Ch.6 - *Thermochemistry*

Ch.6.1: **The Nature of Energy**

**Energy**: An object’s capacity to perform work or produce heat

**Work**: force acting over a distance \( W = F \times d \).

**Heat**: energy transferred between objects because of temperature difference.

**Potential Energy**: Energy due to position or composition (chemical bonds).

**Kinetic Energy**: Energy due to the motion of the object \( KE = \frac{1}{2}mv^2 \)

**Law of Conservation of Energy**: Energy can neither be created nor destroyed, but can be converted between forms

**The First Law of Thermodynamics**: The total energy content of the universe is constant
$\Delta E = q + W$  
$\Delta E$ = change in internal energy of a system

$q =$ heat flowing into or out of the system

$+q :$ energy is absorbed from the surroundings (*endothermic*)

$-q :$ energy is released to the surroundings (*exothermic*)

$W =$ work done by, or on, the system

$+w :$ work is done on the system by the surroundings

$-w :$ work is done by the system on the surroundings

Question 1: Determine the change in the internal energy for a system undergoing an endothermic process in which 15.6 kJ of heat flows and where 1.4 kJ of work is done on the system.

$\Delta E = (+15.6) + (+1.4) \Rightarrow \Delta E = + 17.0 \text{ kJ}$
Work, Pressure, and Volume

**Expansion**

\[ +\Delta V \text{ (increase)} \]

- \( w \) results

- \( E_{\text{system}} \) decreases

Work is done by the system on the surroundings

**Compression**

\[ -\Delta V \text{ (decrease)} \]

+ \( w \) results

+ \( E_{\text{system}} \) increases

Work is done on the system by the surroundings

Q2: Calculate the work resulting from the expansion of a gas from 46 L to 64 L at a constant external pressure of 15 atm.

\[ w = -(15)(64 - 46) = -270 \text{ atm-L} \]

\[ w = -27351 \text{ J} \]
Question 3: Which of the following performs more work? (1 L-atm = 101.3 J)
(a) A gas expanding against a pressure of 2.0 atm from 1.0 L to 4.0 L
(b) A gas expanding against a pressure of 3.0 atm from 1.0 L to 3.0 L

\[ w_a = -(2.0)(4.0 - 1.0) = -6.0 \text{ atm-L} = -607.8 \text{ J} \]

\[ w_b = -(3.0)(3.0 - 1.0) = -6.0 \text{ atm-L} = -607.8 \text{ J} \]

Question 4: Determine the sign of \( \Delta E \) for each of the following processes with the listed conditions:
(a) An endothermic process that performs work on the surroundings.
   a.1) work > heat \( \Rightarrow \) \( \Delta E \) : (+) or, (-) \[ \Delta E = +q + (-w) \]
   a.2) work < heat \( \Rightarrow \) \( \Delta E \) : (+) or, (-)
(b) Work is done on a gas and the process is exothermic.
   b.1) work > heat \( \Rightarrow \) \( \Delta E \) : (+) or, (-) \[ \Delta E = -q + (+w) \]
   b.2) work < heat \( \Rightarrow \) \( \Delta E \) : (+) or, (-)
**Energy Change in Chemical Processes**

*Chemical Reactions:* Energy is absorbed to break chemical bonds & energy is released when new bonds are formed.

**Exothermic Reactions:**

Reactions in which energy flows out of the system and to the surroundings as the reaction proceeds.

Much more energy is released during the creation of new chemical bonds in the products than is used in breaking the old bonds in the reactants.

![Energy change diagram](image)

\[
\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + \text{Heat}
\]
Energy Change in Chemical Processes

**Endothermic Reactions**

Reactions in which energy flows from the surroundings into the system as the reaction proceeds.

Much more energy is used for breaking the chemical bonds in the reactants than it is released during the creation of the new bonds in the products.

\[
+ q_{\text{system}} - q_{\text{surroundings}}
\]

\[
\text{N}_2 + \text{O}_2 \xrightarrow{\text{heat}} 2\text{NO}
\]
Question 5: You strike a match and it burns.
a) Explain the energy transfers of this scenario using the terms exothermic, endothermic, system, surrounding, chemical potential energy, and kinetic energy in your discussion.

b) Sketch an energy level diagram for the process.

