CMB311 Fall 2017, Monday Sep 25

Lecture 9 Notes

Introduction to Enzyme Kinetics

In this lecture we began our discussion of Enzymes and their behavior. This is one of the more challenging aspects of this course, but is important to understand how Enzymes work and why they are so important to biochemistry. Without Enzymes, biology would be impossible.

Introduction

Many chemical reactions that are potentially useful to biology are extremely slow. For instance, the conversion of sucrose (table sugar) to CO_2 and H_2O is extremely favorable *thermodynamically*, releasing lots of energy, but the 44 grams of sugar in your can of soda isn't likely to burst into flames anytime soon. Even though the equilibrium for the process strongly favors the conversion, the reactions required are very slow *kinetically*. In order to take advantage of such energy sources, biology has evolved Enzymes, or *catalysts* that greatly accelerate such reactions. Enzymes are noted for being both *extremely efficient* and *highly specific* for the substrates on which they act.

- Biological chemistry involves a vast number of chemical reactions.
  - breakdown of nutrients
  - biosynthesis of precursors
  - energy yielding reactions
- Many of these reactions occur too slowly to be biologically useful.
- Some reactions are *not* thermodynamically favorable.
  - biosynthesis of macromolecules
  - ATP synthesis
- Unfavorable reactions can be coupled to favorable ones to make them go forward.
- Thermodynamically favorable reactions can be accelerated by enzymes.

Thermodynamics

A few points to introduce important thermodynamic concepts.

- Thermodynamics is concerned with the relationship between heat and its ability to do work within a system.
- It deals with properties of macroscopic systems, but we will use it to describe processes occurring at the molecular level.
- It is critical to understand the distinction between *thermodynamics* and *kinetics*; kinetics describes the rate at which reactions go, while thermodynamics whether they will go.

Here are some conventions that are important to know when we discuss thermodynamics.

- We generally use the absolute temperature scale, Kelvin (abbreviated K, and not °K).
- Remember that the Kelvin scale is absolute, 0 K = -273.15 °C = -459.67 °F, where all molecular motion ceases.
- The absolute value of Kelvin is equal to a Celsius degree (that is, an increase in temperature of 1 K is equal to an increase of 1 °C).
- Energy in thermodynamics is expressed in terms of the unit joules, with the symbol J, or more conveniently for biochemists, kJ for kilojoules. 1 kcal = 4.184 kJ.
- As an example of a biochemical reaction, the hydrolysis of ATP (adenosine triphosphate) to ADP (adenosine diphosphate) and P_i (inorganic phosphate) yields 30.5 kJ mol^{-1} of energy. We'll see later on in the course how important this particular reaction is.

As a reminder, here are some points about Entropy

- Commonly thought of as the level of *disorder* in a system.
- Changes in entropy of a system require the transfer of energy.
- Think of entropy as the number of possible ways (microstates) molecules in a system can be arranged. Remember Ludwig Boltzmann and his famous equation

$$S = k_B \ln W$$

where:
S = entropy
k_B = Boltzmann's constant = 1.38065×10^{-23} J/K
W = the number of microstates corresponding to a given macrostate
(Incidentally, W stands for *Wahrscheinlichkeit*, German for probability)

The main aspects of thermodynamics to understand are:

G = Gibbs free energy

- The amount of energy capable of doing work during a reaction at constant temperature and pressure
- The *change* in free energy of a system as a reaction proceeds is defined as ΔG

H **= Enthalpy**

- The heat content of a reacting system, or potential energy found in chemical bonds.
- The *change* in heat (released or absorbed) during a reaction is defined as ΔH

S **= Entropy**

- The randomness or disorder in a system
- The *change* in entropy of a system during a reaction is defined as ΔS

The relationship between Gibbs free energy, enthalpy and entropy is given by the following equation:

$$G = H - TS$$

For our purposes, we are interested in *changes* in Gibbs free energy as a reaction proceeds. This is given by the following equations:

$$\Delta G = G_{products} - G_{reactants}$$

and

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

Units of ΔG and ΔH are J/mol (or kJ/mol)

Units of ΔS are kJ/mol•K

T = temperature in Kelvin; 25 °C = 298 K

Under 'standard' conditions (all components at 1M concentration, reaction at room temperature = 25 °C = 298 K), we use the equation for 'standard-state' free energy:

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

ΔG = The actual free energy change under specified conditions, including concentration

ΔG° = Standard Free Energy change, all reactants and products in their standard states, i.e. 1M; however, at 1M $[H^+]$, pH = 0, which is not compatible with biochemical processes. Therefore...

$\Delta G^{\circ\prime}$ = Standard Free Energy change for the biochemical standard state, all reactants and products at 1M except at pH = 7 and T = 298 K (delta-G-naught-prime). We will use this term from now on.

$$\Delta G = \Delta G^{\circ\prime} + RT \ln Q$$

Remember that since

$$\Delta G^{\circ\prime} = -RT \ln K_{eq},$$

$$\text{if } \Delta G = 0 \text{ then } Q = K_{eq}$$

- The **reaction quotient**, ***Q***, is the ratio of the concentrations of the products over the concentrations of the reactants *not necessarily at equilibrium*.
- This is distinct from ***K_{eq}*** which is the ratio of products over reactants *at equilibrium*.
- If $Q > K_{eq}$ then the reaction favors the reactants.
- If $Q < K_{eq}$ then the reaction favors the products.
- if $Q = K_{eq}$ then the reaction is already at equilibrium and $\Delta G = 0$.