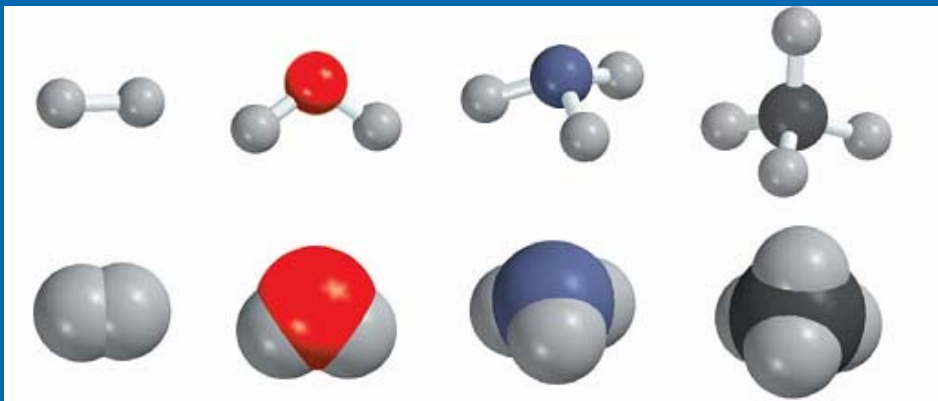


Welcome to Chemistry



If anyone wishes to search out the truth of things in serious earnest, he ought not to select one special science; for all the sciences are conjoined with each other and interdependent.

Descartes, Rules for the Direction of the Mind

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Program of course “**MEDICAL CHEMISTRY**” for the first-year students of the School of Medicine

Module 2. “Equilibrium in biological systems at the phase interface”. 3 lectures.

1. Chemical thermodynamics. Thermochemistry.
2. Kinetics of biochemical processes.
3. Electrochemical phenomena in biological processes.
4. Colloidal solutions and colloid stability.
5. Electric double layer and electrokinetic phenomena. Adsorption. Chromatography.
6. Macromolecules. Properties of polymers.

Practice for Module 2: 4 seminars, 4 tests, 4 praxis and final test.

Grades for Module 2

Work	Grade
Praxis:	15 points
Test	15 points
Final test	80 points
Total:	<p><u>For Module 2:</u></p> <p>60 points = 4 praxis × 15 points 60 points = 4 tests × 15 points 80 points = final test All: up to 200 points</p> <p>Final Graded Credit according to the “Four-level Grade” is calculated as average points of two Modules (No. 1 and No. 2) and placed to the student’s Diploma Appendix.</p>



V.N. Karazin Kharkiv National University

Medical Chemistry

Module 2. Lecture 4

Fundamentals of chemical thermodynamics
and bioenergetics.

Physicochemical fundamentals of kinetics and
catalysis

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Department of Physical Chemistry

October 12 , 2021

Lecture topics. Part I

- √ Chemical thermodynamics
- √ Energy, Work, Heat
- √ The first law of thermodynamics
- √ The thermochemistry
- √ The Hess's law
- √ The second law of thermodynamics
- √ The third law of thermodynamics
- √ Entropy
- √ Calculation of the entropy changes
- √ The Gibbs energy

Chemical thermodynamics

Thermodynamics is an extensive scientific discipline that deals with the interconversion of heat and other forms of energy.

Bioenergetics – the deployment of energy in living organisms.

The laws of thermodynamics provide useful guidelines for understanding the energetics and directions of processes.

All objects in thermodynamics are considered as a part of a system or the surroundings.

The system is the set of objects under consideration that separated from the surroundings by visual or imaginary boundary.



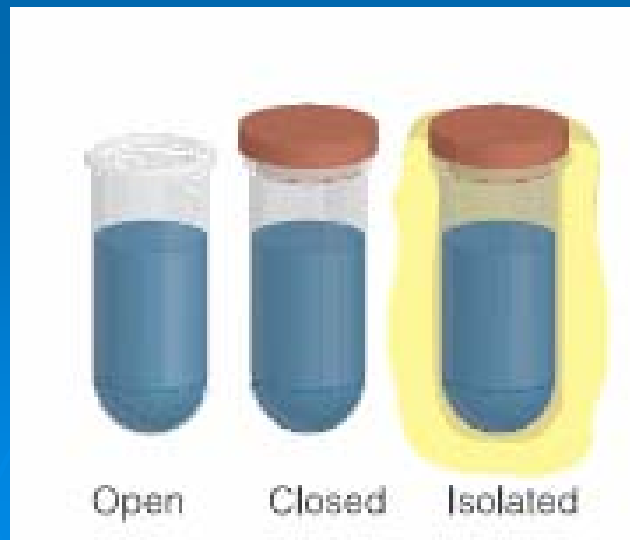
Chemical thermodynamics

All systems can be divided into 3 types according to *the flow of matter and energy* in and out of systems:

Open – both energy and matter of the system can be exchanged with the surroundings.

Closed – can be exchanged between the system and surroundings but not matter.

Isolated – neither matter nor energy can transfer between the system and surroundings.



Chemical thermodynamics

In thermodynamics, we study changes in the state of a system, which is defined by the values of all relevant macroscopic properties, for example, composition, energy, temperature, pressure, and volume.

Energy, pressure, volume, and temperature are said to be state functions – properties that are determined by the state of the system, regardless of how that condition was achieved.

In other words, when the state of a system changes, the magnitude of change depends only on the initial and final states of the system and not on how the change is accomplished.

Chemical thermodynamics

State functions depends only on the initial and final states of the system

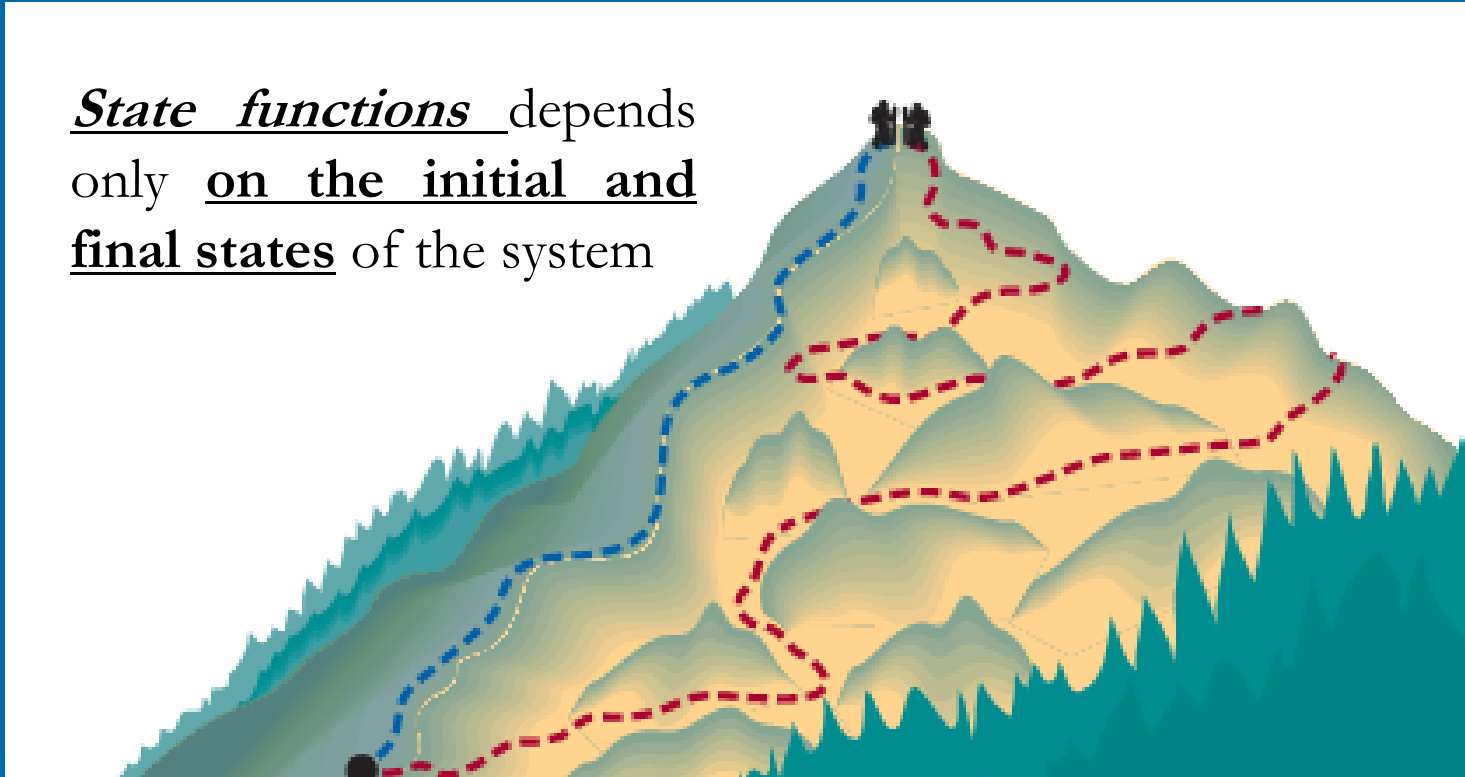


Figure. The gain in gravitational potential energy that occurs when a person climbs from the base to the top of a mountain is independent of the path taken.