Substance on the tip of the match has chemical potential energy stored in its chemical bonds. When friction increases the kinetic energy of the atoms causing those bonds to break, the substance, which is the system, burns releasing heat energy to the surrounding air. This process is exothermic because more energy is released during combustion than it is absorbed during the breaking of the initial bonds.

Question 6: Is the freezing of water an endothermic or exothermic process? Explain.

The freezing process is exothermic because the water molecules have to lose energy to the surroundings in order for them to change from a liquid state (high internal kinetic energy!) to a solid state (lower kinetic energy!).
Question 7: Classify each process as exothermic or endothermic. Explain.

a) **Your hand** gets cold when you touch ice.
   
   Exothermic because your hand loses thermal energy to the ice. Heat energy always flow from the hot to the cold object.

b) The **ice** gets warmer when you touch it.
   
   Endothermic because the ice is absorbing energy from your hand.

c) **Water** boils in a kettle being heated on a stove.
   
   Endothermic because the water is absorbing energy from the material that the kettle is made of.

d) **Water** vapor condenses on a cold pipe.
   
   Exothermic because the water vapor is releasing energy to the pipe and cooling off to the point at which vapor molecules change to a liquid state.

e) **Ice cream** melts.
   
   Endothermic because the ice cream is absorbing energy from the surrounding and the ice is changing into liquid state.
Question 8: For each of the following, define a system and its surroundings and give the direction of energy transfer (exo- or endo-thermic).

a) Methane is burning in a Bunsen burner in a laboratory.
   Methane is the system; air the surroundings; exothermic.

b) Water drops, sitting on your skin after swimming, evaporate.
   Water is the system; skin is the surroundings; endothermic because liquid water absorbs energy from your skin in order to change to vapor.

c) Two chemicals mixing in a beaker give off heat.
   Chemicals is the system; beaker is the surroundings; exothermic because the chemicals release heat to the beaker which gets warmer.
Question 9: Hydrogen gas and oxygen gas react violently to form water.
a) Which is lower in energy: a mixture of hydrogen and oxygen gases, or water? 
b) Sketch an energy level diagram for this reaction.

Water is lower in energy because when hydrogen and oxygen react, heat energy is released: exothermic process.

Question 10: Gas $A_2$ reacts with gas $B_2$ react to form gas $AB$. The bond energy of $AB$ is much greater than the bond energy of $A_2$ or $B_2$.
a) Is the formation of $AB$ endothermic or exothermic?
b) Which is lower in energy, the reactants or the products?

Energy is absorbed to break the initial bonds and released when new bonds are formed. If the bond energy of $A-B$ is greater than the bond energy of $A-A$ and $B-B$, much more energy is released forming $AB$ than absorbed to break the bonds in $A_2$ and $B_2 \rightarrow$ exothermic process.

In an exothermic process, the products ($AB$) are lower in energy.
Ch. 6.2: Enthalpy of Reaction and Calorimetry

**Enthalpy** ($H$)

- $H = E + PV$
- System at constant pressure: $\Delta H = \Delta E + P \Delta V$
- The heat at constant pressure $q_p$ can be calculated from:
  - $\Delta E = q_p + W = q_p - P \Delta V$
  - $q_p = \Delta E + P \Delta V \Rightarrow q_p = \Delta H$
- $\Delta H =$ heat transfer to/from system at constant pressure
- Chemical Rxns.: $\Delta H = H_{products} - H_{reactants}$
  - $H_{products} > H_{reactants} \Rightarrow +\Delta H \Rightarrow$ Endothermic rxn.
  - $H_{products} < H_{reactants} \Rightarrow -\Delta H \Rightarrow$ Exothermic rxn.
**Change in Enthalpy in Chemical Reactions**

**Problem 1**: Methane (CH\(_4\)) is burned at constant pressure, releasing 890 kJ/mol of energy as heat.

a) Write the balanced thermochemical equation for the reaction.

