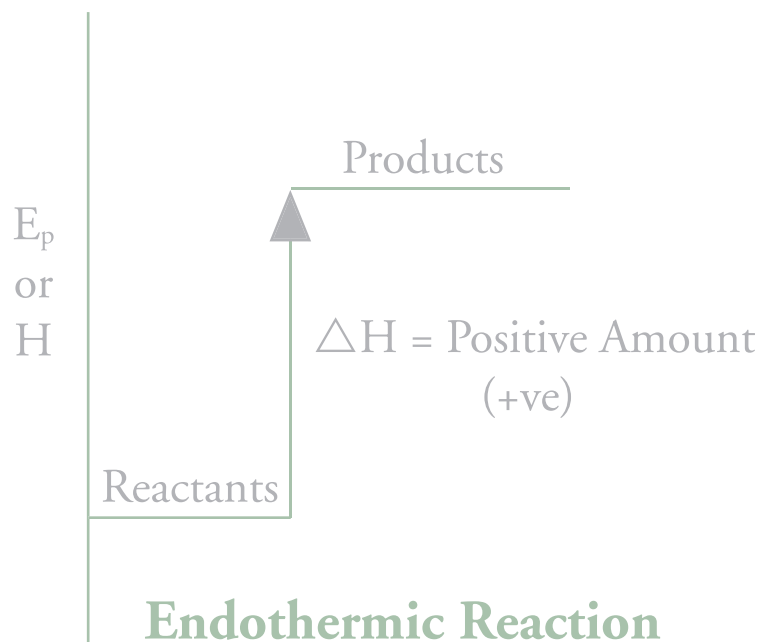
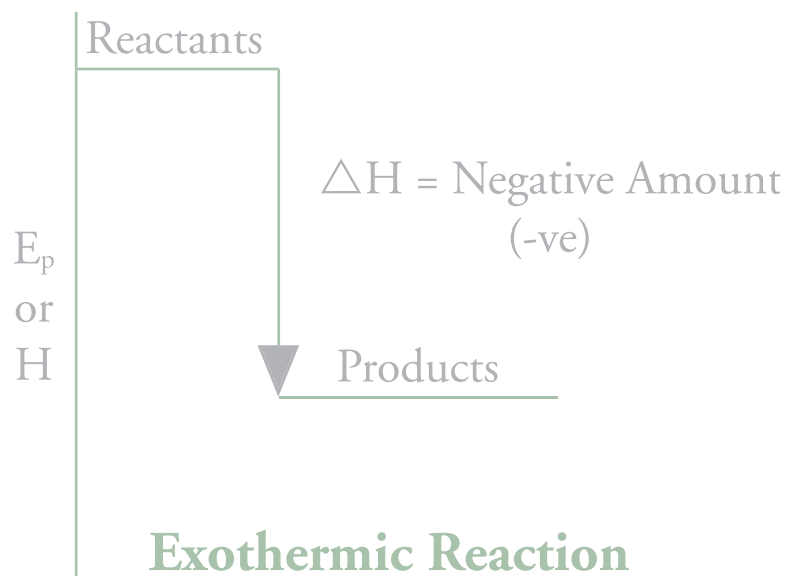


UNIT 1

THERMOCHEMISTRY



THERMOCHEMISTRY LEARNING OUTCOMES

Students will be expected to:

THERMOCHEMISTRY STSE

- ✓ analyse why scientific and technological activities take place in a variety individual and group settings (117-6)
- ✓ analyse from a variety of perspectives the risks and benefits to society and the environment by applying thermochemistry (118-2)
- ✓ distinguish between questions that can be answered using thermochemistry and those that cannot, and problems that can be solved by technology and those that cannot (118-8)
- ✓ compare the molar enthalpies of several combustion reactions involving organic compounds (324-7)
- ✓ write and balance chemical equations for combustion reactions of alkanes, including energy amounts (324-1)
- ✓ propose courses of action on social issues related to science and technology, taking into account an array of perspectives, including that of technology (118-10)

EXPERIMENTS WITH ENERGY CHANGES

- ✓ define endothermic reaction, exothermic reaction, specific heat, enthalpy, bond energy, heat of reaction, and molar enthalpy (324-2)
- ✓ calculate and compare the energy involved in a changes of state in chemical reactions (324-3)
- ✓ design a thermochemistry experiment identifying and controlling major variables (212-3)
- ✓ work co-operatively with team members to develop and carry out thermochemistry experiments (215-6)
- ✓ evaluate and select appropriate instruments for collecting evidence and appropriate processes for problem solving, inquiring (212-8)

- ✓ determine experimentally the changes in energy of various chemical reactions (324-6)
- ✓ analyse the knowledge and skills acquired in their study of thermochemistry to identify areas of further study related to science and technology (117-9)
- ✓ propose alternative solutions to solving energy problems and identify the potential strengths and weaknesses of each (214-15)

THERMOCHEMISTRY AND POTENTIAL ENERGY

- ✓ illustrate changes in energy of various chemical reactions, using potential energy diagrams (324-5)
- ✓ compile and display evidence and information on heats of formation in a variety of formats, including diagrams, flow charts, tables, and graphs (214-3)

BONDING AND HESS'S LAW

- ✓ calculate the changes in energy of various chemical reactions using bond energy, heats of formation, and Hess's Law (324-4)
- ✓ apply one of the methods of predicting heats of reactions to your experimentally determined lab values (214-6)
- ✓ analyse and describe examples where technologies were developed based on understanding thermochemistry (116-4)

SCIENCE DECISIONS INVOLVING THERMOCHEMISTRY

- ✓ describe the importance of peer review in the development of your knowledge about thermochemistry (114-5)
- ✓ use library and electronic research tools to collect information on a given topic (213-6)
- ✓ select and integrate information from various print and electronic sources or from several parts of the same source (213-7)
- ✓ identify multiple perspectives that influence a science-related decision or issue involving your thermochemistry project (215-4)

INTRODUCTION TO THERMOCHEMISTRY LESSON 1

Thermochemistry is the study of the change in *thermal energy* (energy due to the motion of particles). It takes place during physical and chemical changes. *Energy* is defined as the ability to do work. *Work* is defined as a force exerted over a distance. There are two types of energy, kinetic and potential. *Kinetic energy* is the energy of motion. *Potential energy* is stored energy. There are many different forms of energy, such as chemical, thermal, nuclear, electrical, mechanical, sound, solar, magnetic, etc.

Energy can be converted from one form to another. For example, electrical to heat; electrical to light; electrical to sound; chemical to mechanical; chemical to electrical; chemical to thermal.

Chemical reactions require a minimum amount of energy to take place. This energy is needed to break certain chemical bonds of the reactants. Sometimes the energy is thermal energy from the motion of the particles of the reactants, or from a Bunsen burner. Other times, it is supplied as ultraviolet light energy or as electrical energy from passing an electrical current through the chemical system.

Chemical reactions can also be used to produce energy. Some examples would be thermal (heat) energy, light energy, mechanical energy, and electrical energy (dry cell batteries). The energy

involved in chemical reactions is from breaking and making chemical bonds.

Internal energy is the sum of all of the energy contained in a chemical system. The energy of all of the atoms, molecules, or ions within a system. It is made up of both kinetic and potential energies. The kinetic energy is from the translational, rotational, and vibrational motions of the particles. The potential energy is from the positions of the particles (attractive and repulsive forces). An increase in the internal energy of a chemical system (the chemicals we are focused on) has three possible outcomes:

1. The temperature can increase
2. There could be a change of state
3. There could be a chemical reaction

Two terms that students find confusing are heat and temperature. *Heat* is the energy transferred between two objects or systems that are at different temperatures. *Temperature* is a measure of the average kinetic energy of a substance's particles. Temperature is measured with a thermometer, while heat transfer is measured with a calorimeter.

ADDITIONAL READINGS

Types of Changes in Matter and *The Kinetic Molecular Theory* on pages 131-133. *10: Energy Changes* and *10.1: Classifying Energy Changes* on pages 338-339.

LESSON EXERCISES

Exercise 3; page 364

1. Define thermal energy, kinetic energy, and potential energy.
2. What types of energy can be used to start a chemical reaction?
3. What types of energy can be produced as a result of a chemical reaction?

JOURNAL ENTRY**** LEARNING JOURNAL**

You are expected to keep a **Learning Journal** throughout the course. The purpose of this **Learning Journal** is to provide an opportunity for you to reflect upon and expand upon your own learning as you work through the units of the course. For example, you may make notes and observations to ensure you have resolved previous problems. You may have questions to jot down for future reference. You will be reminded to write a journal entry after you finish each lesson when you see the "Journal Entry" heading. Your journal will have a value of 15 per cent of your course mark.