Chemical thermodynamics

Energy, Work, Heat

The energy of a system may be defined as the capacity of the system to perform work.

The work is the process of moving against an opposing force. Work can be defined as force F multiplied by distance d :

$$W = F \times d$$

In thermodynamics, work has a broader meaning that includes mechanical work (for example, a crane lifting a steel beam), electrical work (a battery supplying electrons to light the bulb of a flashlight), and surface work (blowing up a soap bubble). When gas expands against a constant external pressure p the work done by the gas on the surroundings is $W = p \times \Delta V$, where ΔV – the change in gas volume.

The heat is the transfer of thermal energy between two bodies that are at different temperatures. Heat is denoted by Q .

However, we should always remember that 'work' and 'heat' are modes of transfer of energy, not forms of energy.

In general, heat and work are not state functions because they are not properties of a system.

Chemical thermodynamics

Internal energy

Any system possesses an internal energy which has two major components: kinetic energy of the motion of the system's particles (translations, rotations, vibrations) and potential energy of the interaction of all the atoms, ions, and molecules in the system and energy of *chemical bonds*.

Internal energy is designated by U . The internal energy of a system cannot be measured directly and there is not a way to calculate absolute value of U from other measured properties. It is possible to find only a change in U value by measurements.

When a process proceeds in a system it changes the state of the system from state 1 to state 2, as a result the internal energy changes from U_1 to U_2 , and the difference is

$$\Delta U = U_2 - U_1$$

The value of internal energy of a system depends on the temperature and the pressure.

Internal energy is state function, this means that the value of U depends only on state of the system and does not depend on the process path.

The first law of thermodynamics

The first law of thermodynamics is based on the law of conservation of energy:

energy can be converted from one form to another, but cannot be created or destroyed.

The first law of thermodynamics:

the increase in internal energy of a system is equal to the sum of heat transferred to it and work done on it.

This law may be written in form: $\Delta U = Q + W$

According this law, a perpetuum-motion machine, which produces work without consuming an equivalent amount of energy, is an impossible device to construct.

In biological organisms, energy in the form of food and nutrients is required for sustenance in order to perform the work required to live.

Chemistry in action: Making Snow

The secret of snowmaking is in the equation $\Delta U = Q + W$. A snowmaking machine contains a mixture of compressed air and water vapor at about 20 atm. Because of the large difference in pressure between the tank and the outside atmosphere, when the mixture is sprayed into the atmosphere it expands so rapidly that, as a good approximation, no heat exchange occurs between the system (air and water) and its surroundings; that is, $Q = 0$. (In thermodynamics, such a process is called an *adiabatic process*.) Thus, we write $\Delta U = Q + W = W$.

Because the system does work on the surroundings, W is a negative quantity, and there is a decrease in the system's energy. The change in energy ΔU is given by $\Delta U = C\Delta T$, where C is the proportionality constant. Because ΔU is negative, ΔT must also be negative, and it is this cooling effect (or the decrease in the kinetic energy of the water molecules) that is responsible for the formation of snow. Although we need only water to form snow, the presence of air, which also cools on expansion, helps to lower the temperature of the water vapor.



The thermochemistry

The thermochemistry deals with the heat changes accompanying chemical reactions.

If heat is liberated in the reaction the process is said to be exothermic, but if heat is absorbed it is described as endothermic.

Many reactions normally occur at constant (atmospheric) pressure so it is usual practice to record heat of the process by quoting the value of Q_p – the heat absorbed at constant pressure.

The first law of thermodynamics for the case when only work of gas expansion done gives:

$$Q_p = \Delta U - p\Delta V = (U_2 + pV_2) - (U_1 + pV_1)$$

The thermochemistry

Introducing new function of a system $H = U + pV$ one may rearrange previous equation to form:

$$Q_p = \Delta H$$

Function H is the state function called enthalpy.

Although heat is not a state function, the heat of a process at constant pressure is equal to change of state function because in this case “path” is defined and therefore it have only one specific value.

Enthalpy changes of constant-pressure processes are sometimes called “heats” of the processes.

At constant volume the heat of a process is equal to the change in internal energy:

$$Q_V = \Delta U$$

The thermochemistry

Heat of reaction is the difference in the enthalpies of the reaction products and of the reactants, at constant pressure, and at a definite temperature, with every substance in a definite physical state:

$$\Delta H_{\text{reaction}} = \sum H(\text{products}) - \sum H(\text{reactants})$$

Consider an arbitrary reaction of the type:



$$\Delta H_{\text{reaction}} = (cH_C + dH_D) - (aH_A + bH_B)$$

The thermochemistry

When for a process ΔH is negative the heat is actually evolved and this process is exothermic;

if ΔH is positive the heat is absorbed and the process is endothermic.

The chemical equation, that is accompanied by corresponding enthalpy change, is known as thermochemical equation.

This equation show the enthalpy changes as well as the mass relationships.

When writing thermochemical equations, one must always specify the physical states of all reactants and products, because they help determine the actual enthalpy changes.

For example, the thermochemical equation for the combustion of methane is:



The thermochemistry

The heat of formation ΔH_f of a compound is usually defined as the change of enthalpy when 1 mol of the substance is formed from the elements.

By convention, the heat of formation of any element in its most stable form is taken as zero.

The enthalpy change accompanying of complete combustion in oxygen of 1 mol of a compound is called the heat of combustion (ΔH_C).

The thermochemistry

The important law of thermochemistry was discovered experimentally by G.H. Hess. It is known as Hess's law or the law of constant heat summation.

The law states that the resultant heat change in a chemical reaction is the same whether it takes place in one step or several stages.

This means that the net heat of reaction, at constant pressure or constant volume, depends only on the initial and final states, and not on the intermediate states through which the system may pass.

The thermochemistry

It follows from Hess's law that the heat of a reaction is equal to the difference between the heats of formation of all the substances on the right-hand side of the equation of the reaction and the heats of formation of all the substances on the left-hand side (each multiplied by the stoichiometric coefficient of the equation).

For an arbitrary reaction:



It follows from Hess's law:

$$\Delta H_{reaction} = [c\Delta H_f(C) + d\Delta H_f(D)] - [a\Delta H_f(A) + b\Delta H_f(B)]$$

$$\Delta H_{reaction} = [a\Delta H_c(A) + b\Delta H_c(B)] - [c\Delta H_c(C) + d\Delta H_c(D)]$$

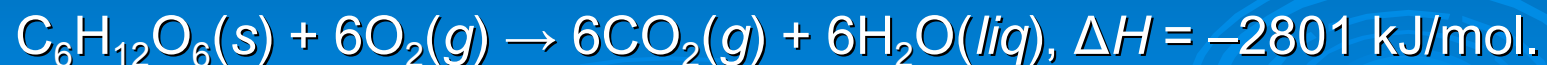
The thermochemistry

The food which we eat is metabolized in our bodies through several stages by a number of complex biological molecules called enzymes.

The energy released at each stage is captured for body function and growth.

The overall change in energy at metabolism is the same as it is in combustion.

For example, the total enthalpy change for the conversion of glucose ($C_6H_{12}O_6$) to carbon dioxide and water is the same whether we burn the substance in air or assimilate it in our bodies. In all these cases the process is represented by thermochemical equation:



The thermochemistry

Various foods have different compositions and hence different energy contents.

The energy content of food is generally measured in calories.
The calorie (cal) is a non-SI energy unit that is equivalent to 4.184 J.

Energy content of foods, that we can find on food wrappers, are the enthalpies of combustion.

Because the composition of particular foods is often not known, energy content values are expressed in terms of kJ/g rather than kJ/mol. In the context of nutrition we often use a “big calorie” instead of calorie that is equal to a kilocalorie:

$$1 \text{ Cal} = 1000 \text{ cal} = 4184 \text{ J.}$$

Note, we use capital “C” for representation the “big calorie.”

