Unit 5 Thermochemical Changes
Chapter 10 Theories of Energy and Chemical Changes
Solutions to Practice Problems

1.
Problem
Ethene, $C_2H_4(g)$, reacts with water to form ethanol, $C_2H_5OH(\ell)$, as shown below:

$$C_2H_4(g) + H_2O(\ell) \rightarrow C_2H_5OH(\ell)$$

Determine the enthalpy change of this reaction, given the following thermochemical equations.

(1) $C_2H_5OH(\ell) + 3O_2(g) \rightarrow 3H_2O(\ell) + 2CO_2(g)$ $\Delta_cH^o = -1366.8$ kJ/mol
(2) $C_2H_4(g) + 3O_2(g) \rightarrow 2H_2O(\ell) + 2CO_2(g)$ $\Delta_cH^o = -1411.0$ kJ/mol

What is Required?
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

What is Given?
The thermochemical equations:
(1) $C_2H_5OH(\ell) + 3O_2(g) \rightarrow 3H_2O(\ell) + 2CO_2(g)$ $\Delta_cH^o = -1366.8$ kJ/mol
(2) $C_2H_4(g) + 3O_2(g) \rightarrow 2H_2O(\ell) + 2CO_2(g)$ $\Delta_cH^o = -1411.0$ kJ/mol

Plan Your Strategy
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change, Δ_rH^o, is the algebraic sum of the enthalpy changes for the reactions used.

Act on Your Strategy
(1) × −1 $3H_2O(\ell) + 2CO_2(g) \rightarrow C_2H_5OH(\ell) + 3O_2(g)$ $\Delta_cH^o = +1366.8$ kJ/mol
(2) $C_2H_4(g) + 3O_2(g) \rightarrow 2H_2O(\ell) + 2CO_2(g)$ $\Delta_cH^o = -1411.0$ kJ/mol

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$C_2H_4(g) + H_2O(\ell) \rightarrow C_2H_5OH(\ell)$ $\Delta_rH^o, = -44.2$ kJ/mol

Check Your Solution
The answer is reasonable for this reaction. The unit is correct (kJ/mol of ethanol) and the answer has the correct number of significant digits after the decimal (1).

2.
Problem
A typical automobile uses a lead-acid battery. During discharge, the following chemical reaction takes place:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(\ell) \rightarrow 2PbSO_4(aq) + 2H_2O(\ell)$$

Determine the enthalpy change of this reaction, given the following equations:
(1) $Pb(s) + PbO_2(s) + 2SO_3(g) \rightarrow PbSO_4(aq)$ $\Delta H^o = -773$ kJ
(2) $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(\ell)$ $\Delta H^o = -133$ kJ

What is Required?
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

What is Given?
The thermochemical equations:

(1) $Pb(s) + PbO_2(s) + 2SO_3(g) \rightarrow PbSO_4(aq)$ $\Delta H^\circ = -773$ kJ

(2) $SO_3(g) + H_2O(\ell) \rightarrow H_2SO_4(\ell)$ $\Delta H^\circ = -133$ kJ

Plan Your Strategy
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change, $\Delta_r H^\circ$, is the algebraic sum of the enthalpy changes for the reactions used.

Act on Your Strategy

(1) $Pb(s) + PbO_2(s) + 2SO_3(g) \rightarrow PbSO_4(aq)$ $\Delta H^\circ = -773$ kJ

(2) $\times -2$ $2H_2SO_4(\ell) \rightarrow 2SO_3(g) + 2H_2O(\ell)$ $\Delta H^\circ = +266$ kJ

$Pb(s) + PbO_2(s) + 2H_2SO_4(\ell) \rightarrow 2PbSO_4(aq) + 2H_2O(\ell)$ $\Delta_r H^\circ = -507$ kJ

Check Your Solution
The reaction has a negative enthalpy change and therefore it is an exothermic reaction, as expected. The unit is correct (kJ) and the answer has the correct number of digits after the decimal (0).

