

**Unit #3, Chapter 5 Outline**  
**Energy and Change: Thermochemistry**

Lesson	Topics Covered	Reference	Homework Questions and Assignments
1	<p><b>Introduction to Thermochemistry</b></p> <ul style="list-style-type: none"> <li>• definitions: kinetic and potential energy</li> <li>• thermal kinetic energy, temperature and heat (know what these terms mean)</li> </ul>	<p><b>Note:</b> Intro to Thermochemistry</p> <p><b>Handout:</b> Understanding Thermal Energy, Temp and Heat</p> <p><b>Text:</b> 220-222</p>	<ol style="list-style-type: none"> <li>1. Define energy, thermal kinetic energy, temperature, and heat.</li> <li>2. Clearly distinguish between heat and temperature.</li> <li>3. Which has more thermal energy: 40 L of boiling water or Lake Ontario in the winter? Explain.</li> </ol>
2	<p><b>Enthalpy and Energy Changes in Chemical Reactions</b></p> <ul style="list-style-type: none"> <li>• definition of enthalpy, units</li> <li>• factors affecting enthalpy (number and type of bonds)</li> <li>• system vs. surroundings</li> </ul> <p><b>Exothermic Reactions</b></p> <ul style="list-style-type: none"> <li>• enthalpy level diagram</li> <li>• thermochemical equations</li> <li>• <math>\Delta H</math> is negative</li> </ul> <p><b>Endothermic Reactions</b></p> <ul style="list-style-type: none"> <li>• enthalpy level diagram</li> <li>• thermochemical equations</li> <li>• <math>\Delta H</math> is positive</li> </ul> <p><b>Comparing the energy changes of chemical, physical and nuclear changes</b></p>	<p><b>Handout:</b> Enthalpy and Energy Changes in Chemical Reactions</p> <p>Pages 222-224</p> <p>Pages 226-229</p>	<p>Page 226 Q 1 a, b 2 a, b 3 a, c 4 a</p> <p>Read pages 226 – 228: Heat Changes and Physical Changes</p> <p>Read page 229, the first 4 paragraphs of Energy and Nuclear Reactions</p> <p>Page 232: Look at Concept Organizer</p>
3 & 4	<p><b>Thermochemical Equations</b></p> <ul style="list-style-type: none"> <li>• enthalpy changes (<math>\Delta H</math>) in chemical reactions</li> <li>• stoichiometric calculations with <math>\Delta H</math></li> </ul> <p><b>Introduction to Calorimetry</b></p> <ul style="list-style-type: none"> <li>• measuring <math>\Delta H</math> by experiment</li> <li>• heat lost by system equals heat gained by surroundings (and visa versa)</li> <li>• <math>-\Delta H = Q</math> for systems at constant pressure (rxns that do not involve gases)</li> <li>• <math>Q = m c \Delta T</math></li> <li>• definition of specific heat capacity (c)</li> <li>• for water, <math>c = 4.184 \text{ J/g } ^\circ\text{C}</math> and <math>1.00 \text{ mL H}_2\text{O} = 1.00 \text{ g}</math></li> </ul> <p><b>Demo of Calorimetry Experiment</b></p> <ul style="list-style-type: none"> <li>• sample calculations for lab #4</li> </ul>	<p>Pages 224-225</p> <p>Pages 234-238</p> <p><b>Handout:</b> Lab #4: Finding <math>\Delta H</math> for Physical and Chemical Changes</p>	<p>Page 226 Q 1 c, d 2 c 3 b 4 b,c</p> <p>Read Pages 234 to 238, but ignore heat capacity on page 235, we will only use specific heat capacity</p> <p>Page 235 Q 7, 6, 8, 5 (in this order)</p> <p>Page 238-239 Q 10, 9, 11, 12 (in this order)</p> <p>Preparation for Lab #4: Read through notes on assumptions in calorimetry on page 239 (under the problems) Prepare a data table to record your observations.</p>

