

## Section 5.5: Standard Enthalpies of Formation

### Tutorial 1 Practice, page 323

1. **Given:** from Table 1,  $\Delta H_f^\circ \text{C}_2\text{H}_2(\text{g}) = 228.2 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ , and  $\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$

**Required:**  $\Delta H_r^\circ$  for the combustion of acetylene gas

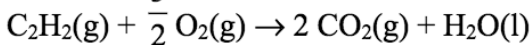
**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since  $\text{O}_2(\text{g})$  is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_2(\text{g})} \times \Delta H_f^\circ \text{C}_2\text{H}_2(\text{g})$$

**Solution:**

Step 1: Write a balanced chemical equation for the combustion of acetylene gas so that acetylene gas has a coefficient of 1.



Step 2: Substitute the appropriate values for standard enthalpy of formation into the equation and solve.

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_2(\text{g})} \times \Delta H_f^\circ \text{C}_2\text{H}_2(\text{g}) \\ &= [2(-393.5 \text{ kJ}) + (-285.8 \text{ kJ})] - (228.2 \text{ kJ}) \\ &= -787.0 \text{ kJ} - 285.8 \text{ kJ} - 228.2 \text{ kJ} \end{aligned}$$

$$\Delta H_r^\circ = -1301.0 \text{ kJ}$$

**Statement:** Since the reaction is for 1 mol of acetylene gas, the standard enthalpy of combustion for acetylene gas is  $-1301.0 \text{ kJ/mol}$ .

2. **Given:** from Question 1,  $\Delta H_r^\circ$  for combustion of acetylene gas =  $-1301 \text{ kJ/mol}$ ;

from Table 1,  $\Delta H_f^\circ \text{C}_3\text{H}_8(\text{g}) = -104.7 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ , and

$$\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}; \Delta H_f^\circ \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_r^\circ$  per 1.00 g acetylene gas and per 1.00 g propane gas

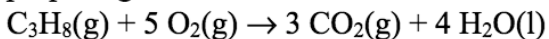
**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since  $\text{O}_2(\text{g})$  is in its standard state, the equation for propane gas can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_3\text{H}_8(\text{g})} \times \Delta H_f^\circ \text{C}_3\text{H}_8(\text{g})$$

**Solution:**

Step 1: Write a balanced chemical equation for the combustion of propane gas so that propane gas has a coefficient of 1.



Step 2: Substitute the appropriate values for standard enthalpy of formation into the equation and solve.

$$\begin{aligned}\Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_3\text{H}_8(\text{g})} \times \Delta H_f^\circ \text{C}_3\text{H}_8(\text{g}) \\ &= [3(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})] - (-104.7 \text{ kJ}) \\ &= -1180 \text{ kJ} - 1143.2 \text{ kJ} + 104.7 \text{ kJ}\end{aligned}$$

$$\Delta H_r^\circ = -2218.5 \text{ kJ}$$

Step 3: Convert to enthalpy per gram of acetylene and enthalpy per gram of propane.

$$M_{\text{C}_2\text{H}_2(\text{g})} = 26.04 \text{ g/mol and } M_{\text{C}_3\text{H}_8(\text{g})} = 44.11 \text{ g/mol}$$

For acetylene:

$$\begin{aligned}\Delta H_r^\circ \text{ per gram C}_2\text{H}_2(\text{g}) &= \frac{\Delta H_r^\circ \text{ per mole C}_2\text{H}_2(\text{g})}{M_{\text{C}_2\text{H}_2(\text{g})}} \\ &= \frac{-1301 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{26.04 \text{ g}}\end{aligned}$$

$$\Delta H_r^\circ \text{ per gram C}_2\text{H}_2(\text{g}) = -50.0 \text{ kJ/g}$$

For propane:

$$\begin{aligned}\Delta H_r^\circ \text{ per gram C}_3\text{H}_8(\text{g}) &= \frac{\Delta H_r^\circ \text{ per mole C}_3\text{H}_8(\text{g})}{M_{\text{C}_3\text{H}_8(\text{g})}} \\ &= \frac{-2218.5 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{44.11 \text{ g}}\end{aligned}$$

$$\Delta H_r^\circ \text{ per gram C}_3\text{H}_8(\text{g}) = -50.3 \text{ kJ/g}$$

**Statement:** The enthalpy of combustion of 1.00 g of acetylene gas is  $-50.0 \text{ kJ}$  and that of 1.00 g of propane gas is  $-50.3 \text{ kJ}$ . For 1.00 g of the gas, propane releases  $0.3 \text{ kJ}$  more energy than acetylene.

