

## Selected End-of-Chapter Problems for Ch. 6

34. Calculate the internal energy change for each of the following.

- One hundred (100.) joules of work is required to compress a gas. At the same time, the gas releases 23 J of heat.
- A piston is compressed from a volume of 8.30 L to 2.80 L against a constant pressure of 1.90 atm. In the process, there is a heat gain by the system of 350. J.
- A piston expands against 1.00 atm of pressure from 11.2 L to 29.1 L. In the process, 1037 J of heat is absorbed.

Answer:

a.  $\Delta E = q + w = -23 \text{ J} + 100. \text{ J} = 77 \text{ J}$

b.  $w = -P\Delta V = -1.90 \text{ atm}(2.80 \text{ L} - 8.30 \text{ L}) = 10.5 \text{ L atm} \times \frac{101.3 \text{ J}}{\text{L atm}} = 1060 \text{ J}$

$$\Delta E = q + w = 350. \text{ J} + 1060 = 1410 \text{ J}$$

c.  $w = -P\Delta V = -1.00 \text{ atm}(29.1 \text{ L} - 11.2 \text{ L}) = -17.9 \text{ L atm} \times \frac{101.3 \text{ J}}{\text{L atm}} = -1810 \text{ J}$

$$\Delta E = q + w = 1037 \text{ J} - 1810 \text{ J} = -770 \text{ J}$$

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36. A piston performs work of 210. L · atm on the surroundings, while the cylinder in which it is placed expands from 10. L to 25 L. At the same time, 45 J of heat is transferred from the surroundings to the system. Against what pressure was the piston working?

Answer:

$$w = -210. \text{ J} = -P\Delta V, -210 \text{ J} = -P(25 \text{ L} - 10. \text{ L}), P = 14 \text{ atm}$$

38. As a system increases in volume, it absorbs 52.5 J of energy in the form of heat from the surroundings. The piston is working against a pressure of 0.500 atm. The final volume of the system is 58.0 L. What was the initial volume of the system if the internal energy of the system decreased by 102.5 J?

Answer:

$$\Delta E = q + w, -102.5 \text{ J} = 52.5 \text{ J} + w, w = -155.0 \text{ J} \times \frac{1 \text{ L atm}}{101.3 \text{ J}} = -1.530 \text{ L atm}$$

$$w = -P\Delta V, -1.530 \text{ L atm} = -0.500 \text{ atm} \times \Delta V, \Delta V = 3.06 \text{ L}$$

$$\Delta V = V_f - V_i, 3.06 \text{ L} = 58.0 \text{ L} - V_i, V_i = 54.9 \text{ L} = \text{initial volume}$$

40. One mole of  $\text{H}_2\text{O}(g)$  at 1.00 atm and  $100.^\circ\text{C}$  occupies a volume of 30.6 L.

When 1 mole of  $\text{H}_2\text{O}(g)$  is condensed to 1 mole of  $\text{H}_2\text{O}(l)$  at 1.00 atm and  $100.^\circ\text{C}$ , 40.66 kJ of heat is released. If the density of  $\text{H}_2\text{O}(l)$  at this temperature and pressure is  $0.996 \text{ g/cm}^3$ , calculate  $\Delta E$  for the condensation of 1 mole of water at 1.00 atm and  $100.^\circ\text{C}$ .

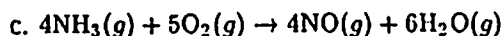
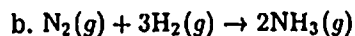
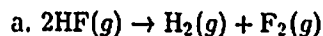
Answer:  $\text{H}_2\text{O}(g) \rightarrow \text{H}_2\text{O}(l)$ :  $\Delta E = q + w$ ;  $q = -40.66 \text{ kJ}$ ;  $w = -P\Delta V$

$$\text{Volume of 1 mol } \text{H}_2\text{O}(l) = 1.000 \text{ mol } \text{H}_2\text{O}(l) \times \frac{18.02 \text{ g}}{\text{mol}} \times \frac{1 \text{ cm}^3}{0.996 \text{ g}} = 18.1 \text{ cm}^3 = 18.1 \text{ mL}$$

$$w = -P\Delta V = -1.00 \text{ atm} \times (0.0181 \text{ L} - 30.6 \text{ L}) = 30.6 \text{ L atm} \times \frac{101.3 \text{ J}}{\text{L atm}} = 3.10 \times 10^3 \text{ J} \\ = 3.10 \text{ kJ}$$

$$\Delta E = q + w = -40.66 \text{ kJ} + 3.10 \text{ kJ} = -37.56 \text{ kJ}$$

50. For the following reactions at constant pressure, predict if  $\Delta H > \Delta E$ ,  $\Delta H < \Delta E$ , or  $\Delta H = \Delta E$ .



