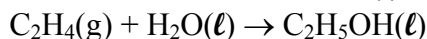


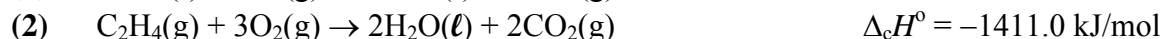
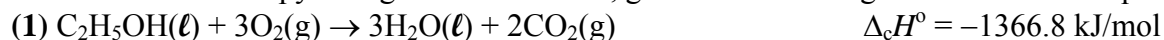
**Unit 5 Thermochemical Changes**  
**Chapter 10 Theories of Energy and Chemical Changes**  
**Solutions to Practice Problems**

**1.****Problem**

Ethene,  $\text{C}_2\text{H}_4(\text{g})$ , reacts with water to form ethanol,  $\text{C}_2\text{H}_5\text{OH}(\ell)$ , as shown below:



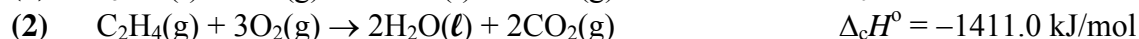
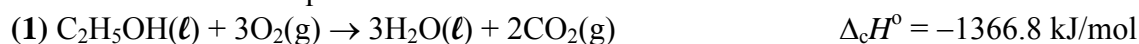
Determine the enthalpy change of this reaction, given the following thermochemical equations.

**What is Required?**

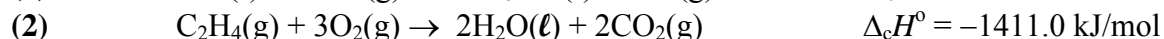
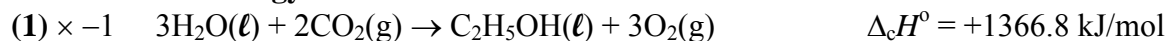
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

**What is Given?**

The thermochemical equations:

**Plan Your Strategy**

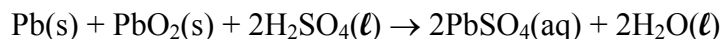
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change,  $\Delta_r H^\circ$ , is the algebraic sum of the enthalpy changes for the reactions used.

**Act on Your Strategy****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:



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

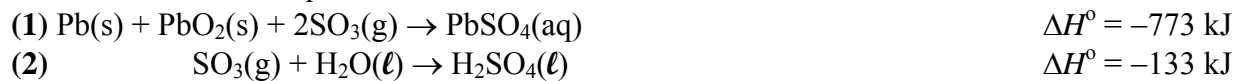


**What is Required?**

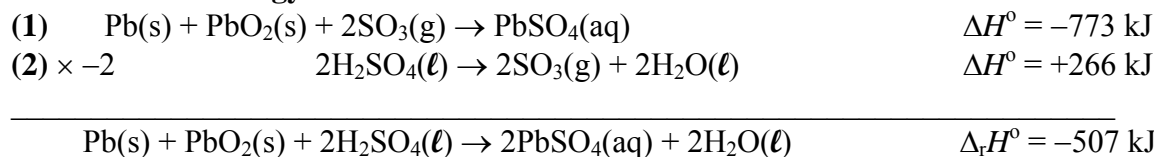
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

**What is Given?**

The thermochemical equations:

**Plan Your Strategy**

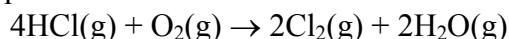
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change,  $\Delta_r H^\circ$ , is the algebraic sum of the enthalpy changes for the reactions used.

**Act on Your Strategy****Check Your Solution**

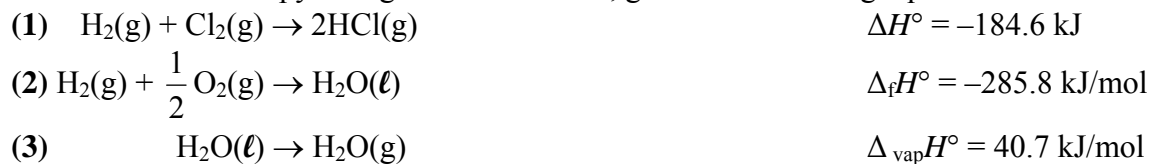
The reaction has a negative enthalpy change and therefore it is an exothermic reaction, as expected. The unit is correct (kJ) and the answer has the correct number of digits after the decimal (0).

