



CHAPTER 6

Energy and Metabolism

Chapter Contents

- 6.1 The Flow of Energy in Living Systems
- 6.2 The Laws of Thermodynamics and Free Energy
- 6.3 ATP: The Energy Currency of Cells
- 6.4 Enzymes: Biological Catalysts
- 6.5 Metabolism: The Chemical Description of Cell Function

Introduction

Life can be viewed as a constant flow of energy, channeled by organisms to do the work of living. Each of the significant properties by which we define life—order, growth, reproduction, responsiveness, and internal regulation—requires a constant supply of energy. Both the lion and the giraffe need to eat to provide energy for a wide variety of cellular functions. Deprived of a source of energy, life stops. Therefore, a comprehensive study of life would be impossible without discussing bioenergetics, the analysis of how energy powers the activities of living systems. In this chapter, we focus on energy—what it is and how it changes during chemical reactions.

6.1 The Flow of Energy in Living Systems

Learning Outcomes

1. Differentiate between kinetic and potential energy.
2. Identify the source of energy for the biosphere.
3. Describe the nature of redox reactions.

Thermodynamics is the branch of chemistry concerned with energy changes. Cells are governed by the laws of physics and chemistry, so we must understand these laws in order to understand how cells function.

Energy can take many forms

Energy is defined as the capacity to do work. We think of energy as existing in two states: kinetic energy and potential energy (figure 6.1).

Kinetic energy is the energy of motion. Moving objects perform work by causing other matter to move. **Potential energy** is stored energy. Objects that are not actively moving but have the capacity to do so possess potential energy. A boulder perched on a hilltop has gravitational potential energy. As it begins to roll downhill, some of its potential energy is converted into kinetic energy. Much of the work that living organisms carry out involves transforming potential energy into kinetic energy.

Energy can take many forms: mechanical energy, heat, sound, electric current,

light, or radioactivity. Because it can exist in so many forms, energy can be measured in many ways. Heat is the most convenient way of measuring energy because all other forms of energy can be converted into heat. In fact, the term *thermodynamics* means “heat changes.”

The unit of heat most commonly employed in biology is the kilocalorie (kcal). One kilocalorie is equal to 1000 calories (cal). One calorie is the heat required to raise the temperature of one gram of water one degree Celsius ($^{\circ}\text{C}$). (You are probably more used to seeing the term *Calorie* with a capital C. This is used on food labels and is actually the same as kilocalorie.) Another energy unit, often used in physics, is the *joule*; one joule equals 0.239 cal.

The Sun provides energy for living systems

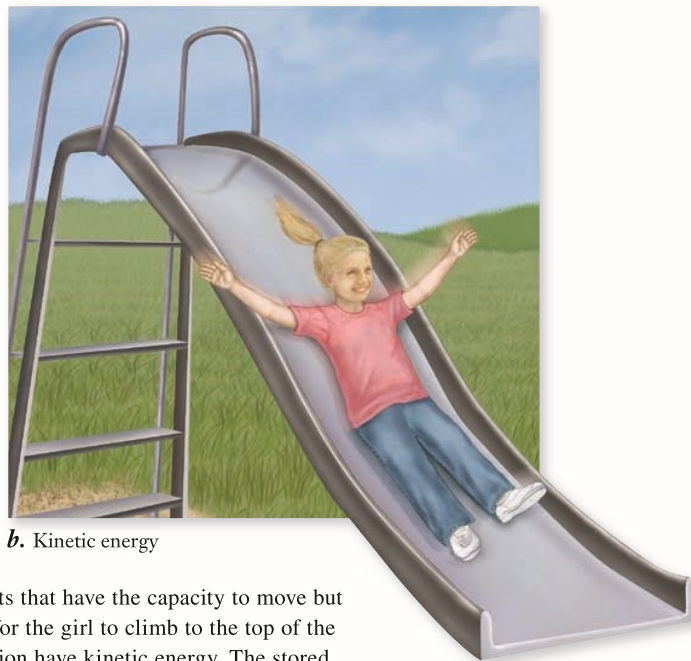
Energy flows into the biological world from the Sun. It is estimated that sunlight provides the Earth with more than 13×10^{23} calories per year, or 40 million billion calories per second! Plants, algae, and certain kinds of bacteria capture a fraction of this energy through photosynthesis.