Answer:  $\Delta H = \Delta E + P\Delta V$ ; from this equation,  $\Delta H > \Delta E$  when  $\Delta V > 0$ ,  $\Delta H < \Delta E$  when  $\Delta V < 0$ , and  $\Delta H = \Delta E$  when  $\Delta V = 0$ . Concentrate on the moles of gaseous products versus the moles of gaseous reactants to predict  $\Delta V$  for a reaction.

a. There are 2 moles of gaseous reactants converting to 2 moles of gaseous products, so  $\Delta V = 0$ . For this reaction,  $\Delta H = \Delta E$ .

b. There are 4 moles of gaseous reactants converting to 2 moles of gaseous products, so  $\Delta V < 0$  and  $\Delta H < \Delta E$ .

c. There are 9 moles of gaseous reactants converting to 10 moles of gaseous products, so  $\Delta V > 0$  and  $\Delta H > \Delta E$ .

52. The specific heat capacity of silver is  $0.24 \text{ J/}^\circ\text{C} \cdot \text{g}$ .

- Calculate the energy required to raise the temperature of  $150.0 \text{ g Ag}$  from  $273 \text{ K}$  to  $298 \text{ K}$ .
- Calculate the energy required to raise the temperature of  $1.0 \text{ mole of Ag}$  by  $1.0^\circ\text{C}$  (called the *molar heat capacity* of silver).
- It takes  $1.25 \text{ kJ}$  of energy to heat a sample of pure silver from  $12.0^\circ\text{C}$  to  $15.2^\circ\text{C}$ . Calculate the mass of the sample of silver.

Answer:

a.  $s = \text{specific heat capacity} = \frac{0.24 \text{ J}}{^\circ\text{C g}} = \frac{0.24 \text{ J}}{\text{K g}}$  since  $\Delta T(\text{K}) = \Delta T(^\circ\text{C})$

$$\text{Energy} = s \times m \times \Delta T = \frac{0.24 \text{ J}}{^\circ\text{C g}} \times 150.0 \text{ g} \times (298 \text{ K} - 273 \text{ K}) = 9.0 \times 10^3 \text{ J}$$

b.  $\text{Molar heat capacity} = \frac{0.24 \text{ J}}{^\circ\text{C g}} \times \frac{107.9 \text{ g Ag}}{\text{mol Ag}} = \frac{26 \text{ J}}{^\circ\text{C mol}}$

c.  $1250 \text{ J} = \frac{0.24 \text{ J}}{^\circ\text{C g}} \times m \times (15.2^\circ\text{C} - 12.0^\circ\text{C})$ .  $m = \frac{1250}{0.24 \times 3.2} = 1.6 \times 10^3 \text{ g Ag}$

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56. A biology experiment requires the preparation of a water bath at  $37.0^\circ\text{C}$  (body temperature). The temperature of the cold tap water is  $22.0^\circ\text{C}$ , and the temperature of the hot tap water is  $55.0^\circ\text{C}$ . If a student starts with  $90.0 \text{ g}$  cold water, what mass of hot water must be added to reach  $37.0^\circ\text{C}$ ?

Answer:

Heat loss by hot water = heat gain by cold water

Since  $q$  is defined as "heat gained", we need to put a minus sign in front of "heat loss of hot water".

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

$$-s \times m_{\text{hot}} \times (T_f - T_{i,\text{hot}}) = s \times m_{\text{cold}} \times (T_f - T_{i,\text{cold}})$$

$$-\frac{4.18 \text{ J}}{^\circ\text{C g}} \times m_{\text{hot}} \times (37.0^\circ\text{C} - 55.0^\circ\text{C}) = \frac{4.18 \text{ J}}{^\circ\text{C g}} \times 90.0 \text{ g} \times (37.0^\circ\text{C} - 22.0^\circ\text{C})$$

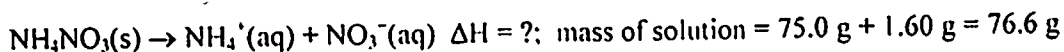
This is equivalent to keeping all quantities positive, which some may find easier to deal with:

$$\frac{4.18 \text{ J}}{^\circ\text{C g}} \times m_{\text{hot}} \times (55.0^\circ\text{C} - 37.0^\circ\text{C}) = \frac{4.18 \text{ J}}{^\circ\text{C g}} \times 90.0 \text{ g} \times (37.0^\circ\text{C} - 22.0^\circ\text{C})$$

$$m_{\text{hot}} = \frac{90.0 \text{ g} \times 15.0^\circ\text{C}}{18.0^\circ\text{C}} = 75.0 \text{ g hot water needed}$$