b) Calculate the change in the enthalpy for a process in which 5.8g of methane is burned at constant pressure.

\[
\begin{align*}
\text{(a)} & \quad \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + 890 \text{ kJ/mol} \quad (\Delta H = -890 \text{ kJ/mol}) \\
\text{(b)} & \quad \Delta H_{\text{orxn}} = 5.8 \text{ g CH}_4 \left(\frac{1 \text{ mol}}{16.04 \text{ g}}\right) \left(-\frac{890 \text{ kJ}}{1 \text{ mol}}\right) = -321.8 \approx -320 \text{ kJ}
\end{align*}
\]

**Problem 2**: The combustion of ethanol (C\(_2\)H\(_5\)OH) releases 1200 kJ/mol.

a) Write the balanced thermochemical equation for the reaction.

b) If we need to generate 6200 kJ of energy, how many grams of ethanol do we need to burn?

\[
\begin{align*}
\text{(a)} & \quad \text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} + 1200 \text{ kJ/mol} \quad (\Delta H = -1200 \text{ kJ/mol}) \\
\text{(b)} & \quad X \text{ grams C}_2\text{H}_5\text{OH} = -6200 \text{ kJ} \left(\frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{-1200 \text{ kJ}}\right) \left(\frac{46.07 \text{ g}}{1 \text{ mol}}\right) = 240 \text{ g C}_2\text{H}_5\text{OH}
\end{align*}
\]
Calorimetry

- **Calorimeter**: Instrument used for measuring the amount of heat absorbed or released during a physical change or chemical reaction.
- Substance releasing or absorbing the heat of reaction is usually water.
- Heat is determined by the change in temperature of a known quantity of water

\[
q = s \cdot m \cdot \Delta T
\]

- **Specific heat capacity**: amount of heat required to raise the temperature of 1.0g of a substance by 1.0°C [J/g-°C]
- ** Heat of Reaction**: \(\Delta H_{\text{Rxn}} = -(q_{\text{water}} + q_{\text{calorimeter}})\)
- \(q_{\text{calorimeter}} = C_{\text{calorimeter}} \cdot \Delta T\)
- **C_{\text{calorimeter}}**: calorimeter heat capacity [J/°C]
- **Molar Heat of Reaction**: \(\Delta H/mol = \Delta H_{\text{Rxn}} / \text{moles of reactant/product}\)
Problem 3: The specific heat of graphite is 0.71 J/g°C. Calculate the energy needed to raise the temperature of 75 kg of graphite from 294 K to 348 K.

\[ Q = s \cdot m \cdot \Delta T \]
\[ Q = (0.71 \text{ J/g·°C}) (75 \times 10^3 \text{ g}) (348 - 294) = 2875500 \text{ J} = 2.9 \times 10^6 \text{ J} \]

Problem 4: A 46.2 g sample of copper is heated to 95.4°C and then placed in a calorimeter containing 75.0 g of water at 19.6°C. The final temperature of both the water and the copper is 21.8°C. What is the specific heat of copper?
Water specific heat capacity is 4.184 J/g·°C
(Hint: At final equilibrium temperature, all the heat energy lost by the metal was absorbed by the water!)

\[ - Q_{\text{Cu lost}} = + Q_{\text{water gain}} \]
\[ - (s_{\text{Cu}} \cdot m_{\text{Cu}} \cdot \Delta T_{\text{Cu}}) = + (s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}}) \]
\[ - (s_{\text{Cu}}) (46.2 \text{ g Cu}) (21.8 - 95.4) = (4.184 \text{ J/g·°C}) (75.0 \text{ g water}) (21.8 - 19.6) \]
\[ - (s_{\text{Cu}}) (-3400.32) = 690.26 \Rightarrow s_{\text{Cu}} = 0.203 \text{ J/g·°C} \]
Problem 5: For the equation: \( \text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{NaOH(aq)} + \frac{1}{2} \text{H}_2(\text{aq}) \)

When 5.0 grams of sodium is placed into 100. mL of water, the temperature of the solution increases by 7.0°C. Calculate \( \Delta H_{\text{rxn}}/\text{mol of Na} \).