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Lesson	Topics Covered	Reference	Homework Questions and Assignments
5	<b>Lab #4: Calorimetry Lab</b>		1. Complete lab calculations and lab write-up. Lab report due:  2. Page 238-239 Q9 – 12 (*remember to use the total volume of the solutions to find the mass of the surroundings) * to find # of moles (n) for a solution, $n = CV$ , where C is molar conc'n in mol/L
6	<b>State Functions:</b> <ul style="list-style-type: none"> <li>definitions and examples</li> </ul> <b>Calculating <math>\Delta H</math> using Hess's Law of Heat Summation</b> <ul style="list-style-type: none"> <li>definition</li> <li>rules for using Hess's Law</li> <li>example calculations</li> </ul>	Pages 243-246  <b>Note and then Handout:</b> Hess's Law of Heat Summation	1. Read pages 243 – 246 2. Answer questions 13 – 16 on page 247 3. Complete additional Hess's Law problems on Homework handout 4. Read through Lab #5: Using Hess's Law to Calculate the Heat of Combustion ( $\Delta H$ ) of Magnesium in preparation for the lab tomorrow. Begin your lab report, including an observation table as directed in the lab handout.
7	<b>Lab #5: Hess's Law</b>	<b>Lab #5: Hess's Law</b>	Complete lab calculations and lab write-up. Lab report due:
8	<b>Calculating <math>\Delta H</math> using Standard Molar Enthalpies of Formation</b> <ul style="list-style-type: none"> <li>standard molar enthalpies of formation (<math>\Delta H_f^\circ</math>)</li> <li>standard states: metals, H, O, Br, F, N, Cl</li> <li>elements and elements with allotropes: S, P, C, O</li> <li>writing formation equations</li> <li>sample calculations</li> </ul>	<b>Handout:</b> Calculating Enthalpy Change ( $\Delta H$ ) using Standard Enthalpies of Formation	1. Read pages 250 – 253. 2. On page 251, do questions 17 – 20. On page 254, do questions 21 – 24. For questions 21 and 22, you do not need to calculate $\Delta H^\circ$ , just rearrange the equations to arrive at the target equation. 3. Do questions 3,4,5 on page 255 and question 8 on page 262
9	<b>Calculating <math>\Delta H</math> using Bond Energies</b> <ul style="list-style-type: none"> <li>definition of bond energy</li> <li><math>\Delta H = \Sigma \text{BE of reactants} - \Sigma \text{BE of products}</math></li> <li>sample calculations</li> </ul>	<b>Handout:</b> Calculating $\Delta H$ using Bond Energies	1. Complete problems on Handout 2. (Average bond dissociation energies or "bond energies" are given on page 599) 3. Create a summary note about the four ways we have learned to calculate $\Delta H$ . Be able to recognize the clues to each problem type. 4. Begin Chapter 5 review on web-page 5. When you are ready, try the practice test which is on the web-page. 6. If you really feel you need it, you can do the additional review, page 263-264: 1 - 6, 8 – 15, 16 a, b
	<b>Unit Test on Chapter 5</b>		Date:

## Unit 3, Lesson 01: Understanding Thermal Energy, Temperature and Heat

A model may help illustrate the difference between thermal energy, temperature and heat:





Imagine cars travelling on the 401. Each car represents a single molecule (particle).

- The cars are vibrating, their tires are rotating and the whole car is translating (moving from one place to another).
- All of the cars are moving at different speeds. The average speed of all of the cars represents **temperature**.
- When traffic is moving smoothly, the average speed is about 110 km/h. This would represent a high temperature.
- When there is a traffic jam, the average speed is about 20 km/h. This would represent a low temperature.
- If there are a lot of cars on the 401, even if they are moving slowly, there will be a lot of total movement. The total amount of movement represents **thermal energy**. The more cars there are and the faster they are moving, the greater the thermal energy (or total movement).
- If a fast-moving car hits a slow-moving car, it will make the slower car speed up. Some of the kinetic energy of the fast car is transferred to the slow car. This represents **heat** (energy transfer).
- If the cars are not moving and have their engines turned off- they have no kinetic energy. This represents absolute zero, when all particle motion has stopped and thermal energy equals zero.

### In summary:

1. **Thermal energy** is the total amount of movement (kinetic energy) of the particles in a system. The total amount of thermal energy depends on how many particles there are (mass or # of moles) and their temperature (how fast, on average, they are moving).
2. **Temperature** (T) is the measure of the average speed (kinetic energy) of the particles in the substance.
3. **Heat** (Q) refers to the transfer of thermal energy from one object to another, for example, by the collisions of their particles.

Compare the thermal energy and temperature of the following (not to scale:)

<p>A large Tim Hortons Iced Cap</p> 	<p>A small Tim Hortons coffee</p> 	<p>A large Tim Hortons coffee</p> 	<p>A bathtub full of warm water</p> 
<p>Temp: 0 °C Speed of particles:</p>	<p>Temp: 87 °C Speed of particles:</p>	<p>Temp: 87 °C Speed of particles:</p>	<p>Temp: 27 °C Speed of particles:</p>
<p>Mass: 500 g (ish)</p>	<p>Mass: 250 g (ish)</p>	<p>Mass: 500 g (ish)</p>	<p>Mass: 225000 g (ish)</p>
<p>Total Thermal energy:</p>	<p>Total Thermal energy:</p>	<p>Total Thermal energy:</p>	<p>Total Thermal energy:</p>

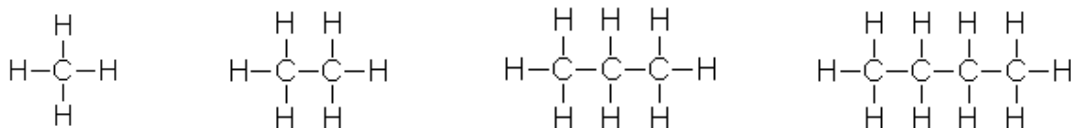
## Unit 3, Lesson 02: Enthalpy and Energy Changes in Chemical Reactions

There is chemical potential energy “stored” in the \_\_\_\_\_ of a molecule.

