POTENTIAL AND KINETIC ENERGY

Energy, Cells, and the Creation of Order

Do you think of yourself as the peaceful, laid-back type? Think again. Every second of your life is a standoff of epic proportions. In fact, your body is trying to defy some of the most fundamental laws of physics right now. "But wait," you ask, "Shmoop, why are we talking about physics in a biology unit?"

Because, whether or not you would like to admit it, physics and chemistry are so intimately connected to biology—ooh la la!—that it is impossible to understand biology without learning about physics and chemistry, too. Therefore, young Jedis, before we get into the nitty gritty of the epic standoff between your life and a world craving **disorder**, we must step back for a second and discuss some physics basics. It will prepare you well for what is to come. (Really, though, we knew you'd had enough chemistry from the last unit, and physics was the only other topic we could plausibly throw in here to increase the suspense of the good stuff that comes later on. Kidding, kidding. Maybe).

You know all about **energy**. Energy is what you never have enough of, what power companies provide you with and charge you lots for, and what politicians ramble on and on about. All of that is correct, but for now, you need to familiarize yourself with the physics-based definition of energy.

Here it is:

Energy is defined as the capacity to **work**, or rather, the capacity to create **change**.

What is **work**? **Work** is defined as what happens when energy is transferred from one system to another. In the physical sense, if we have two objects, and the first object transfers energy to the second object, the first object is said to have done work on the second object.

Back to energy. There are many different types of **energy**:

- **Potential energy**
- **Kinetic energy**
- **Thermal energy** (aka heat)
- Electromagnetic radiation (aka light, microwaves, radio waves, TV waves, X-rays, gamma rays)
- Chemical energy

Kinetic energy is the energy that an object has because of its **movement**, while **thermal energy**, or **heat**, is released when you burn wood in a bonfire (yum, s'mores).

Potential energy is used to describe "unused" energy that has the ability to accomplish work but currently isn't. A good example of potential energy is an object that has the ability to do work because of its position in the gravitational field.

Pretend you are at Disneyland, or Disney World (count us in; we can be packed in under 30 minutes, and if you bring the tickets, we'll bring our savant-like knowledge of hidden Mickey locations). Picture yourself in the railroad car on the Big Thunder Mountain Railroad ride. Clankity-clank-clank-clank-clank, and you are soon at the top of the hill at the start. You know that second where you know you are about to go down, and the fun is about to begin? At that exact moment, the train has a lot of potential energy because gravity will soon get to work, speeding you downhill and through the rock formations and caverns of the mine.

Another example of potential energy is a rock that has been pulled back in a slingshot. Or, better yet, a bird in a slingshot, pointed at some green pigs with eggs under blocks of wood and rocks. How did we create such a scenario? Beats us. Either way, that little red birdie is full of potential energy before he is shot through the air.

We will talk about energy in more depth later, but the important concept to grab here is that energy is the workhorse of change.

Brain Snack

You have probably heard that ATP is the "energy currency" in the cell. How is the potential energy stored in ATP? As it turns out, the ability to release the phosphate group at the end of ATP is what makes it so high in energy.

THE SECOND LAW OF THERMODYNAMICS

You may have figured out that no good standoff, either between biology and physics or you and your folks on Saturday night, can happen without a good amount of energy. What is this biological standoff we speak of? Only one of the most fundamental laws of physics, referred to as the **second law of thermodynamics**. Don't you feel silly? Why would we discuss the first law of thermodynamics, first? Making sense is overrated.

The second law of thermodynamics states that the degree of **disorder** is always increasing in the universe. The amount of disorder in a system can be quantified, and this measurement is called **entropy**. You will often see our new friend entropy expressed as the letter *S*. Don't be fooled; it's still entropy. As you might imagine, creating order in a universe that is always headed toward greater disorder requires a good deal of energy. The universe has made her decision, and there is no talking her out of it; at this point, there isn't much to do but stay strong and keep going.

Now you know why life comes down to an everlasting standoff between order and disorder. Life strives to create ordered structures while the disorder-loving universe rains on his parade as much as possible. Not cool, Universe. Not. Cool. (Not to be a Debbie Downer, either, but the universe generally wins these standoffs when all is said and done. We'll let you know if the outcome is ever in Life's favor.)

