Chapter 5: Thermochemistry

5.1 Energy Basics

Question 28-1.

A burning match and a bonfire may have the same temperature, yet you would not sit around a burning match on a fall evening to stay warm. Why not?

Solution

The temperature of 1 gram of burning wood is approximately the same for both a match and a bonfire. This is an intensive property and depends on the material (wood). However, the overall amount of produced heat depends on the amount of material; this is an extensive property. The amount of wood in a bonfire is much greater than that in a match; the total amount of produced heat is also much greater, which is why we can sit around a bonfire to stay warm, but a match would not provide enough heat to keep us from getting cold.

Question 28-2.

Prepare a table identifying several energy transitions that take place during the typical operation of an automobile.

Solution

Chemical (potential) energy of the car battery is converted into mechanical (kinetic) energy of the engine starter when the ignition switch is turned or pressed; chemical (potential) energy in the gasoline is converted into heat and mechanical energy through its combustion in the auto's engine; kinetic energy of the automobile is converted into heat created by friction of the brake pads during braking.

Question 28-3.

Explain the difference between heat capacity and specific heat of a substance.

Solution

Heat capacity refers to the heat required to raise the temperature of the mass of the substance 1 degree; specific heat refers to the heat required to raise the temperature of 1 gram of the substance 1 degree. Thus, heat capacity is an extensive property, and specific heat is an intensive one.

Question 28-4.

Calculate the heat capacity, in joules and in calories per degree, of the following:
   (a) 28.4 g of water
   (b) 1.00 oz of lead

Solution

(a) 28.4 g × 4.184 J/g °C = 119 J/°C; 28.4 g × 1.000 cal/g °C = 28.4 cal/°C; (b) 28.4 g × 0.129 J/g °C = 3.66 J/°C; $28.4 \text{ g} \times 0.129 \text{ J g}^{-1} \text{ °C}^{-1} \times \dfrac{(1 \text{ cal})}{4.184 \text{ J}} = 0.882 \text{ cal °C}^{-1}$

Question 28-5.

Calculate the heat capacity, in joules and in calories per degree, of the following:
      (a) 45.8 g of nitrogen gas
      (b) 1.00 pound of aluminum metal

Solution

(a) 45.8 g × 1.04 J/g°C = 47.6 J/°C; 45.8 g × 1.04 J g^{-1} °C^{-1} × $\dfrac{1 \text{ cal}}{4.184 \text{ J}}$ = 11.38 J °C^{-1}; (b)

454 g × 0.897 J/g °C = 407 J/°C; 454 g × 0.897 J g^{-1} °C^{-1} × $\dfrac{1 \text{ cal}}{4.184 \text{ J}}$ = 97.3 cal °C^{-1}

Question 28-6.

How much heat, in joules and in calories, must be added to a 75.0-g iron block with a specific heat of 0.449 J/g °C to increase its temperature from 25 °C to its melting temperature of 1535 °C?

Solution

$q = cm\Delta T$; $q = 0.449$ J/g °C × 75.0 g × (1535 − 25) °C= 5.08 × 10^4 J; the conversion is 4.184 J =

1 cal; 5.08 × 10^4 J × $\dfrac{1 \text{ cal}}{4.184 \text{ J}}$ = 1.21 × 10^4 cal

Question 28-7.

How much heat, in joules and in calories, is required to heat a 28.4-g (1-oz) ice cube from –23.0 °C to –1.0 °C?

Solution

$q = cm\Delta T$; $q = 2.093$ J/g °C × 28.4 g × [–1 – (–23)] °C = 1308 J (or 1310 with the correct number of significant digits); the conversion factor is 4.184 J = 1 cal; 1308 J × 1 cal / 4.184 J = 312.6 cal (313 calories with the correct number of significant digits)

Question 28-8.

How much would the temperature of 275 g of water increase if 36.5 kJ of heat were added?

Solution

$q = cm\Delta T$

36,500 J = 4.184 J g^{-1} °C^{-1} × 2.75 × 10^2 g × ΔT

$\Delta T = \dfrac{36,500 \text{ °C}}{4.184 \times 2.75 \times 10^2}$ = 31.7 °C

Question 28-9.

If 14.5 kJ of heat were added to 485 g of liquid water, how much would its temperature increase?

Solution

$q = cm\Delta T$

14,500 J = 4.184 J g^{-1} °C × 4.85 × 10^2 g × ΔT

$\Delta T = \dfrac{14,500 \text{ °C}}{4.184 \times 4.85 \times 10^2}$ = 7.15 °C

Question 28-10.

A piece of unknown substance weighs 44.7 g and requires 2110 J to increase its temperature from 23.2 °C to 89.6 °C.

(a) What is the specific heat of the substance?

(b) If it is one of the substances found in Table 5.1, what is its likely identity?

Solution

(a) $q = cm\Delta T$, $c = q/(m\Delta T) = 2110$ J / [44.7 g (89.6 – 23.2] °C) = 0.711 J/g °C; (b) According to the data in the table, silicon seems to be a likely candidate.

Question 28-11.

A piece of unknown solid substance weighs 437.2 g, and requires 8460 J to increase its temperature from 19.3 °C to 68.9 °C.

(a) What is the specific heat of the substance?

(b) If it is one of the substances found in Table 5.1, what is its likely identity?

Solution

(a) $q = cm\Delta T$, $c = q/(m\Delta T) = 8460$ J / [437.2 g (68.9 – 19.3) °C] = 0.390 J/g °C; (b) Copper is a likely candidate.

Question 28-12.

An aluminum kettle weighs 1.05 kg.

(a) What is the heat capacity of the kettle?

(b) How much heat is required to increase the temperature of this kettle from 23.0 °C to 99.0 °C?

(c) How much heat is required to heat this kettle from 23.0 °C to 99.0 °C if it contains 1.25 L of water (density of 0.997 g/mL and a specific heat of 4.184 J/g °C)?

Solution

(a) Heat capacity can be used in several ways. The molar heat capacity is the amount of heat necessary to raise 1 molar mass of the substance 1 °C. If the term *heat capacity* is used without qualifiers, it means the heat necessary to raise the temperature of the body 1 °C:

Heat capacity = mc = 1050 g × 0.897 J/g °C = 942 J/°C = 9.42 × 10² J/°C; (b) use $q = mc\Delta T$, where mc is known from part (a):

$q = 942$ J/°C($T_f – T_i$)

$= 942$ J/°C(99.0 °C – 23.0 °C) = 7.2 × 10⁴ J;

(c) Two items of different heat capacity must be heated, so the heat required for each must be added. Use the density to determine the mass of the water being heated:

1.25 L × 1000 mL/L × 0.997 g/mL = 1246 g

$q_{kettle} = mc(T_f – T_i)$

$= 1050$ g × 0.897 J/g °C(99.0 °C – 23.0 °C) = 71,600 J

$q_{water} = mc(T_f – T_i)$

$= 1246$ g × 4.184 J/g °C(99.0 °C – 23.0 °C) = 396.2 J

$q_{kettle} + q_{water} = 7.16 × 10⁴$ J + 3.962 × 10⁵ J = 4.68 × 10⁵ J.

Most people find waterbeds uncomfortable unless the water temperature is maintained at about 85°F. Unless it is heated, a waterbed that contains 892 L of water cools from 85 °F to 72 °F in 24 hours. Estimate the amount of electrical energy required over 24 hours, in kWh, to keep the bed from cooling. Note that 1 kilowatt-hour (kWh) = 3.6×10^6 J, and assume that the density of water is 1.0 g/mL (independent of temperature). What other assumptions did you make? How did they affect your calculated result (i.e., were they likely to yield 'positive' or 'negative' errors)?

Solution

We assume that the density of water is 1.0 g/cm³(1 g/mL) and that it takes as much energy to keep the water at 85 °F as to heat it from 72 °F to 85 °F. We also assume that only the water is going to be heated. First, find the mass of water in the bed. The volume is 72 in. × 84 in. × 9 in.

$$= 54{,}432 \text{ in.}^3 \left(\frac{2.54 \text{ cm}}{1 \text{ in.}} \right)^3 = 8.9 \times 10^5 \text{ cm}^3 \longrightarrow 8.9 \times 10^5 \text{ g} ;$$

second, convert the change of 13 °F to °C:

$q = cm\Delta T = 4.184 \text{ J/g °C}(8.92 \times 10^5 \text{ g}) \times 7.22 \text{ °C}$

$°C = \dfrac{5}{9} °F = \dfrac{5}{9} \times 13 = 7.22 °C$

$= 2.69 \times 10^7 \text{ J}$

$$\text{Energy required} = \frac{2.689 \times 10^7 \text{ J}}{3.6 \times 10^6 \text{ kWh}^{-1}} = 7.47 \text{ kWh}$$

5.2 Calorimetry

A 500-mL bottle of water at room temperature and a 2-L bottle of water at the same temperature were placed in a refrigerator. After 30 minutes, the 500-mL bottle of water had cooled to the temperature of the refrigerator. An hour later, the 2-L of water had cooled to the same temperature. When asked which sample of water lost the most heat, one student replied that both bottles lost the same amount of heat because they started at the same temperature and finished at the same temperature. A second student thought that the 2-L bottle of water lost more heat because there was more water. A third student believed that the 500-mL bottle of water lost more heat because it cooled more quickly. A fourth student thought that it was not possible to tell because we do not know the initial temperature and the final temperature of the water. Indicate which of these answers is correct and describe the error in each of the other answers.

Solution

The 2-L bottle of water lost more heat than the 500-mL bottle; the other statements are incorrect. The first student did not realize that the heat lost depends on the mass of the two objects, as it is an extensive property. The intensive property would be heat lost per unit mass of water, and it should be the same for both objects. The third student did not understand that, regardless of the rate of heat exchange, only the mass of water and the temperature change matter. The fourth student did not realize that once the initial and final temperatures are set for a specific substance, only the mass of the two bodies determines how much heat is transferred. Also, the fourth student is wrong because the problem states that the temperatures would be the same.

Question 29-2.

Would the amount of heat measured for the reaction in Example 5.5 be greater, lesser, or remain the same if we used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

Solution

lesser; more heat would be lost to the coffee cup and the environment and so ΔT for the water would be lesser and the calculated q would be lesser

Question 29-3.

Would the amount of heat absorbed by the dissolution in Example 5.6 appear greater, lesser, or remain the same if the experimenter used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

Solution

lesser; more heat would enter the reaction solution from the environment and ΔT would be lesser

Question 29-4.

Would the amount of heat absorbed by the dissolution in Example 5.6 appear greater, lesser, or remain the same if the heat capacity of the calorimeter were taken into account? Explain your answer.

Solution

greater, since taking the calorimeter's heat capacity into account will compensate for the thermal energy transferred to the solution from the calorimeter; this approach includes the calorimeter itself, along with the solution, as "surroundings:" $q_{rxn} = -(q_{solution} + q_{calorimeter})$; since both $q_{solution}$ and $q_{calorimeter}$ are negative, including the latter term (q_{rxn}) will yield a greater value for the heat of the dissolution

Question 29-5.

How many milliliters of water at 23 °C with a density of 1.00 g/mL must be mixed with 180 mL (about 6 oz) of coffee at 95 °C so that the resulting combination will have a temperature of 60 °C? Assume that coffee and water have the same density and the same specific heat.

Solution

From the law of conservation of energy:

$q_{water} + q_{coffee} = 0$

$q_{water} = -q_{coffee}$

$cm_{water}\Delta T = -cm_{coffee}\Delta T.$

Since the specific heats are the same, they cancel on both sides of the equation. Also, 1.00 g = 1.00 mL:

$m_{water} (60 - 23) = -180(60 - 95)$

$$m_{water} = \frac{180(35)}{37} = 170 \text{ g} \times \frac{1.00 \text{ mL}}{1.00 \text{ g}} = 170 \text{ mL} = 1.7 \times 10^2 \text{ mL}.$$

Question 29-6.

