Ch. 8. Thermochemistry

E NERGY involves many aspects of our everyday life. Chemical reactions happening in our bodies consume or release energy as we walk, study, and even breathe. We also use energy at home to warm food or turn on the lights, drive our cars, and go to work. The energy needed for our body to function comes from food. If we do not eat for a while, we run low on energy. Similarly, the burning of fossil fuels such as oil, propane, or gasoline provides enough energy to maintain our homes. On the other hand, how do we measure the energy released or consumed in a chemical process? This chapter will answer this and other questions as it covers different aspects of the interplay between chemistry and energy. You will learn about temperature, heat, and how an energy exchange implies changes in temperature.

8.1 Energy

When you are running, walking, dancing, or thinking, you are using energy to do work. Energy is defined as the ability to do work. Suppose you are climbing a steep hill. Perhaps you become too tired to go on; you do not have sufficient energy to do any more work. Now suppose you sit down and have lunch. In a while, you will have obtained energy from the food, and you will be able to do more work and complete the climb. Similarly, chemical energy is the energy stored in the structure of chemicals and it depends on the arrangement of molecules and the nature of these molecules.

Potential & Kinetic Energy: heat Energy can be classified as potential energy or kinetic energy. Kinetic energy is the energy of motion and any fast-moving object has kinetic energy. Think about a bullet coming out of a gun; as the bullet moves very fast it contains kinetic energy that can be released when it collides with a target. Potential energy is energy stored in objects located at a certain height. A boulder resting on top of a mountain has potential energy because of its location. If the boulder rolls down the mountain, the potential energy becomes kinetic energy. Water stored in a reservoir has potential energy. When the water goes over the dam, the potential energy is converted to kinetic energy. The potential energy resulting from the interaction of charged particles is called electrostatic energy. Heat refers to thermal energy, which is associated with the random motion of particles in a substance and therefore is related to kinetic energy. A frozen pizza feels cold because the particles in the pizza are moving slowly. As the pizza receives heat, the motion of the particles increase. Eventually, the particles have enough energy to make the pizza hot and ready to be eaten. When a substance receives heat it gets warmer and it raises its temperature, whereas if it loses heat it gets cooler and its temperature decreases.

Mechanical energy: work The sum of potential and kinetic energy is called mechanical energy. Mechanical energy refers to the ability to do work. Examples of work are a car engine moving or a balloon expanding its volume.

The law of conservation energy In this chapter, we will analyze energy changes associated with chemical reactions. To do this, we need to define the system and its surroundings. The system will be the chemical reaction often happening in a beaker, whereas the surroundings would be the area surrounding the beaker. The system plus its surroundings are called the universe. The beaker may lose energy, and in that case, energy will flow from the system to the surroundings. Similarly, the system may gain energy, flowing from the surroundings to the beaker. In a closed system, the energy is conserved and when one type of energy disappears, a different type of energy will appear. As an example, if you drop an object from the top of a building, originally the object had potential energy that converts into kinetic energy as the object gains speed. This is called the law of conservation of energy.

Energy units Two different units of energy are often employed: calories (cal) and joules (J). Joule is the SI unit of energy equal to $kg \cdot m^2/s^2$. One can transform calories into joules and joules into calories using the following conversion factor:

$$\boxed{1cal = 4.184J} \qquad \text{or} \qquad \boxed{\frac{1cal}{4.184J}} \tag{8.1}$$

As a note, often you will read on food labels the caloric content of certain foods. In these labels, they use the unit Calorie, with capitalized C, which is not the same as a calorie. One Calorie represents a kilocalorie and contains 1000 calories.

Sample Problem 1

Convert the following energy values:

(a) 50000 cal to Kcal   (b) 48001 J to cal

SOLUTION

We will use the conversion factor for kilo and the relationship between calorie and joule:

(a) $50000cal \times \frac{1kcal}{1000cal} = 50kcal$; (b) $48000J \times \frac{1cal}{4.184J} = 11472cal$.

◈ **STUDY CHECK**

Convert the following energy units: (a) 200 cal to Kcal (b) 7000 J to cal

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Thermodynamics Thermochemistry is a subject of a broader field called thermodynamics, which studies the interconversion of energy (heat and other types) and mass. Thermodynamics study systems, like chemical reactions. The term system, refers to the part of the universe being study. Systems can be classified as: open, closed and isolated systems. An open system can exchange mass and energy with its surroundings, whereas a closed system can only exchange mass and not energy. Isolated systems cannot exchange neither mass nor energy with its surroundings. The state of a system is characterized by the values of its volume, pressure, temperature, energy and composition, so that if a system receives heat it will change its state. Energy, volume, pressure and temperature are called state functions–or state properties–, as these properties are determined by the state of the system, independently of the path used to reach that state. In another words, these properties are path-independent. For example, in a building, the floor location of a person would be a state function, as it would not matter the path the

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▼a bullet has kinetic energy
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▼water on a dam has potential energy
© wikipedia

▼thermal energy refers to heat
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person took to reach that state. In contrast the amount of effort to make it to a specific floor will not be a state function, as it changes depending on the path used.

8.2 Temperature

Temperature indicates how hot or cold a substance is compared to another substance. Heat always flows from a substance with a higher temperature to a substance with a lower temperature until the temperatures of both are the same. When you drink hot coffee or touch a hot pan, heat flows to your mouth or hand, which is at a lower temperature. When you touch an ice cube, it feels cold because heat flows from your hand to the colder ice cube. Three units of temperature often employed are celsius (°C, T_C), Fahrenheit (°F, T_F) or Kelvin (K, T_K). If you need to convert temperature units from Fahrenheit to celsius or from celsius to Fahrenheit you need to use the formulas below:

$$\boxed{T_F = 1.8T_C + 32} \qquad \boxed{T_F = 1.8T_K - 459.4} \qquad \boxed{T_K = T_C + 273} \qquad (8.2)$$

Sample Problem 2

Convert 25 °C to °F.
SOLUTION

1 **Step one:** list of the given variables.

Analyze the Problem	Given	Asking
	$T_c = 25°C$	T_F

2 **Step two:** use the formula $T_F = 1.8T_C + 32$ to convert from °C to °F.

3 **Step three:** solve for $T_F = 1.8 \times 25 + 32 = 77°F$.

◆ **STUDY CHECK**
Convert 200°C to K.

8.3 From energy to temperature

Heat transforms in a temperature change. Some substances like metals can increase their temperature very quickly with a small amount of heat received, whereas others need a larger amount of heat to rise their temperature. Think about why you use oil to deep fry food. Why not use water? First of all, oil can raise its temperature very quickly and on top of that it does not boil easily.

Heat capacity The heat capacity c of a material is defined as:

$$c = \frac{\text{heat adsorbed}}{\text{temperature increase}} \qquad (8.3)$$

This is a characteristic property of each material that indicates the energy required to rise its temperature and can be expressed in cal/°C or J/°C units. As this property depends on the amount of matter, oftentimes the heat capacity is expressed per mass as the specific heat capacity also known as *specific heat* (c_e) or mole unit as the *molar heat capacity* c_m. For example, the specific heat of water is 1cal/g°C that is the same as 4.184J/g°C. That means that we need to give 1 calorie to warm up one gram of water $1°C$. Similarly, the specific heat of aluminum, a metal, is 0.2cal/g°C or 0.89J/g°C; that means the energy needed to raise the temperature of an aluminum gram is 0.2 calories of 0.89 J. Mind the difference between these two values: we need to give 1 cal to increase the temperature of a gram of water in $1°C$, whereas we need to give 0.2 cal to increase the temperature of a gram of aluminum in $1°C$. Why are these two numbers so different? The answer is that water and aluminum are different materials. Normally metals warp up very easily, that is, they need less heat to increase their temperature, whereas liquids need more heat to increase their temperature. That is why pans and cooking pots tend to be metallic. Mind the specific heat if water is a well know value that you need to be familiar with:

$$c_e^{H_2O} = 4.184J/g°C \quad or \quad c_e^{H_2O} = 1cal/g°C \tag{8.4}$$

Table 8.1 Values of specific heat for different materials

Material	Specific heat (J/g°C)	Material	Specific heat (J/g°C)
$H_2O_{(l)}$	4.184	$Fe_{(s)}$	0.444
ethyl alcohol$_{(l)}$	2.460	$Au_{(s)}$	0.129
vegetable oil$_{(l)}$	1.790	$Cu_{(s)}$	0.385
$NH_{3(l)}$	4.700	$H_2O_{(s)}$	2.010
Dry Air$_{(g)}$	1.0035	$CO_{2(g)}$	0.839