What does it mean that the amount of disorder is increasing? Let's think about your…biology binder. The three-ring binder keeps all the pages in the correct order (or it would, if you ever put anything in it). **The second law of thermodynamics** as applied here means that if you accidentally drop your binder, and the rings pop open, all of the pages will be mixed together. Obviously.

Think about why this is. According to the biology teacher and the class, there is only one correct order for the pages to be in, and many more ways for the dropped pieces of paper to land on the ground. Putting the papers back in the right order takes energy.

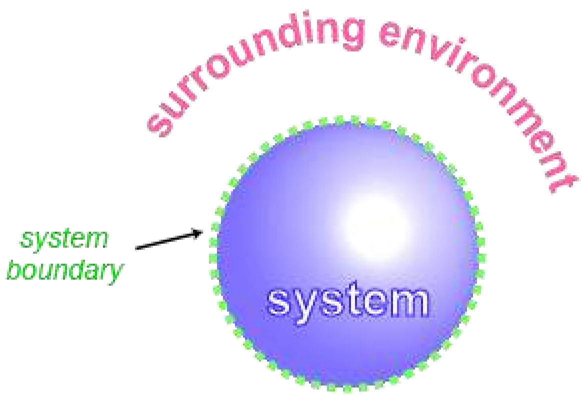
We can also think of the second law of thermodynamics in terms of organizing your bedroom. Your room naturally tends toward **disorder** (pretty crazy that we knew that, right?), or room arrangements with greater entropy, and keeping it all tidy and clean takes a lot of work. The next time that someone, anyone for that matter, tells you to clean your room, tell her that you are letting the **second law of thermodynamics** run its course. It's best to be one with Nature, after all.

This person that rudely commented on the state of your belongings will probably be quick to realize that your excuse for not cleaning your room is a poor one (no offense). Unfortunately for you, the second law of thermodynamics applies to *the whole universe*, and your room is one small component of a ginormous universe. Therefore, you *can* clean your room, and as long as the entire universe is still headed toward disorder, all is right in the world. Sorry to burst your bubble.

Actually, it comes down to basic math. All the disorder in the rest of the universe is to compensate for the little bit of order in your newly immaculate and well-kept room. Your room is what we call an **open system**, which means that energy, entropy, and even you and your stuff can pass freely from the walls of your room to the rest of the universe. Nice try locking the door.

On the other hand, a **closed system** can exchange energy, but not **matter** (matter = anything that has a mass and occupies space), with its surroundings. An **isolated system** cannot exchange anything with its surroundings. If your room were an isolated system, you might be excused from cleaning, but you would also be grounded. For life. Physics is not cutting you much of a deal here, is he?

Here is a generic picture of a system. The dotted line is the determining factor. Can the system exchange everything? If so, it's an open system. Everything but matter? Then it's a closed system. Nothing? Well, in that case, we're looking at an isolated system.



This might blow your mind, but having a clean room in a disordered universe is similar to the cells in your body. Your body is practically the living definition of organization. First, you have your organs, which are composed of cells, which in turn are organized into compartments like the nucleus, mitochondria, endoplasmic reticulum…and on and on. You might wonder how on Earth your body could possibly become ordered when the disorder of the universe is always increasing, right? The answer is that a cell *takes energy* from its environment to generate order. In turn, the cell releases energy as heat to the environment, giving it what it wants and making it more disordered.

You might also be wondering why adding heat to the environment makes it more disordered. **Heat** is basically the most disordered form of energy. When it is released to the environment, it speeds up the random movement of atoms and molecules in its path. This randomness is an excellent example of increasing **entropy**.

Let's contemplate the opposite scenario, where the temperature of a system decreases because heat flows out of a system. In this case, atoms and molecules move more slowly until they can't move at all. This is in fact a more ordered state of being, and therefore, the entropy of the system has decreased. Adding more heat to the system by increasing the temperature will increase the disorder of this system once again.

Brain Snack

One way that the cell creates order is through the generation of a membrane potential (read: the concentration of hydrogen ions across the membrane). A cell has a membrane potential of about 300 millivolts (mV). This amount may not seem like a lot to you, but at the small scale of a cell, that potential accomplish do a lot.

