Unit 5 Opener

Solutions to Practice Problems

1.

Problem

100 g of ethanol at 25 °C is heated until it reaches 50 °C. How much thermal energy does the ethanol gain?

What is Required?

You must calculate the thermal energy (heat) gained by a sample of ethanol when its temperature is increased.

What is Given?

mass ethanol, $m = 100$ g
initial temperature, $t_i = 25\ ^oC$
final temperature, $t_f = 50\ ^oC$
specific heat capacity for ethanol, $c = 2.44$ J/g ^oC

Plan Your Strategy

Determine the change in temperature, Δt.
Calculate the heat gained using the relationship, $Q = mc\Delta t$

Act on Your Strategy

$\Delta t = t_f - t_i = 50\ ^oC - 25\ ^oC = 25\ ^oC$
$Q = mc\Delta t$
$= 100 \text{ g} \times 2.44 \text{ J/g } ^oC \times 25\ ^oC$
$= 6.1 \times 10^3$ J or 6.1 kJ

Check Your Solution

The temperature increases and therefore the change in temperature is positive. Since heat is absorbed Q is positive and has the correct unit (J) and number of significant digits (2).

2.

Problem

Beaker A contains 50 g of liquid at room temperature. The beaker is heated until the liquid increases in temperature by 10 °C. Beaker B contains 100 g of the same liquid at room temperature. The beaker is also heated until the liquid increases in temperature by 10 °C. In which beaker does the liquid absorb the most heat? Explain.

What is Required?

You must compare the heat energy gained by two samples of liquid that have different masses when both increase in temperature by 10 ^oC.

What is Given?

$m_A = 50$ g

$m_B = 100$ g
$\Delta t_A = \Delta t_B = 10\ ^oC$
The specific heat capacity, c, is the same for both samples because the liquid is the same.

Plan Your Strategy

Use the equation $Q = mc\Delta t$ to compare the heat gained by each sample.

Act on Your Strategy

$$\frac{Q_A}{Q_B} = \frac{m_A c\Delta t}{m_B c\Delta t} = \frac{m_A}{m_B}$$

$$\therefore Q_B = Q_A \times \frac{m_B}{m_A} = Q_A \times \frac{100\text{g}}{50\text{g}}$$

$Q_B = 2Q_A$

Check Your Solution

Since the temperature change is the same for each sample and both samples are the same liquid, the heat gained will be directly proportional to the mass of the samples. Sample B should gain twice as much energy as sample A.

3.

Problem

How much heat is released when the temperature of 789 g of liquid ammonia decreases from 82.7 °C to 25.0 °C?

What is Required?

You must calculate the heat lost by a sample of liquid ammonia when its temperature is decreased.

What is Given?

mass $NH_3(l)$, $m = 789$ g
initial temperature, $t_i = 82.7\ ^oC$
final temperature, $t_f = 25.0\ ^oC$
specific heat capacity $NH_3(l)$, $c = 4.70$ J/g ^oC

Plan Your Strategy

Determine the change in temperature, Δt.
Calculate the heat lost using the equation $Q = mc\Delta t$.

Act on Your Strategy

$\Delta t = t_f - t_i = 25.0\ ^oC - 82.7\ ^oC = -57.7\ ^oC$
$Q = mc\Delta t$
$= 789 \text{ g} \times 4.70 \text{ J/g } ^oC \times -57.7\ ^oC$

$= -2.14 \times 10^5\ \text{J} = -214\ \text{kJ}$

Check Your Solution
The temperature decreases and therefore the change in temperature is negative. Since heat is released, *Q* is negative and has the correct unit (J or kJ) and the correct number of significant digits (3).

4.
Problem
A solid substance has a mass of 250.00 g. It is cooled by 25.00 °C and loses 4.937 kJ of heat. What is the specific heat capacity? Identify the substance using the values in Table P5.1.

What is Required?
You must calculate the specific heat capacity of a sample and use this information to identify the solid.

What is Given?
mass solid, $m = 250.00$ g
$\Delta t = -25.00\ ^{\circ}\text{C}$
$Q = -4.937\ \text{kJ} = -4\,937\ \text{J}$

Plan Your Strategy
Solve for c using the $Q = mc\Delta t$.

Act on Your Strategy

$$c = \frac{Q}{m\Delta t} = \frac{-4937\text{J}}{250.00\ \text{g} \times (-25.00^{\circ}\ \text{C})} = 0.7899\ \text{J/g}\ ^{\circ}\text{C}$$

The substance is granite.

Check Your Solution
The specific heat capacity, *c*, has the correct units (J/g °C) and the correct number of significant digits (4).

Chapter 9 Energy and Chemical Reactions

Solutions to Practice Problems

1.

Problem

Hydrogen gas and oxygen gas react to form 0.534 g of *gaseous* water. What is the enthalpy change?
$\Delta_r H = -241.8$ kJ/mol

What is Required?

You must determine the enthalpy change, $\Delta_r H$, when a given mass of $H_2O(g)$ forms from the elements $H_2(g)$ and $O_2(g)$.

What is Given?

$m H_2O(g) = 0.534$ g; $\Delta_f H^o$ $H_2O(g) = -241.8$ kJ/mol

Plan Your Strategy

Convert 0.534 g $H_2O(g)$ to moles using $n = \frac{m}{M}$. Use $\Delta_f H^o$ $H_2O(g)$ to calculate $\Delta_r H$.