### Section 5.5 Questions, page 324

1. Under SATP conditions, magnesium, Mg, is in solid state and bromine, Br<sub>2</sub>, is in liquid state. So, (b) and (d) are in standard states.

2. (a) **Given:** from Table 1,  $\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{H}_2(\text{g}) = 0 \text{ kJ/mol}$ ;

$$\Delta H_f^\circ \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_r^\circ$

$$\text{Analysis: } \Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$$

Since H<sub>2</sub>(g) and O<sub>2</sub>(g) are in their standard states, the equation can be written as:

$$\Delta H_r^\circ = n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})$$

**Solution:**

From the equation:  $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l})$ ,

$$n_{\text{H}_2\text{O}(\text{l})} = 2$$

$$\begin{aligned} \Delta H_{\text{r}}^{\circ} &= n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}(\text{l}) \\ &= 2(-285.8 \text{ kJ}) \end{aligned}$$

$$\Delta H_{\text{r}}^{\circ} = -571.6 \text{ kJ}$$

**Statement:**  $\Delta H_{\text{r}}^{\circ}$  for the reaction is  $-571.6 \text{ kJ}$ .

**(b) Given:** from Table 1,  $\Delta H_{\text{f}}^{\circ} \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ ;  $\Delta H_{\text{f}}^{\circ} \text{C}(\text{s}) = 0 \text{ kJ/mol}$ ;

$$\Delta H_{\text{f}}^{\circ} \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_{\text{r}}^{\circ}$

$$\textbf{Analysis: } \Delta H_{\text{r}}^{\circ} = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$$

Since  $\text{C}(\text{s})$  and  $\text{O}_2(\text{g})$  are in their standard states, the equation can be written as:

$$\Delta H_{\text{r}}^{\circ} = n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f}}^{\circ} \text{CO}_2(\text{g})$$

**Solution:**

From the equation:  $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$ ,

$$n_{\text{CO}_2(\text{g})} = 1$$

$$\Delta H_{\text{r}}^{\circ} = n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f}}^{\circ} \text{CO}_2(\text{g})$$

$$\Delta H_{\text{r}}^{\circ} = -393.5 \text{ kJ}$$

**Statement:**  $\Delta H_{\text{r}}^{\circ}$  for the reaction is  $-393.5 \text{ kJ}$ .

**(c) Given:** from Table 1,  $\Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$ ;  $\Delta H_{\text{f}}^{\circ} \text{H}_2(\text{g}) = 0 \text{ kJ/mol}$ ;

$$\Delta H_{\text{f}}^{\circ} \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_{\text{r}}^{\circ}$

$$\textbf{Analysis: } \Delta H_{\text{r}}^{\circ} = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$$

Since  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$  are in their standard states, the equation can be written as:

$$\Delta H_{\text{r}}^{\circ} = -n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}(\text{l})$$

**Solution:**

From the equation:  $2 \text{H}_2\text{O}(\text{l}) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$ ,

$$n_{\text{H}_2\text{O}(\text{l})} = 2$$

$$\begin{aligned}\Delta H_{\text{r}}^{\circ} &= -n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}(\text{l}) \\ &= -2(-285.8 \text{ kJ})\end{aligned}$$

$$\Delta H_{\text{r}}^{\circ} = 571.6 \text{ kJ}$$

**Statement:**  $\Delta H_{\text{r}}^{\circ}$  for the reaction is 571.6 kJ.

**(d) Given:** from Table 1,  $\Delta H_{\text{f}}^{\circ} \text{C}_2\text{H}_5\text{OH}(\text{l}) = -235.2 \text{ kJ/mol}$ ,  $\Delta H_{\text{f}}^{\circ} \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ ,  
and  $\Delta H_{\text{f}}^{\circ} \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$ ;  $\Delta H_{\text{f}}^{\circ} \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$