WHAT GOES IN YOUR LEARNING JOURNAL

- ◆ Summarize important ideas or concepts in the lesson

- ◆ Write down connections you see between concepts in the lesson or between concepts in this lesson and concepts from other lessons
- ◆ Write down ideas you find interesting
- ◆ Concepts or calculations that you don't understand and come back to the journal entry later when you do understand, and write down your understanding
- ◆ Write down problem solving techniques that work for you, expressing how you were able to solve a certain problem in your own words
- ◆ Relate the ideas in the lesson to examples in everyday life whenever possible
- ◆ Questions that arose in your mind while going through the lesson
- ◆ Summarize what you need to do in order to finish off the lesson and go on to the next lesson

Please follow these instructions for your **Learning Journal**:

- ◆ Use a three-ring binder.
- ◆ for each entry, note the date, the unit and the lesson to which the entry relates.
- ◆ Give each entry an appropriate title.
- ◆ Double-space all entries.
- ◆ Write on one side of the page only. This will provide space on the back to use at the revision stage for adding further ideas or details.

SUBMIT YOUR JOURNAL ENTRIES AT THE END OF EACH UNIT ALONG WITH YOUR DO AND SEND ASSIGNMENTS.

MEASURING AND CALCULATING THE ENERGY INVOLVED IN A CHEMICAL REACTION LESSON 2

The amount of energy involved in a chemical reaction can be measured using an instrument called a calorimeter. A simple laboratory calorimeter is shown on page 357 of the text. The polystyrene cup acts as an insulator to reduce heat flow to the surroundings. The chemical reaction takes place in the calorimeter and the change of temperature during the reaction is measured. The measurement is the difference between the initial temperature of reactants and the highest temperature reached in the calorimeter. When using a simple calorimeter assumptions are made. These assumptions are:

1. The calorimeter is an isolated system and there is no heat transferred between the calorimeter and its surroundings.
2. The amount of heat absorbed or released by the calorimeter itself is too small to influence calculations of energy involved in the reaction.
3. Any dilute solutions involved in the reaction are treated, for calculation purposes, as if they were pure water with the same density and specific heat capacity as pure water.

The internal energy of a chemical under constant pressure conditions is known as *enthalpy*. The symbol for enthalpy is H . Many chemical reactions

take place in an open beaker under constant pressure of the gases in the room.

The internal energy or enthalpy of a reactant or product can not be measured, but the change in enthalpy can be measured. The symbol for change in enthalpy is ΔH . The relationship is

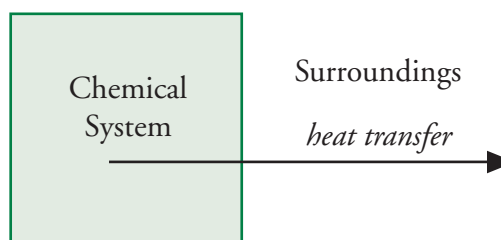
$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

The value of ΔH per mole of the substance involved in the reaction is the molar enthalpy change. When using a calorimeter, ΔH is measured by using the equation

$$\Delta H_{\text{chemical system}} = q_{\text{surroundings}}$$

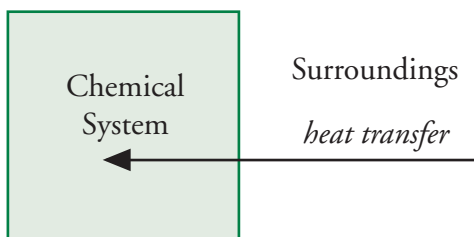
The *chemical system* is the chemicals to be measured in the reaction and the *surroundings* are everything else. The variable q is the heat flow or heat transfer. The change in enthalpy, ΔH equals the heat transferred, q , from the system to the surroundings or from the surroundings to the system. If there is a transfer of heat energy from the chemical system to the surroundings, the reaction is said to be *exothermic*. In an exothermic reaction heat energy is lost by the chemical system to the surroundings.

Exothermic Process

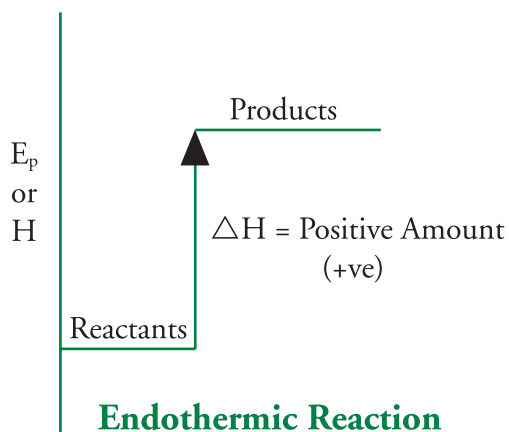
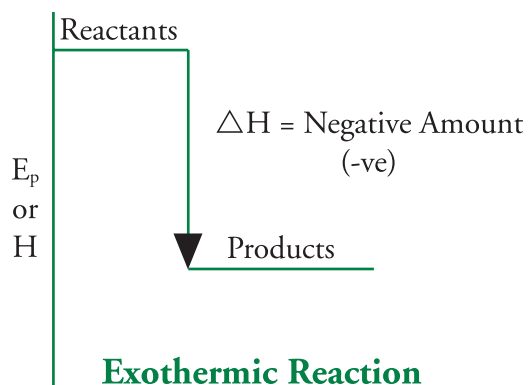


If there is a transfer of heat energy from the surroundings to the chemical system, the reaction is said to be *endothermic*. In an endothermic reaction heat energy is being absorbed by the chemical system from the surroundings.

Endothermic Process



Exothermic and endothermic reactions can be illustrated using potential energy diagrams.



The potential energy diagram for an exothermic reaction demonstrates why exothermic reactions have a negative change in enthalpy (ΔH).

In an exothermic reaction *the change in enthalpy = the enthalpy of the products minus the enthalpy of the reactants* ($\Delta H = H_{\text{products}} - H_{\text{reactants}}$), and *the enthalpy of the reactants is greater than the enthalpy of the products* ($H_{\text{reactants}} > H_{\text{products}}$). Subtracting a larger number from a smaller number will result in a negative answer and a negative change in enthalpy (ΔH).

In an endothermic reaction, bottom left, *the enthalpy of the reactants is less than the enthalpy of the products* ($H_{\text{products}} > H_{\text{reactants}}$). Subtracting a smaller number from a larger number will give a positive answer and a positive change in enthalpy (ΔH).

FORMULA DEFINITIONS

ΔH = heat of reaction in joules

q = variable for heat transfer

m = mass of water or dilute solution of reactants in grams

c_p = specific heat capacity of water or dilute solution of reactant measured in joules per gram multiplied by $^{\circ}\text{C}$

Δt = change in temperature in $^{\circ}\text{C}$

Since $\Delta H_{\text{chemical system}} = q_{\text{surroundings}}$ and since $q = mc_p\Delta t$, the equation that is used to calculate the heat of reaction under constant pressure conditions, the change in enthalpy (ΔH) is $\Delta H = mc_p\Delta t$, where ΔH is the heat of reaction in J; m is the mass of the water or dilute solution of

reactant(s) in g; c_p is the specific heat capacity of the water or dilute solution(s) of reactants in $\text{J/g}\cdot^\circ\text{C}$; and Δt is the change in temperature.

The change in temperature is measured by finding the difference between the initial temperature of the system and the highest temperature reached during the reaction. Remember, any dilute solution that is used is treated like it was pure water and given the same density and specific heat capacity as water.

Specific heat capacity is the amount of energy it takes to raise the temperature of one gram of a given substance by 1.0°C . The specific heat capacity (c_p) of water in its liquid state is $4.18 \text{ J/g}\cdot^\circ\text{C}$. The table on the back cover of the textbook shows the values of c_p for different substances. The *Chemistry Data Booklet* also gives information on specific heat capacity for the elements. If a substance other than water or a dilute solution of a reactant is involved in heat transfer, we must use the value of c_p for that substance.