Table shows the energy content values of some foods and fuels.

Energy content of food and fuels

Food	ΔH of combustion, kJ/g	Fuel	ΔH of combustion, kJ/g
Apple	-2	Charcoal	-35
Beef	-8	Coal	-30
Beer	-1.5	Gasoline	-34
Bread	-11	Kerosene	-37
Butter	-34	Wood	-20
Cheese	-18	Methane	-56
Eggs	-6	Acetylene	-50
Milk	-3	Octane	-48
Potatoes	-3	Methanol	-21

The 2-nd law of thermodynamics

One of the main objectives of thermodynamics is to predict whether or not a reaction will occur when reactants are brought together under a specific set of conditions.

This knowledge is important for the synthesizing of new compounds in a research laboratory, manufacturing chemicals on an industrial scale, or trying to understand the proceeding of biological processes.

A reaction that does occur under the given set of conditions is called a spontaneous reaction. If a reaction does not occur under specified conditions, it is said to be non-spontaneous.



A spontaneous and
a nonspontaneous
process

The 2-nd law of thermodynamics

In order to predict the spontaneity of a process, we need to use special thermodynamic quantity that is a measure of the direction of reaction.

This function is entropy (Greek, *en* – in, *trope* – transforming; was introduced by R. Clausius).

Entropy (designated as *S*) is often described as *a measure of how spread out or dispersed the energy of a system among the different possible ways.*

The definition of a *change in entropy* for a thermodynamic process (ΔS) was introduced by R. Clausius:

$$\Delta S = Q_{rev}/T.$$

That is, the change in entropy of a system is equal to the energy transferred as heat to it *reversibly* (Q_{rev}) divided by the absolute temperature at which the transfer takes place.

The 2-nd law of thermodynamics

L. Boltzmann showed that the entropy of a system is related to the natural logarithm of the number of thermodynamic probability of occurrence of the system state (W):

$$S = k \times \ln W,$$

where $k = 1.38 \cdot 10^{-23}$ J/K – Boltzmann constant.

Thermodynamic probability of occurrence of the system state equals the number of microstates that corresponds to a given macrostate of the system characterized by pressure, temperature and energy.

It follows that a system with fewer microstates has a lower entropy, a system with more microstates has a higher entropy.

The 2-nd law of thermodynamics

The connection between entropy and the spontaneity of a reaction is expressed by the *second law of thermodynamics*: the entropy of an isolated system tends to increase $\Delta S > 0$.

In case of open system the total entropy change for any process is the *sum* of the entropy changes in the system (ΔS_{sys}) and in the surroundings (ΔS_{surr}).

In general, we can mathematically express the second law of thermodynamics as follows:

for a spontaneous process: $\Delta S = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0$,

for an equilibrium process: $\Delta S = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 0$,

According to these equations the 2-nd law may be formulated as: the entropy of the “universe” increases in a spontaneous process and remains constant in an equilibrium process.

The 3-rd law of thermodynamics

Entropy value may be obtained using calorimetric data. This method is based on the third law of thermodynamics:

the entropy of a perfect crystalline substance is equal to zero at the absolute zero of temperature.

Consider a perfect crystalline substance in which every atom is located in a well-defined place and there is no spatial disorder either. In case of $T = 0$, the molecular motions in the substance are kept at a minimum and the number of microstates (W) is one (there is only one way to arrange the atoms or molecules to form a perfect crystal).

Under these conditions the Boltzmann's equation gives:

$$S = k \ln W = k \ln 1 = 0.$$

The 3-rd law of thermodynamics

Absolute entropy values

Starting with the knowledge that the entropy of a pure crystalline substance is zero at absolute zero, we can measure the increase in entropy of the substance when it is heated from 0 K to a given temperature.

The change in entropy, ΔS , is given by:

$$\Delta S = S_T - S_0,$$

$$S_T = S_0 + \Delta S \text{ and } S_T = \Delta S.$$

The entropy of the substance at $T = 298 \text{ K}$ is called the absolute entropy.

Standard entropy is the absolute entropy of a substance at $p = 1 \text{ atm}$ and $T = 298 \text{ K}$, it is designated as S° .

The standard entropy values of a compound is measured in $\text{J}/(\text{K}\cdot\text{mol})$.

Calculation of the entropy changes

The entropy change of a substance may be calculated from the temperature change and heat capacity of the substance, plus any phase changes that substance undergoes during heating or cooling.

If temperature of a substance is changed and no any phase transitions occurred the change in entropy is given by the equation:

$$\Delta S = C \ln(T_2/T_1),$$

where T_2 and T_1 – final and initial temperature of the substance, C – heat capacity of the substance that must be constant over the range of temperatures of interest. If the pressure is during the heating, we use the constant-pressure heat capacity, C_p , and if the volume is constant – constant-volume heat capacity, C_v .

Calculation of the entropy changes

The phase transitions (melting, evaporation, sublimation) are processes that occur at constant temperature and pressure. The change in entropy associated with the phase transition is determined by the equation:

$$\Delta S = \Delta H/T,$$

where ΔH – enthalpy change of the phase transition, T – temperature of the transition.

To calculate the change of entropy for a chemical reaction represented by the following general scheme:



The standard entropy change for the reaction is given by the difference in standard entropies of products and reactants:

$$\Delta S^\circ = (cS^\circ(C) + dS^\circ(D)) - (aS^\circ(A) + bS^\circ(B)).$$

Calculation of the entropy changes

When an isothermic process takes place and the heat is transferred from the system to the surroundings than:

$$Q_{\text{sur}} = -Q_{\text{sys}}.$$

For this heat transfer the change in entropy of the surroundings is:

$$\Delta S_{\text{sur}} = Q_{\text{sur}}/T = -Q_{\text{sys}}/T.$$

For constant-pressure processes the heat change is equal to the enthalpy change of the system, ΔH_{sys} . Therefore, the change in entropy of the surroundings, ΔS_{sur} , is determined by ΔH_{sys} :

$$\Delta S_{\text{sur}} = -\Delta H_{\text{sys}}/T.$$

This equation gives the possibility to calculate the total entropy change using only parameters of the system:

$$\Delta S_{\text{total}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = \Delta S_{\text{sys}} - \Delta H_{\text{sys}}/T.$$

The Gibbs energy

By the definition: $G = H - TS$.

All quantities in this equation pertain to the system.

The unit of G is unit of energy because both H and TS are in energy units. Because H , T , and S are state functions, G is a state function too. A change in Gibbs energy at constant temperature arises from changes in enthalpy and entropy and is:

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

The left side of the above equation equals the total change in entropy at constant temperature and pressure multiplied by temperature

$$\Delta G = -T\Delta S_{\text{total}}.$$

The Gibbs energy

The difference in sign between ΔG and ΔS_{total} implies that the condition for a process being spontaneous changes from:

$$\Delta S_{\text{total}} > 0 \text{ to } \Delta G < 0 .$$

At constant temperature and pressure in a spontaneous process the Gibbs energy decreases:

$$\Delta G_{p,T} < 0 .$$

$\Delta G_{p,T} < 0$ – the reaction is spontaneous in the forward direction,

$\Delta G_{p,T} > 0$ – the reaction of interest is nonspontaneous, the spontaneous is reaction in the opposite direction,

$\Delta G_{p,T} = 0$ – the system is at equilibrium, there is no net change in the system.

Lecture topics. Part II

- √ Chemical kinetics: fundamentals
- √ Rate law or the fundamental equation of chemical kinetics
- √ First order reactions
- √ Second-order reactions
- √ Effect of temperature on the reaction rate
- √ The activation energy
- √ Catalysis: fundamentals
- √ Enzyme catalysis
- √ Enzyme kinetics

Chemical kinetics

Chemical kinetics is the part of chemistry concerned with the rates, at which a chemical reaction occurs and reaction mechanisms.

The word "kinetic" suggests movement or change. Kinetics refers to the reaction rate, which is the change in the concentration of a reactant or a product with time



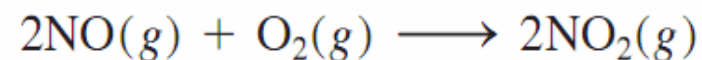
On a practical level, a knowledge of reaction rates is useful in drug design, in pollution control, and in food processing.

