The Energy of Physical, Chemical, and Nuclear Processes

Most physical changes, chemical reactions, and nuclear reactions are accompanied by changes in energy. These energy changes are crucial to life on Earth. For example, chemical reactions in your body generate the heat that helps to regulate your body temperature. Physical changes, such as evaporation, help to keep your body cool. On a much larger scale, there would be no life on Earth without the energy from the nuclear reactions that take place in the Sun.

The study of energy and energy transfer is known as **thermodynamics**. Chemists are interested in the branch of thermodynamics known as **thermochemistry:** the study of energy involved in chemical reactions. In order to discuss energy and its interconversions, thermochemists have agreed on a number of terms and definitions. You will learn about these terms and definitions over the next few pages. Then you will examine the energy changes that accompany chemical reactions, physical changes, and nuclear reactions.

Studying Energy Changes

The **law of conservation of energy** states that the total energy of the universe is constant. In other words, energy can be neither destroyed nor created. This idea can be expressed by the following equation:

$$\Delta E_{universe} = 0$$

Energy can, however, be transferred from one substance to another. It can also be converted into various forms. In order to interpret energy changes, scientists must clearly define what part of the universe they are dealing with. The **system** is defined as the part of the universe that is being studied and observed. In a chemical reaction, the system is usually made up of the reactants and products. By contrast, the **surroundings** are everything else in the universe. The two equations below show the relationship between the universe, a system, and the system's surroundings.

$$\text{Universe} = \text{System} + \text{Surroundings}$$

$$\Delta E_{universe} = \Delta E_{system} + \Delta E_{surroundings} = 0$$

From the relationship, we know that any change in the system is accompanied by an equal and opposite change in the surroundings.

$$\Delta E_{system} = -\Delta E_{surroundings}$$

Look at the chemical reaction that is taking place in the flask in Figure 5.1. A chemist would probably define the system as the contents of the flask—the reactants and products. Technically, the rest of the universe is the surroundings. In reality, however, the entire universe changes very little when the system changes. Therefore, the surroundings are considered to be only the part of the universe that is likely to be affected by the energy changes of the system. In Figure 5.1, the flask, the lab bench, the air in the room, and the student who is carrying out the reaction all make up the surroundings. The system is more likely to significantly influence its immediate surroundings than, say, a mountaintop in Japan (also, technically, part of the surroundings).

Section Preview/ Specific Expectations

In this section, you will

- **write** thermochemical equations, expressing the energy change as a heat term in the equation or as ΔH

- **represent** energy changes using diagrams

- **compare** energy changes that result from physical changes, chemical reactions, and nuclear reactions

- **communicate** your understanding of the following terms: *thermodynamics, thermochemistry, law of conservation of energy, system, surroundings, heat (Q), temperature (T), enthalpy (H), enthalpy change (ΔH), endothermic reaction, exothermic reaction, enthalpy of reaction (ΔH_{rxn}), standard enthalpy of reaction ($\Delta H°_{rxn}$), thermochemical equation, mass defect, nuclear binding energy, nuclear fission, nuclear fusion*

Figure 5.1 The solution in the flask is the system. The flask, the laboratory, and the student are the surroundings.

Heat and Temperature

Heat, **Q**, refers to the transfer of kinetic energy. Heat is expressed in the same units as energy—joules (J). Heat is transferred spontaneously from a warmer object to a cooler object. When you close the door of your home on a cold day to "prevent the cold from getting in," you are actually preventing the heat from escaping. You are preventing the kinetic energy in your warm home from transferring to colder objects, including the cold air, outside.

Temperature, **T**, is a measure of the average kinetic energy of the particles that make up a substance or system. You can think of temperature as a way of quantifying how hot or cold a substance is, relative to another substance.
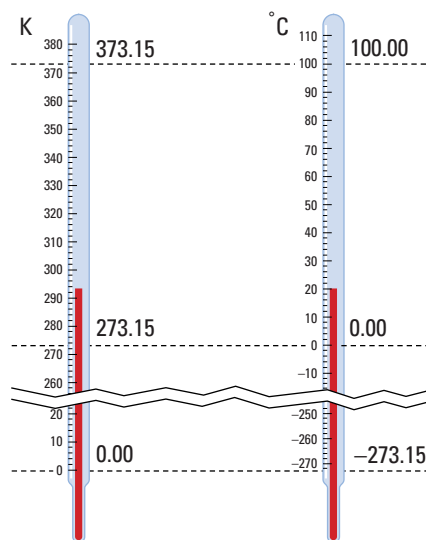
Figure 5.2 Celsius degrees and Kelvin degrees are the same size. The Kelvin scale begins at absolute zero. This is the temperature at which the particles in a substance have no kinetic energy. Therefore, Kelvin temperatures are never negative. By contrast, 0°C is set at the melting point of water. Celsius temperatures can be positive or negative.

Temperature is measured in either Celsius degrees (°C) or kelvins (K). The Celsius scale is a relative scale. It was designed so that water's boiling point is at 100°C and water's melting point is at 0°C. The Kelvin scale, on the other hand, is an absolute scale. It was designed so that 0 K is the temperature at which a substance possesses no kinetic energy. The relationship between the Kelvin and Celsius scales is shown in Figure 5.2, and by the following equation.

Temperature in Kelvin degrees = Temperature in Celsius degrees + 273.15

Enthalpy and Enthalpy Change

Chemists define the total internal energy of a substance at a constant pressure as its **enthalpy**, **H**. Chemists do not work with the *absolute* enthalpy of the reactants and products in a physical or chemical process. Instead, they study the **enthalpy change**, **ΔH**, that accompanies a process. That is, they study the *relative* enthalpy of the reactants and products in a system. This is like saying that the distance between your home and your school is 2 km. You do not usually talk about the *absolute* position of your home and school in terms of their latitude, longitude, and elevation. You talk about their *relative* position, in relation to each other.