**Required:**  $\Delta H_{\text{r}}^{\circ}$

**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$

Since  $\text{O}_2(\text{g})$  is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_5\text{OH}(\text{l})} \times \Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\text{l})$$

**Solution:**

From the equation:  $\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\text{l})$ ,

$$n_{\text{C}_2\text{H}_5\text{OH}(\text{l})} = 1, n_{\text{CO}_2(\text{g})} = 2, \text{ and } n_{\text{H}_2\text{O}(\text{l})} = 3$$

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_5\text{OH}(\text{l})} \times \Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\text{l}) \\ &= [2(-393.5 \text{ kJ}) + 3(-285.8 \text{ kJ})] - (-235.2 \text{ kJ}) \\ &= -787.0 \text{ kJ} - 857.4 \text{ kJ} + 235.2 \text{ kJ} \end{aligned}$$

$$\Delta H_r^\circ = -1409.2 \text{ kJ}$$

**Statement:**  $\Delta H_r^\circ$  for the reaction is  $-1409.2 \text{ kJ}$ .

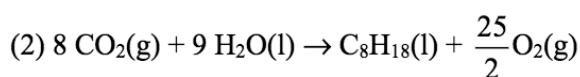
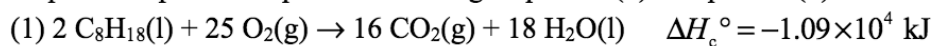
**3. Given:**  $\Delta H_c^\circ = -1.09 \times 10^4 \text{ kJ}$

**Required:**  $\Delta H_r^\circ$  for the required reaction

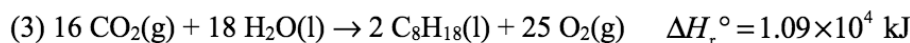
**Analysis:** rules from Hess's law

**Solution:**

Step 1: Compare the equations to change equation (1) to equation (2).



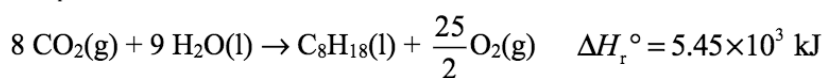
Step 2: Reverse (1) so that reactants and products are on the same side as (2). Reverse the sign of  $\Delta H_c^\circ$ .



Step 3: Equation (3) has 2 mol of  $\text{C}_8\text{H}_{18}(\text{l})$  on the product side, while equation (2) has 1 mol. Multiply equation (3) and its  $\Delta H_r^\circ$  by the factor  $\frac{1}{2}$ .

$$\Delta H_r^\circ = \frac{1}{2} \times 1.09 \times 10^4 \text{ kJ}$$

$$\Delta H_r^\circ = 5.45 \times 10^3 \text{ kJ}$$



**Statement:**  $\Delta H_r^\circ$  for the required reaction is  $5.45 \times 10^3 \text{ kJ}$ .

**4. Given:** from Table 1,  $\Delta H_f^\circ \text{NH}_4\text{ClO}_4(\text{s}) = -295.8 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{Al}_2\text{O}_3(\text{s}) = -1675.7 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{AlCl}_3(\text{s}) = -704.2 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{NO}(\text{g}) = 90.2 \text{ kJ/mol}$ , and  $\Delta H_f^\circ \text{H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$ ;  
 $\Delta H_f^\circ \text{Al}(\text{s}) = 0 \text{ kJ/mol}$

**Required:**  $\Delta H_r^\circ$

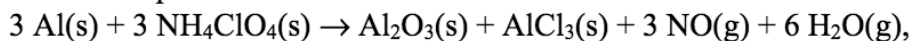
**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since Al(s) is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{Al}_2\text{O}_3(\text{s})} \times \Delta H_{\text{Al}_2\text{O}_3(\text{s})}^\circ + n_{\text{AlCl}_3(\text{s})} \times \Delta H_{\text{AlCl}_3(\text{s})}^\circ + n_{\text{NO}(\text{g})} \times \Delta H_{\text{NO}(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{g})} \times \Delta H_{\text{H}_2\text{O}(\text{g})}^\circ] - n_{\text{NH}_4\text{ClO}_4(\text{s})} \times \Delta H_{\text{NH}_4\text{ClO}_4(\text{s})}^\circ$$