In photosynthesis, energy absorbed from sunlight is used to combine small molecules (water and carbon dioxide) into more complex ones (sugars). This process converts carbon from an inorganic to an organic form. In the process, energy from sunlight is stored as potential energy in the covalent bonds between atoms in the sugar molecules.

Breaking the bonds between atoms requires energy. In fact, the strength of a covalent bond is measured by the amount of energy required to break it. For example, it takes 98.8 kcal to break one mole (6.023×10^{23}) of the carbon–hydrogen (C–H) bonds found in organic molecules. Fat molecules have many C–H bonds,



a. Potential energy



b. Kinetic energy

Figure 6.1 Potential and kinetic energy. *a.* Objects that have the capacity to move but are not moving have potential energy. The energy required for the girl to climb to the top of the slide is stored as potential energy. *b.* Objects that are in motion have kinetic energy. The stored potential energy is released as kinetic energy as the girl slides down.

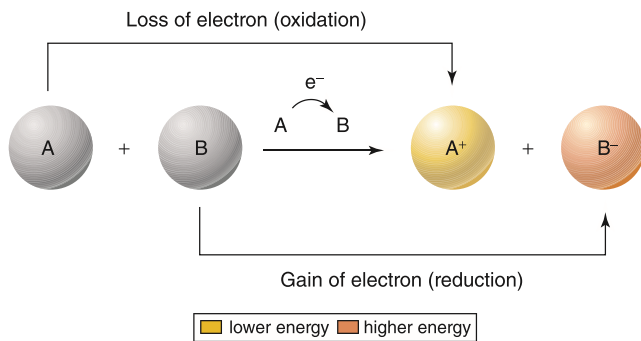


Figure 6.2 Redox reactions. Oxidation is the loss of an electron; reduction is the gain of an electron. In this example, the charges of molecules A and B appear as superscripts in each molecule. Molecule A loses energy as it loses an electron, and molecule B gains that energy as it gains an electron.

and breaking those bonds provides lots of energy. This is one reason animals store fat. The oxidation of one mole of a 16-carbon fatty acid that is completely saturated with hydrogens yields 2340 kcal.

Redox reactions transfer electrons

During a chemical reaction, the energy stored in chemical bonds may be used to make new bonds. In some of these reactions, electrons actually pass from one atom or molecule to another. An atom or molecule that loses an electron is said to be oxidized, and the process by which this occurs is called **oxidation**. The name comes from the fact that oxygen is the most common electron acceptor in biological systems. Conversely, an atom or molecule that gains an electron is said to be reduced, and the process is called **reduction**. The reduced form of a molecule has a higher level of energy than the oxidized form (figure 6.2).

Oxidation and reduction always take place together, because every electron that is lost by one atom through oxidation is gained by another atom through reduction. Therefore, chemical reactions of this sort are called **oxidation–reduction**, or **redox**, reactions. Oxidation–reduction reactions play a key role in the flow of energy through biological systems.

In chapters 7 and 8, you will learn the details of how organisms derive energy from the oxidation of organic compounds via respiration, as well as from the energy in sunlight via photosynthesis.

Learning Outcomes Review 6.1

Energy is defined as the capacity to do work. The two forms of energy are kinetic energy, or energy of motion, and potential energy, or stored energy. The ultimate source of energy for living systems is the Sun. Organisms derive their energy from redox reactions. In oxidation, a molecule loses an electron; in reduction, a molecule gains an electron.

- *What energy source might ecosystems at the bottom of the ocean use?*

6.2 The Laws of Thermodynamics and Free Energy

Learning Outcomes

1. Explain the laws of thermodynamics.
2. Relate free energy changes to the outcome of chemical reactions.
3. Contrast the course of a reaction with and without an enzyme catalyst.

All activities of living organisms—growing, running, thinking, singing, reading these words—involve changes in energy. A set of two universal laws we call the laws of thermodynamics govern all energy changes in the universe, from nuclear reactions to a bird flying through the air.

The First Law states that energy cannot be created or destroyed

The **First Law of Thermodynamics** concerns the amount of energy in the universe. Energy cannot be created or destroyed; it can only change from one form to another (from potential to kinetic, for example). The total amount of energy in the universe remains constant.

The lion eating a giraffe at the beginning of this chapter is acquiring energy. Rather than creating new energy or capturing the energy in sunlight, the lion is merely transferring some of the potential energy stored in the giraffe's tissues to its own body, just as the giraffe obtained the potential energy stored in the plants it ate while it was alive.

