kJ/mol}}{} \times 3 \text{ mol} = 121.8 \text{ kJ} \]

025  10.0 points

Calculate the heat of formation for 2.6 mol of sulfur dioxide (\(\text{SO}_2\)) from its elements, sulfur and oxygen. Use the balanced chemical equation and the following information.

\[ \text{S(s)} + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g}) \]

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

\[ 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{SO}_3(\text{g}) \]

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

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: \(-769.86 \text{ kJ}\).

Explanation:
\( \Delta H_f \) of \(\text{SO}_2\) from \(\text{S}\) and \(\text{O} = ? \)

The balanced equation is

\[ \text{S(s)} + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) \]
Reaction \( \Delta H^0_f \) (kJ/mol)

\[
\begin{array}{l}
S(s) + \frac{3}{2} O_2(g) \rightarrow SO_3(g) & -395.2 \\
\frac{1}{2} \left[2 SO_3(g) \rightarrow 2 SO_2(g) + O_2 \right] & 99.1 \\
S(s) + O_2(g) \rightarrow SO_2(g) & -296.1 \\
\end{array}
\]

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. \( \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(g) \)
B. \( \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s) \)
C. \( \frac{1}{2}\text{H}_2(g) + \frac{1}{2}\text{F}_2(g) \rightarrow \text{HF}(g) \)
\( \Delta H = -271.1 \text{ kJ/mol} \)
D. \( \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \)
E. \( \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO}(g) \)
\( \Delta H = 180.50 \text{ kJ/mol} \)

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 \( \Delta H_{\text{rxn}} = \Delta H_f \) 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
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.

\[
2 \text{CH}_3\text{OH}(g) + 3 \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) + 4 \text{H}_2\text{O}(\ell)
\]

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.

---

**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:**
\[
q = 455 \text{ J} \\
m = 67 \text{ g} \\
\text{SH}_{\text{H}_2\text{O}} = 4.18 \text{ J/g} \cdot \text{K}
\]

\[
\Delta T = \frac{q}{m \cdot \text{SH}_{\text{H}_2\text{O}}} = \frac{455 \text{ J}}{(67 \text{ g})(4.18 \text{ J/g} \cdot \text{K})} = 1.62465 \text{ K} = 1.62465^\circ \text{C}
\]

\[
\Delta T = T_f - T_i \\
T_f = \Delta T + T_i = 1.62465^\circ \text{C} + 7^\circ \text{C} = 8.62465^\circ \text{C}
\]

---

**032 10.0 points**
Which of the reactions below is a formation reaction?

1. \( 2 \text{Fe}(s) + 3 \text{O}(g) \rightarrow \text{Fe}_2\text{O}_3(s) \)
2. \( \text{B}_2(\ell) + 2 \text{I}_2(\ell) + \text{Cl}_2(g) \rightarrow 2 \text{BI}_2\text{Cl}(g) \)
3. \( \text{C}_{\text{diamond}}(s) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{CO}(g) \)
4. \( \text{N}_2(g) + 2 \text{H}_2(g) + \frac{1}{2} \text{O}_2(g) \rightarrow \text{N}_2\text{H}_4\text{O}(g) \) **correct**

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

An important reaction that takes place in the atmosphere is

\[ \text{NO}_2(g) \rightarrow \text{NO}(g) + \text{O}(g) \]

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

<table>
<thead>
<tr>
<th>reaction</th>
<th>( \Delta H^\circ ) (kJ)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{O}_2(g) \rightarrow 2 \text{O}(g) )</td>
<td>+498.4</td>
</tr>
<tr>
<td>( \text{NO}(g) + \text{O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) )</td>
<td>-200.0</td>
</tr>
<tr>
<td>( \frac{1}{2} \text{O}_2(g) \rightarrow \text{O}(g) )</td>
<td>+142.7</td>
</tr>
</tbody>
</table>

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:

\[ \text{NO}_2(g) + \text{O}_2(g) \rightarrow \text{NO}(g) + \text{O}_3(g) + 200 \]
\[ \text{O}_3(g) \rightarrow \frac{3}{2} \text{O}_2(g) \quad -142.7 \]
\[ \frac{1}{2} \text{O}_2(g) \rightarrow \text{O}(g) \quad +249.2 \]

\[ \text{NO}_2(g) \rightarrow \text{NO}(g) + \text{O}(g) \quad 306.5 \]