Act on Your Strategy

$M H_2O(g) = 18.02$ g/mol

$$nH_2O(g) = \frac{0.534\,\text{g}}{18.02\,\frac{\text{g}}{\text{mol}}} = 0.02963 \text{ mol}$$

$0.02963 \text{ mol} \times (-241.8 \text{ kJ/mol}) = -7.17 \text{ kJ}$

Check Your Solution

In each case, the $\Delta_r H$ is negative since energy is released in this reaction. The unit is correct (kJ) and the number of significant digits is correct (3).

2.

Problem

Pentane reacts with an excess of oxygen to produce carbon dioxide and water. What is the enthalpy change or the reaction per mole of each of the following?
a) oxygen
b) carbon dioxide
c) water

What is Required?

You must write the balanced equation for the combustion of pentane.
You must calculate the enthalpy change or the reaction per mole of each of the following: oxygen, carbon dioxide, and water.

What is Given?

Pentane reacts with oxygen to produce carbon dioxide and water.

Plan Your Strategy

Write the balanced equation for the combustion of pentane and then add excess $O_2(g)$.
Using values from Table 9.1, calculate the enthalpy change for the reaction using the formula $\Delta_r H = \Sigma_f H_{products} - \Sigma_f H_{reactants}$.
For each reactant or product required, divide the enthalpy change for the reaction by the number of moles of each reactant and product.

Act on Your Strategy

$C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l)$
$C_5H_{12}(l) = -3244.8$ kJ/mol from Table 9.1.

a) $\Delta H\ O_2(g) = \dfrac{-3244.8\ \text{kJ/mol}\ C_5H_{12}(l)}{8\ \text{mol}\ O_2(g)} = -405.60\ \text{kJ/mol}\ O_2(g)$

b) $\Delta H\ CO_2(g) = \dfrac{-3244.8\ \text{kJ/mol}\ C_5H_{12}(l)}{5\ \text{mol}\ CO_2(g)} = -648.96\ \text{kJ/mol}\ CO_2(g)$

c) $\Delta H\ H_2O(l) = \dfrac{-3244.8\ \text{kJ/mol}\ C_5H_{12}(l)}{5\ \text{mol}\ H_2O(g)} = -540.80\ \text{kJ/mol}\ H_2O(l)$

Check Your Solution

The answers seem reasonable and have the correct units (kJ/mol).

3.

Problem

Determine the thermal energy released by the combustion of each of the following samples of hydrocarbons in an open system:
a) 56.78 g pentane, $C_5H_{12}(l)$
b) 1.36 kg octane, $C_8H_{18}(l)$
c) 2.344×10^4 g heptane, $C_7H_{16}(l)$

What is Required?

For each given mass of hydrocarbon, the thermal energy released during combustion, $\Delta_r H$, must be calculated.

What is Given?

a) $mC_5H_{12}(l) = 56.78$ g; $\Delta_c H\ C_5H_{12}(l) = -3244.8$ kJ/mol
b) $mC_8H_{18}(l) = 1.36$ kg $= 1.36 \times 10^3$ g; $\Delta_c H\ C_8H_{18}(l) = -5074.1$ kJ/mol
c) $mC_7H_{16}(l) = 2.344 \times 10^4$ g; $\Delta_c H\ C_7H_{16}(l) = -4464.7$ kJ/mol

Plan Your Strategy

a) Determine the molar mass, $MC_5H_{12}(l)$, and calculate the moles of $C_5H_{12}(l)$ using

$n = \frac{m}{M}$. Use Δ_cH $C_5H_{12}(l)$ to calculate Δ_rH.

b) Determine the molar mass, $MC_8H_{18}(l)$, and calculate the moles of $C_8H_{18}(l)$ using

$n = \frac{m}{M}$. Use Δ_cH $C_8H_{18}(l)$ to calculate Δ_rH.

c) Determine the molar mass, $MC_7H_{16}(l)$, and calculate the moles of $C_7H_{16}(l)$ using

$n = \frac{m}{M}$. Use Δ_cH $C_7H_{16}(l)$ to calculate Δ_rH.

Act on Your Strategy

a) $MC_5H_{12}(l) =$ 72.17 g/mol

$$nC_5H_{12}(l) = \frac{56.78\text{ g}}{72.17\text{ g/mol}} = 0.78675\text{ mol}$$

$0.78675\text{ mol} \times (-3244.8\text{ kJ/mol}) = -2553\text{ kJ or } -2.553 \times 10^3\text{ kJ}$

b) $MC_8H_{18}(l) =$ 114.26 g/mol

$$nC_8H_{18}(l) = \frac{1.36\times10^3\text{ g}}{114.26\text{g/mol}} = 11.90\text{ mol}$$

$11.90\text{ mol} \times (-5074.1\text{ kJ/mol}) = -6.04 \times 10^4\text{ kJ}$

c) $MC_7H_{16}(l) =$ 100.23 g/mol

$$nC_7H_{16}(l) = \frac{2.344\times10^4\text{ g}}{100.23\text{ g/mol}} = 233.86\text{ mol}$$

$233.86\text{ mol} \times (-4464.7\text{ kJ/mol}) = -1.044 \times 10^6\text{ kJ}$

Check Your Solution

In each case, the Δ_rH is negative because energy is released in this reaction. The units are correct (kJ) and the numbers of significant digits are correct.

4.