3.
Problem
Mixing household cleansers can result in the production of hydrogen chloride gas, HCl(g). Not only is this gas toxic and corrosive, but it also reacts with oxygen to form poisonous chlorine gas according to the following equation:

$$4HCl(g) + O_2(g) \rightarrow 2Cl_2(g) + 2H_2O(g)$$

Determine the enthalpy change of this reaction, given the following equations:

(1) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ $\Delta H^\circ = -184.6$ kJ

(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta_f H^\circ = -285.8$ kJ/mol

(3) $H_2O(\ell) \rightarrow H_2O(g)$ $\Delta_{vap} H^\circ = 40.7$ kJ/mol

What is Required?
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

What is Given?
The thermochemical equations:

(1) $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ $\Delta H^\circ = -184.6$ kJ

(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta_f H° = –285.8$ kJ/mol

(3) $H_2O(\ell) \rightarrow H_2O(g)$ $\Delta_{vap}H° = 40.7$ kJ/mol

Plan Your Strategy
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change, $\Delta_r H°$, is the algebraic sum of the enthalpy changes for the reactions used.

Act on Your Strategy

(1) × –2 $4HCl(g) \rightarrow 2H_2(g) + 2Cl_2(g)$ $\Delta H° = +369.2$ kJ

(2) × 2 $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$ $\Delta H° = –571.6$ kJ

(3) × 2 $2H_2O(\ell) \rightarrow 2H_2O(g)$ $\Delta H° = +81.4$ kJ

$4HCl(g) + O_2(g) \rightarrow 2Cl_2(g) + 2H_2O(g)$ $\Delta_r H° = –121.0$ kJ

Check Your Solution
The reaction has a negative enthalpy change and it is exothermic, as expected. The unit is correct (kJ) and the answer has the correct number of digits after the decimal (1).

4.
Problem
Calculate the enthalpy change of the following reaction between nitrogen gas and oxygen gas, given thermochemical equations (1), (2), and (3):

$$2N_2(g) + 5O_2(g) \rightarrow 2N_2O_5(g)$$

(1) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$ $\Delta H° = –572$ kJ

(2) $N_2O_5(g) + H_2O(\ell) \rightarrow 2HNO_3(aq)$ $\Delta H° = –77$ kJ

(3) $\frac{1}{2}N_2(g) + \frac{3}{2}O_2(g) + \frac{1}{2}H_2(g) \rightarrow HNO_3(\ell)$ $\Delta_f H° = –174$ kJ/mol

What is Required?
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

What is Given?
The thermochemical equations:

(1) $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$ $\Delta H° = –572$ kJ

(2) $N_2O_5(g) + H_2O(\ell) \rightarrow 2HNO_3(aq)$ $\Delta H° = –77$ kJ

(3) $\frac{1}{2}N_2(g) + \frac{3}{2}O_2(g) + \frac{1}{2}H_2(g) \rightarrow HNO_3(\ell)$ $\Delta_f H° = –174$ kJ/mol

Plan Your Strategy
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change, $\Delta_r H°$, is the algebraic sum of the enthalpy changes for the reactions used.

Act on Your Strategy

(1) × −1 $2H_2O(\ell) \rightarrow 2H_2(g) + O_2(g)$ $\Delta H° = +572$ kJ

(2) × −2 $4HNO_3(aq) \rightarrow 2N_2O_5(g) + 2H_2O(\ell)$ $\Delta H° = +154$ kJ

(3) × 4 $2N_2(g) + 6O_2(g) + 2H_2(g) \rightarrow 4HNO_3(\ell)$ $\Delta H° = -696$ kJ

$2N_2(g) + 5O_2(g) \rightarrow 2N_2O_5(g)$ $\Delta_rH° = 30$ kJ

Check Your Solution

The reaction has a positive enthalpy change and it is endothermic, as expected. The unit is correct (kJ) and the answer has the correct number of digits after the decimal (0).

5.

Problem

Ethene, $C_2H_4(g)$, is used in the manufacturing of many polymers, including polyethylene terephthalate (PET), which is used to make pop bottles. Determine the molar enthalpy of formation for ethene, as shown by $2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$, given the following equations:

(1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H° = -393.5$ kJ

(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta H° = -285.8$ kJ

(3) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$ $\Delta H° = -1411.2$ kJ

What is Required?

You must manipulate the given equations to calculate the enthalpy of formation for ethene, $C_2H_4(g)$.

What is Given?

The thermochemical equations:

(1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H° = -393.5$ kJ

(2) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta H° = -285.8$ kJ

(3) $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$ $\Delta H° = -1411.2$ kJ

Plan Your Strategy

Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change is the algebraic sum of the enthalpy changes for the reactions used and represents the enthalpy of formation, $\Delta_fH°$.