The enthalpy change of a process is equivalent to its heat change at constant pressure.

Enthalpy Changes in Chemical Reactions

In chemical reactions, enthalpy changes result from chemical bonds being broken and formed. Chemical bonds are sources of stored energy. *Breaking a bond is a process that requires energy. Creating a bond is a process that releases energy.* For example, consider the combustion reaction that takes place when nitrogen reacts with oxygen.

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$$
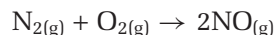
In this reaction, one mole of nitrogen-nitrogen triple bonds and one mole of oxygen-oxygen double bonds are broken. Two moles of nitrogen-oxygen bonds are formed. This reaction absorbs energy. In other words, more energy is released to form two nitrogen-oxygen bonds than is used to break one nitrogen-nitrogen bond and one oxygen-oxygen bond. When a reaction results in a net *absorption* of energy, it is called an **endothermic reaction**.

On the other hand, when a reaction results in a net *release* of energy, it is called an **exothermic reaction**. In an exothermic reaction, more energy is released to form bonds than is used to break bonds. Therefore, energy is released. Figure 5.3 shows the relationship between bond breaking, bond formation, and endothermic and exothermic reactions.
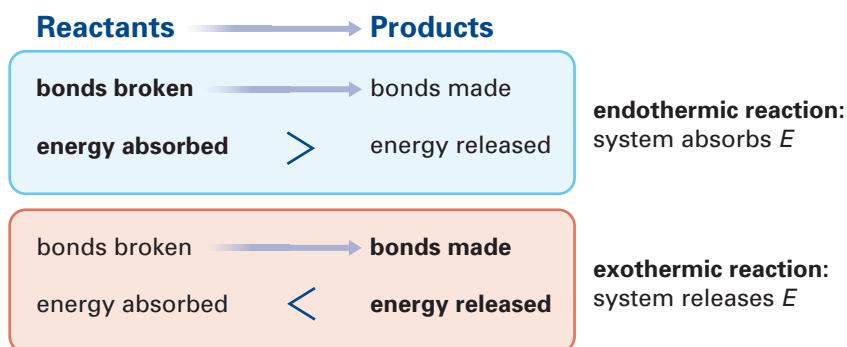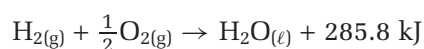
Reactants ⟶ Products

bonds broken	⟶ bonds made	
energy absorbed	$>$	energy released

endothermic reaction: system absorbs E

bonds broken	⟶ **bonds made**	
energy absorbed	$<$	**energy released**

exothermic reaction: system releases E

Figure 5.3 The energy changes that result from bonds breaking and forming determine whether a reaction is exothermic or endothermic.

Representing Enthalpy Changes

The enthalpy change of a chemical reaction is known as the **enthalpy of reaction**, ΔH_{rxn}. The enthalpy of reaction is dependent on conditions such as temperature and pressure. Therefore, chemists often talk about the **standard enthalpy of reaction**, $\Delta H°_{rxn}$: the enthalpy change of a chemical reaction that occurs at SATP (25°C and 100 kPa). Often, $\Delta H°_{rxn}$ is written simply as $\Delta H°$. The ° symbol is called "nought." It refers to a property of a substance at a standard state or under standard conditions. You may see the enthalpy of reaction referred to as the *heat of reaction* in other chemistry books.

CHEM FACT

Chemists use different subscripts to represent enthalpy changes for specific kinds of reactions. For example, ΔH_{comb} represents the enthalpy change of a combustion reaction.

Representing Exothermic Reactions

There are three different ways to represent the enthalpy change of an exothermic reaction. The simplest way is to use a **thermochemical equation**: a balanced chemical equation that indicates the amount of heat that is absorbed or released by the reaction it represents. For example, consider the exothermic reaction of one mole of hydrogen gas with half a mole of oxygen gas to produce liquid water. For each mole of hydrogen gas that reacts, 285.8 kJ of heat is produced. Notice that the heat term is included with the products because heat is produced.

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(\ell)} + 285.8 \text{ kJ}$$

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)}$$
(reactants)

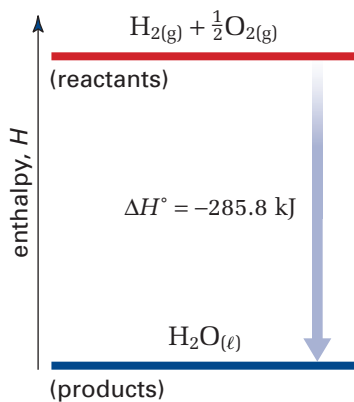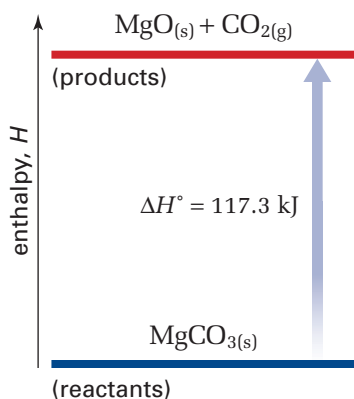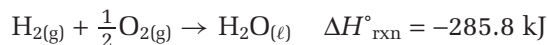enthalpy, H

$\Delta H° = -285.8$ kJ

$H_2O_{(\ell)}$
(products)

Figure 5.4 In an exothermic reaction, the enthalpy of the system decreases as energy is released to the surroundings.

You can also indicate the enthalpy of reaction as a separate expression beside the chemical equation. For exothermic reactions, $\Delta H°$ is always negative.

$$H_{2(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow H_2O_{(\ell)} \quad \Delta H°_{rxn} = -285.8 \text{ kJ}$$

A third way to represent the enthalpy of reaction is to use an enthalpy diagram. Examine Figure 5.4 to see how this is done.

Representing Endothermic Reactions

The endothermic decomposition of solid magnesium carbonate produces solid magnesium oxide and carbon dioxide gas. For each mole of magnesium carbonate that decomposes, 117.3 kJ of energy is absorbed. As for an exothermic reaction, there are three different ways to represent the enthalpy change of an endothermic reaction.