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,

\[ w = -P_{\text{ext}} \Delta V \]
\[ = (-2 \text{ atm})(2 \text{ L} - 11.2 \text{ L}) \]
\[ \times (101.325 \text{ J} \cdot \text{L}^{-1} \cdot \text{atm}^{-1}) \]
\[ = 1.86 \text{ kJ} \]

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 \( \Delta E \) 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)

\[ \text{C}_2\text{H}_4 + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 2 \text{H}_2\text{O} \]

assume all reactants and products are gases, and calculate the \( \Delta H^\circ_{\text{rxn}} \) using bond energies from the list below.

- \( \text{C} \rightarrow \text{H} : 413 \text{ kJ/mol} \)
- \( \text{H} \rightarrow \text{O} : 463 \text{ kJ/mol} \)
- \( \text{O} \rightarrow \text{O} : 146 \text{ kJ/mol} \)
- \( \text{O} \rightarrow \text{O} : 498 \text{ kJ/mol} \)
- \( \text{C} \rightarrow \text{C} : 346 \text{ kJ/mol} \)
- \( \text{C} \rightarrow \text{C} : 602 \text{ kJ/mol} \)
- \( \text{C} \rightarrow \text{O} : 358 \text{ kJ/mol} \)
- \( \text{C} \rightarrow \text{O} : 799 \text{ kJ/mol} \)

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:

\[
\Delta H^\circ_{\text{rxn}} = \sum \text{BE}_{\text{reactants}} - \sum \text{BE}_{\text{products}} \\
= \left[ (\text{C} \equiv \text{C}) + 4 (\text{C} \rightarrow \text{H}) \right. \\
\left. + 3 (\text{O} \equiv \text{O}) \right] \\
- \left[ 4 (\text{C} \equiv \text{O}) + 4 (\text{H} \rightarrow \text{O}) \right] \\
= \left[ \left( 602 \frac{\text{kJ}}{\text{mol}} \right) + 4 \left( 413 \frac{\text{kJ}}{\text{mol}} \right) \right]
\]

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 J/g

scale up to 2.26 g :

432.324(2.26) = 977.052 J

= 977 J released

040 10.0 points
Which of the following reactions is an enthalpy of formation reaction?

1. \( \frac{1}{2} \text{N}_2(\ell) + \frac{3}{2} \text{H}_2(g) \rightarrow \text{NH}_3(g) \)
2. \( \text{NaOH(aq)} + \text{HCl(g)} \rightarrow \text{H}_2\text{O(\ell)} + \text{NaCl(aq)} \)
3. \( 2 \text{Fe(s)} + \frac{3}{2} \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) \) correct
4. \( \text{CH}_4(g) \rightarrow \text{C}_{\text{graphite}} + 2 \text{H}_2(g) \)

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.

\[
w = -P \Delta V = -(330/760)(3.75L)
w = -1.628 L \text{ atm} 
× 101.325 \text{ J/(L atm)} = -165 \text{ J}
\]

042 10.0 points
Calculate the standard reaction enthalpy for the reaction

\[
2\text{H}_2\text{O}(\ell) \rightarrow 2\text{CO}_2(\g) + 3\text{H}_2\text{O}(\ell)
\]

given

\[
\Delta H^\circ = 228 \text{ kJ} \cdot \text{mol}^{-1}
\]

\[
\text{CO}_2(\g) \rightarrow \text{C}_{\text{graphite}}(s) + \text{O}_2(\g) 
\Delta H^\circ = 394 \text{ kJ} \cdot \text{mol}^{-1}
\]

\[
\text{H}_2(\g) + 1/2 \text{O}_2(\g) \rightarrow \text{H}_2\text{O}(\ell) 
\Delta H^\circ = -286 \text{ kJ} \cdot \text{mol}^{-1}
\]

1. \(730\) kJ \cdot \text{mol}^{-1}
2. \(-846\) kJ \cdot \text{mol}^{-1}
3. \(-1,418\) kJ \cdot \text{mol}^{-1} correct
4. \(336\) kJ \cdot \text{mol}^{-1}
5. \(-452\) kJ \cdot \text{mol}^{-1}