**3.****Problem**

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



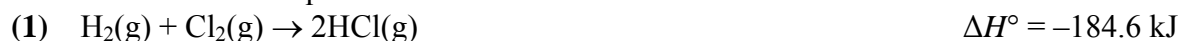
Determine the enthalpy change of this reaction, given the following equations:

**What is Required?**

You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

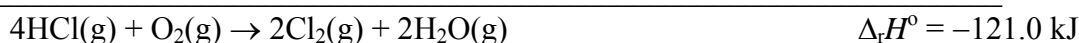
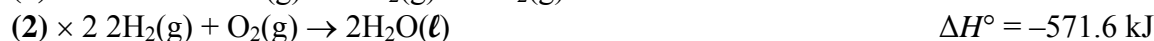
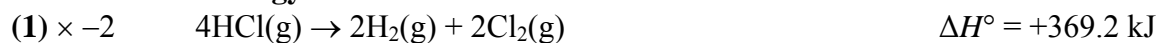
**What is Given?**

The thermochemical equations:



**Plan Your Strategy**

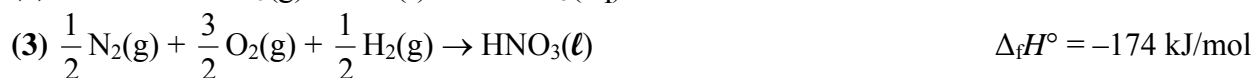
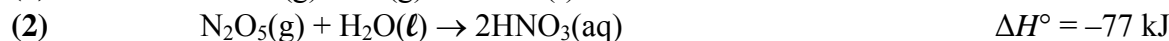
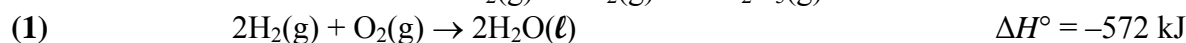
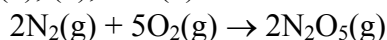
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change,  $\Delta_r H^\circ$ , is the algebraic sum of the enthalpy changes for the reactions used.

**Act on Your Strategy****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**

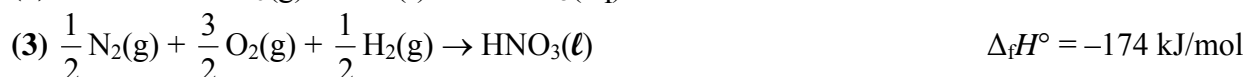
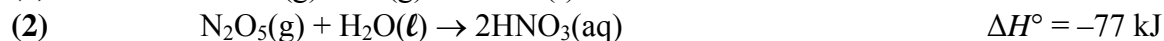
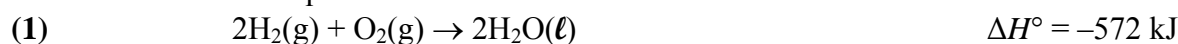
Calculate the enthalpy change of the following reaction between nitrogen gas and oxygen gas, given thermochemical equations (1), (2), and (3):

**What is Required?**

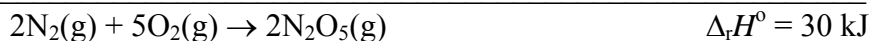
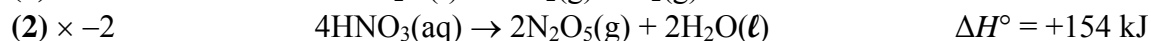
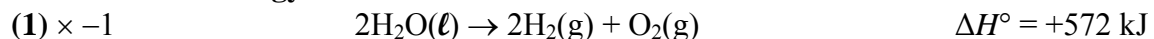
You must manipulate the given equations to calculate the enthalpy change for the overall reaction.

**What is Given?**

The thermochemical equations:

**Plan Your Strategy**

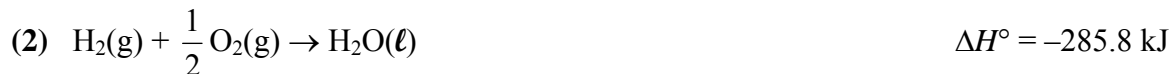
Applying Hess's Law, manipulate the given thermochemical equations to obtain an algebraic sum equivalent to the overall reaction. The overall enthalpy change,  $\Delta_r H^\circ$ , is the algebraic sum of the enthalpy changes for the reactions used.