How much will the temperature of a cup (180 g) of coffee at 95 °C be reduced when a 45-g silver spoon (specific heat 0.24 J/g °C) at 25 °C is placed in the coffee and the two are allowed to reach the same temperature? Assume that the coffee has the same density and specific heat as water.

Solution

Because of the law of conservation of energy, we write:

$q_{spoon} + q_{coffee} = 0$; $q_{spoon} = -q_{coffee}$

$c_{spoon} \times m_{spoon} \times \Delta T = -c_{coffee} \times m_{coffee} \times \Delta T$

0.24 J/g °C × 45 g × $(T_f - 25$ °C$)$ = 4.184 J/g °C × 180 g × $(T_f - 95$ °C$)$

$10.8 T_f - 270 = -753.1 T_f + 71546.4$

$763.9 T_f = 71816.4$

$T_f = 94$ °C.

The temperature of the coffee will drop 1 degree.

Question 29-7.

A 45-g aluminum spoon (specific heat 0.88 J/g °C) at 24 °C is placed in 180 mL (180 g) of coffee at 85 °C and the temperature of the two become equal.

    (a) What is the final temperature when the two become equal? Assume that coffee has the same specific heat as water.

    (b) The first time a student solved this problem she got an answer of 88 °C. Explain why this is clearly an incorrect answer.

Solution

(a) because of the law of conservation of energy, we write:

$q_{spoon} + q_{coffee} = 0$

$q_{spoon} = -q_{coffee}$

$c_{spoon} \times m_{spoon} \times \Delta T = c_{coffee} \times m_{coffee} \times \Delta T$

0.88 J/g °C × 45 g × $(T_f - 24$ °C$)$

$= -4.184$ J/g °C × 180 g × $(T_f - 85$ °C$)$

$39.6 T_f - 950.4 = -753.12 T_f + 64,015$

$753.12 T_f + 39.6 T_f = 64,015 + 950.4$

$792.7 T_f = 64,965.4$

$T_f = 82$ °C; (b) The temperature of the coffee cannot increase when a colder spoon is placed in it.

Question 29-8.

The temperature of the cooling water as it leaves the hot engine of an automobile is 240 °F. After it passes through the radiator it has a temperature of 175 °F. Calculate the amount of heat transferred from the engine to the surroundings by one gallon of water with a specific heat of 4.184 J/g °C.

Solution

First, find the change in temperature of the water in °C:

240 °F − 175 °F = 65 °F.

We are concerned here only with the difference between temperatures and not the conversion from one temperature to the corresponding temperature.

$$°C = \frac{5}{9}(°F) = \frac{5}{9}(65) = 36.1 \text{ °C}$$

Assuming 1 mL of water has a mass of 1 g, 1.0 gal of water has a mass of:

$$1.0 \text{ gal} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{0.9463 \text{ L}}{1 \text{ qt}} \times 1000 \text{ g} = 3785 \text{ g}$$

$q = cm\Delta T = 4.184 \text{ J/g °C} \times 3758 \text{ g} \times 36.1 \text{ °C}$
$= 5.7 \times 10^2 \text{ kJ}$.

Question 29-9.

A 70.0-g piece of metal at 80.0 °C is placed in 100 g of water at 22.0 °C contained in a calorimeter like that shown in Figure 5.12. The metal and water come to the same temperature at 24.6 °C. How much heat did the metal give up to the water? What is the specific heat of the metal?

Solution

$q_{water} = 4.184 \text{ J/g °C} \times 100 \text{ g} \times (24.6 - 22.0) \text{ °C} = 1.1 \text{ kJ}$

$$c_{metal} = \frac{1.1 \text{ kJ} \times 1000 \text{ J kJ}^{-1}}{70.0 \text{ g} (80.0 - 24.6) \text{ °C}}$$

$= 0.28 \text{ J/g °C}$

Question 29-10.

If a reaction produces 1.506 kJ of heat, which is trapped in 30.0 g of water initially at 26.5 °C in a calorimeter like that in Figure 5.12, what is the resulting temperature of the water?

Solution

$q = cm\Delta T = 4.184 \text{ J/g °C} \times 30.0 \text{ g} \times (T_f - 26.5 \text{ °C}) = 1506 \text{ J}$
$T_f - 26.5 \text{ °C} = 1506 \text{ J} / (4.184 \text{ J/g° C} \times 30.0 \text{ g}) = 12.0 \text{ °C}$
$T_f = 26.5 \text{ °C} + 12.0 \text{ °C} = 38.5 \text{ °C}$

Question 29-11.

A 0.500-g sample of KCl is added to 50.0 g of water in a calorimeter (Figure 5.12). If the temperature decreases by 1.05 °C, what is the approximate amount of heat involved in the dissolution of the KCl, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

Solution

Assume that both the KCl and the water are at the same temperature. The entire temperature change can then be attributed to the heat of dissolution:

$q = cm\Delta T = 4.18 \text{ J/g °C} \times (50.0 \text{ g} + 0.500 \text{ g}) \times (-1.05 \text{ °C})$
$= -222 \text{ J}$.

The reaction is endothermic, which means that 222 J has been absorbed from the water and the KCl in order for the KCl to dissolve.

Question 29-12.

Dissolving 3.0 g of $CaCl_2(s)$ in 150.0 g of water in a calorimeter (Figure 5.12) at 22.4 °C causes the temperature to rise to 25.8 °C. What is the approximate amount of heat involved in the dissolution, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

Solution

Assume that the mass of the added $CaCl_2$ in solution must be added to the mass of the water:

$q_{reaction} + q_{solution} = 0$

$q_{reaction} = q_{solution}$

$q_{solution} = cm\Delta T$

$= 4.18 \text{ J/g °C} \times 153.0 \text{ g} \times (25.8 - 22.4) \text{ °C} = 2200 \text{ J}$

$= -2.2 \text{ kJ}.$

The heat produced shows that the reaction is exothermic.

Question 29-13.

When 50.0 g of 0.200 M NaCl(aq) at 24.1 °C is added to 100.0 g of 0.100 M AgNO₃(aq) at 24.1 °C in a calorimeter, the temperature increases to 25.2 °C as AgCl(s) forms. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat in joules produced.
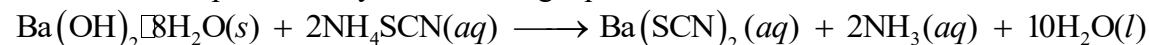
Solution

$q = cm\Delta T$

$= 4.20 \text{ J/g °C} \times 150 \text{ g} \times (25.2 - 24.1) \text{ °C} = 693 \text{ J}$

Question 29-14.

The addition of 3.15 g ofBa(OH)₂·8H₂O to a solution of 1.52 g of NH₄SCN in 100 g of water in a calorimeter caused the temperature to fall by 3.1 °C. Assuming the specific heat of the solution and products is 4.20 J/g°C, calculate the approximate amount of heat absorbed by the reaction, which can be represented by the following equation:

$Ba(OH)_2\square8H_2O(s) + 2NH_4SCN(aq) \longrightarrow Ba(SCN)_2(aq) + 2NH_3(aq) + 10H_2O(l)$

Solution

$q = cm\Delta T$

$= 4.20 \text{ J/g °C} \times (3.15 + 1.52 + 100) \text{ g} \times 3.1 \text{ °C}$

$= 1362.8 \text{ J} = 1.4 \text{ kJ (two significant figures)}$

Question 29-15.

The reaction of 50 mL of acid and 50 mL of base described in Example 5.5 increased the temperature of the solution by 6.9 degrees. How much would the temperature have increased if 100 mL of acid and 100 mL of base had been used in the same calorimeter starting at the same temperature of 22.0 °C? Explain your answer.

Solution

The answer is still 6.9 °C. Twice as much heat would be produced, but it would be absorbed by twice as much solution.

Question 29-16.

If the 3.21 g of NH₄NO₃ in Example 5.6 were dissolved in 100.0 g of water under the same conditions, how much would the temperature change? Explain your answer.

Solution

The heat of the reaction would be approximately the same as before or 1.0×10^3 J. The reaction is endothermic, and the temperature would decrease:

$q_{rxn} = -q_{soln} = -(c \times m \times \Delta T)_{soln}$

$-(4.18 \text{ J/g}° \text{ C}) \times (103.2 \text{ g}) \times (T_f - 24.9 \text{ °C}) = 1000 \text{ J}$

$T_f - 24.9 \text{ °C} = -2.3 \text{ °C}$

$T_f = 22.6$. Since the mass and the heat capacity of the solution is approximately equal to that of the water, the two-fold increase in the amount of water leads to a two-fold decrease of the temperature change.

Question 29-17.

When 1.0 g of fructose, $C_6H_{12}O_6(s)$, a sugar commonly found in fruits, is burned in oxygen in a bomb calorimeter, the temperature of the calorimeter increases by 1.58 °C. If the heat capacity of the calorimeter and its contents is 9.90 kJ/°C, what is q for this combustion?

Solution

$q_{comb} + q_{cal} = 0$

$q_{comb} = -q_{cal}$

$= -9.90 \text{ kJ/°C} (1.58 \text{ °C}) = -15.6 \text{ kJ}$

The negative sign indicates that this was reaction exothermic, as expected. The mass did not directly enter into the calculation because the heat capacity of the system is calculated on a kJ/°C basis. The mass of the calorimeter is constant.

Question 29-18.

When a 0.740-g sample of trinitrotoluene (TNT), $C_7H_5N_2O_6$, is burned in a bomb calorimeter, the temperature increases from 23.4 °C to 26.9 °C. The heat capacity of the calorimeter is 534 J/°C, and it contains 675 mL of water. How much heat was produced by the combustion of the TNT sample?

Solution

The heat absorbed by the calorimeter is $q_1 = 534 \text{ J/°C} \times (26.9 \text{ °C} - 23.4 \text{ °C}) = 1869 \text{ J}$. The heat absorbed by water is $q_1 = 675 \text{ mL} \times 0.997 \text{ g/mL} \times 4.184 \text{ J/g °C} \times (26.9 \text{ °C} - 23.4 \text{ °C}) = 9855 \text{ J}$. The overall amount of heat $q = q_1 + q_1 = 11{,}724 \text{ J}$ or 11.7 kJ with three significant digits.

Question 29-19.

One method of generating electricity is by burning coal to heat water, which produces steam that drives an electric generator. To determine the rate at which coal is to be fed into the burner in this type of plant, the heat of combustion per ton of coal must be determined using a bomb calorimeter. When 1.00 g of coal is burned in a bomb calorimeter (Figure 5.17), the temperature increases by 1.48 °C. If the heat capacity of the calorimeter is 21.6 kJ/°C, determine the heat produced by combustion of a ton of coal (2.000×10^3 pounds).

Solution

$q = mc\Delta T = 21.6 \text{ kJ/°C} \times 1.48 \text{ °C} = 32.0 \text{ kJ}$

This is the heat released by burning 1 g of coal.

Find the mass in kilograms of one ton:

$2.000 \times 10^3 \text{ lb} \times 453.6 \text{ g/lb} = 907.2 \text{ kg}$

Total heat = heat per g × mass of coal

$= 32.0 \text{ kJ/g} \times 1000 \text{ J/kJ} \times 907{,}200 \text{ g} = 2.90 \times 10^{10} \text{ J}$.

Question 29-20.

The amount of fat recommended for someone with a daily diet of 2000 Calories is 65 g. What percent of the calories in this diet would be supplied by this amount of fat if the average number of Calories for fat is 9.1 Calories/g?

Solution

Energy released by burning fat is:

65 g × 9.1 Calories/g = 5.9×10^2 Calories

% Calories from fat = 5.9×10^2 Calories/2000 Calories × 100%

= 30% (one significant figure).

