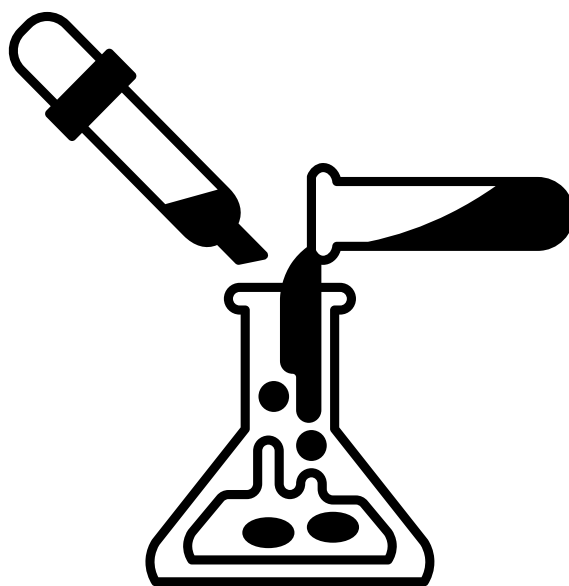


CHAPTER 1

ENERGY AND CHEMICAL CHANGE



CHAPTER SUMMARY

by Abd Alrahman Basim

SECTION 1:
ENERGY

SECTION 2:
HEAT

SECTION 3:
THERMOCHEMICAL EQUATIONS

SECTION 4:
CALCULATING ENTHALPY CHANGE

SECTION 5:
REACTION SPONTANEITY

This file is provided by [InLibrary](#).

ENERGY

The Nature of Energy

Energy is the ability to do work or produce heat. It exists in **two** basic forms:

- **Potential Energy**
is energy due to the **composition** or **position** of an object.
- **Kinetic Energy**
is energy of **motion**.

Chemical systems contain both kinetic energy and potential energy.

The kinetic energy of a substance is directly related to the constant random motion of its representative particles and is **proportional** to temperature. As temperature increases, the motion of submicroscopic particles increases.

The potential energy of a substance depends on its **composition**:

- the type of atoms in the substance,
- the number and type of chemical bonds joining the atoms,
- and the particular way the atoms are arranged.

Law of Conservation of Energy

The **law of conservation of energy** states that in any chemical reaction or physical process, energy can be converted from one form to another, but it is neither created nor destroyed. This is also known as the **first law of thermodynamics**.

Chemical Potential Energy

Chemical potential energy is the energy that is stored in a substance because of its composition.

Chemical potential energy plays an important role in chemical reactions. For example, the chemical potential energy of **propane (C₃H₈)** results from the arrangement of the carbon and hydrogen atoms and the strength of the bonds that join them.

Heat

The principal component of gasoline is **octane (C₈H₁₈)**. When gasoline burns in an automobile's engine, some of the chemical potential energy is converted to the work of moving the wheels. However, much of the chemical potential energy of octane is released as heat.

Heat (q): energy that is in the process of flowing from a warmer object to a cooler object.

ENERGY

Measuring Heat

Calorie (cal) is the amount of energy required to raise the temperature of **1 g** of pure water by **1°C**.

When your body breaks down sugars and fats to form carbon dioxide and water, these exothermic reactions generate heat that can be measured in Calories.

Note that the nutritional Calorie is capitalized. That is because one nutritional Calorie equals 1000 calories, or one kilocalorie (kcal).

Joule (J) is the SI unit of energy and heat.

Table 1 Relationships Among Energy Units	
Relationship	Conversion Factors
1 J = 0.2390 cal	$\frac{1 \text{ J}}{0.2390 \text{ cal}}$ $\frac{0.2390 \text{ cal}}{1 \text{ J}}$
1 cal = 4.184 J	$\frac{1 \text{ cal}}{4.184 \text{ J}}$ $\frac{4.184 \text{ J}}{1 \text{ cal}}$
1 Calorie = 1 kcal	$\frac{1 \text{ Calorie}}{1000 \text{ cal}}$ $\frac{1000 \text{ cal}}{1 \text{ Calorie}}$

Table 1 summarizes the relationships between calories, nutritional Calories, joules, and kilojoules (kJ) and the conversion factors you can use to convert from one unit to another.

Specific Heat

The **specific heat** of any substance is the amount of heat required to raise the temperature of **1 g** of that substance by **1°C**. Because different substances have different compositions, each substance has its own specific heat.

To raise the temperature of water by one degree Celsius, 4.184 J must be absorbed by every gram of water. Much less energy is required to raise the temperature of an equal mass of concrete by one degree Celsius.

