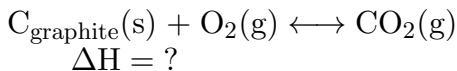
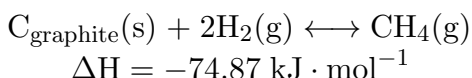
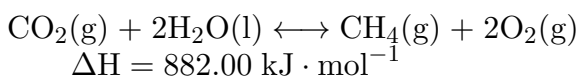
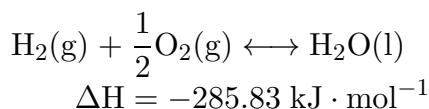


This print-out should have 21 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

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**001 10.0 points**

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



1.  $-385.21 \text{ kJ} \cdot \text{mol}^{-1}$  **correct**
2.  $-515.87 \text{ kJ} \cdot \text{mol}^{-1}$
3.  $521.30 \text{ kJ} \cdot \text{mol}^{-1}$
4.  $-230.04 \text{ kJ} \cdot \text{mol}^{-1}$
5.  $235.47 \text{ kJ} \cdot \text{mol}^{-1}$

**Explanation:**

In order for the three provided reactions to cancel to result in the unknown reaction (the combustion of graphite), the first reaction needs to be reversed and doubled, the second needs to be reversed, and the third will remain unchanged. Consequently, the overall change in enthalpy for the reaction is:  $\Delta\text{H}_{\text{rxn}} = -2 \times -285.83 + -1 \times 882.00 + -74.87 = -385.21 \text{ kJ} \cdot \text{mol}^{-1}$

---

**002 10.0 points**

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

1.  $\text{C}_5\text{H}_{12}(\text{l}) + 8\text{O}_2(\text{g}) \rightarrow 6\text{H}_2\text{O}(\text{g}) + 5\text{CO}_2(\text{g})$  **correct**
2.  $\text{N}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
3.  $\text{Na}^+(\text{g}) + \text{Cl}^-(\text{g}) \rightarrow \text{NaCl}(\text{s})$
4.  $\text{S}_3(\text{g}) + 9\text{F}_2(\text{g}) \rightarrow 3\text{SF}_6(\text{g})$
5.  $2\text{CH}_4(\text{g}) + 2\text{O}_3(\text{g}) \rightarrow 4\text{H}_2\text{O}(\text{g}) + 2\text{CO}(\text{g})$

**Explanation:**

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

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**003 10.0 points**

What is true about the first law of thermodynamics?

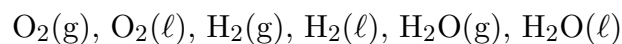
1.  $\Delta E_{\text{univ}} < 0$
2.  $\Delta E_{\text{univ}} = 0$  **correct**
3.  $\Delta E_{\text{sys}} = 0$
4.  $\Delta E_{\text{sys}} < 0$
5.  $\Delta E_{\text{sys}} > 0$
6.  $\Delta E_{\text{univ}} > 0$

**Explanation:**

---

**004 10.0 points**

Which of



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

5.  $\text{O}_2(\text{g}), \text{H}_2(\text{g}), \text{H}_2\text{O}(\text{g})$

**Explanation:**

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

---

**005 10.0 points**

If 25.0 g of water at  $100.0^\circ\text{C}$  are mixed with 15.0 g of water at  $40.0^\circ\text{C}$ , what temperature will the 40.0 g of combined water be at once they reach equilibrium?