THE FIRST LAW OF THERMODYNAMICS

Why can't heat energy be easily converted back into another form of energy? Excellent question. This belatedly introduces **the first law of thermodynamics**. Talk about being late to the party.

The first law of thermodynamics states that energy can be transferred from one form to another, but it can never be created or destroyed. All right, then how can heat energy be easily turned into another useful form of energy? It can't, really. As we mentioned earlier, not all types of energy are created equal. In fact, it is for this reason that we worry about depleting our fossil fuels; once they're gone, they're gone. No do-overs or takebacks.

Think of **heat energy** as basically the lowest guy on the energy ladder. This disordered form of energy cannot be easily converted back into another type of usable energy. Heat is at the bottom and has no promotions in his future. He's like the Milton Waddams of energy.

It is important to remember that although the first law of thermodynamics says that energy cannot be created or destroyed, this law applies to *the total sum of energy in the universe*. Which is a heck of a lot. Since a system is one piece of the universe, the change in energy *in a system* can be either positive, negative, or zero.

This concept can be written as a handy formula: $\Delta E = Q - W$, where the change in energy (ΔE; delta signs, or Δ's, mean "change in") in a system equals the heat energy (Q) minus the work energy (W). If heat energy is lost to the surroundings, like when fossil fuel is burned, the Q of the system will be negative. If the energy content of the fuel is transferred to create work as well, the total change in energy (ΔE) will be negative and equal to the amount of work energy (W) done plus the amount of energy lost as heat (Q).

In summary, then, the first law of thermodynamics states that if a system loses energy, it has lost heat to the surroundings, or done work. In this way, energy is not created or destroyed in the universe as a whole. The catch is that all forms of energy are *not* equal, so the amount of **useful energy** in the universe can still be decreasing. Hey, we warned you that the universe has a pretty remarkable track record when it comes to winning these standoffs. (100% is pretty remarkable, right?)

Brain Snack

Useful energy comes in many different forms. You probably know that you use the energy found in food to convert it to useful work energy in your body. However, other organisms, like bacteria, are able to use energy found in all sorts of things that would be toxic to us, like hydrogen sulfide, uranium, iron oxides, and of course, light.

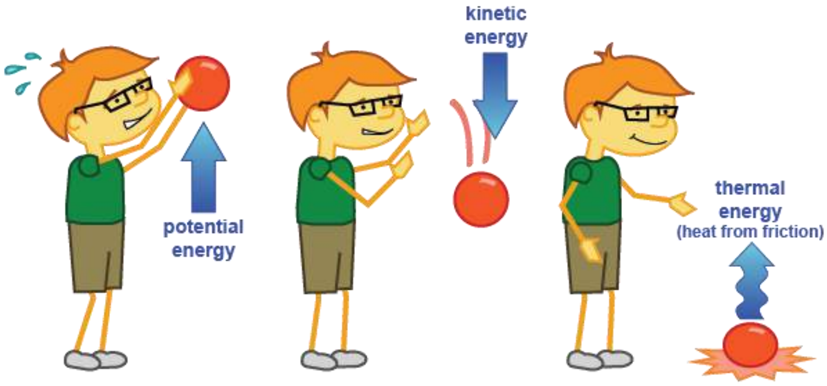
ORDER

What makes your cells amazing is that they use **heat energy** to produce **order**. See what we did there? Brought the biology back full circle—that's Shmoop. One thing that will become clear as we investigate different topics in biology is that a cell's precious life is completely dependent on the transfer of one type of energy to another. The main chemical energy carrier of the cell is **ATP**, short for adenosine triphosphate, and cells convert energy from ATP to many things, including:

- **Mechanical energy** to move organelles (read: the machinery of the cell)
- **Electrical energy** to move ions against a concentration gradient
- **Thermal energy** to release heat

In fact, one of the most important energy transfer events, perhaps **the** most important transfer event, for life on our planet is the biological process called **photosynthesis**, which occurs in most plants, algae, and some bacteria. Photosynthesis is the process wherein **light energy**from the sun is converted into **chemical energy** to make organic compounds, or food (mainly sugars), for the organism.

