Chemical kinetics_part 1

Chemical kinetics

The study of a process rate is called ***kinetics or dynamics***. The branch of kinetics that studies the rates and mechanisms of chemical reactions is chemical kinetics. A reacting system is not in equilibrium, so reaction kinetics is not a part of thermodynamics but is a branch of kinetics.

The rate of a chemical reaction

The rate of a chemical reaction is quantitatively characterized by the change in moles number of the reactants proceeds in a unit of time in a unit of the volume

$$V = \frac{1}{v}\left(\pm \frac{1}{a}\frac{dn}{dt}\right)$$

where t is the time, v is the volume, and a is a stoichiometric number of a substance. In most systems the volume is constant and one can replace the number of moles in a certain volume by concentration and define the reaction rate as the change of the concentration of reactant or product proceeds with time.

The rate of a chemical reaction

One may use the finite changes in concentration $C_2 - C_1$ occurring in the time $t_2 - t_1$ and in this way determine the average rate of the reaction for that time period:

$$\bar{V} = \pm \frac{C_2 - C_1}{t_2 - t_1}$$

The rate of a chemical reaction

Usually the rate of a chemical reaction does not remain constant, but changes with time, so it's more convenient the change in concentration to refer to infinitesimal time and thereby the instantaneous rate of reaction at the given moment can be determined as a derivative of the infinitesimal change of a concentration with respect to infinitesimal time change:

$$V = \pm \frac{dC}{dt}$$

In the following we shall consider only **instantaneous reaction rate**.

The reaction rates in all cases depend on the three basic factors :

- Concentration
- Temperature
- Catalysts

The reaction rate dependence on the reactants concentration

According the rate law, the rate of a given reaction at constant temperature is proportional to the product of the reactant concentrations, each concentrations being raised to a power, according the coefficient in front of the formula of the substance in the equation of the reaction.

A → B

$$-\frac{d[A]}{dt}=k[A]$$

and for reaction as **mA + nB → pC** is:

$$-\frac{d[A]}{dt}=k[A]^{m}[B]^{n}$$

Molecularity of a reaction

The number of molecules that simultaneously react in the elementary reaction determines the **molecularity of a reaction**. Molecularity is defined only for elementary step and should not be used to describe overall reactions that consist of more than one elementary step.

- The elementary reaction A → products is unimolecular ($CaCO_3 \rightarrow CaO + CO_2$);
- A + B → products, and 2A → products are bimolecular ($H_2 + I_2 \leftrightarrow 2HI$);
- A + B + C → products, 2A + B → products, and 3A → products are trimolecular ($2NO + H_2 \rightarrow N_2O + H_2O$).
- No elementary reactions involving more than three molecules are known

Overall reaction order

For a general reaction of the type

$$\mathbf{aA + bB = cC + dD}$$

the rate law takes the form

$$\mathbf{k \bullet [A\,]^{a} \bullet [B\,]^{b} = V.}$$

The sum of the powers to which all reactant concentrations appearing in the rate law are raised is called the **overall reaction order**. The overall order is the sum of the individual orders and this reaction therefore is the second-order. It is most important to distinguish molecularity according to the order.

The order of a reaction

Reactions are classified as

- the first-order
- the second-order
- the third order.

There are also **zero-order** and **fractional order** reactions. The above reaction is the first-order in A as well as in B (if a=1 and b=1). Reaction order is an empirical quantity, obtained from the experimental rate law, and related to all the chemical processes (with all steps).

The order of a reaction

The order of a reaction must be determined by experiment, it cannot be deduced from the overall balanced equation. A useful indication of the rate of chemical reactions is also half-life of a reaction, τ1/2, which is equal to the time required for the concentration of a reactant to decrease to half of its initial concentration.

Zero-order reactions

Some (very rarely) reactions obey a zero-order rate law, and therefore have a rate independent on the concentration of the reactant:

$$-\frac{dC}{dt} = k_o$$

Only a few heterogeneous reactions can have rate law, which is zero-order overall. Thus, the catalytic decomposition of phosphine (PH_3) in hot tungsten at high pressures has the rate law of zero-order. The PH_3 decomposes at a constant rate until it has almost entirely disappeared.