Act on Your Strategy

(1) × 2 $2C(s) + 2O_2(g) \rightarrow 2CO_2(g)$ $\Delta H° = -787.0$ kJ

(2) × 2 $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$ $\Delta H° = -571.6$ kJ

(3) × −1 $2CO_2(g) + 2H_2O(\ell) \rightarrow C_2H_4(g) + 3O_2(g)$ $\Delta H° = +1411.2$ kJ

$2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$ $\Delta_fH° = +52.6$ kJ/mol

Check Your Solution
The unit is correct (kJ/mol) and the answer has the correct number of digits after the decimal (1).

6.
Problem
From the following equations, determine the molar enthalpy of formation for $HNO_2(aq)$, as shown below in the overall equation:

$$\frac{1}{2}H_2(g) + \frac{1}{2}N_2(g) + O_2(g) \rightarrow HNO_2(aq)$$

(1) $NH_4NO_2(aq) \rightarrow N_2(g) + 2H_2O(\ell)$ $\Delta H° = –320.1$ kJ
(2) $NH_3(aq) + HNO_2(aq) \rightarrow NH_4NO_2(aq)$ $\Delta H° = –37.7$ kJ
(3) $2NH_3(aq) \rightarrow N_2(g) + 3H_2(g)$ $\Delta H° = +169.9$ kJ
(4) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta H° = –285.8$ kJ

What is Required?
You must manipulate the given equations to calculate the enthalpy of formation for $HNO_2(aq)$.

What is Given?
The thermochemical equations:
(1) $NH_4NO_2(aq) \rightarrow N_2(g) + 2H_2O(\ell)$ $\Delta H° = –320.1$ kJ
(2) $NH_3(aq) + HNO_2(aq) \rightarrow NH_4NO_2(aq)$ $\Delta H° = –37.7$ kJ
(3) $2NH_3(aq) \rightarrow N_2(g) + 3H_2(g)$ $\Delta H° = +169.9$ kJ
(4) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$ $\Delta H° = –285.8$ kJ

Plan Your Strategy
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change is the algebraic sum of the enthalpy changes for the reactions used and represents the enthalpy of formation, $\Delta_f H^o$ for $HNO_2(aq)$.

Act on Your Strategy
(1) × −1 $N_2(g) + 2H_2O(\ell) \rightarrow NH_4NO_2(aq)$ $\Delta H° = +320.1$ kJ
(2) × −1 $NH_4NO_2(aq) \rightarrow NH_3(aq) + HNO_2(aq)$ $\Delta H° = +37.7$ kJ
(3) × 1/2 $NH_3(aq) \rightarrow \frac{1}{2}N_2(g) + \frac{3}{2}H_2(g)$ $\Delta H° = +85.0$ kJ
(4) × 2 $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$ $\Delta H° = –571.6$ kJ

$\frac{1}{2}H_2(g) + \frac{1}{2}N_2(g) + O_2(g) \rightarrow HNO_2(aq)$ $\Delta_f H^o = −128.8$ kJ/mol

Check Your Solution
The unit is correct (kJ/mol) and the answer has the correct number of digits after the decimal (1).

7.
Problem
Hydrogen can be added to ethene, $C_2H_4(g)$, to obtain ethane, $C_2H_6(g)$:

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

Write thermochemical equations for the formation of both ethene, $C_2H_4(g)$, and ethane, $C_2H_6(g)$. Show that these equations can be algebraically combined to obtain the equation for the addition of hydrogen to ethene. Determine the enthalpy of reaction.

What is Required?
You must write the thermochemical equations representing the formation of ethene, $C_2H_4(g)$, and ethane, $C_2H_6(g)$ and demonstrate that these equations can be combined algebraically to give the overall reaction between ethene and hydrogen to produce ethane. From this answer calculate the enthalpy of reaction.

What is Given?
The balanced equation for the reaction between ethene and hydrogen to produce ethane is given. The heats of formation of ethene, $C_2H_4(g)$, and ethane, $C_2H_6(g)$ can be found in the Chemistry Data Booklet..

Plan Your Strategy
Write the equation for the formation of ethene, $C_2H_4(g)$, and ethane, $C_2H_6(g)$. Apply Hess's Law to obtain the overall reaction between ethene, $C_2H_4(g)$, and ethane, $C_2H_6(g)$.
The overall enthalpy change, Δ_rH^o is the algebraic sum of the enthalpy changes for the reactions used.