1.  $70.0^\circ\text{C}$
2.  $77.5^\circ\text{C}$  correct
3.  $60.0^\circ\text{C}$
4.  $62.5^\circ\text{C}$

**Explanation:**

$$q_{\text{cold}} = -q_{\text{hot}}$$

$$15 \cdot 4.184 \cdot (T_f - 40) = -25 \cdot 4.184 \cdot (T_f - 100)$$

$$15 \cdot (T_f - 40) = -25 \cdot (T_f - 100)$$

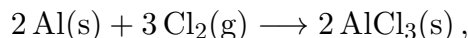
$$40 \cdot T_f = 3100$$

$$T_f = 77.5^\circ\text{C}$$

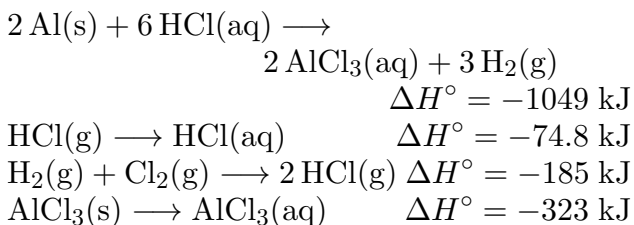
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**006 10.0 points**

Calculate the reaction enthalpy for the formation



of anhydrous aluminum chloride using the data



1.  $-1883.5 \text{ kJ}$
2.  $-1100.36 \text{ kJ}$
3.  $-1225.7 \text{ kJ}$
4.  $-1502.4 \text{ kJ}$
5.  $-1450.85 \text{ kJ}$

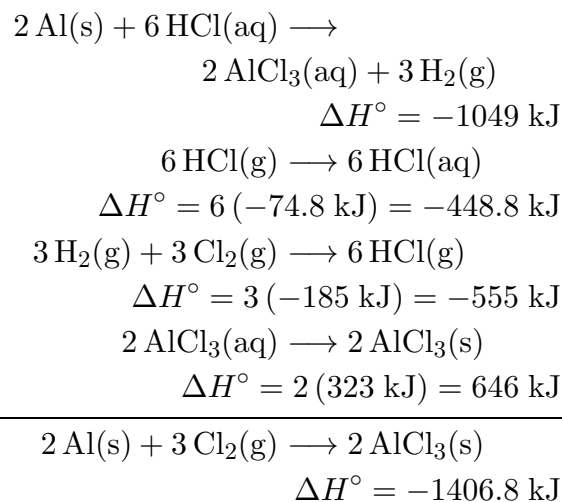
6.  $-1406.8 \text{ kJ}$  correct

7.  $-1826.2 \text{ kJ}$

**Explanation:**

Using Hess' Law:

The desired reaction is obtained by adding reaction 1; 6 times reaction 2; 3 times reaction 3; and 2 times the reverse of reaction 4:




---

**007 10.0 points**

The pressure-volume work done by an ideal gaseous system at constant volume is

1.  $-\frac{\Delta P}{P}$
2. zero correct
3.  $-\Delta E$
4.  $q$
5.  $-V \Delta P$

**Explanation:**

When  $V = \text{constant}$ , nothing moves through a distance and therefore no work is done:  $w = 0$ .

---

**008 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. –700 kJ
2. –200 kJ
3. –800 kJ **correct**
4. –750 kJ
5. +200 kJ

**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.

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

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**009 10.0 points**

Which of the following would probably have a positive  $\Delta S$  value?

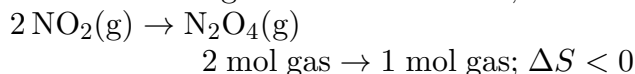
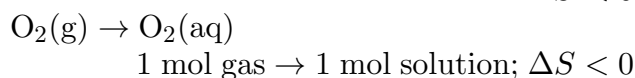
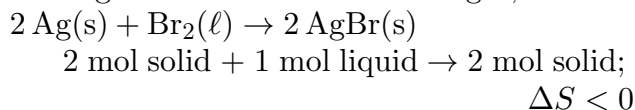
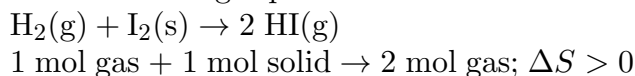
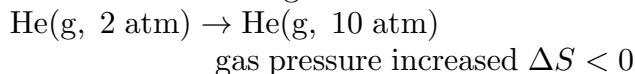
1.  $\text{He(g, 2 atm)} \rightarrow \text{He(g, 10 atm)}$
2.  $\text{H}_2(\text{g}) + \text{I}_2(\text{s}) \rightarrow 2 \text{HI(g)}$  **correct**
3.  $2 \text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$
4.  $\text{O}_2(\text{g}) \rightarrow \text{O}_2(\text{aq})$
5.  $2 \text{Ag(s)} + \text{Br}_2(\ell) \rightarrow 2 \text{AgBr(s)}$