Question 29-21.

A teaspoon of the carbohydrate sucrose (common sugar) contains 16 Calories (16 kcal). What is the mass of one teaspoon of sucrose if the average number of Calories for carbohydrates is 4.1 Calories/g?

Solution

$$\frac{16 \text{ kcal teaspoon}^{-1}}{4.1 \text{ kcal g}^{-1}} = 3.9 \text{ g teaspoon}^{-1}$$

Question 29-22.

What is the maximum mass of carbohydrate in a 6-oz serving of diet soda that contains less than 1 Calorie per can if the average number of Calories for carbohydrates is 4.1 Calories/g?

Solution

The average energy from carbohydrates is 4.1 kcal/g (4.1 Calories/g). The drink cannot contain

more than $\dfrac{1 \text{ Calorie}}{4.1 \text{ Calorie g}^{-1}} = 0.24 \text{ g}$.

Question 29-23.

A pint of premium ice cream can contain 1100 Calories. What mass of fat, in grams and pounds, must be produced in the body to store an extra 1.1×10^3 Calories if the average number of Calories for fat is 9.1 Calories/g?

Solution

The average heat of combustion of fats is –9.1 kcal/g = –9.1 Calories/g:

$$\text{Mass} = 1100 \text{ Calories} \times \frac{1.0 \text{ g}}{9.1 \text{ Calories}} = 121 \text{ g} = 1.2 \times 10^2$$

$$\text{Mass} = 121 \text{ g} \times \frac{10 \text{ lb}}{453.6 \text{ g}} = 0.27 \text{ lb}.$$

Question 29-24.

A serving of a breakfast cereal contains 3 g of protein, 18 g of carbohydrates, and 6 g of fat. What is the Calorie content of a serving of this cereal if the average number of Calories for fat is 9.1 Calories/g, for carbohydrates is 4.1 Calories/g, and for protein is 4.1 Calories/g?

Solution

Protein: 3 g × 4.1 Calories/g   12.3 Calories

Carbohydrates: 18 g × 4.1 Calories/g   73.8 Calories

Fat: 6 g × 9.1 Calories/g   <u>54.6 Calories</u>

                                         140.7 Calories

Total $= 1.4 \times 10^2$ Calories

Question 29-25.

Which is the least expensive source of energy in kilojoules per dollar: a box of breakfast cereal that weighs 32 ounces and costs \$4.23, or a liter of isooctane (density, 0.6919 g/mL) that costs \$0.45? Compare the nutritional value of the cereal with the heat produced by combustion of the isooctane under standard conditions. A 1.0-ounce serving of the cereal provides 130 Calories.

Solution

Convert ounces to grams and liters to grams, and then find the cost per gram. Determine the heat of combustion of isooctane on a per gram basis, and compare it with the cost of the heat produced by cereal.

For isooctane:

1 L = 1000 mL

Mass = 1000 mL × 0.6919 g/mL = 691.9 g

$$\frac{\$0.45}{691.9 \text{ g}} = \$0.000650 \text{ g}^{-1}$$

On a heat-released basis, we must calculate the heat of combustion of isooctane. This value is:

5460 kJ/mol for C_8H_{18} (M.W. = 114.23 g/mol). Division gives $\dfrac{5460 \text{ kJ mol}^{-1}}{114.23 \text{ g mol}^{-1}} = 47.80 \text{ kJ g}^{-1}$ and

from the previous calculation:

$$\frac{47.80 \text{ kJ g}^{-1}}{\$0.000650 \text{ g}^{-1}} = 7.35 \times 10^4 \text{ kJ/\$}.$$

For cereal:

32 oz × 28.35 g/oz = 907.2 g

$$\frac{\$4.23}{907.2 \text{ g}} = \$0.00466 \text{ g}^{-1}$$

For 1.0 oz or 28.35 g, 130 Calories or 130,000 cal are produced. For 1.0 g of cereal:

$$1.30 \times 10^3 \text{ calories} \times \frac{4.184 \text{ J cal}^{-1}}{28.35 \text{ g}} = 19.2 \text{ kJ g}^{-1}.$$
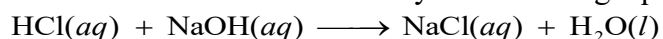
The cost of cereal heat is then:

$$\frac{19.2 \text{ kJ g}^{-1}}{\$0.00466 \text{ g}^{-1}} = 4.12 \times 10^3 \text{ kJ/\$}.$$

Isooctane is the least expensive source.

5.3 Enthalpy

Explain how the heat measured in Example 5.5 differs from the enthalpy change for the exothermic reaction described by the following equation:

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

Solution

The enthalpy change of the indicated reaction is for exactly 1 mol HCL and 1 mol NaOH; the heat in the example is produced by 0.0500 mol HCl and 0.0500 mol NaOH.

Question 30-2.

Using the data in the check your learning section of Example 5.5, calculate ΔH in kJ/mol of AgNO3(aq) for the reaction: $NaCl(aq) + AgNO_3(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$.

Solution

$q_{soln} = cm\Delta T = 4.184$ J/g °C× 200 g× $(23.5 - 21.9)$ °C $= 1.3 \times 10^3$ J

$q_{soln} = -q_{rxn}$

Therefore, $q_{rxn} = -1.3 \times 10^3$ J. At constant pressure, $\Delta H = q_{rxn}$ so $\Delta H = -1.3 \times 10^3$ J or -1.3 kJ

$$0.100 \; \cancel{L} \times \frac{0.200 \text{ mol}}{1 \; \cancel{L}} = 0.02 \text{ mol AgNO}_3$$

ΔH in kJ/mol AgNO3 would be $\dfrac{-1.3 \text{ kJ}}{0.0200 \text{ mol AgNO}_3} = -65$ kJ/mol $AgNO_3$

Question 30-3.

Calculate the enthalpy of solution (ΔH for the dissolution) per mole of NH4NO3 under the conditions described in Example 5.6.

Solution

The molar mass of NH_4NO_3 is 80.0423 g/mol. From the example, 1000 J is required to dissolve 3.21 g of NH_4NO_3. One mole under the same conditions would require

$$\frac{80.0432 \; \cancel{g} \text{ mol}^{-1}}{3.21 \; \cancel{g}} \times 1000 \text{ J} = 25 \text{ kJ mol}^{-1}.$$

(The heat of solution is positive because the process is endothermic.)

Question 30-4.

Calculate ΔH for the reaction described by the equation. (Hint: Use the value for the approximate amount of heat absorbed by the reaction that you calculated in a previous exercise.)

$$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4SCN(aq) \longrightarrow Ba(SCN)_2(aq) + 2NH_3(aq) + 10H_2O(l)$$
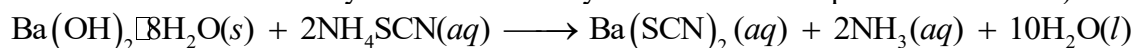
Solution

$Ba(OH)_2 \cdot 8H_2O$: Ba 1 × 137.327 = 137.327

O 10 × 15.9994 = 159.994

H 18 × 1.00794 = <u>18.143</u>

315.464 g mol^{-1}

$2NH_4SCN:$ N $4 \times 14.00674 = 56.02696$

S $2 \times 32.066 = 64.132$

C $2 \times 12.011 = 24.022$

H $8 \times 1.0079 = \underline{8.0632}$

152.244 or 76.122 g mol^{-1}

3.15 g of Ba(OH)$_2 \cdot$8H$_2$O is 0.01 mol and 1.52 g of NH$_4$SCN is 0.02 mol. Assume that a solution of these substances is 100 g and still gives 100 mL of solution, with their concentrations being 0.1 M and 0.2 M, respectively. This relation is the ratio needed to make 0.01 M Ba(SCN)$_2$. Since 1.4 kJ is released for these quantities, 1 mol of Ba(OH)$_2 \cdot$8H$_2$O should release 100×1.4 kJ $= 140$ kJ mol^{-1}.

Question 30-5.

Calculate the enthalpy of solution (ΔH for the dissolution) per mole of CaCl$_2$ (refer to Exercise 25).

Solution

The molar mass of CaCl$_2$ is $40.078 + 2(35.4572) = 110.992$ g/mol. In Exercise 25, 3.0 g of CaCl$_2$ dissolved in water releases a heat of 2.2 kJ. Therefore,

$$\frac{3.0 \text{ g}}{110.992 \text{ g mol}^{-1}} = 0.027 \text{ mol}$$

$$\frac{2.2 \text{ kJ}}{0.027 \text{ mol}} = 81 \text{ kJ mol}^{-1}.$$

Question 30-6.

Although the gas used in an oxyacetylene torch (Figure 5.7) is essentially pure acetylene, the heat produced by combustion of one mole of acetylene in such a torch is likely not equal to the enthalpy of combustion of acetylene listed in Table 5.2. Considering the conditions for which the tabulated data are reported, suggest an explanation.

Solution

The tabulated enthalpy changes are reported for standard state conditions (1 bar pressure, and a specified temperature of 25 °C). The combustion of acetylene in the torch is clearly taking place at a much higher temperature. Also, since the acetylene and oxygen are usually supplied from compressed gas cylinders, the pressures of these gases in the torch combustion zone is very likely greater than 1 bar. Since the conditions for the combustion are far from standard state conditions, it stands to reason the molar enthalpy change will be different from the standard value.

Question 30-7.

How much heat is produced by burning 4.00 moles of acetylene under standard state conditions?

Solution

The heat of combustion is –1301.1 as given in Table 5.2.

Heat released = 4.00 mol × (–1301.1 kJ/mol) = 5204.4 kJ

Question 30-8.

How much heat is produced by combustion of 125 g of methanol under standard state conditions?

Solution

The molar mass of methanol = 32.04 g/mol.

$$\frac{125 \text{ g}}{32.01 \text{ g/mol}} = 3.90 \text{ mol}$$

The heat of combustion is –726.1 as given in Table 5.2. $\Delta H = -726.1 \times 3.90 = 2833$ kJ or 2.83×10^3 kJ.

Question 30-9.

How many moles of isooctane must be burned to produce 100 kJ of heat under standard state conditions?

Solution

The value of $\Delta H_{comb} = -5461$ kJ/mol. To produce 100 kJ requires:

$$\frac{100 \text{ kJ}}{5461 \text{ kJ mol}^{-1}} = 1.83 \times 10^{-2} \text{ mol}.$$

Question 30-10.

What mass of carbon monoxide must be burned to produce 175 kJ of heat under standard state conditions?

Solution

Find the heat generated from burning 1 mol of carbon monoxide. The reaction is:

$$CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO_2(g)$$
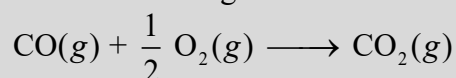
$$\Delta H_f^{\circ}/\text{kJ mol}^{-1} = \quad -110.52 \quad 0 \quad\quad\quad -393.51$$

$$\Delta H_{comb} = -393.51 - (-110.52) = -282.99 \text{ kJ mol}^{-1}$$

$$\text{Moles of CO required} = \frac{175 \text{ kJ}}{282.99 \text{ kJ mol}^{-1}} = 0.618 \text{ mol}$$

The molar mass of CO is 28.01 g/mol, the total mass of CO required is 0.618 mol × 28.01 g/mol = 17.3 g.

Question 30-11.

When 2.50 g of methane burns in oxygen, 125 kJ of heat is produced. What is the enthalpy of combustion per mole of methane under these conditions?

Solution
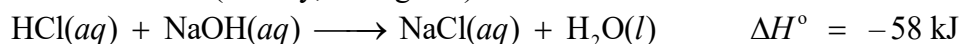
The molar mass of CH$_4$ is 16.04 g/mol.