Explanation:
Values for \(\Delta H_i^\circ\) from external table are in order (from reaction) +33, -286, -207, and +90

\[
\Delta H^\circ_{\text{rxn}} = \left(\sum n \Delta H_j^\circ\right)_{\text{products}} - \left(\sum n \Delta H_j^\circ\right)_{\text{reactants}}
\]

\[
= \left[2 \Delta H_f^\circ, \text{HNO}_3(\text{aq}) + \Delta H_f^\circ, \text{O}_2(\g)\right] - \left[3 \Delta H_f^\circ, \text{NO}_2(\g) + \Delta H_f^\circ, \text{H}_2\text{O}(\ell)\right]
\]

\[
= \left[2 (-207) + 90\right] - \left[3 (33) + (-286)\right]
\]

\[
= -137 \text{ kJ}
\]

043 10.0 points
Calculate the standard reaction enthalpy \((\Delta H^\circ_{\text{rxn}})\) for the final stage in the production of nitric acid, when nitrogen dioxide dissolves in and reacts with water:

\[
3\text{NO}_2(\g) + \text{H}_2\text{O}(\ell) \rightarrow 2\text{HNO}_3(\text{aq}) + \text{NO}(\g)
\]

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

Explanation:

\[
\Delta H^\circ_{\text{rxn}} = \left(\sum n \Delta H_j^\circ\right)_{\text{products}} - \left(\sum n \Delta H_j^\circ\right)_{\text{reactants}}
\]

\[
= \left[2 \Delta H_f^\circ, \text{HNO}_3(\text{aq}) + \Delta H_f^\circ, \text{O}_2(\g)\right] - \left[3 \Delta H_f^\circ, \text{NO}_2(\g) + \Delta H_f^\circ, \text{H}_2\text{O}(\ell)\right]
\]

\[
= \left[2 (-207) + 90\right] - \left[3 (33) + (-286)\right]
\]

\[
= -137 \text{ kJ}
\]

044 10.0 points
The molar heat capacity of \(\text{C}_6\text{H}_6(\ell)\) is 136 J/mol · °C and of \(\text{C}_6\text{H}_6(\g)\) 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 \( \text{C}_6\text{H}_6(s) \) at 5.5°C into benzene vapor \( \text{C}_6\text{H}_6(g) \) 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:**

\[ m_{\text{benzene}} = 234 \text{ g} \]
\[ T_1 = 5.5^\circ \text{C} \]
\[ T_2 = 100.0^\circ \text{C} \]

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

\[ \text{C}_6\text{H}_6(s) \xrightarrow{\text{step 1}} \text{C}_6\text{H}_6(\ell) \xrightarrow{\text{step 2}} \]
\[ \text{C}_6\text{H}_6(\ell) \xrightarrow{\text{step 3}} \text{C}_6\text{H}_6(g) \xrightarrow{\text{step 4}} \text{C}_6\text{H}_6(g) \]

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

Step 2 : \( \frac{136 \text{ J}}{\text{mol} \cdot ^\circ \text{C}} \times (3 \text{ mol}) \times (80.1 - 5.5)^\circ \text{C} \)
\[ = 30436.8 \text{ J} = 30.4368 \text{ kJ} \]

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

Step 4 : \( \frac{81.6 \text{ J}}{\text{mol} \cdot ^\circ \text{C}} \times (3 \text{ mol}) \times (100.0 - 80.1)^\circ \text{C} \)
\[ = 4871.52 \text{ J} = 4.87152 \text{ kJ} \]

**Total** = 29.76 kJ + 30.4368 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

\[ \text{NO}_2(g) \rightarrow \text{NO(g) + O(g)} \]

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

\[ \text{O}_2(g) \rightarrow 2 \text{O(g)} \]
\[ \Delta H^o = +498.4 \text{ kJ/mol} \]

\[ \text{NO(g) + O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) \]
\[ \Delta H^o = -200 \text{ kJ/mol} \]

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:**

\[ \text{O}_2(g) \rightarrow 2 \text{O(g)} \]
\[ \Delta H^o = +498.4 \text{ kJ/mol} \]

\[ \text{NO(g) + O}_3(g) \rightarrow \text{NO}_2(g) + \text{O}_2(g) \]
\[ \Delta H^o = -200 \text{ kJ/mol} \]