The specific heat of concrete is **0.84 J/(g · °C)**, which means that the temperature of concrete increases roughly **five** times more than water's temperature when equal masses of concrete and water absorb the same amount of energy.

Calculating Heat Absorbed

Equation for Calculating Heat

$$q = c \times m \times \Delta T$$

The quantity of heat absorbed or released by a substance is equal to the product of its specific heat, the mass of the substance, and the change in its temperature.

Calculating Heat Released

The same equation for **q**, the quantity of heat, can be used to calculate the energy released by substances when they cool off.

HEAT

Calorimetry

Calorimeter is an insulated device used for measuring the amount of heat absorbed or released during a chemical or physical process.

A known mass of water is placed in an **insulated** chamber to absorb the energy released from the reacting system or to provide the energy absorbed by the system. The data to be collected is the **change in temperature** of this mass of water.

Determining Specific Heat

The foam-cup calorimeters provide satisfactory results. These calorimeters are open to the atmosphere, so reactions carried out in them occur at constant pressure.

They can be used to determine the specific heat of an unknown metal, for **example**:

- 1) An initial temperature of **25.60 °C** is recorded for the **125 g** of water in the calorimeter.
- 2) A **50.0 g** sample of an unknown metal is heated to **115.0 °C** and placed in the calorimeter.
- 3) The metal transfers heat to the water until metal and water are at the same temperature. The final temperature is **29.30 °C**.

Assuming **no** heat is lost to the surroundings, the **heat gained** by the water is **equal** to the **heat lost** by the metal. This quantity of heat can be calculated using the equation in Section 1 ($q = c \times m \times \Delta T$):

- First, calculate the heat gained by the water,

$$q_{\text{water}} = 4.184 \text{ J/(g} \cdot ^\circ\text{C)} \times 125 \text{ g} \times (29.30^\circ\text{C} - 25.60^\circ\text{C})$$

$$q_{\text{water}} = 4.184 \text{ J/(g} \cdot ^\circ\text{C)} \times 125 \text{ g} \times 3.70^\circ\text{C}$$

$$q_{\text{water}} = 1940 \text{ J}$$

- The heat gained by the water, **1940 J**, equals the heat lost by the metal, q_{metal} , so you can write this equation:

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

$$q_{\text{metal}} = -1940 \text{ J}$$

$$c_{\text{metal}} \times m \times \Delta T = -1940 \text{ J}$$

- Solve the equation for the **specific heat** of the metal. Substitute the known values of m and ΔT (50.0 g and -85.7°C) into the equation and solve.

$$c_{\text{metal}} = \frac{-1940 \text{ J}}{m \times \Delta T} \quad c_{\text{metal}} = \frac{-1940 \text{ J}}{(50.0 \text{ g})(-85.7^\circ\text{C})} = 0.453 \text{ J/(g} \cdot ^\circ\text{C)}$$

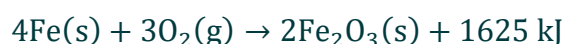
The unknown metal has a specific heat of **0.453 J/(g · °C)**.

HEAT

Chemical Energy and the Universe

Thermochemistry is the study of heat changes that accompany chemical reactions and phase changes.

You might have used a heat pack to warm your hands on a cold day. The energy released by a heat pack is produced by the following reaction and is shown in the equation as one of the products.



Because you are interested in the heat given off by the chemical reaction going on inside the pack, it is convenient to think of the pack and its contents as the system. In **thermochemistry**, the **system** is the specific part of the universe that contains the reaction or process you wish to study.

Everything in the universe other than the system is considered the **surroundings**. Therefore, the **universe** is defined as the system plus the surroundings.

$$\text{Universe} = \text{system} + \text{surroundings}$$

During this **exothermic** heat-pack reaction, heat produced by the reaction flows from the **heat pack** (the system) to your **cold hands** (part of the surroundings).

In an **endothermic reaction**, the flow of heat is **reversed**. Heat flows from the surroundings to the system.

Enthalpy and Enthalpy Changes

The total amount of energy a substance contains depends on many factors, some of which are still not completely understood. Therefore, it is **impossible** to know the total energy content of a substance.

Fortunately, chemists are usually more interested in **changes** in energy during reactions than in the absolute amounts of energy contained in the reactants and products.

For many reactions, the amount of energy lost or gained can be measured conveniently in a calorimeter at constant pressure.

Many reactions take place at constant atmospheric pressure; for example, those that occur in living organisms on Earth's surface, in lakes and oceans, and those that take place in open beakers and flasks in the laboratory.