The total amount of chemical potential energy in a molecule depends, in part, on:

### 1. The number of bonds:

- the \_\_\_\_\_ a molecule has, the more \_\_\_\_\_ it can store



eg. ethane has \_\_\_\_\_ chemical potential energy than methane because ethane has \_\_\_\_\_

eg. butane has \_\_\_\_\_ chemical potential energy than propane because butane has \_\_\_\_\_

### 2. The type and length of the bonds:

- bond length is defined as the \_\_\_\_\_, measured in \_\_\_\_\_, where 1 pm = \_\_\_\_\_
- in general, shorter bonds have \_\_\_\_\_ chemical potential energy
- bond lengths are reported on page \_\_\_\_\_ of your text (McGraw-Hill)

eg. C – H bond length is \_\_\_\_\_, the atoms are fairly \_\_\_\_\_ so chemical potential energy is \_\_\_\_\_

eg. O – H bond length is \_\_\_\_\_, the atoms are \_\_\_\_\_ so chemical potential energy is \_\_\_\_\_

The \_\_\_\_\_ amount of chemical potential energy in a molecule contributes to total energy content or \_\_\_\_\_ (\_\_\_\_) of the substance. Enthalpy includes all of the energy, both \_\_\_\_\_ and \_\_\_\_\_, of the particles in the substance and is reported in \_\_\_\_\_.

There is no way to measure the absolute enthalpy of a substance. Instead, we measure the \_\_\_\_\_ (\_\_\_\_) during a chemical reaction. It is the change in enthalpy that is responsible for energy changes that occur during a chemical reaction, such as \_\_\_\_\_ or \_\_\_\_\_ being given off.

To find the change in enthalpy (\_\_\_\_) during a chemical reaction, we need to consider both the \_\_\_\_\_ and \_\_\_\_\_, so we talk about the “\_\_\_\_\_”. In chemistry, a **system** refers to the \_\_\_\_\_ that we are studying. Everything outside of the system is called the \_\_\_\_\_. Energy can be transferred between the system and its surroundings as \_\_\_\_\_.

During a chemical reaction, the bonds between the atoms of the reactants are \_\_\_\_\_ and new bonds are \_\_\_\_\_ to create the \_\_\_\_\_. Because the \_\_\_\_\_ and \_\_\_\_\_ of bonds change, the \_\_\_\_\_ will also change.

There are two possibilities:

- the enthalpy of the products is \_\_\_\_\_ than the enthalpy of the reactants so energy will be \_\_\_\_\_ by the system and absorbed by the surroundings (an \_\_\_\_\_ reaction); or
- the enthalpy of the products is \_\_\_\_\_ than the enthalpy of the reactants so energy must be \_\_\_\_\_ by the system from the surroundings (an \_\_\_\_\_ reaction).

Let’s look at both situations:

## 1. Exothermic Reactions

Consider the combustion reaction:  $\text{CH}_4(\text{g}) + 2 \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{v})$

### Bond lengths:

_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
}		}	
Total length of bonds in <b>reactants</b> = _____		Total length of bonds in <b>products</b> = _____	

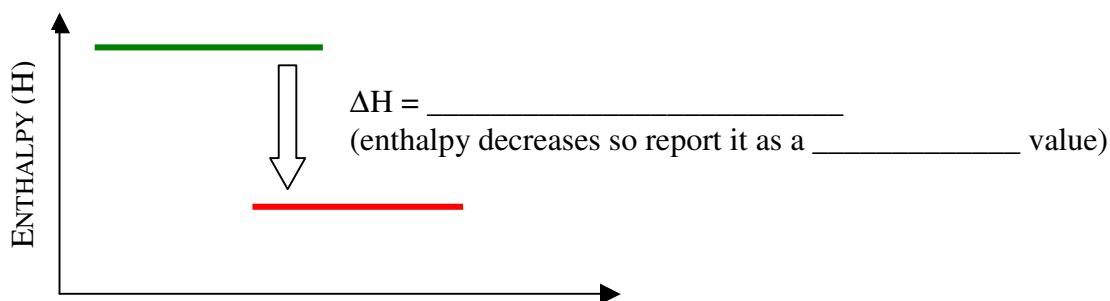
Overall, the atoms in the products are \_\_\_\_\_ than the atoms in the reactants, so the products have \_\_\_\_\_ chemical potential energy and \_\_\_\_\_ enthalpy. Energy is \_\_\_\_\_.