Within any living organism, chemical potential energy stored in some molecules can be shifted to other molecules and stored in different chemical bonds. It can also be converted into other forms, such as kinetic energy, light, or electricity. During each conversion, some of the energy dissipates into the environment as **heat**, which is a measure of the random motion of molecules (and therefore a measure of one form of kinetic energy). Energy continuously flows through the biological world in one direction, with new energy from the Sun constantly entering the system to replace the energy dissipated as heat.

Heat can be harnessed to do work only when there is a heat gradient—that is, a temperature difference between two areas. Cells are too small to maintain significant internal temperature differences, so heat energy is incapable of doing the work of cells. Instead, cells must rely on chemical reactions for energy.

Although the total amount of energy in the universe remains constant, the energy available to do work decreases as more of it is progressively lost as heat.

The Second Law states that some energy is lost as disorder increases

The **Second Law of Thermodynamics** concerns the transformation of potential energy into heat, or random molecular motion during any energy transaction. It states that the disorder in the universe, more formally called **entropy**, is continuously increasing. Put simply, disorder is more likely than order. For example, it is much more likely that a column of bricks will tumble over than that a pile of bricks will arrange themselves spontaneously to form a column.

In general, energy transformations proceed spontaneously to convert matter from a more ordered, less stable form to a less ordered, but more stable form. For this reason, the second law is sometimes called “time’s arrow.” Looking at the photographs in figure 6.3, you could put the pictures into correct sequence using the information that time had elapsed with only natural processes occurring. Although it might be great if our rooms would straighten themselves up, we know from experience how much work it takes to do so.

The Second Law of Thermodynamics can also be stated simply as “entropy increases.” When the universe formed, it held all the potential energy it will ever have. It has become progressively more disordered ever since, with every energy exchange increasing the amount of entropy.

Chemical reactions can be predicted based on changes in free energy

It takes energy to break the chemical bonds that hold the atoms in a molecule together. Heat energy, because it increases atomic motion, makes it easier for the atoms to pull apart. Both chemical bonding and heat have a significant influence on a molecule. Chemical bonding reduces disorder; heat increases it. The net effect, the amount of energy actually available to break and subsequently form other chemical bonds, is called the *free energy* of that

molecule. In a more general sense, **free energy** is defined as the energy available to do work in any system.

For a molecule within a cell, where pressure and volume usually do not change, the free energy is denoted by the symbol G (for “Gibbs free energy”). G is equal to the energy contained in a molecule’s chemical bonds (called **enthalpy** and designated H) together with the energy term (TS) related to the degree of disorder in the system, where S is the symbol for *entropy* and T is the absolute temperature expressed in the Kelvin scale ($K = ^\circ\text{C} + 273$):

$$G = H - TS$$

Chemical reactions break some bonds in the reactants and form new ones in the products. Consequently, reactions can produce changes in free energy. When a chemical reaction occurs under conditions of constant temperature, pressure, and volume—as do most biological reactions—the change symbolized by the Greek capital letter delta, Δ , in free energy (ΔG) is simply:



$$\Delta G = \Delta H - T\Delta S$$

The change in free energy is useful because it allows us to predict whether a particular chemical reaction is spontaneous or not. The change in free energy is calculated as energy of products minus energy of reactants, so if the products have *more* free energy than the reactants, ΔG is positive. Such reactions are not spontaneous because they require an input of energy. This can be because either the bond energy (H) is higher, or the disorder (S) in the system is lower. We call this an **endergonic** reaction. A plot of free energy over the course of reaction shows this graphically (figure 6.4). In the case of an endergonic reaction, it is “uphill” from reactants to products, and not spontaneous.

Conversely, if the products have less free energy than the reactants, ΔG is negative, and the reaction will proceed spontaneously. The negative ΔG can be because the bond energy (H) is lower, or the disorder (S) is higher, or both. These reactions release the excess free energy as heat and are called **exergonic** reactions.