Chemical kinetics

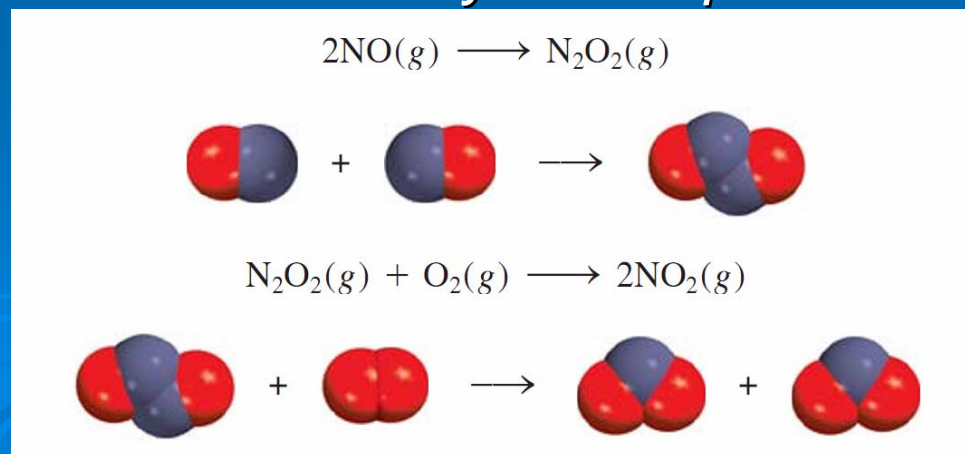
The way of a reaction can be broken into a series of elementary steps at the molecular level, and the sequence of such steps is the mechanism of the reaction. In a kinetics experiment, a chemist attempts to understand the step-by-step transformation of reactants to products.

Taken together these elementary steps gives us the mechanism by which the reaction proceeds.

As an example of a reaction mechanism, let us consider the reaction between nitric oxide and oxygen



Let us assume that the reaction actually takes place via two elementary steps as follows



Chemical kinetics

The molecularity of a reaction. According to the number of particles reacting in each elementary step they can be classified as

- 1) unimolecular, involving one molecule,
- 2) bimolecular, when two molecules react, and
- 3) termolecular in which three molecules react simultaneously, but the latter are very rare.

Tetramolecular reactions are unknown at all.

The rate of a chemical reaction is expressed as a change in number of moles of reactive species with time in the unity of system volume:

$$v = \frac{1}{V} \frac{\Delta n}{\Delta t}$$

here V – volume of the system in which reaction proceeds, Δn – change in number of moles of reagent or product species (molecules, atoms or ions), Δt – time that corresponds to this change.

Chemical kinetics

If volume kept constant the rate equals to the change in concentration (Δc) of some species with time:

$$v = \left(\frac{\Delta c}{\Delta t} \right)_V$$

or in more precise form

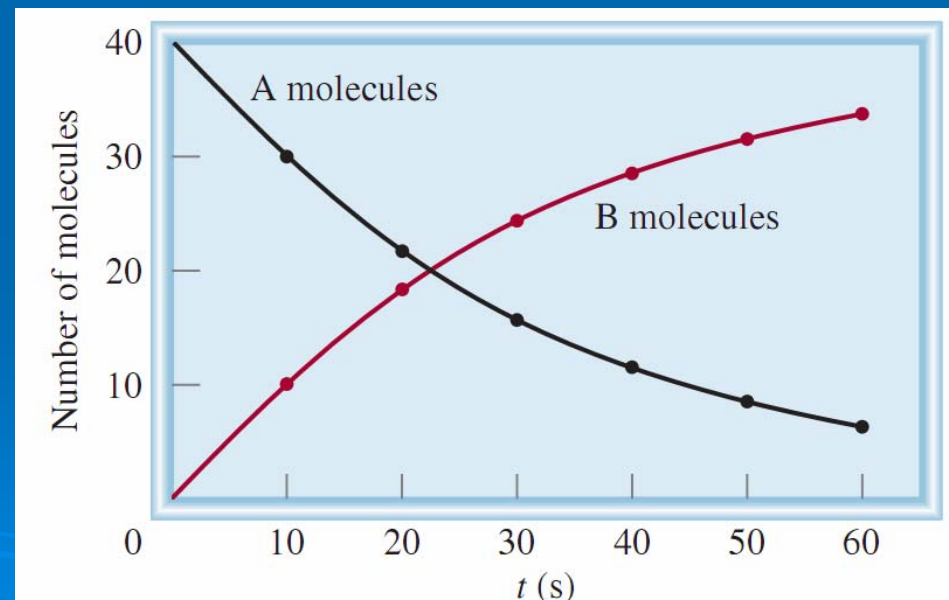
$$v = \left(\frac{dc}{dt} \right)_V$$

The units of the rate must be those of concentration divided by time: mol/(liter·sec) or mol/(liter·min), etc.

For example, a simple reaction in which A molecules are converted to B molecules:



The decrease in the number of A molecules and the increase in the number of B molecules with time are shown in Figure.



Chemical kinetics

If we consider the generalized reaction proceeding according to equation: $aA + bB \rightarrow pP$,

the concentration of reagents (A and B) is decreased as they are consumed and the changes in their concentrations will be negative, the concentration of the product (P) is increased with time and the change in its concentration will be positive. Consequently for the reaction rate we may write:

$$v = -\frac{1}{a} \frac{dc_A}{dt} = -\frac{1}{b} \frac{dc_B}{dt} = \frac{1}{p} \frac{dc_P}{dt}$$

Note the appearance of negative signs and stoichiometric coefficients in this equation; “-“ corresponds to change in concentrations of reagents and “+” – to change in concentrations of products. By defining the rate of reaction in this way the rate will always be a positive number that is independent of whether we follow a reactant or a product and that is independent of the overall balanced chemical equation.

Chemical kinetics

In aqueous solutions, molecular bromine reacts with formic acid (HCOOH) as follows:



Molecular bromine is reddish-brown in color. All the other species in the reaction are colorless. As the reaction progresses, the concentration of Br₂ steadily decreases and its color fades (Figure).



Figure. From left to right: the decrease in bromine concentration as time elapses shows up as a loss of color.

Rate law or the fundamental equation of chemical kinetics

The disappearance of reactant over time is proportional to the concentration of each reactant raised to some power. This power known as the order of reaction with respect that reactant. The sum of the individual orders is the overall order of the reaction. The order of reaction with respect to each reactant cannot be determined from the balanced chemical equation, it must be found experimentally. This statement can be written as:

$$v = kC_A^{n_A}C_B^{n_B}$$

where k – rate constant, n_A and n_B – orders of reaction on A and B substances respectively. This equation is known as the reaction's rate law and is the fundamental equation of chemical kinetics.

Chemical kinetics

Rate constant (k) is a constant of proportionality between the reaction rate and the concentrations of reactants. It equals the rate of reaction at reagent concentrations of 1 mol/l. The units of the rate constant vary with the order of the reaction.

Rate constants for the reactions are always determined experimentally:

using the dependence of concentrations of reactants on time we can determine the reaction order and then the rate constant of the reaction.

Chemical kinetics

Note that:

- a) the order of a reactant is not related to the stoichiometric coefficient of the reactant in the overall balanced equation;
- b) reaction order can also be zero and a fraction;
- c) reaction order is always defined in terms of reactant (not product) concentrations.

First order reactions

Consider a reaction at constant temperature of a single reactant:



There are many first-order reactions. An example is the decomposition of ethane (C_2H_6) into highly reactive fragments called methyl radicals (CH_3):



If there is no significant reverse reaction and if the reaction is first order, the rate law is:

$$v = -\frac{dc_A}{dt} = k_I c_A$$

where k_I – first-order rate constant, it has units of reciprocal time. This equation is a differential equation that can be solved by separation of variables.

First order reactions

Solving it one obtain:

$$\ln \frac{c_{A,t}}{c_{A,0}} = -k_I t$$

$$k_I = \frac{1}{t} \ln \frac{c_{A,0}}{c_{A,t}}$$

where $C_{A,0}$ – initial reagent concentration, $C_{A,t}$ – reagent concentration a moment of time t .