**Solution:**

From the equation:



$$n_{\text{NH}_4\text{ClO}_4(\text{s})} = 3, n_{\text{Al}_2\text{O}_3(\text{s})} = 1, n_{\text{AlCl}_3(\text{s})} = 1, n_{\text{NO}(\text{g})} = 3, \text{ and } n_{\text{H}_2\text{O}(\text{g})} = 6$$

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{Al}_2\text{O}_3(\text{s})} \times \Delta H_{\text{Al}_2\text{O}_3(\text{s})}^\circ + n_{\text{AlCl}_3(\text{s})} \times \Delta H_{\text{AlCl}_3(\text{s})}^\circ + n_{\text{NO}(\text{g})} \times \Delta H_{\text{NO}(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{g})} \times \Delta H_{\text{H}_2\text{O}(\text{g})}^\circ] - n_{\text{NH}_4\text{ClO}_4(\text{s})} \times \Delta H_{\text{NH}_4\text{ClO}_4(\text{s})}^\circ \\ &= [(-1675.7 \text{ kJ}) + (-704.2 \text{ kJ}) + 3(90.2 \text{ kJ}) + 6(-241.8 \text{ kJ})] - 3(-295.8 \text{ kJ}) \\ &= [-1675.7 \text{ kJ} - 704.2 \text{ kJ} + 270.6 \text{ kJ} - 1450.8 \text{ kJ}] + 887.4 \text{ kJ} \end{aligned}$$

$$\Delta H_r^\circ = -2672.7 \text{ kJ}$$

**Statement:**  $\Delta H_r^\circ$  for the reaction of a mixture of aluminum and ammonium perchlorate is  $-2672.7 \text{ kJ}$ .

**5. Given:** from Table 1,  $\Delta H_{\text{f}^\circ}^{\text{NH}_3(\text{g})} = -45.9 \text{ kJ/mol}$  and  $\Delta H_{\text{f}^\circ}^{\text{HF}(\text{g})} = -273.3 \text{ kJ/mol}$ ;

$$\Delta H_{\text{f}^\circ}^{\text{N}_2(\text{g})} = -1196 \text{ kJ/mol}; \Delta H_{\text{f}^\circ}^{\text{N}_2(\text{g})} = 0 \text{ kJ/mol}; \Delta H_{\text{f}^\circ}^{\text{Cl}_2(\text{g})} = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_{\text{f}^\circ}^{\text{ClF}_3(\text{g})}$

**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since  $\text{N}_2(\text{g})$  and  $\text{Cl}_2(\text{g})$  are in their standard states, the equation can be written as:

$$\Delta H_r^\circ = n_{\text{HF}(\text{g})} \times \Delta H_{\text{HF}(\text{g})}^\circ - [n_{\text{ClF}_3(\text{g})} \times \Delta H_{\text{ClF}_3(\text{g})}^\circ + n_{\text{NH}_3(\text{g})} \times \Delta H_{\text{NH}_3(\text{g})}^\circ]$$

**Solution:**



$$n_{\text{ClF}_3(\text{g})} = 2, n_{\text{NH}_3(\text{g})} = 2, \text{ and } n_{\text{HF}(\text{g})} = 6$$

$$\begin{aligned} \Delta H_r^\circ &= n_{\text{HF}(\text{g})} \times \Delta H_{\text{HF}(\text{g})}^\circ - [n_{\text{ClF}_3(\text{g})} \times \Delta H_{\text{ClF}_3(\text{g})}^\circ + n_{\text{NH}_3(\text{g})} \times \Delta H_{\text{NH}_3(\text{g})}^\circ] \\ -1196 \text{ kJ} &= 6(-273.3 \text{ kJ}) - [2 \text{ mol} \times \Delta H_{\text{ClF}_3(\text{g})}^\circ + 2(-45.9 \text{ kJ})] \\ -1196 \text{ kJ} &= -1693.8 - 2 \text{ mol} \times \Delta H_{\text{ClF}_3(\text{g})}^\circ + 91.8 \text{ kJ} \end{aligned}$$