Problem

What is the enthalpy change for the combustion of a 1.00 g sample of methane, $CH_4(g)$, in an open system?

What is Required?

Calculate Δ_rH, released during the combustion of 1.00 g of methane.

What is Given?

$m\text{CH}_4\text{(g)} = 1.00\text{ g}$; $\Delta_cH\,\text{CH}_4\text{(g)} = -802.5\text{ kJ/mol}$

Plan Your Strategy

Convert 1.00 g $\text{CH}_4\text{(g)}$ to moles using $n = \frac{m}{M}$. Use $\Delta_cH\,\text{CH}_4\text{(g)}$ to calculate Δ_rH.

Act on Your Strategy

$M\text{CH}_4\text{(g)} = 16.05\text{ g/mol}$

$$n\text{CH}_4\text{(g)} = \frac{1.00\,\text{g}}{16.05\,\text{g/mol}} = 0.06230\text{ mol}$$

$0.06230\text{ mol} \times (-802.5\text{ kJ/mol}) = -50.0\text{ kJ}$

Check Your Solution

The Δ_rH is negative since energy is released in this reaction. The unit is correct (kJ) and the number of significant digits is correct (3).

5.

Problem

The molar enthalpy of combustion of methanol, $\text{CH}_3\text{OH}(l)$, is –725.9 kJ. What mass of methanol must be burned to generate -2.34×10^4 kJ of thermal energy?

What is Required?

You must calculate the mass of $\text{CH}_3\text{OH}(l)$ that forms from its elements when the enthalpy of reaction is given.

What is Given?

$\Delta_rH\,\text{CH}_3\text{OH}(l) = -2.34 \times 10^4\text{ kJ}$

$\Delta_cH\,\text{CH}_3\text{OH}(l) = -725.9\text{ kJ}$

Plan Your Strategy

Compare $\Delta_cH^\circ\,\text{CH}_3\text{OH}(l)$ to the given $\Delta_rH\,\text{CH}_3\text{OH}(l)$ and calculate the number of moles of $\text{CH}_3\text{OH}(l)$ that reacted. Convert the moles of $\text{CH}_3\text{OH}(l)$ using the equation $m = n \times M$.

Act on Your Strategy

$$\frac{\Delta_r H}{\Delta_c H^\circ} = \frac{-2.34\times10^4\,\text{kJ}}{-725.9\,\text{kJ/mol}} = 32.23585\text{ mol}\,\text{CH}_3\text{OH}(l)$$

$M\text{CH}_3\text{OH}(l) = 32.05\text{ g/mol}$

$m\text{CH}_3\text{OH}(l) = 32.23585\text{ mol} \times 32.05\text{ g/mol} = 1033.159\text{ g} = 1033\text{ g} = 1.03\text{ kg CH}_3\text{OH}(l)$

Check Your Solution

The answer is reasonable for this data. The unit is correct (g or kg) and the number of significant digits is correct (3).

6.

Problem

The enthalpy change for the following reaction, as written is –906 kJ.

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

What is the molar enthalpy change of the reaction for each of the following?

a) ammonia
b) oxygen
c) nitrogen monoxide
d) water

What is Required?

You must calculate the molar enthalpy change of the reaction for each of the following: ammonia, oxygen, nitrogen monoxide, and water.

What is Given?

The enthalpy change for the reaction as written is –906 kJ.
The balanced reaction is given.

Plan Your Strategy

Using values from Tables 9.1 and 9.2, calculate the enthalpy change for the reaction using the formula $\Delta_r H = \Sigma_f H_{\text{products}} - \Sigma_f H_{\text{reactants}}$.
For each reactant or product required, divide the enthalpy change for the reaction by the number of moles of each reactant and product.

Act on Your Strategy

a) $\Delta H \ NH_3(g) = \dfrac{-906 \text{ kJ/mol}}{4} = -226.5 \text{ kJ/mol} = -227 \text{ kJ/mol}$

b) $\Delta H \ O_2(g) = \dfrac{-906 \text{ kJ/mol}}{5} = -181.2 \text{ kJ/mol} = -181 \text{ kJ/mol}$

b) $\Delta H \ NO(g) = \dfrac{-906 \text{ kJ/mol}}{4} = -226.5 \text{ kJ/mol} = -227 \text{ kJ/mol}$

c) $\Delta H \ H_2O(l) = \dfrac{-906 \text{ kJ/mol}}{6} = -151 \text{ kJ/mol}$

Check Your Solution

The answers seem reasonable, have the correct number of significant digits (3), and have the correct units (kJ/mol).

7.

Problem

A chemist wants to determine the enthalpy of neutralization for the following reaction:

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

The chemist uses a simple calorimeter to neutralize completely 61.1 mL of 0.543 mol/L HCl(aq) with 42.6 mL of sufficiently concentrated NaOH(aq). The initial temperature of both solutions is

17.80 °C. After neutralization, the highest recorded temperature is 21.60 °C. Calculate the enthalpy of neutralization in kJ/mol of HCl(aq). Assume that the density of both solutions is 1.00 g/mL.

What is Required?
You must calculate the enthalpy of reaction for the neutralization reaction between HCl(aq) and NaOH(aq).