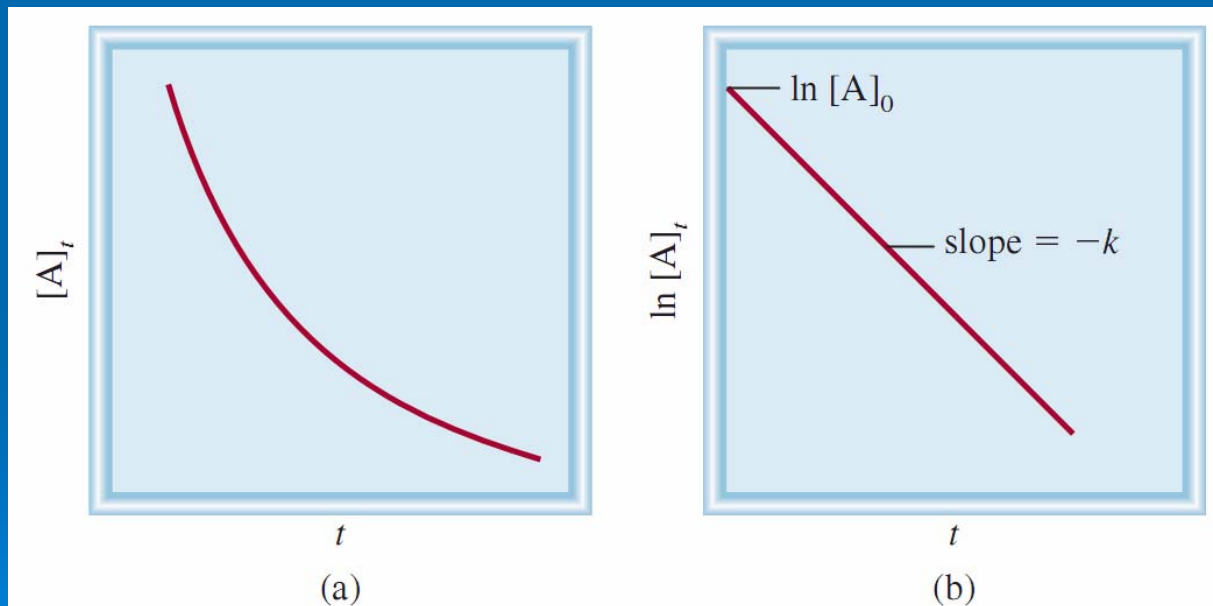


Figure. First-order reaction characteristics: (a) the exponential decrease of reactant concentration with time; (b) a plot of $\ln [A]_t$ versus t . The slope of the line is equal to $-k$.

First order reactions

As a reaction proceeds, the concentration of the reactant(s) decreases. Another measure of the rate of a reaction, relating concentration to time, is the half-life, $\tau_{1/2}$, which is the time required for the concentration of a reactant to decrease to half of its initial concentration:

$$\tau_{1/2} = \frac{\ln 2}{k_1}$$

Second-order reactions

As a second-order reaction may be considered some reactions that involves two reagents, A and B, and proceeds to form the products:



The reaction of this type is saponification of ester by alkali:



Second-order reactions

If initial concentrations of the reagents A and B, are equal
the rate law may be written in form:

$$v = -\frac{dc_t}{dt} = k_{II}c_t^2$$

where k_{II} – second-order rate constant with units of reciprocal time and concentration, c_t – concentration of the reagents that corresponds to time t .

Integration of this equation gives:

$$\frac{1}{c_t} - \frac{1}{c_0} = k_{II}t$$



$$k_{II} = \frac{1}{t} \left(\frac{1}{c} - \frac{1}{c_0} \right)$$

$$c = \frac{c_0}{1 + c_0 k_{II} t}$$

This equation is the result of:

$$\int_{[A]_0}^{[A]_t} \frac{d[A]}{[A]^2} = -k \int_0^t dt$$

Second-order reactions

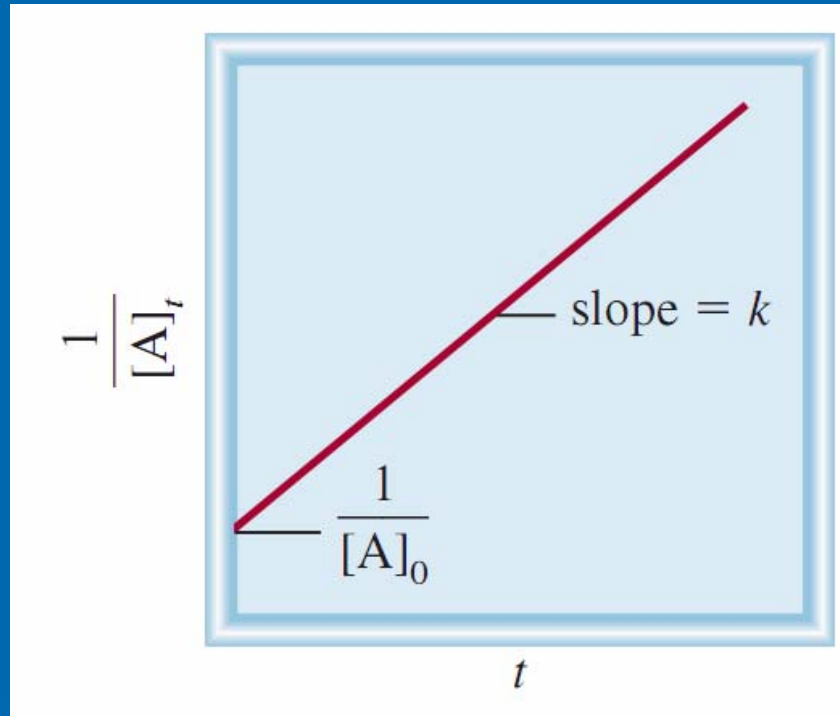


Figure. A plot of $1/[A]_t$ versus t for a second-order reaction. The slope of the line is equal to k .

For half-life: $\tau_{1/2} = \frac{1}{c_0 k_{II}}$

Summary of the Kinetics of Zero-Order, First-Order, and Second-Order Reactions

Order	Rate Law	Concentration-Time Equation	Half-Life
0	Rate = k	$[A]_t = -kt + [A]_0$	$\frac{[A]_0}{2k}$
1	Rate = $k[A]$	$\ln \frac{[A]_t}{[A]_0} = -kt$	$\frac{0.693}{k}$
2 [†]	Rate = $k[A]^2$	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$	$\frac{1}{k[A]_0}$

[†]A \longrightarrow product.

Effect of temperature on the reaction rate

The rate constants of the chemical reactions depend on the nature of reagents and for a given reaction changes with temperature, solvent and the presence of catalyst.

Reaction rates increase with increasing temperature.

The temperature dependence of rate constants

follow to approximate van't Hoff's rule:

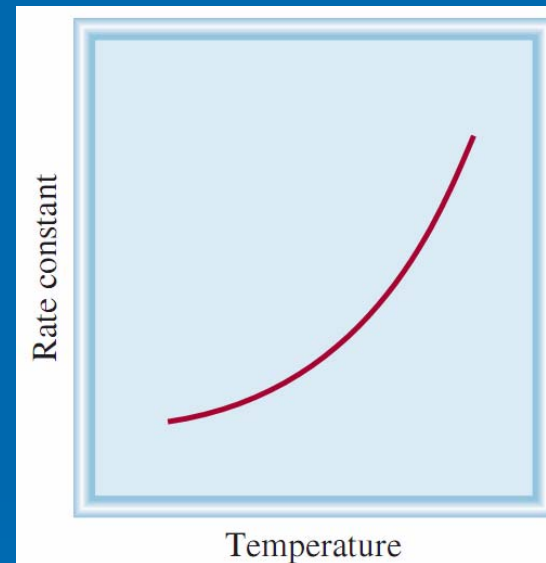
$$k_{T_2} = k_{T_1} \cdot \gamma^{\frac{T_2 - T_1}{10}}$$

where k_{T_2} and k_{T_1} – rate constants at different temperatures T_2 and T_1 ,

γ – temperature coefficient that has a value from 2 to 4.

More precisely the temperature dependence of the rate constants can be expressed by the following equation, known as the Arrhenius equation:

$$\ln k = A - \frac{B}{T}$$



Effect of temperature on the reaction rate

Theoretical interpretation of this dependence gives:

$$k = k_0 e^{-E_a/RT}$$

E_a – activation energy, which is the minimum amount of energy required to initiate a chemical reaction; k_0 – preexponential factor – rate constant that would be observed if $E_a = 0$.

The equation shows that *the rate constant is directly proportional to k_0* ; due to the minus sign associated with the exponent E_a/RT , the rate constant decreases with increasing of activation energy and increases with increasing temperature.

The activation energy

In order to react, the colliding molecules must have a total kinetic energy equal to or greater than the activation energy (E_a), which is the minimum amount of energy required to initiate a chemical reaction. When molecules collide they form an activated complex (also called the transition state), a temporary species formed by the reactant molecules as a result of the collision before they form the product:



where AB^{\ddagger} denotes an activated complex formed by the collision between A and B.