$$2 \text{ mol} \times \Delta H_f^\circ \text{ ClF}_3(\text{g}) = -352 \text{ kJ}$$

$$\Delta H_f^\circ \text{ ClF}_3(\text{g}) = \frac{-352 \text{ kJ}}{2 \text{ mol}}$$

$$\Delta H_f^\circ \text{ ClF}_3(\text{g}) = -176 \text{ kJ/mol}$$

**Statement:**  $\Delta H_f^\circ$  for  $\text{ClF}_3(\text{g})$  is  $-176 \text{ kJ/mol}$ .

**6. Given:** from Table 1,  $\Delta H_f^\circ \text{ H}_2\text{O}(\text{g}) = -241.8 \text{ kJ/mol}$  and  $\Delta H_f^\circ \text{ CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ ;

$\Delta H_f^\circ \text{ N}_2(\text{g}) = 0 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{ N}_2\text{H}_3\text{CH}_3(\text{l}) = 53 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{ N}_2\text{O}_4(\text{l}) = -20 \text{ kJ/mol}$

**Required:**  $\Delta H_r^\circ$

**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since  $\text{N}_2(\text{g})$  is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f CO}_2(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{g})} \times \Delta H_{\text{f H}_2\text{O}(\text{g})}^\circ] - [n_{\text{N}_2\text{H}_3\text{CH}_3(\text{l})} \times \Delta H_{\text{f N}_2\text{H}_3\text{CH}_3(\text{l})}^\circ + n_{\text{N}_2\text{O}_4(\text{l})} \times \Delta H_{\text{f N}_2\text{O}_4(\text{l})}^\circ]$$

**Solution:**

From the equation:  $4 \text{N}_2\text{H}_3\text{CH}_3(\text{l}) + 5 \text{N}_2\text{O}_4(\text{l}) \rightarrow 12 \text{H}_2\text{O}(\text{g}) + 9 \text{N}_2(\text{g}) + 4 \text{CO}_2(\text{g})$ ,

$$n_{\text{N}_2\text{H}_3\text{CH}_3(\text{l})} = 4, n_{\text{N}_2\text{O}_4(\text{l})} = 5, n_{\text{H}_2\text{O}(\text{g})} = 12, \text{ and } n_{\text{CO}_2(\text{g})} = 4$$

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f CO}_2(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{g})} \times \Delta H_{\text{f H}_2\text{O}(\text{g})}^\circ] - \\ &\quad [n_{\text{N}_2\text{H}_3\text{CH}_3(\text{l})} \times \Delta H_{\text{f N}_2\text{H}_3\text{CH}_3(\text{l})}^\circ + n_{\text{N}_2\text{O}_4(\text{l})} \times \Delta H_{\text{f N}_2\text{O}_4(\text{l})}^\circ] \\ &= [4(-393.5 \text{ kJ}) + 12(-241.8 \text{ kJ})] - [4(53 \text{ kJ}) + 5(-20 \text{ kJ})] \\ &= [-1574.0 \text{ kJ} - 2901.6 \text{ kJ}] - [212 \text{ kJ} - 100 \text{ kJ}] \end{aligned}$$

$$\Delta H_r^\circ = -4587.6 \text{ kJ}$$

**Statement:**  $\Delta H_r^\circ$  for the required reaction is  $-4587.6 \text{ kJ}$ .

**7. Given:** from Table 1,  $\Delta H_{\text{f CO}_2(\text{g})}^\circ = -393.5 \text{ kJ/mol}$  and  $\Delta H_{\text{f H}_2\text{O}(\text{l})}^\circ = -285.8 \text{ kJ/mol}$ ;

$$\Delta H_{\text{c}}^\circ = -1411.1 \text{ kJ/mol}; \Delta H_{\text{f O}_2(\text{g})}^\circ = 0 \text{ kJ/mol}$$

**Required:**  $\Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ$

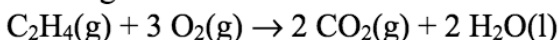
**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$

Since  $\text{O}_2(\text{g})$  is in its standard state, the equation can be written as:

$$\Delta H_{\text{c}}^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f CO}_2(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_{\text{f H}_2\text{O}(\text{l})}^\circ] - n_{\text{C}_2\text{H}_4(\text{g})} \times \Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ$$

**Solution:**

Step 1: Write a balanced chemical equation for the combustion of ethene gas so that ethene gas has a coefficient of 1.