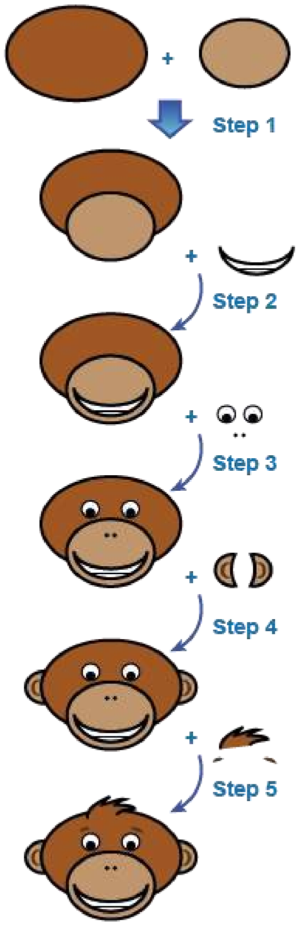
Behold, the different types of energy in context! We think our friend needs to work on his upper body strength.



Photosynthetic organisms make their own nutrients, but animals like us depend completely on taking in organic compounds from our environment (delicious food). Eating isn't *that* bad, right? In fact, it's quite a tasty experience as long as the food is good. But, our need to find food outside our bodies means that photosynthesis also fuels our existence, albeit a little less directly.

Through a fundamental life process called **respiration**, organisms, including plants, obtain their energy from the organic compounds they consume. To do this, said organisms allow their carbon and hydrogen compounds to combine with oxygen in the environment to produce carbon dioxide (CO_2) and water (H_2O). This actually happens in numerous steps through a number of biochemical reactions collectively known as a cell's **metabolism**. These different biological reactions can be grouped together into **pathways**. A key feature of a **metabolic pathway** is that each intermediate is used in the next step, which eventually results in the production of the final product.

A pathway can be thought of as individual reactions that each work toward a common goal. Take, for example, this lump of brown clay.



What an adorable little monkey.

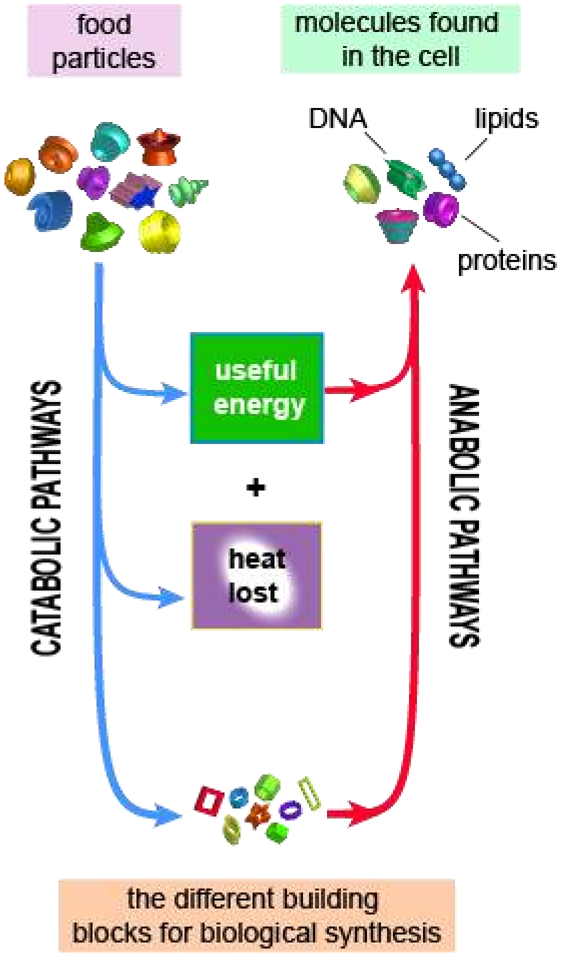
Metabolism can be further divided into two *opposing* biological processes called **anabolism**and **catabolism**. **Catabolism** refers to the processes, or **catabolic pathways**, by which cells break down food into smaller molecules, in turn releasing energy. This energy is used in **anabolism**, or anabolic pathways, to build new and larger molecules. More specifically, the smaller molecules generated in catabolic pathways are used in anabolic pathways as building blocks to create large molecules such as proteins, lipids, and DNA.

