

20. Ball A: $PE = mgz = 2.00 \text{ kg} \times \frac{9.8 \text{ m}}{\text{s}^2} \times 10.0 \text{ m} = \frac{196 \text{ kg m}^2}{\text{s}^2} = 196 \text{ J}$

At Point I: All of this energy is transferred to Ball B. All of B's energy is kinetic energy at this point. $E_{\text{total}} = KE = 196 \text{ J}$. At point II, the sum of the total energy will equal 196 J.

At Point II: $PE = mgz = 4.00 \text{ kg} \times \frac{9.8 \text{ m}}{\text{s}^2} \times 3.00 \text{ m} = 118 \text{ J}$

$$KE = E_{\text{total}} - PE = 196 \text{ J} - 118 \text{ J} = 78 \text{ J}$$

22. a. $\Delta E = q + w = -47 \text{ kJ} + 88 \text{ kJ} = 41 \text{ kJ}$

b. $\Delta E = 82 + 47 = 129 \text{ kJ}$

c. $\Delta E = 47 + 0 = 47 \text{ kJ}$

d. When the surroundings deliver work to the system, $w > 0$. This is the case for a and b.

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

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

$$w = -P\Delta V = -1.00 \text{ atm} \times (0.0181 \text{ L} - 30.6 \text{ L}) = 30.6 \text{ L atm} \times \frac{101.3 \text{ J}}{\text{L atm}} = 3.10 \times 10^3 \text{ J} = 3.1 \text{ kJ}$$

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

34. $4 \text{ Fe(s)} + 3 \text{ O}_2\text{(g)} \rightarrow 2 \text{ Fe}_2\text{O}_3\text{(s)}$ $\Delta H = -1652 \text{ kJ}$; Note that 1652 kJ of heat are released when 4 mol Fe reacts with 3 mol O_2 to produce 2 mol Fe_2O_3 .

a. $4.00 \text{ mol Fe} \times \frac{-1652 \text{ kJ}}{4 \text{ mol Fe}} = -1650 \text{ kJ heat released}$

b. $1.00 \text{ mol Fe}_2\text{O}_3 \times \frac{-1652 \text{ kJ}}{2 \text{ mol Fe}_2\text{O}_3} = -826 \text{ kJ heat released}$

c. $1.00 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g}} \times \frac{-1652 \text{ kJ}}{4 \text{ mol Fe}} = -7.39 \text{ kJ heat released}$

d. $10.0 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g}} = 0.179 \text{ mol Fe}$; $2.00 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g}} = 0.0625 \text{ mol O}_2$

$0.179 \text{ mol Fe} / 0.0625 \text{ mol O}_2 = 2.86$; The balanced equation requires a 4 mol Fe/3 mol $\text{O}_2 = 1.33$ mol ratio. O_2 is limiting since the actual mol Fe/mol O_2 ratio is less than the required mol ratio.

$$0.0625 \text{ mol O}_2 \times \frac{-1652 \text{ kJ}}{3 \text{ mol O}_2} = -34.4 \text{ kJ heat released}$$

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

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

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

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

42. Heat loss by Ni = Heat gain by water; Keeping ΔT values positive to avoid sign errors:

$$\frac{0.444 \text{ J}}{\text{g } ^\circ\text{C}} \times 15.0 \text{ g} \times (100.0^{\circ}\text{C} - T_f) = \frac{4.18 \text{ J}}{\text{g } ^\circ\text{C}} \times 55.0 \text{ g} \times (T_f - 23.0^{\circ}\text{C})$$

$$666 - 6.66 T_f = 230. T_f - 5290, T_f = \frac{5960}{237} = 25.1^{\circ}\text{C}$$

48. $0.100 \text{ L} \times \frac{0.500 \text{ mol HCl}}{\text{L}} = 5.00 \times 10^{-2} \text{ mol HCl}$

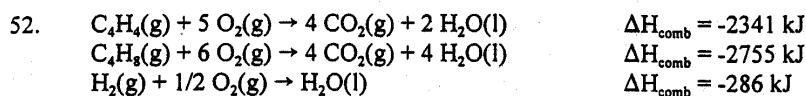
$$0.300 \text{ L} \times \frac{0.500 \text{ mol Ba(OH)}_2}{\text{L}} = 0.150 \text{ mol Ba(OH)}_2$$

To react with all the HCl present, $5.00 \times 10^{-2} / 2 = 2.50 \times 10^{-2} \text{ mol Ba(OH)}_2$ are required. Since 0.150 mol Ba(OH)₂ are present, then HCl is the limiting reactant.