Step 2: Substitute the appropriate values for standard enthalpy into the equation and solve.

$$\begin{aligned} \Delta H_{\text{c}}^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_{\text{f CO}_2(\text{g})}^\circ + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_{\text{f H}_2\text{O}(\text{l})}^\circ] - n_{\text{C}_2\text{H}_4(\text{g})} \times \Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ \\ -1411.1 \text{ kJ} &= [2(-393.5 \text{ kJ}) + 2(-285.8 \text{ kJ})] - \Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ \\ -1411.1 \text{ kJ} &= [-787 \text{ kJ} - 571.6 \text{ kJ}] - 1 \text{ mol} \times \Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ \end{aligned}$$

$$1 \text{ mol} \times \Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ = -1385.6 \text{ kJ} + 1411.1 \text{ kJ}$$

$$\Delta H_{\text{f C}_2\text{H}_4(\text{g})}^\circ = 52.5 \text{ kJ/mol}$$

**Statement:**  $\Delta H_{\text{f}}^\circ$  for ethene gas is  $52.5 \text{ kJ/mol}$ .



**8. Given:** from Table 1,  $\Delta H_f^\circ \text{C}_2\text{H}_5\text{OH(l)} = -235.2 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{CO}_2\text{(g)} = -393.5 \text{ kJ/mol}$ , and  $\Delta H_f^\circ \text{H}_2\text{O(l)} = -285.8 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{O}_2\text{(g)} = 0 \text{ kJ/mol}$

**Required:**  $\Delta H_r^\circ$  per gram of liquid ethanol

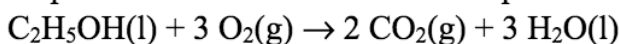
**Analysis:**  $\Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}} - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}$

Since  $\text{O}_2\text{(g)}$  is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_5\text{OH}(\text{l})} \times \Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\text{l})$$

**Solution:**

Step 1: Write a balanced chemical equation for the combustion of liquid ethanol.



Step 2: Substitute the appropriate values for standard enthalpy of formation into the equation and solve.

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{C}_2\text{H}_5\text{OH}(\text{l})} \times \Delta H_f^\circ \text{C}_2\text{H}_5\text{OH}(\text{l}) \\ &= [2(-393.5 \text{ kJ}) + 3(-285.8 \text{ kJ})] - (-235.2 \text{ kJ}) \\ &= -787.0 \text{ kJ} - 857.4 \text{ kJ} + 235.2 \text{ kJ} \end{aligned}$$

$$\Delta H_r^\circ = -1409.2 \text{ kJ}$$

Step 3: Convert to enthalpy per gram of liquid ethanol.

$$M_{\text{C}_2\text{H}_5\text{OH}(\text{l})} = 46.08 \text{ g/mol}$$

$$\begin{aligned} \Delta H_r^\circ \text{ per gram C}_2\text{H}_5\text{OH}(\text{l}) &= \frac{\Delta H_r^\circ \text{ per mole C}_2\text{H}_5\text{OH}(\text{l})}{M_{\text{C}_2\text{H}_5\text{OH}(\text{l})}} \\ &= \frac{-1409.2 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{46.08 \text{ g}} \end{aligned}$$

$$\Delta H_r^\circ \text{ per gram C}_2\text{H}_5\text{OH}(\text{l}) = -30.58 \text{ kJ/g}$$

**Statement:** The standard enthalpy of combustion per gram of liquid ethanol is  $-30.58 \text{ kJ/g}$ .

**9. (a) Given:** from Table 1,  $\Delta H_f^\circ \text{CH}_3\text{OH}(\text{l}) = -239.1 \text{ kJ/mol}$ ,  $\Delta H_f^\circ \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$ , and  $\Delta H_f^\circ \text{H}_2\text{O}(\text{l}) = -285.8 \text{ kJ/mol}$ ;  $\Delta H_f^\circ \text{O}_2(\text{g}) = 0 \text{ kJ/mol}$