The standard formation of ozone is

\[ \frac{3}{2} \text{O}_2(g) \rightarrow \text{O}_3(g) \]
\[ \Delta H = +142.7 \text{ kJ/mol} \]

We calculate the \( \Delta H_{\text{rxn}} \) 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:

\[ \text{NO}_2(g) + \text{O}_2(g) \rightarrow \text{NO(g) + O}_3(g) \]
\[ \Delta H = +200 \text{ kJ/mol} \]

\[ \text{O}_3(g) \rightarrow \frac{3}{2} \text{O}_2(g) \]
\[ \Delta H = -142.7 \text{ kJ/mol} \]

\[ \frac{1}{2} \text{O}_2(g) \rightarrow \text{O(g)} \]
\[ \Delta H = \frac{1}{2}(498.4 \text{ kJ/mol}) \]

\[ \text{NO}_2(g) \rightarrow \text{NO(g) + O(g)} \]
\[ \Delta H_{\text{rxn}} = 306.5 \text{ kJ/mol} \]
046 10.0 points
A coffee cup calorimeter measures the heat at constant pressure whereas a bomb calorimeter measures the heat at constant volume.

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 J/g°C and the other is benzene, \( \text{C}_6\text{H}_6 \), with a specific heat of 1.74 J/g°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 \( (\Delta T) \) 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, \( \Delta H \), for this gas phase reaction.

\[
\begin{align*}
\text{H}_3\text{C} & \quad \text{CH} - \text{CH}_2 - \text{Br} \\
\text{H}_3\text{C} & \quad \text{C} = \text{CH}_2 + \text{HBr}
\end{align*}
\]

1. \(-270 \text{ kJ/mol}\)

2. \(+270 \text{ kJ/mol}\)

3. \(-76 \text{ kJ/mol}\)

4. \(+337 \text{ kJ/mol}\)

5. \(+76 \text{ kJ/mol}\) correct

6. \(-337 \text{ kJ/mol}\)

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. correct
## Explanation:

For which of the following chemical equations would $\Delta H^\circ_{\text{rxn}} = \Delta H^\circ_f$?

1. $\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\ell)$ **correct**

2. $\text{C}(\text{s, graphite}) + \frac{3}{2}\text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$

3. $\text{CO}(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$

4. $\text{N}_2(\ell) + 3\text{F}_2(\text{g}) \rightarrow 2\text{NF}_3(\ell)$

### Explanation:

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

## Explanation:

The combustion of methane gas ($\text{CH}_4$) 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 $\Delta H^0_f$ data to help:

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

1. 1566.75
2. 1513.34
3. 1290.79
4. 1459.93
5. 1726.99
6. 1175.06
7. 1424.32
8. 1201.77
9. 1121.65
10. 1762.6

Correct answer: 1762.6 kJ.

### Explanation:

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

$\text{CH}_4(\text{g}) + 2\text{O}_2 \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\ell)$

$\Delta H^0$ for combustion of $\text{CH}_4 = ?$

### Reaction

| $\text{CH}_4(\text{g}) \rightarrow \text{C}(\text{s}) + 2\text{H}_2(\text{g})$ | 74.9 kJ/mol |
| $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ | -393.5 kJ/mol |
| $2\text{H}_2(\text{g}) + \text{O}_2 \rightarrow 2\text{H}_2\text{O}(\ell)$ | $2(-285.8) = -571.6$ kJ/mol |
| $\text{CH}_4(\text{g}) + 2\text{O}_2 \rightarrow \text{CO}_2(\text{g}) + \text{H}_2(\ell)$ | -890.2 kJ/mol |

$\Delta H^0 = (890.2 \text{ kJ/mol}) (1.98 \text{ mol}) = 1762.6 \text{ kJ}$

## Explanation:

Calculate the standard reaction enthalpy for the reaction.