Assume density and specific heat capacity of the solution to be same as water.

\[
\Delta H_{\text{rxn}} = -Q_{\text{water}} = -(100 \text{ g})(4.184 \text{ J/g-°C})(7.0 \text{ °C}) = -2929 \text{ J}
\]

\[
\text{# mol Na} = \frac{5.0 \text{ g}}{22.99 \text{ g-mol}^{-1}} = 0.217 \text{ mol Na}
\]

\[
\Delta H_{\text{rxn}} = \frac{-2929 \text{ J}}{0.217 \text{ mol Na}} = -13.5 \text{ kJ/mol Na}
\]

Problem 6: For the equation: \( \text{Ca(s)} + 2\text{HCl (aq)} \rightarrow \text{CaCl}_2 (\text{aq}) + \text{H}_2 (\text{g}) \)

If 2.0 grams of calcium is placed into 75 mL of 1.0M HCl solution at 18°C and the final temperature of the system reaches 23°C, calculate \( \Delta H_{\text{rxn}}/\text{mole Ca} \).

Assume density and specific heat capacity of the solution to be same as water.

\[
\Delta H_{\text{rxn}} = -Q_{\text{water}} = -(75 \text{ g})(4.184 \text{ J/g-°C})(5 \text{ °C}) = -1569 \text{ J}
\]

\[
\text{# mol Ca} = \frac{2.0 \text{ g}}{40.08 \text{ g-mol}^{-1}} = 0.0499 \text{ mol Ca}
\]

\[
\Delta H_{\text{rxn}} = \frac{-1569 \text{ J}}{0.0499 \text{ mol Ca}} = -31.4 \text{ kJ/mol Ca}
\]
Bomb Calorimetry

- Constant volume calorimeter is called a bomb calorimeter.
- Material is put in a container with pure oxygen.
- Wires are used to start the combustion.
- The reaction container is put into a container of water.
- The heat capacity of the calorimeter is known and tested.

Heat of Reaction: \( \Delta H_{\text{Rxn}} = - (q_{\text{water}} + q_{\text{calorimeter}}) \)

\[ q_{\text{calorimeter}} = C_{\text{calorimeter}} \cdot \Delta T \]

\( C_{\text{calorimeter}} \) : heat capacity of the calorimeter [J/°C]
Problem 7: You burn 0.300 g of C in an excess of O\(_2\) in a bomb calorimeter. The temperature of the calorimeter which contains 775 grams of water increases from 25°C to 27.38°C. The heat capacity of the calorimeter is 893 J/K. What is \(\Delta H_{\text{rxn}}\) per mole of carbon?

\[
\Delta H_{\text{rxn}} = - (q_{\text{water}} + q_{\text{bomb}}) =
\]

\[
q_{\text{water}} = s_{\text{water}} \ m_{\text{water}} \ \Delta T_{\text{water}} \quad \& \quad q_{\text{bomb}} = C_{\text{calorimeter}} \ \Delta T_{\text{bomb}}
\]

\[
q_{\text{water}} = (4.184 \ J/g\text{-}°C)(775 \ g)(2.38 \ °C) = 7717.39 \ J
\]

\[
q_{\text{bomb}} = (893 \ J/K)(2.38 \ K)] = 2125.34 \ J
\]

\[
Q_{\text{rxn}} = - (7717.39 + 2125.34) = -9842.73 = -9840 \ J = -9.84 \ kJ
\]

\# moles C = 0.300 g C *(1 mole/12.01 g) = 0.0250 moles C

\[
\Delta H_{\text{rxn}} / \text{mol C} = -9.84 / 0.0250 = -394 \ kJ/mol C
\]
Prbl. 8: A 1.5886g sample of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) was ignited in a bomb calorimeter. The temperature increased by 3.682°C. The heat capacity of the calorimeter was 3.562kJ/°C, and the calorimeter contained 1.000 kg of water. Write the balanced thermochemical equation and find the molar heat of reaction (i.e., kJ/mole of glucose).