You can include the enthalpy of reaction as a heat term in the chemical equation. Because heat is absorbed in an endothermic reaction, the heat term is included on the reactant side of the equation.
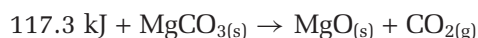
$$117.3 \text{ kJ} + MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$$

You can also indicate the enthalpy of reaction as a separate expression beside the chemical reaction. For endothermic reactions, the enthalpy of reaction is always positive.

$$MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)} \quad \Delta H°_{rxn} = 117.3 \text{ kJ}$$

Finally, you can use a diagram to show the enthalpy of reaction. Figure 5.5 shows how the decomposition of solid magnesium carbonate can be represented graphically.

$$MgO_{(s)} + CO_{2(g)}$$
(products)

enthalpy, H

$\Delta H° = 117.3$ kJ

$MgCO_{3(s)}$
(reactants)

Figure 5.5 In an endothermic reaction, the enthalpy of the system increases as heat energy is absorbed from the surroundings.

Stoichiometry and Thermochemical Equations

The thermochemical equation for the decomposition of magnesium carbonate, shown above, indicates that 117.3 kJ of energy is absorbed when one mole, or 84.32 g, of magnesium carbonate decomposes. The decomposition of two moles of magnesium carbonate absorbs twice as much energy, or 234.6 kJ.

$$MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)} \quad \Delta H°_{rxn} = 117.3 \text{ kJ}$$
$$2MgCO_{3(s)} \rightarrow 2MgO_{(s)} + 2CO_{2(g)} \quad \Delta H°_{rxn} = 234.6 \text{ kJ}$$

Enthalpy of reaction is *linearly dependent* on the quantity of products. That is, if the amount of products formed doubles, the enthalpy change also doubles. Figure 5.6 shows the relationship between the stoichiometry of a reaction and its enthalpy change. Because of this relationship, an exothermic reaction that is relatively safe on a small scale may be extremely dangerous on a large scale. One of the jobs of a chemical engineer is to design systems that allow exothermic reactions to be carried out safely on a large scale. For example, the blast furnaces used in steel making must withstand temperatures of up to 2000°C, produced by the exothermic combustion reaction of coal with oxygen.
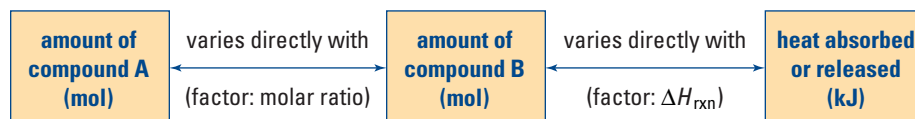
amount of compound A (mol)	varies directly with (factor: molar ratio)	amount of compound B (mol)	varies directly with (factor: ΔH_{rxn})	heat absorbed or released (kJ)

Figure 5.6 This diagram summarizes the relationship between the stoichiometry of a reaction and ΔH.

Stoichiometry and Thermochemical Reactions

Problem

Aluminum reacts readily with chlorine gas to produce aluminum chloride. The reaction is highly exothermic.

$$2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)} \quad \Delta H°_{rxn} = -1408 \text{ kJ}$$

What is the enthalpy change when 1.0 kg of Al reacts completely with excess Cl_2?

What Is Required?

You need to calculate the enthalpy change, ΔH, when the given amount of Al reacts.

What Is Given?

You know the enthalpy change for the reaction of two moles of Al with one mole of Cl_2. From the periodic table, you know the molar mass of Al.

$$2Al_{(s)} + 3Cl_{2(g)} \rightarrow 2AlCl_{3(s)} \quad \Delta H°_{rxn} = -1408 \text{ kJ}$$
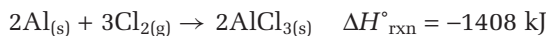
$M_{Al} = 26.98 \text{ g/mol}$

Plan Your Strategy

Convert the given mass of Al to moles. The enthalpy change is linearly dependent on the quantity of reactants. Therefore, you can use a ratio to determine the enthalpy change for 1.0 kg of Al reacting with Cl_2.

Act on Your Strategy

Determine the number of moles of Al in 1 kg. Remember to convert to grams.

$$n \text{ mol Al} = \frac{m_{Al}}{M_{Al}}$$

$$= \frac{1.0 \times 10^3 \text{ g}}{26.98 \text{ g/mol}}$$

$$= 37 \text{ mol}$$

Use ratios to compare the reference reaction with the known enthalpy change (ΔH_1) to the reaction with the unknown enthalpy change (ΔH_2).
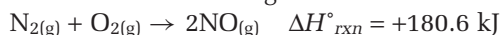
$$\frac{\Delta H_2}{\Delta H_1} = \frac{n_2 \text{ mol Al}}{n_1 \text{ mol Al}}$$

$$\frac{\Delta H_2}{-1408 \text{ kJ}} = \frac{37 \text{ mol Al}}{2 \text{ mol Al}}$$

$$\Delta H_2 = -2.6 \times 10^4 \text{ kJ}$$

Check Your Solution

The sign of the answer is negative, which corresponds to an exothermic reaction. The 1 kg sample contained about 20 times more moles of Al. Therefore, the enthalpy change for the reaction should be about 20 times greater, and it is.

1. Consider the following reaction.

 $N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$ $\Delta H°_{rxn} = +180.6$ kJ

   (a) Rewrite the thermochemical equation, including the standard enthalpy of reaction as either a reactant or a product.

   (b) Draw an enthalpy diagram for the reaction.

   (c) What is the enthalpy change for the formation of one mole of nitrogen monoxide?

   (d) What is the enthalpy change for the reaction of 1.000×10^2 g of nitrogen with sufficient oxygen?