Heat When a material receives heat, that heat normally becomes temperature as the temperature of the material increases. For example, if you warm milk in a microwave, the milk's temperature increases from room temperature (25°C) to a higher temperature. How to estimate the temperature increase given the heat received? Or how to estimate the heat needed to increase the temperature of an object? We can use the following formula:

$$Q = m \cdot c_e \cdot (T_f - T_i) \tag{8.5}$$

where:

Q is the amount of heat received, either in cal or J.

m is the mass of material in grams

c_e is the specific heat of the material (in cal/g°C or J/g°C)

$T_f - T_i = \Delta T$, is the temperature change from the initial to the final temperature

A system can receive or give away heat and this is indicated by the sign of Q. The sign convention for heat is:

$Q > 0$ the system receives heat $\quad and \quad Q < 0$ the system gives away heat

Sample Problem 3

How many calories are absorbed by a 45.2g piece of aluminum ($c_e = 0.214 \frac{cal}{g°C}$) if its temperature rises from 25°C to 50°C.

SOLUTION

① **Step one:** list of the given variables.

Given	Asking
$c_e = 0.214 \frac{cal}{g°C}$ $m = 45.2g$ $T_{initial} = 25°C$ $T_{final} = 50°C$	Q

② **Step two:** use the formula $Q = m \cdot c_e \cdot (T_{final} - T_{initial})$ to transform the temperature increase into heat absorbed. Mind this formula depends on the mass involved and the specific heat of the material, in this case, aluminum.

③ **Step three:** solve $Q = 45.2 \cdot 0.214 \cdot (50 - 25) = 241.82 cal$.

◈ **STUDY CHECK**

How many calories are absorbed by 100g of Gold ($c_e = 0.0308 \frac{cal}{g°C}$) if its temperature rises from 25°C to 100°C.

In the previous example you needed to convert temperature into heat. In the next example, the heat is given and you need to calculate the final temperature of an object after it receives a certain amount of heat.

Sample Problem 4

A 50g piece of aluminum ($c_e = 0.214 \frac{cal}{g·°C}$) initially at 25°C absorbs 100cal. Calculate the final temperature of the aluminum piece.

SOLUTION

① **Step one:** list of the given variables.

Given	Asking
$c_e = 0.214 \frac{cal}{g·°C}$ $m = 50g$ $T_{initial} = 25°C$ $Q = 100cal$	T_{final}

② **Step two:** use the formula $Q = m \cdot c_e \cdot (T_{final} - T_{initial})$ that converts temperature increase to heat absorbed.

③ **Step three:** solve $100 = 50 \cdot 0.214 \cdot (T_{final} - 25)$ for T_{final}:

$$100 = 50 \cdot 0.214 \cdot (T_{final} - 25) \qquad \text{divide by 50 in both sides}$$

$$\frac{100}{50} = 0.214 \cdot (T_{final} - 25) \qquad \text{divide by 0.214 in both sides}$$

$$\frac{100}{50 \cdot 0.214} = (T_{final} - 25)$$

$$9.34 = (T_{final} - 25) \qquad \text{add 25 in both sides}$$

$$9.34 + 25 = T_{final}$$

$$34.34 = T_{final}$$

The final temperature of the aluminum piece is $34.34°C$.

◈ **STUDY CHECK**

A 200g piece of iron ($c_e = 0.1 \frac{cal}{g \cdot °C}$) initially at 15°C absorbs 1000cal. Calculate the final temperature of the metal piece in °C.

8.4 The first law of thermodynamics

Materials can absorb heat and receive work. On one hand, think about a pizza in your oven, or a cup of milk in the microwave. These substances receive heat from the oven or in form of microwaves and they become hot. Heat transforms in an increase in temperature. On the other hand, if you hammer a wall, the wall receives work from you, and this work is translated to energy as the wall may break. Work and heat are both combined in a property called internal energy, E.

Work Think about what happens to a car's engine. In an engine, chemical energy is converted into movement and with this movement, a car can carry work. Work (W) is force (F) applied over a distance (Δh):

$$W = F \cdot \Delta h = (P \cdot A) \cdot \Delta h$$

For the case of a gas confined in a cylindrical container, force is related to pressure times area (A). Therefore, if the pressure is constant as $A \cdot \Delta h$ equals to volume (V), we have that

$$\boxed{W = -P \cdot \Delta V} \qquad (8.6)$$

where:

W represents work

P is the pressure

ΔV is the change of volume calculated as $V_{final} - V_{initial}$

The minus sign is just a convention, as in chemistry the work done by the surroundings to a system is considered positive, as the system gains work (energy). Therefore, when ΔV is negative because the system receives work, the value of W has to be positive. In another word, the sign convention for work is:

$W > 0$ the system receives work   *and* $W < 0$ the system gives away work

As a final note, the type of work that involves a change in volume at constant pressure is normally called PV work.

Sample Problem 5

Calculate the work (in L · atm) involved in the expansion of a gas from 10L to 20L at constant external pressure of 2atm. Convert the value in joules using 101.3J=1L · atm.

SOLUTION

We will use Equation 8.6 that related work with pressure and volume change:

$$W = -P \cdot \Delta V = -2\text{atm} \cdot (20 - 10)\text{L} = -20\text{L} \cdot \text{atm}$$

If we convert this value into J:

$$-20\text{L} \cdot \text{atm} \times \frac{101.3 J}{1\text{L} \cdot \text{atm}} = -2026 J$$

As the value is negative, it means that the system gives away work.

◈ **STUDY CHECK**

Calculate the work involved in the compresion of a gas from 10L to 5L at constant external pressure of 5atm (a) in L · atm units, and (b) in joules, using 101.3J=1L · atm.

First law of thermodynamics: the internal energy The combination of work (W) and heat (Q) is called internal energy (ΔE):

$$\boxed{\Delta E = Q + W} \tag{8.7}$$

The first law of thermodynamics–the law of conservation of energy–states that the energy of the universe is constant.

Sample Problem 6

When a hot balloon inflates and deflates in order to change its height. It receives 10^3J of heat and its volume increases from 3.0×10^5L to 3.5×10^5L at fixed external pressure of 1atm. Calculate the internal energy of the hot ballon, using 1J=101.3L · atm.

SOLUTION

We will use Equation 8.6 that related work with pressure and volume change. As the resulting unit of work will be L · atm, we will directly convert the value into J:

$$W = -P \cdot \Delta V = -1\text{atm} \cdot (3.5 \times 10^5 - 3.0 \times 10^5)\text{L} \times \frac{1\text{J}}{101.3\text{L} \cdot \text{atm}} = -493.6\text{J}$$

Now we will add the value of heat (10^3J) to the value of work to calculate the internal energy:

$$\Delta E = Q + W = 10^3 + (-493.6) = 1493.6\text{J}$$

Overall, the hot balloon receives more heat that the work it gives away and hence the resulting internal energy is positive–the system gains energy.