(a) $\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{O}_2 \rightarrow 6\text{CO}_2 + 6\text{H}_2\text{O}$

(b) $\Delta H_{rxn} = -(q_{\text{water}} + q_{\text{bomb}}) =$

$q_{\text{water}} = s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}}$  \&  $q_{\text{bomb}} = C_{\text{calorimeter}} \cdot \Delta T_{\text{bomb}}$

$q_{\text{water}} = (4.184 \text{ J/g} \cdot \text{°C})(1000 \text{ g})(3.682 \text{ °C}) = 15405.49 \text{ J}$

$q_{\text{bomb}} = (3562 \text{ J/K})(3.682 \text{ K})] = 13115.28 \text{ J}$

$Q_{rxn} = -(15405.49 + 13115.28) = -28520.77 = -28520 \text{ J} = -28.52 \text{ kJ}$

$\# \text{ moles } \text{C}_6\text{H}_{12}\text{O}_6 = 1.5886 \text{ g} \text{C}_6\text{H}_{12}\text{O}_6 \times (1 \text{ mole/180.2 g}) = 0.008816 \text{ mol C}_6\text{H}_{12}\text{O}_6$

$\Delta H_{rxn} / \text{ mol C}_6\text{H}_{12}\text{O}_6 = -28.52 / 0.008816 = -3235 \text{ kJ/mol C}_6\text{H}_{12}\text{O}_6$
Problem 9: A 100.0 g sample of water at 90°C is added to a 100.0 g sample of water at 10°C. The final temperature of the water is:

a) Between 50°C and 90°C  

b) 50°C  

c) Between 10°C and 50°C

Calculate the final temperature of the water.

\[- Q_{\text{water lost}} = + Q_{\text{water gain}} \Rightarrow - (s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}}) = + (s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}})\]

\[- (T_f - 90) = + (T_f - 10)\]

\[- T_f + 90 = T_f - 10 \Rightarrow -2T_f = -100 \Rightarrow T_f = 50°C\]

Problem 10: A 100.0 g sample of water at 90°C is added to a 500.0 g sample of water at 10°C. The final temperature of the water is:

a) Between 50°C and 90°C  

b) 50°C  

c) Between 10°C and 50°C

Calculate the final temperature of the water.

\[- Q_{\text{water lost}} = + Q_{\text{water gain}} \Rightarrow - (s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}}) = + (s_{\text{water}} \cdot m_{\text{water}} \cdot \Delta T_{\text{water}})\]

\[- (100.0) (T_f - 90) = + (500.0) (T_f - 10)\]

\[- 100.0 T_f + 9000 = 500.0 T_f - 5000 \Rightarrow -600.0T_f = -14000 \Rightarrow T_f = 23.3°C\]
Problem 11: You have a Styrofoam cup with 50.0 g of water at 10°C. You add a 50.0 g iron ball at 90°C to the water. The final temperature of the water is:

a) Between 50°C and 90°C  

b) 50°C  

c) Between 10°C and 50°C

Calculate the final temperature of the water.

Specific heat of iron = 0.4456 J/g-°C

\[ - Q_{Fe\ lost} = + Q_{water\ gain} \quad \Rightarrow \quad -(s_{Fe} \ m_{Fe} \ \Delta T_{Fe}) = + (s_{water} \ m_{water} \ \Delta T_{water}) \]

\[ - (0.4456) (50.0) (T_f - 90) = (4.184) (50.0) (T_f - 10) \]

\[ - 0.4456 \ T_f + 40.104 = 4.184 \ T_f - 41.84 \]

\[ - 4.630 \ T_f = -81.94 \quad \Rightarrow \quad T_f = 17.7 \ °C \]
Ch. 6.3: Hess’s Law