**Required:**  $\Delta H_r^\circ$  per gram of liquid methanol

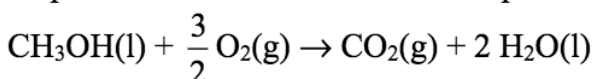
$$\text{Analysis: } \Delta H_r^\circ = \sum n_{\text{products}} \times \Delta H_{\text{products}}^\circ - \sum n_{\text{reactants}} \times \Delta H_{\text{reactants}}^\circ$$

Since  $\text{O}_2(\text{g})$  is in its standard state, the equation can be written as:

$$\Delta H_r^\circ = [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{CH}_3\text{OH}(\text{l})} \times \Delta H_f^\circ \text{CH}_3\text{OH}(\text{l})$$

**Solution:**

Step 1: Write a balanced chemical equation for the combustion of liquid ethanol.



Step 2: Substitute the appropriate values for standard enthalpy of formation into the equation and solve.

$$\begin{aligned} \Delta H_r^\circ &= [n_{\text{CO}_2(\text{g})} \times \Delta H_f^\circ \text{CO}_2(\text{g}) + n_{\text{H}_2\text{O}(\text{l})} \times \Delta H_f^\circ \text{H}_2\text{O}(\text{l})] - n_{\text{CH}_3\text{OH}(\text{l})} \times \Delta H_f^\circ \text{CH}_3\text{OH}(\text{l}) \\ &= [(-393.5 \text{ kJ}) + 2(-285.8 \text{ kJ})] - (-239.1 \text{ kJ}) \\ &= -393.5 \text{ kJ} - 571.6 \text{ kJ} + 239.1 \text{ kJ} \end{aligned}$$

$$\Delta H_r^\circ = -726.0 \text{ kJ}$$

Step 3: Convert to enthalpy per gram of liquid ethanol.

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

$$\begin{aligned}\Delta H_{\text{r}}^{\circ} \text{ per gram CH}_3\text{OH(l)} &= \frac{\Delta H_{\text{r}}^{\circ} \text{ per mole CH}_3\text{OH(l)}}{M_{\text{CH}_3\text{OH(l)}}} \\ &= \frac{-726.0 \text{ kJ}}{1 \text{ mol}} \times \frac{1 \text{ mol}}{32.05 \text{ g}}\end{aligned}$$

$$\Delta H_{\text{r}}^{\circ} \text{ per gram CH}_3\text{OH(l)} = -22.65 \text{ kJ/g}$$

**Statement:** The standard enthalpy of combustion per gram of liquid methanol is  $-22.65$  kJ/g.

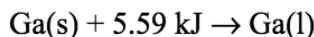
**(b)** The standard enthalpy of combustion per gram of liquid ethanol is  $-30.58$  kJ and that of liquid methanol is  $-22.65$  kJ. For 1 gram of the liquid fuel, liquid ethanol releases 7.93 kJ more energy than liquid methanol.

**(c)** The fuel that would be the most convenient source of energy for a vehicle is ethanol because for the mass of fuel burned, liquid ethanol releases 7.93 kJ/g more energy than methanol. That means a lighter mass of fuel would be needed to produce the same amount of energy.

**10. (a)** A thermochemical equation for the melting of gallium is:



or



**(b)** Since the standard state of gallium is solid,  $\Delta H_{\text{f}}^{\circ} \text{ Ga(s)} = 0 \text{ kJ/mol}$ .

From the thermochemical equation, the enthalpy of formation of liquid gallium can be calculated.

$$\Delta H_{\text{r}}^{\circ} = \sum n_{\text{Ga(l)}} \times \Delta H_{\text{Ga(l)}} - \sum n_{\text{Ga(s)}} \times \Delta H_{\text{Ga(s)}}$$

$$5.59 \text{ kJ/mol} = \Delta H_{\text{f}}^{\circ} \text{ Ga(l)}$$

The enthalpy of formation of liquid gallium is  $\Delta H_{\text{f}}^{\circ} \text{ Ga(l)} = 5.59 \text{ kJ/mol}$ , which is different from that of solid gallium.