**Act on Your Strategy****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**

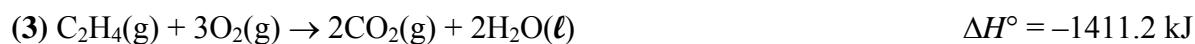
Ethene,  $\text{C}_2\text{H}_4(\text{g})$ , is used in the manufacturing of many polymers, including polyethylene terephthalate (PET), which is used to make pop bottles. Determine the molar enthalpy of formation for ethene, as shown by  $2\text{C}(\text{s}) + 2\text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_4(\text{g})$ , given the following equations:

**What is Required?**

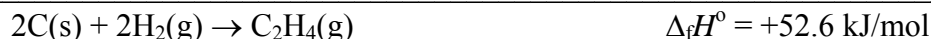
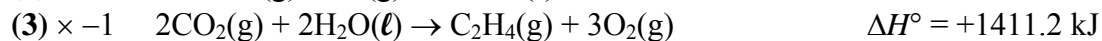
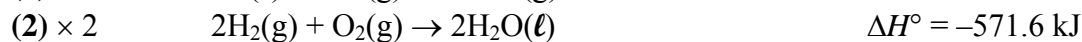
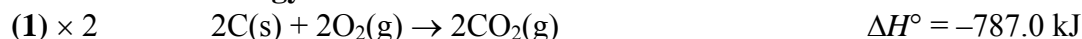
You must manipulate the given equations to calculate the enthalpy of formation for ethene,  $\text{C}_2\text{H}_4(\text{g})$ .

**What is Given?**

The thermochemical equations:

**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,  $\Delta_f H^\circ$ .

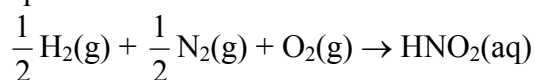
**Act on Your Strategy**

**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  $\text{HNO}_2(\text{aq})$ , as shown below in the overall equation:



- |     |  |                                      |
|-----|--|--------------------------------------|
| (1) | $\text{NH}_4\text{NO}_2(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell)$ | $\Delta H^\circ = -320.1 \text{ kJ}$ |
| (2) | $\text{NH}_3(\text{aq}) + \text{HNO}_2(\text{aq}) \rightarrow \text{NH}_4\text{NO}_2(\text{aq})$ | $\Delta H^\circ = -37.7 \text{ kJ}$  |
| (3) | $2\text{NH}_3(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$               | $\Delta H^\circ = +169.9 \text{ kJ}$ |
| (4) | $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\ell)$   | $\Delta H^\circ = -285.8 \text{ kJ}$ |

**What is Required?**

You must manipulate the given equations to calculate the enthalpy of formation for  $\text{HNO}_2(\text{aq})$ .

**What is Given?**

The thermochemical equations:

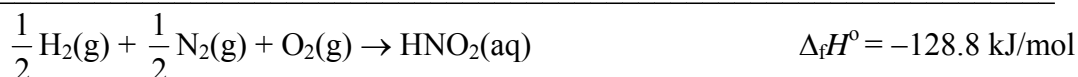
- |     |  |                                      |
|-----|--|--------------------------------------|
| (1) | $\text{NH}_4\text{NO}_2(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell)$ | $\Delta H^\circ = -320.1 \text{ kJ}$ |
| (2) | $\text{NH}_3(\text{aq}) + \text{HNO}_2(\text{aq}) \rightarrow \text{NH}_4\text{NO}_2(\text{aq})$ | $\Delta H^\circ = -37.7 \text{ kJ}$  |
| (3) | $2\text{NH}_3(\text{aq}) \rightarrow \text{N}_2(\text{g}) + 3\text{H}_2(\text{g})$               | $\Delta H^\circ = +169.9 \text{ kJ}$ |
| (4) | $\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\ell)$   | $\Delta H^\circ = -285.8 \text{ kJ}$ |