2. The reaction of iron with oxygen is very familiar. You can see the resulting rust on buildings, vehicles, and bridges. You may be surprised, however, at the large amount of heat that is produced by this reaction.

   $4Fe_{(s)} + 3O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 1.65 \times 10^3$ kJ

   (a) What is the enthalpy change for this reaction?

   (b) Draw an enthalpy diagram that corresponds to the thermochemical equation.

   (c) What is the enthalpy change for the formation of 23.6 g of iron(III) oxide?

3. Consider the following thermochemical equation.

   25.9 kJ $+ \frac{1}{2}H_{2(g)} + \frac{1}{2}I_{2(g)} \rightarrow HI_{(g)}$
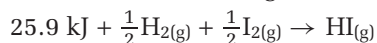
   (a) What is the enthalpy change for this reaction?

   (b) How much energy is needed for the reaction of 4.57×10^{24} molecules of iodine, I_2, with excess hydrogen, H_2?

   (c) Draw and label an enthalpy diagram that corresponds to the given thermochemical equation.

4. Tetraphosphorus decoxide, P_4O_{10}, is an acidic oxide. It reacts with water to produce phosphoric acid, H_3PO_4, in an exothermic reaction.

   $P_4O_{10(s)} + 6H_2O_{(\ell)} \rightarrow 4H_3PO_{4(aq)}$ $\Delta H°_{rxn} = -257.2$ kJ

   (a) Rewrite the thermochemical equation, including the enthalpy change as a heat term in the equation.

   (b) How much energy is released when 5.00 mol of P_4O_{10} reacts with excess water?

   (c) How much energy is released when 235 g of $H_3PO_{4(aq)}$ is formed?

Heat Changes and Physical Changes

Enthalpy changes are associated with physical changes as well as with chemical reactions. You have observed examples of these enthalpy changes in your daily life. Suppose that you want to prepare some pasta. You put an uncovered pot of water on a stove element. The heat from the element causes the water to become steadily hotter, until it reaches 100°C (the boiling point of water at 100 kPa). At this temperature, heat is still being added to the water. The average kinetic energy of the liquid water molecules does not increase, however. Instead, the energy is used to break the intermolecular bonds between the water molecules as they change from liquid to vapour. The temperature of the liquid water remains at

100°C until all the water has been vaporized. If you add heat to the vapour, the temperature of the vapour will increase steadily.

When you heat ice that is colder than 0°C, a similar process occurs. The temperature of the ice increases until it is 0°C (the melting point of water). If you continue to add heat, the ice remains at 0°C but begins to melt, as the bonds between the water molecules in the solid state begin to break.

Figure 5.7 shows the relationship between temperature and heat for a solid substance that melts and then vaporizes as heat is added to it.
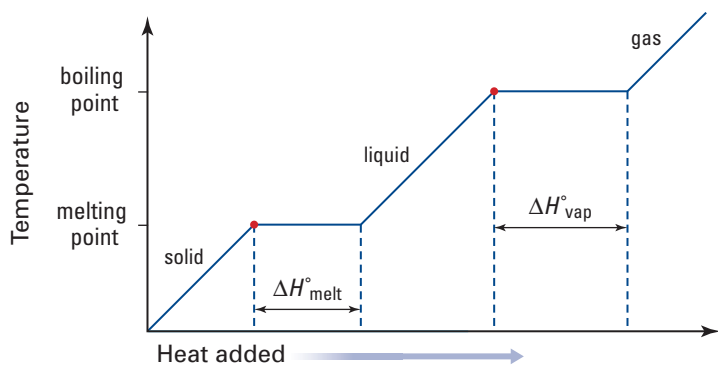
You can represent the enthalpy change that accompanies a phase change—from liquid to solid, for example—just like you represented the enthalpy change of a chemical reaction. You can include a heat term in the equation, or you can use a separate expression of enthalpy change. For example, when one mole of water melts, it absorbs 6.02 kJ of energy.

$$H_2O_{(s)} + 6.02 \text{ kJ} \rightarrow H_2O_{(\ell)}$$

$$H_2O_{(s)} \rightarrow H_2O_{(\ell)} \quad \Delta H = 6.02 \text{ kJ}$$
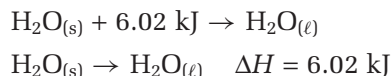
Normally, however, chemists represent enthalpy changes associated with phase changes using modified ΔH symbols. These symbols are described below.

- *enthalpy of vaporization*, ΔH_{vap} : the enthalpy change for the phase change from liquid to gas
- *enthalpy of condensation*, ΔH_{cond}: the enthalpy change for the phase change of a substance from gas to liquid
- *enthalpy of melting*, ΔH_{melt}: the enthalpy change for the phase change of a substance from solid to liquid
- *enthalpy of freezing*, ΔH_{fre}: the enthalpy change for the phase change of a substance from liquid to solid

Vaporization and condensation are opposite processes. Thus, the enthalpy changes for these processes have the same value but opposite signs. For example, 6.02 kJ is needed to vaporize one mole of water. Therefore, 6.02 kJ of energy is released when one mole of water freezes.

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

Similarly, melting and freezing are opposite processes.

$$\Delta H_{melt} = -\Delta H_{fre}$$

Several enthalpies of melting and vaporization are shown in Table 5.1. Notice that the same units (kJ/mol) are used for the enthalpies of melting, vaporization, condensation, and freezing. Also notice that energy changes associated with phase changes can vary widely.

CHEM FACT

The process of melting is also known as *fusion*. Therefore, you will sometimes see the enthalpy of melting referred to as the *enthalpy of fusion*.

