Selected End-of-Chapter Problems for Ch. 6

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

    a. One hundred (100.) joules of work is required to compress a gas. At the same time, the gas releases 23 J of heat.

    b. 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.

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

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

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

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

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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:

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

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

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

40. One mole of $H_2O(g)$ at 1.00 atm and 100.°C occupies a volume of 30.6 L. When 1 mole of $H_2O(g)$ is condensed to 1 mole of $H_2O(l)$ at 1.00 atm and 100.°C, 40.66 kJ of heat is released. If the density of $H_2O(l)$ at this temperature and pressure is 0.996 g/cm^3, calculate ΔE for the condensation of 1 mole of water at 1.00 atm and 100.°C.

Answer: $H_2O(g) \rightarrow H_2O(l)$: $\Delta E = q + w$; $q = -40.66\ \text{kJ}$; $w = -P\Delta V$

Volume of 1 mol $H_2O(l)$ = $1.000\ \text{mol}\ H_2O(l) \times \dfrac{18.02\ \text{g}}{\text{mol}} \times \dfrac{1\ \text{cm}^3}{0.996\ \text{g}} = 18.1\ \text{cm}^3 = 18.1\ \text{mL}$

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

$= 3.10\ \text{kJ}$

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

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

a. $2HF(g) \rightarrow H_2(g) + F_2(g)$

b. $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

c. $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

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 ΔV 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 J/°C · g.

a. Calculate the energy required to raise the temperature of 150.0 g Ag from 273 K to 298 K.

b. Calculate the energy required to raise the temperature of 1.0 mole of Ag by 1.0°C (called the *molar heat capacity* of silver).

c. It takes 1.25 kJ of energy to heat a sample of pure silver from 12.0°C to 15.2°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 $\Delta T(\text{K}) = \Delta T(^\circ\text{C})$

$$\text{Energy} = s \times m \times \Delta T = \frac{0.24\text{ J}}{^\circ\text{C g}} \times 150.0\text{ g} \times (298\text{ K} - 273\text{ K}) = 9.0 \times 10^2\text{ J}$$

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

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

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

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

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

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

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

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

64. In a coffee-cup calorimeter, 1.60 g NH_4NO_3 is mixed with 75.0 g water at an initial temperature of 25.00°C. After dissolution of the salt, the final temperature of the calorimeter contents is 23.34°C. Assuming the solution has a heat capacity of 4.18 J/°C · g and assuming no heat loss to the calorimeter, calculate the enthalpy change for the dissolution of NH_4NO_3 in units of kJ/mol.

Answer:

$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$ ΔH = ?; mass of solution = 75.0 g + 1.60 g = 76.6 g

Heat lost by solution = heat gained as NH_4NO_3 dissolves. To keep things simple, we will keep all quantities positive (q and ΔT) and then deduce the correct sign for ΔH at the end of the problem. Here, because temperature decreases as NH_4NO_3 dissolves, heat is absorbed as NH_4NO_3 dissolves, so this is an endothermic process (ΔH is positive).

$$\text{Heat lost by solution} = \frac{4.18\ \text{J}}{°\text{C g}} \times 76.6\ \text{g} \times (25.00 - 23.34)°\text{C} = 532\ \text{J}$$

$$= \text{heat gained by } NH_4NO_3 \text{ as it dissolves}$$

$$\Delta H = \frac{532\ \text{J}}{1.60\ \text{g}\ NH_4NO_3} \times \frac{80.05\ \text{g}\ NH_4NO_3}{\text{mol}\ NH_4NO_3} \times \frac{1\ \text{kJ}}{1000\ \text{J}} = 26.6\ \text{kJ/mol}\ NH_4NO_3 \text{ dissolving}$$

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68. The combustion of 0.1584 g benzoic acid increases the temperature of a bomb calorimeter by 2.54°C. Calculate the heat capacity of this calorimeter. (The energy released by combustion of benzoic acid is 26.42 kJ/g.) A 0.2130-g sample of van illin ($C_8H_8O_3$) is then burned in the same calorimeter, and the temperature increases by 3.25°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.

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

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

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

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

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 kJ/mol), C_4H_8 (−2755 kJ/mol), and H_2 (−286 kJ/mol), calculate ΔH for the reaction

$$C_4H_4(g) + 2H_2(g) \rightarrow C_4H_8(g)$$

Answer:

Given:

$C_4H_4(g) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 2\ H_2O(l)$	$\Delta H_{comb} = -2341$ kJ
$C_4H_8(g) + 6\ O_2(g) \rightarrow 4\ CO_2(g) + 4\ H_2O(l)$	$\Delta H_{comb} = -2755$ kJ
$H_2(g) + 1/2\ O_2(g) \rightarrow H_2O(l)$	$\Delta H_{comb} = -286$ kJ

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:

$C_4H_4(g) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 2\ H_2O(l)$	$\Delta H_1 = -2341$ kJ
$4\ CO_2(g) + 4\ H_2O(l) \rightarrow C_4H_8(g) + 6\ O_2(g)$	$\Delta H_2 = -(-2755$ kJ$)$
$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$	$\Delta H_3 = 2(-286$ kJ$)$
$C_4H_4(g) + 2\ H_2(g) \rightarrow C_4H_8(g)$	$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 = -158$ kJ

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73. Given the following data

$2O_3(g) \rightarrow 3O_2(g)$	$\Delta H = -427$ kJ
$O_2(g) \rightarrow 2O(g)$	$\Delta H = 495$ kJ
$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$	$\Delta H = -199$ kJ

calculate ΔH for the reaction

$$NO(g) + O(g) \rightarrow NO_2(g)$$

Answer:

$NO + O_3 \rightarrow NO_2 + O_2$	$\Delta H = -199$ kJ
$3/2\ O_2 \rightarrow O_3$	$\Delta H = -1/2(-427$ kJ$)$
$O \rightarrow 1/2\ O_2$	$\Delta H = -1/2(495$ kJ$)$
$NO(g) + O(g) \rightarrow NO_2(g)$	$\Delta H = -233$ kJ

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

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

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

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

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

Answer:

$$4\ \text{Na(s)} + \text{O}_2\text{(g)} \rightarrow 2\ \text{Na}_2\text{O(s)} \quad \Delta H° = 2\ \text{mol}\left(\frac{-416\ \text{kJ}}{\text{mol}}\right) = -832\ \text{kJ}$$

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

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

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

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

In Reactions 2 and 3, sodium metal reacts with the "extinguishing agent." Both reactions are exothermic, and each reaction produces a flammable gas, H_2 and CO, respectively.

From Appendix 4:

Substance and State	ΔH_f° (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