Chemistry 30

20 max

# Unit 1: Energy Changes in Chemical Reactions

## Assignment 3: Sections 2-4 to 2-7

1. When iron rusts in air, the following reaction occurs:

   $4\ Fe_{(s)} + 3\ O_{2\ (g)} \rightarrow 2\ Fe_2O_{3\ (s)}$  $\Delta H = -1643\ kJ$

   What is the heat of formation of $Fe_2O_3$?

   **A heat of formation reaction describes the *production* of *one mole* of the compound *from its elements.* In the reaction given above, *two moles* of the compound are formed from the elements. Therefore,**

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$$\Delta H_f = \frac{-1643}{2} = -821.5 kJ$$

2. The equation for the decomposition of mercury(II) oxide is as follows:

   $2\ HgO_{(s)} + 182\ kJ \rightarrow 2\ Hg_{(l)} + O_{2\ (g)}$

   Determine $\Delta H_f$ for $HgO_{(s)}$?

   **The reaction shown in the question is the *reverse* of a formation reaction. As well, it involves *two* moles of the compound HgO, whereas heat of formation is concerned with *one* mole of the compound.**

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**The equation needs to be reversed, which will change the sign of ΔH from +182 kJ to -182 kJ. As well, moles need to be changed to represent *one* mole, not two.**

**Therefore, $\Delta H_f$ =** $\frac{-182}{2} = -91 kJ$

3. Using a table of thermochemical data, calculate ΔH° for the following reactions.

   **Look up heats of formation from the table of thermochemical data. Multiply by coefficients from the balanced equation. Find the total sum of the reactants, and of the products. Then use the formula**

$$\mathbf{\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}}$$

| | | | | | |
|---|---|---|---|---|---|
| a. | CO (g) | + | ½ $O_2$ (g) | $\rightarrow$ | $CO_2$ (g) |
| | -110.5 | + | ½ (0) | | -393.5 |
| | | -110.5 | | | -393.5 |

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$$\mathbf{\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}}$$

**ΔH = -393.5 – (-110.5) = <u>- 283 kJ</u>**

| b. | $CH_4$ (g) | + | 2 $O_2$ (g) | → | $CO_2$ (g) | + | 2 $H_2O$(g) |
|---|---|---|---|---|---|---|---|
| | -74.8 | + | 2 (0) | | -393.5 | + | 2(-241.8) |
| | | -74.8 | | | | -877.1 | |

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$$\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$$

**ΔH = -877.1 – (-74.8) = <u>- 802.3 kJ</u>**

| c. | $CS_2$(g) | + | 2 $H_2O$(l) | → | $CO_2$ (g) | + | 2 $H_2S$(g) |
|---|---|---|---|---|---|---|---|
| | + 117.4 | + | 2 (-285.8) | | -393.5 | + | 2 (-20.6) |
| | | -454.2 | | | | -434.7 | |

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$$\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$$

**ΔH = -434.7 – (-454.2) = <u>+ 19.5 kJ</u>**

4. Calculate ΔH° for the process:

   $2\ Al(s) + Fe_2O_3\ (s) \rightarrow Al_2O_3\ (s) + 2\ Fe(s)$

   Given that $\Delta H_f°$ of $Fe_2O_3$ = -813.0 kJ/mole and $\Delta H_f°$ of $Al_2O_3$ is -1,655.0 kJ/mol

| 2 Al(s) | + | $Fe_2O_3$ (s) | → | $Al_2O_3$ (s) | + | 2 Fe(s) |
|---|---|---|---|---|---|---|
| 2 (0) | + | -813.0 | | -1655.0 | + | 2 (0) |

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$$\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$$

**ΔH = - 1655.0 – (- 813.0) = <u>- 842.0 kJ</u>**

5. The standard heats of formation for $CH_4$, $CHCl_3$ and HCl are -74.8, -132, -92.0 kJ/mole, respectively. Use this information to calculate the heat of reaction for the following reaction:

   $CH_4 + 3\ Cl_2 \rightarrow CHCl_3 + 3\ HCl$

| $CH_4$ | + | 3 $Cl_2$ | → | $CHCl_3$ | + | 3 HCl |
|---|---|---|---|---|---|---|
| - 74.8 | + | 3 (0) | | - 132.0 | + | 3 (-92.0) |
| | -74.8 | | | | -408.0 | |

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$$\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}$$

**ΔH = - 408.0 – ( - 74.8 ) = <u>- 333.2 kJ</u>**

6. The standard heat of formation, $\Delta H_f$, for $C_2H_4$ is +52.3 kJ/mol. If $C_2H_4$ (ethylene) reacts with $H_2$ to produce $C_2H_6$ (ethane) according to the following equation:

$C_2H_4 + H_2 \rightarrow C_2H_6$   $\Delta H = -137$ kJ

what is the heat of formation of $C_2H_6$?