Table 5.1 Enthalpies of Melting and Vaporization for Several Substances

Substance	Enthalpy of melting, ΔH_{melt} (kJ/mol)	Enthalpy of vaporization, ΔH_{vap} (kJ/mol)
argon	1.3	6.3
diethyl ether	7.3	29
ethanol	5.0	40.5
mercury	23.4	59
methane	8.9	0.94
sodium chloride	27.2	207
water	6.02	40.7

Hot Packs and Cold Packs: Using the Energy of Physical Changes

You just learned about the enthalpy changes that are associated with phase changes. Another type of physical change that involves a heat transfer is dissolution. When a solute dissolves in a solvent, the enthalpy change that occurs is called the *enthalpy of solution*, ΔH_{soln}. Dissolution can be either endothermic or exothermic.

Manufacturers take advantage of endothermic dissolution to produce cold packs that athletes can use to treat injuries. One type of cold pack contains water and a salt, such as ammonium nitrate, in separate compartments. When you crush the pack, the membrane that divides the compartments breaks, and the salt dissolves. This dissolution process is endothermic. It absorbs heat for a short period of time, so the cold pack feels cold. Figure 5.8 shows how a cold pack works.
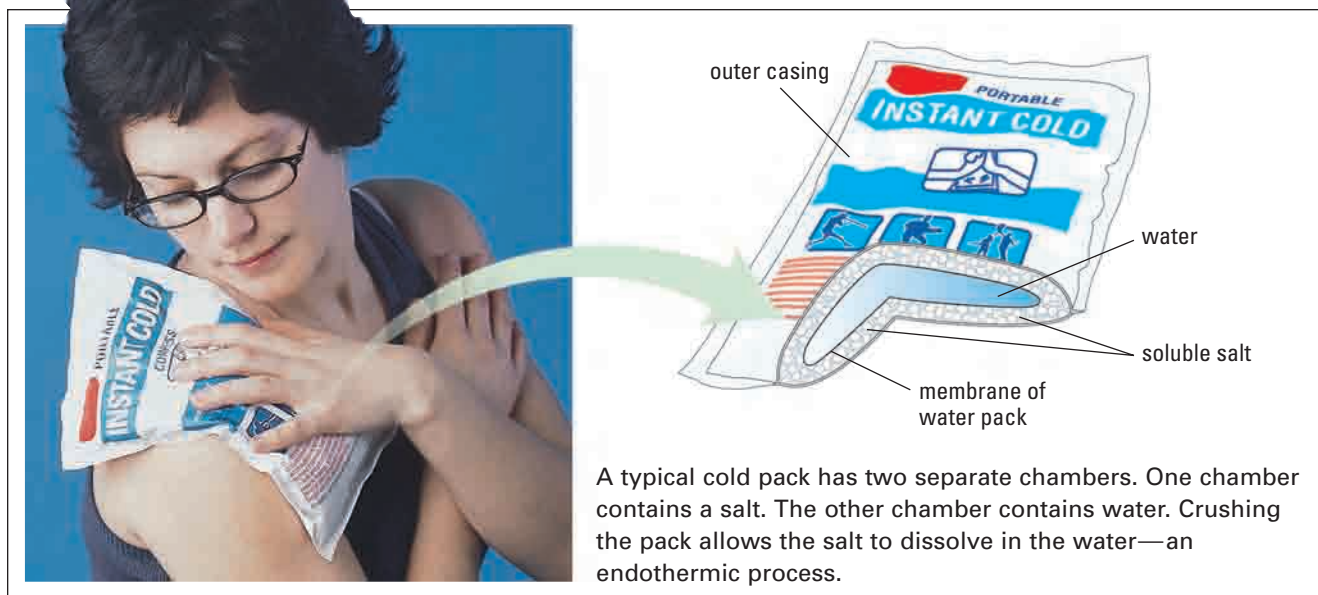

outer casing

water

soluble salt

membrane of water pack

A typical cold pack has two separate chambers. One chamber contains a salt. The other chamber contains water. Crushing the pack allows the salt to dissolve in the water—an endothermic process.

Figure 5.8 This person's shoulder was injured. Using a cold pack helps to reduce the inflammation of the joint.

Some types of hot packs are constructed in much the same way as the cold packs described above. They have two compartments. One compartment contains a salt, such as calcium chloride. The other compartment contains water. In hot packs, however, the dissolution process is exothermic. It releases heat to the surroundings.

Energy and Nuclear Reactions

The energy that is released by a physical change, such as the dissolution of calcium chloride, can warm your hands. The energy that is released by a chemical reaction, such as the formation of water, can power a rocket. The energy that is released by a nuclear reaction, such as the nuclear reactions in the Sun, however, can provide enough heat to fry an egg on a sidewalk that is 150 000 000 km away from the surface of the Sun.

From previous science courses, you will recall that nuclear reactions involve changes in the nuclei of atoms. Often nuclear reactions result in the transformation of one or more elements into one or more different elements.

Like physical changes and chemical reactions, nuclear reactions are accompanied by energy changes. Nuclear reactions, however, produce significantly more energy than physical and chemical processes. *In nuclear reactions, a significant amount of the mass of the reactants is actually converted into energy*.

Ever since Albert Einstein devised his famous equation, $E = mc^2$, we have known that mass and energy are interconvertible. In Einstein's equation, E is energy in $kg \cdot m^2/s^2$ (J), m is the mass in kg, and c^2 is the square of the speed of light.

$$c^2 = (3.0 \times 10^8 \text{ m/s})^2$$
$$= 9.0 \times 10^{16} \text{ m}^2/\text{s}^2$$

As you can see, c^2 is an enormous number. Therefore, even a very tiny amount of matter is equivalent to a significant amount of energy.