- Enthalpy is a state function.
- Enthalpy of reaction does not depend on number of steps or pathway.
- From a particular set of reactants to a particular set of products, $\Delta H_{Rxn}$ is the same whether the reaction takes place in one step or in a series of steps.
- We can add simpler, known reaction equations to come up with the desired final equation.
- Two important rules:
  a) If one step reaction is reversed the sign of $\Delta H_{Rxn}$ is also reversed.
  b) If a step reaction is multiplied by a factor, so is $\Delta H_{Rxn}$

\[
\begin{align*}
N_2 + 2O_2 & \rightarrow 2NO_2 \quad \Delta H_{Rxn} = ? \\
N_2 + O_2 & \rightarrow 2NO \quad \Delta H_{Rxn} = 180kJ \\
2NO + O_2 & \rightarrow 2NO_2 \quad \Delta H_{Rxn} = -112kJ
\end{align*}
\]
**Standard Enthalpy of Reaction** ($\Delta H^o_{Rxn}$)

- The $\Delta H_{Rxn}$ at standard conditions (25°C, 1 atm, 1 M solutions)

**Hess’s Law Problem 1:** Given the following known reactions,

\[ C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \quad \Delta H^o = -1300. \text{kJ} \]

\[ C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H^o = -394 \text{kJ} \]

\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H^o = -286 \text{kJ} \]

Find $\Delta H^o_{Rxn}$ for the following reaction: \[ 2C(s) + H_2(g) \rightarrow C_2H_2(g) \]

\[ 2CO_2(g) + H_2O(l) \rightarrow C_2H_2(g) + \frac{5}{2}O_2(g) \quad \Delta H^o = +1300. \text{kJ} \]

\[ 2C(s) + 2O_2(g) \rightarrow 2CO_2(g) \quad \Delta H^o = -788 \text{kJ} \]

\[ H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H^o = -286 \text{kJ} \]

\[ 2C(s) + H_2(g) \rightarrow C_2H_2(g) \quad \Delta H^o = +226 \text{kJ} \]
Hess’s Law Problem 1: Given the following known reactions,

\[ \text{O}_2(g) + \text{H}_2(g) \rightarrow 2\text{OH}(g) \quad \Delta H^\circ = +77.9\text{kJ} \]

\[ \text{O}_2(g) \rightarrow 2\text{O}(g) \quad \Delta H^\circ = +495\text{kJ} \]

\[ \text{H}_2(g) \rightarrow 2\text{H}(g) \quad \Delta H^\circ = +435.9\text{kJ} \]

Find \( \Delta H^\circ_{\text{Rxn}} \) for the following reaction: \( \text{O}(g) + \text{H}(g) \rightarrow \text{OH}(g) \)

\[ \frac{1}{2}\text{O}_2(g) + \frac{1}{2}\text{H}_2(g) \rightarrow \text{OH}(g) \quad \Delta H^\circ = +38.95\text{kJ} \]

\[ \text{O}(g) \rightarrow \frac{1}{2}\text{O}_2(g) \quad \Delta H^\circ = -247.5\text{kJ} \]

\[ \text{H}(g) \rightarrow \frac{1}{2}\text{H}_2(g) \quad \Delta H^\circ = -217.95\text{kJ} \]

\[ \text{O}(g) + \text{H}(g) \rightarrow \text{OH}(g) \quad \Delta H^\circ = -426.5\text{kJ} \]
Ch. 6.4: *Standard Enthalpies of Formation* ($\Delta H_f^\circ$)

- $\Delta H_f^\circ$ of a compound: amount of heat needed for the formation of 1 mole of a compound from its elements in their standard states.

- Standard state conditions are: 1 atm, 1M concentration and 25°C

- $\Delta H_f^\circ = 0$ for all substances in their elemental states: H$_2$(g), O$_2$(g), N$_2$(g), Na(s), K(s), Fe(s), Al(s), etc.