Find the mole of CH$_4$ present: $\dfrac{2.50 \text{ g}}{16.04 \text{ g mol}^{-1}} = 0.15586 \text{ mol}$

$$\Delta H_{comb} = -\frac{125 \text{ kJ}}{0.15586 \text{ mol}} = -802 \text{ kJ mol}^{-1}.$$

Question 30-12.

How much heat is produced when 100 mL of 0.250 M HCl (density, 1.00 g/mL) and 200 mL of 0.150 M NaOH (density, 1.00 g/mL) are mixed?

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l) \qquad \Delta H° = -58 \text{ kJ}$$

If both solutions are at the same temperature and the specific heat of the products is 4.19 J/g °C, how much will the temperature increase? What assumption did you make in your calculation?

Solution
for HCl: 0.100 L × 0.250 M = 0.0250 mol;
for NaOH, 0.200 L × 0.150 M = 0.0300 mol;
Only a total of 0.0250 mol of NaCl can be produced. To find the heat produced, multiply the moles of product times the heat produced per mole:
0.0250 mol × 58 kJ/mol = 1.5 kJ.
This solution assumes that $\Delta H°$ is not different from ΔH under the conditions of the reaction (1 °C temperature increase and assuming volumes are additive).
Assuming the overall mass of the final system is approximately 300 g, the increase of temperature will be $\dfrac{1500 \text{ J}}{(300 \text{ g} \times 4.19 \text{ J g}^{-1} \text{ °C}^{-1})} = 1.2 \text{ °C}$.

Question 30-13.

A sample of 0.562 g of carbon is burned in oxygen in a bomb calorimeter, producing carbon dioxide. Assume both the reactants and products are under standard state conditions, and that the heat released is directly proportional to the enthalpy of combustion of graphite. The temperature of the calorimeter increases from 26.74 °C to 27.93 °C. What is the heat capacity of the calorimeter and its contents?

Solution
The heat produced is found from the enthalpy of combustion:
$$\frac{0.562 \text{ g °C}}{12.011 \text{ g mol}^{-1} \text{ °C}} \times -393.5 \text{ kJ mol}^{-1} = -18.4 \text{ kJ}.$$
This is the heat released.
q = heat capacity × ΔT
18.4 kJ = heat capacity × (27.93 °C – 26.74 °C) = specific heat × 1.19 °C
specific heat = 15.5 kJ/°C

Question 30-14.

Before the introduction of chlorofluorocarbons, sulfur dioxide (enthalpy of vaporization, 6.00 kcal/mol) was used in household refrigerators. What mass of SO_2 must be evaporated to remove as much heat as evaporation of 1.00 kg of CCl_2F_2 (enthalpy of vaporization is 17.4 kJ/mol)?

The vaporization reactions for SO_2 and CCl_2F_2 are $SO_2(l) \longrightarrow SO_2(g)$ and $CCl_2F_2(l) \longrightarrow CCl_2F_2(g)$, respectively.
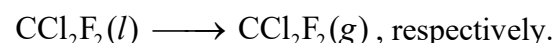
Solution
Convert enthalpy of vaporization in kilocalories to kilojoules:
6.00 kcal/mol × 4.18 kJ/kcal = 25.1 kJ/mol.
Convert 1.00 kg CCl_2F_2 into moles of CCl_2F_2. Let the change in temperature be 1.0 °C.

Molar mass: For CCl_2F_2, $12.011 + 2(35.4527) + 2(18.9984) = 120.913$ g/mol;
for SO_2, $32.066 + 2(15.9994) = 64.065$ g/mol

$$\frac{1.00 \times 10^3 \text{ g}}{120.913 \text{ g mol}^{-1}} = 8.27 \text{ mol}$$

g SO_2 = mol SO_2 × 25.1 kJ/mol = mol CCl_2F_2 × 17.4 kJ/mol

$$\text{mol SO}_2 = \frac{8.27 \text{ mol} \times 17.4}{25.1} = 5.73 \text{ mol}$$

g SO_2 = 5.73 mol × 64.065 g/mol = 367 g

Question 30-15.

Homes may be heated by pumping hot water through radiators. What mass of water will provide the same amount of heat when cooled from 95.0 to 35.0 °C, as the heat provided when 100 g of steam is cooled from 110 °C to 100 °C.

Solution

$q = mc\Delta T$. The amount of heat transferred from steam is:

q_{steam} = 100 g × 1.864 J/(g °C) × (110 − 100) °C = 1864 J

$q_{water} = m_{water}$× 4.184 J/(g °C) × (95 − 35) °C = m_{water}× 251.04 J/g = q_{steam} = 1864 J

$$m_{water} = \frac{1864 \text{ J}}{251.04 \text{ J/g}} = 7.43 \text{ g}.$$
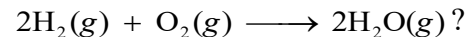
Question 30-16.

Which of the enthalpies of combustion in the table are also standard enthalpies of formation?

Solution

Substance	Combustion/Formation Reaction
carbon	$C(s) + \dfrac{1}{2}O_2(s) \longrightarrow CO(g)$ $C(s) + O_2(g) \longrightarrow CO_2(g)$
hydrogen	$H_2(g) + \dfrac{1}{2}O_2(g) \longrightarrow H_2O(l)$
magnesium	$Mg(s) + \dfrac{1}{2}O_2(g) \longrightarrow MgO(s)$
sulfur	$S(s) + O_2(g) \longrightarrow SO_2(g)$

Question 30-17.

Does the standard enthalpy of formation of $H_2O(g)$ differ from $\Delta H°$ for the reaction
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)?$$

Solution
Yes. The heat of this reaction is defined for 2 moles of $H_2O(g)$, thus the heat of formation is half of the heat of the reaction.

Question 30-18.

Joseph Priestly prepared oxygen in 1774 by heating red mercury(II) oxide with sunlight focused through a lens. How much heat is required to decompose exactly 1 mole of red $HgO(s)$ to $Hg(l)$ and $O_2(g)$ under standard conditions?

Solution

The reaction $HgO(s) \longrightarrow Hg(l) + \frac{1}{2}O_2(g)$ under standard state conditions is the reverse of the reaction that forms 1 mol $HgO(s)$ from the elements in their most stable states under standard state conditions. Thus, $\Delta H° = -\Delta H°_{f_{HgO(s)}} = -(-90.83 \text{ kJ}) = 90.83 \text{ kJ}$.
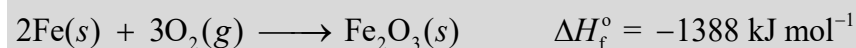
Question 30-19.

How many kilojoules of heat will be released when exactly 1 mole of manganese, Mn, is burned to form $Mn_3O_4(s)$ at standard state conditions?

Solution

This process requires 3 mol of Mn. For 1 mol, $\frac{1}{3}(-1378.83 \text{ kJ}) = 459.6 \text{ kJ}$.

Question 30-20.

How many kilojoules of heat will be released when exactly 1 mole of iron, Fe, is burned to form $Fe_2O_3(s)$ at standard state conditions?
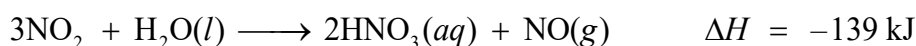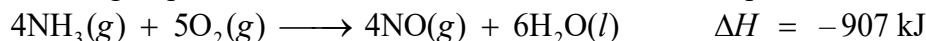
Solution

$2Fe(s) + 3O_2(g) \longrightarrow Fe_2O_3(s) \qquad \Delta H°_f = -1388 \text{ kJ mol}^{-1}$

$\Delta H°_f = -824.2 \text{ kJ/mol}$

This process requires 2 mol of Fe. For 1 mol, $\frac{1}{2}(-1388 \text{ kJ}) = 462.7 \text{ kJ}$.

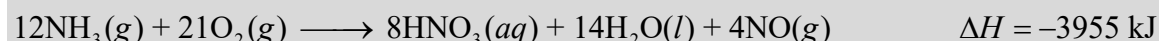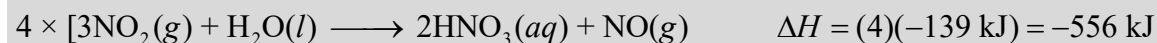$\frac{1}{2}(-8.24.2) = 412.1 \text{ kJ}$ of heat will be released.

Question 30-21.

The following sequence of reactions occurs in the commercial production of aqueous nitric acid:

$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l) \qquad \Delta H = -907 \text{ kJ}$

$2NO(g) + O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H = -113 \text{ kJ}$

$3NO_2 + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g) \qquad \Delta H = -139 \text{ kJ}$

Determine the total enthalpy change for the production of one mole of aqueous nitric acid by this process. Coproducts of the net reaction include water and nitrogen monoxide.

Solution

First, appropriately sum these three equations to yield a balanced equation for the overall reaction. Although fractional coefficients may be used in thermochemical calculations such as this, the approach shown here uses the smallest possible integer coefficients:
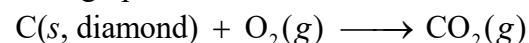
$3 \times [4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(l)]$ $\qquad$ $\Delta H = (3)(-907 \text{ kJ}) = -2721 \text{ kJ}$

$6 \times [2NO(g) + O_2(g) \longrightarrow 2NO_2(g)]$ $\qquad$ $\Delta H = (6)(-113 \text{ kJ}) = -678 \text{ kJ}$

$4 \times [3NO_2(g) + H_2O(l) \longrightarrow 2HNO_3(aq) + NO(g)$ $\qquad$ $\Delta H = (4)(-139 \text{ kJ}) = -556 \text{ kJ}$

$12NH_3(g) + 21O_2(g) \longrightarrow 8HNO_3(aq) + 14H_2O(l) + 4NO(g)$ $\qquad$ $\Delta H = -3955 \text{ kJ}$

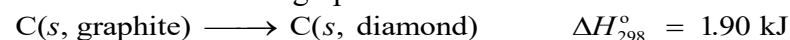The energy change shown is for production of 8 moles of aqueous nitric acid; therefore, production of 1 mole is
$(1/8)(-3995 \text{ kJ}) = -494 \text{ kJ}$.

Question 30-22.

Both graphite and diamond burn.

$C(s, \text{diamond}) + O_2(g) \longrightarrow CO_2(g)$

For the conversion of graphite to diamond:

$C(s, \text{graphite}) \longrightarrow C(s, \text{diamond})$ $\qquad$ $\Delta H^\circ_{298} = 1.90 \text{ kJ}$

Which produces more heat, the combustion of graphite or the combustion of diamond?

Solution

Heat is consumed by graphite as it is converted into diamond. Therefore, the enthalpy of diamond is higher than that of graphite. The heat of formation of C(s, graphite) is 0, the heat of formation of C(s, diamond) is 1.90 kJ/mol. The combustion reaction is:

$C(s) + O_2(g) \longrightarrow CO_2(g)$,
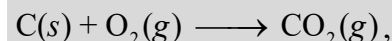
and the heat of combustion may be calculated using heats of formation as:

$\Delta H^\circ = \Delta H^\circ_f(CO2) - \Delta H^\circ_f(C(s))$. Since $\Delta H^\circ_f(C(s))$ is higher for diamond than for graphite, the heat of this reaction will be lower (more negative) for diamond, thus more heat will be released in burning of diamond than in burning of graphite.
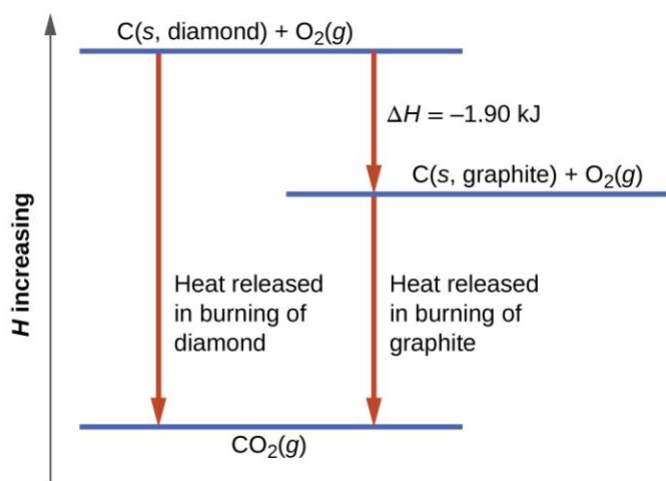
Question 30-23.