SAMPLE PROBLEM 1

A chemical reaction was carried out in a calorimeter. The mass of the water in the calorimeter was 650 g. The initial temperature of the water was 20°C . The highest temperature reached during the reaction was 55°C . Calculate how much energy was given off in the reaction in joules.

$$\Delta H = mc_p\Delta t = 650 \text{ g} \times 4.18 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times \overset{(55^\circ - 20^\circ)}{\downarrow} 35^\circ\text{C} = 9.5 \times 10^4 \text{ J} \text{ or } 9.5 \times 10^4 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 95 \text{ kJ}$$

SAMPLE PROBLEM 2

A chemical reaction released 35 kJ of energy. We assume that all of this energy went into heating a sample of water from 21.0°C to 60.0°C . How many g of water was heated?

$$\Delta H = mc_p\Delta t = 35\,000 \text{ J} = m \times 4.18 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times \overset{(60^\circ - 21^\circ)}{\downarrow} 39^\circ\text{C} \quad m = 215 \text{ g}$$

SAMPLE PROBLEM 3

A hot metal having a mass of 75.3 g at a temperature of 97.5°C was placed in a container of 100.3 g of water at an initial temperature of 21.4°C . The final temperature of the system was 25.8°C . Calculate the specific heat capacity, c_p , of the metal.

We will assume that all of the energy lost by the hot metal will be gained by the water.

Heat lost by the metal = Heat gained by the water

$$(q_{\text{metal}} = q_{\text{water}})$$

CHEMISTRY 12

The heat gained by the water is:

$$q_{\text{water}} = mc_p \Delta t = 100.3 \text{ g} \times 4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \times \frac{(25.8^\circ - 21.4^\circ)}{4.4^\circ\text{C}} = 1840 \text{ J}$$

Heat lost by the metal is equal to 1840 J

$$q_{\text{metal}} = mc_p \Delta t; \quad c_p = \frac{q_{\text{metal}}}{m \Delta t} = \frac{1840 \text{ J}}{75.3 \text{ g} \times 71.7^\circ\text{C}} = 0.341 \text{ J/g} \cdot ^\circ\text{C}$$

ADDITIONAL READINGS

Heat and Enthalpy Changes on pages 340-346.

LESSON EXERCISES

Exercises 29, 30, 31, 32; pages 357-365

Exercise 15; page 365

Exercises 6, 7, 10, 11; page 342

JOURNAL ENTRY

MEASURING ΔH USING A BOMB CALORIMETER

LESSON 3

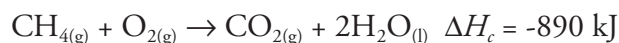
An important class of chemical reactions is combustion reactions. A *bomb calorimeter* is used to measure the energy in combustion reactions. A diagram and explanation of the operation of a bomb calorimeter is shown on page 361 of the textbook. Combustion reactions are all exothermic and transfer heat to the surroundings.

The surroundings in a bomb calorimeter include water, a stirrer, a thermometer, and the calorimeter's bomb itself. The calorimeter is usually calibrated, showing the heat capacity of the entire bomb calorimeter. Knowing the heat capacity of the calorimeter in $J/^\circ C$ and knowing the temperature change, Δt , we can calculate the energy transferred from the combustion reaction to the entire calorimeter.

The bomb calorimeter can be used to measure the molar heat of combustion, ΔH_c , of various fuels. The table below compares the heats of combustion of some chemicals.

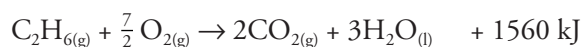
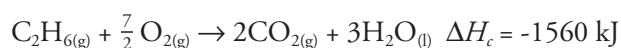
SUBSTANCE	ΔH_c (kJ/mol)
Carbon, C (graphite)	-394
Hydrogen, $H_{2(g)}$	-286
Acetone, $CH_3COCH_{3(l)}$	-1790
Benzene, $C_6H_{6(l)}$	-3268
Ethane, $C_2H_{6(g)}$	-1560
Ethyl alcohol, $C_2H_5OH_{(l)}$	-1367
Methane, $CH_{4(g)}$	-890
Octane, $C_8H_{18(l)}$	-5451
Propane, $C_3H_{8(g)}$	-2220
Sucrose, $C_{12}H_{22}O_{11(s)}$	-5641

Alkanes are hydrogen-carbon compounds that can be used as fuels because they all undergo combustion reactions and release heat. They all react with oxygen, $O_{2(g)}$, to produce $CO_{2(g)}$ and $H_2O_{(l)}$ when undergoing complete combustion. For example,



SAMPLE PROBLEM 1

Write the balanced thermochemical equation (includes either ΔH after the balanced chemical equation or the energy amount as part of the balanced chemical equation) for the molar heat of combustion of ethane, $C_2H_{6(g)}$, using the table of molar heats of combustion below, left. Note the two ways of writing the equations



SAMPLE PROBLEM 2

A bomb calorimeter was used to measure the heat of combustion of naphthalene, $C_{10}H_8$, in oxygen. The heat capacity of the calorimeter is $10.13 \text{ kJ}/^\circ C$. The reaction of 0.640 g of the naphthalene raised the temperature of the calorimeter and its contents by $2.54^\circ C$. Calculate the molar heat of combustion of naphthalene.

$$nH_c = C\Delta t$$

$$q = 10.13 \text{ kJ}/^\circ C \times 2.54^\circ C = 25.7 \text{ kJ}$$

$$q = C\Delta t$$

CHEMISTRY 12

This is the amount of energy released during the combustion of 0.640 g of naphthalene. The molar mass of naphthalene is 128 g. The number of mol of naphthalene used is:

$$\frac{0.640 \text{ g naphthalene} \times 1 \text{ mol naphthalene}}{128 \text{ g naphthalene}} = 0.00500 \text{ mol naphthalene.}$$

The heat released during the combustion of 1.0 mol of naphthalene is:

$$\frac{1.0 \text{ mol naphthalene}}{0.005 \text{ mol naphthalene}} \times 25.7 \text{ kJ} = 5.14 \times 10^1 \text{ kJ}$$

The molar heat of combustion for naphthalene is $\Delta H_c = -5.14 \times 10^1 \text{ kJ}$.

SAMPLE PROBLEM 3

A 1.298 g sample of benzoic acid, $\text{C}_6\text{H}_5\text{COOH}_{(s)}$, was burned in a bomb calorimeter. The temperature change was 4.32°C . The heat of combustion of benzoic acid is $-26\,440 \text{ J/g}$. Calculate the heat capacity of the calorimeter. The energy released by the combustion of 1.298 g of benzoic acid is equal to the energy absorbed by the calorimeter:

$$1.298 \text{ g benzoic acid} \times 26\,440 \text{ J/g} = 3.432 \times 10^4 \text{ J}$$

$$q_{\text{calorimeter}} = C\Delta t$$

where $q_{\text{calorimeter}}$ is the energy absorbed by the calorimeter; C is the heat capacity of the calorimeter; and Δt is the temperature change.

$$3.432 \times 10^4 \text{ J} = C \times 4.32^\circ\text{C}$$

$$C = 7940 \text{ J/}^\circ\text{C}$$

The heat capacity of the calorimeter is $7940 \text{ J/}^\circ\text{C}$

SAMPLE PROBLEM 4

Using the table of molar heats of combustion in this section, calculate which of the following is the best chemical fuel based solely on amount of energy released per g during combustion:

Methane, CH_4	$890 \text{ kJ}/16.0 \text{ g} = 55.6 \text{ kJ/g}$
Ethane, C_2H_6	$1560 \text{ kJ}/30.0 \text{ g} = 52.0 \text{ kJ/g}$
Propane, C_3H_8	$2220 \text{ kJ}/44.0 \text{ g} = 50.4 \text{ kJ/g}$

The best chemical fuel of the three, based solely on energy per gram of fuel, is methane, CH_4 . In determining which one is the best fuel overall, look at other considerations such as availability, cost, effect on the environment, etc.

ADDITIONAL READINGS

Bomb Calorimeters and Heat Capacity on pages 361-363 and the *Summary* on page 364.

LESSON EXERCISES

Exercises 33, 34; page 362

Exercise 26; page 366

JOURNAL ENTRY

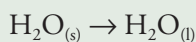
CALCULATING THE ENERGY INVOLVED IN CHANGES OF STATE (PHASE CHANGES)

LESSON 4

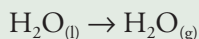
During a chemical reaction, there is a change in composition when reactants change into products. There is also a change in temperature during a chemical reaction. However, during a *change of state* or *phase change*, there is no change in temperature and no substances are produced. Therefore, there is no change in composition.

To calculate the energy involved in a change of state we need to know the mass or number of moles of the substance undergoing the change of state and the molar enthalpy (ΔH) value for the change of state. Changes in enthalpy, ΔH , for changes of state are usually given in kJ per mole of substance changed at constant temperature. Examples and values of changes of state for water are:

Heat of fusion (melting): $\Delta H_{fus} = +6.03$ kJ/mol



Heat of vaporization (boiling): $\Delta H_{vap} = +40.8$ kJ/mol



Note that the sign of ΔH changes for the reverse change of state.