◈ **STUDY CHECK**

When a hot balloon deflates, it receives 10^7J of work from the external atmo-

sphere and its temperature change from 90°C to 25°C. Given that the air initially contained in the balloon has a mass of 3×10^5g and a specific heat of $1\text{J}/\text{g}°\text{C}$. Calculate the internal energy of the hot ballon.

8.5 Calorimetry

A calorimeter is a tool used to measure the exchange of heat happening in chemical reactions and calorimetry is the science that measures heat exchange by using calorimeters. There are two types of calorimeters, very fancy and expensive ones called constant-volume calorimeters, and more affordable ones, in the form of a coffee cup, called constant-pressure calorimeters.

The calorimeter A calorimeter is a device used to measure the energy exchanged in a chemical reaction–we call this molar heat of reaction, ΔQ_r. In essence, a calorimeter is a closed system that does not let the heat come through its walls. By measuring the temperature change inside the calorimeter we can compute the energy exchange in a chemical reaction happening inside the calorimeter. If the temperature inside a calorimeter increases, this means that the reaction releases energy. On the contrary, if the temperature inside a calorimeter decreases, this means the reaction consumes energy. There are two different types of calorimeters: constant pressure and constant-volume calorimeters. As a note, ΔQ_r is called the molar heat of the reaction as it represents energy per mole, with units of kJ/mol.

Constant-pressure calorimeter A constant-pressure calorimeter is the simplest of all calorimeters and is called constant pressure as the pressure inside the calorimeter is constant and equal to the atmospheric pressure. A constant-pressure calorimeter is just a double coffee cup covered with a lit. Inside this cup, a chemical reaction occurs in a liquid phase. If the reaction produces any gases as the cup is just covered with a lit, the pressure will always be equal to the atmospheric pressure as the gas can escape through the lit. The formula used in calorimetry with a constant-pressure calorimeter has only two terms. Let us use a reaction that produced heat as an example. Inside a constant-pressure calorimeter, you introduced two reagents, and a reaction happens, producing heat. The heat exchanged from the reaction (first term) changes the temperature of the liquid inside the calorimeter (second term). At the same time, we assume that the walls of the calorimeter do not absorb heat. The formula used in calorimetry with a constant-pressure calorimeter is the following:

$$0 = n \cdot \Delta Q_r + \Delta Q_{water} \quad \text{constant pressure}$$

where:

$$n \cdot \Delta Q_r \text{ is the heat exchanged due to a chemical reaction}$$
$$\Delta Q_{water} \text{ is the heat received or released by water}$$

The water contribution is given by the heat formula given above. After we plug the formula of the heat into the formula above we arrive at the constant-pressure calorimetry formula:

$$\boxed{0 = n \cdot \Delta Q_r + V \cdot d \cdot c_e^{\text{sol}} \cdot (T_f - T_i)} \quad \text{constant pressure}$$

where:

ΔQ_r is the heat exchanged due to a chemical reaction in J/mol

n is the number of moles of the limiting reagent

V is the volume of water in mL contained in the calorimeter

d is the density of the solution in g/mL

c_e^{sol} is the specific heat of the solution: tend to be similar to water, 4.184J/g°C

$T_f - T_i = \Delta T$, is the temperature change from the initial to the final temperature

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Sample Problem 7

We mix 5mL of NaOH 0.5M with 5mL of HCl 0.5M both at 25°C in a constant-pressure calorimeter. The final temperature inside the calorimeter is 27°C. Calculate the heat of reaction if the solution density is 1g/mL and the specific heat of the solution is 4.184J/°C.

SOLUTION

Given	Asking
$n = 5 \times 10^{-3}L \cdot$ $0.5M{=}2.5{\times}10^{-3}$moles $V = 10mL$ $d = 1g/mL$ $T_f = 27°C$ $T_i = 25°C$ $c_e^{\text{H}_2\text{O}} =$ 4.184J/g°C	ΔQ_r

We have all data needed to solve the calorimetry formula. We have the moles of the limiting reagent, the overall volume of the mixture, the density of the mixture, the temperature change and the specific heat of the solution. Plugging all values into the calorimetry formula we have:

$$0 = n \cdot \Delta Q_r + V \cdot d \cdot c_e^{\text{sol}} \cdot (T_f - T_i)$$

$$0 = 2.5 \times 10^{-3}mol \cdot \Delta Q_r + 10g \cdot 4.184\text{J/g°C} \cdot (27°C - 25°C)$$

Solving for ΔQ_r we obtain $-33472 J/mol$ that is the same as $-33.5KJ/mol$. As the value is negative, it means that the reaction produced energy and hence is exothermic.

◈ STUDY CHECK

We mix 2.5mL of NaOH 0.5M with 2.5mL of HCl 0.5M both at 25°C in a constant-pressure calorimeter. The heat of reaction is -40KJ/mol. Calculate the final temperature inside the calorimeter, if the solution density is 1g/mL and the specific heat of the solution is 4.184J/g°C.

Constant-volume calorimeter A constant-volume calorimeter–also known as a bomb calorimeter–is a more complex and costly calorimeter in which normally gas phase reactions occur. This type of calorimeter is rigid and even if gas is produced

the volume of the container will not change–that is why is called a constant-volume calorimeter. Constant-volume calorimeters are used to calculate the energy value of food–and for example, calculate the calories in a bag of chips. The formula to carry calorimetric calculations with a constant-volume calorimeter has three terms: the first term represents the energy exchanged due to the reaction, the second term represents the energy exchanged by water in the calorimeter, and the last term represents the heat exchanged by the walls of the calorimeter. The formula used in a constant-volume calorimeter is the following:

$$0 = n \cdot \Delta Q_r + \Delta Q_{water} + \Delta Q_{walls}$$

where:

$n \cdot \Delta Q_r$ is the heat exchanged due to a chemical reaction

ΔQ_{water} is the heat received or released by water

ΔQ_{walls} is the heat absorbed by the walls

$$\boxed{0 = n \cdot \Delta Q_r + m \cdot c_e^{\text{H}_2\text{O}} \cdot (T_f - T_i) + c^{cal} \cdot (T_f - T_i)} \quad \text{Constant-volume}$$

where:

ΔQ_r is the heat exchanged due to a chemical reaction in J/mol

n is the number of moles that reacted inside the calorimeter

m is the mass of water contained in the calorimeter

$c_e^{\text{H}_2\text{O}}$ is the specific heat absorbed of water: $4.184\text{J/g}°\text{C}$

c^{cal} is the heat capacity of the calorimeter also known as the calorimeter factor

$T_f - T_i = \Delta T$, is the temperature change from the initial to the final temperature

Sample Problem 8

A 3 mol-sample of a chemical is burned in a constant-volume calorimeter with 10g of water and a heat capacity of $10KJ/°\text{C}$. Calculate the heat of reaction knowing that the initial temperature of the water inside the calorimeter is $25°C$ and the final $40°C$.

SOLUTION

	Given	Asking
Analyze the Problem	$n = 3mol$ $m = 10g$ $T_f = 40°C$ $T_i = 25°C$ $c^{cal} = 10^4\text{J/}°\text{C}$ $c_e^{\text{H}_2\text{O}} = 4.184\text{J/g}°\text{C}$	ΔQ_r

We have all data needed to solve the calorimetry formula. We have the moles of

chemical inside the calorimeter, the heat capacity of the calorimeter, the initial and final temperature of water, and the amount of water. Mind that the specific heat of water is always given and you need to remember the value. Also and more importantly mind that the units of the heat capacity of the calorimeter are $KJ/°C$, whereas the units of the specific heat of water are J/g°C and hence, we need to convert KJ into J; that is the reason we use $10000\text{J}/°C$ as the heat capacity of the calorimeter. Plugging all values into the calorimetry formula we have:

$$0 = 3mol \cdot \Delta Q_r + 10g \cdot 4.184\text{J/g}°C \cdot (40°C - 25°C)$$
$$+ 10000\text{J}/°C \cdot (40°C - 25°C)$$

Solving for ΔQ_r we obtain $-50209 J/mol$ that is the same as $-50.209 KJ/mol$. As the value is negative, it means that the reaction produced energy and hence is exothermic.