$\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$

1. +155.5 kJ mol$^{-1}$
2. +206.1 kJ mol$^{-1}$ **correct**
3. +216 kJ mol$^{-1}$
4. +412.1 kJ mol$^{-1}$
5. +42.0 kJ mol$^{-1}$

### Explanation:

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

- $\text{CH}_3\text{OH}(\ell) \rightarrow 2\text{H}_2(\text{g}) + \text{CO}(\text{g})$
  $\Delta H = 128.3 \text{ kJ/mol}$

- $\text{CH}_4(\text{g}) + 0.5\text{O}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\ell)$
  $\Delta H = -164.05 \text{ kJ/mol}$

- $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2(\text{g}) + 0.5\text{O}_2(\text{g})$
  $\Delta H = +241.8 \text{ kJ/mol}$

- $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\ell) \rightarrow \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$
  $\Delta H = 206.05 \text{ kJ/mol}$
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:
\[ \Delta U = q + w = 237 \text{ J} + (-435 \text{ J}) = -198 \text{ J} \]

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 kJ·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\(_3\)COCH\(_3\)
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:
\[ \Delta H = q = n \Delta H_{\text{cond}} = n(-\Delta H_{\text{vap}}) = (17.2177 \text{ mol})(-29.1 \text{ kJ/mol}) = -501.035 \text{ kJ} \]

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

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

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

The value of \( \Delta H \) for the reaction
\[ \text{C}_3\text{H}_8(g) + 5 \text{O}_2(g) \rightarrow 3 \text{CO}_2(g) + 4 \text{H}_2\text{O}(\ell) \]
is −2220 kJ/mol rxn. How much heat is given off when 33.0 g of propane gas (C\(_3\)H\(_8\)) 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:
\[ \Delta H = -2220 \text{ kJ/mol} \]
\[ m_{C_3H_8} = 33.0 \text{ g} \]

\[ q = \left( \frac{-2220 \text{ kJ}}{\text{mol rxn}} \right) \left( \frac{1 \text{ mol rxn}}{1 \text{ mol C}_3\text{H}_8} \right) \times \left( \frac{1 \text{ mol C}_3\text{H}_8}{44 \text{ g C}_3\text{H}_8} \right) (33 \text{ g C}_3\text{H}_8) \]

\[ = -1665 \text{ kJ} \text{ or } 1665 \text{ kJ released} \]

057 10.0 points

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

1. \( 2\text{NO(g)} + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \) correct

2. Neither reaction has \( \Delta H < \Delta U \).

3. Both reactions have \( \Delta H < \Delta U \).

4. \( \text{N}_2(g) + \text{O}_2(g) \rightarrow 2\text{NO(g)} \)

Explanation:

\[ \Delta H^0_{\text{rxn}} = \sum n H^0_{\text{prod}} - \sum n H^0_{\text{rct}} \]

\[ = \left( -13.8 \frac{\text{kJ}}{\text{mol}} - 6.3 \frac{\text{kJ}}{\text{mol}} \right) \]

\[ - \left( -22.5 \frac{\text{kJ}}{\text{mol}} \right) \]

\[ = 2.4 \frac{\text{kJ}}{\text{mol}} \]

\( \Delta H \) is positive.

058 10.0 points

Consider the following reaction

\[ \text{H}_2(g) + \text{CO}_2(g) \rightarrow \text{H}_2\text{O}(g) + \text{CO}(g) \]

\( \Delta H^\ddagger \) for \( \text{CO}_2(g) \) is \(-22.5 \text{ kJ/mol}\);

\( \Delta H^\ddagger \) for \( \text{CO}(g) \) is \(-6.3 \text{ kJ/mol}\);

\( \Delta H^\ddagger \) for \( \text{H}_2\text{O}(g) \) is \(-13.8 \text{ kJ/mol}\).

1. \( \Delta H \) of the reaction is negative.

2. \( \Delta H \) of the reaction is zero.

3. \( \Delta H \) of the reaction is positive. correct

Explanation:

Reactants:

\( \Delta H^\ddagger_{\text{CO}_2(g)} = -22.5 \text{ kJ/mol} \)

Products:

\( \Delta H^\ddagger_{\text{CO}(g)} = -6.3 \text{ kJ/mol} \)

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

059 10.0 points

Consider the following specific heats: copper, 0.384 J/g·°C; lead, 0.159 J/g·°C; water, 4.18 J/g·°C; glass, 0.502 J/g·°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) \( \Delta H \) is...