$$5.00 \times 10^{-2} \text{ mol HCl} \times \frac{118 \text{ kJ}}{2 \text{ mol HCl}} = 2.95 \text{ kJ of heat is evolved by reaction.}$$

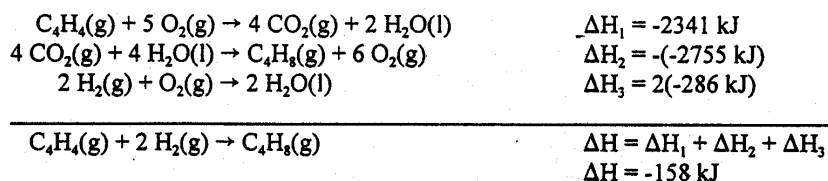
$$\text{Heat gained by solution} = 2.95 \times 10^3 \text{ J} = \frac{4.18 \text{ J}}{\text{g } ^\circ\text{C}} \times 400.0 \text{ g} \times \Delta T$$

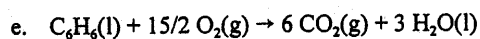
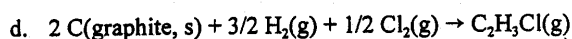
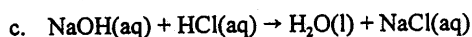
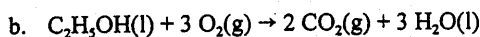
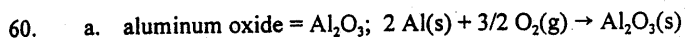
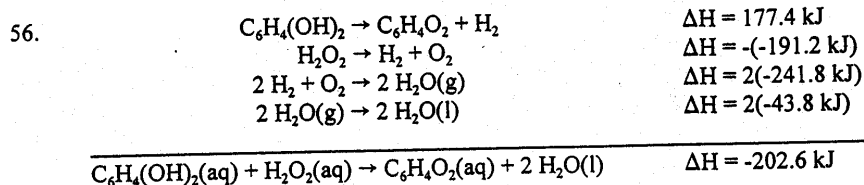
$$\Delta T = 1.76^{\circ}\text{C} = T_f - T_i = T_f - 25.0^{\circ}\text{C}, T_f = 26.8^{\circ}\text{C}$$



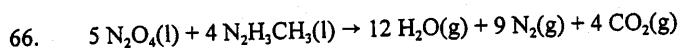
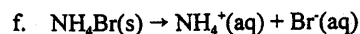
By convention, H₂O(l) is produced when enthalpies of combustion are given and since per mol quantities are given, the combustion reaction refers to 1 mol of that quantity reacting with O₂(g).

Using Hess's Law to solve:

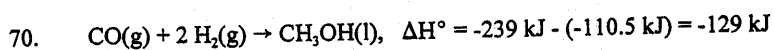




Note: ΔH_{comb} values assume one mol of compound combusted.



$$\Delta H^\circ = \left[12 \text{ mol} \left(\frac{-242 \text{ kJ}}{\text{mol}} \right) + 4 \text{ mol} \left(\frac{-393.5 \text{ kJ}}{\text{mol}} \right) \right] - \left[5 \text{ mol} \left(\frac{-20. \text{ kJ}}{\text{mol}} \right) + 4 \text{ mol} \left(\frac{54 \text{ kJ}}{\text{mol}} \right) \right] = -4594 \text{ kJ}$$



74. Mass of $H_2O = 1.00 \text{ gal} \times \frac{3.785 \text{ L}}{\text{gal}} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.00 \text{ g}}{\text{mL}} = 3790 \text{ g } H_2O$

Energy required (theoretical) = $s \times m \times \Delta T = \frac{4.18 \text{ J}}{\text{g } ^\circ\text{C}} \times 3790 \text{ g} \times 10.0 \text{ } ^\circ\text{C} = 1.58 \times 10^5 \text{ J}$

For an actual (80.0% efficient) process, more than this quantity of energy is needed since heat is always lost in any transfer of energy. The energy required is:

$$1.58 \times 10^5 \text{ J} \times \frac{100. \text{ J}}{80.0 \text{ J}} = 1.98 \times 10^5 \text{ J}$$

$$\text{Mass of } C_2H_2 = 1.98 \times 10^5 \text{ J} \times \frac{1 \text{ mol } C_2H_2}{1300. \times 10^3 \text{ J}} \times \frac{26.04 \text{ g } C_2H_2}{\text{mol } C_2H_2} = 3.97 \text{ g } C_2H_2$$