For example, compare the mass of 1 mol of carbon-12 atoms with the mass of the individual nucleons in 1 mol of carbon-12 atoms. The mass of 6 mol of hydrogen-1 atoms (one proton and one electron each) and 6 mol of neutrons is 12.098 940 g. The mass of 1 mol of carbon-12 atoms is exactly 12 g. Note that the mass of the electrons does not change in a nuclear reaction.

$$
\begin{array}{r}
12.098\ 940 \text{ g/mol} \\
-\ 12.000\ 000 \text{ g/mol} \\
\hline
0.098\ 940 \text{ g/mol}
\end{array}
$$

The difference in mass is significant. It would show up on any reasonably precise balance. Thus, the mass of the nucleus of carbon-12 is significantly less than the mass of its component nucleons. The difference in mass between a nucleus and its nucleons is known as the **mass defect**. What causes this mass defect? It is caused by the **nuclear binding energy**: the energy associated with the strong force that holds a nucleus together.

$$\text{Nucleus} + \text{Nuclear binding energy} \rightarrow \text{Nucleons}$$

You can use Einstein's equation to calculate the nuclear binding energy for carbon-12.

$$\Delta E = \Delta mc^2$$
$$= (9.89 \times 10^{-5} \text{ kg/mol})(9.0 \times 10^{16} \text{ m}^2/\text{s}^2)$$
$$= 8.9 \times 10^{12} \text{ J/mol}$$
$$= 8.9 \times 10^9 \text{ kJ/mol}$$

Clearly, the energy associated with the bonds that hold a nucleus together is much greater than the energy associated with chemical bonds, which are usually only a few hundred kJ/mol.

The higher the binding energy of a nucleus, the more stable the nucleus is. Nuclei with mass numbers (A) that are close to 60 are the most stable. Nuclear reactions, in which nuclei break apart or fuse, tend to form nuclei that are more stable than the reactant nuclei. Figure 5.9 illustrates the relative stability of various nuclei.

Figure 5.9 This graph shows the relative stability of nuclei. It indicates whether nuclei are more likely to split or fuse in a nuclear reaction. Notice that the helium-4 nucleus is unusually stable.
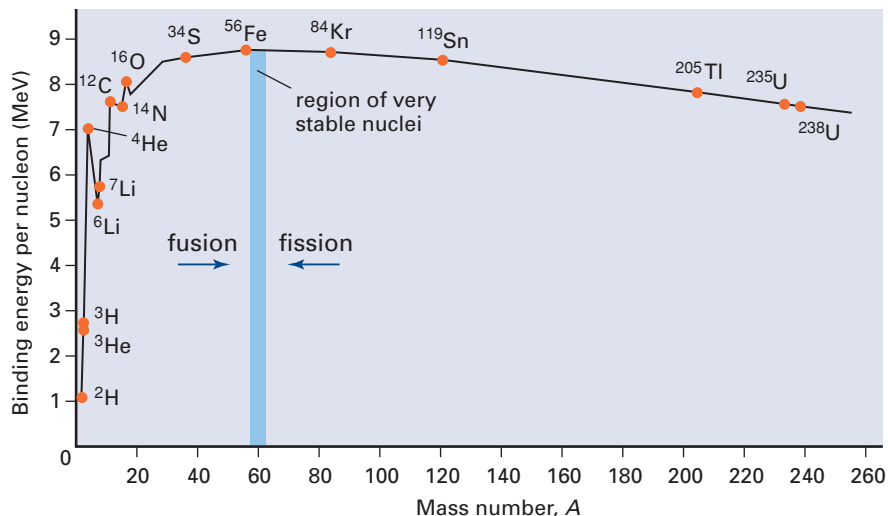
The difference between the nuclear binding energy of the reactant nuclei and the product nuclei represents the energy change of the nuclear reaction.

Nuclear Fission

A heavy nucleus can split into lighter nuclei by undergoing **nuclear fission**. Nuclear power plants use controlled nuclear fission to provide energy. Uncontrolled nuclear fission is responsible for the massive destructiveness of an atomic bomb.

The most familiar fission reactions involve the splitting of uranium atoms. In these reactions, a uranium-235 atom is bombarded with neutrons. The uranium nucleus then splits apart into various product nuclei. Two examples of fission reactions that involve uranium-235 are shown in Figure 5.10.

Electronic Learning Partner

To learn more about nuclear fission, go to the Chemistry 12 Electronic Learning Partner.

$$^{1}_{0}n + {}^{235}_{92}U \longrightarrow {}^{92}_{36}Kr + {}^{141}_{56}Ba + 3\,{}^{1}_{0}n$$

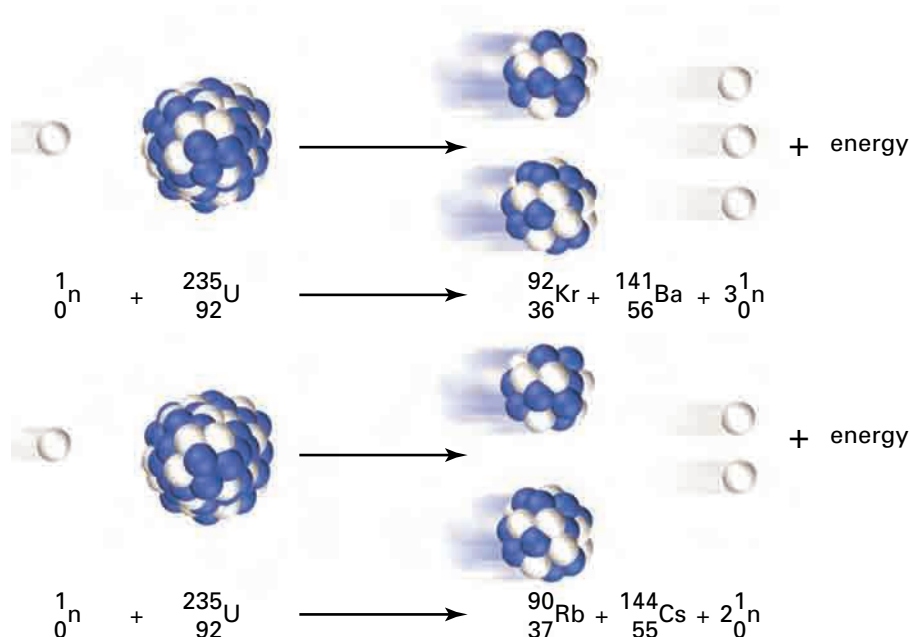$$^{1}_{0}n + {}^{235}_{92}U \longrightarrow {}^{90}_{37}Rb + {}^{144}_{55}Cs + 2\,{}^{1}_{0}n$$

Figure 5.10 Uranium can undergo fission in numerous different ways, producing various product nuclei. Two examples are shown here.