First-order reactions

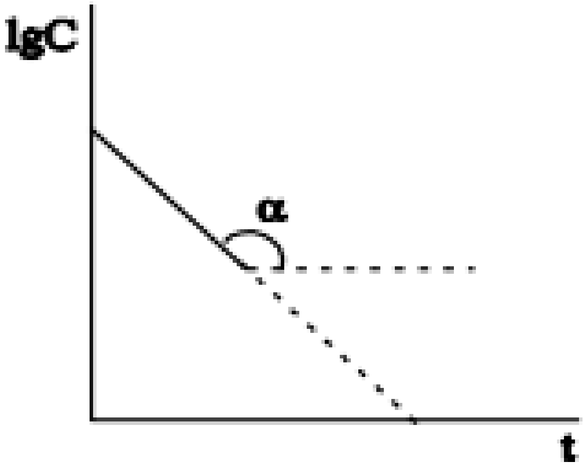
For the first-order reaction as **A → B** the rate dependence on the concentration has the following form:

$$-\frac{dC}{dt} = k_1 C$$

$$-\frac{dC}{C} = k_1 dt$$

$$-\ln C = k_1 t + const$$

$$k_1 = \frac{1}{t}\ln\frac{C_o}{C}$$



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

Second-order reactions

The rate of such reaction depends on reactant concentration raised to the second power or on the concentrations of two different reactants, each raised to the first power:

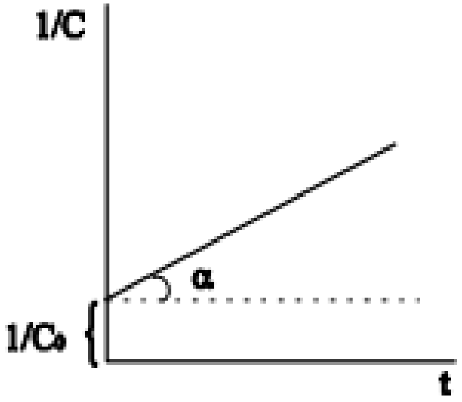
$$-\frac{dC}{dt} = k_2 C_1 C_2$$

In case $C_1 = C_2$

$$-\frac{dC}{dt} = k_2 C^2 \quad \text{or} \quad -\frac{dC}{C^2} = k_2 dt$$

$$\frac{1}{C} = k_2 t + \frac{1}{C_o} \quad \text{from which}$$

$$k_2 = \frac{1}{t}\frac{C_o - C}{C_o C}$$



In case $C_1 \neq C_2$

$$\frac{dx}{dt} = k_2(a-x)(b-x) \quad or \quad \frac{dx}{(a-x)(b-x)} = k_2 dt$$

$$k_2 = \frac{1}{t}\frac{1}{\left(C_{0(1)} - C_{0(2)}\right)} \ln \frac{C_{0(2)} C_1}{C_{0(1)} C_2}$$

$$\tau_{1/2} = \frac{1}{k_2}\frac{C_o - 1/2C_o}{C_o 1/2C_o} = \frac{1}{k_2 C_o}$$

Determination of the reaction’s order

- The method of substitution. For a certain reaction the concentrations of the reactants is determined during the reaction (in certain intervals) and then their meanings in the kinetic equations (separately in the first-, second-, and third order) are put. Where we obtain the same meanings of the rate constant the reaction is that order.
- The graphic method. Having carried out the same determination as in the first method one can constructs a diagram of lnC, 1/C, and $1/C_2$ versus t. Where dependence is shown by straight line, the order is the same.
- 3. The half-life determination.

Complex reactions

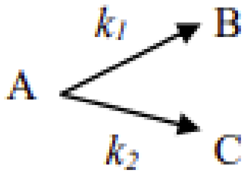
Complex reactions are those consisting of one or more simple reactions related to one another in one way or another depending on the nature of their interrelation and on the ratio of their reaction rates. Species formed in an earlier elementary step and consumed in a later elementary step are called intermediates. They appear in the mechanism of the reaction (that is, the elementary steps) but not in the overall balanced equation.