1. positive

2. less than \( \Delta U \)

3. zero correct
4. greater than $\Delta U$
5. negative

**Explanation:**
Since the initial and final states of the system are identical, and since the value of a state function such as $\Delta H$ is dependent only on the state of the system, we can conclude that $\Delta H$ 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:**

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

1. $\Delta E = +3.34 \text{ kJ/mol}$
2. $\Delta E = -40.1 \text{ kJ/mol}$
3. $\Delta E = -285. \text{ kJ/mol}$
4. $\Delta E = -3.34 \text{ kJ/mol}$
5. $\Delta E = -401.0 \text{ kJ/mol}$ correct

**Explanation:**

$m_{\text{graphite}} = 0.100 \text{ g}$
$m_{\text{water}} = 3000 \text{ g}$
$\Delta T = 0.21 \text{ °C}$
$\Delta H_{\text{water}} = 4.184 \text{ J/g} \cdot \text{°C}$

HC $= 3.344 \text{ kJ/°C}$

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

$$q = \left( \frac{4.184 \text{ J}}{\text{g} \cdot \text{°C}} \right) (3000 \text{ g})(0.21 \text{ °C})(\frac{1 \text{ kJ}}{1000 \text{ J}})$$

$$= 2.6359 \text{ kJ}$$

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

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

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

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

062 10.0 points
When 0.100 g of graphite is burned completely in a bomb calorimeter (heat capacity $= 3.344 \text{ kJ} / \text{°C}$), containing 3000 g of water, a temperature rise of 0.21 °C is observed. What is $\Delta E$ for the combustion of graphite? The specific heat of liquid water is 4.184 $\text{J} / \text{g} \cdot \text{°C}$.

1. $+5 \text{ kJ}$, $+0.2 \text{ kJ}$
2. $-5 \text{ kJ}$, $-4.8 \text{ kJ}$
3. $+5 \text{ kJ}$, $+5.2 \text{ kJ}$ correct
4. $+5.2 \text{ kJ}$, $+5 \text{ kJ}$
5. $+5 \text{ kJ}$, $+4.8 \text{ kJ}$
6. $-4.8 \text{ kJ}$, $+0.2 \text{ kJ}$
7. $-5 \text{ kJ}$, $-5.2 \text{ kJ}$

**Explanation:**

$m_{\text{graphite}} = 0.100 \text{ g}$
$m_{\text{water}} = 3000 \text{ g}$
$\Delta T = 0.21 \text{ °C}$
$SH_{\text{water}} = 4.184 \text{ J/g} \cdot \text{°C}$

HC $= 3.344 \text{ kJ/°C}$

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

$$q = \left( \frac{4.184 \text{ J}}{\text{g} \cdot \text{°C}} \right) (3000 \text{ g})(0.21 \text{ °C})(\frac{1 \text{ kJ}}{1000 \text{ J}})$$

$$= 2.6359 \text{ kJ}$$

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

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

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

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

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 $\Delta H$ and $\Delta U$, respectively?

1. $+5 \text{ kJ}$, $+0.2 \text{ kJ}$
2. $-5 \text{ kJ}$, $-4.8 \text{ kJ}$
3. $+5 \text{ kJ}$, $+5.2 \text{ kJ}$ correct
4. $+5.2 \text{ kJ}$, $+5 \text{ kJ}$
5. $+5 \text{ kJ}$, $+4.8 \text{ kJ}$
6. $-4.8 \text{ kJ}$, $+0.2 \text{ kJ}$
7. $-5 \text{ kJ}$, $-5.2 \text{ kJ}$

**Explanation:**
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?

\[ \text{SiO}_2(g) + 3 \text{C(s)} \rightarrow \text{SiC(s)} + 2 \text{CO(g)} \]

\( \Delta H \) for the reaction is +624.7 kJ/mol rxn.