### What happens to the chemical potential energy that is released?

- It is converted to \_\_\_\_\_ and the particles begin to \_\_\_\_\_
- The \_\_\_\_\_ of the system goes up and heat is \_\_\_\_\_ to the surroundings
- Because heat “\_\_\_\_\_” or leaves the system, the reaction is said to be \_\_\_\_\_
- Energy can be considered to be a \_\_\_\_\_ of the reaction

We can represent the energy change of an **exothermic** reaction in three different ways:

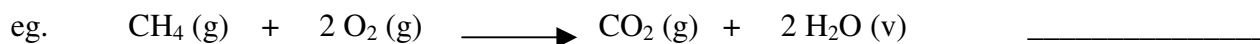
a) an enthalpy level diagram that shows the difference in the enthalpy of the reactants and products graphically:



b) a \_\_\_\_\_ equation that shows the amount of energy released as a \_\_\_\_\_



a) a balanced chemical equation that is followed by a separate \_\_\_\_\_



Because energy is \_\_\_\_\_ by an exothermic reaction, it is shown on the \_\_\_\_\_ side of the thermochemical equation. If written as a separate energy term, the \_\_\_\_\_ value of  $\Delta H$  indicates an \_\_\_\_\_ reaction.

## 2. Endothermic Reactions

Consider the reaction:  $2 \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2\text{O}(\text{g})$

### Bond lengths:

_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

Total length of bonds in **reactants** = \_\_\_\_\_

Total length of bonds in **products** = \_\_\_\_\_

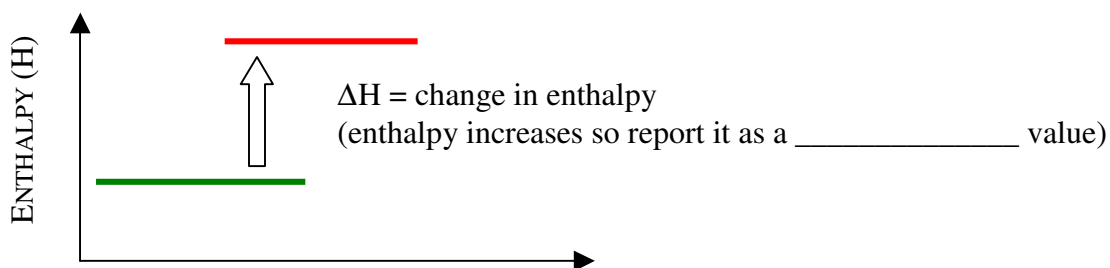
Overall, the atoms in the products are \_\_\_\_\_ than the atoms in the reactants, so the products have \_\_\_\_\_ chemical potential energy and \_\_\_\_\_ enthalpy. Energy is \_\_\_\_\_.

### Where did the extra chemical potential energy come from?

- It is \_\_\_\_\_ from the \_\_\_\_\_, so the temperature of the surroundings \_\_\_\_\_
- Because heat \_\_\_\_\_ the system, the reaction is said to be \_\_\_\_\_ (“endo” means \_\_\_\_\_ or \_\_\_\_\_)
- Energy can be considered to be a \_\_\_\_\_ for the reaction

We can represent the energy change of an **endothermic** reaction in three different ways:

- a) an enthalpy level diagram that shows the difference in the enthalpy of the reactants and products graphically:



- b) a \_\_\_\_\_ equation that shows the amount of energy absorbed as a \_\_\_\_\_  
eg. \_\_\_\_\_ +  $2 \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2\text{O}(\text{g})$

- c) a balanced chemical equation that is followed by a separate \_\_\_\_\_  
eg.  $2 \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2\text{O}(\text{g})$  \_\_\_\_\_

Because energy is \_\_\_\_\_ by an endothermic reaction, it is shown on the \_\_\_\_\_ of the thermochemical equation. If written as a separate energy term, the \_\_\_\_\_ value of  $\Delta H$  indicates an \_\_\_\_\_ reaction.

### Sample Calculation for Calorimetry: Using Q to Calculate $\Delta H$

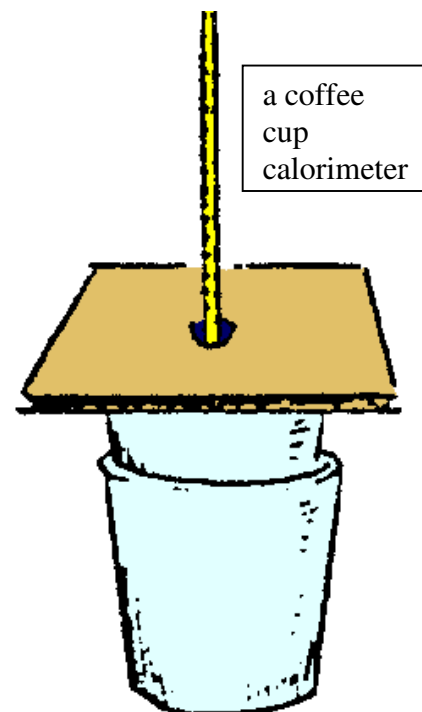
- the heat ( ) lost or gained during a chemical reaction at constant pressure is equal to the \_\_\_\_\_ ( ) for the reaction
- in an aqueous system, when a reaction produces energy (is \_\_\_\_\_), the released energy will be absorbed by the water and its temperature will \_\_\_\_\_
- in an aqueous system, when a reaction requires energy (is \_\_\_\_\_) the required energy will be absorbed from the water, and the water temperature will \_\_\_\_\_
- by measuring the temperature change of the water, we can calculate the amount of energy transferred (heat, Q) to or from the water
- to calculate Q, we use the equation:



where  $m$  = the \_\_\_\_\_ of whatever surrounds the reaction and changes temperature by providing or absorbing heat. In aqueous systems, it is \_\_\_\_\_.

$c$  = the specific heat capacity of whatever surrounds the reaction and changes temperature.  
= \_\_\_\_\_ for an aqueous system

$\Delta T$  = the change in temperature of whatever surrounds the reaction



The purpose of the calorimeter is to ensure that all of the energy lost or gained by the reaction is transferred to the water, so the temperature change of the water is entirely due to the reaction. The heat lost outside of the calorimeter will be negligible.

### Sample Calculation: The Molar Heat of Solution of NaOH in Water

#### Observations: Data Table for Calorimetry Calculation

Solute	Exact Mass of Solute (3 decimals)	Volume of Water ( $\pm 0.1$ mL)	Initial Temperature of Water ( $\pm 0.1$ °C)	Final Temperature of Water ( $\pm 0.1$ °C)	Specific Heat Capacity of Water ( $c = 4.184$ J/g°C)
NaOH (s)					

**Reaction (the system):**

**Calculations:**

1. Mass of water ( $m$ ):

2. Calculate the temperature change of water ( $\Delta T$ ):

3. Calculate the amount of heat ( $Q$ ) transferred to/from water:
  
4. Determine the sign for  $\Delta H$ :
  - when no gases are produced by a reaction (no energy is lost by the gases expanding): \_\_\_\_\_
  - if  $Q$  is positive (the temperature of the water \_\_\_\_\_), then  $\Delta H$  is \_\_\_\_\_
  - if  $Q$  is negative (the temperature of the water \_\_\_\_\_), then  $\Delta H$  is \_\_\_\_\_
  - so, when NaOH dissolves in water,  $\Delta H$  is \_\_\_\_\_
5. Calculate the amount of energy released when **one mole** of NaOH dissolves ( $\Delta H$ ):
  - the amount of heat ( $Q$ ) calculated in part 3 was released when \_\_\_\_\_ g of NaOH dissolved
  - convert to heat change per mole of NaOH:

When we do calorimetry experiments and calculations, we make several assumptions. The validity of these assumptions may affect the accuracy of the results and introduce some degree of error:

1. Assume that no heat is lost to, or gained from, the calorimeter or anything outside it. Because styrofoam is a good insulator and the chemical reactions occur relatively quickly, this is an acceptable assumption.
2. Assume that all of the energy released by the reaction is transferred as heat to the solution.
3. Assume that the specific heat capacity ( $c$ ) of a dilute aqueous solution is the same as the specific heat capacity of pure water. If dilute solutions are used ( $\leq 1.0\text{ M}$ ), this is an acceptable assumption.
4. Assume that the density ( $D$ ) of a dilute aqueous solution is the same as the density of pure water. If dilute solutions are used ( $\leq 1.0\text{ M}$ ), this is an acceptable assumption to calculate the mass of the solution.
5. If a reaction produces gases, assume that the energy lost by the expansion of the gas is negligible, so that  $Q = -\Delta H$ . The energy loss is usually small, so this is only a small source of error.
6. Assume that there is no transfer of mechanical energy to the water as heat when the solution is stirred. This is reasonable as long as the stirring is minimal and gentle.



### Unit 3, Lesson 06: Hess's Law of Heat Summation

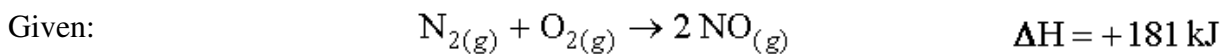
**Hess's Law** states that when a reaction can be expressed as the algebraic sum of two or more reactions, then the total enthalpy change is equal to the algebraic sum of the enthalpy changes for the intermediate reactions.

#### Rules for using Hess's Law

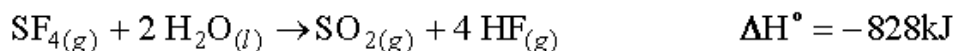
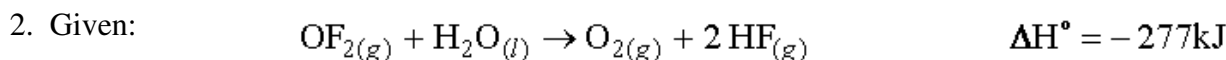
1. If the reaction is multiplied (or divided) by some factor,  $\Delta H$  must also be multiplied (or divided) by that same factor.
2. If the reaction is reversed (flipped), the sign of  $\Delta H$  must be changed from a negative to positive, or visa versa.