❖ STUDY CHECK

A 2 mol-sample of a chemical reacts in a constant-volume calorimeter with 20g of water and a heat capacity of $11KJ/°C$. Calculate the heat of reaction knowing that the temperature of water inside the calorimeter rises $10°C$.

▼A constant-volume calorimeter   ▼A constant-pressure calorimeter


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Figure 8.1 Different types of calorimeters. A constant-volume calorimeter is also called a bomb calorimeter. A constant-pressure calorimeter is also called a coffee-cup calorimeter

8.6 Reactions and energy

Exothermic and endothermic reactions Some reactions release heat; these are called exothermic reactions. Others absorb heat and are called endothermic reactions. Think for example the combustion of the gas in a cooking stove. This reaction produces heat and hence it is exothermic. Differently, bread needs heat to rise. Hence, the chemical reaction involved in bread making should be endothermic. Similarly, if you melt an ice cube you need to give energy to the cube so that it becomes water. The melting of ice is an endothermic reaction. Endothermic reactions have positive heat of

reaction ΔH_R and are normally accompanied by an indication of heat in the reactants side of the reaction before the arrow.
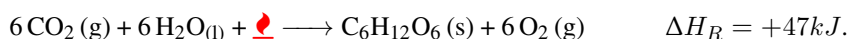
Reactants + 🔥 $\longrightarrow$ Products $\Delta H_R > 0$
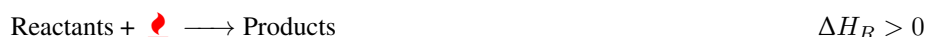
Differently, exothermic reactions normally accompanied by an expression of heat in the products side have negative ΔH_R.

Reactants $\longrightarrow$ Products + 🔥 $\Delta H_R < 0$

Sample Problem 9

Classify the following reactions as exothermic and endothermic:

$$2\,H_2\,(g) + O_{2\,(g)} \longrightarrow 2\,H_2O\,(g) + 🔥 \qquad \Delta H_R = -572 kJ.$$

$$6\,CO_2\,(g) + 6\,H_2O_{(l)} + 🔥 \longrightarrow C_6H_{12}O_6\,(s) + 6\,O_2\,(g) \qquad \Delta H_R = +47 kJ.$$
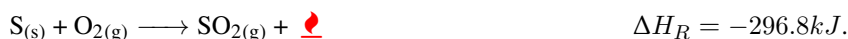
SOLUTION

The first reaction is exothermic as it produces hear. Differently, the second reaction is endothermic and absorbs energy.

◈ **STUDY CHECK**

Classify the following reactions as exothermic and endothermic:

$$C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 🔥 \qquad \Delta H_R = -393.5 kJ.$$

$$C_{(s)} + 2\,S_{(s)} + 🔥 \longrightarrow CS_{2\,(l)} \qquad \Delta H_R = +87.9 kJ.$$

$$S_{(s)} + O_{2\,(g)} \longrightarrow SO_{2\,(g)} + 🔥 \qquad \Delta H_R = -296.8 kJ.$$

Heat-Mole conversions Remember that a chemical reaction can be translated into a series of conversion factors that relate the moles of reactants with the products or with other reactants. At the same time, a chemical reaction involving heat can be converted into a series of conversion factors that are related to energy and the moles of reactants and products. For the exothermic reaction:
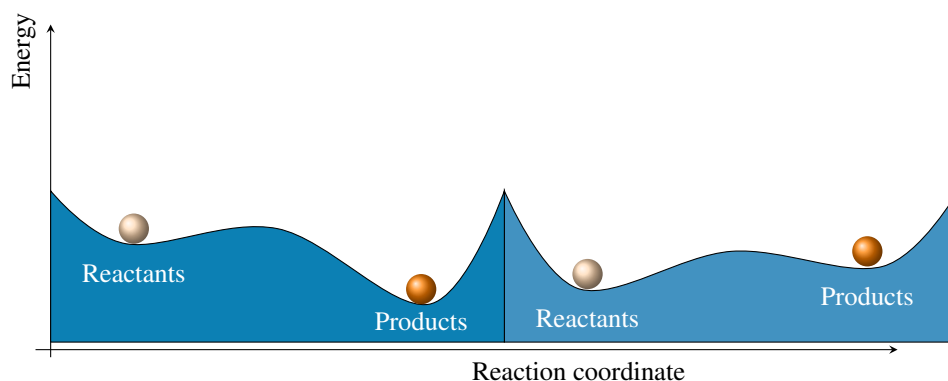
$$2\,H_2\,(g) + O_{2\,(g)} \longrightarrow 2\,H_2O\,(g) \qquad \Delta H_R = -572 KJ.$$

the moles of hydrogen are related to heat:

$$\frac{2 \text{ moles of } H_2}{\text{-572 KJ}} \quad \text{or} \quad \frac{\text{-572 KJ}}{2 \text{ moles of } H_2}$$

Similarly, we can relate energy with moles of O_2 or moles of water:

$$\frac{1 \text{ moles of } O_2}{\text{-572 KJ}} \quad \text{or} \quad \frac{\text{-572 KJ}}{1 \text{ moles of } O_2} \qquad \frac{2 \text{ moles of } H_2O}{\text{-572 KJ}} \quad \text{or} \quad \frac{\text{-572 KJ}}{2 \text{ moles of } H_2O}$$
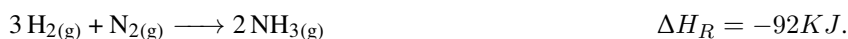
We will use these relationships to convert moles of reactants or products into heat.

Figure ?? An exothermic (left) reaction and and endothermic (right) reaction.

Sample Problem 10

Hydrogen reacts with nitrogen to produce ammonia (NH_3) according to the following reaction

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)} \qquad\qquad \Delta H_R = -92KJ.$$

Calculate: (a) the enthalpy of reaction; (b) indicate whether the reaction is endo or exothermic; (c) calculate the heat exchanged when produced 5 moles of ammonia.

SOLUTION

(a) the heat of reaction is -92KJ, and (b) the reaction is exothermic as the heat appears as a product. This means the reaction produces heat. (c) We will use the conversion factor that relates ammonia with heat and will set up the moles of ammonia on the bottom of the conversion factor so that the units will cancel and energy will remain

$$5\;\text{moles of } NH_3 \times \frac{\text{-92KJ}}{2\;\text{moles of } NH_3} = -230KJ,$$

that is: 5 moles of ammonia produce -230KJ. The fact that this value is negative means that heat will be released.

◈ **STUDY CHECK**

Given the reaction below, calculate the number of hydrogen moles needed to generate -200KJ.

$$3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)} \qquad\qquad \Delta H_R = -92KJ.$$

8.7 Energy value of food

Calories in food   How much food do you eat? How many calories do you ingest a day? When you are watching your food intake, the Calories you are counting are kilocalories (1000cal, Kcal, or Cal). In the field of nutrition, it is common to use the Calorie, Cal (with an uppercase C) to indicate 1000 cal or 1 kcal.

$$\boxed{1Cal = 1000cal} \qquad \text{or} \qquad \boxed{\frac{1Cal}{1000cal}} \qquad \text{or} \qquad \boxed{\frac{1000cal}{1Cal}} \qquad (8.8)$$

Table 8.2 Energy value of food	
Food Type	Energy value ($\frac{kcal}{g}$)
Carbohydrates	4
Fat	9
Protein	4

Energy values Do you ever eat pasta? Think about how does your body feel after you eat pasta? Normally, whenever you eat pasta in a few hours you need to eat again more food. Differently, whenever you eat meat, that is enough to keep you going for a longer time. Similarly, eating a salad for lunch brings you less energy than a pizza slice. This is because each type of food–each ingredient–contains different energy. We refer to this as the energy value of food ϵ. To compute the energy (E) provided by a certain mass of food (m) we need to multiply the mass times the energy value (ϵ):

$$E = m \cdot \epsilon \tag{8.9}$$

For example, the energy value of fat ϵ_{fat} is $9\frac{kcal}{g}$, which means that if you eat three grams of fat that will bring you a given amount of energy E_{fat}:

$$E_{fat} = 3\cancel{g} \times 9\frac{kcal}{\cancel{g}} = 18kcal$$

Normally, when you eat a food plate, you ingest energy from the different types of ingredients of that plate: fat, carbs, or protein.