From the molar heats of formation in Appendix G, determine how much heat is required to evaporate one mole of water: $H_2O(l) \longrightarrow H_2O(g)$

Solution

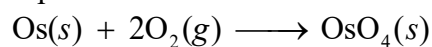The heat of formation of $H_2O(l)$ is –285.83 kJ/mol. The heat of formation of $H_2O(g)$ is –241.82 kJ/mol. For the reaction:

$H_2O(l) \longrightarrow H_2O(g)$,
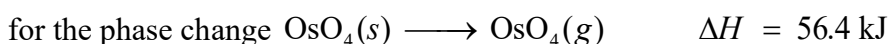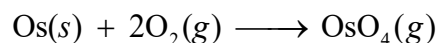
the enthalpy of the reaction is the difference of heats of formation of the reactant and the product; thus, it is $-241.82 - (-285.83) = 44.01$ kJ/mol.

Question 30-24.

Which produces more heat?

$$Os(s) + 2O_2(g) \longrightarrow OsO_4(s)$$

or

$$Os(s) + 2O_2(g) \longrightarrow OsO_4(g)$$

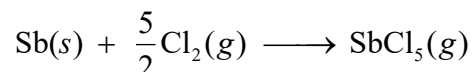for the phase change $OsO_4(s) \longrightarrow OsO_4(g)$ $\Delta H = 56.4$ kJ

Solution

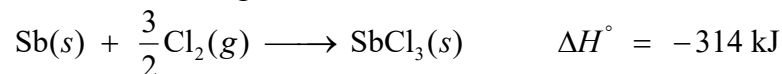Heat must be added to make the transition to the gas phase. Therefore, the formation of the solid produces more heat.
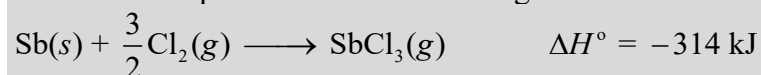
Question 30-25.

Calculate $\Delta H°$ for the process

$$Sb(s) + \frac{5}{2}Cl_2(g) \longrightarrow SbCl_5(g)$$

from the following information:

$$Sb(s) + \frac{3}{2}Cl_2(g) \longrightarrow SbCl_3(s) \qquad \Delta H° = -314 \text{ kJ}$$

$$SbCl_3(s) + Cl_2(g) \longrightarrow SbCl_5(s) \qquad \Delta H° = -80 \text{ kJ}$$

Solution

Add the two equations and their heat together.

$$Sb(s) + \frac{3}{2}Cl_2(g) \longrightarrow SbCl_3(g) \qquad \Delta H° = -314 \text{ kJ}$$

$$\underline{SbCl_3(s) + Cl_2(g) \longrightarrow SbCl_5(s) \qquad \Delta H° = -80 \text{ kJ}}$$

$$Sb(s) + \frac{5}{2}Cl_2(g) \longrightarrow SbCl_5(s) \qquad \Delta H° = -394 \text{ kJ}$$
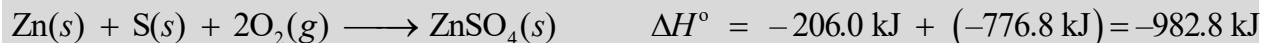
Question 30-26.

Calculate $\Delta H°_{298}$ for the process $Zn(s) + S(s) + 2O_2(g) \longrightarrow ZnSO_4(s)$ from the following information:

$$Zn(s) + S(s) \longrightarrow ZnS(s) \qquad \Delta H° = -206.0 \text{ kJ}$$

$$ZnS(s) + 2O_2(g) \longrightarrow ZnSO_4(s) \qquad \Delta H° = -776.8 \text{ kJ}$$
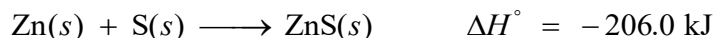
Solution
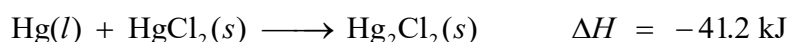
Addition of the two reaction equations as though they were mathematical equation gives:

$$Zn(s) + S(s) + 2O_2(g) \longrightarrow ZnSO_4(s) \qquad \Delta H° = -206.0 \text{ kJ} + (-776.8 \text{ kJ}) = -982.8 \text{ kJ}$$

.

Question 30-27.
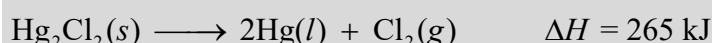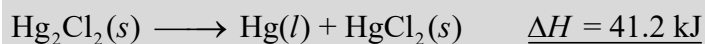
Calculate ΔH for the process $Hg_2Cl_2(s) \longrightarrow 2Hg(l) + Cl_2(g)$
from the following information:

$$Hg(l) + Cl_2(g) \longrightarrow HgCl_2(s) \qquad \Delta H = -224 \text{ kJ}$$

$$Hg(l) + HgCl_2(s) \longrightarrow Hg_2Cl_2(s) \qquad \Delta H = -41.2 \text{ kJ}$$
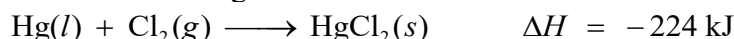
Solution

Reverse the direction of both equations and add the new equations and enthalpies.

$$HgCl_2(s) \longrightarrow Hg(l) + Cl_2(l) \qquad \Delta H = 224 \text{ kJ}$$

$$\underline{Hg_2Cl_2(s) \longrightarrow Hg(l) + HgCl_2(s) \qquad \Delta H = 41.2 \text{ kJ}}$$

$$Hg_2Cl_2(s) \longrightarrow 2Hg(l) + Cl_2(g) \qquad \Delta H = 265 \text{ kJ}$$

Question 30-28.
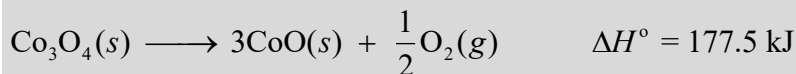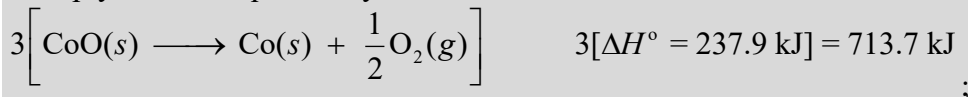
Calculate $\Delta H°_{298}$ for the process $Co_3O_4(s) \longrightarrow 3Co(s) + 2O_2(g)$
from the following information:

$$Co(s) + \frac{1}{2}O_2(g) \longrightarrow CoO(s) \qquad \Delta H° = -237.9 \text{ kJ}$$

$$3CoO(s) + \frac{1}{2}O_2(g) \longrightarrow Co_3O_4(s) \qquad \Delta H° = -177.5 \text{ kJ}$$
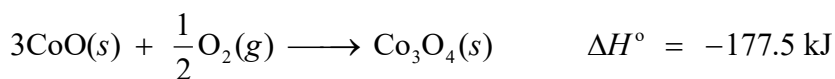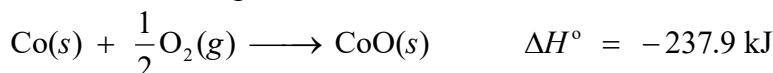
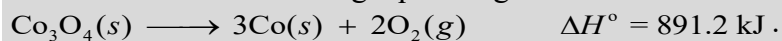**Solution**

Reverse the directions of the two equations so that Co₃O₄(s) will be on the left-hand side and multiply the first equation by 3:

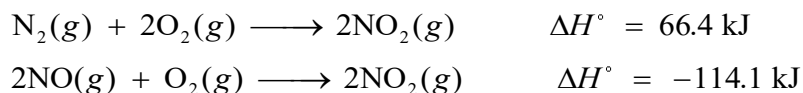$$3\left[CoO(s) \longrightarrow Co(s) + \frac{1}{2}O_2(g) \right] \qquad 3[\Delta H° = 237.9 \text{ kJ}] = 713.7 \text{ kJ}$$

;

$$Co_3O_4(s) \longrightarrow 3CoO(s) + \frac{1}{2}O_2(g) \qquad \Delta H° = 177.5 \text{ kJ}$$

addition of the two resulting equation gives:

$$Co_3O_4(s) \longrightarrow 3Co(s) + 2O_2(g) \qquad \Delta H° = 891.2 \text{ kJ} .$$

Question 30-29.

Calculate the standard molar enthalpy of formation of NO(g) from the following data:

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H° = 66.4 \text{ kJ}$$

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H° = -114.1 \text{ kJ}$$

Solution

Hess's law can be applied to the two equations by reversing the direction of the second equation. The first equation is a formation reaction and is so indicated by writing $\Delta H°$.

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H° = 66.4 \text{ kJ}$$

$$2NO_2(g) \longrightarrow 2NO(g) + O_2(g) \qquad \Delta H° = 114.1 \text{ kJ}$$
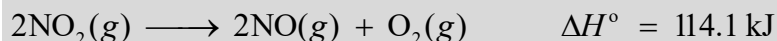
Adding the equations yields: $N_2(g) + O_2(g) \longrightarrow 2NO(g) \qquad \Delta H° = 180.5 \text{ kJ}$.

This is the heat of formation of 2 mol of NO. For 1 mol,

$$\Delta H° = \frac{180.5 \text{ kJ}}{2} = 90.3 \text{ kJ/mol of NO}.$$

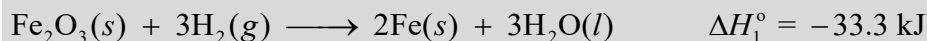## Question 30-30.

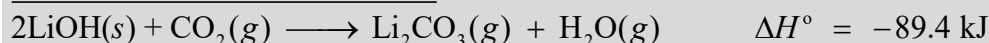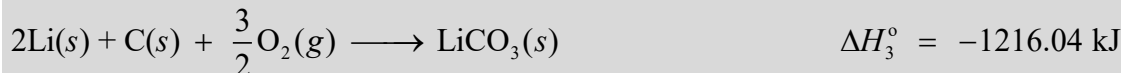Using the data in Appendix G, calculate the standard enthalpy change for each of the following reactions:

(a) $N_2(g) + O_2(g) \longrightarrow 2NO(g)$

(b) $Si(s) + 2Cl_2(g) \longrightarrow SiCl_4(g)$

(c) $Fe_2O(s) + 3H_2(g) \longrightarrow 2Fe(s) + 3H_2O(l)$

(d) $2LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(g)$

Solution

(a) $N_2(g) + O_2(g) \longrightarrow 2NO(g) \qquad \Delta H° = 2(90.25 \text{ kJ}) = 180.50 \text{ kJ}$, the reaction is twice the reaction of formation of NO(g). (b) This reaction is simply the reaction of formation of SiCl4(g); thus, its enthalpy is the enthalpy of formation of SiCl4(g), which is –662.75 (from Appendix G). (c)

$$3H_2(g) + \frac{3}{2}O_2(g) \longrightarrow 3H_2O(l) \qquad \Delta H_1° = 3(-285.83 \text{ kJ}) = -857.49 \text{ kJ}$$

$$\underline{Fe_2O_3(s) \longrightarrow 3Fe(s) + \frac{3}{2}O_2(g)} \qquad \underline{\Delta H_1° = -(-824.2 \text{ kJ})} \qquad ;$$

$$Fe_2O_3(s) + 3H_2(g) \longrightarrow 2Fe(s) + 3H_2O(l) \qquad \Delta H_1° = -33.3 \text{ kJ}$$