**Explanation:**

We can predict the sign of  $\Delta S$  by noting the relative order of entropy: solids (lowest) < liquids < solutions < gases (highest) and the number of moles of each type. Also, if

a gas is compressed (pressure increased) its entropy decreases.

For the reactions given we have




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**010 10.0 points**

Which of the following statements is always true?

1. An exothermic reaction is spontaneous.
2. If the number of moles of gas does not change in a chemical reaction, then  $\Delta S^0 = 0$ .
3. A reaction for which delta  $S^0$  is positive is spontaneous.
4. If  $\Delta H^0$  and  $\Delta S^0$  are both positive,  $\Delta G^0$  will decrease as the temperature increases. **correct**

**Explanation:**

$$\Delta H = (+)$$

$$\Delta S = (+)$$

$$\begin{aligned}\Delta G &= \Delta H - T \Delta S \\ &= (+) - T (+) \\ &= (+) - T\end{aligned}$$

As  $T$  increases,  $\Delta G$  decreases.

---

**011 10.0 points**

For the reaction



$\Delta H_r^\circ = +198 \text{ kJ} \cdot \text{mol}^{-1}$  and  $\Delta S_r^\circ = 190 \text{ J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$  at 298 K. The forward reaction will be spontaneous at

1. temperatures above 1042 K. **correct**
2. all temperatures.
3. temperatures above 1315 K.
4. temperatures below 1042 K.
5. no temperature.

**Explanation:**

$$\Delta H_r^\circ = +198 \text{ kJ/mol} \quad \Delta S_r^\circ = 0.019 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}$$

$$\Delta G^0 = \Delta H^0 - T \Delta S$$

$\Delta G^0 < 0$  for a spontaneous reaction, so

$$\begin{aligned} 0 &> \Delta H^0 - T \Delta S \\ T &> \frac{\Delta H^0}{\Delta S^0} \\ &= \frac{198 \text{ kJ/mol}}{0.019 \frac{\text{kJ}}{\text{mol} \cdot \text{K}}} = 1042.11 \text{ K} \end{aligned}$$

Thus the temperature would need to be  $> 1042.11 \text{ K}$ .

---

**012 10.0 points**

Which of the following is **not** a definition of enthalpy or change in enthalpy?

1. A measure of the motional energy of a system. **correct**
2. A correction for internal energy that accounts for pressure-volume work.
3. A measure of the heat of a system at constant pressure.
4. A measure of a system's ability to change the entropy of its surroundings.

**Explanation:**

Total motional energy is a a measure of the internal energy ( $U$ , or sometimes  $E$ ) of a system, not the enthalpy ( $H$ ).

---

**013 10.0 points**

Which of the following quantities is **not** path independent?

1. entropy (S)
2. volume (V)
3. pressure (P)
4. heat (q) **correct**
5. enthalpy (H)

**Explanation:**

Heat quantitatively describes a transition **between** states, but not a state itself, and is thus a process quantity (path function), not a state function.

---

**014 10.0 points**

Which of the following would experience the smallest increase in temperature if 1 kJ of heat were added to it?

1. 1 g of copper metal
2. 10 g of water **correct**
3. 10 g of copper metal
4. 1 g of water

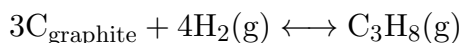
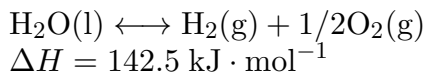
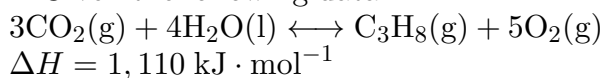
**Explanation:**

Heat capacity is an extensive property, that reflects the amount of energy required to raise the temperature of an object. Because water has a much higher heat capacity than metals (to some extent a consequence of IMF), and because 10 grams is more than 1 gram, the 10 gram sample of water would experience a much smaller increase in temperature than the other choices.