In this way, anabolic and catabolic pathways are **interdependent**. During the chemical reactions that comprise a cell's metabolism, **electrons** are transferred from one chemical reaction to another.

- When electrons are removed from one biological molecule, that molecule is said to be **oxidized**.
- When electrons are transferred to a biological molecule, that molecule is said to be **reduced**.

All in all, together anabolic and catabolic pathways generate the order that drives life. We will discuss metabolic pathways in further detail, as well as the enzymes that make them possible, later on in this unit. We bet you can't wait.

Here is a simple demonstration of how food molecules, cell components, and biological molecules all relate through catabolic and anabolic pathways.



Is anybody else getting a Perfection flashback after scoping out those particle synthesis building blocks? Just us then. Great.

Ready for a random question? Good. How do the processes going on inside a cell contribute to the rest of the universe? We have just informed you that when you eat your double bacon cheeseburger, your body uses catabolic pathways to take the energy from the food and convert it into useful chemical molecules and cellular order. It turns out that some of this energy from the breakdown of food is also released back into the environment as **heat**, CO_2, and H_2O, which in turn adds to the increasing disorder of the universe. Didn't see that coming, didya? Universe, 2, You, 0.

Brain Snack

Some organisms possess the ability to convert chemical energy into light energy. An example? The firefly.

ENTHALPY

Looking at Cells Under the Energy Microscope

When we discussed heat and entropy, it became crystal clear that thermodynamics is critically important to cells and cell function. Now, let's take a closer look at the involvement of thermodynamics in cellular biochemical reactions. The second law of thermodynamics (yep, it's back for more) can be interpreted in the context of chemical reactions, where **energy transfer events** tend to progress downhill—that is, products have less energy than reactants—and any extra energy produced during chemical reactions is lost as heat, which ultimately increases the disorder in the universe. In other words, once an energy transfer event occurs, there is less energy left in the product to do additional work.

Cellular reactions that occur spontaneously will proceed to a more disordered, but lower energy state. This may seem counterintuitive if you were thinking that increasing disorder always means increasing energy in the products, but this is not the case. Remember that cellular reactions release heat to the surroundings as they form products that are themselves in a lower energy state.

Proteins that have been broken down into their individual amino acids (see the Biomolecules and the Chemistry of Life unit if this is not ringing a bell) by a cellular reaction will not suddenly spontaneously re-form intact proteins. Instead, any **uphill reaction** is a reaction that involves a large input of energy, such as the reactions that build proteins or nucleic acids. However, when we start trying to predict if a reaction will occur, or why a protein takes on a certain structure, it is not enough to rely on the second law of thermodynamics. Instead, we need to combine **the first and the second laws of thermodynamics**.

The first and second laws of thermodynamics were combined into one equation by Josiah Willard Gibbs (not Barry, Robin, or Maurice) in the late 1800s. The equation is

$\Delta H = \Delta G + T\Delta S$,

or the change in something called **enthalpy** (ΔH) equals the change in something called **free energy** (ΔG) plus the **absolute temperature** (T, in degrees Kelvin, where x degrees Kelvin = $y$degrees Celsius + 273) multiplied by the change in **entropy** (ΔS). Did you catch all that?

This equation is useful (it is; stop shaking your head) because it allows scientists to find out information about a reaction while only knowing a few details about the system. This equation is not nearly as equation-y as it may seem at first glance.

You might say, "But Shmoop, what exactly *is* **enthalpy**?" Sorry, you are not privileged enough to receive such information. In all seriousness, you will learn more about enthalpy in chemistry, but for our purposes, **enthalpy** *is equal to the total amount of energy in a system*.

If a system gains heat from a chemical reaction, ΔH will be positive. Alternatively, if heat is lost from a system, ΔH will be negative. Not hard, right? Then, there is the term $T\Delta S$. We know that entropy is a measurement of disorder. In this case, $T\Delta S$ takes into account *only the entropy change of the system*.

- If a reaction increases the disorder in a system, the entropy term $T\Delta S$ will be positive.
- If a reaction decreases the disorder in a system, the entropy term will be negative.

Now, what about free energy? You wish gasoline were free... Oh, right, ΔG. Coming right up.