Act on Your Strategy
(1) formation of $C_2H_4(g)$: $2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$ $\Delta_fH^o = +52.4$ kJ/mol
(2) formation of $C_2H_6(g)$: $2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$ $\Delta_fH^o = -84.0$ kJ/mol

(1) $\times -1$	$C_2H_4(g) \rightarrow 2C(s) + 2H_2(g)$	$\Delta_fH^o = -52.4$ kJ/mol
(2)	$2C(s) + 3H_2(g) \rightarrow C_2H_6(g)$	$\Delta_fH^o = -84.0$ kJ/mol
	$C_2H_4(g) + 3H_2(g) \rightarrow C_2H_6(g)$	$\Delta_rH^o = -136.4$ kJ

Check Your Solution
The enthalpy of reaction is reasonable. The answer has the correct unit (kJ) and the correct number of digits after the decimal (1)

8.
Problem
Zinc sulfide reacts with oxygen gas to produce zinc oxide and sulfur dioxide:

$$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$$

Calculate the enthalpy change of this reaction by using enthalpies of formation from Appendix G or your Chemistry Data Booklet.

What is Required?
You must calculate the enthalpy change for the given reaction using standard enthalpies of formation of the reactants and products.

What is Given?
The balanced equation for the reaction is given and the standard enthalpies of formation are in Appendix J.

Plan Your Strategy
Look up the standard enthalpies of formation, $\Delta_f H^o$, for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression $\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$

Act on Your Strategy
$2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$
$\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$
$\Delta_r H^o = [(2 \text{ mol}) (\Delta_f H^o\ ZnO(s)) + (2 \text{ mol})(\Delta_f H^o\ SO_2(g))] - [(2 \text{ mol}) (\Delta_f H^o\ ZnS(s)) +$
$(3 \text{ mol}) (\Delta_f H^o\ O_2(g))]$
$\Delta_r H^o = [(2 \text{ mol}) (-350.5 \text{ kJ/mol}) + (2 \text{ mol}) (-296.8 \text{ kJ/mol})] - [(2 \text{ mol}) (-206.0 \text{ kJ/mol}) + (3 \text{ mol}) (0)]$
$\Delta_r H^o = [-701.0 \text{ kJ} + (-593.6)] - [-412.0 \text{ kJ}]$
$\Delta_r H^o = -882.6 \text{ kJ}$

Check Your Solution
The answer has the correct unit (kJ) and the correct number of digits after the decimal (1).

9.
Problem
Small amounts of oxygen gas can be produced in a laboratory by heating potassium chlorate, $KClO_3(s)$:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Calculate the enthalpy change of this reaction by using enthalpies of formation.

What is Required?
You must calculate the enthalpy change for the given reaction using standard enthalpies of formation of the reactants and products.

What is Given?
The balanced equation for the reaction is given and the standard enthalpies of formation are in the Chemistry Data Booklet.

Plan Your Strategy
Look up the standard enthalpies of formation, $\Delta_f H^o$, for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression $\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$

Act on Your Strategy
$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
$\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$
$\Delta_r H^o = [(2 \text{ mol}) (\Delta_f H^o \text{ KCl(s)}) + (3 \text{ mol}) (\Delta_f H^o \text{ O}_2\text{(g)})] - [(2 \text{ mol}) (\Delta_f H^o \text{ KClO}_3\text{(s)})]$
$\Delta_r H^o = [(2 \text{ mol}) (-436.5 \text{ kJ/mol}) + (3 \text{ mol}) (0)] - [(2 \text{ mol}) (-397.7 \text{ kJ/mol})]$
$\Delta_r H^o = [-873.0 \text{ kJ}] - [-795.4 \text{ kJ}] = -77.6 \text{ kJ}$

Check Your Solution
The answer has the correct unit (kJ) and the correct number of digits after the decimal (1).

10.
Problem
Use the following equation to answer the questions below:

$$CH_3OH(\ell) + O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

a) Calculate the molar enthalpy of combustion of methanol by using enthalpies of formation.
b) Use your answer from **a)** to determine how much energy is released when 125 g of methanol undergoes complete combustion.