What is Given?
The mole ratio between the reactants is known from the balanced equation.
VHCl(aq) = 61.1 mL = 0.0611 L
VNaOH(aq) = 42.6 mL
Vtotal = 61.1 mL + 42.6 mL =103.7 mL
Since density of solutions = 1.00 g/mL, $m$total = 103.7 g
cHCl(aq) = 0.543 mol/L
$t_i = 17.8\ ^oC$
$t_f = 21.6\ ^oC$
specific heat capacity of solution = $cH_2O(l) = 4.19$ J/g ^oC

Plan Your Strategy
Determine the temperature change, Δt, of the surroundings (solutions), and calculate $Q_{surroundings}$ using the equation $Q_{surroundings} = mc\Delta t$, where m = total mass of solution.
$\Delta H = -Q_{surroundings}$
Since the reaction is complete, $\Delta_r H$ can be calculated per mole of HCl(aq).
Calculate the number of moles of HCl(aq) using the equation $n = c \times V$ and determine $\Delta_r H$ per mole of HCl(aq).

Act on Your Strategy
$\Delta t = t_f - t_i = 21.6\ ^oC - 17.8\ ^oC = 3.8\ ^oC$
$\Delta H = -mc\Delta t$
$= -(103.7\text{ g} \times 4.19\text{ J/g }^oC \times 3.8\ ^oC)$
$= -1.6511 \times 10^3\text{ J}$

nHCl(aq) = 0.543 mol/L × 0.0611 L = 0.033177 mol

$$\Delta_r H \text{ per mole of HCl(aq)} = \frac{-1.6511 \times 10^3\text{ J}}{0.033177\text{ mol}}$$
$$= -49766.3\text{ J/mol HCl(aq)}$$
$$= -49.8\text{ kJ/mol HCl(aq)}$$

Check Your Solution
The enthalpy of reaction is negative since heat is given off in this reaction. The value calculated for $\Delta_r H$ has the correct unit (kJ/mol) and the correct number of significant digits (2).

8.

Problem

A chemist wants to determine empirically the enthalpy change for the following reaction as written:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The chemist uses a simple calorimeter to react 0.50 g of magnesium ribbon with 100 mL of 1.00 mol/L HCl(aq). The initial temperature of the HCl(aq) is 20.40 °C. After the reaction, the highest recorded temperature is 40.70 °C.

a) Calculate the enthalpy change, in kJ/mol of Mg(s), for the reaction as written.

b) State all assumptions that you made in order to determine the enthalpy change.

What is Required?

You must calculate the enthalpy of reaction, $\Delta_r H$, for the reaction that occurs when magnesium metal is added to a hydrochloric acid solution.

What is Given?

VHCl(aq) = 100 mL

Assume that density of HCl(aq) = 1.00 g/mL

mHCl(aq) = 100 g

mMg(s) = 0.50 g

$t_i = 20.4\ ^oC$

$t_f = 40.7\ ^oC$

specific heat capacity of HCl(aq) = cH$_2$O(l) = 4.19 J/g °C

Plan Your Strategy

Determine the temperature change, Δt, of the surroundings (HCl(aq)), and calculate $Q_{surroundings}$ using the equation $Q_{surroundings} = mc\Delta t$.

$\Delta H = -Q_{surroundings}$

Calulate the amount of HCl(aq) in mol, and determine the limiting reactant.

Calculate the amount of Mg(s) in mol that reacted and determine $\Delta_r H$ per mole of Mg(s).

Act on Your Strategy

a) $\Delta t = t_f - t_i = 40.7\ ^oC - 20.4\ ^oC = 20.3\ ^oC$

$$\begin{aligned}\Delta H &= -mc\Delta t \\ &= -(100\ g \times 4.19\ J/g\ ^oC \times 20.3\ ^oC) \\ &= -8.506 \times 10^3\ J\end{aligned}$$

MMg(s) = 24.31 g/mol

$$nMg(s) = \frac{m}{M} = \frac{0.50 g}{24.31 \frac{g}{mol}} = 0.0206\ mol$$

nHCl(aq) = 0.100 L × 1.00 mol/L = 0.100 mol

Mg(s) is the limiting reactant.

$$\Delta_r H \text{ per mole of Mg(s)} = \frac{-8.506\times10^3\ \text{J}}{0.0206\ \text{mol}}$$

$$= -4.1\times10^5\ \text{J/mol Mg(s)}$$

$$= -4.1\times10^2\ \text{kJ/mol Mg(s)}$$

b) Assumptions made in **a)**:
(i) HCl(aq) has a density of 1.00 g/mL.
(ii) HCl(aq) has $c = 4.19$ J/g °C.
(iii) All of the Mg(s) reacted.
(iv) No heat lost to the surroundings.
(v) No error in calibration of thermometer or in reading the thermometer.
(vi) Mg(s) is pure.
(vii) Reaction proceeds to completion.

Check Your Solution

The enthalpy of reaction is negative because heat is given off in this reaction. The value calculated for $\Delta_r H$ has the correct unit (kJ/mol) and the correct number of significant digits (2).

9.

Problem

Sodium reacts violently to form sodium hydroxide when placed in water, as shown in the equation below:

$$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$$

Determine an experimental value for the molar enthalpy of reaction for sodium given the following data:
mass of sodium, Na(s): 0.37 g
mass of water in calorimeter: 175 g
initial temperature of water: 19.30 °C
final temperature of mixture: 25.70 °C

What is Required?

You must calculate the enthalpy of reaction, $\Delta_r H$, for the reaction that occurs when sodium is added to water.

What is Given?