Complex reactions

- Simultaneous,
- Consecutive
- Conjugate
- Reversible
- Chain
- Photochemical reactions

Simultaneous (parallel) reactions

Simultaneous reactions are those in which a species can react in different ways to give a variety of products.



$$V_1 = \frac{dx_1}{dt} = k_1(a-x) \quad and \quad V_2 = \frac{dx_2}{dt} = k_2(a-x)$$

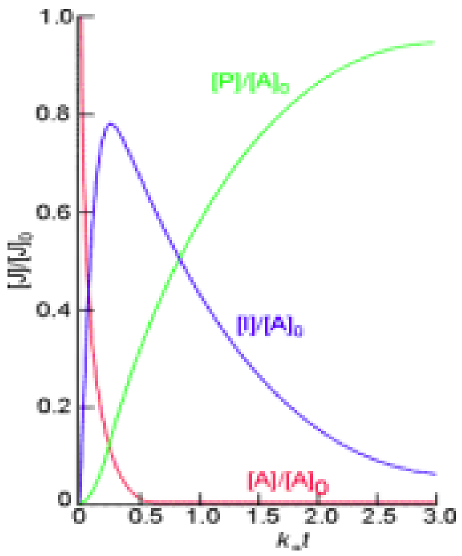
$$V = V_1 + V_2 = k_1(a-x) + k_2(a-x) = (k_1 + k_2)(a-x) \quad \text{or}$$

$$\frac{dx}{dt} = (k_1 + k_2)(a-x)$$

$$(k_1 + k_2) = \frac{1}{t}\ln\frac{a}{a-x}$$

Consecutive reactions

Frequently the product of one reaction becomes the reactant in the subsequent reaction. This is true in multistep reaction mechanisms. In the scheme above, I is formed from A and is called an intermediate product, which decays to P.



$$A \xrightarrow{k_1} I \xrightarrow{k_2} P$$

$$-\frac{dA}{dt} = k_A[A]$$

$$\frac{d[B]}{dt} = k_A[A] - k_B[B]$$

$$\frac{d[C]}{dt} = k_B[B]$$

Coupled reactions

In this case two chemical reactions occurred in a system when a chemical species takes part in both reactions. It could be defined as

$$A + B \rightarrow M$$

$$A + C \rightarrow N$$

Those reactions take place only together, since one reaction influences the equilibrium position of the second reaction, i.e., is induced by the other. This phenomenon is called chemical induction. In such case substance **C** is an **inductor** of the first reaction, **A**, common to both reactions, is an **actor**, and substance **B** is an **acceptor**.

The reason of coupled reactions is that the second reaction occurs with absorption of energy, which arrived from the first reaction.

Reversible reactions

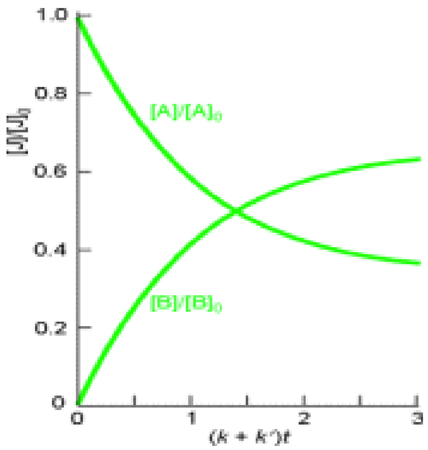
They are those taking place both in forward and reverse directions:

A ↔ B

The rate of a reversible reaction equals to the difference between the rates of the forward and reverse reactions.

$$\frac{dx}{dt} = k_1(a-x) - k_2(b+x) = k_1a - k_1x - k_2b - k_2x = k_1a - k_2b - x(k_1 - k_2)$$

$$\frac{dx}{dt} = \left(\frac{k_1a - k_2b}{k_1 + k_2} - x\right)(k_1 + k_2) \qquad \frac{k_1a - k_2b}{k_1 + k_2} = x_\infty$$