What is Required?
a) You must calculate the enthalpy change for the given reaction using standard enthalpies of formation of the reactants and products.
b) You must use the answer from **a)** and the balanced equation to calculate the energy released when 125 g of methanol is burned.

What is Given?
a) The balanced equation for the reaction is given and the standard enthalpies of formation are in the Chemistry Data Booklet.
b) You know that 125 g of methanol is burned.

Plan Your Strategy
a) Look up the standard enthalpies of formation, $\Delta_f H^o$, for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression $\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$
b) Determine the molar mass (M) of methanol, $CH_3OH(\ell)$, and convert 125 g to moles using $n = \frac{m}{M}$. Use the mole ratio in the balanced equation to calculate the energy released from the burning of this number of moles of methanol.

Act on Your Strategy

a) $CH_3OH(\ell) + \frac{3}{2} O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

$\Delta_r H^o = \Sigma(n\Delta_f H^o \text{ products}) - \Sigma(n\Delta_f H^o \text{ reactants})$

$\Delta H^o = [(1 \text{ mol}) (\Delta_f H^o\, CO_2(g)) + (2 \text{ mol}) (\Delta_f H^o\, H_2O(g))] - [(1 \text{ mol}) (\Delta_f H^o\, CH_3OH(\ell)) + \frac{3}{2} \Delta_f H^o$
$O_2(g))]$

$\Delta_r H^o = [(1 \text{ mol}) (-393.5 \text{ kJ/mol}) + (2 \text{ mol}) (-241.8 \text{ kJ/mol})] - [(1 \text{ mol}) (-239.2 \text{ kJ/mol}) + \frac{3}{2} \text{ mol}$
$(0)]$

$\Delta_r H^o = [-877.1 \text{ kJ}] - [-239.2 \text{ kJ}]$

$\Delta_r H^o = -637.9$ kJ per mol of $CH_3OH(\ell)$

b) $M CH_3OH(\ell) = 32.05$ g/mol

$$nCH_3OH(\ell) = \frac{m}{M} = \frac{125\,\text{g}}{32.05\,\text{g/mol}} = 3.90 \text{ mol}$$

energy released = 3.90 mol $CH_3OH(\ell) \times (-637.9 \text{ kJ/mol}) = -2.49 \times 10^3$ kJ

Check Your Solution

a) The enthalpy of reaction is negative and the reaction is exothermic as expected. The answer has the correct unit (kJ per mol of $CH_3OH(\ell)$), and the correct number of digits after the decimal (1).

b) The heat given off has the correct unit (kJ) and it has the correct number of significant digits (3).

11.

Problem

Consider the following equation representing the reaction of methane and chlorine to form chloroform, $CHCl_3(g)$:

$$CH_4(g) + 3Cl_2(g) \rightarrow CHCl_3(g) + 3HCl(g) \qquad \Delta H° = -305.0 \text{ kJ}$$

Use standard molar enthalpies of formation to determine the molar enthalpy of formation for chloroform, $CHCl_3(g)$.

What is Required?

You must calculate the enthalpy of formation, $\Delta_f H^o$, of chloroform, $CHCl_3(g)$.

What is Given?

The balanced equation for the reaction is given. The enthalpy of reaction, $\Delta_r H^o$, is known and the standard enthalpies of formation are in the Chemistry Data Booklet.

Plan Your Strategy

Look up the standard enthalpies of formation, $\Delta_f H^o$, for $CH_4(g)$ and $HCl(g)$. Use these values and the given enthalpy of reaction, $\Delta_r H^o$, in the expression

$\Delta_rH^o = \Sigma(n\Delta_fH^o \text{ products}) - \Sigma(n\Delta_fH^o \text{ reactants})$. Solve for $\Delta_fH°$ of $CHCl_3(g)$.