$m\text{H}_2\text{O}(l) = 175$ g
$m\text{Na(s)} = 0.37$ g
$t_i = 19.3$ °C
$t_f = 25.7$ °C
specific heat capacity, c, of $H_2O(l) = 4.19$ J/g °C

Plan Your Strategy

Determine the temperature change, Δt, of the surroundings (water), and calculate $Q_{\text{surroundings}}$ using the equation $Q_{\text{surroundings}} = mc\Delta t$.
$\Delta H = -Q_{\text{surroundings}}$

Calculate the number of moles of Na(s) and determine $\Delta_r H$ per mole of Na(s).

Act on Your Strategy

$\Delta t = t_f - t_i = 25.7\ ^oC - 19.3\ ^oC = 6.4\ ^oC$

$$\begin{aligned}\Delta H &= -mc\Delta t \\ &= -(175\text{ g} \times 4.19\text{ J/g }^oC \times 6.4\ ^oC) \\ &= -4.69 \times 10^3\text{ J}\end{aligned}$$

MNa(s) = 22.99 g/mol

$$n\text{Na(s)} = \frac{m}{M} = \frac{0.37\text{g}}{22.99\text{g/mol}} = 0.0161\text{ mol}$$

$$\begin{aligned}\Delta_r H \text{ per mole of Na(s)} &= \frac{-4.69\times10^3\text{ J}}{0.0161\text{ mol}} \\ &= -2.9 \times 10^5\text{ J/mol Na(s)} \\ &= -2.9 \times 10^2\text{ kJ/mol Na(s)}\end{aligned}$$

Check Your Solution

The enthalpy of reaction is negative because heat is given off in this reaction. The value calculated for $\Delta_r H$ has the correct unit (kJ/mol) and the correct number of significant digits (2).

10.

Problem

Nitric acid is neutralized with potassium hydroxide in the following reaction:

$$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l) \quad \Delta_r H = -53.4\text{ kJ/mol } HNO_3(aq)$$

55.0 mL of 1.30 mol/L solutions of both reactants, at 21.40 °C, are mixed in a calorimeter. What is the final temperature of the mixture? Assume that the density of both solutions is 1.00 g/mL. Also assume that the specific heat capacity of both solutions is the same as the specific heat capacity of water. No heat is lost to the calorimeter itself.

What is Required?

You must calculate the final temperature after a reaction occurs between HNO_3(aq) and KOH(aq).

What is Given?

The mole ratio between the reactants is known from the balanced equation.

$V HNO_3(aq) = V KOH(aq) = 55.0\text{ mL} = 0.0550\text{ L}$

$c HNO_3(aq) = c KOH(aq) = 1.30\text{ mol/L}$

$V(\text{total}) = 55.0\text{ mL} + 55.0\text{ mL} = 110.0\text{ mL}$

Since density of solutions = 1.00 g/mL, m (total) = 110.0 g

$t_i = 21.4\ ^oC$

specific heat capacity, $cH_2O(l) = 4.19\text{ J/g }^oC$

$\Delta_r H = -53.4\text{ kJ/mol}$

Plan Your Strategy

Since the mol ratio between reactants is 1:1 and both solutions have equal volumes and equal concentrations, neither reactant is in excess. $\Delta_r H = -53.4$ kJ/mol per either reactant. Calculate the number of moles of HNO_3(aq) using the equation $n = c \times V$.

Calculate the enthalpy change, ΔH, for this amount of HNO_3(aq)

$\Delta H = -Q_{\text{surroundings}} = -m \times c \times \Delta t$, where m = total mass of solution.

Solve this equation for Δt and calculate the final temperature using $\Delta t = t_f - t_i$

Act on Your Strategy

amount of HNO_3(aq) = 1.30 mol/L × 0.0550 L = 0.0715 mol

$\Delta_r H = -53.4 \text{ kJ/mol} = -5.34 \times 10^4 \text{ J/mol}$

heat given off when 0.0715 HNO_3 reacts, $(\Delta_r H) = 0.0715 \text{ mol} \times (-5.34 \times 10^4 \text{ J/mol})$

$= -3.818 \times 10^3 \text{ J}$

$\Delta_r H = -m \times c \times \Delta t$

$-3.818 \times 10^3 \text{ J} = -110.0 \text{ g} \times 4.19 \text{ J/g }^\circ\text{C} \times \Delta t$

$\Delta t = 8.3\ ^\circ\text{C}$

$\Delta t = t_f - t_i$

$8.3\ ^\circ\text{C} = t_f - 21.4\ ^\circ\text{C}$

$t_f = 29.7\ ^\circ\text{C}$

Check Your Solution

The final temperature is greater than the initial temperature, which is expected when heat is given off. The final temperature has the correct unit (°C) and the correct number of significant digits (3).

11.

Problem

A student uses a simple calorimeter to determine the enthalpy of reaction for hydrobromic acid and aqueous potassium hydroxide. The student mixes 100.0 mL of 0.50 mol/L HBr(aq) at 21.00 °C with 100.0 mL of 0.50 mol/L KOH(aq), also at 21.00 °C. The highest temperature that is reached is 24.40 °C. Write a thermochemical equation for the reaction.

What is Required?

You must calculate the enthalpy of reaction in kJ/ mol of reactant and use this value to write the thermochemical equation for the reaction between HBr(aq) and KOH(aq).

What is Given?

The mole ratio between the reactants is known from the balanced equation.