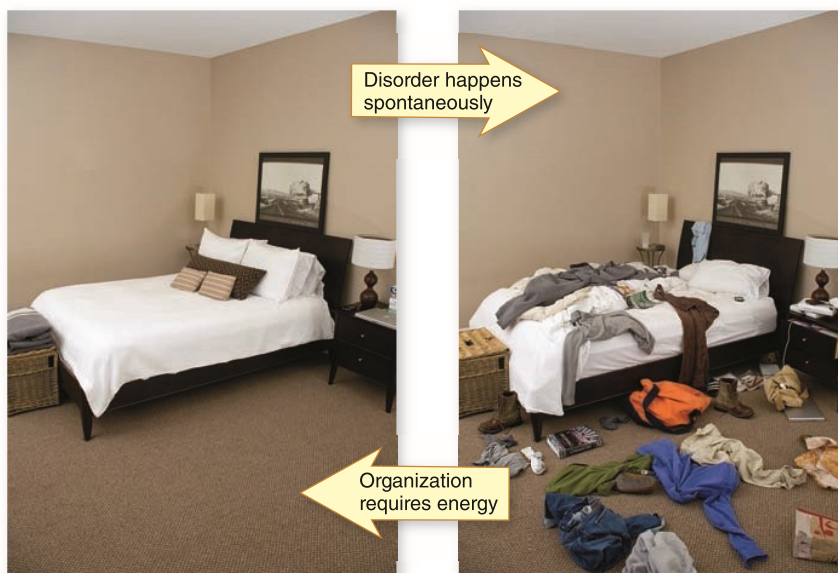


Figure 6.3 Entropy in action. As time elapses, the room shown at right becomes more disorganized. Entropy has increased in this room. It takes energy to restore it to the ordered state shown at left.

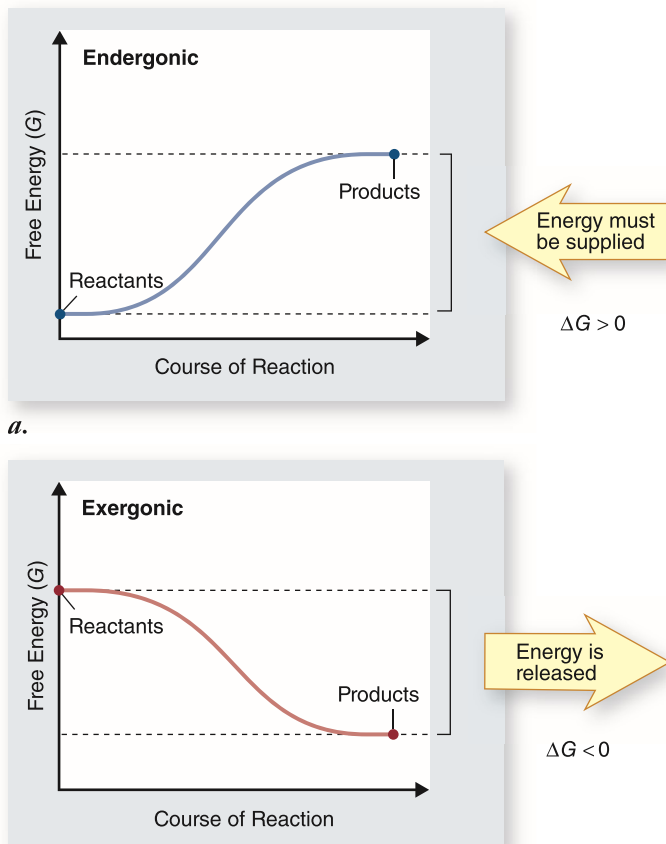


Figure 6.4 Energy in chemical reactions. *a.* In an endergonic reaction, the products of the reaction contain more energy than the reactants, and the extra energy must be supplied for the reaction to proceed. *b.* In an exergonic reaction, the products contain less energy than the reactants, and the excess energy is released.

Our plot of free energy for the reaction is now “downhill” and the reaction will be spontaneous (figure 6.4). Note that *spontaneous* does not mean the same thing as *instantaneous*. A spontaneous reaction may proceed very slowly.

Because chemical reactions are reversible, a reaction that is exergonic in the forward direction will be endergonic in the reverse direction. For each reaction, an equilibrium exists at some point between the relative amounts of reactants and products. This equilibrium has a numeric value and is called the *equilibrium constant*. This characteristic of reactions provides us with another way to think about free energy changes: An exergonic reaction has an equilibrium favoring the products, and an endergonic reaction has an equilibrium favoring the reactants.

Spontaneous chemical reactions require activation energy

If all chemical reactions that release free energy tend to occur spontaneously, why haven’t all such reactions already occurred? Consider the gasoline tank of your car: The oxidation of the hydrocarbons in gasoline is an exergonic reaction, but your gas tank does not spontaneously explode. One reason is that most reactions

require an input of energy to get started. In the case of your car, this input consists of the electrical sparks in the engine’s cylinders, producing a controlled explosion.