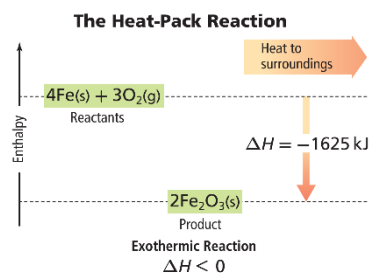
To more easily measure and study the energy changes that accompany such reactions, chemists have defined a property called enthalpy. **Enthalpy (H)** is the heat content of a system at constant pressure.

Although you cannot measure the actual energy or enthalpy of a substance, you can measure the **change in enthalpy**, which is the heat absorbed or released in a chemical reaction.

HEAT

Enthalpy (heat) of reaction (ΔH_{rxn}) is the change in enthalpy for a reaction. ΔH_{rxn} is the difference between the enthalpy of the substances that exist at the **end** of the reaction and the enthalpy of the substances present at the **start**.

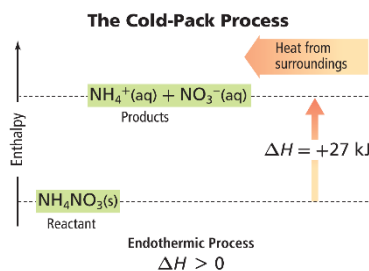
$$\Delta H_{\text{rxn}} = H_{\text{products}} - H_{\text{reactants}}$$



Recall the heat-pack reaction. According to the equation, the reactants in this **exothermic** reaction **lose** heat. Enthalpy changes for exothermic reactions are always negative. The equation for the heat-pack reaction and its enthalpy change are usually written as shown.



Now, recall the cold-pack process. For this **endothermic** process, $H_{\text{products}} > H_{\text{reactants}}$. Therefore, when $H_{\text{reactants}}$ is subtracted from the larger H_{products} , a positive value for ΔH_{rxn} is obtained. Chemists write the equation for the cold-pack process and its enthalpy change in the following way.



In this process, the enthalpy of the **products** is **27 kJ** greater than the enthalpy of the **reactant** because energy is **absorbed**. Thus, the sign of ΔH_{rxn} for this and all **endothermic** reactions and processes is **positive**.

The **enthalpy change**, ΔH , is **equal** to q_p , **the heat gained or lost in a reaction or process carried out at constant pressure**. Because all reactions presented in this textbook occur at constant pressure, you can assume that $q = \Delta H_{\text{rxn}}$.

THERMOCHEMICAL EQUATIONS

Writing Thermochemical Equations

A thermochemical equation is a **balanced** chemical equation that includes the **physical states** of all reactants and products and the **energy change**, usually expressed as the change in enthalpy, ΔH .

The **enthalpy (heat) of combustion** (ΔH_{comb}) of a substance is the enthalpy change for the complete burning of one mole of the substance.

Standard enthalpy changes have the symbol ΔH° . The zero superscript tells you that the enthalpy changes were determined with all reactants and products at **standard conditions**.

Standard conditions are 1 atm pressure and 298 K (25°C) and should not be confused with standard temperature and pressure (STP).

Changes of State

Molar enthalpy (heat) of vaporization (ΔH_{vap}) is the heat required to vaporize one mole of a liquid.

Molar enthalpy (heat) of fusion (ΔH_{fus}) is the heat required to melt one mole of a solid substance.

Because vaporizing a liquid and melting a solid are **endothermic processes**, their ΔH values are **positive**.

Thermochemical Equations for Changes of State

The **vaporization** of water and the **melting** of ice can be described by the following equations:

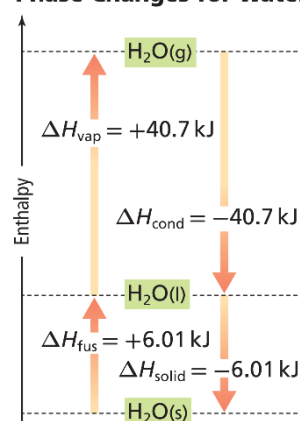


What happens in the reverse processes is that the same amounts of energy are released with the same numerical values but with opposite signs as shown in the figure.

$$\Delta H_{\text{vap}} = -\Delta H_{\text{cond}}$$

$$\Delta H_{\text{fus}} = -\Delta H_{\text{solid}}$$

Phase Changes for Water



THERMOCHEMICAL EQUATIONS

When a mole of glucose is burned in the body, 2808 kJ of energy is released. The process takes place in every cell of your body in a series of complex steps in which glucose is broken down and carbon dioxide and water are released.

