

## 5.1 Thermochemistry: Energy Changes (p. 283 onwards)

The ability to harness energy is one of the most important scientific endeavors. Think in your life all of the energy devices you use and how you get your own personal energy. Below fill out the table and also create an annotated diagram to show all of the interactions, conversions of energy.

Concept	Definition	Everyday example of the concept in action.
Thermochemistry		
Energy		
Work		
Potential energy		
Kinetic energy		
Thermal energy		
Heat		
Temperature		
Chemical system		
Surroundings		
Open system		
Closed system		

SCH4U

Unit #3 Energy

Name: \_\_\_\_\_

Isolated system.		
------------------	--	--

**Introduction to Thermochemistry.**

1. Recall the first law of thermodynamics, or the Law of Conservation of Energy.

A. How does our use of petroleum reserves connect to that law?	B. Pose 3-5 questions other than A you would like to ask about the petroleum industry, future supply, market, or effects of its use.

-Describe 2-3 examples of energy transformations involving chemical reactions – use technologies or biological examples.

--

In this unit, we will learn how to predict energy outputs or requirements of different chemical reactions.

- Think of 2-3 examples of how that kind of knowledge could be useful to chemists or people working in industry. To whom could that kind of knowledge be useful to other than straight-up chemistry professors and researchers?

Recall the events in a chemical reaction, like the exothermic combustion of methane. Use a balanced chemical equation, potential energy graphs, and energy ideas related to bond breaking and bond formation. Connect to the idea of why atoms react.

Use a concept map or other visual organizer and examples, illustrations, and definitions to show how the following ideas are related:

- Energy, work, potential energy, kinetic energy, nuclear potential, electric potential, chemical potential, thermal energy, heat, temperature, thermal energy, exothermic, endothermic and a couple of other terms you'd like to connect to these as well

## 5.2 Calorimetry and Enthalpy

In this section you will define the following terms: specific heat capacity, calorimetry and calorimeter (p. 292).

Do Practice p. 297, # 1 and 2.

Define Enthalpy and enthalpy change ( $\Delta H$ ), p. 298. Do Practice #1 and 2.

Do Practice p. 304, #1 and 2, 3 and 4.

## 5.3 Bond Energies

- Define: bond dissociation energy (p. 307).
- Define the formula that is used to calculate enthalpy change for a reaction (p. 308).

$(\Delta H) =$

Go through Tutorial #1 and write out each step. You do not solve the problem because it is already solved but you need to analyze the steps. Ensure you understand what GRAS means

Do Practice p. 312, # 1-4.

## 5.4 Hess's Law

1. Read page 314, define what are a single-step and two step chemical equations.
2. Define Hess's Law (p. 314)
3. What are the rules for enthalpy changes (p. 315)
4. Analyze Tutorial # 1, Sample Problem #1.
  - a. What are the steps for solving Hess's law problems?
  - b. How were elimination of products and reactants carried out so the only remaining substance was C(graphite) and C (diamond) ?
5. Do Practice Problems, p. 317, # 1 and 2.

- The heat of a reaction can be calculated using the standard enthalpies of formation for the reactants and products.
- A standard enthalpy of formation is the quantity of energy absorbed or released by the formation of **one mole** of a compound from its elements in their standard states.
- The symbol for standard enthalpy of formation is  $\Delta H_f^\circ$
- The change in enthalpy is calculated by determining the difference in total molar enthalpy between the products and reactants:

$$\Delta H_r = \sum n\Delta H_{f \text{ products}} - \sum n\Delta H_{f \text{ reactants}}$$

**Note:** 'n' represents the total number of moles of each product/reactant

$\Delta H_f^\circ$  = **molar enthalpy of formation** = the quantity of energy that is absorbed or released when one mole of a *compound* is formed directly from its elements in their standard states.

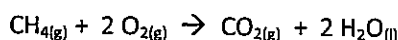
By definition, the enthalpy of formation of an element in its standard state is zero (even diatomic elements).

**NOTES:**

- Consult Table 1 on page 320 for Standard Enthalpies of Formation of Common Compounds
- To determine standard states of elements, consult the periodic table. If the element is not in its standard state, it will have a  $\Delta H_f^\circ$  value.
- Before summing the  $\Delta H_f^\circ$  values in the formula, they must be multiplied first by the coefficients from the **BALANCED** chemical equation, n.
- Be sure the **STATES** match. For example, water can be found as (s), (l) or (g) and each will have a different  $\Delta H_f^\circ$  value.

**Example 1:** Calculate  $\Delta H$  for the combustion of methane,  $\text{CH}_4$ , (in kJ/mol) using standard enthalpies of formation.

**Step 1:** Write a balanced chemical equation.



**Step 2:** Use the formula to calculate  $\Delta H$

$$\begin{aligned} \Delta H_r &= \sum n\Delta H_{f \text{ products}} - \sum n\Delta H_{f \text{ reactants}} \\ &= (\Delta H_f^\circ \text{CO}_{2(g)} + 2 \Delta H_f^\circ \text{H}_2\text{O}_{(l)}) - (\Delta H_f^\circ \text{CH}_{4(g)} + 2 \Delta H_f^\circ \text{O}_{2(g)}) \\ &= [(1 \text{ mol } (-393.5 \text{ kJ/mol}) + (2 \text{ mol } (-285.8 \text{ kJ/mol})))] \\ &\quad - \\ &\quad [(1 \text{ mol } (-74.4 \text{ kJ/mol}) + (2 \text{ mol } (0 \text{ kJ/mol})))] \\ &= (-965.1 \text{ kJ}) - (-74.4 \text{ kJ}) \\ &= -890.7 \text{ kJ} \end{aligned}$$