Activation energy

Before new chemical bonds can form, even bonds that contain less energy, existing bonds must first be broken, and that requires energy input. The extra energy needed to destabilize existing chemical bonds and initiate a chemical reaction is called **activation energy** (figure 6.5).

The rate of an exergonic reaction depends on the activation energy required for the reaction to begin. Reactions with larger activation energies tend to proceed more slowly because fewer molecules succeed in getting over the initial energy hurdle. The rate of reactions can be increased in two ways: (1) by increasing the energy of reacting molecules or (2) by lowering activation energy. Chemists often drive important industrial reactions by increasing the energy of the reacting molecules, which is frequently accomplished simply by heating up the reactants. The other strategy is to use a catalyst to lower the activation energy.

How catalysts work

Stressing particular chemical bonds can make them easier to break. The process of influencing chemical bonds in a way that lowers the activation energy needed to initiate a reaction is called **catalysis**, and substances that accomplish this are known as *catalysts* (figure 6.5).

Catalysts exert their action by affecting an intermediate stage in a reaction—the transition state. The energy needed to reach this transition state is the activation energy. Catalysts stabilize this transition state, thus lowering activation energy.

Catalysts cannot violate the basic laws of thermodynamics; they cannot make an endergonic reaction proceed spontaneously. By reducing the activation energy, a catalyst accelerates both the forward and the reverse reactions by exactly the same amount.

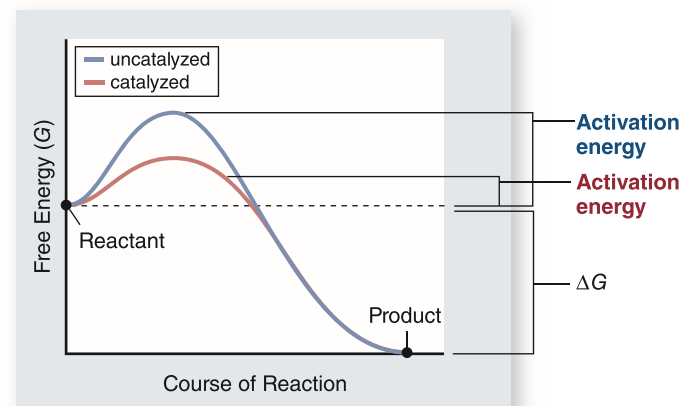


Figure 6.5 Activation energy and catalysis. Exergonic reactions do not necessarily proceed rapidly because activation energy must be supplied to destabilize existing chemical bonds. Catalysts accelerate particular reactions by lowering the amount of activation energy required to initiate the reaction. Catalysts do not alter the free-energy change produced by the reaction.

Therefore, a catalyst does not alter the proportion of reactant ultimately converted into product.

To understand this, imagine a bowling ball resting in a shallow depression on the side of a hill. Only a narrow rim of dirt below the ball prevents it from rolling down the hill. Now imagine digging away that rim of dirt. If you remove enough dirt from below the ball, it will start to roll down the hill—but removing dirt from below the ball will *never* cause the ball to roll up the hill. Removing the lip of dirt simply allows the ball to move freely; gravity determines the direction it then travels.

Similarly, the direction in which a chemical reaction proceeds is determined solely by the difference in free energy between reactants and products. Like digging away the soil below the bowling ball on the hill, catalysts reduce the energy barrier that is preventing the reaction from proceeding. Only exergonic reactions can proceed spontaneously, and catalysts cannot change that. What catalysts *can* do is make a reaction proceed much faster. In living systems, enzymes act as catalysts.

Learning Outcomes Review 6.2

The First Law of Thermodynamics states that energy cannot be created or destroyed. The Second Law states that disorder, or entropy, is increasing. Free-energy changes (ΔG) can predict whether chemical reactions take place. Reactions with a negative ΔG occur spontaneously, and those with a positive ΔG do not. Energy needed to initiate a reaction is termed activation energy. Catalysts stabilize an intermediate transition state, lowering activation energy and accelerating reactions.

- Can an enzyme make an endergonic reaction exergonic?