Figure below shows two different potential energy profiles for the reaction:

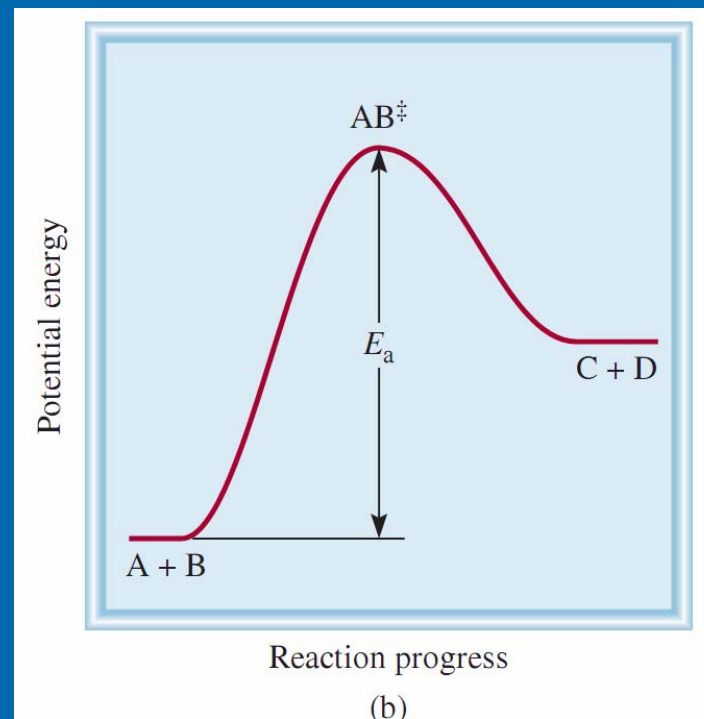
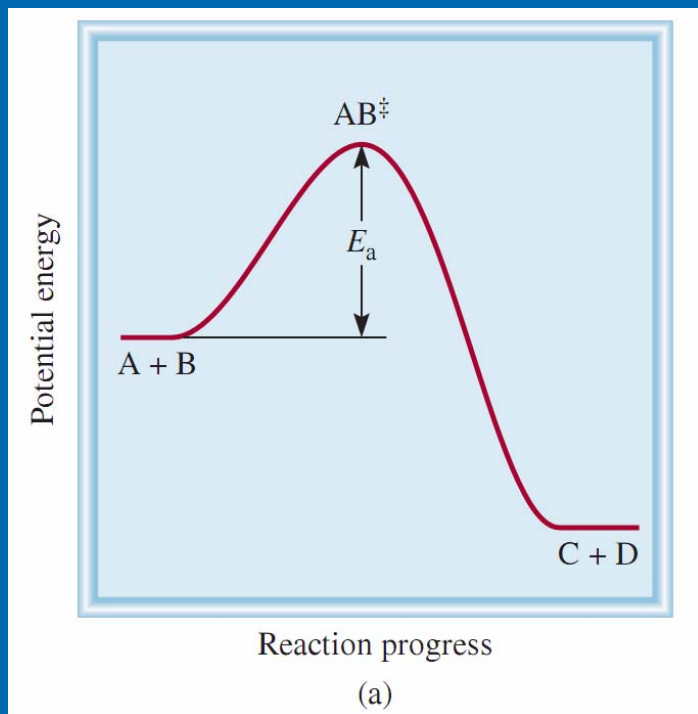


Figure. Potential energy profiles for (a) exothermic and (b) endothermic reactions.

These plots show the change in potential energy as reactants A and B are converted to products C and D. The activated complex (AB^\ddagger) is a highly unstable species with a high potential energy. The activation energy is defined for the forward reaction in both (a) and (b).

Note that the products C and D are more stable than the reactants in (a) and less stable than those in (b).

Effect of temperature on the reaction rate

It follows from the Arrhenius equation the interrelation between rate constants for a reaction proceeding at two different temperatures:

$$\ln \frac{k_{T_2}}{k_{T_1}} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Thus, a plot of $\ln(k_{T_2}/k_{T_1})$ versus $1/T$ gives a straight line whose slope is equal to E_a/R .

Catalysis

A substance that increases the rate of a chemical reaction but does not appear in its stoichiometric equation is called a *catalyst*.

Catalysis can be divided into three classes.

- 1) In homogeneous catalysis all substances involved in the reaction, including the catalyst, occur in the same phase.
- 2) In heterogeneous catalysis the catalyzed reaction occurs at the boundary between two phases (usually on the surface of a solid catalyst).
- 3) Enzyme catalysis is a special case of homogeneous catalysis.

Catalysis

In *catalysis* the increasing of the reaction rate occurs by the lowering the activation energy E_a .

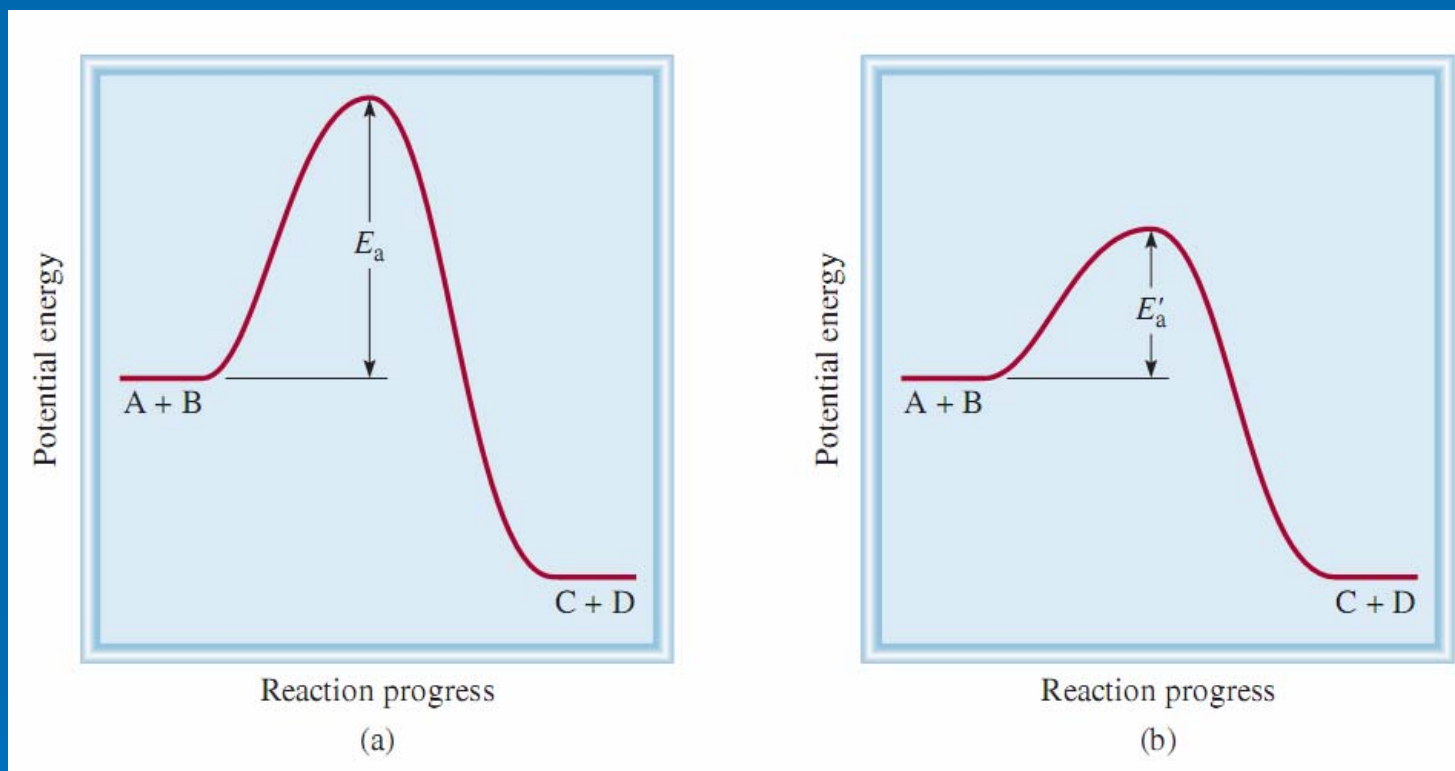


Figure. Comparison of the activation energy barriers of an uncatalyzed reaction (a) and the same reaction with a catalyst (b). The catalyst lowers the energy barrier but does not affect the actual energies of the reactants or products. Although the reactants and products are the same in both cases, the reaction mechanisms and rate laws are different in (a) and (b).