**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,  $\Delta_f H^\circ$  for  $\text{HNO}_2(\text{aq})$ .

**Act on Your Strategy**

- |                  |  |                                      |
|------------------|--|--------------------------------------|
| (1) $\times -1$  | $\text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4\text{NO}_2(\text{aq})$         | $\Delta H^\circ = +320.1 \text{ kJ}$ |
| (2) $\times -1$  | $\text{NH}_4\text{NO}_2(\text{aq}) \rightarrow \text{NH}_3(\text{aq}) + \text{HNO}_2(\text{aq})$         | $\Delta H^\circ = +37.7 \text{ kJ}$  |
| (3) $\times 1/2$ | $\text{NH}_3(\text{aq}) \rightarrow \frac{1}{2} \text{N}_2(\text{g}) + \frac{3}{2} \text{H}_2(\text{g})$ | $\Delta H^\circ = +85.0 \text{ kJ}$  |
| (4) $\times 2$   | $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\ell)$                     | $\Delta H^\circ = -571.6 \text{ kJ}$ |



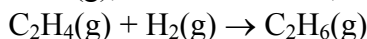
**Check Your Solution**

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,  $\text{C}_2\text{H}_4(\text{g})$ , to obtain ethane,  $\text{C}_2\text{H}_6(\text{g})$ :



Write thermochemical equations for the formation of both ethene,  $\text{C}_2\text{H}_4(\text{g})$ , and ethane,  $\text{C}_2\text{H}_6(\text{g})$ . 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,  $\text{C}_2\text{H}_4(\text{g})$ , and ethane,  $\text{C}_2\text{H}_6(\text{g})$  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.

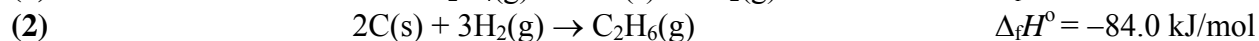
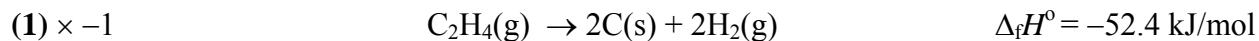
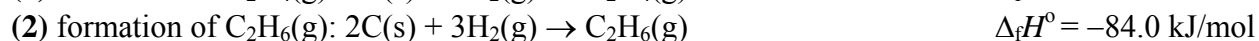
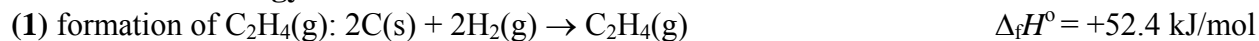
**What is Given?**

The balanced equation for the reaction between ethene and hydrogen to produce ethane is given. The heats of formation of ethene,  $\text{C}_2\text{H}_4(\text{g})$ , and ethane,  $\text{C}_2\text{H}_6(\text{g})$  can be found in the Chemistry Data Booklet.

**Plan Your Strategy**

Write the equation for the formation of ethene,  $\text{C}_2\text{H}_4(\text{g})$ , and ethane,  $\text{C}_2\text{H}_6(\text{g})$ . Apply Hess's Law to obtain the overall reaction between ethene,  $\text{C}_2\text{H}_4(\text{g})$ , and ethane,  $\text{C}_2\text{H}_6(\text{g})$ .

The overall enthalpy change,  $\Delta_r H^\circ$  is the algebraic sum of the enthalpy changes for the reactions used.

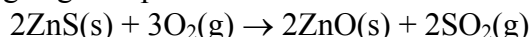
**Act on Your Strategy****Check Your Solution**

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:



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.

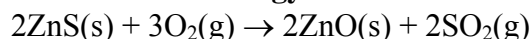
### What is Given?