Brain Snack

Spontaneous combustion—when something starts to be consumed by fire without external ignition—is real! And, in reference to our enthalpy discussion above, spontaneous combustion occurs if the heat in the system increases without being able to escape.

FREE ENERGY

We only talk about free energy when it changes (ΔG). Much like how you only talk about your friends when they aren't listening. Basically, ΔG describes the change in **useful energy** that can be extracted from a particular reaction to do **work**. Let's rewrite the Gibbs equation in terms of ΔG:

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

Now the Gibbs equation says that the amount of useful, or free, energy (ΔG) that can be extracted from a reaction equals the total energy change (ΔH) minus the energy that cannot do work because it increased the disorder of the system ($T\Delta S$). Make a little more sense? Good.

Why is solving for ΔG useful? ΔG tells us whether a reaction will proceed or not, which is a nice thing to know.

- If ΔG is negative, the reaction is said to be **exergonic**, meaning that the process will occur spontaneously and is **thermodynamically favorable**. Heat will often be released, which in turn creates disorder.

- If ΔG is positive, then the reaction is said to be **endergonic**, meaning that it is **thermodynamically unfavorable**. In other words, this type of reaction needs the input of energy to make it happen.

The only way that an endergonic reaction can occur spontaneously is if it is linked to an exergonic reaction such that the sum of the two ΔG reactions is negative. A good analogy would be when an introvert breaks out his party hat whenever his extroverted buddy stops by.

This process of hooking together is called **coupling**, and your cells use coupling all the time to perform many of the energetically unfavorable reactions in a cell. We can think of coupling as using the energy gained from a **downhill reaction** (a spontaneous and favorable one), to help power an **uphill reaction** (a nonspontaneous and energy-needy one).

What happens when $\Delta G = 0$? This type of reaction is considered to be at **equilibrium**, and therefore, there is no net change of energy going in or out of the system. We will go into the ins and outs of achieving equilibrium a little later. The major take-home message for understanding the Gibbs equation is that it lets you predict the outcome of a reaction. Think of it as ESP in the thermodynamic dimension.

If you are having a hard time distinguishing the difference between the change in entropy ($T\Delta S$) and the change in free energy (ΔG), this paragraph is for you. At first, it may seem confusing. Contradictory even. A reaction that increases the disorder in a system is said to increase the entropy. You already know that the universe is relentlessly moving toward greater entropy. Your Spidey sense is right if you think that a reaction that causes a system to become extremely disordered (positive $T\Delta S$) is likely to be spontaneous (negative ΔG—dangerous!). Unfortunately, knowing only the change in energy in a system or knowing only the change in entropy in a system is not enough to predict whether a reaction will occur spontaneously. You need to know **both** factors. This need is what makes free energy such an important factor in the equation. Hehe. A reaction with a negative ΔG *always* occurs spontaneously.

Time to think about making popsicles. Never done it? You are missing out. You pour some juice into a mold, and freeze it. Easy as pie. Actually, way easier than pie. Whoever came up with that idiom had obviously never made popsicles. Freezing the juice decreases the **entropy**of the system because the molecules become more rigid and in an ordered state. For juice to freeze, a large amount of heat needs to be released from the system. When do you think that the popsicles will freeze? If you said "when you put it in the freezer," congratulations, you win! But, why will the popsicles freeze only at temperatures below 0 °C? Because, only at temperatures below 0 °C will ΔH be negative *enough* to compensate for the negative $T\Delta S$ value.

Here is a table to make your life a tiny bit easier.

Type of Reaction	**Spontaneous Reaction?**	**Free Energy Change (ΔG)**	**Entropy Change ($T \Delta S$)**	**Enthalpy Change (ΔH)**
Exergonic	Yes	Negative, and $\Delta G < 0$	Positive, or increasing, and $T \Delta S$ 0	System may absorb or release heat, so $\Delta H > 0$ OR $\Delta H < 0$ If $\Delta H > 0$, then $\Delta H < T \Delta S$
Exergonic	Yes	Negative, and $\Delta G < 0$	Negative, or decreasing, and $T \Delta S < 0$	System must release heat, so $\Delta H < 0$, and $\|\Delta H > \| T \Delta S \|$
Endergonic	No	Positive, and $\Delta G > 0$	Positive, or increasing, and $T \Delta S > 0$	System must absorb heat, so $\Delta H > 0$, and $\Delta H > T \Delta S$
Endergonic	No	Positive, and $\Delta G > 0$	Negative, or decreasing and $T \Delta S < 0$	System may absorb or release heat, so $\Delta H > 0$ OR $\Delta H < 0$ If $\Delta H < 0$, then $\|\Delta H< \| T\Delta S \|$

What? We *did* say "tiny bit.