Heat of freezing:



Heat of condensation:



SAMPLE PROBLEM 1

How much energy is needed to melt 25.0 g of ice?

$$\Delta H = n\Delta H_{fus}$$

In this formula, n is the number of moles of ice and ΔH_{fus} is the molar heat of fusion for water.

$$\Delta H = 25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol}}{18.01 \text{ g}} \times 6.03 \text{ kJ/mol} = 8.38 \text{ kJ}$$

SAMPLE PROBLEM 2

If 15.0 kJ of energy is available to boil water, how many g of water can be boiled?

$$\Delta H = n\Delta H_{vap}$$

$$15.0 \text{ kJ} = n \times 40.8 \text{ kJ/mol}$$

$$n = 0.368 \text{ mol} \quad \text{since } n = \frac{\text{mass}}{\text{molar mass}}$$

$$\text{mass} = 0.368 \text{ mol H}_2\text{O} \times \frac{18.01 \text{ g H}_2\text{O}}{1.0 \text{ mol H}_2\text{O}} = 6.63 \text{ g H}_2\text{O}$$

SAMPLE PROBLEM 3

If ΔH_{fus} for lead is +5.10 kJ/mol, how much energy is released when 500 g of lead freezes.

$$\Delta H = n\Delta H$$

$$= 500 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} \times 5.10 \text{ kJ/mol} = 12.3 \text{ kJ}$$

ADDITIONAL READINGS

10.2 *Phase Changes of a System* on pages 346-349.

LESSON EXERCISE

Exercises 14, 15, 16, 19; page 349

Exercises 21, 22 a); pages 351-352

JOURNAL ENTRY

CALCULATING THE TOTAL ENERGY CHANGE FOR A SYSTEM

LESSON 5

When a given amount of energy is available to heat a substance, sometimes there is more than one change of state involved as the substance goes from an initial to a final temperature. Sometimes it is necessary to determine the amount of energy that is required to take a substance at a certain temperature and state to a much higher temperature.

This requires energy calculations for changing the temperature in a certain state and then calculations for a change of state. To calculate the energy involved in a change of temperature of a substance without a change of state or chemical reaction we would use the following formula

$$q = mc_p\Delta t$$

Where q is the flow of heat or energy involved, m is the mass of the substance losing or gaining heat, c_p is the specific heat capacity of the substance, and Δt is the change in temperature of the substance.

To calculate the energy involved in a change of state we would use:

$$\Delta H = n\Delta H_x$$

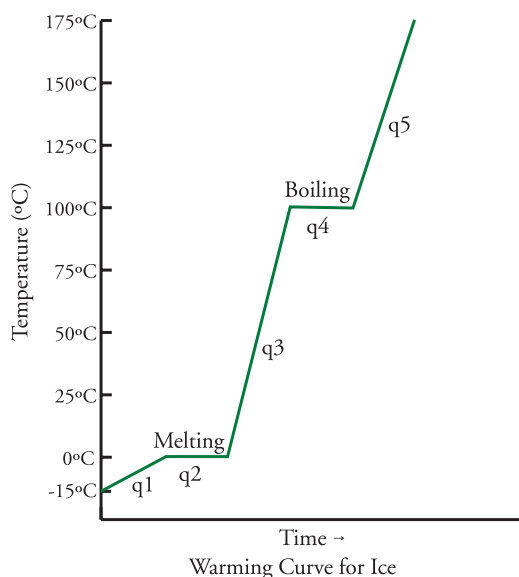
Where ΔH is the energy involved in the change of state, n is the number of moles of the substance, ΔH_x is the molar enthalpy for the change of state, i.e., heat of fusion ΔH_{fus} or heat of vaporization ΔH_{vap} .

SAMPLE PROBLEM 1

How much energy is required to raise the temperature of 50.0 g of water (ice) at -15.0°C to a temperature of 175°C ? In order to calculate the total amount of energy required we need to analyse what energy requirements are involved in taking water (ice) from a temperature of -15.0°C to water (vapour) at a temperature of 175°C . We would look for changes of state as well as instances where we are just warming up the substance without a change of state.

For this sample problem the total energy required is the sum of five separate energy calculations. These include the energy required to change the temperature of the ice from -15.0°C to 0°C , q_1 ; the energy involved in the change of state as the ice melts at 0°C , q_2 ; the energy involved to heat the liquid water from 0°C to 100°C , q_3 ; the energy involved in the change of state as the water boils, q_4 ; and finally the energy required to heat water vapour at 100°C to 175°C , q_5 .

$$q_{total} = q_1 + q_2 + q_3 + q_4 + q_5$$



$$q1 = mc_p\Delta t = 50.0 \text{ g} \times (2.06 \text{ J/g}\cdot^\circ\text{C}) \times 15.0^\circ\text{C} = 1540 \text{ J or } 1.54 \text{ kJ}$$

$$q2 = n\Delta H_{fus} = 50.0 \text{ g} \times \frac{1 \text{ mol}}{18.01 \text{ g}} \times 6.03 \text{ kJ/mol} = 16.7 \text{ kJ}$$

$$q3 = mc_p\Delta t = 50.0 \text{ g} \times (4.18 \text{ J/g}\cdot^\circ\text{C}) \times 100.0^\circ\text{C} = 20\,900 \text{ J or } 20.9 \text{ kJ}$$

$$q4 = n\Delta H_{vap} = 50.0 \text{ g} \times \frac{1 \text{ mol}}{18.01 \text{ g}} \times 40.8 \text{ kJ/mol} = 113 \text{ kJ}$$

$$q5 = mc_p\Delta t = 50.0 \text{ g} \times (2.02 \text{ J/g}\cdot^\circ\text{C}) \times 75.0^\circ\text{C} = 7\,580 \text{ J or } 7.58 \text{ kJ}$$

$$q_{total} = q1 + q2 + q3 + q4 + q5 = 1.54 \text{ kJ} + 16.7 \text{ kJ} + 20.9 \text{ kJ} + 113 \text{ kJ} + 7.58 \text{ kJ} = 160 \text{ kJ}$$

Note, in these calculations a different value of c_p for water was used for each different state of water. For this problem 160 kJ of energy must be supplied. If the problem had been reversed, starting with 50.0 g of water vapour at 175°C and cooling it down to -15.0°C, a total of 160 kJ of energy would have been released.

SAMPLE PROBLEM 2

Lead melts at 328°C and boils at 1740°C. How much energy is required to change the temperature of 75.0 g of lead at 200°C to a temperature of 380°C? The specific heat capacity, c_p , for lead is 0.159 J/g·°C and ΔH_{fus} is 5.10 kJ/mol. In analysing this problem, there are three calculations. The energy required to heat the 75.0 g of lead from 200°C to its melting point at 328°C: $q1$. The energy required to melt the lead: $q2$. Finally the energy required to heat the lead once it has melted from 328°C to 380°C: $q3$.

$$q1 = mc_p\Delta t = 75.0 \text{ g} \times (0.159 \text{ J/g}\cdot^\circ\text{C}) \times 128^\circ\text{C} = 1\,530 \text{ J or } 1.53 \text{ kJ}$$

$$q2 = n\Delta H_{fus} = 75.0 \text{ g} \times \frac{1.0 \text{ mol}}{207.2 \text{ g}} \times 5.10 \text{ kJ/mol} = 1.85 \text{ kJ}$$

$$q3 = mc_p\Delta t = 75.0 \text{ g} \times (0.159 \text{ J/g}\cdot^\circ\text{C}) \times 52.0^\circ\text{C} = 620 \text{ J or } 0.620 \text{ kJ}$$

$$q_{total} = q1 + q2 + q3 = 1.53 \text{ kJ} + 1.85 \text{ kJ} + 0.620 \text{ kJ} = 4.00 \text{ kJ}$$

ADDITIONAL READINGS

Total Energy Changes of a System on pages 350-351

LESSON EXERCISES

Exercises 18, 19, 20, 21,
23, 24; page 365
Exercise 22.b); page 352

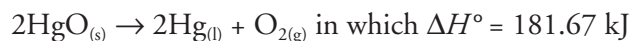
JOURNAL ENTRY

WRITING THERMOCHEMICAL EQUATIONS LESSON 6

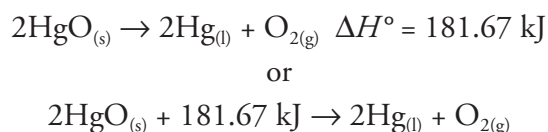
A *thermochemical equation* is a balanced chemical equation that includes ΔH for the reaction and shows the states of all reactants and products. The ΔH can be included as either $\Delta H = \text{some value}$ after the equation or the actual value of ΔH can be included. The placement of ΔH on the reactant or product side depends on whether ΔH is exothermic or endothermic. It is assumed that the value of ΔH is for the molar amounts indicated in the balanced chemical equation. A superscript for ΔH , ΔH° , indicates at standard conditions.