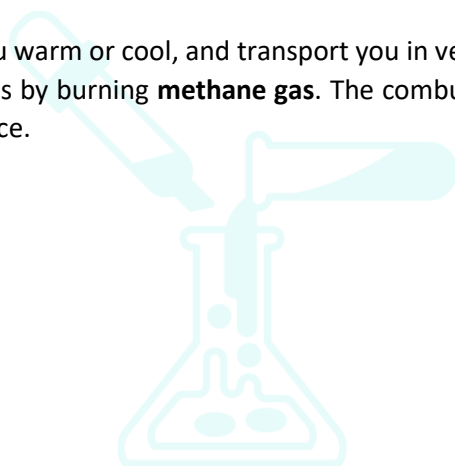
The energy released is stored as **chemical potential energy** in the **bonds** of molecules of **adenosine triphosphate** (ATP). When energy is needed by any part of the body, molecules of ATP release their energy.

Combustion Reactions

Combustion is the reaction of a fuel with oxygen. In biological systems, food is the fuel.

Most vehicles—cars, airplanes, boats, and trucks—run on the combustion of gasoline, which is mostly octane (C_8H_{18}).

Combustion reactions keep you warm or cool, and transport you in vehicles. One way you might heat your home or cook your food is by burning **methane gas**. The combustion of hydrogen provides the energy to lift a shuttle into space.



CALCULATING ENTHALPY CHANGE

Hess's Law

Sometimes it is impossible or impractical to measure the ΔH of a reaction by using a calorimeter. The conversion of carbon in its allotropic form, diamond, to carbon in its allotropic form, graphite, occurs so slowly that measuring the enthalpy change is impossible.

Other reactions occur under conditions difficult to duplicate in a laboratory. For these reactions, chemists use a theoretical way to determine ΔH .

Hess's law states that if you can add two or more thermochemical equations to produce a final equation for a reaction, then the sum of the enthalpy changes for the individual reactions is the enthalpy change for the final reaction.

In other words, the heat change in a chemical reaction is the same regardless of the number of stages in which the reaction is affected.

Applying Hess's Law

Follow these steps to calculate the energy change for the reaction that produces SO_3 using Hess's law:



Step 1

Chemical equations that contain the substances found in the desired equation and have known enthalpy changes are needed. The following equations contain S, O_2 , and SO_3 .

- a. $\text{S}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) \quad \Delta H = -297 \text{ kJ}$
- b. $2\text{SO}_3(\text{g}) \rightarrow 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \quad \Delta H = 198 \text{ kJ}$

Step 2

The desired equation shows **two** moles of **sulfur** reacting, so rewrite Equation (a) for two moles of sulfur by **multiplying** the coefficients by **two**.

Double the enthalpy change ΔH because twice the energy will be released if two moles of sulfur react. With these changes, Equation (a) becomes the following (**Equation c**).

- c. $2\text{S}(\text{s}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{SO}_2(\text{g}) \quad \Delta H = 2(-297)\text{kJ} = -594 \text{ kJ}$

Step 3

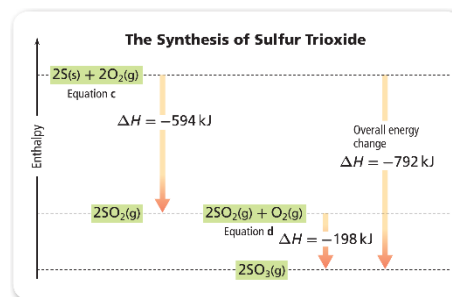
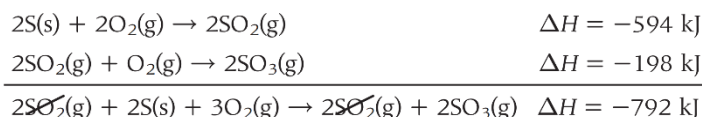
In the desired equation, **sulfur trioxide** is a **product** rather than a reactant, so **reverse** Equation (b). When you reverse an equation, you must also **change** the sign of its ΔH . Equation (b) then becomes (**Equation d**).

- d. $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g}) \quad \Delta H = -198 \text{ kJ}$

CALCULATING ENTHALPY CHANGE

Step 4

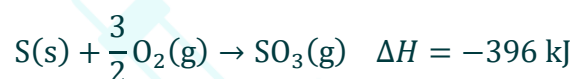
Add **Equations (c)** and **(d)** to obtain the desired reaction. Add the corresponding ΔH values. Cancel any terms that are common to both sides of the combined equation.



The **thermochemical equation** for the burning of **sulfur** to form **sulfur trioxide** is as follows.



Thermochemical equations are usually written and balanced for **one mole** of product. Often, that means that **fractional coefficients** must be used. For example, the thermochemical equation for the reaction between sulfur and oxygen to form one mole of sulfur trioxide is the following.