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,  $\Delta_f H^\circ$ , for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression  $\Delta_r H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$

### Act on Your Strategy



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

$$\Delta_r H^\circ = [(2 \text{ mol}) (\Delta_f H^\circ \text{ ZnO}(s)) + (2 \text{ mol}) (\Delta_f H^\circ \text{ SO}_2(g))] - [(2 \text{ mol}) (\Delta_f H^\circ \text{ ZnS}(s)) + (3 \text{ mol}) (\Delta_f H^\circ \text{ O}_2(g))]$$

$$\Delta_r H^\circ = [(2 \text{ mol}) (-350.5 \text{ kJ/mol}) + (2 \text{ mol}) (-296.8 \text{ kJ/mol})] - [(2 \text{ mol}) (-206.0 \text{ kJ/mol}) + (3 \text{ mol}) (0)]$$

$$\Delta_r H^\circ = [-701.0 \text{ kJ} + (-593.6)] - [-412.0 \text{ kJ}]$$

$$\Delta_r H^\circ = -882.6 \text{ kJ}$$

### 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,  $\text{KClO}_3(s)$ :



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,  $\Delta_f H^\circ$ , for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression  $\Delta_r H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$

**Act on Your Strategy**

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

$$\Delta_r H^\circ = [(2 \text{ mol}) (\Delta_f H^\circ \text{ KCl}(\text{s})) + (3 \text{ mol}) (\Delta_f H^\circ \text{ O}_2(\text{g}))] - [(2 \text{ mol}) (\Delta_f H^\circ \text{ KClO}_3(\text{s}))]$$

$$\Delta_r H^\circ = [(2 \text{ mol}) (-436.5 \text{ kJ/mol}) + (3 \text{ mol}) (0)] - [(2 \text{ mol}) (-397.7 \text{ kJ/mol})]$$

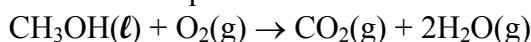
$$\Delta_r H^\circ = [-873.0 \text{ kJ}] - [-795.4 \text{ kJ}] = -77.6 \text{ kJ}$$

**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:



**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**

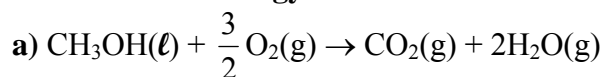
**a)** Look up the standard enthalpies of formation,  $\Delta_f H^\circ$ , for each reactant and product in this reaction and calculate the enthalpy of reaction by substituting these values into the expression  $\Delta_r H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ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

**b)** Determine the molar mass ( $M$ ) of methanol,  $\text{CH}_3\text{OH}(\ell)$ , 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**

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

$$\Delta H^\circ = [(1 \text{ mol}) (\Delta_f H^\circ \text{CO}_2(\text{g})) + (2 \text{ mol}) (\Delta_f H^\circ \text{H}_2\text{O}(\text{g}))] - [(1 \text{ mol}) (\Delta_f H^\circ \text{CH}_3\text{OH}(\ell)) + \frac{3}{2} \Delta_f H^\circ \text{O}_2(\text{g})]$$

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

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

$$\Delta_r H^\circ = -637.9 \text{ kJ per mol of CH}_3\text{OH}(\ell)$$

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

$$n\text{CH}_3\text{OH}(\ell) = \frac{m}{M} = \frac{125 \text{ g}}{32.05 \text{ g/mol}} = 3.90 \text{ mol}$$

$$\text{energy released} = 3.90 \text{ mol CH}_3\text{OH}(\ell) \times (-637.9 \text{ kJ/mol}) = -2.49 \times 10^3 \text{ kJ}$$

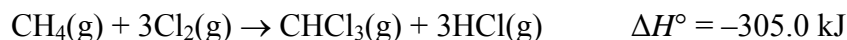
**Check Your Solution**

a) The enthalpy of reaction is negative and the reaction is exothermic as expected. The answer has the correct unit (kJ per mol of  $\text{CH}_3\text{OH}(\ell)$ ), 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,  $\text{CHCl}_3(\text{g})$ :



Use standard molar enthalpies of formation to determine the molar enthalpy of formation for chloroform,  $\text{CHCl}_3(\text{g})$ .

**What is Required?**

You must calculate the enthalpy of formation,  $\Delta_f H^\circ$ , of chloroform,  $\text{CHCl}_3(\text{g})$ .

**What is Given?**

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

**Plan Your Strategy**

Look up the standard enthalpies of formation,  $\Delta_f H^\circ$ , for  $\text{CH}_4(\text{g})$  and  $\text{HCl}(\text{g})$ . Use these values and the given enthalpy of reaction,  $\Delta_r H^\circ$ , in the expression

$\Delta_r H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants})$ . Solve for  $\Delta_f H^\circ$  of  $\text{CHCl}_3(\text{g})$ .