6.3 ATP: The Energy Currency of Cells

Learning Outcomes

1. Describe the role of ATP in short-term energy storage.
2. Distinguish which bonds in ATP are “high energy.”

The chief “currency” all cells use for their energy transactions is the nucleotide *adenosine triphosphate* (ATP). ATP powers almost every energy-requiring process in cells, from making sugars, to supplying activation energy for chemical reactions, to actively transporting substances across membranes, to moving through the environment and growing.

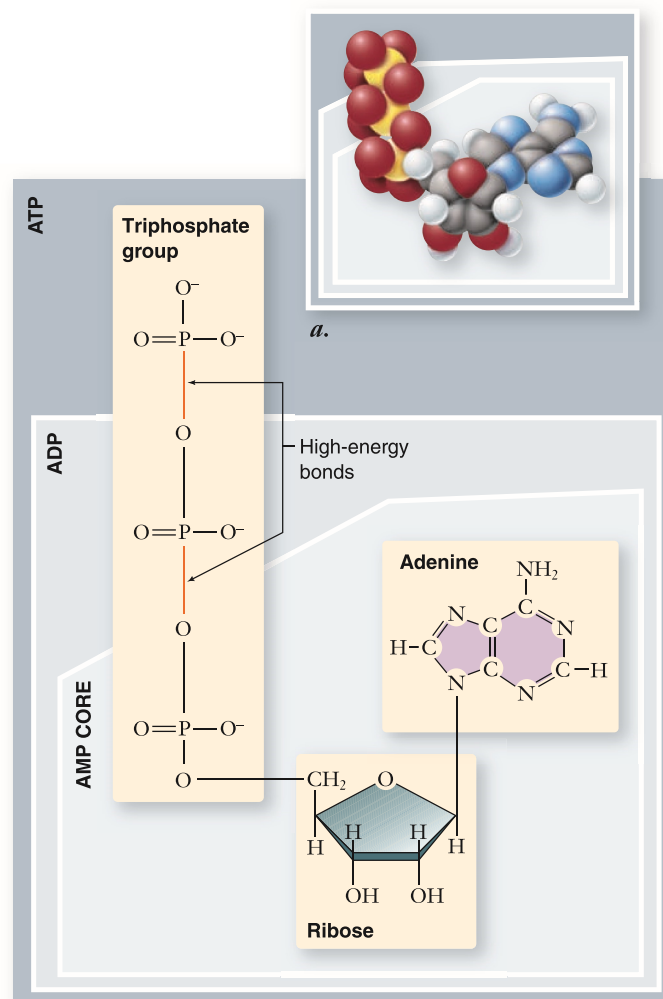
Cells store and release energy in the bonds of ATP

You saw in chapter 3 that nucleotides serve as the building blocks for nucleic acids, but they play other cellular roles as well. ATP is used as a building block for RNA molecules, and it also has a

critical function as a portable source of energy on demand for endergonic cellular processes.

The structure of ATP

ATP is composed of three smaller components (figure 6.6). The first component is a 5-carbon sugar, ribose, which serves as the framework to which the other two subunits are attached. The second component is adenine, an organic molecule composed of two carbon–nitrogen rings. Each of the nitrogen atoms in the ring has an unshared pair of electrons and weakly attracts hydrogen ions, making adenine chemically a weak base. The third component of ATP is a chain of three phosphates, thus adenosine *triphosphate*.



b.

Figure 6.6 The ATP molecule. The model (a) and the structural diagram (b) both show that ATP has a core of AMP. Addition of one phosphate to AMP yields ADP, and addition of a second phosphate yields ATP. These two terminal phosphates are attached by high-energy bonds so that removing either by hydrolysis is an exergonic reaction that releases energy. ADP, adenosine diphosphate; AMP, adenosine monophosphate; ATP, adenosine triphosphate

How ATP stores energy

The key to how ATP stores energy lies in its triphosphate group. Phosphate groups are highly negatively charged, and thus they strongly repel one another. This electrostatic repulsion makes the covalent bonds joining the phosphates unstable. The molecule is often referred to as a “coiled spring,” with the phosphates straining away from one another.

The unstable bonds holding the phosphates together in the ATP molecule have a low activation energy and are easily broken by hydrolysis. When they break, they can transfer a considerable amount of energy. In other words, the hydrolysis of ATP has a negative ΔG , and the energy it releases can be used to perform work.

In most reactions involving ATP, only the outermost high-energy phosphate bond is hydrolyzed, cleaving off the phosphate group on the end. When this happens, ATP becomes *adenosine diphosphate (ADP)* plus an **inorganic phosphate (P_i)**, and energy equal to 7.3 kcal/mol is released under standard conditions. The liberated phosphate group usually attaches temporarily to some intermediate molecule. When that molecule is dephosphorylated, the phosphate group is released as P_i .