Standard Enthalpy (Heat) of Formation

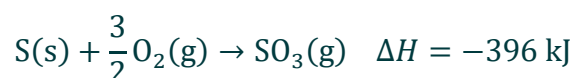
Hess's law allows you to calculate **unknown** ΔH values using known reactions and their experimentally determined ΔH values. However, recording ΔH values for all known chemical reactions would be a huge and unending task.

Instead, scientists record and use **enthalpy changes** for only one type of reaction – a reaction in which a compound is formed from its elements in their **standard states**.

The **standard state** of a substance means the normal physical state of the substance at **1 atm** and **298 K (25°C)**. For solids, **endothermic processes**, their ΔH values are **positive**. For example, in their standard states,

- iron is a solid,
- mercury is a liquid,
- and oxygen is a diatomic gas.

The **standard enthalpy (heat) of formation** (ΔH_f°) is defined as the change in enthalpy that accompanies the formation of one mole of the compound in its standard state from its elements in their standard states. A typical standard heat of formation reaction is the formation of one mole of SO_3 from its elements.



The product of this equation is SO_3 , a suffocating gas that produces acid rain when mixed with moisture in the atmosphere. Acid precipitation slowly destroys trees and property

CALCULATING ENTHALPY CHANGE

Where do standard heats of formation come from?

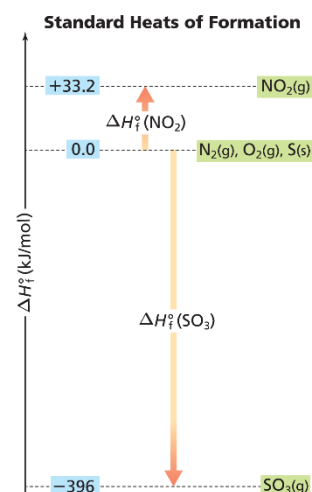
Standard enthalpies of formation are stated based on the following assumption: Elements in their standard state have a ΔH_f° of **0.0 kJ**. With zero as the starting point, the experimentally determined enthalpies of formation of compounds can be placed on a scale above and below the elements in their standard states.

Standard enthalpies of formation of many compounds have been measured experimentally. For example, consider the equation for the formation of nitrogen dioxide.



The elements **nitrogen** and **oxygen** are diatomic gases in their standard states, so their standard enthalpies of formation are zero.

When nitrogen and oxygen gases react to form one mole of **nitrogen dioxide**, the experimentally determined ΔH for the reaction is **+33.2 kJ**. That means that 33.2 kJ of energy is absorbed in this endothermic reaction.

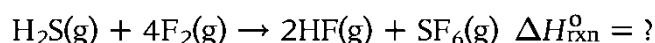


The energy content of the product **NO₂** is **33.2 kJ** greater than the energy content of the reactants. On a scale on which ΔH_f° of reactants is **0.0 kJ**, ΔH_f° of **NO₂** is **+33.2 kJ**.

Sulfur trioxide (SO₃) is placed **396 kJ** below zero on the scale because the formation of **SO₃(g)** is an **exothermic** reaction. The energy content of the sulfur trioxide, ΔH_f° , is **-396 kJ**.

Using standard enthalpies of formation

Suppose you want to calculate $\Delta H_{\text{rxn}}^\circ$ for a reaction that produces **sulfur hexafluoride**.



Compound	Formation Equation	ΔH_f° (kJ/mol)
H ₂ S(g)	H ₂ (g) + S(s) → H ₂ S(g)	-21
HF(g)	$\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{F}_2(\text{g}) \rightarrow \text{HF}(\text{g})$	-273
SO ₃ (g)	S(s) + $\frac{3}{2}\text{O}_2(\text{g}) \rightarrow \text{SO}_3(\text{g})$	-396
SF ₆ (g)	S(s) + 3F ₂ (g) → SF ₆ (g)	-1220

Step 1

Refer to **Table 5** to find an equation for the formation of each of the **three** compounds in the desired equation—**HF**, **SF₆**, and **H₂S**.

- $\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{F}_2(\text{g}) \rightarrow \text{HF}(\text{g}) \quad \Delta H_f^\circ = -273 \text{ kJ}$
- $\text{S}(\text{s}) + 3\text{F}_2(\text{g}) \rightarrow \text{SF}_6(\text{g}) \quad \Delta H_f^\circ = -1220 \text{ kJ}$
- $\text{H}_2(\text{g}) + \text{S}(\text{s}) \rightarrow \text{H}_2\text{S}(\text{g}) \quad \Delta H_f^\circ = -21 \text{ kJ}$

CALCULATING ENTHALPY CHANGE

Step 2

Equations **a** and **b** describe the formation of the products **HF** and **SF₆** in the desired equation, so use Equations **a** and **b** in the direction in which they are written.