(d)

$$2LiOH(s) \longrightarrow 2Li(s) + O_2(g) \qquad \Delta H_1° = -2(-487.23 \text{ kJ}) = 975 \text{ kJ}$$

$$CO_2(g) \longrightarrow C(s) + O_2(g) \qquad \Delta H_2° = -(-393.51 \text{ kJ})$$

$$2Li(s) + C(s) + \frac{3}{2}O_2(g) \longrightarrow LiCO_3(s) \qquad \Delta H_3° = -1216.04 \text{ kJ}$$

$$\underline{H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)} \qquad \underline{\Delta H_4° = -241.82 \text{ kJ}}$$

$$2LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(g) + H_2O(g) \qquad \Delta H° = -89.4 \text{ kJ}$$
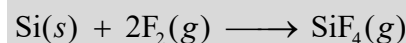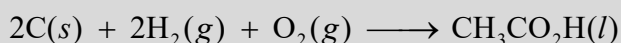
Using the data in Appendix G, calculate the standard enthalpy change for each of the following reactions:

(a) $Si(s) + 2F_2(g) \longrightarrow SiF_4(g)$

(b) $2C(s) + 2H_2(g) + O_2(g) \longrightarrow CH_3CO_2H(l)$

(c) $CH_4(g) + N_2(g) \longrightarrow HCN(g) + NH_3(g)$

(d) $CS_2(g) + 3Cl_2(g) \longrightarrow CCl_4(g) + S_2Cl_2(g)$

Solution

(a)

$Si(s) + 2F_2(g) \longrightarrow SiF_4(g)$

$\Delta H^\circ_{reaction} = \Delta H^\circ_{products} - \Delta H^\circ_{reactants}$;

$= \Delta H^\circ_{SiF_4(g)} - \Delta H^\circ_{Si(s)} - 2\Delta H^\circ_{F_2(g)}$

$= -1614.9 - (0) - 2(0) = -1615.0 \text{ kJ mol}^{-1}$

$2C(s) + 2H_2(g) + O_2(g) \longrightarrow CH_3CO_2H(l)$

(b)
$\Delta H^\circ_{reaction} = \Delta H^\circ_{products} - \Delta H^\circ_{reactants}$;

$= \Delta H^\circ_{CH_3CO_2H(l)} - 2\Delta H^\circ_{C(s)} - 2\Delta H^\circ_{H_2(g)} - \Delta H^\circ_{O_2(g)}$

$= -484.5 - 2(0) - 2(0) - (0) = -484.3 \text{ kJ mol}^{-1}$

(c)
$CH_4(g) \longrightarrow C(s) + 2H_2(g)$	$\Delta H^\circ_1 = -(-74.6 \text{ kJ})$
$\frac{1}{2}H_2(g) + C(s) + \frac{1}{2}N_2(g) \longrightarrow HCN(g)$	$\Delta H^\circ_2 = 135.5 \text{ kJ}$;
$\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \longrightarrow NH_3(g)$	$\Delta H^\circ_3 = -45.9 \text{ kJ}$
$CH_4(g) + N_2(g) \longrightarrow HCN(g) + NH_3(g)$	$\Delta H^\circ = 164.2 \text{ kJ}$
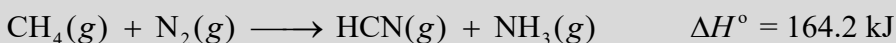
(d)
$CS_2(g) \longrightarrow C(s) + 2S(s)$	$\Delta H^\circ_1 = -(116.9 \text{ kJ})$
$C(s) + 2Cl_2(g) \longrightarrow CCl_4(g)$	$\Delta H^\circ_2 = -95.7 \text{ kJ}$
$2S(s) + Cl_2(g) \longrightarrow S_2Cl_2(g)$	$\Delta H^\circ_3 = -19.50 \text{ kJ}$
$CS_2(g) + 3Cl_2(g) \longrightarrow CCl_4(g) + S_2Cl_2(g)$	$\Delta H^\circ = -232.1 \text{ kJ}$
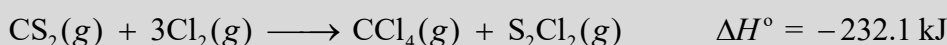
The following reactions can be used to prepare samples of metals. Determine the enthalpy change under standard state conditions for each.

(a) $2Ag_2O(s) \longrightarrow 4Ag(s) + O_2(g)$

(b) $SnO(s) + CO(g) \longrightarrow Sn(s) + CO_2(g)$

(c) $Cr_2O_3(s) + 3H_2(g) \longrightarrow 2Cr(s) + 3H_2O(l)$

(d) $2Al(s) + Fe_2O_3(s) \longrightarrow Al_2O_3(s) + 2Fe(s)$

Solution

(a) $\Delta H^\circ = -2\Delta H^\circ_{f(Ag_2O(s))} = -2(-31.5\ kJ) = 63.0\ kJ$; (b)

$\Delta H^\circ = \Delta H^\circ_{(CO_2(g))} - \Delta H^\circ_{(SnO(s))} - \Delta H^\circ_{(CO(g))} = -393.51\ kJ - (-285.8\ kJ) - (-110.52\ kJ) = 2.8\ kJ$

; (c) $\Delta H^\circ = 3\Delta H^\circ_{(H_2O(l))} - \Delta H^\circ_{(Cr_2O_3(s))} = 3(-285.83) - (-1139.7) = 282.2\ kJ$; (d)

$\Delta H^\circ = \Delta H^\circ_{(Al_2O_4(s))} - \Delta H^\circ_{(Fe_2O_3(s))} = -1676 - (-824.2) = -852\ kJ$

Question 30-33.

The decomposition of hydrogen peroxide, H_2O_2, has been used to provide thrust in the control jets of various space vehicles. Using the data in Appendix G, determine how much heat is produced by the decomposition of exactly 1 mole of H_2O_2 under standard conditions.

$$2H_2O_2(l) \longrightarrow 2H_2O(g) + O_2(g)$$
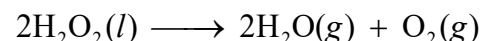
Solution

$$\Delta H^\circ_{reaction} = \Delta H^\circ_{products} - \Delta H^\circ_{reactants}$$
$$= 2\Delta H^\circ_{H_2O(g)} + \Delta H^\circ_{O_2(g)} - 2\Delta H^\circ_{H_2O_2(l)}$$
$$= 2(-241.82) + 0 - 2(-187.78) = -108.08\ kJ$$

The value relates to the decomposition of 2 mol of hydrogen peroxide. For 1 mol,

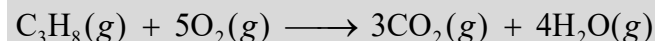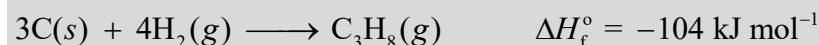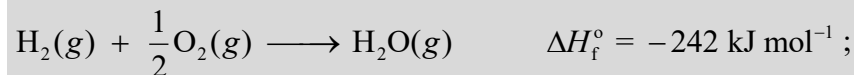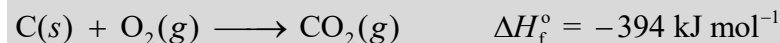$$\frac{-108.0}{2} = -54.04\ kJ\ mol^{-1}.$$

Question 30-34.

Calculate the enthalpy of combustion of propane, $C_3H_8(g)$, for the formation of $H_2O(g)$ and $CO_2(g)$. The enthalpy of formation of propane is $-104\ kJ/mol$.
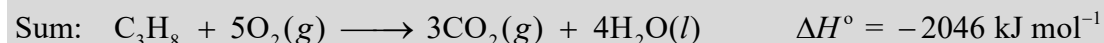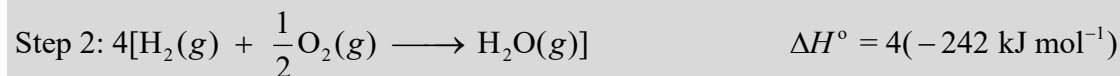
Solution

The equation for the combustion of $C_3H_8(g)$ is:

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

Thus, in order to work this problem, we have to have three equations with known ΔH°_f :

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H^\circ_f = -394\ kJ\ mol^{-1}$$

$$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g) \qquad \Delta H^\circ_f = -242\ kJ\ mol^{-1}\ ;$$

$$3C(s) + 4H_2(g) \longrightarrow C_3H_8(g) \qquad \Delta H^\circ_f = -104\ kJ\ mol^{-1}$$

We want C_3H_8 as a reactant, so we will reverse the third reaction (along with making the ΔH positive). The other two reactions will have to be multiplied appropriate factors. The answers are shown to four significant figures.

Step 1: $3[C(s) + O_2(g) \longrightarrow CO_2(g)]$ $\qquad\qquad \Delta H^\circ = 3(-394\ kJ\ mol^{-1})$

Step 2: $4[H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)]$ $\qquad\quad \Delta H^\circ = 4(-242\ kJ\ mol^{-1})$

Step 3: $\underline{C_3H_8 \longrightarrow 3C(s) + 4H_2(g)}$ $\qquad\qquad\qquad\ \underline{\Delta H^\circ = +104\ kJ\ mol^{-1}}$

Sum: $\ C_3H_8 + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$ $\qquad \Delta H^\circ = -2046\ kJ\ mol^{-1}$

Question 30-35.

Calculate the enthalpy of combustion of butane, $C_4H_{10}(g)$ for the formation of $H_2O(g)$ and $CO_2(g)$. The enthalpy of formation of butane is -126 kJ/mol.

Solution

Step 1: $4[C(s) + O_2(g) \longrightarrow CO_2(g)]$ $\qquad\qquad$ $\Delta H° = 4(-394 \text{ kJ mol}^{-1})$

Step 2: $5[H_2(g) + \dfrac{1}{2}O_2(g) \longrightarrow H_2O(g)]$ $\qquad$ $\Delta H° = 5(-242 \text{ kJ mol}^{-1})$

Step 3: $\underline{C_4H_{10} \longrightarrow 4C(s) + 5H_2(g)}$ $\qquad\qquad$ $\underline{\Delta H^v = +126 \text{ kJ mol}^{-1}}$

Sum: $\quad C_4H_{10} + \dfrac{13}{2}O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$ $\qquad$ $\Delta H° = -2660 \text{ kJ mol}^{-1}$
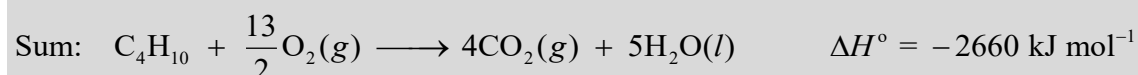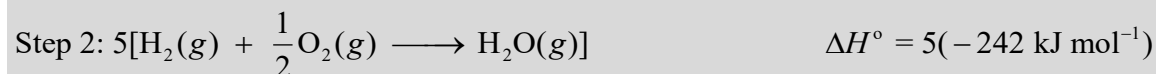
Question 30-36.

Both propane and butane are used as gaseous fuels. Which compound produces more heat per gram when burned?

Solution

The molar masses of the two compounds are $C_3H_8 = 44.1$ g/mol and $C_4H_{10} = 58.1$ g/mol. The kilojoules per gram for each compound is:

$C_3H_8:$ $\quad \dfrac{-2046 \text{ kJ}}{\text{mol}} \times \dfrac{1 \text{ mol}}{44.1 \text{ g}} = -46.4 \text{ kJ g}^{-1}$

$C_4H_{10}:$ $\quad \dfrac{-2660 \text{ kJ}}{\text{mol}} \times \dfrac{1 \text{ mol}}{58.1 \text{ g}} = -45.8 \text{ kJ g}^{-1}$
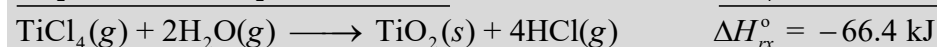
Therefore, propane, C_3H_8, gives more heat per gram when combusted.