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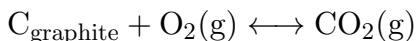
**015 10.0 points**

Given the following data:



$$\Delta H = -52 \text{ kJ} \cdot \text{mol}^{-1}$$

calculate  $\Delta H$  for the reaction



1.  $-592 \text{ kJ} \cdot \text{mol}^{-1}$
2.  $-543 \text{ kJ} \cdot \text{mol}^{-1}$
3.  $-197 \text{ kJ} \cdot \text{mol}^{-1}$  **correct**
4.  $543 \text{ kJ} \cdot \text{mol}^{-1}$
5.  $1202 \text{ kJ} \cdot \text{mol}^{-1}$

**Explanation:**

$$\begin{aligned} \Delta H_{\text{rxn}} &= \frac{11}{3}(1,110 \text{ kJ} \cdot \text{mol}^{-1}) \\ &- \frac{4}{3}(142.5 \text{ kJ} \cdot \text{mol}^{-1}) - \frac{1}{3}(-52 \text{ kJ} \cdot \text{mol}^{-1}) \\ &= -197 \text{ kJ} \cdot \text{mol}^{-1} \end{aligned}$$

---

**016 10.0 points**

Rank the following reactions from least to greatest in terms of change in entropy ( $\Delta S_{\text{rxn}}$ ):

- a)  $\text{KNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \longrightarrow \text{NaNO}_3(\text{aq}) + \text{KCl}(\text{aq})$
- b)  $4\text{Ag}(\text{s}) + \text{O}_2(\text{g}) \longrightarrow 2\text{Ag}_2\text{O}(\text{s})$
- c)  $2\text{H}_2\text{O}_2(\ell) \longrightarrow 2\text{H}_2\text{O}(\text{g}) + \text{O}_2(\text{g})$
- d)  $\text{NaHCO}_3(\text{s}) \longrightarrow \text{NaOH}(\text{s}) + \text{CO}_2(\text{g})$

1.  $b < a < c < d$
2.  $a < b < d < c$
3.  $c < d < b < a$
4.  $b < a < d < c$  **correct**
5.  $a < d < b < c$
6.  $d < c < a < b$

**Explanation:**

Reaction b has a decrease in total moles and a decrease in moles of gas;  $\Delta S_{\text{rxn}}$  is a negative number. In reaction a, no changes in phase and no changes in number of moles

occur;  $\Delta S_{\text{rxn}}$  will be nearly zero. Reaction d evolves a single mole of gas;  $\Delta S_{\text{rxn}}$  will be a small positive number. Reaction c evolves 3 moles of gas;  $\Delta S_{\text{rxn}}$  will be a large positive number.

---

**017 10.0 points**

Phosphine (the common name for  $\text{PH}_3$ , a highly toxic gas used for fumigation), has a  $\Delta H_{\text{vap}}^\circ = 14.6 \text{ kJ} \cdot \text{mol}^{-1}$  and a  $S_{\text{vap}}^\circ = 78.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$ . What is the normal boiling point of phosphine expressed in centigrade?

1.  $185.2 \text{ }^\circ\text{C}$
2.  $-0.2 \text{ }^\circ\text{C}$
3.  $-87.8 \text{ }^\circ\text{C}$  **correct**
4.  $273 \text{ }^\circ\text{C}$

**Explanation:**

Because boiling is an equilibrium process,  $\Delta G_{\text{vap}}^\circ = 0 = \Delta H_{\text{vap}}^\circ - T\Delta S_{\text{vap}}^\circ$ .

$$\begin{aligned} \text{And so } T\Delta S_{\text{vap}}^\circ &= \Delta H_{\text{vap}}^\circ \\ \text{and } T &= \Delta H_{\text{vap}}^\circ / \Delta S_{\text{vap}}^\circ \\ &= 14,600 \text{ J} \cdot \text{mol}^{-1} / 78.83 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \\ &= 185.2 \text{ K} \\ &= -87.8 \text{ }^\circ\text{C} \end{aligned}$$

---

**018 10.0 points**

Which of the following combustion reactions does not occur at high temperature?