Sample Problem 11

A Big Mac from McDonalds contains 28g of fat ($9kcal/g$), 46g of carbs ($4kcal/g$) and 25 g of protein ($4kcal/g$), where the caloric values are indicated in parenthesis. What is the total energy content of a Big Mac?

SOLUTION

① **Step one:** list of the given information and the unknown.

Given	Asking
$m_{fat} = 28g$ $\epsilon_{fat} = 9\frac{kcal}{g}$ $m_{carb} = 46g$ $\epsilon_{carb} = 4\frac{kcal}{g}$ $m_{prot} = 25g$ $\epsilon_{prot} = 4\frac{kcal}{g}$	Energy Content

Analyze the Problem

② **Step two:** use the formula $E_{fat} = m_{fat} \cdot \epsilon_{fat}$ to calculate the energy coming from fat. And do the same for carbs and protein.

③ **Step three:** Compute the energy coming from each ingredient and add all the values:

$$E_{fat} = m_{fat} \cdot \epsilon_{fat} = 28\cancel{g} \times 9\frac{kcal}{\cancel{g}} = 252kcal \qquad \text{energy from fat}$$

$$E_{carb} = m_{carb} \cdot \epsilon_{carb} = 46\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 184kcal \qquad \text{energy from carbs}$$

$$E_{prot} = m_{prot} \cdot \epsilon_{prot} = 25\cancel{g} \times 4\frac{kcal}{\cancel{g}} = 100kcal \qquad \text{energy from protein}$$

The total energy content of a Big Mac is: $E_{fat} + E_{carb} + E_{prot} = 532kcal$

◆ **STUDY CHECK**

A pepperoni pizza slice contains 11g of fat ($9kcal/g$), 36g of carbs ($4kcal/g$) and 14 g of protein ($4kcal/g$), where the caloric values are indicated in parenthesis. What is the total energy content of a pizza slice?

8.8 Enthalpy

In the last section, we have seen that when a chemical reaction proceeds it exchanges energy with the surroundings. This energy can be measured in many different conditions. When it is measured at constant pressure–these are regular conditions in chemistry, think about a reaction happening at a beaker–this energy change has a different name: it is called enthalpy and is represented with the symbol ΔH_f°. In this section, we will cover the different types of enthalpies depending on the type of reaction–formation or reaction–and we will find out how to compute the enthalpy change for a reaction using tables of standard enthalpies given at the end of the chapter.

What is enthalpy? You want to think about enthalpy as heat. There are different ways to measure the heat exchanged in a system–as constant-pressure heat or constant-volume heat. Enthalpy is the constant-pressure heat. It is important in chemistry, as many chemical reactions happen at constant pressure, that is, in open containers in contact with the atmosphere. Enthalpy is indeed related to the internal energy

$$H = U + PV$$

Working under constant pressure we have

$$\Delta H = \Delta U + P\Delta V$$

as $\Delta U = Q - P\Delta V$ we have

$$\Delta H = Q_p$$

where Q_p means heat measured at constant pressure.

Table of standard enthalpies The values of enthalpies are tabulated at the end of the chapter. The term standard refers to standard pressure conditions (1 atm) and is indicated by a degree sign on the top right side (ΔH°). Let us see how to use this table. If you look for the standard enthalpy of C–an element–from the table you might find several values. The values of graphite carbon are $\Delta H_f^\circ = 0$KJ/mol. Differently, the values for diamond carbon are different than zero, being $\Delta H_f^\circ = 1.0$KJ/mol. Similarly, the value for gas carbon is not zero also, being $\Delta H_f^\circ = 716.67$KJ/mol. This is because the natural state of carbon is in the form of graphite. That is, the most common way in which we find carbon in nature is in the form of graphite and not diamond or gas. Let us find the standard enthalpy for molecular nitrogen, N_2(g)–another element. If you look into the table you will find a value of $\Delta H_f^\circ = 0$KJ/mol, again because the natural state of nitrogen is in the form of gas N_2. How much is the standard enthalpy of gas hydrogen, H_2? If you look in the table, the value is also $\Delta H_f^\circ = 0$KJ/mol. The rule of thumb is: elements in their natural state have zero H_f°. Below we will explain more about the meaning of the natural state. Now, look for the standard enthalpy of carbon monoxide gas, CO(g). The value should not be zero, as carbon dioxide is not an element and is made of two different types of atoms. Indeed, in the table we find $\Delta H_f^\circ(CO_{(g)}) = -110.5$KJ/mol.

Using the enthalpy tables at the end of the chapter, locate the enthalpy values for: $CaS_{(s)}$, $AlCl_{3(s)}$ and $H_2O_{(l)}$.

SOLUTION

The enthalpy of calcium sulfide in solid state ($CaS_{(s)}$) is -482.4KJ/mol. For the case of aluminum chloride in solid state too ($AlCl_{3(s)}$), the enthalpy is -705.63KJ/mol. Finally, the enthalpy of liquid water is -285.8KJ/mol.

◈ **STUDY CHECK**

Using the enthalpy tables at the end of the chapter, locate the enthalpy values for: $I_{2(aq)}$, $F_{2(g)}$ and $C_{diamond(s)}$.

Natural state of an element The natural state of an element is the most stable state in which we find this element in nature. For example, you can find Aluminum as a solid, liquid, or gas. However, its natural state is solid, as we find Al in nature as a solid. That is the reason why $\Delta H_f^\circ(Al_{(g)}) = 314$KJ/mol, whereas $\Delta H_f^\circ(Al_{(s)}) = 0$KJ/mol. In general, metals, in their natural state is solid. For non-metallic elements, such as hydrogen, oxygen, nitrogen, fluorine, and chlorine, its natural state is in the form of a diatomic gas molecule. For example, $\Delta H_f^\circ(H_{2(g)}) = 0$KJ/mol, $\Delta H_f^\circ(N_{2(g)}) = 0$KJ/mol or $\Delta H_f^\circ(O_{2(g)}) = 0$KJ/mol. For the case of carbon, its natural state is graphite, $\Delta H_f^\circ(C_{graphite(s)}) = 0$KJ/mol. Molecules such as H_2O or NO have standard enthalpy different than zero. Mind that molecules are not elements, and hence are made of different elements.