VHBr(aq) = VKOH(aq) = 50.0 mL = 0.0500 L

Vtotal = 50.0 mL + 50.0 mL =100.0 mL

Since density of solutions = 1.00 g/mL, m (total) = 100.0 g

concentration, cHBr(aq) = cKOH(aq) = 0.50 mol/L

$t_i = 21.0\ ^\circ\text{C}$

$t_f = 24.4\ ^\circ\text{C}$

specific heat capacity, $cH_2O(l) = 4.19 \text{ J/g }^\circ\text{C}$

Plan Your Strategy

Since the mol ratio between reactants is 1:1 and both solutions have equal volumes and equal concentrations, neither reactant is in excess. $\Delta_r H$ can be calculated per mole of either reactant. Determine the temperature change, Δt, of the surroundings (solutions), and calculate $Q_{\text{surroundings}}$ using the equation $Q_{\text{surroundings}} = mc\Delta t$, where m = total mass of solution.

$\Delta H = -Q_{\text{surroundings}}$

Calculate the amount of HBr(aq) using the equation $n = c \times V$ and determine $\Delta_r H$ per mole of HBr(aq).

Act on Your Strategy

$\Delta t = t_f - t_i = 24.4\ ^{\circ}\text{C} - 21.0\ ^{\circ}\text{C} = 3.4\ ^{\circ}\text{C}$

$\Delta H = -mc\Delta t$

$= -100.0\ \text{g} \times 4.19\ \text{J/g}\ ^{\circ}\text{C} \times 3.4\ ^{\circ}\text{C}$

$= -1.42 \times 10^3\ \text{J}$

amount HBr(aq) = 0.50 mol/L × 0.0500 L = 0.0250 mol

$$\Delta_r H \text{ per mole of HBr(aq)} = \frac{-1.42\times10^3\ \text{J}}{0.0250\ \text{mol}}$$

$= -5.7 \times 10^4$ J/mol per HBr(aq)

$= -57$ kJ/mol per HBr(aq)

The thermochemical equation is:

$\text{HBr(aq)} + \text{KOH(aq)} \rightarrow \text{KBr(aq)} + \text{H}_2\text{O}(\ell) \qquad \Delta_r H = -57\ \text{kJ/mol}$

Check Your Solution

The enthalpy of reaction is negative since heat is given off in this reaction. The value calculated for $\Delta_r H$ has the correct unit (kJ/mol) and the correct number of significant digits (2).

12.

Problem

In a simple calorimeter, 150 mL of 1.000 mol/L NaOH(aq) is mixed with 150 mL of 1.000 mol/L HCl(aq). If both solutions were initially at 25.00 ºC and after mixing the temperature rose to 30.00 ºC, what is the enthalpy change of the reaction as written?

What is Required?

You must calculate the enthalpy of reaction, $\Delta_r H$, for the reaction between HCl(aq) and NaOH(aq)

What is Given?

The mole ratio between the reactants is known from the balanced equation.

$v(\text{HCl(aq)}) = v(\text{NaOH(aq)}) = 150\ \text{mL} = 0.150\ \text{L}$

$v(\text{total}) = 0.150\ \text{L} + 0.150\ \text{L} = 0.300\ \text{L}$

Since density of solutions = 1.00 g/mL, $m(\text{total}) = 300$ g

concentration, $c(\text{HCl(aq)}) = c(\text{NaOH(aq)}) = 1.000$ mol/L

$t_i = 25\ ^{\circ}\text{C}$

$t_f = 27\ ^oC$
specific heat capacity, $c(\text{solutions}) = c(H_2O(l)) = 4.19$ J/g ^oC

Plan Your Strategy
Since the mole ratio between reactants is 1:1 and both solutions have equal volumes and equal concentrations, neither reactant is in excess. Δ_rH can be calculated per mole of either reactant. Determine the temperature change, Δt, of the surroundings (solutions), and calculate $Q_{\text{surroundings}}$ using the equation $Q_{\text{surroundings}} = mc\Delta t$, where m = total mass of solution.
$\Delta H = -Q_{\text{surroundings}}$
Calculate the number of moles of HCl(aq) using the equation $n = c \times v$ and determine Δ_rH per mole of HCl(aq).

Act on Your Strategy
$\Delta t = t_f - t_i = 27\ ^oC - 25\ ^oC = 2\ ^oC$
$\Delta H = -mc\Delta t$
$= -300\text{ g} \times 4.19\text{ J/g }^oC \times 2\ ^oC$
$= -2.514 \times 10^3\text{ J}$
moles HCl(aq) = 1.000 mol/L × 0.150 L = 0.150 mol

$$\Delta_rH \text{ per mole of HCl(aq)} = \frac{-2.514 \times 10^3\text{ J}}{0.150\text{ mol}}$$
$$= -16.76 \times 10^5\text{ J/mol HCl(aq)}$$
$$= -1.676 \times 10^2\text{ kJ/mol HCl(aq)}$$

Check Your Solution
The enthalpy of reaction is negative because heat is released in this reaction. The value calculated for Δ_rH has the correct unit (kJ/mol) and the correct number of significant digits (2). The same values apply if Δ_rH per mole of NaOH is calculated.

13.
Problem
Predict the final temperature of a 500 g iron ring that is initially at 25.0 °C and is heated by combusting 4.95 g of ethanol, $\Delta_cH = -1234.8$ kJ/mol $C_2H_5OH(l)$ in an open system.