#### Examples:

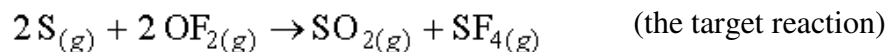
1. Calculate  $\Delta H$  for the reaction:  $\text{N}_{2(g)} + 2 \text{O}_{2(g)} \rightarrow 2 \text{NO}_{2(g)}$  (the "target" reaction)



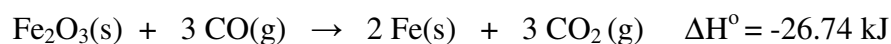
Adding these reactions, we get:



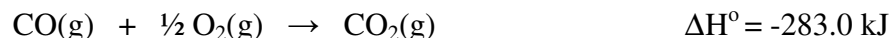
Calculate the value of  $\Delta H^\circ$  for the reaction:



3. Carbon monoxide (CO) is often used in metallurgy to remove oxygen from metal oxides and purify the free metal. The thermochemical equation for the reaction of CO with iron (III) oxide, Fe<sub>2</sub>O<sub>3</sub>, is:



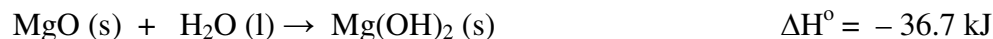
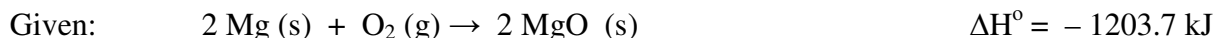
The equation for the combustion of CO (g) is:



Calculate the value of  $\Delta H^\circ$  for the reaction:



4. Calculate the  $\Delta H$  for:  $\text{Mg}(\text{s}) + \text{O}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{Mg}(\text{OH})_2(\text{s})$



**Final notes:**

- The state of the reactants and products is important because it affects enthalpy. If necessary, convert substances to the correct state using a thermochemical reaction and include the  $\Delta H$  value for the phase change (change of state) in your calculation. The text refers to H<sub>2</sub>O (v) as H<sub>2</sub>O (g).
- $\Delta H^\circ$  is often listed as just  $\Delta H$ . If you are given  $\Delta H$ , assume that it was measured at standard conditions

### Unit 3, Lesson 06: Hess's Law of Heat Summation

#### Homework:

1. Read pages 243 – 246 in McGraw-Hill
2. Answer questions 13 – 16 on page 247
3. Read through Lab #5: Using Hess's Law to Calculate  $\Delta H$  in preparation for the lab tomorrow. Begin your lab report, including an observation table as directed in the lab handout.

#### Hess's Law Problems:

4. Given:
- |  |                                    |
|--|------------------------------------|
| $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$   | $\Delta H^\circ = 175 \text{ kJ}$  |
| $\text{Ca}(\text{OH})_2(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{H}_2\text{O}(\text{l})$                                 | $\Delta H^\circ = 67 \text{ kJ}$   |
| $\text{Ca}(\text{OH})_2(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$ | $\Delta H^\circ = -198 \text{ kJ}$ |

Find the  $\Delta H^\circ$  for:  $\text{CaCO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$

5. Calculate the heat of reaction ( $\Delta H^\circ$ ) for:  $\text{N}_2\text{H}_4(\text{l}) + 2 \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{N}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{g})$

- Given:
- |  |                                    |
|--|------------------------------------|
| $\text{N}_2\text{H}_4(\text{l}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{NO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$ | $\Delta H^\circ = -466 \text{ kJ}$ |
| $\text{H}_2\text{O}(\text{l}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}_2(\text{l})$                   | $\Delta H^\circ = 98.0 \text{ kJ}$ |
| $\frac{1}{2} \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g})$                                    | $\Delta H^\circ = 34.0 \text{ kJ}$ |
| $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$  | $\Delta H^\circ = 44.0 \text{ kJ}$ |

6. Find the heat of reaction ( $\Delta H^\circ$ ) for:  $4 \text{P}(\text{s}) + 5 \text{O}_2(\text{g}) \rightarrow \text{P}_4\text{O}_{10}(\text{s})$

- Given:
- |   |                                      |
|---|--------------------------------------|
| $2 \text{PCl}_3(\text{l}) + \text{O}_2(\text{g}) \rightarrow 2 \text{POCl}_3(\text{l})$               | $\Delta H^\circ = -587.4 \text{ kJ}$ |
| $\text{P}_4\text{O}_{10}(\text{s}) + 6 \text{PCl}_5(\text{s}) \rightarrow 10 \text{POCl}_3(\text{l})$ | $\Delta H^\circ = -418.8 \text{ kJ}$ |
| $2 \text{P}(\text{s}) + 3 \text{Cl}_2(\text{g}) \rightarrow 2 \text{PCl}_3(\text{l})$                 | $\Delta H^\circ = -685.8 \text{ kJ}$ |
| $2 \text{P}(\text{s}) + 5 \text{Cl}_2(\text{g}) \rightarrow 2 \text{PCl}_5(\text{s})$                 | $\Delta H^\circ = -892.0 \text{ kJ}$ |