Fission reactions produce vast quantities of energy. For example, when one mole of uranium-235 splits, it releases 2.1×10^{13} J. By contrast, when one mole of coal burns, it releases about 3.9×10^6 J. Thus, the combustion of coal releases about five million times fewer joules of energy per mole than the fission of uranium-235.

Nuclear Fusion

Two smaller nuclei can fuse to form a larger nucleus, in what is called a **nuclear fusion** reaction. You and all other life on Earth would not exist without nuclear fusion reactions. These reactions are the source of the energy produced in the Sun.

One example of a fusion reaction is the fusion of deuterium and tritium.

$$\mathrm{^{2}_{1}H + {}^{3}_{1}H \rightarrow {}^{4}_{2}He + {}^{1}_{0}n}$$
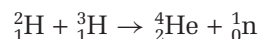
The seemingly simple reaction between deuterium and tritium produces 1.7×10^{12} J of energy for each mole of deuterium. This is about 10 times fewer joules of energy than are produced by the fission of one mole of uranium. It is still, however, 500 000 times more energy than is produced by burning one mole of coal.
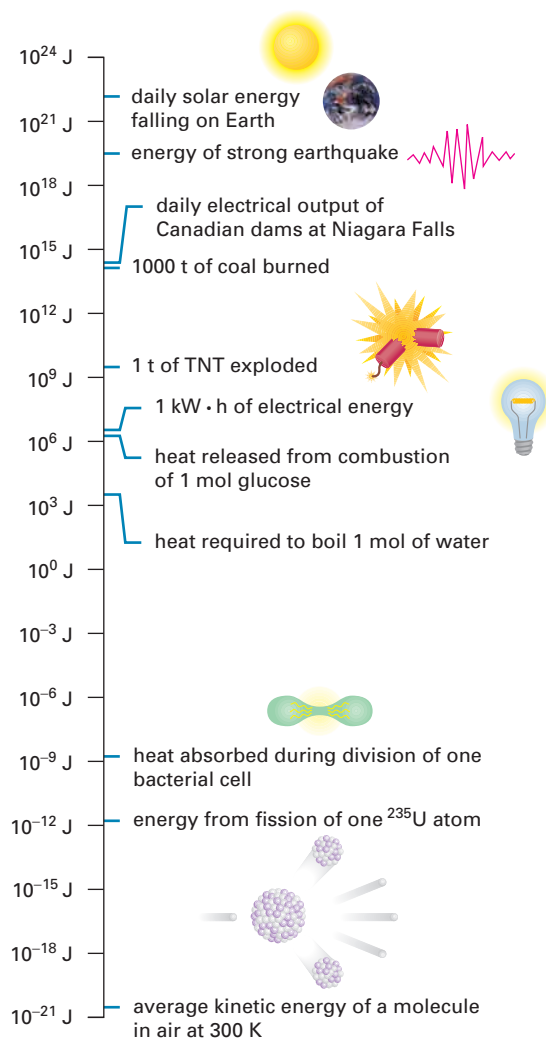
Scientists are searching for a way to harness the energy from fusion reactions. Fusion is a more desirable way to produce energy than fission. The main product of fusion, helium, is relatively harmless compared with the radioactive products of fission. Unfortunately, fusion is proving more difficult than fission to harness. Fusion will not proceed at a reasonable rate without an enormous initial input of energy. This is not a problem in the core of the Sun, where the temperature ranges from 7 500 000°C to 15 000 000°C. It is a problem in industry. Scientists are working on safe and economical ways to provide the high-temperature conditions that are needed to make fusion a workable energy source.

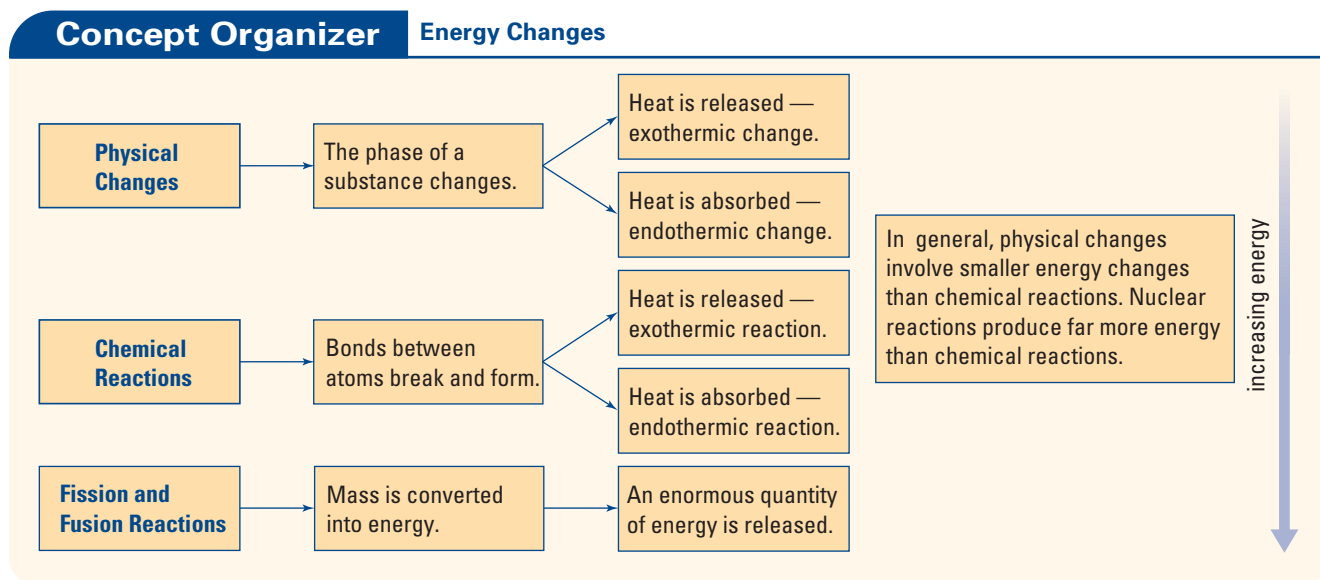## Comparing the Energy of Physical, Chemical, and Nuclear Processes

In this section, you learned that physical changes, chemical reactions, and nuclear reactions all involve energy changes. You also learned that the energy changes have some striking differences in magnitude. Figure 5.11 shows energy changes for some physical, chemical, and nuclear processes. Some other interesting energy statistics are included for reference.