- I)  $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + 6\text{O}_2(\text{g}) \longleftrightarrow 6\text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{g})$
- II)  $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longleftrightarrow 2\text{H}_2\text{O}(\text{g})$
- III)  $\text{C}_3\text{H}_8(\text{g}) + 5\text{O}_2(\text{g}) \longleftrightarrow 3\text{CO}_2(\text{g}) + 4\text{H}_2\text{O}(\text{g})$

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

4. I, II

5. I, II, III

6. I, III

7. I only

**Explanation:**

Only reaction II results in a decrease in entropy, and is thus becomes less spontaneous and eventually non-spontaneous as temperature increases.

---

**019 10.0 points**

(For this problem assume that the calorimeter itself absorbs no heat, the density of water is  $1.00 \text{ g} \cdot \text{mL}^{-1}$ , and the specific heat capacity of water is  $4.184 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1}$ .)

1.14 g of octane ( $\text{C}_8\text{H}_{18}$ ) is combusted in a bomb calorimeter surrounded by 1 L of water. The initial and final temperatures of the water are  $25^\circ\text{C}$  and  $38^\circ\text{C}$  respectively. Determine the molar enthalpy of combustion of octane.

1.  $-5,440 \text{ kJ} \cdot \text{mol}^{-1}$  **correct**2.  $-54,400 \text{ kJ} \cdot \text{mol}^{-1}$ 3.  $-54.4 \text{ kJ} \cdot \text{mol}^{-1}$ 4.  $-544 \text{ kJ} \cdot \text{mol}^{-1}$ 5.  $-5.44 \text{ kJ} \cdot \text{mol}^{-1}$ **Explanation:**

$\Delta T = T_f - T_i = 38^\circ\text{C} - 25^\circ\text{C} = 13^\circ\text{C} = 12.98 \text{ K}$

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

$$n = 1.14 \text{ g octane} \cdot \frac{114 \text{ g}}{\text{mol}} = 0.01 \text{ mol}$$

$$\Delta H_{rxn} = -\Delta H_{cal} = -mc\Delta T$$

$$\frac{-mc\Delta T}{n} = -\frac{1000 \text{ g} \cdot \frac{4.184 \text{ J}}{\text{g}\cdot\text{K}} \cdot 13 \text{ K}}{0.01 \text{ mol}}$$

$$= -5,440 \text{ kJ} \cdot \text{mol}^{-1}$$

---

**020 10.0 points**

If both  $\Delta n_{\text{gas}}$  and  $\Delta n$  for a particular reaction are zero, which statement below would be the most accurate prediction we could make about the change in entropy?

1. It is definitely positive.

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

3. It is definitely negative.

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

**Explanation:**

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

---

**021 10.0 points**

Which of the following statements concerning internal energy is/are true?

I) If the expansion work is small,  $\Delta H$  and  $\Delta U$  are close in value.II)  $\Delta U$  for a system is equal to  $q$  at constant volume.III) Assuming no heat is exchanged, when pressure-volume work is done on the system,  $\Delta U$  is positive.1. I, II, III **correct**

2. II only

3. III only

4. I, III

5. I only

6. II, III

**7. I, II****Explanation:**

Statement I follows from the identity  $\Delta H = \Delta U + p\Delta V$ , because  $p\Delta V$  is the expansion work. Statement II follows from  $\Delta U = q + w$ , because  $w = 0$  for processes that occur at constant volume. Statement III also follows from  $\Delta U = q + w$ , because  $w$  is the pressure-volume work. (Note: for reversible processes, expansion work and pressure-volume work are identical.)