(**Hint**: Solve this question without looking up $\Delta H_f°$ for $C_2H_6$ in the Table of Thermochemical Data, although you may want to in order to check your answer. Solve this using the standard formula, $\Delta H_{reaction} = \Sigma\Delta H_{products} - \Sigma\Delta_{reactants}$. However, this time you know $\Delta H_{reaction}$ and need to solve for a substance on the product side of the equation.)

| | | | | |
|---|---|---|---|---|
| $C_2H_4$ | + | $H_2$ | $\rightarrow$ | $C_2H_6$ |
| 52.3 | + | 0 | | ***x*** |

$$\mathbf{\Delta H = \Sigma\Delta H_{products} - \Sigma\Delta H_{reactants}}$$

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**-137 = *x* – 52.3**

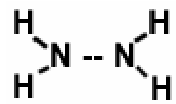
**-137 + 52.3 = *x* – 52.3 + 52.3**

**-84.7 kJ = x = $\Delta H_{C2H6}$**

7. Using bond enthalpies, calculate $\Delta H$ for the following reaction:

$N_2 + 2\ H_2 \rightarrow N_2H_4$

The structural formulas for all reaction participants are shown here:

$N \equiv N \ + \ 2\ H — H \ \rightarrow$ 

**You need to refer to the table of bond enthalpies. Determine how many of each type of chemical bond are broken (on the reactant side of the equation) and how many bonds of each type are formed (on the product side of the equation).**

**Bonds Broken (energy required):**

| | | | |
|---|---|---|---|
| **1** $N \equiv N$ | $1 \times 941$ | **=** | 941 kJ |
| 2 H – H | $2 \times 436$ | **=** | 872 kJ |
| **TOTAL** | | | 1813 kJ |

**Bonds Formed (energy released):**

| | | | |
|---|---|---|---|
| **4 N – H** | $4 \times 391$ | **=** | 1564 kJ |
| **1 N – N** | $1 \times 163$ | **=** | 163 kJ |
| **TOTAL** | | | 1727 kJ |

**The net energy change (heat of reaction) is the difference between how much energy is required to break bonds and how much energy is released when bonds form:**

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**$\Delta H = \Sigma$(energy required) - $\Sigma$(energy released)**

**$\Delta H$ = 1813 – 1727 = <u>+86 kJ</u>**

**There is a net requirement of 86 kJ in this endothermic reaction**

*Note: Solving for $\Delta H$ using bond enthalpies will not always agree completely with values from a Table of Thermochemical Data due to difficulties in accurately measuring bond enthalpies.*

8. The energy from the combustion of hydrazine, $N_2H_4$, is used to power rockets into space in the reaction:

$N_2H_4\ (g) + O_2\ (g) \rightarrow N_2\ (g) + 2\ H_2O\ (l) \qquad \Delta H^\circ = -627.6\ kJ$

How many kilograms of hydrazine would be necessary to produce $1.0 \times 10^8$ kJ of energy?

Hint: **One mole** of $N_2H_4$ produces 627.6 kJ of energy. How many moles (and then grams) are required to produce $1.0 \times 10^8$ kJ of energy?

**Use unit analysis to set up and solve this equation. Since the question concerns mass, you will need to find the molar mass of $N_2H_4$ as part of the solution:**

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**2 N = 2 × 14.0 = 28.0**
**4 H = 4 × 1.0 = 4.0**

**molar mass of $N_2H_4$ = 32.0 g/mol**

$$kg = \frac{1kg}{1{,}000g} \times \frac{32.0g}{mol} \times \frac{1.0 \times 10^8 kJ}{1} \times \frac{1mol}{627.6kJ} = 5.1 \times 10^3 kg$$