Math ⟶ LINK

The fission of one mole of uranium-235 produces more energy than the fusion of one mole of deuterium with one mole of tritium. What if you compare the energy that is produced in terms of *mass* of reactants? Calculate a ratio to compare the energy that is produced from fusion and fission, per gram of fuel. What practical consequences arise from your result?

10^{24} J — daily solar energy falling on Earth
10^{21} J — energy of strong earthquake
10^{18} J — daily electrical output of Canadian dams at Niagara Falls
10^{15} J — 1000 t of coal burned
10^{12} J
10^{9} J — 1 t of TNT exploded
— 1 kW·h of electrical energy
10^{6} J — heat released from combustion of 1 mol glucose
10^{3} J — heat required to boil 1 mol of water
10^{0} J
10^{-3} J
10^{-6} J
10^{-9} J — heat absorbed during division of one bacterial cell
10^{-12} J — energy from fission of one ^{235}U atom
10^{-15} J
10^{-18} J
10^{-21} J — average kinetic energy of a molecule in air at 300 K

Figure 5.11 The energy changes of physical, nuclear, and chemical processes vary widely. In general, however, chemical reactions are associated with greater energy changes than physical changes. Nuclear reactions are associated with far greater energy changes than chemical reactions.

The following Concept Organizer summarizes what you learned about the energy changes associated with physical changes, chemical reactions, and nuclear reactions.

Concept Organizer   Energy Changes

Physical Changes	The phase of a substance changes.	Heat is released — exothermic change.
		Heat is absorbed — endothermic change.
Chemical Reactions	Bonds between atoms break and form.	Heat is released — exothermic reaction.
		Heat is absorbed — endothermic reaction.
Fission and Fusion Reactions	Mass is converted into energy.	An enormous quantity of energy is released.

In general, physical changes involve smaller energy changes than chemical reactions. Nuclear reactions produce far more energy than chemical reactions.

increasing energy

Section Summary

In section 5.1, you learned about the energy changes that accompany physical changes, chemical reactions, and nuclear reactions. You learned how to represent energy changes using thermochemical equations and diagrams. In the next section, you will determine the enthalpy of a reaction by experiment.

Section Review

1 **K/U** In your own words, explain why exothermic reactions have $\Delta H < 0$.

2 **K/U** Label each thermochemical equation with the most specific form(s) of ΔH. Remember to pay attention to the *sign* of ΔH.

(a) $Ag_{(s)} + \frac{1}{2}Cl_{2(g)} \rightarrow AgCl_{(s)} + 127.0$ kJ (at 25°C and 100 kPa)

(b) 44.0 kJ $+ H_2O_{(\ell)} \rightarrow H_2O_{(g)}$ (at 25°C and 100 kPa)

(c) $C_2H_{4(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 2H_2O_{(g)} +$ energy

3 **C** Suppose that one of your friends was absent on the day that you learned about labelling ΔH. To help your friend, create a table that summarizes the different ways to label ΔH, and their meanings.

4 **I** Calcium oxide, CaO, reacts with carbon in the form of graphite. Calcium carbide, CaC_2, and carbon monoxide, CO, are produced in an endothermic reaction.

$CaO_{(s)} + 3C_{(s)} + 462.3$ kJ $\rightarrow CaC_{2(s)} + CO_{(g)}$

(a) 246.7 kJ of energy is available to react. What mass of calcium carbide is produced, assuming sufficient reactants?

(b) What is the enthalpy change for the reaction of 46.7 g of graphite with excess calcium oxide?

(c) 1.38×10^{24} formula units of calcium oxide react with excess graphite. How much energy is needed?

5 **I** Acetylene, C_2H_2, undergoes complete combustion in oxygen. Carbon dioxide and water are formed. When one mole of acetylene undergoes complete combustion, 1.3×10^3 kJ of energy is released.

(a) Write a thermochemical equation for this reaction.

(b) Draw a diagram to represent the thermochemical equation.

(c) How much energy is released when the complete combustion of acetylene produces 1.50 g of water?

6 **C** Write an equation to represent each phase change in Table 5.1 on page 228. Include the enthalpy change as a heat term in the equation.

7 **MC** When one mole of gaseous water forms from its elements, 241.8 kJ of energy is released. In other words, when hydrogen burns in oxygen or air, it produces a great deal of energy. Since the nineteenth century, scientists have been researching the potential of hydrogen as a fuel. One way in which the energy of the combustion of hydrogen has been successfully harnessed is as rocket fuel for aircraft.

(a) Write a thermodynamic equation for the combustion of hydrogen.

(b) Describe three reasons why hydrogen gas is a desirable rocket fuel.

(c) Suggest challenges that engineers might have had to overcome in order to make hydrogen a workable rocket fuel for aircraft.

(d) Use print and electronic resources to find out about research into hydrogen as a fuel. Create a time line that shows significant events and discoveries in this research.

8 **MC** A healthy human body maintains a temperature of about 37.0°C. Explain how physical, chemical, and nuclear processes all contribute, directly or indirectly, to keeping the human body at a constant temperature.