SAMPLE PROBLEM 1

Write the thermochemical equation for the following reaction:



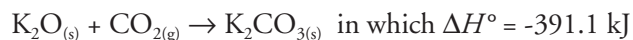
The thermochemical equation can be written as either:



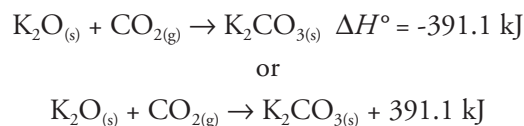
Since ΔH is positive, it is an endothermic reaction and energy is absorbed by the reactants. This is why the value of ΔH is on the chemical reactant side of the chemical equation.

SAMPLE PROBLEM 2

Write the thermochemical equation for the following reaction:



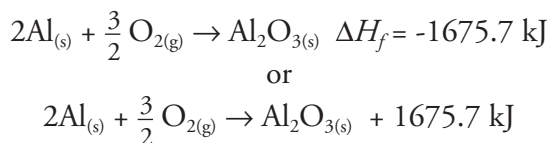
The thermochemical equation can be written as either:



Since ΔH is negative, it is an exothermic reaction and energy is released as the products are formed. This is why the value of ΔH is on the product side of the chemical equation.

SAMPLE PROBLEM 3

Write the thermochemical equation for the heat of formation of $\text{Al}_2\text{O}_{3(s)}$. From the *Chemistry Data Booklet*, ΔH_f° for $\text{Al}_2\text{O}_{3(s)}$ is -1675.7 kJ . The definition for *heat of formation* is the energy absorbed or released when a mole of a compound is formed from its elements in their standard states.



Note, when the energy term is put in as part of the equation on the reactant or product side, it will always be positive. The sign of the energy (- or +) is only written when it is in the form $\Delta H =$.

ADDITIONAL READINGS

11: Reaction Enthalpies and *11.1 Communicating Enthalpy Changes* on pages 369-371.

LESSON EXERCISES

Exercises 1, 2, 4. a) and 5. a); Page 373

JOURNAL ENTRY

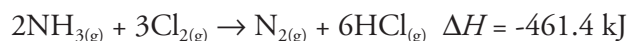
CALCULATING THE ENERGY INVOLVED IN A CHEMICAL REACTION FROM A BALANCED THERMOCHEMICAL EQUATION

LESSON 7

The value of ΔH in thermochemical equations represents the energy involved for the molar amounts indicated in the balanced chemical equation for the chemical reaction. This information can be used to calculate the energy involved in the reaction for any amount of reactants reacted or products produced.

SAMPLE PROBLEM 1

Given the reaction:



How much energy would be released if 35.7g of $\text{NH}_{3(g)}$ were completely reacted? From the thermochemical equation, when 2 mol of NH_3 is completely reacted 461.4 kJ of energy is released. The energy released when 35.7 $\text{NH}_{3(g)}$ is reacted is:

$$35.7 \text{ g NH}_3 \times \frac{1 \text{ mol}}{17.04 \text{ g}} \times \frac{461.4 \text{ kJ}}{2 \text{ mol NH}_3} = 483 \text{ kJ}$$

SAMPLE PROBLEM 2

Given the reaction:



How much energy must be absorbed when 150 g of $\text{I}_{2(s)}$ is completely reacted? From the thermochemical equation, when 1 mol of $\text{I}_{2(s)}$ is

reacted, 53.0 kJ is absorbed. The energy absorbed when 150 g of $\text{I}_{2(s)}$ is completely reacted is:

$$150 \text{ g} \times \frac{1 \text{ mol I}_2}{253.8 \text{ g}} \times \frac{53.0 \text{ kJ}}{1 \text{ mol I}_2} = 31.3 \text{ kJ}$$

LESSON EXERCISES

Questions 14.c), 15.c) page 395

Exercise 19; page 391

JOURNAL ENTRY

DRAWING POTENTIAL ENERGY DIAGRAMS

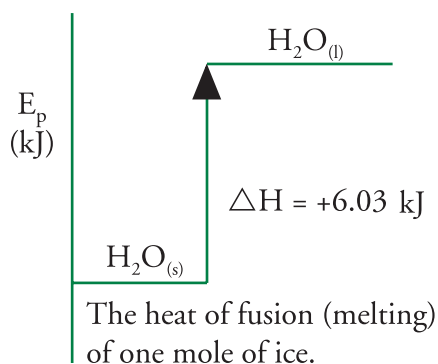
LESSON 8

The *potential energy* of a substance is related to the position of its atoms, molecules, or ions. This affects the magnitude of the forces between these particles. When a change of state takes place, the distance between particles change. When a chemical reaction takes place there is also a change in position and the type of particles as the reactants are changed into products and chemical bonds are broken and formed.

The potential energy of a substance cannot be measured. The change in potential energy of a substance during a change of state or chemical reaction can be measured. The change in potential energy of the particles of a substance is equal to the change in enthalpy (ΔH) under constant pressure conditions. In a *potential energy diagram* the relative potential energy, E_p , of a substance is indicated on the vertical axis and the difference in potential energies is indicated by ΔH .

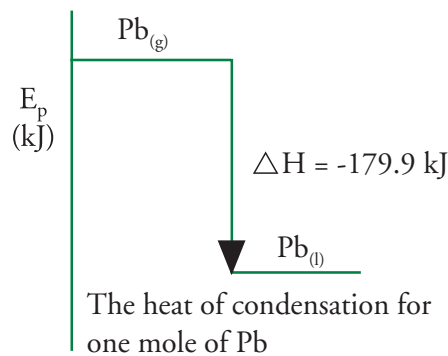
SAMPLE PROBLEM 1

Draw a potential energy diagram to illustrate the melting of one mole of ice. From the *Chemistry Data Booklet*, ΔH_{fus} for water is +6.03 kJ/mol.



SAMPLE PROBLEM 2

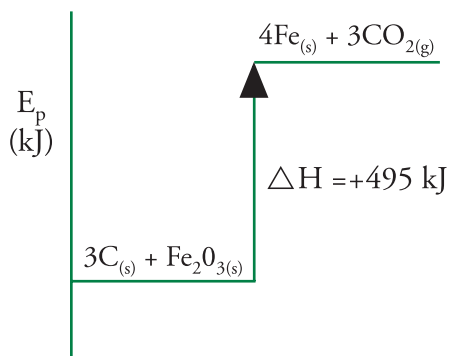
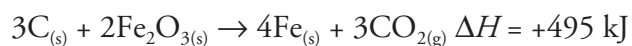
Draw a potential energy diagram to show the condensation of one mole of Pb. The value of ΔH_{vap} for Pb is +179.9 kJ/mol.



Note, since the reverse of vaporization (condensation) is taking place, the sign of ΔH must change.

SAMPLE PROBLEM 3

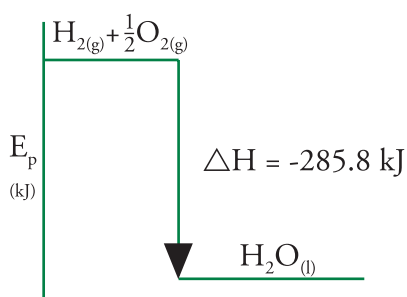
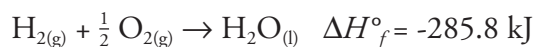
Draw a potential energy diagram for:



SAMPLE PROBLEM 4

Draw a potential energy diagram for the heat of formation, ΔH_f , for $\text{H}_2\text{O}_{(l)}$. From the *Chemistry Data Booklet*, the value of ΔH_f for $\text{H}_2\text{O}_{(l)}$ is -285.8 kJ/mol .

The thermochemical equation for the reaction is

**ADDITIONAL READINGS**

Potential Energy Diagrams on page 354 and *Method 4: Potential Energy Diagrams* and the *Summary* on pages 371-373.

LESSON EXERCISES

Exercises 23, 24, 25, 26; pages 354-355

Exercise 6; page 355

JOURNAL ENTRY

CALCULATING ΔH FOR A CHEMICAL REACTION USING AVERAGE BOND ENERGIES LESSON 9

The ΔH for a chemical reaction is related to the making and breaking of chemical bonds. The action of breaking chemical bonds in reactants requires an input of energy and is endothermic. Whenever chemical bonds are formed in products energy is always released and is an exothermic process. The average bond energies shown below will be used to calculate the ΔH for chemical reactions.