Act on Your Strategy

$CH_4(g) + 3Cl_2(g) \rightarrow CHCl_3(g) + 3HCl(g)$ $\Delta H° = -305.0$ kJ

$\Delta_rH^o = \Sigma\ (n\Delta_fH^o \text{ products}) - \Sigma(n\Delta_fH^o \text{ reactants})$

$\Delta_rH^o = [(1 \text{ mol})\ (\Delta_fH^o\ CHCl_3(g)) + (3 \text{ mol})\ (\Delta_fH^o\ HCl(g)] - [(1 \text{ mol})\ (\Delta_fH^o\ CH_4(g)) + (3 \text{ mol})\ \Delta_fH^o\ Cl_2(g))]$

$-305.0 \text{ kJ} = [(1 \text{ mol})\ (\Delta_fH^o\ CHCl_3(g)) + (3 \text{ mol})\ (-92.3 \text{ kJ/mol})] - [(1 \text{ mol})\ (-74.6 \text{ kJ/mol}) + (3 \text{ mol})\ (0)]$

$(1 \text{ mol})\Delta_fH^o\ CHCl_3(g) = -305.0 \text{ kJ} + 276.9 \text{ kJ} -74.6 \text{ kJ}$

$\Delta_rH^o = -102.7$ kJ/mol

Check Your Solution

This answer is reasonable, has the correct unit (kJ/mol) and the correct number of digits after the decimal (1).

12.

Problem

The molar enthalpy of combustion of heptane, $C_7H_{16}(\ell)$, in a bomb calorimeter is –4816.7 kJ/mol of heptane. Using this and $\Delta_fH°$ data, determine the molar enthalpy of formation of heptane.

What is Required?

You must calculate the enthalpy of formation, Δ_fH^o, of heptane, $C_7H_{16}(\ell)$.

What is Given?

The balanced equation for the reaction is given. The enthalpy of reaction, Δ_rH^o, is known and the standard enthalpies of formation are in the Chemistry Data Booklet.

Plan Your Strategy

Write the balanced equation for the combustion of heptane, $C_7H_{16}(\ell)$, to $CO_2(g)$ and $H_2O(\ell)$. Look up the standard enthalpies of formation, Δ_fH^o, for $CO_2(g)$ and $H_2O(\ell)$. Use these values and the given enthalpy of reaction, Δ_rH^o, in the expression

$\Delta_cH^o = \Sigma(n\Delta_fH^o \text{ products}) - \Sigma(n\Delta_fH^o \text{ reactants})$. Solve for $\Delta_fH°\ C_7H_{16}(\ell)$.

Act on Your Strategy

$C_7H_{16}(\ell) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(\ell)$ $\Delta_cH^o = -4816.7$ kJ/mol

$\Delta_cH^o = \Sigma(n\Delta_fH^o \text{ products}) - \Sigma(n\Delta_fH^o \text{ reactants})$

$\Delta_cH^o = [(7 \text{ mol})(\Delta_fH^o\ CO_2(g)) + (8 \text{ mol})(\Delta_fH^o\ H_2O(\ell)] - [(1 \text{ mol})\ (\Delta_fH^o\ C_7H_{16}(\ell)) + ((11 \text{ mol})\ \Delta_fH^o\ O_2(g))]$

$-4816.7 \text{ kJ} = [(7 \text{ mol})\ (-393.5 \text{ kJ})] + (8 \text{ mol})\ (-285.8 \text{ kJ/mol})] - [(1 \text{ mol})\ (\Delta_fH^o\ C_7H_{16}(\ell)) + (11 \text{ mol})\ (0)]$

$-4816.7 \text{ kJ} = [(-2754.5 \text{ kJ.mol}) + (-2286.4 \text{ kJ})] - [(1 \text{ mol})\ (\Delta_fH^o\ C_7H_{16}(\ell))]$

$(1 \text{ mol})\ (\Delta_fH^o\ C_7H_{16}(\ell)) = -2754.5 - 2286.4 \text{ kJ} + 4816.7 \text{ kJ}$

$\Delta_rH^o = -224.2$ kJ/mol

Check Your Solution
This answer is reasonable, has the correct unit (kJ/mol) and the correct number of digits after the decimal (1).

13.
Problem
Using the data for the molar enthalpy of combustion of butane from Table 9.1 (page 347), determine the efficiency of a lighter as a heating device if 0.70 g of butane is required to raise the temperature of a 250 g stainless steel spoon (c = 0.503 J/g • °C) by 45 °C.

What is Required?
You must determine the efficiency of a butane lighter by comparing the heat released by the butane as it burns with the heat gained by a stainless steel pot.