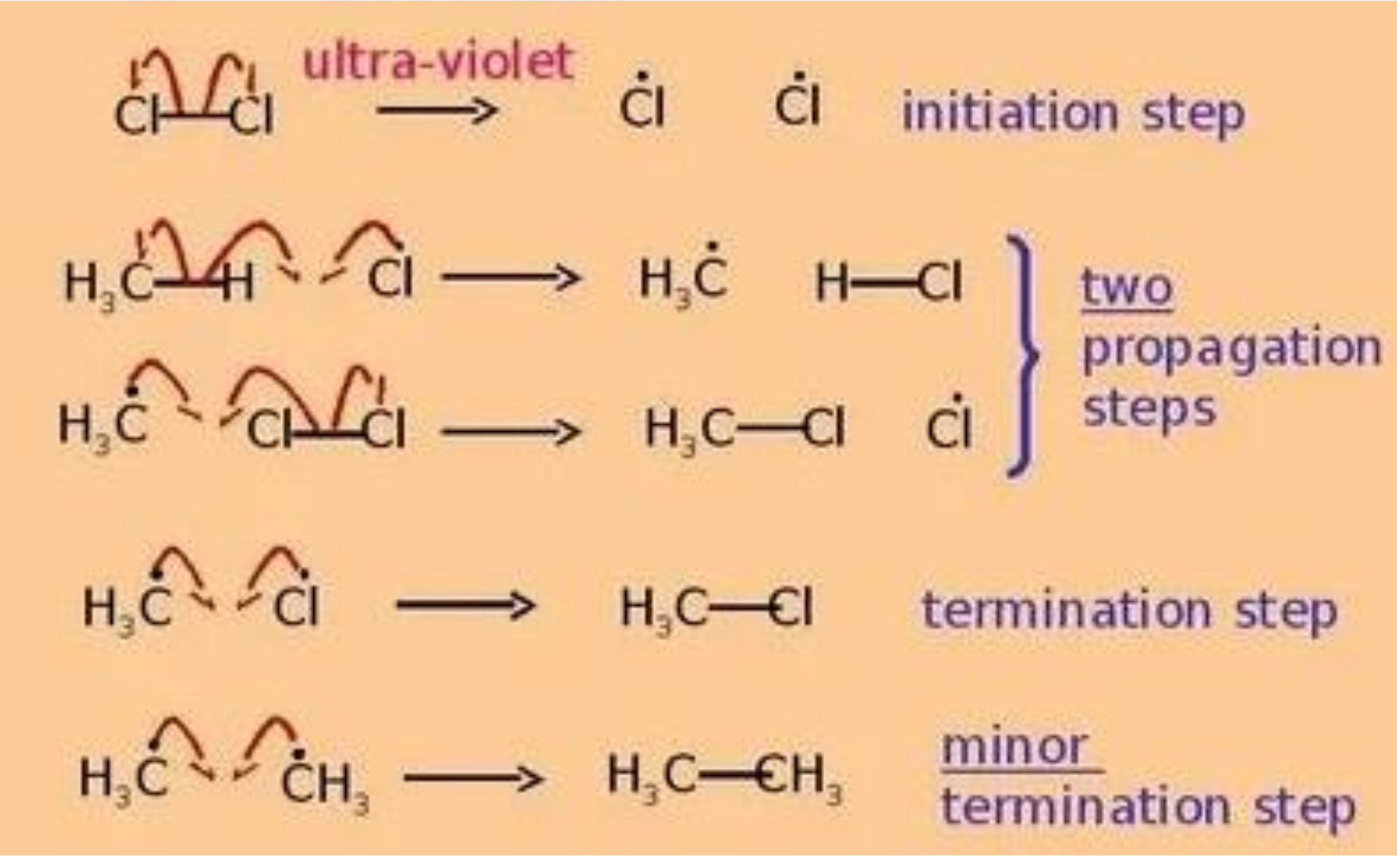


where x is the equilibrium concentration of the reactant

$$\frac{dx}{dt} = (x_\infty - x)(k_1 + k_2)$$

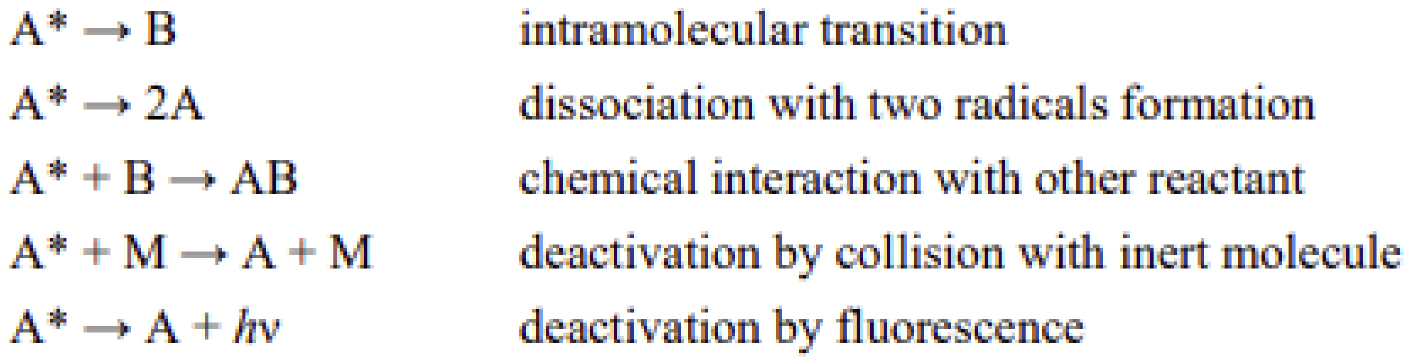
$$(k_1 + k_2) = \frac{1}{t}\ln\frac{x_\infty}{x_\infty - x}$$

Chain reactions



Photochemical reactions

- These reactions are produced by light. Absorption of a photon of light may raise a molecule to an excited electronic state, where it will be more likely to react than in the ground electronic state. The number of photons absorbed equals the number of molecules making a transition to an excited electronic state.
- The initial step of a photochemical reaction is $A + h\nu \rightarrow A^*$ in which the absorption of a photon occurs by a molecule to raise it to an excited electronic state. Then this excited molecule could take part in some primary reactions:

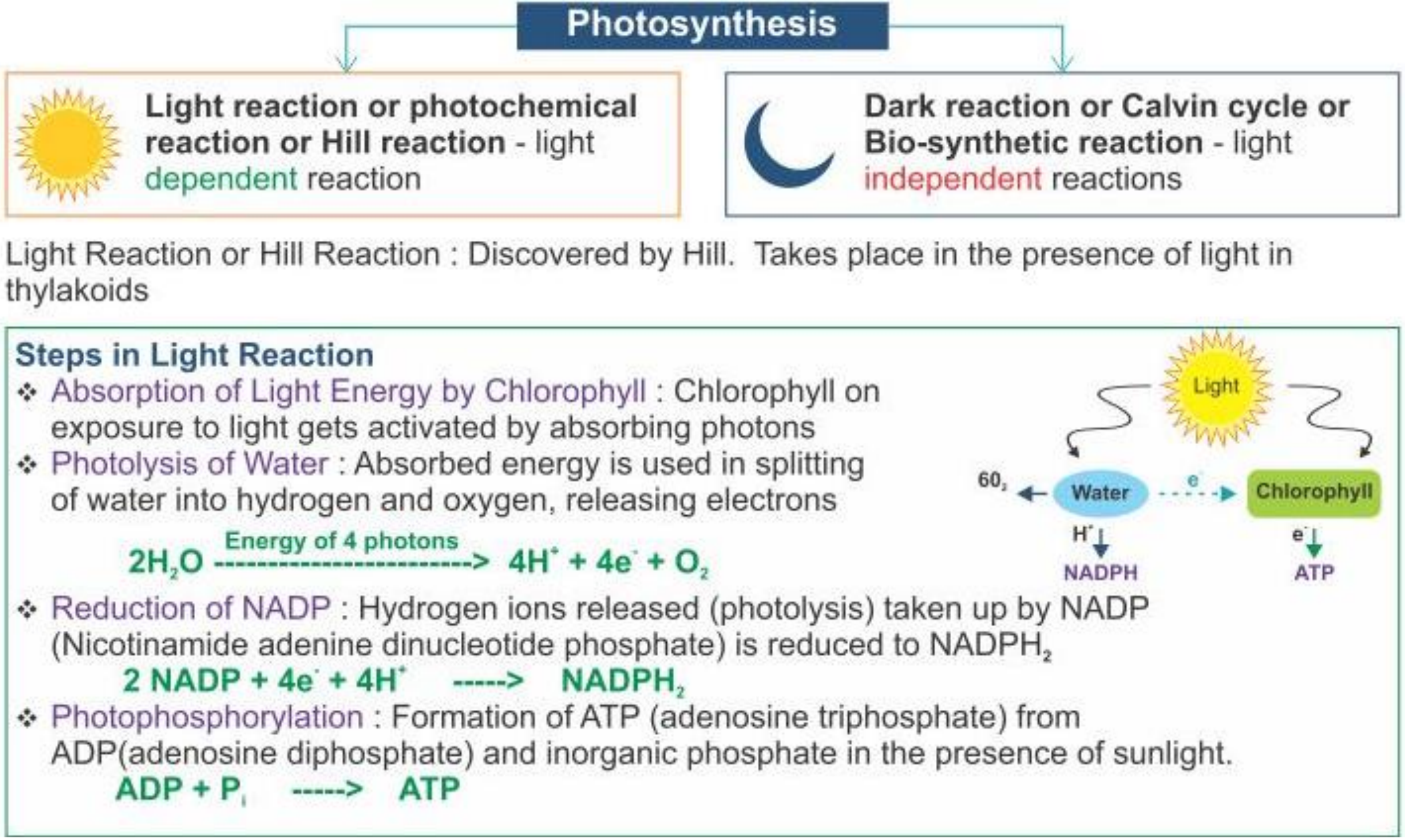


$A^* \rightarrow B$	intramolecular transition
$A^* \rightarrow 2A$	dissociation with two radicals formation
$A^* + B \rightarrow AB$	chemical interaction with other reactant
$A^* + M \rightarrow A + M$	deactivation by collision with inert molecule
$A^* \rightarrow A + h\nu$	deactivation by fluorescence

Photochemical reactions

- The quantitative characteristic of these primary reactions are the primary quantum yield which is equal to the number of excited moles, taking part in the primary reactions, divided by the number of absorbed photons. It couldn’t be more then 1 (varies between 0 and 1), because the number of excited moles couldn’t be more then the number of absorbed photons.
- The products generated in the primary reactions, could take part in the secondary (following) reactions. For example, the radicals formed from dissociation of excited molecule could take part in the chain reaction. When the primary and secondary reactions occur, the completely photochemical process is characterized by overall quantum yield, which is equal to the number of product formed in the secondary reaction divided by the number of absorbed photons.

Photochemical reactions



Photosynthesis

Light reaction or photochemical reaction or Hill reaction - light dependent reaction

Dark reaction or Calvin cycle or Bio-synthetic reaction - light independent reactions

Light Reaction or Hill Reaction : Discovered by Hill. Takes place in the presence of light in thylakoids

Steps in Light Reaction

- Absorption of Light Energy by Chlorophyll : Chlorophyll on exposure to light gets activated by absorbing photons
- Photolysis of Water : Absorbed energy is used in splitting of water into hydrogen and oxygen, releasing electrons

$$2H_2O \xrightarrow{\text{Energy of 4 photons}} 4H^+ + 4e^- + O_2$$

- Reduction of NADP : Hydrogen ions released (photolysis) taken up by NADP (Nicotinamide adenine dinucleotide phosphate) is reduced to $NADPH_2$