What is Required?
You must calculate the final temperature of an iron ring after it has been heated by burning a sample of alcohol.

What is Given?
$t_i = 25.0\ ^oC$
$\Delta_cH\ (C_2H_5OH(l)) = -1234.8$ kJ/mol
specific heat capacity, $c(\text{iron}) = 0.449$ J/g ^oC
mass Fe(s) = 500 g
mass $(C_2H_5OH(l)) = 4.95$ g

Plan Your Strategy

Calculate the amount of $C_2H_5OH(\ell)$ and use this information together with $\Delta_cH\,(C_2H_5OH(\ell))$ to calculate the heat given off when the alcohol burns.

$\Delta H = -Q_{\text{surroundings}} = -mc\Delta t$, where m = mass of Fe(s).

Solve this equation for Δt and calculate the final temperature using $\Delta t = t_f - t_i$

Act on Your Strategy

$M(C_2H_5OH(\ell)) = 46.08$ g/mol

amount of $C_2H_5OH(\ell) = \dfrac{m}{M} = \dfrac{4.95\text{g}}{46.08\text{g/mol}} = 0.1074$ mol

heat given off, $\Delta H = 0.1074 \text{ mol } C_2H_5OH(\ell) \times (-1234.8 \text{ kJ/mol})$

$= -132.6$ kJ

$= -1.326 \times 10^5$ J

$\Delta H = -Q_{\text{surroundings}} = -mc\Delta t$

$-1.326 \times 10^5 \text{ J} = -500 \text{ g} \times 0.449 \text{ J/g}^{o}\text{C} \times \Delta t$

$\Delta t = 591\ ^{o}\text{C}$

$\Delta t = t_f - t_i$

$591\ ^{o}\text{C} = t_f - 25.0\ ^{o}\text{C}$

$t_f = 616\ ^{o}\text{C}$

Check Your Solution

The final temperature is greater than the initial temperature, which is expected after heating the iron. The final temperature has the correct unit (^{o}C) and the correct number of significant digits (3). You would expect the actual final temperature to be less than this since heat will be lost to the surroundings as the alcohol burns.

14.

Problem

Calculate the molar enthalpy of combustion of octane if 0.53 g of the fuel increased the temperature of a coffee can calorimeter (13 g of aluminium and 250 mL of water) by 17.2 °C. Remember to include the heat gained by not only the water but also by the aluminium can.

What is Required?

You must determine the heat of combustion of octane.

What is Given?

mass of octane = 0.53 g

$\Delta t = 17.2\ ^{o}\text{C}$

mass of aluminium calorimeter = 13 g

mass of water = 250 g (density of water = 1.0 g/mL)

specific heat capacity, $c(\text{Al(s)}) = 0.900$ J/g °C
specific heat capacity, $c(H_2O(\ell)) = 4.19$ J/g °C

Plan Your Strategy

$Q_{\text{surroundings}} = Q_{\text{water}} + Q_{\text{cal}}$ (where $Q = mc\Delta t$)
$Q_{\text{system}} = -Q_{\text{surroundings}}$
Calculate the number of moles of octane.

$$\text{enthalpy of combustion of octane} = \frac{Q_{\text{system}}}{\text{mol of fuel}}$$

Act on Your Strategy

$$\begin{aligned} Q_{\text{surroundings}} &= [(mc\Delta t)_{\text{water}} + (mc\Delta t)_{\text{cal}}] \\ &= [(250\text{ g} \times 4.19\text{ J/g }^{\circ}\text{C} \times 17.2\,^{\circ}\text{C}) + (13\text{ g} \times 0.900\text{ J/g }^{\circ}\text{C} \times 17.2\,^{\circ}\text{C})] \\ &= 18\,017\text{ J} + 201.2\text{ J} \\ &= 1.8218 \times 10^4\text{ J} \\ &= 18.22\text{ kJ} \end{aligned}$$

$\therefore Q_{\text{system}} = -18.22$ kJ
$M(\text{octane}) = 114.26$ g/mol

$$\text{moles of octane} = \frac{m}{M} = \frac{0.53\text{ g}}{114.26\text{ g/mol}} = 0.004639\text{ mol}$$

$$\text{heat of combustion of octane} = \frac{Q_{\text{system}}}{\text{moles of octane}} = \frac{-18.218\text{kJ}}{0.04639\text{ mol}} = -3.9 \times 10^3\text{ kJ/mol } C_8H_{18}(l)$$

Check Your Solution

The calculated value for the heat of combustion of octane has the correct unit (kJ/mol) and the correct number of significant digits for this data (2)

15.

Problem

How much propane (in grams) would have to be burned in an open system to raise the temperature of 300 mL of water from 20.0 °C to its boiling point? (The molar enthalpy of combustion of propane may be found in Table 9.1 on page 347).

What is Required?

You must calculate the mass of propane that will raise the temperature of a given mass of water to its boiling point.

What is Given?