Equation **c** describes the formation of a product, **H₂S**, but in the desired equation, **H₂S** is a reactant. **Reverse** Equation **c** and change the sign of its ΔH_f° .



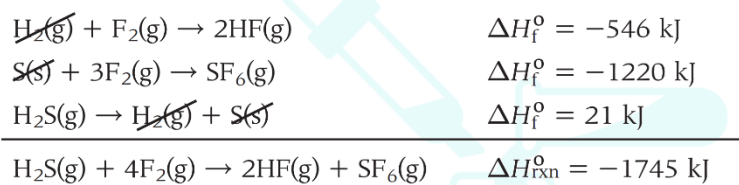
Step 3

Two moles of **HF** are required. Multiply Equation **a** and its enthalpy change by **two**.



Step 4

Add the three equations and their enthalpy changes. The elements **H₂** and **S** cancel.



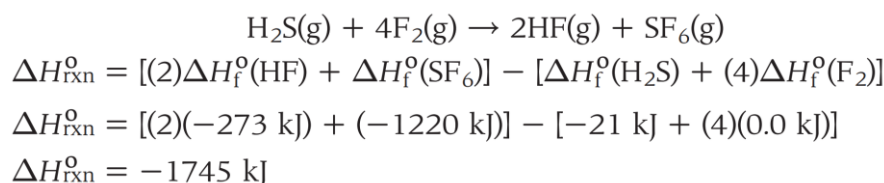
The Summation Equation

The stepwise procedure you have just read about shows how standard heats of formation equations combine to produce the desired equation and its $\Delta H_{\text{rxn}}^\circ$. The procedure can be summed up in the following formula.

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$$

- $\Delta H_{\text{rxn}}^\circ$ represents the standard enthalpy of the reaction.
- Σ represents the sum of the terms.
- ΔH_f° (products) and ΔH_f° (reactants) represent the standard enthalpies of formation of all the products and all the reactants.

You can see how this formula applies to the reaction between **hydrogen sulfide** and **fluorine**.



REACTION SPONTANEITY

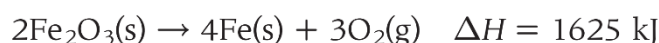
Spontaneous Processes

The heat pack goes into action the moment you activate it. Similarly, unprotected iron objects rust whether you want them to or not. Rusting is spontaneous.



Spontaneous process is any physical or chemical change that once begun, occurs with no outside intervention. However, for many spontaneous processes, some energy from the surroundings must be supplied to get the process started.

Suppose you reverse the direction of the equation for the rusting of iron. Recall that when you change the direction of a reaction, the sign of ΔH changes. The reaction becomes endothermic.



Reversing the equation will not make rust decompose spontaneously into iron and oxygen under ordinary conditions. The equation represents a reaction that is not spontaneous.

The formation of rust on iron is an **exothermic** and **spontaneous** reaction. The **reverse** reaction is **endothermic** and **nonspontaneous**.

You might conclude that all exothermic processes are spontaneous, and all endothermic processes are nonspontaneous. But remember that **ice melting** at room temperature is a spontaneous, endothermic process.

Entropy is another factor other than ΔH that plays a role in determining whether a chemical process occurs spontaneously under a given set of conditions.

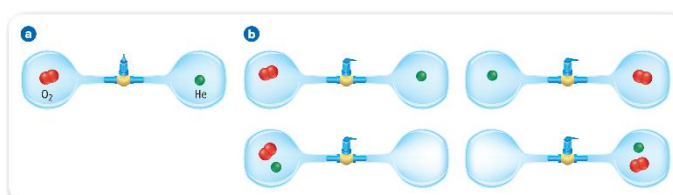
What is Entropy?

When gases spread out, a system reaches a state of **maximum entropy**.

Entropy (S) is

- a measure of the number of possible ways that the energy of a system can be distributed,
- and this is related to the freedom of the system's particles to move
- and number of ways they can be arranged.

Consider the two bulbs in the figure shown. As the number of particles increases, the number of possible arrangements for a group of particles increases dramatically.



REACTION SPONTANEITY

In general, the number of possible arrangements available to a system increases under the following conditions:

- when volume increases,
- when energy increases,
- when the number of particles increases,
- or when the particles' freedom of movement increases.

The Second Law of Thermodynamics

The tendency toward increased entropy is summarized in the **second law of thermodynamics**, which states that spontaneous processes always proceed in such a way that the entropy of the universe increases.