7. Given:
- |  |                                    |
|--|------------------------------------|
| $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{HCl}(\text{g})$                                | $\Delta H^\circ = -184 \text{ kJ}$ |
| $\text{HI}(\text{g}) \rightarrow \text{HI}(\text{aq})$   | $\Delta H^\circ = -80 \text{ kJ}$  |
| $\text{KOH}(\text{aq}) + \text{HCl}(\text{aq}) \rightarrow \text{KCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ | $\Delta H^\circ = -157 \text{ kJ}$ |
| $\text{KOH}(\text{aq}) + \text{HI}(\text{aq}) \rightarrow \text{KI}(\text{aq}) + \text{H}_2\text{O}(\text{l})$   | $\Delta H^\circ = -140 \text{ kJ}$ |
| $\text{HCl}(\text{g}) \rightarrow \text{HCl}(\text{aq})$   | $\Delta H^\circ = -75 \text{ kJ}$  |
| $2 \text{KI}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{KCl}(\text{aq}) + \text{I}_2(\text{s})$      | $\Delta H^\circ = -219 \text{ kJ}$ |

Calculate the heat of reaction ( $\Delta H^\circ$ ) for:  $\text{H}_2(\text{g}) + \text{I}_2(\text{s}) \rightarrow 2 \text{HI}(\text{g})$

#### Answers:

4. -90. kJ (the decimal indicates that the zero is significant, so 2 sig digs)
5. -642 kJ
6. -3271.2 kJ or -3271 kJ (both are acceptable answers for sig digs)
7. +11 kJ

### Unit 3, Lesson 08: Calculating Enthalpy Change ( $\Delta H$ ) using Standard Enthalpies of Formation

$\Delta H$  can be measured or calculated using:

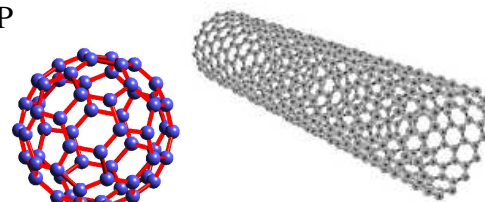
1. **Calorimetry data** for chemical reactions that occur at \_\_\_\_\_ (systems that do not include \_\_\_\_\_ as products or reactants). At constant pressure,  $-\Delta H = Q =$  \_\_\_\_\_.
2. **Hess's Law** when you know  $\Delta H$  values for other chemical reactions that can be added to give you the target chemical reaction.
3. **Standard Enthalpies of Formation ( $\Delta H^\circ_f$ )**
  - chemists have measured "standard enthalpies of formation" for many compounds ( see page 597)
  - defined as the amount of energy released or absorbed when \_\_\_\_\_ of a compound is formed directly from its \_\_\_\_\_, in their \_\_\_\_\_, at SATP.

The **standard state** for an element is its most stable or most common form at SATP.

**For metals:** the standard state is defined as neutral atoms eg. \_\_\_\_\_

**For non-metals:**

- a) Noble gases are present as neutral atoms in the gas state eg. \_\_\_\_\_
- b) "HOB rFINCl" elements: these elements are present as \_\_\_\_\_ molecules at SATP and their state is indicated on the Periodic Table eg. \_\_\_\_\_
- c) allotropes are non-metals that are found in more than one form at SATP
  - i) **carbon** can be found in its pure state as graphite, diamond, buckminsterfullerene (buckyballs), or even carbon nanotubes.  
The standard state of carbon is defined as \_\_\_\_\_
  - ii) **oxygen** can be found in its pure state as oxygen gas  $O_2$  (g) or ozone  $O_3$  (g). The standard state of oxygen at SATP is defined as \_\_\_\_\_
  - iii) **sulfur** can be found as S (s) or  $S_8$  (s). The standard state is \_\_\_\_\_
  - iv) **phosphorus** can be red phosphorus, P (s) or white phosphorus,  $P_4$  (s). The standard state is \_\_\_\_\_.