Now that your head is full of *H*s, *G*s, *S*s, and *T*s (HGST!), we will finally look at some biology, specifically, cellular reactions from a thermodynamic perspective.

First, know that all reactions are referred to as having directions: forward, neutral, or backward.

Second, know that all biochemical reactions are **reversible** even if one direction of the reversible reaction rarely occurs.

We like to refer to these reactions as **the forward reaction** and **the reverse reaction**, with **reactants** (that…react) on one side of the reaction equation, and products (that are…produced) on the other side of the reaction equation.

Important: If the ΔG of the forward reaction is negative, the reaction will proceed **spontaneously** forward in the direction of the products.

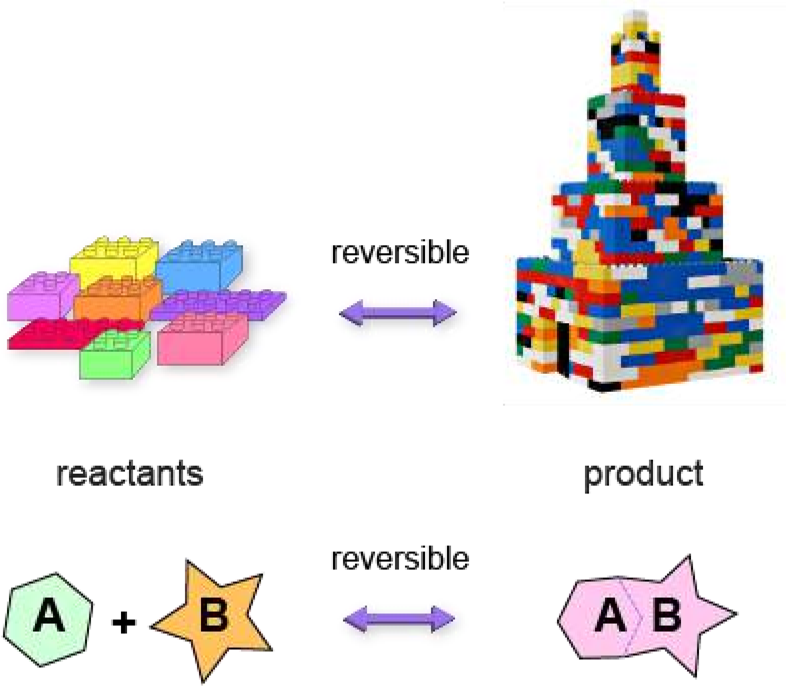
Example Reactions

$A + B \Rightarrow AB$

A and B are reactants, and AB is the product. This reaction proceeds forward (double arrow).

$A + B \Leftrightarrow C + D$

A and B are reactants, and C and D are products, OR C and D are reactants, and A and B are products (two-way arrow).



We know; we lied. We said that this would be all physics and no chemistry. Honestly, it's not our fault that everything in the universe is so well connected.

These tidbits are all fine and dandy, but (there is always a "but") knowing that a reaction is spontaneous is not enough to gauge *how* fast a reaction will occur. Sure, a reaction with a negative Δ*G* will occur, but it might be hundreds of years until you see any noticeable amount of product. If you think back to our earlier example, it would be like your parents saying that they will pay your allowance…sometime. Timing is pretty darn important.

What determines how fast a reaction will occur? **(We refer to this as the rate of the reaction.)**

There are two factors we will discuss:

1. Equilibrium
2. Temperature

Equilibrium

One important factor is the **concentration** of the *reactants and products*. In fact, chemical reactions will proceed until they reach **equilibrium**, or the point where the forward and reverse reactions are occurring *at the same rate*. At equilibrium, the amount of reactants used to make the product is the same as the amount of reactants produced from the product. That is, the forward reaction rate *equals* the reverse reaction rate. In *other* other words, at equilibrium, there is a constant ratio of reactants to products. Examples are 1:1, 2:1, and so on.