Average Bond Energies (kJ/mol)

	H	F	Cl	Br	I	C	N	O	S	Si
H-	432	565	427	363	295	413	391	467	347	393
F-		154	253	237		485	272	190	327	
Cl-			239	218	208	339	200	203	253	
Br-				193	175	276	243		218	
I-					149	240		234		
C-						347	305	358	259	360
N-							160	201		
O-								146		452
S-									266	
Si-										340
C=						614		799		
N=						615	418	607		
O=								495		
C \equiv						839	891			
N \equiv							941			

$\Delta H = \text{sum of reaction heats from breaking all the bonds in reactants} - (\text{minus}) \text{sum of reaction heats from forming all the bonds in products}$

or

$\Delta H = \text{sum of bond energies of all bonds in reactants} - (\text{minus}) \text{sum of bond energies of all bonds in products}$

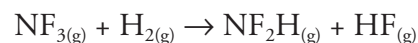
By writing the equation as bond energies of reactants minus bond energies of products, it recognizes that ΔH s for bond energies for reactants are endothermic and positive and ΔH s for bond energies for products are exothermic and negative.

These bond energies are in kJ/mol and must be multiplied by the actual number of moles of bonds indicated in the balanced equation for the reaction.

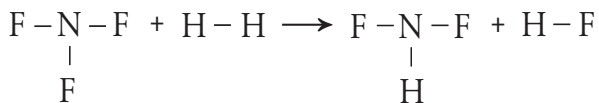
Draw the structural formulas for all of the reactants and products in the balanced chemical equation (if they are not provided) to determine the types of bonds and the number of moles of each type of bond.

SAMPLE PROBLEM 1

Calculate the ΔH for the following chemical reaction using the table of average bond energies.



First, draw structural formulas for all the reactants and products.



For reactants there are:

3 mol of N–F bonds

1 mol of H–H bonds

For products there are:

2 mol of N–F bonds

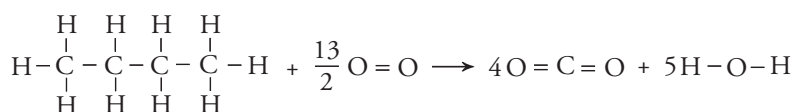
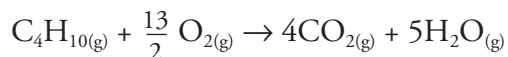
1 mol of N–H bonds

1 mol of H–F bonds

$$\Delta H = (3 \times 272 \text{ kJ} + 1 \times 432 \text{ kJ}) - (2 \times 272 \text{ kJ} + 1 \times 391 \text{ kJ} + 1 \times 565 \text{ kJ}) = -252 \text{ kJ}$$

SAMPLE PROBLEM 2

Calculate ΔH for the following chemical reaction using the table of average bond energies.



For reactants there are:

3 mol of C–C bonds

10 mol of H–C bonds

$\frac{13}{2}$ mol of O=O bonds

For products there are:

8 mol of C=O bonds

10 mol of H–O bonds

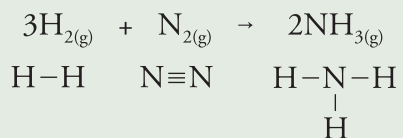
$$\Delta H = (3 \times 347 \text{ kJ} + 10 \times 413 \text{ kJ} + \frac{13}{2} \times 495 \text{ kJ}) - (8 \times 799 \text{ kJ} + 10 \times 467 \text{ kJ}) = -2\,673 \text{ kJ}$$

ADDITIONAL READINGS

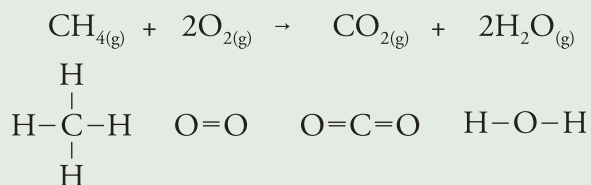
8.4: *Energy Changes* on pages 278-279.

LESSON EXERCISES

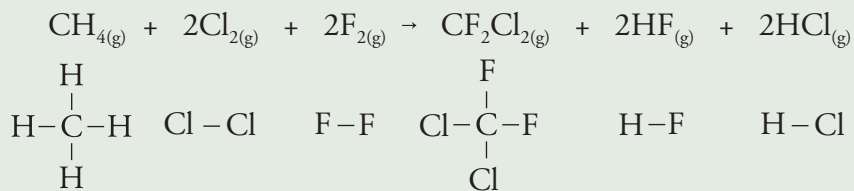
1. Calculate ΔH for the following reaction using average bond energies:



2. Calculate ΔH for the following reaction using average bond energies:



3. Calculate ΔH for the following reaction using average bond energies:

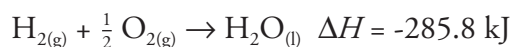


JOURNAL ENTRY

CALCULATING ΔH FOR A CHEMICAL REACTION USING HESS'S LAW LESSON 10

Hess's law states that if a chemical reaction can be written as the sum of two or more simple chemical reactions, then the ΔH for the chemical reaction is the sum of the ΔH s for the two or more simple reactions. For that reason, the ΔH for a reaction can be determined without actually performing the reaction. This can be helpful when the ΔH being determined is a dangerous reaction.

Solving the ΔH for a reaction using Hess's law requires setting up the two or more equations that will combine to give the reaction in question. Sometimes an equation must be reversed. When an equation is reversed the products become reactants and the reactants become products, the sign of ΔH for the reaction is also reversed. For example, to reverse:



gives

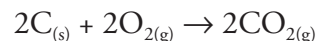


Another way of manipulating simple equations is to multiply an equation by a whole number. ΔH must also be multiplied by this whole number. This is done when the correct number of moles of reactants or products can not be achieved by adding the simple equations together. For example, if we needed 2 moles of $\text{CO}_{2(\text{g})}$ in our

final overall equation and can not get it by adding or subtracting the two or more simple equations, the equation can be multiplied like the one below.



multiplied by 2 to give



$$\Delta H = 2 \times \Delta H = 2(-394 \text{ kJ}) = -788 \text{ kJ}$$

Sometimes nothing needs to be done to the simple equations and the equation is used as is. When adding the equations, treat them as algebraic equations and cancel chemicals if they are the same chemicals, in the same states, in the same amounts, on opposite sides of the arrows of the equations.

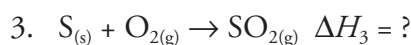
When calculating ΔH using Hess's law, remember to focus on obtaining the correct reactants and products in the final equation and the correct number of moles of reactants and products. This determines whether to reverse an equation or to multiply it by a whole number. Occasionally, the coefficients in a balanced chemical equation must be divided by some whole number.

SAMPLE PROBLEM 1

Given the following equations:

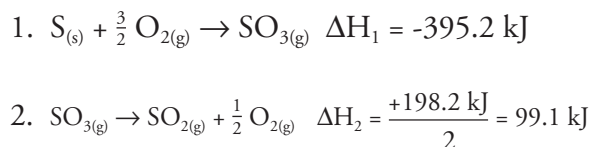


Calculate ΔH for the reaction

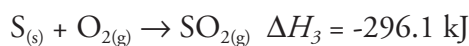


CHEMISTRY 12

In *equation 1* and *equation 3*, $S_{(s)}$ is one of the reactants and the correct number of moles. *Equation 1* is left alone. SO_2 is a reactant in *equation 2*, and a product in *equation 3*. *Equation 2* must be reversed. *Equation 3* has 1 mol of SO_2 and *Equation 2* has 2. *Equation 2* must be divided by 2. This gives:

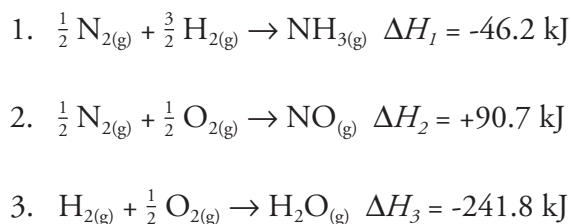


Adding equations 1 and 2 together, including ΔH_1 and ΔH_2 , gives:

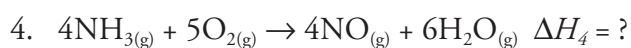


SAMPLE PROBLEM 2

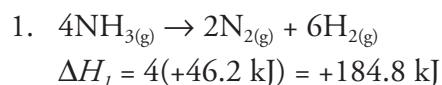
Given the following equations:



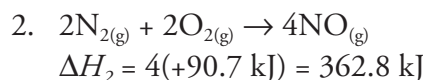
Calculate ΔH for the reaction:



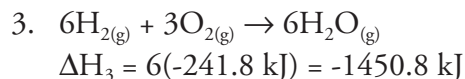
In *equation 1*, $NH_{3(g)}$ is on the wrong (product) side of the equation compared to *Equation 4* and you cannot get 4 moles of NH_3 by adding or subtracting equations. Reverse *equation 1* and multiply it by 4 to give:



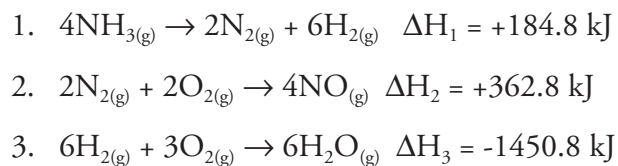
In *equation 2*, $NO_{(g)}$ is on the proper (product) side of the equation compared to *equation 4*. *Equation 2* does not have to be reversed. However, there are 4 moles of NO in *equation 4*. You cannot obtain 4 moles of NO by adding or subtracting equations, so multiply *equation 2* by 4 to give:



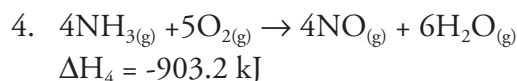
In *equation 3*, $H_2O_{(g)}$ is on the proper (product) side of the equation compared to *equation 4*. *Equation 3* does not have to be reversed. However, there are 6 moles of H_2O in *equation 4*. You cannot get 6 moles of H_2O by adding or subtracting equations. Multiply *equation 3* by 6 to give:



The three manipulated equations now appear as:



Adding the three equations together and cancelling, adding and reducing terms, gives:



ADDITIONAL READINGS

11.2 *Predicting Enthalpy Changes* on pages 374-376.

LESSON EXERCISES

Exercises 7, 8, 9, 10; pages 377-378

JOURNAL ENTRY

CALCULATING ΔH FOR A CHEMICAL REACTION USING STANDARD MOLAR HEATS OF FORMATION LESSON 11

Standard molar heat of formation is the energy absorbed or released when a mole of a compound is formed from its elements in their standard states. These standard states are known as SATP, *standard ambient temperature and pressure*, which is a temperature of 25°C and an air pressure of 100 kPa. A table of standard molar heats (enthalpies) of formation is found on page 609 of the textbook and in the *Chemistry Data Booklet*. Standard molar heats of formation can be used to calculate the ΔH for a chemical reaction. Use the following equation:

$$\Delta H = \sum n\Delta H_{f(\text{product})}^{\circ} - \sum n\Delta H_{f(\text{reactant})}^{\circ}$$

ΔH is the heat of reaction for a given chemical reaction. \sum indicates the sum of. $\Delta H_{f(\text{product})}^{\circ}$ is the standard heat of formation of a product. $\Delta H_{f(\text{reactant})}^{\circ}$ is the standard heat of formation of a reactant, and n is the number of moles of each reactant and product. In calculating the ΔH for a chemical reaction keep in mind that the ΔH_f° for any element is equal to zero. The ΔH_f° value, which is for one mole, must be multiplied by the number of moles of the compound in the balanced equation.

SAMPLE PROBLEM 1

Calculate ΔH for the following reaction using standard molar heats of formation, ΔH_f° s.



From the *standard molar enthalpies of formation table* on page 609 of the textbook, we would find the following values:

$$\Delta H_f^{\circ} \text{ for } \text{NH}_3(\text{g}) = -45.9 \text{ kJ/mol}$$

$$\Delta H_f^{\circ} \text{ for } \text{HCl}(\text{g}) = -92.3 \text{ kJ/mol}$$

$$\Delta H_f^{\circ} \text{ for } \text{Cl}_2(\text{g}) \text{ and } \text{N}_2(\text{g}) \text{ is } 0$$

$$\Delta H = \sum n\Delta H_{f(\text{product})}^{\circ} - \sum n\Delta H_{f(\text{reactant})}^{\circ}$$

$$\Delta H = (0 + 6(-92.3 \text{ kJ})) - (2(-45.9 \text{ kJ}) + 0)$$

$$= (-553 \text{ kJ}) - (-91.8 \text{ kJ}) = -461.2 \text{ kJ}$$

SAMPLE PROBLEM 2

Calculate the ΔH for the following reaction using standard molar heats of formation, ΔH_f° s:



From the *standard molar enthalpies of formation table* on page 609 of the textbook:

$$\Delta H_f^{\circ} \text{ for } \text{CH}_3\text{OH}(\text{l}) = -239.1 \text{ kJ/mol}$$

$$\Delta H_f^{\circ} \text{ for } \text{CO}_2(\text{g}) = -393.5 \text{ kJ/mol}$$

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

CHEMISTRY 12

ΔH_f° for $\text{O}_{2(g)}$ is 0.

$$\begin{aligned}\Delta H &= \sum n\Delta H_f^\circ(\text{product}) - \sum n\Delta H_f^\circ(\text{reactant}) \\ &= (2(-393.5 \text{ kJ}) + 4(-285.8 \text{ kJ})) - (2(-239.1 \text{ kJ}) + 0) \\ &= -1452.0 \text{ kJ}\end{aligned}$$

ADDITIONAL READINGS

Predicting ΔH_r Using Formation Reactions on pages 380-383.

LESSON EXERCISES

Exercises 11, 12, 13; pages 383-384

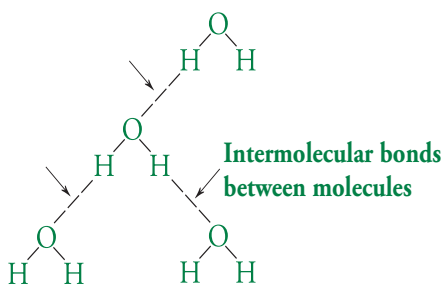
Exercises 13, 14.a), b), 15; page 395

JOURNAL ENTRY

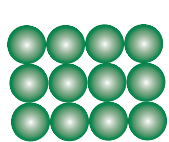
A COMPARISON OF THE ENERGIES INVOLVED IN PHASE CHANGES, CHEMICAL CHANGES, AND NUCLEAR REACTIONS

LESSON 12

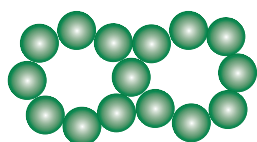
The energy involved in chemical reactions is approximately 10–100 times greater than the energy involved in phase changes. Comparing similar masses reacting, the energy from nuclear reactions are about 1 million times greater than the energy involved in chemical reactions. During a phase change, the *intermolecular bonds* between molecules must be overcome to separate the molecules when changing from a solid to a liquid or a liquid to a gas.



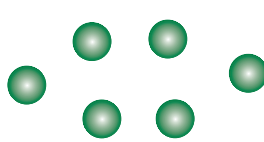
The particles (molecules) in the gas state are completely separated from each other but not all molecules are separated from each other in the liquid state.



Solid

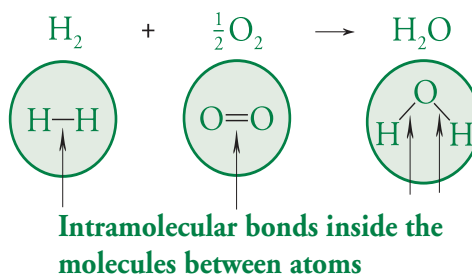


Liquid



Gas

In chemical reactions, the *intramolecular bonds* between atoms within molecules must be broken during the reaction. These intramolecular bonds are much stronger than the intermolecular bonds between molecules.



Nuclear reactions involve separating the particles inside the nucleus of an atom to create a different element(s). These forces of attraction are much stronger than chemical bonds. In all three types of change, energies must be absorbed (endothermic) to separate particles at the beginning of the change. The same energies are released (exothermic) in the reverse process.

ADDITIONAL READINGS

11.3: *Nuclear Reactions* on pages 385–386.