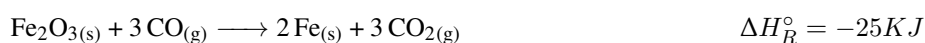| Table 8.3 Standard states for different elements. For all $\Delta H_f^\circ = 0$**KJ/mol**			
Element	Standard state	Element	Standard state
Hydrogen	$H_{2(g)}$	Oxygen	$O_{2(g)}$
Nitrogen	$N_{2(g)}$	Chlorine	$Cl_{2(g)}$
Iron	$Fe_{(s)}$	Aluminium	$Al_{(s)}$
Carbon	$C_{graphite(s)}$	Phosphorus	$P_{4(s)}$
Fluorine	$F_{2(g)}$	Bromine	$Br_{(l)}$
Mercury	$Hg_{(l)}$	Sulfur	$S_{8(s)}$
Iodine	$I_{2(s)}$	Silicon	$Si_{(s)}$

Formation reactions: standard enthalpy of formation The formation reaction starts with natural-state elements (e.g. H_2, N_2, etc.) to form a chemical. For example, the formation reaction of NO is given by:

$$N_{2(g)} + O_{2(g)} \longrightarrow 2\,NO_{(g)}$$
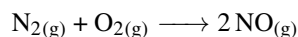
On the products side, we have the chemical formed (NO) whereas on the reactant side, we have the elements that make NO in its natural state (H_2 and N_2) The enthalpy associated with this reaction is called standard enthalpy of formation $\Delta H_f^\circ(NO_{(g)})$. This value is often listed on the right of the reaction:

$N_{2(g)} + O_{2(g)} \longrightarrow 2\,NO_{(g)}$ $\hspace{2cm}$ $\Delta H_f^\circ = 181KJ$

ΔH_R° *and* ΔH_f° Consider the following two reactions:

$N_{2(g)} + O_{2(g)} \longrightarrow 2\,NO_{(g)}$ $\hspace{2cm}$ $\Delta H_f^\circ = 181KJ$

$Fe_2O_{3(s)} + 3\,CO_{(g)} \longrightarrow 2\,Fe_{(s)} + 3\,CO_{2(g)}$ $\hspace{2cm}$ $\Delta H_R^\circ = -25KJ$

The first example represents a formation reaction and thus the enthalpy is labeled as ΔH_f°. In contrast, the second reaction is not a formation reaction. This is because the reactants are not elements in their natural state: $CO_{(g)}$ and $Fe_2O_{3(s)}$ have enthalpies different than zero. For this reason, the second enthalpy is labeled as ΔH_R° and is called the standard enthalpy of the reaction.

Standard enthalpy change for a reaction To calculate the standard enthalpy for a reaction you need to use the following formula:

$$\boxed{\Delta H_R^\circ = \Delta H_{products}^\circ - \Delta H_{reactants}^\circ} \quad \text{\textcolor{blue}{Enthalpy change}}$$
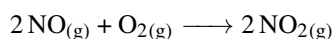
where:

ΔH_R° is the standard enthalpy change of the reaction

$\Delta H_{products}^\circ$ is the standard enthalpy of all products

$\Delta H_{reactants}^\circ$ is the standard enthalpy of all reactants

Now, imagine we need to calculate the change of standard enthalpy for the following reaction:

$$2\,NO_{(g)} + O_{2(g)} \longrightarrow 2\,NO_{2(g)}$$

We need to locate three enthalpies from the table: $\Delta H_f^\circ(NO_{(g)})$, $\Delta H_f^\circ(O_{2(g)})$, $\Delta H_f^\circ(NO_{2(g)})$. If you locate these values in the table you will see $\Delta H_f^\circ(O_{2(g)}) = 0$KJ/mol, whereas $\Delta H_f^\circ(NO_{(g)}) = 90.29$KJ/mol and $\Delta H_f^\circ(NO_{2(g)}) = 33.2$KJ/mol. Using the formula for ΔH_R° we have:
$\Delta H_R^\circ = \Delta H_{products}^\circ - \Delta H_{reactants}^\circ =$

$$= \left[2 \cdot \Delta H_f^\circ(NO_{2(g)})\right] - \left[2 \cdot \Delta H_f^\circ(NO_{(g)}) + \Delta H_f^\circ(O_{2(g)})\right] =$$
$$= \left[2 \cdot 33.2\right] - \left[2 \cdot 90.29 + 0\right] = -114KJ$$

This reaction is exothermic and releases heat.

Sample Problem 13

Using the enthalpy table, calculate ΔH_R° for the following reactions:

(a) $4\,H_{2(g)} + O_{2(g)} \longrightarrow 2\,H_2O_{(l)}$

(b) $3\,H_{2(g)} + N_{2(g)} \longrightarrow 2\,NH_{3(g)}$

(c) $2\,Al_{(s)} + 3\,Cl_{2(g)} \longrightarrow 2\,AlCl_{3(s)}$

SOLUTION

In order to answer all questions, we need a set of ΔH_f° values: $\Delta H_f^\circ(H_{2(g)})$, $\Delta H_f^\circ(O_{2(g)})$, $\Delta H_f^\circ(N_{2(g)})$, $\Delta H_f^\circ(Al_{(s)})$ are all zero, whereas $\Delta H_f^\circ(H_2O_{(l)}) = -285.8KJ/mol$, $\Delta H_f^\circ(NH_{3(g)}) = -45.0KJ/mol$ and $\Delta H_f^\circ(AlCl_{3(s)}) = -705.63KJ/mol$. For the first example, we have:

$$\Delta H_R^\circ = \left[2 \cdot \Delta H_f^\circ(H_2O_{(l)})\right] - \left[4 \cdot \Delta H_f^\circ(H_{2(l)}) + \Delta H_f^\circ(O_{2(g)})\right]$$
$$= \left[2 \cdot -285.8\right] - \left[4 \cdot 0 + 0\right] = -572KJ$$

For the second example:

$$\Delta H_R^\circ = \left[2 \cdot \Delta H_f^\circ(\text{NH}_{3\,(g)})\right] - \left[2 \cdot \Delta H_f^\circ(\text{Al}_{(s)}) + 3 \cdot \Delta H_f^\circ(\text{Cl}_{2\,(g)})\right]$$
$$= \left[2 \cdot -45\right] - \left[2 \cdot 0 + 3 \cdot 0\right] = -90 KJ$$

Finally, for the last reaction:

$$\Delta H_R^\circ = \left[2 \cdot \Delta H_f^\circ(\text{AlCl}_{3\,(s)})\right] - \left[3 \cdot \Delta H_f^\circ(\text{H}_{2\,(g)}) + \Delta H_f^\circ(\text{N}_{2\,(g)})\right]$$
$$= \left[2 \cdot -705.63\right] - \left[3 \cdot 0 + 0\right] = -1411 KJ$$

◈ **STUDY CHECK**

Using the enthalpy table, calculate ΔH_R° for the following reaction:

$$\text{Fe}_2\text{O}_{3\,(s)} + 3\,\text{CO}_{(g)} \longrightarrow 2\,\text{Fe}_{(s)} + 3\,\text{CO}_{2\,(g)}$$
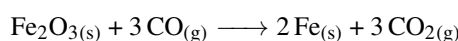
8.9 Hess's Law: Manipulating reaction enthalpies

In the previous section, we relied on a table el standard enthalpies of formation to compute enthalpy changes in general reactions. This enthalpy change ΔH_R° is related to the heat exchanged in the reaction. In this section, we will not use the tables of enthalpy anymore. Imagine you do not have access to this table. And we will find alternative ways to predict ΔH_f° given a series of reactions with known enthalpies. In short, you will have to identify the enthalpies that are zero–the enthalpies corresponding to an element in its natural state– and set up an equation that helps you find out the missing enthalpy.

Reverting reactions Imagine they give you the following reaction:

$$\text{N}_{2\,(g)} + \text{O}_{2\,(g)} \longrightarrow 2\,\text{NO}_{(g)} \qquad\qquad \Delta H_1^\circ = -114 KJ$$

and you need to calculate the enthalpy change for this other reaction:

$$2\,\text{NO}_{(g)} \longrightarrow \text{N}_{2\,(g)} + \text{O}_{2\,(g)} \qquad\qquad \Delta H_2^\circ$$

If you compare both reactions you will see the second reaction equals the first reaction but reverted. If you revert a reaction, the enthalpy change changes sign. Therefore, $\Delta H_2^\circ = 114 KJ$.