Equilibrium likes to make sure that sharing is at 100% and that there are no squabbles. Equilibrium is an awesome pie-slicer.

Temperature

Another important factor for determining the reaction rate is the **temperature** of the reaction. It turns out that most of the reactions in your body require surprisingly high temperatures to occur quickly—much higher than your body temperature of 37 degrees Celsius or 98.6 degrees Fahrenheit. What? How can that be possible?

Enzymes are the answer (finally, some biology). Enzymes are usually proteins, but not any old run-of-the-mill protein earns the enzyme label. Enzymes **catalyze**, or *accelerate*, reactions. Unlike **reactants** and **products**, which are either directly used or produced in a chemical reaction, enzymes change the rate the reaction *but are not consumed or produced in the reaction*.

Enzymes are conservative little fellows that won't change for anybody or anything. Enzymes will not change based on the thermodynamics of a reaction. They won't add energy, entropy, or enthalpy, or temperature changes to a reaction, and they absolutely will not determine whether a reaction is thermodynamically favorable or not. They get down to bizness and make reactions occur much faster. Enzymes are the focus of our next section.

Well, that's a wrap for us on free energy. You have accomplished quite a lot by tolerating the physics presented here: by understanding thermodynamics, you will now be able to understand the biology of the cell on a whole new and improved level. Energy is not only what makes a cell run smoothly, but it also governs the basics of cell morphology (shape) and cellular biochemical reactions. If you so desired, you could look at all biological structures from an energy perspective.

Many of the "Why?" questions in biology can be answered by gaining a better understanding of thermodynamics. For example, "Why does a plasma membrane behave the way it does?" Answer: Check out the Big Theme section in this unit on structure and function.

Brain Snack

What does the word "enzyme" mean? Its origin is Greek, and it means "in yeast." In the late 1800s, there was a debate about how alcohol was produced from yeast. Louis Pasteur thought that there was something special about the yeast that allowed the reaction to occur. In 1897, the Buchner brothers ground yeast cells and showed that Pasteur was right—there was something special biological molecules in the yeast cells that promoted fermentation: enzymes. You'll probably appreciate that more in a few years.

COMMON MISTAKES

Energy, Cells, and the Creation of Order

- Sometimes, it is easy to think that energy gets completely used; however, in truth, while one form of energy may disappear, the amount of energy in the universe stays constant. The "vanished" energy has converted into a new energy form.

- The concept of entropy can be pretty confusing. Try thinking about entropy as a measurement, like when you measure distance or size. Your pile of laundry in the corner, for example, has more entropy than the clothes folded nicely in your drawers. What's that? Your drawers are missing, off their dresser tracks, or full of comic books? Nevermind. Go clean your room.

- The formula for the first law of thermodynamics is useful, but don't over-interpret the value of ΔE. Knowing whether ΔE is positive or negative does not tell you how likely a reaction is to occur.

Looking at Cells Under the Energy Microscope

- Do not confuse *enthalpy* with *entropy*. Entropy is the measurement of disorder in a system while enthalpy is, for our purposes, basically the total energy content of a system, including the energy that can be used as work + the heat energy that increases the disorder of the universe.

- If you are having a hard time figuring out whether a reaction will proceed spontaneously or not, plug the values into the Gibbs equation. If ΔG is negative, the reaction will occur spontaneously.

- We often get caught in the trap of thinking that if ΔG of a reaction is very negative, the reaction will proceed rapidly. Do not fall for this nonsense! ΔG tells you *nothing* about how fast a reaction will occur. Nothing.

Enzymes in Depth

- Often, people think that enzymes change a reaction that normally wouldn't occur (positive ΔG) to a spontaneous reaction (negative ΔG). This isn't the case. Enzymes lower the activation energy of the same reaction such that the reaction will have a negative ΔG and occurs more quickly.
- Remember that enzymes are like the fast-forward button on your remote. They don't alter the content of the movie; they increase the speed of the playback.