$VH_2O(l) = 300$ mL
$mH_2O(l) = 300$ g (density of water = 1.00 g/mL)
$t_i = 20.0\,^{\circ}\text{C}$

$t_f = 100.0\ ^oC$
specific heat capacity, $cH_2O(l) = 4.19$ J/g °C
$\Delta_c H\ C_3H_8(g) = –2043.9$ kJ/mol

Plan Your Strategy

Determine the temperature change, Δt, of the surroundings (water), and calculate $Q_{surroundings}$ using the equation $Q_{surroundings} = mc\Delta t$
Since $\Delta_c H = –Q_{surroundings}$, calculate the number of moles of propane that reacted.
Calculate the mass of $C_3H_8(g)$, $m = n \times M$

Act on Your Strategy

$\Delta t = t_f - t_i = 100.0\ ^oC – 20.0\ ^oC = 80.0\ ^oC$
$Q_{surroundings} = mc\Delta t = 300\text{ g} \times 4.19\text{ J/g }^oC \times 80.0\ ^oC = 1.006 \times 10^5\text{ J} = 1.006 \times 10^2\text{ kJ}$

$$\text{moles of } C_3H_8(g) = \frac{-1.006\times10^2\text{ kJ}}{-2043.9\text{ kJ/mol}} = 0.04920\text{ mol}$$

mass of $C_3H_8(g) = 0.04920\text{ mol} \times 44.11\text{ g/mol} = 2.17\text{ g}$

Check Your Solution

The calculate mass of propane has the correct unit (g) and the correct number of significant digits (3)

16.

Problem

A lab technician places a 5.00 g food sample into a bomb calorimeter that is calibrated at 9.23 kJ/°C. The initial temperature of the calorimeter system is 21.0 °C. After burning the food, the final temperature of the system is 32.0 °C. What is the enthalpy of combustion of the food in kJ/g?

What is Required?

You must determine the enthalpy of combustion of a food sample in kJ/g.

What is Given?

heat capacity of bomb calorimeter = $C = 9.23$ kJ/°C
$t_i = 21.0\ ^oC$
$t_f = 32.0\ ^oC$
mass of food sample, m, = 5.00 g

Plan Your Strategy

Determine the temperature change, Δt, of the calorimeter system.
Calculate the heat absorbed by the calorimeter using $Q_{cal} = C\Delta t$
Thermal energy released when food burns = thermal energy absorbed by calorimeter
$Q_{cal} = C\Delta t$
$Q_{sample} = –Qcal$

heat of combustion of food sample = $\dfrac{Q_{\text{sample}}}{m}$

Act on Your Strategy

$\Delta t = t_f - t_i = 32.0\ ^{o}\text{C} - 21.0\ ^{o}\text{C} = 11.0\ ^{o}\text{C}$
$Q_{\text{cal}} = C\Delta t = 9.23\ \text{kJ}/^{o}\text{C} \times 11.0\ ^{o}\text{C} = 101.5\ \text{kJ}$
$Q_{\text{sample}} = -Q_{\text{cal}} = -101.5\ \text{kJ}$

enthalpy of combustion of food sample = $\dfrac{Q_{\text{sample}}}{m} = \dfrac{-101.5\ \text{kJ}}{5.00\ \text{g}} = -20.3\ \text{kJ/g}$

Check Your Solution

The value for the heat of combustion has the correct unit (kJ/g) and the correct number of significant digits (3)

17.

Problem

Determine the enthalpy of combustion of an unknown fuel if a 2.75 g sample increased the temperature of 500 mL of hot chocolate (c = 3.75 J/g • °C) in a 150 g glass mug (c = 0.84 J/g • °C) from 10 °C to 45 °C. Express the value for enthalpy of combustion in appropriate units.

What is Required?

You must determine the enthalpy of combustion of a fuel.

What is Given?

mass of fuel = 2.75 g
$t_i = 10\ ^{o}\text{C}$
$t_f = 45\ ^{o}\text{C}$
mass of mug = 150 g
volume of hot chocolate = 500 mL
mass of hot chocolate = 500 g (assume density of hot chocolate = 1.0 g/mL)
specific heat capacity, c (glass) = 0.84 J/g ^{o}C
specific heat capacity, c (hot chocolate) = 3.75 J/g ^{o}C

Plan Your Strategy

$Q_{\text{surroundings}} = Q_{\text{water}} + Q_{\text{mug}}$ (where $Q = mc\Delta t$)
$Q_{\text{system}} = -Q_{\text{surroundings}}$

heat of combustion of fuel = $\dfrac{Q_{\text{system}}}{\text{mass of fuel}}$

Act on Your Strategy

$\Delta t = t_f - t_i = 45\ ^{o}\text{C} - 10\ ^{o}\text{C} = 35\ ^{o}\text{C}$

$$\begin{aligned} Q_{\text{surroundings}} &= [(mc\Delta t)_{\text{hot chocolate}} + (mc\Delta t)_{\text{mug}}] \\ &= [(500\text{ g} \times 3.75\text{ J/g }^{\circ}\text{C} \times 35\ ^{\circ}\text{C}) + (150\text{ g} \times 0.84\text{ J/g }^{\circ}\text{C} \times 35\ ^{\circ}\text{C})] \\ &= 65625\text{ J} + 4410\text{ J} \\ &= 7.00 \times 10^{4}\text{ J} = 70.0\text{ kJ} \end{aligned}$$

$\therefore Q_{\text{system}} = -70.0\text{ kJ}$

$$\text{heat of combustion of octane} = \frac{Q_{\text{system}}}{\text{mass of fuel}} = \frac{-70.0\text{ kJ}}{2.75\text{ g}} = -25\text{ kJ/g of fuel}$$

Check Your Solution

The calculated value for the heat of combustion of the fuel has the correct unit (kJ/mol) and the correct number of significant digits for this data (2).