64. In a coffee-cup calorimeter, 1.60 g  $\text{NH}_4\text{NO}_3$  is mixed with 75.0 g water at an initial temperature of  $25.00^\circ\text{C}$ . After dissolution of the salt, the final temperature of the calorimeter contents is  $23.34^\circ\text{C}$ . Assuming the solution has a heat capacity of  $4.18 \text{ J}/^\circ\text{C} \cdot \text{g}$  and assuming no heat loss to the calorimeter, calculate the enthalpy change for the dissolution of  $\text{NH}_4\text{NO}_3$  in units of  $\text{kJ/mol}$ .

Answer:



Heat lost by solution = heat gained as  $\text{NH}_4\text{NO}_3$  dissolves. To keep things simple, we will keep all quantities positive ( $q$  and  $\Delta T$ ) and then deduce the correct sign for  $\Delta H$  at the end of the problem. Here, because temperature decreases as  $\text{NH}_4\text{NO}_3$  dissolves, heat is absorbed as  $\text{NH}_4\text{NO}_3$  dissolves, so this is an endothermic process ( $\Delta H$  is positive).

$$\begin{aligned} \text{Heat lost by solution} &= \frac{4.18 \text{ J}}{^\circ\text{C} \cdot \text{g}} \times 76.6 \text{ g} \times (25.00 - 23.34)^\circ\text{C} = 532 \text{ J} \\ &= \text{heat gained by } \text{NH}_4\text{NO}_3 \text{ as it dissolves} \end{aligned}$$

$$\Delta H = \frac{532 \text{ J}}{1.60 \text{ g } \text{NH}_4\text{NO}_3} \times \frac{80.05 \text{ g } \text{NH}_4\text{NO}_3}{\text{mol } \text{NH}_4\text{NO}_3} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 26.6 \text{ kJ/mol } \text{NH}_4\text{NO}_3 \text{ dissolving}$$

68. The combustion of 0.1584 g benzoic acid increases the temperature of a bomb calorimeter by  $2.54^\circ\text{C}$ . Calculate the heat capacity of this calorimeter. (The energy released by combustion of benzoic acid is  $26.42 \text{ kJ/g}$ .) A 0.2130-g sample of vanillin ( $\text{C}_8\text{H}_8\text{O}_3$ ) is then burned in the same calorimeter, and the temperature increases by  $3.25^\circ\text{C}$ . What is the energy of combustion per gram of vanillin? Per mole of vanillin?

Answer:

First, we need to get the heat capacity of the calorimeter from the combustion of benzoic acid. Heat lost by combustion = heat gained by calorimeter.

$$\text{Heat loss} = 0.1584 \text{ g} \times \frac{26.42 \text{ kJ}}{\text{g}} = 4.185 \text{ kJ}$$

$$\text{Heat gain} = 4.185 \text{ kJ} = C_{\text{cal}} \times \Delta T. \quad C_{\text{cal}} = \frac{4.185 \text{ kJ}}{2.54^\circ\text{C}} = 1.65 \text{ kJ}/^\circ\text{C}$$

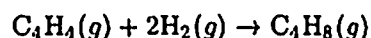
Now we can calculate the heat of combustion of vanillin. Heat loss = heat gain.

$$\text{Heat gain by calorimeter} = \frac{1.65 \text{ kJ}}{^\circ\text{C}} \times 3.25^\circ\text{C} = 5.36 \text{ kJ}$$

Heat loss = 5.36 kJ, which is the heat evolved by combustion of the vanillin.

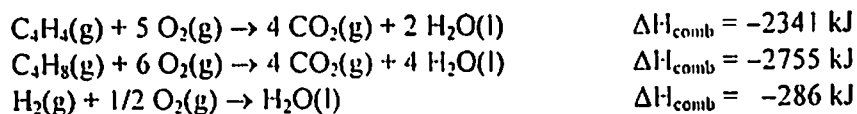
$$\Delta E_{\text{comb}} = \frac{-5.36 \text{ kJ}}{0.2130 \text{ g}} = -25.2 \text{ kJ/g}; \quad \Delta E_{\text{comb}} = \frac{-25.2 \text{ kJ}}{\text{g}} \times \frac{152.14 \text{ g}}{\text{mol}} = -3830 \text{ kJ/mol}$$