### Act on Your Strategy



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

$$\Delta_r H^\circ = [(1 \text{ mol}) (\Delta_f H^\circ \text{ CHCl}_3(\text{g})) + (3 \text{ mol}) (\Delta_f H^\circ \text{ HCl}(\text{g}))] - [(1 \text{ mol}) (\Delta_f H^\circ \text{ CH}_4(\text{g})) + (3 \text{ mol}) (\Delta_f H^\circ \text{ Cl}_2(\text{g}))]$$

$$-305.0 \text{ kJ} = [(1 \text{ mol}) (\Delta_f H^\circ \text{ CHCl}_3(\text{g})) + (3 \text{ mol}) (-92.3 \text{ kJ/mol})] - [(1 \text{ mol}) (-74.6 \text{ kJ/mol}) + (3 \text{ mol}) (0)]$$

$$(1 \text{ mol})\Delta_f H^\circ \text{ CHCl}_3(\text{g}) = -305.0 \text{ kJ} + 276.9 \text{ kJ} - 74.6 \text{ kJ}$$

$$\Delta_r H^\circ = -102.7 \text{ kJ/mol}$$

### 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,  $\text{C}_7\text{H}_{16}(\ell)$ , in a bomb calorimeter is  $-4816.7 \text{ kJ/mol}$  of heptane. Using this and  $\Delta_f H^\circ$  data, determine the molar enthalpy of formation of heptane.

### What is Required?

You must calculate the enthalpy of formation,  $\Delta_f H^\circ$ , of heptane,  $\text{C}_7\text{H}_{16}(\ell)$ .

### What is Given?

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

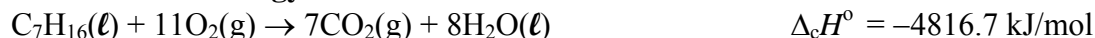
### Plan Your Strategy

Write the balanced equation for the combustion of heptane,  $\text{C}_7\text{H}_{16}(\ell)$ , to  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\ell)$ .

Look up the standard enthalpies of formation,  $\Delta_f H^\circ$ , for  $\text{CO}_2(\text{g})$  and  $\text{H}_2\text{O}(\ell)$ . Use these values and the given enthalpy of reaction,  $\Delta_r H^\circ$ , in the expression

$$\Delta_c H^\circ = \Sigma(n\Delta_f H^\circ \text{ products}) - \Sigma(n\Delta_f H^\circ \text{ reactants}). \text{ Solve for } \Delta_f H^\circ \text{ C}_7\text{H}_{16}(\ell).$$

### Act on Your Strategy



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

$$\Delta_c H^\circ = [(7 \text{ mol})(\Delta_f H^\circ \text{ CO}_2(\text{g})) + (8 \text{ mol})(\Delta_f H^\circ \text{ H}_2\text{O}(\ell))] - [(1 \text{ mol}) (\Delta_f H^\circ \text{ C}_7\text{H}_{16}(\ell)) + ((11 \text{ mol}) (\Delta_f H^\circ \text{ O}_2(\text{g})))]$$

$$-4816.7 \text{ kJ} = [(7 \text{ mol}) (-393.5 \text{ kJ}) + (8 \text{ mol}) (-285.8 \text{ kJ/mol})] - [(1 \text{ mol}) (\Delta_f H^\circ \text{ C}_7\text{H}_{16}(\ell)) + (11 \text{ mol}) (0)]$$

$$-4816.7 \text{ kJ} = [(-2754.5 \text{ kJ}\cdot\text{mol}) + (-2286.4 \text{ kJ})] - [(1 \text{ mol}) (\Delta_f H^\circ \text{ C}_7\text{H}_{16}(\ell))]$$

$$(1 \text{ mol}) (\Delta_f H^\circ \text{ C}_7\text{H}_{16}(\ell)) = -2754.5 - 2286.4 \text{ kJ} + 4816.7 \text{ kJ}$$

$$\Delta_f H^\circ = -224.2 \text{ kJ/mol}$$

**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 ( $c = 0.503 \text{ J/g} \cdot ^\circ\text{C}$ ) by  $45^\circ\text{C}$ .