Enzyme catalysis

Enzymes are biological catalysts.

Enzymes can increase the rate of biochemical reactions by factors ranging from 10^6 to 10^{18} and they are also highly specific.

An enzyme acts only on certain molecules, called substrates (that is, reactants), while leaving the rest of the system unaffected.

It has been estimated that an average living cell may contain some 3000 different enzymes, each of them catalyze a specific reaction in which a substrate is converted into the appropriate products.

An enzyme is typically a large protein molecule that contains one or more active sites where interactions with substrates take place.

Enzyme catalysis

In fact, the notion of a rigid enzyme structure that binds only to molecules whose shape exactly matches that of the active site was the basis of an early theory of enzyme catalysis, the so-called lock-and-key theory.

Active sites of the enzymes are structurally compatible with specific substrate molecules, in much the same way as a key fits a particular lock.

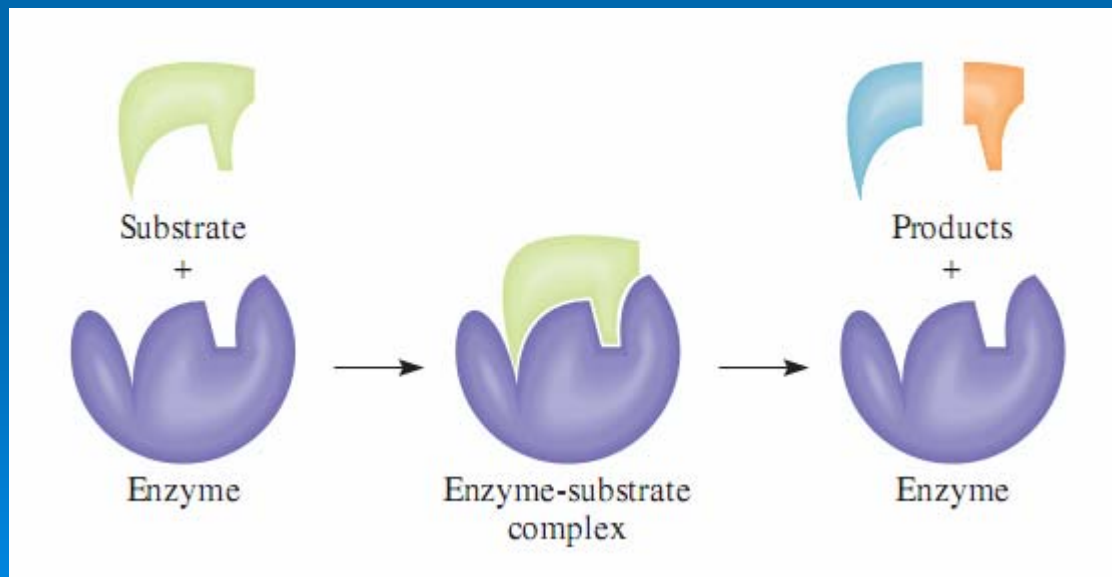
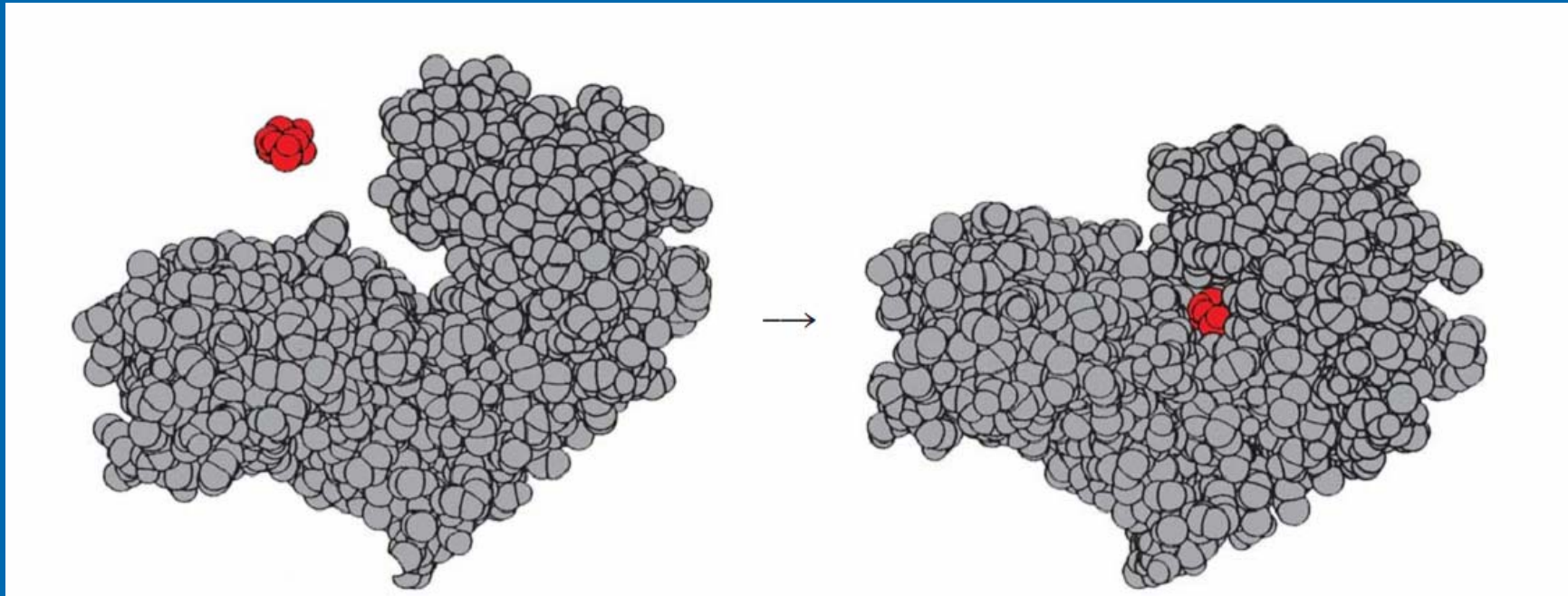


Figure. The lock-and-key model of an enzyme's specificity for substrate molecules.

Enzyme catalysis

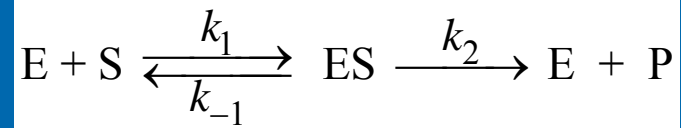
Figure shows a molecular model of an enzyme in action.



The binding of glucose molecule (red) to hexokinase (an enzyme in the metabolic pathway). Note how the region at the active site closes around glucose after binding. Frequently, the geometries of both the substrate and the active site are altered to fit each other.

Enzyme kinetics

The mathematical treatment of enzyme kinetics is quite complex, even when we know the basic steps involved in the reaction. *A simple theory of enzyme action was proposed by L. Michaelis and M. Menten in 1913.* The enzyme catalysed reaction is given by the following scheme:



where E, S, and P represent enzyme, substrate, and product, and ES is the enzyme-substrate intermediate;

k_1 , k_{-1} , k_2 – rate constants for the corresponding elementary steps.

Enzyme kinetics

It is assumed that the formation of intermediate ES and its decomposition back to enzyme and substrate molecules occur rapidly and that the rate-determining step is the formation of product.

Because the rate-limiting step is the second reaction, the rate of the overall reaction is determined by the second step $ES \xrightarrow{k_2} E + P$ and is proportional to the concentration of the complex:

$$v = k_2 c_{ES}$$

Experimentally, enzyme reactions are probed by measuring the initial velocity, when the concentration of the substrate is much greater than the concentration of the enzyme. The use of the initial velocity allows the use of the assumption that changes of the substrate concentration are negligible.

Enzyme kinetics

The initial velocity usually has a linear dependence on the substrate concentration at low substrate concentrations and approaches an asymptotic value denoted by v_{\max} at high concentrations (Figure).