70. Combustion reactions involve reacting a substance with oxygen. When compounds containing carbon and hydrogen are combusted, carbon dioxide and water are the products. Using the enthalpies of combustion for  $C_4H_4$  ( $-2341 \text{ kJ/mol}$ ),  $C_4H_8$  ( $-2755 \text{ kJ/mol}$ ), and  $H_2$  ( $-286 \text{ kJ/mol}$ ), calculate  $\Delta H$  for the reaction



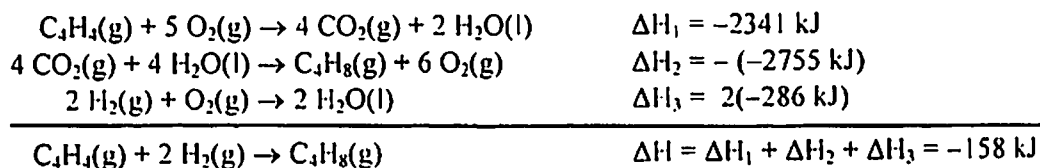
Answer:

Given:

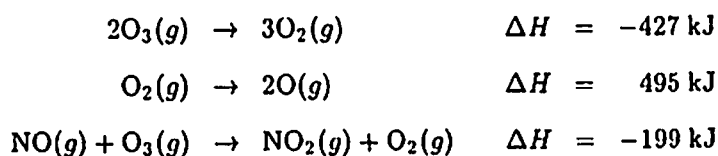


By convention,  $H_2O(l)$  is produced when enthalpies of combustion are given, and because per-mole quantities are given, the combustion reaction refers to 1 mole of that quantity reacting with  $O_2(g)$ .

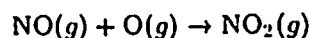
Using Hess's law to solve:



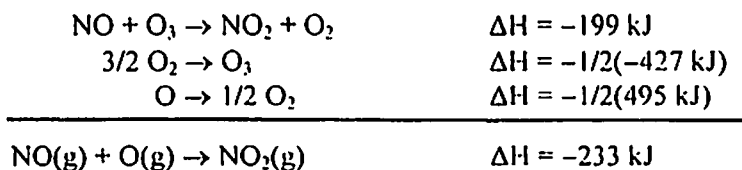
73. Given the following data



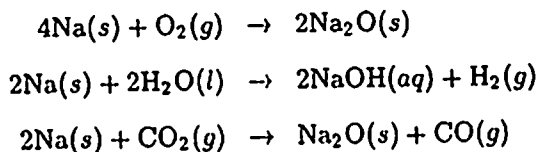
calculate  $\Delta H$  for the reaction



Answer:

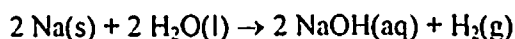
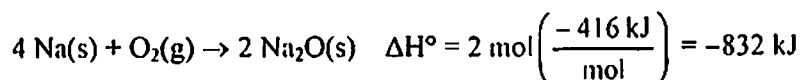


82. Calculate  $\Delta H^\circ$  for each of the following reactions using the data in Appendix 4:

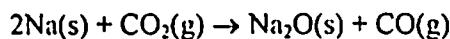


Explain why a water or carbon dioxide fire extinguisher might not be effective in putting out a sodium fire.

Answer:



$$\Delta H^\circ = \left[ 2 \text{ mol} \left( \frac{-470. \text{ kJ}}{\text{mol}} \right) \right] - \left[ 2 \text{ mol} \left( \frac{-286 \text{ kJ}}{\text{mol}} \right) \right] = -368 \text{ kJ}$$



$$\Delta H^\circ = \left[ 1 \text{ mol} \left( \frac{-416 \text{ kJ}}{\text{mol}} \right) + 1 \text{ mol} \left( \frac{-110.5 \text{ kJ}}{\text{mol}} \right) \right] - \left[ 1 \text{ mol} \left( \frac{-393.5 \text{ kJ}}{\text{mol}} \right) \right] = -133 \text{ kJ}$$

In Reactions 2 and 3, sodium metal reacts with the "extinguishing agent." Both reactions are exothermic, and each reaction produces a flammable gas,  $\text{H}_2$  and  $\text{CO}$ , respectively.

From Appendix 4:

Substance and State	$\Delta H_f^\circ$ (kJ/mol)
$\text{CO}(g)$	-110.5
$\text{CO}_2(g)$	-393.5
$\text{H}_2(g)$	0
$\text{H}_2\text{O}(l)$	-286
$\text{Na}(s)$	0
$\text{Na}_2\text{O}(s)$	-416
$\text{NaOH}(aq)$	-470
$\text{O}_2(g)$	0