Question 30-37.

The white pigment TiO_2 is prepared by the reaction of titanium tetrachloride, $TiCl_4$, with water vapor in the gas phase: $TiCl_4(g) + 2H_2O(g) \longrightarrow TiO_2(s) + 4HCl(g)$.

How much heat is evolved in the production of exactly 1 mole of $TiO_2(s)$ under standard state conditions?

Solution

$TiCl_4(g) \longrightarrow Ti(s) + 2Cl_2(g)$ $\qquad\qquad$ $\Delta H_1° = -(-763.2 \text{ kJ})$

$2H_2O(g) \longrightarrow 2H_2(g) + O_2(g)$ $\qquad\qquad$ $\Delta H_2° = -2(-241.82 \text{ kJ}) = 483.6 \text{ kJ}$

$Ti(s) + O_2(g) \longrightarrow TiO_2(s)$ $\qquad\qquad$ $\Delta H_3° = -944.0 \text{ kJ}$

$\underline{2H_2(g) \longrightarrow 2Cl_2(g) + 4HCl(g)}$ $\qquad\qquad$ $\underline{\Delta H_4° = 4(-92.3 \text{ kJ}) = -369.2 \text{ kJ}}$

$TiCl_4(g) + 2H_2O(g) \longrightarrow TiO_2(s) + 4HCl(g)$ $\qquad$ $\Delta H_{rx}° = -66.4 \text{ kJ}$

Thus, 66.4 kJ of heat is evolved.

Question 30-38.

Water gas, a mixture of H_2 and CO, is an important industrial fuel produced by the reaction of steam with red hot coke, essentially pure carbon: $C(s) + H_2O(g) \longrightarrow CO(g) + H_2(g)$.

(a) Assuming that coke has the same enthalpy of formation as graphite, calculate $\Delta H°$ for this reaction.

(b) Methanol, a liquid fuel that could possibly replace gasoline, can be prepared from water gas and additional hydrogen at high temperature and pressure in the presence of a suitable catalyst: $2H_2(g) + CO(g) \longrightarrow CH_3OH(g)$.

Under the conditions of the reaction, methanol forms as a gas. Calculate $\Delta H°$ for this reaction and for the condensation of gaseous methanol to liquid methanol.

(c) Calculate the heat of combustion of 1 mole of liquid methanol to $H_2O(g)$ and $CO_2(g)$.

Solution

(a) $\Delta H° = \Delta H_f°(CO(g)) - \Delta H_f°(H_2O(g)) = -110.52 - (-241.82) = 131.30 \text{ kJ}$;

(b) $\Delta H° = \Delta H_f°(CH_3OH(g)) - \Delta H_f°(CO(g)) = -201.0 - (-110.52) = -90.5 \text{ kJ}$;

For the condensation,

$CH_3OH(g) \longrightarrow CH_3OH(l)$

$\Delta H° = \Delta H_f°(CO_2(g)) - \Delta H° = \Delta H_f°(CH_3OH(g)) = -239.2 - (-201.0) = -38.2 \text{ kJ}$;

$CH_3OH(l) + \dfrac{3}{2}O_2 \longrightarrow CO_2(g) + 2H_2O(g)$
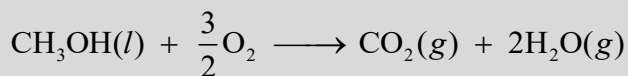
(c) $\Delta H° = \Delta H_f°(CO_2(g)) + 2\Delta H_f°(H_2O(g)) - \Delta H_f°(CH_3OH(l))$

$= -393.51 + 2(-241.82) - (-239.2) = -638.0 \text{ kJ}$

Question 30-39.

In the early days of automobiles, illumination at night was provided by burning acetylene, C_2H_2. Though no longer used as auto headlamps, acetylene is still used as a source of light by some cave explorers. The acetylene is (was) prepared in the lamp by the reaction of water with calcium carbide, CaC_2:

$CaC_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(s) + C_2H_2(g)$.
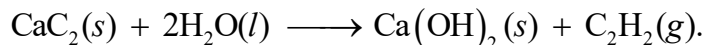
Calculate the standard enthalpy of the reaction. The $\Delta H_f°$ of CaC_2 is –15.14 kcal/mol.

Solution

convert –15.14 kcal to kJ:

–15.14 kcal × 4.184 kJ/kcal = 63.35 kJ/mol

$\Delta H°_{reaction} = \Delta H°_{products} - \Delta H°_{reactants}$

$= \Delta H°_{Ca(OH)_2(s)} + \Delta H°_{C_2H_2(g)} - \Delta H°_{CaC_2(s)} - 2\Delta H°_{H_2O(l)}$

$= -985.2 + 227.4 - (-63.35) - 2(-285.83) = -122.8 \text{ kJ}$

Question 30-40.

From the data in Table 5.2, determine which of the following fuels produces the greatest amount of heat per gram when burned under standard conditions: $CO(g)$, $CH_4(g)$, or $C_2H_2(g)$.

Solution

$$CO(g): \quad \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{28.01 \text{ g}} = -10.20 \text{ kJ/g}$$

$$CH_4(g): \quad \frac{-809.8 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{16.04 \text{ g}} = -55.54 \text{ kJ/g}$$

$$C_2H_2(g): \quad \frac{-1301.1 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{26.04 \text{ g}} = -49.97 \text{ kJ/g}$$

$CH_4(g)$ produces the greatest heat per gram.

Question 30-41.

The enthalpy of combustion of hard coal averages -35 kJ/g, that of gasoline, -1.28×10^5 kJ/gal. How many kilograms of hard coal provide the same amount of heat as is available from 1.0 gallon of gasoline? Assume that the density of gasoline is 0.692 g/mL (the same as the density of isooctane).

Solution

the amount of heat produced by burning of 1.0 gallon of gasoline is:

$q = 1.0$ gallon $\times$ (-1.28×10^5 kJ/gal) $= -1.28 \times 10^5$ kJ

Mass $\times$ (-35 kJ/g) $= -1.28 \times 10^5$ kJ

Mass $= 3657$ g or 3.7 kg

Question 30-42.

Ethanol, C_2H_5OH, is used as a fuel for motor vehicles, particularly in Brazil.

(a) Write the balanced equation for the combustion of ethanol to $CO_2(g)$ and $H_2O(g)$, and, using the data in Appendix G, calculate the enthalpy of combustion of 1 mole of ethanol.

(b) The density of ethanol is 0.7893 g/mL. Calculate the enthalpy of combustion of exactly 1 L of ethanol.

(c) Assuming that an automobile's mileage is directly proportional to the heat of combustion of the fuel, calculate how much farther an automobile could be expected to travel on 1 L of gasoline than on 1 L of ethanol. Assume that gasoline has the heat of combustion and the density of n–octane, C_8H_{18} ($\Delta H_f^\circ = -208.4$ kJ/mol; density = 0.7025 g/mL).

Solution

(a) $C_2H_5OH(l) + 3O_2 \longrightarrow 3H_2O(g) + 2CO_2(g)$;
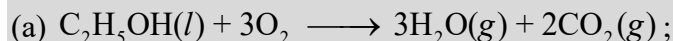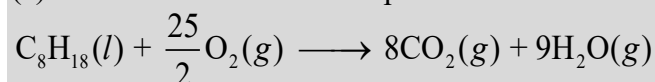
(b) the molecular weight of ethanol is 46.077 g/mol, then:

$$\frac{46.07 \text{ g mol}^{-1}}{0.7839 \text{ g mL}^{-1}} = 58.37 \text{ mL mol}^{-1}$$

$$\Delta H_{comb} = \frac{-1234.9 \text{ g mol}^{-1}}{58.37 \text{ g mL}^{-1}} = \frac{-21.17 \text{ kJ}}{\text{mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = -2.17 \times 10^4 \text{ kJ/L} \quad ;$$

(c) the heat of combustion expression is

$$C_8H_{18}(l) + \frac{25}{2}O_2(g) \longrightarrow 8CO_2(g) + 9H_2O(g)$$

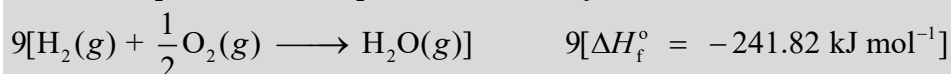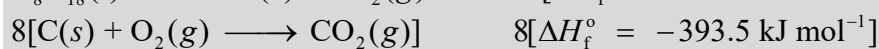$$\Delta H^\circ = 8\Delta H_f^\circ(CO_2(g)) + 9\Delta H_f^\circ(H_2O(g)) - \Delta H_f^\circ(C_8H_{18}(l))$$

and the

$$= 8(-393.51) + 9(-241.82) - (-208.4) = -5116.1 \text{ kJ/mol}$$

molar mass of C_8H_{18} is 114.23 g/mol; $\dfrac{114.23 \text{ g/mol}}{0.7025 \text{ g/mL}} = 162.6 \text{ mL/mol}$

$$\dfrac{-5116.1 \text{ kJ/mol}}{162.6 \text{ mL/mol}} = 31.46 \text{ kJ/mol} = -3.146 \times 10^4 \text{kJ/L}$$

$$\dfrac{-3.146 \times 10^4 \text{ kJ/L}}{-2.17 \times 10^4 \text{ kJ/L}} = 1.45 ;$$

gasoline will make the automobile travel 1.45 times farther than ethanol; from the heats of formation we have:

$$C_8H_{18}(l) \longrightarrow 8C(s) + 9H_2(g) \qquad\qquad -[\Delta H_f^\circ = -208.4 \text{ kJ mol}^{-1}]$$
$$8[C(s) + O_2(g) \longrightarrow CO_2(g)] \qquad\qquad 8[\Delta H_f^\circ = -393.5 \text{ kJ mol}^{-1}]$$
$$9[H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(g)] \qquad\qquad 9[\Delta H_f^\circ = -241.82 \text{ kJ mol}^{-1}]$$

$$\Delta H_{comb} = \dfrac{-5116.0 \text{ kJ mol}^{-1}}{162.49 \text{ mL mol}^{-1}} = -31.485 \text{ kJ mL}$$
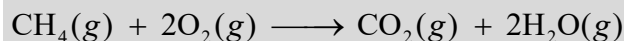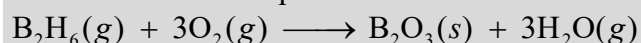
$$\dfrac{\Delta H_{comb}(octane)}{\Delta H_{comb}(ethanol)} = \dfrac{-31.485 \text{ kJ/mL}}{-21.15 \text{ kJ/mL}} = 1.489 \text{ times farther}$$

Question 30-43.

Among the substances that react with oxygen and that have been considered as potential rocket fuels are diborane [B_2H_6, produces $B_2O_3(s)$ and $H_2O(g)$], methane [CH_4, produces $CO_2(g)$ and $H_2O(g)$], and hydrazine [N_2H_4, produces $N_2(g)$ and $H_2O(g)$]. On the basis of the heat released by 1.00 g of each substance in its reaction with oxygen, which of these compounds offers the best possibility as a rocket fuel? The ΔH_f° of $B_2H_6(g)$, $CH_4(g)$, and $N_2H_4(l)$ may be found in Appendix G.

Solution

Write the balanced equation for each reaction.

$$B_2H_6(g) + 3O_2(g) \longrightarrow B_2O_3(s) + 3H_2O(g)$$

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

$$N_2H_4(l) + O_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

Calculate the heat released per mole; then, calculate the heat per gram.

$$\Delta H_{rxn}^\circ = \Delta H_P^\circ - \Delta H_R^\circ$$

$$\Delta H_{comb}^\circ = -1273.5 + 3(-241.82) - 36.4 = -2035.36 \text{ kJ mol}^{-1}$$

$$\Delta H_{comb}^\circ = -393.51 + 2(-241.82) - (-74.6) = -802.6 \text{ kJ mol}^{-1}$$

$$\Delta H_{comb}^\circ = 2(-241.82) - (50.63) = -534.27 \text{ kJ mol}^{-1}$$

Calculate the heat per mole released per gram.