The standard heat of formation of an element in its standard state at SATP ( $\Delta H^\circ_f$ ) is \_\_\_\_\_.



eg. Write the formation equation for the following compounds. Include  $\Delta H^\circ_f$  for each reaction (on page 597):

- a) carbon dioxide, \_\_\_\_\_
- b) sodium chloride, \_\_\_\_\_
- c) solid tetraphosphorus decaoxide, \_\_\_\_\_
- d) ammonium nitrate, \_\_\_\_\_

We can calculate the  $\Delta H^\circ$  for a reaction from the standard heats of formation ( $\Delta H^\circ_f$ ) of the reactants and products, using Hess's Law.

eg. Calculate  $\Delta H^\circ$  for the combustion of one mole of ethyne using  $\Delta H^\circ_f$

**Step 1:** Write the target equation:

**Step 2:** Write the equations for the formation reactions of the reactants and products in the target equation. Write the values for  $\Delta H^\circ_f$ . You do NOT need to write equations for elements that are in their standard states because the  $\Delta H^\circ_f$  for these elements is \_\_\_\_\_.

**Step 3:** Rearrange the equations using Hess's Law to give the target equation. Add up the values for  $\Delta H^\circ_f$  to give the overall  $\Delta H^\circ$  for the target reaction.

If the values for  $\Delta H^\circ_f$  for all of the reactants and products is known, we can use a mathematical shortcut for the above:

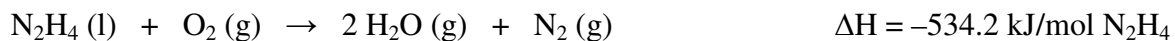
$$\Delta H^\circ \text{ for the target reaction} = \sum n (\Delta H^\circ_f \text{ of the products}) - \sum n (\Delta H^\circ_f \text{ of the reactants})$$

eg. Calculate  $\Delta H^\circ$  for the combustion of one mole of ethyne using  $\Delta H^\circ_f$

$$\Delta H^\circ \text{ for the target reaction} = \sum n (\Delta H^\circ_f \text{ of the products}) - \sum n (\Delta H^\circ_f \text{ of the reactants})$$

Similarly, if the  $\Delta H^\circ$  for a reaction is known (perhaps from calorimetry), the equation above can be rearranged to “work backwards” to calculate the heat of formation of either a reactant or a product.

eg. Hydrazine is used as rocket fuel for the space shuttle. Calculate the heat of formation ( $\Delta H^\circ_f$ ) of liquid hydrazine ( $\text{N}_2\text{H}_4$ ), given that the heat of combustion of one mole of hydrazine is  $-534.2 \text{ kJ/mol}$ :



eg. Calculate the heat of formation ( $\Delta H^\circ_f$ ) of propanone, given that the heat of combustion of one mole of propanone is  $-1816.5 \text{ kJ/mol}$ .

### Homework:

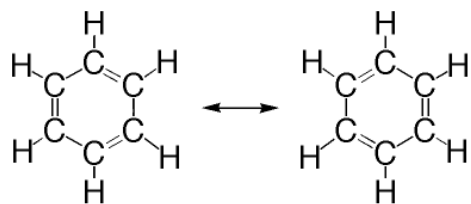
1. Read pages 250 – 253.
2. On page 251, do questions 17 – 20. On page 254, do questions 21 – 24. For questions 21 and 22, you do not need to calculate  $\Delta H^\circ$ , just rearrange the equations to arrive at the target equation.
3. Do questions 3,4,5 on page 255 and question 8 on page 262

### Unit 3, Lesson 09: Estimating $\Delta H$ Using Bond Energies

#### Homework:

1. Read pages 226 to 229 in McGraw-Hill, on physical and nuclear changes
2. Use bond energies to estimate the heat of reaction ( $\Delta H$ ) for the following:
  - a)  $C_5H_8 + 2 H_2 \rightarrow C_5H_{12}$  ( $C_5H_8$  is 1,3-pentadiene)
  - b)  $C_2H_5OH + 3 O_2 \rightarrow 2 CO_2 + 3 H_2O$
  - c)  $C_3H_4 + Cl_2 \rightarrow C_3H_4Cl_2$  ( $C_3H_4$  is cyclopropene)
  - d) the complete combustion of butane
  - e)  $2 NH_3 + 3 HOCl \rightarrow N_2 + 3 H_2O + 3 HCl$

3. The bonding in benzene is unusual. While the structure is often shown with alternating single and double bonds around the carbon ring, evidence shows that all six bonds are the same and are neither single nor double bonds.



The complete combustion of benzene produces 3260 kJ of energy per mole of benzene burned.

Use bond energies to “work backward” to calculate the average bond energy of the six carbon-carbon bonds in benzene. What does the value suggest about the “nature” of these bonds?

#### Answers:

- 2a)  $\Delta H = + 40$  or  $40$ . kJ
  - 2b)  $\Delta H = - 1391$  kJ
  - 2c)  $\Delta H = - 291$  kJ
  - 2d)  $\Delta H = - 2902$  kJ
  - 2e)  $\Delta H = - 780$ . kJ
3. The BE of the carbon-carbon bond in benzene is + 446 kJ. This is half-way in between the BE for C – C and C = C bonds, and supports the idea that the carbon-carbon bond in benzene is a resonance structure and is like a  $1 \frac{1}{2}$  bond)