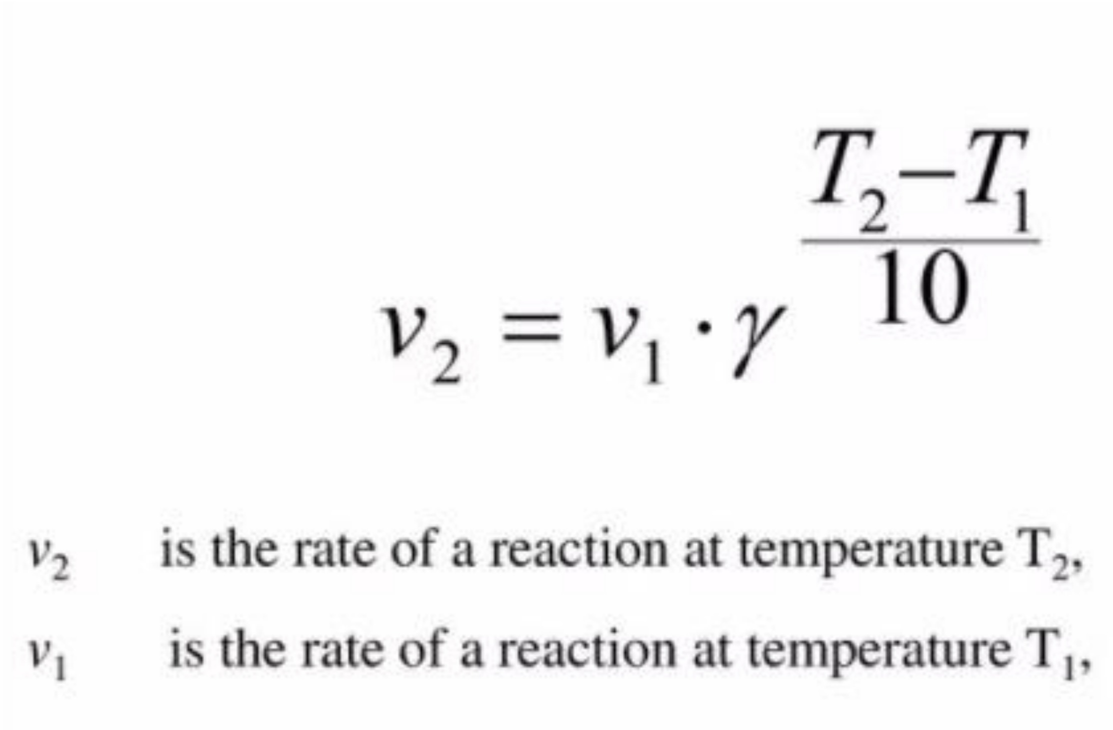
$$2\ NADP + 4e^- + 4H^+ \longrightarrow NADPH_2$$

- Photophosphorylation : Formation of ATP (adenosine triphosphate) from ADP(adenosine diphosphate) and inorganic phosphate in the presence of sunlight.

$$ADP + P_i \longrightarrow ATP$$

The reaction rate dependence on temperature

The rate of most chemical reaction increases with an increase in temperature. There is a rough **rule of van’t Hoff** according to which the rate of a reaction increases about two to fourfold with each 10^{o}C rise in temperature.