Timing reactions by a number Imagine they give you the following reaction:

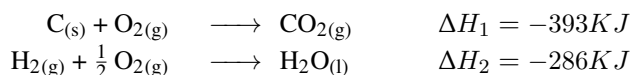$$\text{N}_{2\,(g)} + \text{O}_{2\,(g)} \longrightarrow 2\,\text{NO}_{(g)} \qquad\qquad \Delta H_1^\circ = -114 KJ$$

and you need to calculate the enthalpy change for this other reaction:

$$2\,\text{N}_{2\,(g)} + 2\,\text{O}_{2\,(g)} \longrightarrow 4\,\text{NO}_{(g)} \qquad\qquad \Delta H_2^\circ$$

If you compare both reactions you will see the second reaction equals the first reaction time by two. If you time a reaction by two, the enthalpy change should also be timed by two. Therefore, $\Delta H_2^\circ = 2 \cdot -114 = -228 KJ$.
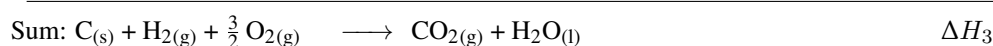
Combining reactions Imagine they give you the following two reactions:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H_1 = -393KJ$$
$$H_{2(g)} + \tfrac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)} \qquad \Delta H_2 = -286KJ$$

and then ask the enthalpy change for the following reaction:

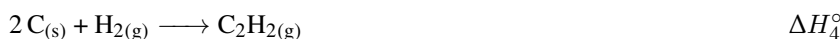$$C_{(s)} + H_{2(g)} + \tfrac{3}{2} O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(l)} \qquad\qquad\qquad \Delta H_3^\circ$$
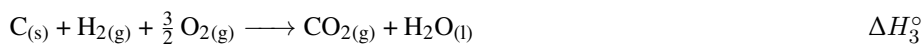
If you look closely at the last reaction, you will see it results from adding the first two reactions, so that:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad\qquad\qquad \Delta H_1 = -393KJ$$
$$H_{2(g)} + \tfrac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)} \qquad\qquad\qquad \Delta H_2 = -286KJ$$

Sum: $C_{(s)} + H_{2(g)} + \tfrac{3}{2} O_{2(g)} \longrightarrow CO_{2(g)} + H_2O_{(l)} \qquad\qquad \Delta H_3$
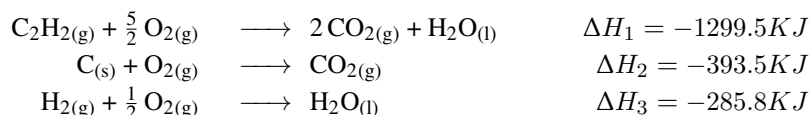
Therefore, $\Delta H_3 = \Delta H_1 + \Delta H_2 = -679KJ$. When adding two chemical reactions the resulting enthalpy is the result of adding the enthalpy of both reactions.

Sample Problem 14

Calculate the enthalpy for this reaction:

$$2\,C_{(s)} + H_{2(g)} \longrightarrow C_2H_{2(g)} \qquad\qquad\qquad \Delta H_4^\circ$$

Given the following thermochemical equations:

$$C_2H_{2(g)} + \tfrac{5}{2} O_{2(g)} \longrightarrow 2\,CO_{2(g)} + H_2O_{(l)} \qquad \Delta H_1 = -1299.5KJ$$
$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H_2 = -393.5KJ$$
$$H_{2(g)} + \tfrac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)} \qquad \Delta H_3 = -285.8KJ$$

SOLUTION

In order to get the enthalpy for reaction (4) we will have to combine reactions (1), (2) and (3), by adding, subtracting, or multiplying by a number so that the results adds up to reaction (4). A trick to do this is compare molecule by molecule in reaction (4) and see in which reaction we can find the same one. For example, reaction (4) contains $2C_{(s)}$ in the reactant side. $C_{(s)}$ can also be found in (2) also as reactant. However, in (2) $C_{(s)}$ is not timed by 2. There we will use two times reaction (2):

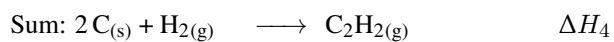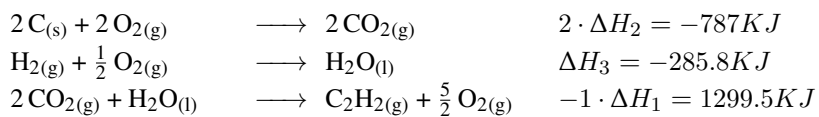$$2\,C_{(s)} + 2\,O_{2(g)} \longrightarrow 2\,CO_{2(g)} \qquad\qquad 2 \cdot \Delta H_2^\circ = -787$$

Reaction (4) also contains $H_{2(g)}$, which can be found in (3). There we will use (3) as it is:

$$H_{2(g)} + \tfrac{1}{2} O_{2(g)} \longrightarrow H_2O_{(l)} \qquad\qquad \Delta H_3^\circ = -285.8KJ$$

Reaction (4) also contains $C_2H_{2(g)}$ as a product. We can find the same chemical in (1) but as a reactant. There we will have to invert (1):

$$2\,CO_{2(g)} + H_2O_{(l)} \longrightarrow C_2H_{2(g)} + \tfrac{5}{2} O_{2(g)} \qquad\qquad -1 \cdot \Delta H_1 = 1299.5KJ$$
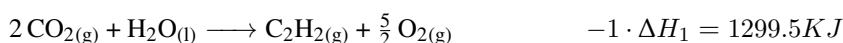
If we add the three previous reactions, we have:

$$2\,C_{(s)} + 2\,O_{2(g)} \longrightarrow 2\,CO_{2(g)} \qquad 2 \cdot \Delta H_2 = -787 KJ$$

$$H_{2(g)} + \tfrac{1}{2}\,O_{2(g)} \longrightarrow H_2O_{(l)} \qquad \Delta H_3 = -285.8 KJ$$

$$2\,CO_{2(g)} + H_2O_{(l)} \longrightarrow C_2H_{2(g)} + \tfrac{5}{2}\,O_{2(g)} \qquad -1 \cdot \Delta H_1 = 1299.5 KJ$$

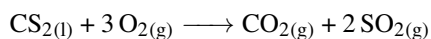$$\text{Sum: } 2\,C_{(s)} + H_{2(g)} \longrightarrow C_2H_{2(g)} \qquad \Delta H_4$$

Therefore in the enthalpy for the reaction (4) will be:

$$\Delta H_4^{\circ} = 2 \cdot \Delta H_2 + \Delta H_3 - 1 \cdot \Delta H_1 = 226.7 KJ$$

⬙ STUDY CHECK

Calculate the enthalpy for this reaction:

$$CS_{2(l)} + 3\,O_{2(g)} \longrightarrow CO_{2(g)} + 2\,SO_{2(g)}$$

Given the following thermochemical equations:

$$C_{(s)} + O_{2(g)} \longrightarrow CO_{2(g)} \qquad \Delta H_1 = -393.5 KJ$$

$$S_{(s)} + O_{2(g)} \longrightarrow SO_{2(g)} \qquad \Delta H_2 = -296.8 KJ$$

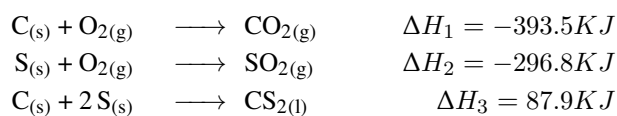$$C_{(s)} + 2\,S_{(s)} \longrightarrow CS_{2(l)} \qquad \Delta H_3 = 87.9 KJ$$

Table 8.4 Standard enthalpy table at 1atm and 298K in KJ/mol.

Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
Aluminum					
$Al_{(s)}$	0	$AlCl_{3(s)}$	-705.63	$Al_2O_{3(s)}$	-1675.5
$Al(OH)_{3(s)}$	-1277	$Al_2(SO_4)_{3(s)}$	-3440	$NH_{3(aq)}$	-80.8
$NH_{3(g)}$	-46.1	$NH_4NO_{3(s)}$	-365.6	$Al_{(g)}$	314
Barium					
$BaCl_{2(s)}$	-858.6	$BaCO_{3(s)}$	-1213	$Ba(OH)_{2(s)}$	-944.7
$BaO_{(s)}$	-548.1	$BaSO_{4(s)}$	-1473.2	$BaSO_{4(s)}$	-1473.2
Boron					
$BCl_{3(s)}$	-402.96				
Bromine					
$Br_{2(l)}$	0	$Br^-_{(aq)}$	-121	$Br_{(g)}$	111.884
$Br_{2(g)}$	30.91	$BrF_{3\,(g)}$	-255.60	$HBr_{(g)}$	-36.29
Cadmium					
$Cd_{(s)}$	0	$CdO_{(s)}$	-258	$Cd(OH)_{2(s)}$	-561
$CdS_{(s)}$	-162	$CdSO_{4(s)}$	-935		
Calcium					
$Ca_{(s)}$	0	$Ca_{(g)}$	178.2	$Ca^{2+}_{(g)}$	1925.90
$CaC_{2(s)}$	-59.8	$CaCO_{3(s)}$	-1206.9	$CaCl_{2(s)}$	-795.8
$CaCl_{2(aq)}$	-877.3	$Ca_3(PO_4)_{2(s)}$	-4132	$CaF_{2(s)}$	-1219.6
$CaH_{2(s)}$	-186.2	$Ca(OH)_{2(s)}$	-986.09	$Ca(OH)_{2(aq)}$	-1002.82
$CaO_{(s)}$	-635.09	$CaSO_{4(s)}$	-1434.52	$CaS_{(s)}$	-482.4
$CaSiO_{3(s)}$	-1630				
Caesium					
$Cs_{(s)}$	0	$Cs_{(g)}$	76.50	$Cs_{(l)}$	2.09
$Cs^+_{(g)}$	457.964	$CsCl_{(s)}$	-443.04		
Carbon					
$C_{graphite(s)}$	0	$C_{diamond(s)}$	1.9	$C_{(g)}$	716.67
$CO_{2(g)}$	-393.509	$CS_{2(l)}$	89.41	$CS_{2(g)}$	116.7
$CO_{(g)}$	-110.525	$COCl_{2(g)}$	-218.8	$CO_{2(aq)}$	-419.26
$HCO_3^-{}_{(aq)}$	-689.93	$CO_3^{2-}{}_{(aq)}$	-675.23		
Chlorine					
$Cl_{(g)}$	121.7	$Cl^-_{(aq)}$	-167.2	$Cl_{2(g)}$	0
Chromium					
$Cr_{(s)}$	0				
Copper					
$Cu_{(s)}$	0	$CuO_{(s)}$	-155.2	$CuSO_{4(aq)}$	-769.98
Fluorine					
$F_{2(g)}$	0				
Hydrogen					
$H_{(g)}$	218	$H_{2(g)}$	0	$H_2O_{(g)}$	-241.818
$H_2O_{(l)}$	-285.8	$H^+_{(aq)}$	0	$OH^-_{(aq)}$	-230
H_2O_2	-187.8	$H_3PO_{4(l)}$	-1288	$HCN_{(g)}$	130.5
$HBr_{(l)}$	-36.3	$HCl_{(g)}$	-92.30	$HCl_{(aq)}$	-167.2
$HF_{(g)}$	-273.3	$HI_{(g)}$	26.5		
Iodine					
$I_{2(s)}$	0	$I_{2(g)}$	62.438	$I_{2(aq)}$	23
$I^-_{(aq)}$	-55				

(cont.) Standard enthalpy table at 1atm and 298K.

Substance	ΔH_f°	Substance	ΔH_f°	Substance	ΔH_f°
Iron					
$Fe_{(s)}$	0	$Fe_3C_{(s)}$	5.4	$FeCO_{3(s)}$	-750.6
$FeCl_{3(s)}$	-399.4	$FeO_{(s)}$	-272	$Fe_3O_{4(s)}$	-1118.4
$Fe_2O_{3(s)}$	-824.2	$FeSO_{4(s)}$	-929	$Fe_2(SO_4)_{3(s)}$	-2583
$FeS_{(s)}$	-102	$FeS_{2(s)}$	-178		
Lead					
$Pb_{(s)}$	0	$PbO_{2(s)}$	-277	$PbS_{(s)}$	-100
$PbSO_{4(s)}$	-920	$Pb(NO_3)_{2(s)}$	-452	$PbO_{(s)}$	-276.6
Magnesium					
$Mg_{(s)}$	0	$Mg^{2+}_{(aq)}$	-466.85	$MgCO_{3(s)}$	-1095.7
$MgO_{(s)}$	-601.6	$MgSO_{4(s)}$	-1278.2	$MgCl_{2(s)}$	-641.8
Manganese					
$Mn_{(s)}$	0	$MnO_{(s)}$	-384.9	$MnO_{2(s)}$	-519.7
$Mn_2O_{3(s)}$	-971	$Mn_3O_{4(s)}$	-1387	$MnO_4{-}_{(aq)}$	-543
Mercury					
$HgO_{(s)}$	-90.83	$HgS_{(s)}$	-58.2		
Nitrogen					
$N_{2(g)}$	0	$NH_{3(aq)}$	-80.8	$NH_{3(g)}$	-45.90
NH_4Cl	-314.55	$NO_{2(g)}$	33.2	$N_2O_{(g)}$	82.05
$NO_{(g)}$	90.29	$N_2O_{4(g)}$	9.16	$N_2O_{5(s)}$	-43.1
Oxygen					
$O_{(g)}$	249	$O_{2(g)}$	0	$O_{3(g)}$	143
Phosphorus					
$P_{4(s)}$	0	$P_{red(s)}$	-17.4	$P_{black(s)}$	-39.3
$PCl_{3(l)}$	-319.7	$PCl_{3(g)}$	-278	$PCl_{5(s)}$	-440
$PCl_{5(g)}$	-321	$P_2O_{5(s)}$	-1505.5		
Potassium					
$KBr_{(s)}$	-392.2	$K_2CO_{3(s)}$	-1150	$KClO_{3(s)}$	-391.4
$KCl_{(s)}$	-436.68	$KF_{(s)}$	-562.6	$K_2O_{(s)}$	-363
$KClO_{4(s)}$	-430.12				
Silicon					
$Si_{(g)}$	368.2	$SiC_{(s)}$	-74.4	$SiCl_{4(l)}$	-640.1
$SiO_{2(s)}$	-910.86				
Silver					
$AgBr_{(s)}$	-99.5	$AgCl_{(s)}$	-127.01	$AgI_{(s)}$	-62.4
$Ag_2O_{(s)}$	-31.1	$Ag_2S_{(s)}$	-31.8		
Sodium					
$Na_{(s)}$	0	$Na_{(g)}$	+107.5	$NaHCO_{3(s)}$	-950.8
$Na_2CO_{3(s)}$	-1130.77	$NaCl_{(aq)}$	-407.27	$NaCl_{(s)}$	-411.12
$NaF_{(s)}$	-569.0	$NaOH_{(aq)}$	-469.15	$NaOH_{(s)}$	-425.93
$Na_2O_{(s)}$	-414.2				
Sulfur					
$S_{8monoclinic(s)}$	0.3	$S_{8rhombic(s)}$	0	$H_2S_{(g)}$	-20.63
$SO_{2(g)}$	-296.84	$SO_{3(g)}$	-395.7	$H_2SO_{4(l)}$	-814
Titanium					
$Ti_{(s)}$	0	$Ti_{(g)}$	468	$TiCl_{4(g)}$	-763.2
$TiCl_{4(l)}$	-804.2	$TiO_{2(s)}$	-944.7		
Zinc					
$Zn_{(g)}$	130.7	$ZnCl_{2(s)}$	-415.1	$ZnO_{(s)}$	-348.0