What is Given?
Energy Input
$mC_4H_{10}(g) = 0.70$ g
$\Delta_c H^o = -2\ 657.3$ kJ/mol (from Table 9.1)
Energy output
$m_{spoon} = 250$ g
$c_{spoon} = 0.503$ J/g • °C
$\Delta t = 45.0$ °C

Plan Your Strategy
Determine the molar mass, M, for $C_4H_{10}(g)$ and calculate the number of moles of

$$nC_4H_{10}(g) = \frac{m}{M}.$$

Calculate the energy input.
Energy input = $Q = -n\Delta_c H^o$

Calculate the energy output.
Energy output = $Q = mc\Delta t$

Calculate the efficiency.

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\%$$

Act on Your Strategy
$MC_4H_{10}(g) = 58.14$ g/mol

$$nC_4H_{10}(g) = \frac{m}{M} = \frac{0.70\text{ g}}{58.14\text{ g/mol}} = 0.0120 \text{ mol}$$

Energy input = 0.0120 mol $C_4H_{10}(g)$ × (−2657.3 kJ/mol) = −31.99 kJ

Energy output = $mc\Delta t$ = 250 g × 0.503 J/g °C × 45.0 °C = 5.659 J = 5.659 kJ

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\% = \frac{5.659\ \text{kJ}}{31.99\ \text{kJ}} \times 100\% = 18\%$$

Check Your Solution
The efficiency is low, as expected when heating a spoon in this manner. The answer has the correct unit (%) and the correct number of significant digits (2).

14.
Problem
A camping fuel has a posted energy content of 50 kJ/g. Determine its efficiency if a 2.50 g piece was required to raise the temperature of 500 g of soup (c = 3.77 J/g • °C) in a 50 g aluminium pot by 45 °C.

What is Required?
You must determine the efficiency of a fuel by comparing the heat released by the fuel as it burns with the heat gained by an aluminium pot and the soup in the pot.

What is Given?
Energy Input
m_{fuel} = 2.50 g
energy content of fuel = 50 kJ/g
Energy output
m_{soup} = 500 g
c_{soup} = 3.77 J/g °C
$m_{Al\ pot}$ = 20 g
$c_{Al\ pot}$ = 0.897 J/g °C
Δt = 45.0°C

Plan Your Strategy
Calculate the heat output of the fuel using the energy content and the mass of fuel.
Energy input = $Q = -m$(energy content)

Calculate the energy output.
Energy output = $Q_{soup} + Q_{aluminium}$
= $(mc\Delta t)_{soup} + (mc\Delta t)_{aluminium}$

Calculate the efficiency.

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\%$$

Act on Your Strategy

Energy input = –2.50 g × 50 kJ/g = –125 kJ

Energy output = $(mc\Delta t)_{soup} + (mc\Delta t)_{aluminium}$
= 500 g × 3.77 J/g °C ×45 °C + 50 g × 0.897 J/g °C × 45 °C
= 84825 J + 2018 J
= 86843 J
= 86.8 kJ

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\% = \frac{86.8\text{ kJ}}{125\text{ kJ}} \times 100\% = 69\%$$

Check Your Solution

Considering heat loss to the surroundings, this is a reasonable answer. The answer has the correct unit (%) and the correct number of significant digits (2).

15.

Problem

What mass of pentane, $C_5H_{12}(g)$, would have to be burned in an open system to heat 250 g of hot chocolate (c = 3.59 J/g • °C) from 20.0 ºC to 39.8 ºC if the energy conversion is 45% efficient?

What is Required?

You must calculate the mass of pentane, $C_5H_{12}(g)$, needed to heat a cup of hot chocolate.

What is Given?

$m_{\text{hot chocolate}}$ = 250 g
$c_{\text{hot chocolate}}$ = 3.59 J/g • °C
t_i = 20.0 °C
t_f = 39.8 °C
$\Delta t = t_f - t_i$ = 39.8 °C – 20.0 °C = 19.8 °C
$\Delta_c H^o$ $C_5H_{12}(g)$ = –3244.8 kJ/mol (from Table 9.1)
Efficiency = 45 %
$M C_5H_{12}(g)$ = 72.17 g/mol

Plan Your Strategy

Calculate energy output.
Energy output = Heat absorbed = $Q_{\text{hot chocolate}}$
= $(mc\Delta t)_{\text{hot chocolate}}$

Use the expression for efficiency to calculate energy input.