*Standard Enthalpy of Reaction* ($\Delta H_{Rxn}^\circ$)

$\Delta H_{Rxn}^\circ$ can be calculated from enthalpies of formation of each reactant and product.

$$\Delta H_{Rxn}^\circ = [\Sigma n_p \Delta H_f^\circ(\text{products})] - [\Sigma n_r \Delta H_f^\circ(\text{reactants})]$$
Problem 1: For each of the substances below, write a balanced equation showing the formation of 1 mole of the compound from its elements in their standard states. Look up the value for $\Delta H_f^\circ$ in the Appendix of the textbook.

a) $\text{Al}_2\text{O}_3$

b) $\text{Mg(OH)}_2$

c) $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

d) $\text{NaHCO}_3$
Problem 2: Use the standard enthalpy of formation of the compounds to calculate the standard heat of reaction ($\Delta H^\circ_{Rxn}$):

$$C_2H_5OH (l) + 3O_2 (g) \rightarrow 2CO_2 (g) + 3H_2O (g)$$

Problem 3: Nitroglycerin is a powerful explosive that decomposes by the following equation while giving off 4200 kJ of heat.

$$C_3H_5(NO_3)_3 \rightarrow 3/2N_2 (g) + \frac{1}{4}O_2 (g) + 3CO_2 (g) + 5/2H_2O (g)$$

Calculate the $\Delta H_f^\circ$ of nitroglycerin.
Using the following data, calculate the heat of formation of the compound ICl (g) from its elements at 25°C.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>ΔH° (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( Cl_2 (g) \rightarrow 2Cl (g) )</td>
<td>242.3</td>
</tr>
<tr>
<td>( I_2 (g) \rightarrow 2I (g) )</td>
<td>151.0</td>
</tr>
<tr>
<td>( ICl (g) \rightarrow I (g) + Cl (g) )</td>
<td>211.3</td>
</tr>
<tr>
<td>( I_2 (s) \rightarrow I_2 (g) )</td>
<td>62.8</td>
</tr>
</tbody>
</table>
**Bond Energies and Enthalpy of Reaction ($\Delta H_{Rxn}$)**

- Using published value of atomic bond energies, we can find the change in the enthalpy for a reaction.
- In order to break existing bonds, energy must be added to the compounds.
- When new bonds are formed, energy is released to the surrounding.
- If we add up all the energy it took to break the bonds, and subtract all the energy released forming the bonds we get the $\Delta H_{Rxn}$.

$$\Delta H_{Rxn} = \left[ \sum D(\text{bonds broken}) \right] - \left[ \sum D(\text{bonds formed}) \right]$$

$D$: bond energy/mole [kJ/mol]
Problem 1: The bond energy (kJ) for H-H, F-F, and H-F are 432, 154 and 565 kJ/mol respectively, calculate the enthalpy of the reaction: \( \text{H}_2 + \text{F}_2 \rightarrow 2\text{HF} \)

Problem 2: Calculate the heat of reaction (\( \Delta H_{\text{Rxn}} \)) for: \( 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \)
Bond energies: H-H: 432 kJ/mol, O═O: 495 kJ/mol, O-H: 467 kJ/mol
**Bond Energies and Enthalpy of Reaction ($\Delta H_{Rxn}$)**

Problem 3: Calculate the enthalpy of reaction ($\Delta H_{Rxn}$):
\[
\text{CH}_4 + 3\text{Cl}_2 \rightarrow \text{CHCl}_3 + 3\text{HCl}
\]
Bond energies: C-H: 413 kJ/mol, Cl-Cl: 239 kJ/mol, C-Cl: 339 kJ/mol, H-Cl: 427 kJ/mol

Problem 4: Calculate $\Delta H_{Rxn}$ for the following reaction given the bond energies below:
\[
\text{C}_2\text{H}_4 + \text{H}_2 \rightarrow \text{C}_2\text{H}_6
\]
Bond energies: C=C: 614 kJ/mol, C-H: 413 kJ/mol, H-H: 432 kJ/mol, C-C: 347 kJ/mol
Use the value of the bond energies to find the energy for this reaction:
Is this an endothermic or exothermic reaction?

\[ 2\text{CH}_2 = \text{CHCH}_3 + 2\text{NH}_3 + \text{O}_2 \rightarrow 2\text{CH}_2 = \text{CHC} \equiv \text{N} + 6\text{H}_2\text{O} \]