LESSON EXERCISES

Exercises 7, 8, 9; pages 364–365

JOURNAL ENTRY

DO AND SEND

THERMOCHEMISTRY ASSIGNMENT QUESTIONS 70 POINTS

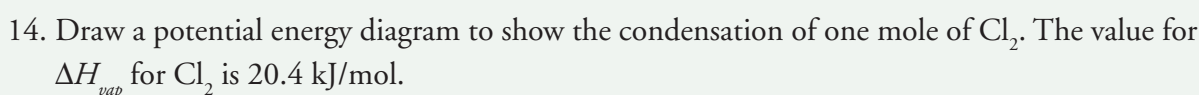
1. Explain the differences between the terms “heat” and “temperature”.
2. How is chemical energy released?
3. A chemical reaction was carried out in a calorimeter. The mass of the water in the calorimeter was 500 g. The initial temperature of the calorimeter was 18.0°. The highest temperature reached during the reaction was 60.0°C. Calculate how much energy was given off in the reaction in joules.
4. A chemical reaction released 50.0 kJ of energy as it heated a sample of water from 17.0°C to 75.0°C. What mass of water was heated?
5. Write the balanced thermochemical equation for the molar heat of combustion for propane, $C_3H_{8(g)}$ using the table of molar heats of combustion.
6. A 0.5269 g sample of octane, $C_8H_{18(g)}$ was burned in a bomb calorimeter. The bomb calorimeter had a heat capacity of 11.3 kJ/°C. The temperature increased by 2.25°C. Calculate the molar heat of combustion of octane.
7. If ΔH_{fus} for NaCl is 28.0 kJ/mol, how much energy is required to melt 350 g of NaCl at its melting point?
8. Calculate the total energy required to heat 65.0 g of sodium, Na, from 20.0°C to 1000°C. The melting point of Na is 98.0°C and its boiling point is 883°C. The molar heat of fusion, ΔH_{fus} , for sodium is 2.60 kJ/mol, and molar heat of vaporization, ΔH_{vap} of sodium is 101.0 kJ/mol. The specific heat capacity, c_p for Na is 1.226 J/g•°C.
9. Write the thermochemical equation for the following reaction: $Xe_{(g)} + 2F_{2(g)} \rightarrow XeF_{4(s)}$ in which $\Delta H = -251$ kJ.
10. Write the thermochemical equation for the standard molar heat of formation for $BaSO_{4(s)}$.



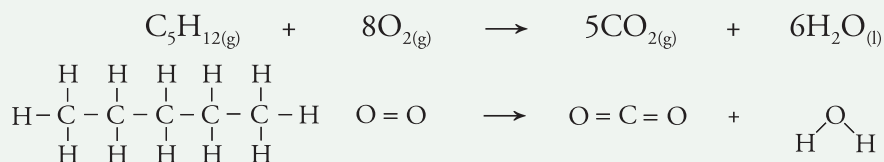
How much energy would be released if 20.0 g of $O_{2(g)}$ was completely reacted?



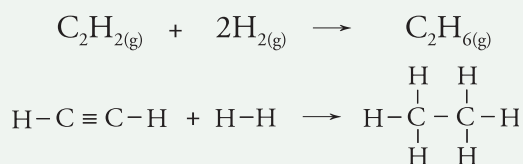
How much energy would be released if 35.0 g of $NH_{3(g)}$ was completely reacted?



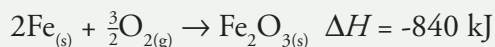
15. Calculate the ΔH for the following reaction using average bond energies:



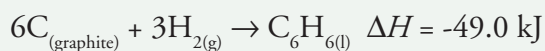
16. Calculate the ΔH for the following reaction using average bond energies:



17. Use Hess's law to calculate ΔH for: $3C_{(s)} + 2Fe_2O_{3(s)} \rightarrow 4Fe_{(s)} + 3CO_{2(g)}$ given:



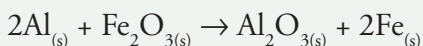
18. Use Hess's law to calculate ΔH for: $C_6H_6(l) + 15/2O_{2(g)} \rightarrow 6CO_{2(g)} + 3H_2O_{(l)}$ given:



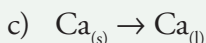
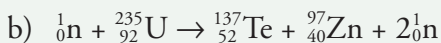
19. Calculate ΔH for the following reaction using standard molar heats of formation, ΔH_f° s.



20. Calculate ΔH for the following reaction, using standard molar heats of formation, ΔH_f° s.



21. Rank the energy involved for the following three changes, from the least energetic to the most energetic and explain.



DO AND SEND

THERMOCHEMISTRY RESEARCH ASSIGNMENT 30 POINTS

a) Gasoline is burned in millions of vehicles every day. Consider the effects gasoline consumption has from different points of view, like cost, convenience, exhausting resources etc. What are six problems, specific to Nova Scotia, with the gasoline we use? *200 words*

b) From the following six suggestions for alternatives to gasoline

- | | |
|---------------------------|------------------------|
| 1. "clean gasoline" | 4. methanol |
| 2. ethanol | 5. electricity |
| 3. compressed natural gas | 6. solar-hydrogen cell |

decide which alternative we should use now, in Nova Scotia, to replace gasoline fuel for automobiles. Present arguments explaining why this alternative fuel should be used. Include all **references**. There is a great deal of information about these six alternative fuels on the Internet and in science resource books. *150 words*

c) Suggest another fuel or combination of fuels for automobiles in Nova Scotia besides the six mentioned. Describe that fuel or combination of fuels and why we should use it (them). Use the Internet and/or other resources, and **give all of your references**. *150 words*

EXPERIMENT 1

DO AND SEND

TITLE

Measuring The Energy Given Off By A Burning Candle

PURPOSE

To measure the energy released by a candle as it burns

CHEMICALS

water, candle wax

EQUIPMENT

thermometer, small empty can, measuring cup, candle and holder, spoon, oven mitts or tongs

DATA TABLE

Volume of water added to can (mL)	mL
Initial temperature of the water	°C
Highest temperature reached by the water	°C

SAFETY

Use an oven mitt or tongs to hold the can with water being heated if you feel the can is getting too hot to hold.

PROCEDURE

Part 1: Measure out 100 mL of tap water into a small can.

Part 2: Set up a candle in a candle holder in a sink and light the candle.

Part 3: Place a thermometer in the water in the can. The bulb of the thermometer should be completely below the surface of the water. If more water is needed to completely cover the thermometer bulb, add an additional, measured amount of water so that you know the total volume of water in the can.

Part 4: After the thermometer has been in the water for 30 seconds, observe and record the temperature of the water.

Part 5: Holding the can at the top, place the bottom of the can just above the candle flame. Stir the water with the thermometer, but do not rest it on the bottom of the can. Heat the water in the can until the temperature rises by about 15°C.

Part 6: Blow out the candle. Stir the water in the can and observe and record the highest temperature reached by the water before the temperature of the water starts to fall.

Part 7: Observe the bottom of the can.

CLEAN UP

Empty the water from the can into the sink.

QUESTIONS

We are going to use the equation $q = mc_p\Delta t$ to calculate the energy given off by the burning candle and absorbed by the water.

CHEMISTRY 12

1. What assumptions are made in doing this experiment?
2. Calculate the energy q , transferred to the water by the burning candle.
3. Explain what you observed on the bottom of the can.
4. Will soot on the bottom of the can cause a smaller or larger value of calculated energy released by the burning candle? Explain.
5. Comment on the sources of error in this experiment, since we did not use a sophisticated calorimeter to measure the energy given off.
6. What three questions do you have as a result of doing this experiment?

CONCLUSION

What conclusions did you make as a result of doing this experiment?

EXPERIMENT 2

DO AND SEND

TITLE

Measuring The Energy Transferred By Copper Pennies

PURPOSE

To measure the energy transferred by hot copper pennies to a known amount of water.

CHEMICALS

water, 10 copper pennies

EQUIPMENT

thermometer, small heat-safe glass container, measuring cup, source of energy to boil water, small pot

DATA TABLE

Volume of water added to container (mL)	mL
Initial temperature of the water	°C
Highest temperature reached by the water	°C

SAFETY

Use caution when working with the boiling water. Do not pick up hot pennies with your bare hands.

PROCEDURE

Part 1: Measure out 100 mL of tap water into a small glass container.

Part 2: Half fill a pot with water and add 10 pennies. Heat the water to boiling on the stove. Once the water is boiling vigorously, turn off the stove.

Part 3: Measure and record the temperature of the 100 mL of tap water that you added to the small glass container.

Part 4: Carefully pour the pot of water and pennies into a strainer in the sink. Quickly dump the pennies into the 100 mL of water in the small container. Stir with a thermometer. Measure and record the highest temperature reached by the 100 mL of water before the temperature of the water starts to fall.

CLEAN UP

Empty the water from the small container into the sink. Wash the pot and the container.

QUESTIONS

We are going to use the equation $q = mc_p\Delta t$ to calculate the energy given off by the copper pennies and absorbed by the water.

1. What assumptions are made in doing this experiment?
2. Calculate the energy q , transferred to the water by the copper pennies.
3. What was the temperature of the copper pennies when they were added to the 100 mL of water?

CHEMISTRY 12

4. What information is required in order to calculate the specific heat capacity of the copper penny?
5. Comment on the sources of error in this experiment, since we did not use a sophisticated calorimeter to measure the energy given off.
6. What three questions do you have as a result of doing this experiment?

CONCLUSION

What conclusions did you make as a result of doing this experiment?