Both of the two terminal phosphates can be hydrolyzed to release energy, leaving *adenosine monophosphate (AMP)*, but the third phosphate is not attached by a high-energy bond. With only one phosphate group, AMP has no other phosphates to provide the electrostatic repulsion that makes the bonds holding the two terminal phosphate groups high-energy bonds.

ATP hydrolysis drives endergonic processes

Cells use ATP to drive endergonic processes. Remember that endergonic reactions do not proceed spontaneously because their products possess more free energy than their reactants; ΔG for an endergonic reaction is positive. If the ATP hydrolysis releases more energy than the other reaction consumes, then coupling the two reactions will produce an overall ΔG for the coupled reactions that is negative. That is, energy released by the hydrolysis of ATP can supply the energy needed by the endergonic reaction. Because most endergonic reactions in cells require less energy than is released by ATP hydrolysis, ATP can provide most of the energy a cell needs. ATP also powers other cellular functions such as generating force in muscles, or creating concentration gradients of important ions, both processes that require energy.

Data analysis Consider the reaction: glutamate + $\text{NH}_3 \longrightarrow$ glutamine ($\Delta G = +3.4$ kcal/mol). If this reaction is coupled to ATP hydrolysis ($\Delta G = -7.3$ kcal/mol), what would be the overall ΔG ? Would this process be endergonic or exergonic?

ATP cycles continuously

The same feature that makes ATP an effective energy donor—the instability of its phosphate bonds—make it a poor long-term energy-storage molecule. Fats and carbohydrates serve that function better.

Cells make and use ATP cyclically: Cells use exergonic reactions to provide energy to synthesize ATP from $\text{ADP} + P_i$; they then use the hydrolysis of ATP to provide energy to drive

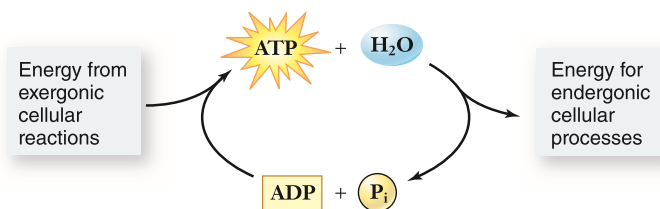


Figure 6.7 The ATP cycle. ATP is synthesized and hydrolyzed in a cyclic fashion. The synthesis of ATP from $\text{ADP} + P_i$ is endergonic and is powered by exergonic cellular reactions. The hydrolysis of ATP to $\text{ADP} + P_i$ is exergonic, and the energy released is used to power endergonic cellular functions such as muscle contraction. ADP, adenosine diphosphate; ATP, adenosine triphosphate; P_i , inorganic phosphate

endergonic processes (figure 6.7). Most cells do not maintain large stockpiles of ATP. Instead, they typically have only a few seconds' supply of ATP at any given time, and they continually produce more from ADP and P_i . It is estimated that even a sedentary individual turns over an amount of ATP in one day roughly equal to his or her body weight. This statistic makes clear the importance of ATP synthesis. In chapters 7 and 8 we will explore in detail the cellular mechanisms for synthesizing ATP.

Learning Outcomes Review 6.3

ATP is a nucleotide with three phosphate groups. Endergonic cellular processes can be driven by coupling to the exergonic hydrolysis of the two terminal phosphates. The bonds holding the terminal phosphate groups together are easily broken, releasing energy like a coiled spring. The cell is constantly building ATP using exergonic reactions and breaking it down to drive endergonic reactions.

- If the molecular weight of ATP is 507.18 g/mol, and the ΔG for hydrolysis is -7.3 kcal/mol, how much energy is released over the course of the day by a 100-kg man?

6.4 Enzymes: Biological Catalysts

Learning Outcomes

1. Discuss the specificity of enzymes.
2. Explain how enzymes bind to their substrates.
3. List the factors that influence the rate of enzyme-catalyzed reactions.

The chemical reactions within living organisms are regulated by controlling the points at which catalysis takes place. Life itself, therefore, can be seen as regulated by catalysts. This is possible because the catalysts in living systems, enzymes, are also highly specific. Most enzymes are proteins, although RNA molecules can also have catalytic activity.