1. 31.2 kJ correct
2. \( 1.33 \times 10^4 \) kJ
3. 5.06 kJ
4. 366 kJ
5. \( 1.13 \times 10^5 \) kJ

066 10.0 points
Using bond energies, estimate the enthalpy change for the reaction between hydrogen peroxide (H$_2$O$_2$) and carbon disulfide (CS$_2$) to produce carbon dioxide (CO$_2$) and hydrogen disulfide (H$_2$S$_2$) according to the balanced equation:

\[ \text{H}_2\text{O}_2 + \text{CS}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{S}_2 \]

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:
2 Fe₂O₃(s) + 3 C(graphite) → 4 Fe(s) + 3 CO₂(g)

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

ΔHₜ for CO₂(g) = −393.51 kJ/mol, and
ΔHₜ for Fe₂O₃(s) = −824.2 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:

\[ \Delta H_{\text{rxn}} = \Delta H_{\text{products}} - \Delta H_{\text{reactants}} \]
\[ = (-393.51)(3) - (-824.2)(2) \]
\[ = 467.9 \text{ kJ} \]

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

068 10.0 points

Based on thermodynamic table data calculate ∆Hrxn for

2 H₂O(ℓ) + 2 SO₂(g) ←→ 2 H₂S(g) + 3 O₂(g)

1. 560 kJ · mol⁻¹
2. −560 kJ · mol⁻¹
3. 1120 kJ · mol⁻¹ correct
4. −1120 kJ · mol⁻¹

Explanation:

\[ \Delta H_{\text{rxn}} = \sum n \Delta H_{\text{f, products}} - \sum n \Delta H_{\text{f, reactants}} \]
\[ = [2(-21 \text{ kJ} \cdot \text{mol}^{-1}) + 3(0)] \]
\[ = 2(-286 \text{ kJ} \cdot \text{mol}^{-1}) + 2(-297 \text{ kJ} \cdot \text{mol}^{-1}) \]
\[ = 1,120 \text{ kJ} \cdot \text{mol}^{-1} \]

069 10.0 points

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

CaCO₃(s) + 2 HCl(aq) → CaCl₂(aq) + H₂O(ℓ) + CO₂(g)

The standard enthalpies of formation are:
for CaCl₂(aq) : −877.1 kJ/mol;
for H₂O(ℓ) : −285.83 kJ/mol;
for CO₂(g) : −393.51 kJ/mol;
for CaCO₃(s) : −1206.9 kJ/mol;
and for HCl(aq) : −167.16 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:

\[ \Delta H^° = \sum n \Delta H_{\text{f, products}} - \sum n \Delta H_{\text{f, reactants}} \]
\[ = \Delta H_{\text{f, CaCl}_2(aq)} + \Delta H_{\text{f, H}_2\text{O}(\ell)} + \Delta H_{\text{f, CO}_2(g)} - [\Delta H_{\text{f, CaCO}_3(s)} + 2 \Delta H_{\text{f, HCl(aq)}}] \]
\[ = -877.1 \text{ kJ/mol} + (-285.83 \text{ kJ/mol}) \]
\[ + (-393.51 \text{ kJ/mol}) \]
\[ - [(-1206.9 \text{ kJ/mol}) + 2 (-167.16 \text{ kJ/mol})] \]
\[ = -15.22 \text{ kJ/mol} \]
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.

071 10.0 points
Calculate the standard reaction enthalpy for the reaction

\( \text{H}_2\text{S}(g) + 2\text{O}_2(g) \rightarrow \text{SO}_3(g) + \text{H}_2\text{O}(\ell) \)

given

\[ \text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(\ell) \quad \Delta H^\circ = -11.0 \text{ kJ} \cdot \text{mol}^{-1} \]
\[ \text{H}_2\text{SO}_4(\ell) \rightarrow \text{H}_2\text{S}(g) + 2\text{O}_2(g) \quad \Delta H^\circ = +78.5 \text{ kJ} \cdot \text{mol}^{-1} \]
\[ \text{H}_2\text{SO}_4(\ell) \rightarrow \text{SO}_3(g) + \text{H}_2\text{O}(g) \quad \Delta H^\circ = +20.5 \text{ kJ} \cdot \text{mol}^{-1} \]

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 \( \Delta H_{\text{rxn}} \) for the below reaction?

\( 4\text{Na}(s) + \text{O}_2(g) \rightarrow 2\text{Na}_2\text{O}(s) \)

1. -1682 kJ/mol
2. -152 kJ/mol
3. -152 kJ/mol

4. -828 kJ/mol correct
5. -168.2 kJ/mol

Explanation:
None