**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

$$m_{\text{C}_4\text{H}_{10}(\text{g})} = 0.70 \text{ g}$$

$$\Delta_c H^\circ = -2\,657.3 \text{ kJ/mol (from Table 9.1)}$$

Energy output

$$m_{\text{spoon}} = 250 \text{ g}$$

$$c_{\text{spoon}} = 0.503 \text{ J/g} \cdot ^\circ\text{C}$$

$$\Delta t = 45.0^\circ\text{C}$$

**Plan Your Strategy**

Determine the molar mass,  $M$ , for  $\text{C}_4\text{H}_{10}(\text{g})$  and calculate the number of moles of

$$n_{\text{C}_4\text{H}_{10}(\text{g})} = \frac{m}{M}$$

Calculate the energy input.

$$\text{Energy input} = Q = -n\Delta_c H^\circ$$

Calculate the energy output.

$$\text{Energy output} = Q = mc\Delta t$$

Calculate the efficiency.

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

**Act on Your Strategy**

$$M_{\text{C}_4\text{H}_{10}(\text{g})} = 58.14 \text{ g/mol}$$

$$n_{\text{C}_4\text{H}_{10}(\text{g})} = \frac{m}{M} = \frac{0.70 \text{ g}}{58.14 \text{ g/mol}} = 0.0120 \text{ mol}$$

$$\text{Energy input} = 0.0120 \text{ mol C}_4\text{H}_{10}(\text{g}) \times (-2657.3 \text{ kJ/mol}) = -31.99 \text{ kJ}$$

$$\text{Energy output} = mc\Delta t = 250 \text{ g} \times 0.503 \text{ J/g } ^\circ\text{C} \times 45.0 \text{ } ^\circ\text{C} = 5.659 \text{ J} = 5.659 \text{ kJ}$$

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

### 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 ( $c = 3.77 \text{ J/g} \cdot ^\circ\text{C}$ ) in a 50 g aluminium pot by  $45 \text{ } ^\circ\text{C}$ .

### 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

$$m_{\text{fuel}} = 2.50 \text{ g}$$

$$\text{energy content of fuel} = 50 \text{ kJ/g}$$

Energy output

$$m_{\text{soup}} = 500 \text{ g}$$

$$c_{\text{soup}} = 3.77 \text{ J/g } ^\circ\text{C}$$

$$m_{\text{Al pot}} = 20 \text{ g}$$

$$c_{\text{Al pot}} = 0.897 \text{ J/g } ^\circ\text{C}$$

$$\Delta t = 45.0 \text{ } ^\circ\text{C}$$

### Plan Your Strategy

Calculate the heat output of the fuel using the energy content and the mass of fuel.

$$\text{Energy input} = Q = -m(\text{energy content})$$

Calculate the energy output.

$$\begin{aligned} \text{Energy output} &= Q_{\text{soup}} + Q_{\text{aluminium}} \\ &= (mc\Delta t)_{\text{soup}} + (mc\Delta t)_{\text{aluminium}} \end{aligned}$$

Calculate the efficiency.

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

**Act on Your Strategy**

$$\text{Energy input} = -2.50 \text{ g} \times 50 \text{ kJ/g} = -125 \text{ kJ}$$

$$\begin{aligned} \text{Energy output} &= (mc\Delta t)_{\text{soup}} + (mc\Delta t)_{\text{aluminium}} \\ &= 500 \text{ g} \times 3.77 \text{ J/g} \cdot ^\circ\text{C} \times 45 \text{ }^\circ\text{C} + 50 \text{ g} \times 0.897 \text{ J/g} \cdot ^\circ\text{C} \times 45 \text{ }^\circ\text{C} \\ &= 84825 \text{ J} + 2018 \text{ J} \\ &= 86843 \text{ J} \\ &= 86.8 \text{ kJ} \end{aligned}$$

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

**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,  $\text{C}_5\text{H}_{12}(\text{g})$ , would have to be burned in an open system to heat 250 g of hot chocolate ( $c = 3.59 \text{ J/g} \cdot ^\circ\text{C}$ ) from  $20.0 \text{ }^\circ\text{C}$  to  $39.8 \text{ }^\circ\text{C}$  if the energy conversion is 45% efficient?

**What is Required?**

You must calculate the mass of pentane,  $\text{C}_5\text{H}_{12}(\text{g})$ , needed to heat a cup of hot chocolate.