For B_2H_6: $\dfrac{-2035.36 \text{ kJ}}{27.610 \text{ g}} = -73.72 \text{ kJ g}^{-1}$

For CH4: $\dfrac{-802.34 \text{ kJ}}{16.043 \text{ g}} = -50.03 \text{ kJ g}^{-1}$

For N2H4: $\dfrac{-534.27 \text{ kJ}}{32.0452 \text{ g}} = 16.67 \text{ kJ g}^{-1}$

On the assumption that the best rocket fuel is the one that gives off the most heat, B2H6 is the prime candidate. Other things must be considered, however. For example, the moles of gaseous product formed are related to the specific impulse of the fuel, toxicity of products, cost, and ability to contain original fuel (stability and corrosiveness).

Question 30-44.

How much heat is produced when 1.25 g of chromium metal reacts with oxygen gas under standard conditions?

Solution

The heat of formation of Cr2O3 is –1139.7 kJ/mol. Then, the heat released is 0.0120 ~~mol~~ × –1140 kJ/~~mol~~ = –13.7 kJ.

Question 30-45.

Ethylene, C2H4, a byproduct from the fractional distillation of petroleum, is fourth among the 50 chemical compounds produced commercially in the largest quantities. About 80% of synthetic ethanol is manufactured from ethylene by its reaction with water in the presence of a suitable catalyst. $C_2H_4(g) + H_2O(g) \longrightarrow C_2H_5OH(l)$

Using the data in the table in Appendix G, calculate $\Delta H°$ for the reaction.

Solution

$$\Delta H^o_{rxn} = \Delta H^o_{C_2H_5OH(l)} - \Delta H^o_{H_2O(g)} - \Delta H^o_{C_2H_4(g)}$$

$$= -277.7 - (-241.82) - (52.4)$$

$$= -88.2 \text{ kJ}$$

Question 30-46.

The oxidation of the sugar glucose, C6H12O6, is described by the following equation:

$$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l) \qquad \Delta H = -2816 \text{ kJ}$$
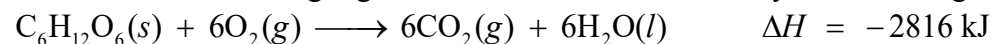
The metabolism of glucose gives the same products, although the glucose reacts with oxygen in a series of steps in the body.

    (a) How much heat in kilojoules can be produced by the metabolism of 1.0 g of glucose?

    (b) How many Calories can be produced by the metabolism of 1.0 g of glucose?

Solution

(a) The molar mass of glucose is 180.158 g/mol

$$\dfrac{\Delta H_{comb}}{g} = \dfrac{-2816 \text{ kJ } \text{mol}^{-1}}{180.16 \text{ g } \text{mol}} = -15.63 \text{ kJ g}^{-1} \text{ or } -16 \text{ kJ g}^{-1}$$

to the correct number of significant figures. Thus, 16 kJ of heat are released.

(b) $\dfrac{-15.63 \text{ kJ}}{1 \text{ g}} \times \dfrac{1000 \text{ J}}{1 \text{ kJ}} \times \dfrac{1 \text{ cal}}{4.184 \text{ J}} \times \dfrac{1 \text{ Cal}}{1000 \text{ cal}} = 3.7 \text{ Calories g}^{-1}.]$

Propane, C_3H_8, is a hydrocarbon that is commonly used as a fuel.

(a) Write a balanced equation for the complete combustion of propane gas.

(b) Calculate the volume of air at 25 °C and 1.00 atmosphere that is needed to completely combust 25.0 grams of propane. Assume that air is 21.0 percent O_2 by volume. (Hint: We will see how to do this calculation in a later chapter on gases—for now use the information that 1.00 L of air at 25 °C and 1.00 atm contains 0.275 g of O_2.)

(c) The heat of combustion of propane is –2219.2 kJ/mol. Calculate the heat of formation, ΔH_f°, of propane given that ΔH_f° of $H_2O(l) = -285.8$ kJ/mol and ΔH_f° of $CO_2(g) = -393.5$ kJ/mol.

(d) Assuming that all of the heat released in burning 25.0 grams of propane is transferred to 4.00 kilograms of water, calculate the increase in temperature of the water.

Solution

(a) $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$
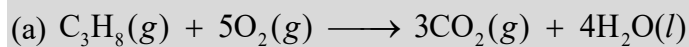
(b) Determine the number of moles of O_2 required and from that the number of grams. Then use the density to find the volume.

$$\text{Mass } O_2 = 25.0 \text{ g } C_3H_8 \times \frac{1 \text{ mol } C_3H_8}{44.097 \text{ g}} \times \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8(g)} \times \frac{31.999 \text{ g}}{1 \text{ mol } O_2} = 90.7 \text{ g}$$

$$\text{Volume air} = \frac{90.7 \text{ g}}{0.275 \text{ g } L^{-1}} = 330 \text{ L}$$

(c)

$$\Delta H^\circ_{comb}(C_3H_6(g)) = -2219.2 \text{ kJ mol}^{-1} = 3\Delta H_f^\circ(CO_2(g)) + 4\Delta H_f^\circ(H_2O(l)) - \Delta H_f^\circ(C_3H_6(g))$$

$$\Delta H_f^\circ(C_3H_6(g)) = [2220.1 + 3(-393.5) + 4(-285.8)] \text{ kJ mol}^{-1}$$

$$= -104.5 \text{ kJ mol}^{-1}$$

$$\frac{25.0 \text{ g}}{44.097 \text{ g mol}^{-1}} \times 2219.2 \text{ kJ mol}^{-1} = 1258 \text{ kJ} = 1.26 \times 10^6 \text{ J}$$

(d) Heat $= 1.26 \times 10^6$ J $= 4.18$ J g^{-1} °C^{-1} $\times 4.00 \times 10^3 \times \Delta T$

$\Delta T = 75.4$ °C

During a recent winter month in Sheboygan, Wisconsin, it was necessary to obtain 3500 kWh of heat provided by a natural gas furnace with 89% efficiency to keep a small house warm (the efficiency of a gas furnace is the percent of the heat produced by combustion that is transferred into the house).

(a) Assume that natural gas is pure methane and determine the volume of natural gas in cubic feet that was required to heat the house. The average temperature of the natural gas was 56 °F; at this temperature and a pressure of 1 atm, natural gas has a density of 0.681 g/L.

(b) How many gallons of LPG (liquefied petroleum gas) would be required to replace the natural gas used? Assume the LPG is liquid propane [C_3H_8: density, 0.5318 g/mL; enthalpy

of combustion, 2219 kJ/mol for the formation of $CO_2(g)$ and $H_2O(l)$] and the furnace used to burn the LPG has the same efficiency as the gas furnace.

(c) What mass of carbon dioxide is produced by combustion of the methane used to heat the house?

(d) What mass of water is produced by combustion of the methane used to heat the house?

(e) What volume of air is required to provide the oxygen for the combustion of the methane used to heat the house? Air contains 23% oxygen by mass. The average density of air during the month was 1.22 g/L.

(f) How many kilowatt-hours (1 kWh = 3.6×10^6 J) of electricity would be required to provide the heat necessary to heat the house? Note electricity is 100% efficient in producing heat inside a house.

(g) Although electricity is 100% efficient in producing heat inside a house, production and distribution of electricity is not 100% efficient. The efficiency of production and distribution of electricity produced in a coal-fired power plant is about 40%. A certain type of coal provides 2.26 kWh per pound upon combustion. What mass of this coal in kilograms will be required to produce the electrical energy necessary to heat the house if the efficiency of generation and distribution is 40%?

Solution

(a) The total heat that has to be obtained from combustion of methane is

$$q = 3500 \text{ kWh} \times 1000 \text{ W/kW} \times \frac{3600 \text{ J/Wh}}{0.89} = \frac{1.26 \times 10^{10} \text{ kJ}}{0.89} = 1.42 \times 10^7 \text{ kJ}.$$

Methane's heat of combustion is –890.8 kJ/mol as given in Table 5.2. The number of moles of

methane that have to be burned is $\dfrac{1.42 \times 10^7 \text{ kJ}}{890.8 \text{ kJ/mol}} = 15{,}940 \text{ mol}$. The molar mass of methane is

16.04 g/mol, so the total mass of methane needed is 15940 mol × 16.04 g/mol = 255,700 g. The volume of this mass of methane is

$$\frac{255700 \text{ g}}{0.681 \text{ g/L}} = 375500 \text{ L} . \quad \frac{375500 \text{ L}}{28.316 \text{ L/cubic foot}} = 13{,}260 \text{ cubic feet (the conversion factor is}$$

taken from Appendix C). (b) We found in (a) that the total amount of heat to be obtained is $q = $

1.42×10^7 kJ. The number of moles of propane required is $\dfrac{1.42 \times 10^7 \text{ kJ}}{2219 \text{ kJ/mol}} = 6399 \text{ mol}$. Given

the molar mass of propane, 44.1 g/mol, the mass of propane needed is 6399 mol × 44.1 g/mol =

282,200 g. The volume of LPG is $\dfrac{28220 \text{ g}}{0.5318 \text{ g/mL}} = 530700 \text{ mL}$ or 530.7 L or

$530.7 \text{ L} \times \dfrac{1 \text{ gallon}}{(4 \times 0.9463 \text{ L})} = 140.2 \text{ gallons of LPG}$. With three significant figures, the answer is

140 gallons of LPG. (c) The reaction is $CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$.We have

determined in part (a) that 15,940 mol of methane are needed. According to the reaction equation, this leads to production of the same amount of CO_2. Given that the molar mass of CO_2 is 44.01 g/mol, the mass of produced CO_2 is 15940 mol × 44.01 g/mol = 701,500 g or 701.5 kg. Leaving only three significant digits, the resulting mass of CO_2 is 702 kg. (d) The reaction equation is given in part (c). 15,940 mol of methane are needed. According to the reaction equation, this leads to the production of twice the amount of H_2O or 31,880 mol of H_2O. Given

that the molar mass of H_2O is 18.015 g/mol, the mass of produced H_2O is 31880 mol × 18.015 g/mol = 574,300 g or 574.3 kg. Leaving only three significant digits, the resulting mass of H_2O is 574 kg. (e) Two moles of oxygen are consumed in combustion of each mole of methane. Thus, the overall amount of oxygen needed is 2 × 15,940 mol = 31,880 mol. The molar mass of oxygen is 32.0 g/mol; thus, the mass of oxygen required is 31,880 mol × 32.0 g/mol = 1,020,000 g. Given that only 23% of air is oxygen, the overall mass of air is $\dfrac{1020000 \text{ g}}{0.23} = 4435000 \text{ g}$. The volume of air required is $\dfrac{4435000 \text{ \cancel{g}}}{1.22 \text{ \cancel{g}/L}} = 3635000$ L or 3635 m³. Keeping only three significant figures, we obtain a result of 3640 m³ of air. (f) Since the efficiency is 100% in this case, the amount of energy is the same as specified in the problem or 3500 kWh. (g) The overall amount of heat to be obtained from the burning of coal is $\dfrac{3500 \text{ kWh}}{0.4} = 8750 \text{ kWh}$. The mass of coal required for this process is $\dfrac{8750 \text{ \cancel{kWh}}}{2.26 \text{ \cancel{kWh}/lb}} = 3872 \text{ lb}$. The mass of coal in kilograms is 3872 lb × 0.4536 kg/lb = 1620 kg.