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\%$$

Use the expression for energy input to calculate the number of moles of $C_5H_{12}(g)$.

$$n = \frac{-Q}{-\Delta_c H^\circ}$$

$$mC_5H_{12}(g) = n \times M$$

Act on Your Strategy

Energy output = $(mc\Delta t)_{\text{hot chocolate}}$

$= 250\ \text{g} \times 3.59\ \text{J/g}\ {}^\circ\text{C} \times 19.8\ {}^\circ\text{C} = 17770\ \text{J} = 17.77\ \text{kJ}$

$$\text{Energy input} = \frac{\text{Energy output}}{\text{Efficiency}} \times 100\% = \frac{17.77\ \text{kJ}}{45\%} \times 100\% = 39.49\ \text{kJ}$$

$$nC_5H_{12}(g) = \frac{-Q}{-\Delta_c H^\circ} = \frac{-39.49\ \text{kJ}}{-3244.8\ \text{kJ}} = 0.01216\ \text{mol}$$

$$mC_5H_{12}(g) = n \times M = 0.01216\ \text{mol} \times 72.17\ \text{g/mol} = 0.88\ \text{g}$$

Check Your Solution

The answer is reasonable for this efficiency. The answer has the correct unit (g) and number of significant digits (2).

16.

Problem

Determine the efficiency of a heating device that burns methanol, $CH_3OH(\ell)$, given the following information:

Data for Determining the Efficiency of a Methanol-Burning Heater

Quantity being measured	Data
Initial mass of burner	38.37 g
Final mass of burner	36.92 g
Mass of aluminium can	257.36 g
Mass of aluminium can and water	437.26 g
Initial temperature of water	10.45 °C
Final temperature of water	23.36 °C
$\Delta_c H(CH_3OH)$	$726.1\ \frac{\text{kJ}}{\text{mol}}$

What is Required?

You must determine the efficiency of a methanol burner by comparing the heat released by the fuel as it burns with the heat gained by an aluminium can and the water in the can.

What is Given?
Energy Input
initial mass of methanol burner = 38.37 g
final mass of methanol burner = 36.92 g
$m\text{CH}_3\text{OH}(\ell)$ = 38.37 g – 36.92 g = 1.45 g
$\Delta_c H^o$ $\text{CH}_3\text{OH}(\ell)$ = –725.9 kJ/mol (from Chemistry Data Booklet)
$M\text{CH}_3\text{OH}(\ell)$ = 32.05 g/mol
Energy output
$m\text{H}_2\text{O}(\ell)$ = [m(Al can + water)] – [m(Al can)] = 437.26 g – 257.36 g = 179.9 g
$c\text{H}_2\text{O}$ (ℓ) = 4.19 J/g °C
$m_{\text{Al can}}$ = 257.36 g
$c_{\text{Al can}}$ = 0.897 J/g °C
t_i = 10.45 °C
t_f = 23.36 °C
$\Delta t = t_f - t_i$ = 23.36 °C – 10.45 °C = 12.91 °C

Plan Your Strategy

$$n\text{CH}_3\text{OH}(\ell) = \frac{m}{M}$$

Calculate the energy input.
Energy Input = $Q = -n\Delta_c H^o$

Calculate the energy output.
Energy output = Heat absorbed = $Q_{\text{water}} + Q_{\text{aluminium}}$
= $(mc\Delta t)_{\text{water}} + (mc\Delta t)_{\text{aluminium}}$

Calculate the efficiency.

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\%$$

Act on Your Strategy

$$n\text{CH}_3\text{OH}(\ell) = \frac{m}{M} = \frac{1.45\,\text{g}}{32.05\frac{\text{g}}{\text{mol}}} = 0.04524\ \text{mol}$$

Energy input = 0.04524 mol × –725.9 kJ/mol = –32.84 kJ

Energy output = $(mc\Delta t)_{\text{water}} + (mc\Delta t)_{\text{alumium}}$
= 179.9 g × 4.19 J/g • °C × 12.91 °C + 257.36 g × 0.897 J/g • °C × 12.91 °C
= 9731.3 J + 2980.3 J
= 12711.6 J
= 12.71 kJ

$$\text{Efficiency} = \frac{\text{Energy output}}{\text{Energy input}} \times 100\% = \frac{12.71\,\text{kJ}}{32.84\,\text{kJ}} \times 100\% = 38.7\%$$

Check Your Solution

Considering heat loss to the surroundings, this is a reasonable answer. The answer has the correct unit (%) and the correct number of significant digits (3).