Entropy is sometimes considered to be a measure of the **disorder** or **randomness** of the particles that make up a system. Particles that are **more spread** out are said to be **more disordered**, causing the system to have greater entropy than when the particles are closer together.

Predicting Changes in Entropy

The **change in entropy** (ΔS) during a reaction or process is

$$\Delta S_{\text{system}} = S_{\text{products}} - S_{\text{reactants}}$$

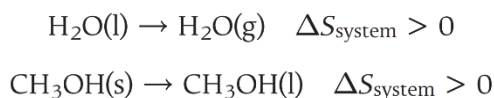
- If the entropy of a system **increases** during a reaction or process, ΔS_{system} is **positive**.
- Conversely, if the entropy of a system **decreases** during a reaction or process, ΔS_{system} is **negative**.

You can sometimes predict if ΔS_{system} is positive or negative by examining the equation for a reaction or process.

1) Entropy changes associated with changes in state can be predicted.

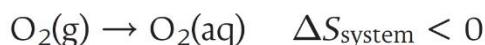
In solids, molecules have **limited** movement. In liquids, they have **some** freedom to move, and in gases, molecules can move **freely** within their container.

Thus, entropy increases as a substance changes from a solid to a liquid and from a liquid to a gas. ΔS_{system} is positive as water vaporizes and methanol melts.



2) The dissolving of a gas in a solvent always results in a decrease in entropy.

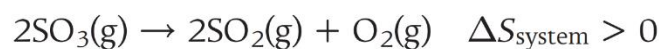
Gas particles have **more** entropy when they can move **freely** than when they are **dissolved** in a liquid or solid that limits their movements and randomness. ΔS_{system} is negative for the dissolving of oxygen in water.



REACTION SPONTANEITY

- 3) *Assuming no change in physical state occurs, the entropy of a system usually increases when the number of gaseous **PRODUCT** particles is greater than the number of gaseous **REACTANT** particles.*

For the following reaction, ΔS_{system} is **positive** because **two** molecules of gas react and **three** molecules of gas are produced.



- 4) *With some exceptions, entropy increases when a solid or a liquid dissolves in a solvent.*

The solute particles, which are together before dissolving, become **dispersed** throughout the solvent. The solute particles have **more** freedom of movement. For the dissolving of sodium chloride in water ΔS_{system} is **positive**.



- 5) *The random motion of the particles of a substance increases as its **TEMPERATURE** increases.*

Increased kinetic energy means **faster** movement and more possible arrangements of particles. Therefore, the entropy of any substance increases as its temperature increases. ΔS_{system} is **positive**.

Earth's Spontaneous Processes

Volcanoes, fumaroles, hot springs, and geysers are evidence of geothermal energy in Earth's interior.

- **Volcanoes** are vents in Earth's crust from which molten rock (magma), steam, and other materials flow.
- **Hot spring** is a naturally occurring spring of water that emerges due to heated groundwater.
- **Geysers** are hot springs that spout hot water and steam into the air.
- **Fumaroles** emit steam and other gases, such as hydrogen sulfide.

These geothermal processes are obviously spontaneous.

Entropy, the Universe, and Free Energy

The following is true for any spontaneous process.

$$\Delta S_{\text{universe}} > 0$$

Any change in the entropy of the universe is the sum of changes occurring in the system and surroundings.

$$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

REACTION SPONTANEITY

In nature, $\Delta S_{\text{universe}}$ tends to be **positive** for reactions and processes under the following conditions.

- **The reaction or process is exothermic, which means ΔH_{system} is negative.**
The heat released by an exothermic reaction **raises** the temperature of the surroundings and thereby increases the entropy of the surroundings. $\Delta S_{\text{surroundings}}$ is **positive**.
- **The entropy of the system increases, so ΔS_{system} is positive.**

Thus, **exothermic** chemical reactions accompanied by an **increase** in entropy are all **spontaneous**.

Free Energy

Can you definitely determine if a reaction is spontaneous? In 1878, **J. Willard Gibbs**, a physicist at Yale University, defined a combined **enthalpy-entropy** function called **Gibbs free energy** that answers that question.

For reactions or processes that take place at **constant pressure** and **temperature**, Gibbs free energy (G_{system}), commonly called **free energy**, **is energy that is available to do work**. Thus, free energy is **useful energy**.

In contrast, some entropy is associated with energy that is **spread out** into the surroundings as, for example, random molecular motion, and cannot be recovered to do useful work.

Gibbs Free Energy Equation

$$\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

The **free energy** released or absorbed in a chemical reaction is equal to the difference between the **enthalpy change** and the product of the **change in entropy** (in joules per kelvin) and the **temperature** (in kelvins).