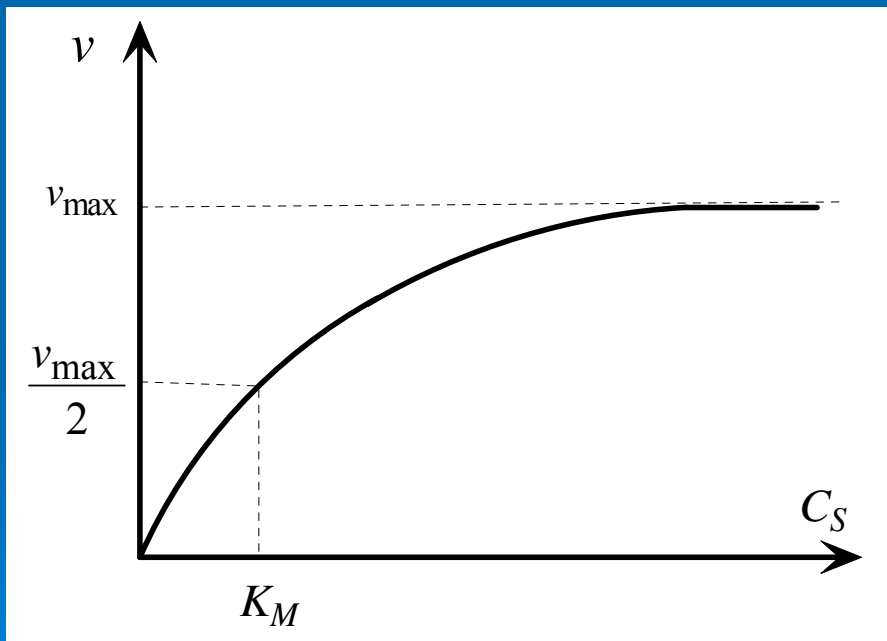


Figure. The effect of substrate concentration on the initial velocity for an enzyme-catalyzed reaction

Enzyme kinetics

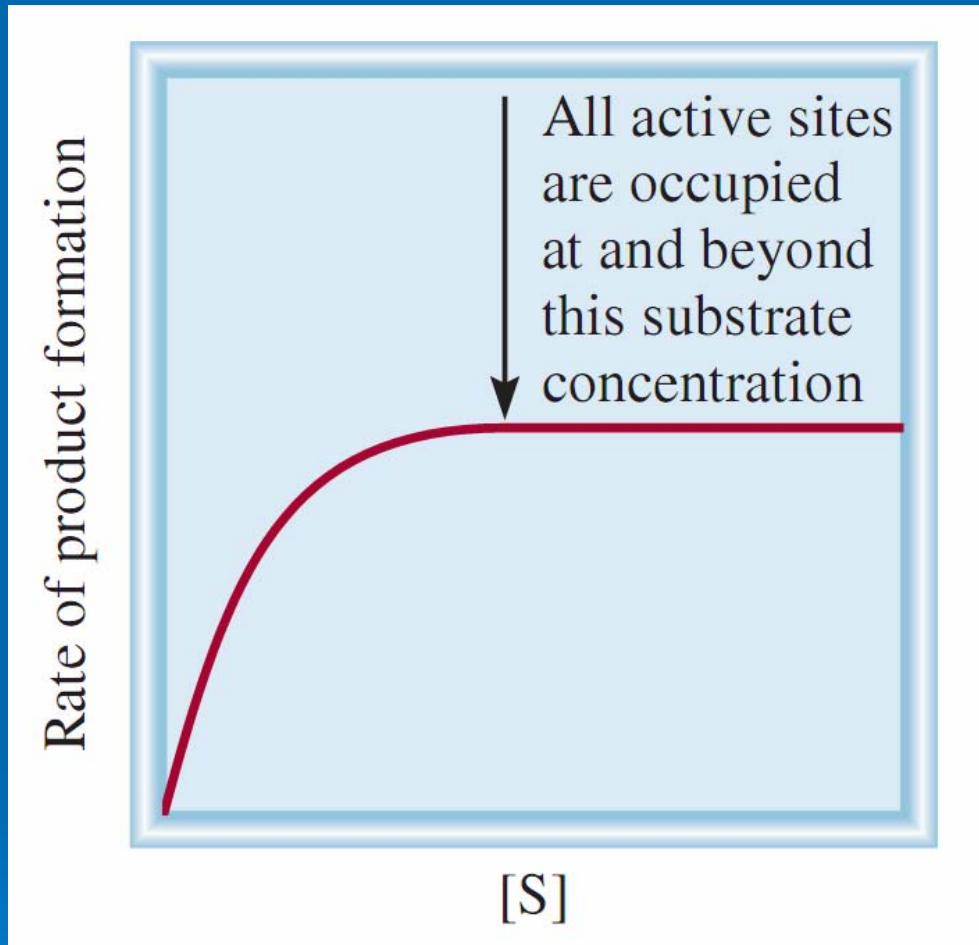


Figure. Plot of the rate of product formation versus substrate concentration in an enzyme-catalyzed reaction

Enzyme kinetics

This dependence of the initial velocity on the substrate concentration can be qualitatively understood in terms of the Michaelis–Menten model.

At any given time, the enzyme is present as a free enzyme and as a complex with the substrate. At low substrate concentrations, most of the enzyme is present in the free form and the rate is proportional to the substrate concentration because the complex formation is favored as the substrate concentration is increased.

The maximum velocity is reached at high substrate concentrations when all of the enzyme is present as the intermediate complex ES. Under these conditions the enzyme is said to be saturated with its substrate, so the changes in the substrate concentration have no effect.

Enzyme kinetics

This dependence can also be described using the *Michaelis–Menten equation*:

$$v = \frac{dc_P}{dt} = \frac{v_{\max} c_{S,0}}{c_{S,0} + K_M}$$

$$v_{\max} = k_2 c_{E,0}$$

$$K_M = (k_{-1} + k_2) / k_1$$

$c_{S,0}$ – initial substrate concentration, $c_{E,0}$ – total enzyme concentration, v_{\max} – maximum velocity, K_M – Michaelis constant.

Enzyme kinetics

For enzyme catalysed reactions following the Michaelis–Menten mechanism, a plot of the reciprocal of the initial velocity against the reciprocal of the substrate concentration, a so-called double-reciprocal plot, produces a straight line:

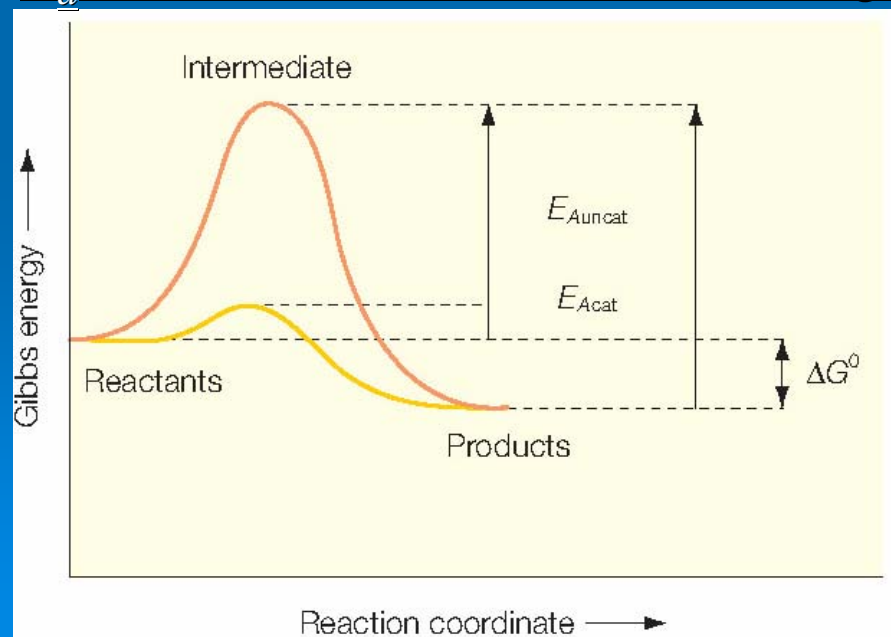
$$\frac{1}{v} = \frac{1}{v_{\max}} + \frac{K_M}{v_{\max}} \cdot \frac{1}{c_{S,0}}$$

This equation is known as Lineweaver–Burk equation.
The y-intercept of the plot gives the value of the maximum velocity and the slope of the plot gives the value of K_M .

Enzyme kinetics

Enzymes accelerate reactions that have a substantial activation energy by modifying the reaction rates.

The Gibbs energy difference between the initial and final states is not altered and the equilibrium is not changed. Rather, enzymes alter the transitional state of the reaction such that the activation energy is significantly decreased. Since the rate is exponentially dependent upon the activation energy, reduction of E_a leads to substantial increasing in the rate.



$$\frac{k_{cat}}{k_{uncat}} = e^{\frac{\Delta E_a}{RT}}$$

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