**What is Given?**

$$m_{\text{hot chocolate}} = 250 \text{ g}$$

$$c_{\text{hot chocolate}} = 3.59 \text{ J/g} \cdot ^\circ\text{C}$$

$$t_i = 20.0 \text{ }^\circ\text{C}$$

$$t_f = 39.8 \text{ }^\circ\text{C}$$

$$\Delta t = t_f - t_i = 39.8 \text{ }^\circ\text{C} - 20.0 \text{ }^\circ\text{C} = 19.8 \text{ }^\circ\text{C}$$

$$\Delta_c H^\circ \text{C}_5\text{H}_{12}(\text{g}) = -3244.8 \text{ kJ/mol (from Table 9.1)}$$

$$\text{Efficiency} = 45 \%$$

$$M_{\text{C}_5\text{H}_{12}(\text{g})} = 72.17 \text{ g/mol}$$

**Plan Your Strategy**

Calculate energy output.

$$\begin{aligned} \text{Energy output} &= \text{Heat absorbed} = Q_{\text{hot chocolate}} \\ &= (mc\Delta t)_{\text{hot chocolate}} \end{aligned}$$

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

$$\begin{aligned} \text{Energy output} &= (mc\Delta t)_{\text{hot chocolate}} \\ &= 250 \text{ g} \times 3.59 \text{ J/g } ^\circ\text{C} \times 19.8 \text{ } ^\circ\text{C} = 17770 \text{ J} = 17.77 \text{ kJ} \end{aligned}$$

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

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

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

### 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,  $CH_3OH(l)$ , 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

 $m_{\text{CH}_3\text{OH}(\ell)} = 38.37 \text{ g} - 36.92 \text{ g} = 1.45 \text{ g}$  $\Delta_c H^\circ \text{CH}_3\text{OH}(\ell) = -725.9 \text{ kJ/mol}$  (from Chemistry Data Booklet) $M_{\text{CH}_3\text{OH}(\ell)} = 32.05 \text{ g/mol}$ 

Energy output

 $m_{\text{H}_2\text{O}(\ell)} = [m(\text{Al can} + \text{water})] - [m(\text{Al can})] = 437.26 \text{ g} - 257.36 \text{ g} = 179.9 \text{ g}$  $c_{\text{H}_2\text{O}(\ell)} = 4.19 \text{ J/g } ^\circ\text{C}$  $m_{\text{Al can}} = 257.36 \text{ g}$  $c_{\text{Al can}} = 0.897 \text{ J/g } ^\circ\text{C}$  $t_i = 10.45 \text{ } ^\circ\text{C}$  $t_f = 23.36 \text{ } ^\circ\text{C}$  $\Delta t = t_f - t_i = 23.36 \text{ } ^\circ\text{C} - 10.45 \text{ } ^\circ\text{C} = 12.91 \text{ } ^\circ\text{C}$ **Plan Your Strategy**

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

Calculate the energy input.

$$\text{Energy Input} = Q = -n\Delta_c H^\circ$$

Calculate the energy output.

$$\begin{aligned} \text{Energy output} = \text{Heat absorbed} &= Q_{\text{water}} + Q_{\text{aluminium}} \\ &= (mc\Delta t)_{\text{water}} + (mc\Delta t)_{\text{aluminium}} \end{aligned}$$

Calculate the efficiency.

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

**Act on Your Strategy**

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

$$\text{Energy input} = 0.04524 \text{ mol} \times -725.9 \text{ kJ/mol} = -32.84 \text{ kJ}$$

$$\begin{aligned} \text{Energy output} &= (mc\Delta t)_{\text{water}} + (mc\Delta t)_{\text{aluminium}} \\ &= 179.9 \text{ g} \times 4.19 \text{ J/g} \cdot ^\circ\text{C} \times 12.91 \text{ } ^\circ\text{C} + 257.36 \text{ g} \times 0.897 \text{ J/g} \cdot ^\circ\text{C} \times 12.91 \text{ } ^\circ\text{C} \\ &= 9731.3 \text{ J} + 2980.3 \text{ J} \\ &= 12711.6 \text{ J} \\ &= 12.71 \text{ kJ} \end{aligned}$$

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

**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).