To calculate Gibbs free energy, it is usually necessary to **convert** units because ΔS is usually expressed in **J/K**, whereas ΔH is expressed in **kJ**.

The Sign of Free Energy

When a reaction or process occurs under **standard conditions** (298 K and 1 atm), the standard free energy change can be expressed as follows.

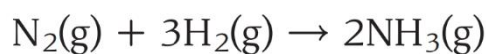
$$\Delta G^{\circ}_{\text{system}} = \Delta H^{\circ}_{\text{system}} - T\Delta S^{\circ}_{\text{system}}$$

- If the sign of the free energy change is **negative**, the reaction is **spontaneous**.
- If the sign of the free energy change is **positive**, the reaction is **nonspontaneous**.

Recall that free energy is energy that is available to do work. In contrast, energy related to entropy is **useless** because it is **dispersed** and cannot be harnessed to do work.

REACTION SPONTANEITY

Calculating Free Energy Change



$$\Delta H_{\text{system}}^{\circ} = -91.8 \text{ kJ} \quad \Delta S_{\text{system}}^{\circ} = -197 \text{ J/K}$$

The **entropy** of the system **decreases** because **4 mol** of gaseous molecules react and only **2 mol** of gaseous molecules are produced. Therefore, $\Delta S_{\text{system}}^{\circ}$ is **negative**.

A decrease in the entropy of the system tends to make the reaction **nonspontaneous**, but the reaction is exothermic ($\Delta H_{\text{system}}^{\circ}$ is **negative**), which tends to make the reaction **spontaneous**.

To determine which of the two tendencies predominates, you must calculate $\Delta G_{\text{system}}^{\circ}$ for the reaction. First, convert $\Delta S_{\text{system}}^{\circ}$ to kilojoules.

$$\Delta S_{\text{system}}^{\circ} = -197 \text{ J/K} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = -0.197 \text{ kJ/K}$$

Now, substitute $\Delta H_{\text{system}}^{\circ}$, T , and $\Delta S_{\text{system}}^{\circ}$ into the equation for $\Delta G_{\text{system}}^{\circ}$.

$$\Delta G_{\text{system}}^{\circ} = \Delta H_{\text{system}}^{\circ} - T\Delta S_{\text{system}}^{\circ}$$

$$\Delta G_{\text{system}}^{\circ} = -91.8 \text{ kJ} - (298 \text{ K})(-0.197 \text{ kJ/K})$$

$$\Delta G_{\text{system}}^{\circ} = -91.8 \text{ kJ} + 58.7 \text{ kJ} = -33.1 \text{ kJ}$$

$\Delta G_{\text{system}}^{\circ}$ for this reaction is **negative**, so the reaction is **spontaneous**.

The reaction between **nitrogen** and **hydrogen** demonstrates that the entropy of a system **can decrease** during a **spontaneous** process. However, it can do so only **if** the entropy of the surroundings **increases** more than the entropy of the system decreases.

Thus, the entropy of the universe (system + surroundings) **always increases** in any spontaneous process.

Table 6 Reaction Spontaneity $\Delta G_{\text{system}} = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$

ΔH_{system}	ΔS_{system}	ΔG_{system}	Reaction Spontaneity
negative	positive	always negative	always spontaneous
negative	negative	negative or positive	spontaneous at lower temperatures
positive	positive	negative or positive	spontaneous at higher temperatures
positive	negative	always positive	never spontaneous

CHAPTER 1

ENERGY AND CHEMICAL CHANGE

CHAPTER SUMMARY

Resources

- *Energy and Chemical Change, Chapter 15, from Glencoe Chemistry: Matter and Change ©2017*

THIS FILE IS FOR EDUCATIONAL PURPOSES ONLY.

This file is under Fair Use:

Copyright Disclaimer under section 107 of the Copyright Act of 1976, allowance is made for "fair use" for purposes such as criticism, comment, news reporting, teaching, scholarship, education and research. Fair use is a use permitted by copyright statute that might otherwise be infringing.

All rights and credit go directly to its rightful owners. No copyright infringement Intended.

Related Files:

Chemistry Student Book EN T1 2022-2023

[<https://t.me/InLibrary/1045>]

Chemistry Teacher Guide EN T1 2022-2023

[<https://t.me/InLibrary/1113>]

Chemistry Student Book AR T1 2022-2023

[<https://t.me/InLibrary/1041>]

Chemistry Teacher Guide AR 2022-2023

[<https://t.me/InLibrary/1100>]

InLibrary – An All-in-One Library for Students

[<https://t.me/InLibrary>]

Copyright © 2022 InLibrary. All rights reserved.