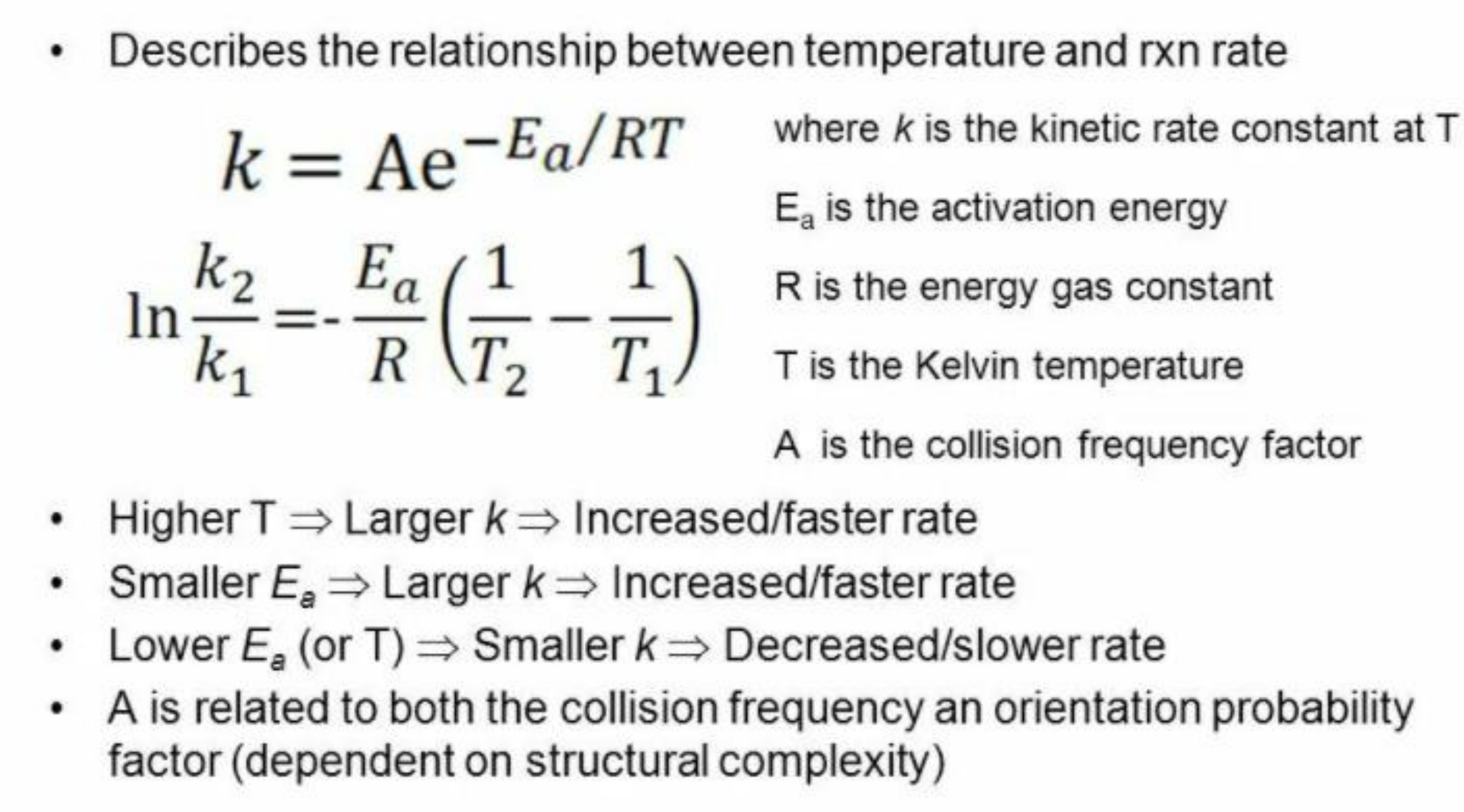


$$v_2 = v_1 \cdot \gamma^{\frac{T_2 - T_1}{10}}$$

v_2 is the rate of a reaction at temperature T_2,

v_1 is the rate of a reaction at temperature T_1,

Arrhenius equation



- Describes the relationship between temperature and rxn rate

$$k = \mathrm{A}e^{-E_a/RT}$$

$$\ln\frac{k_2}{k_1} =- \frac{E_a}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

where *k* is the kinetic rate constant at T

E_a is the activation energy

R is the energy gas constant

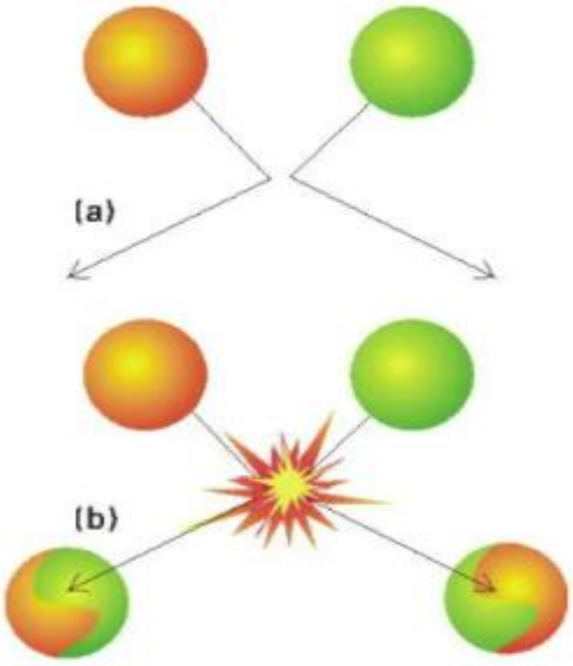
T is the Kelvin temperature

A is the collision frequency factor

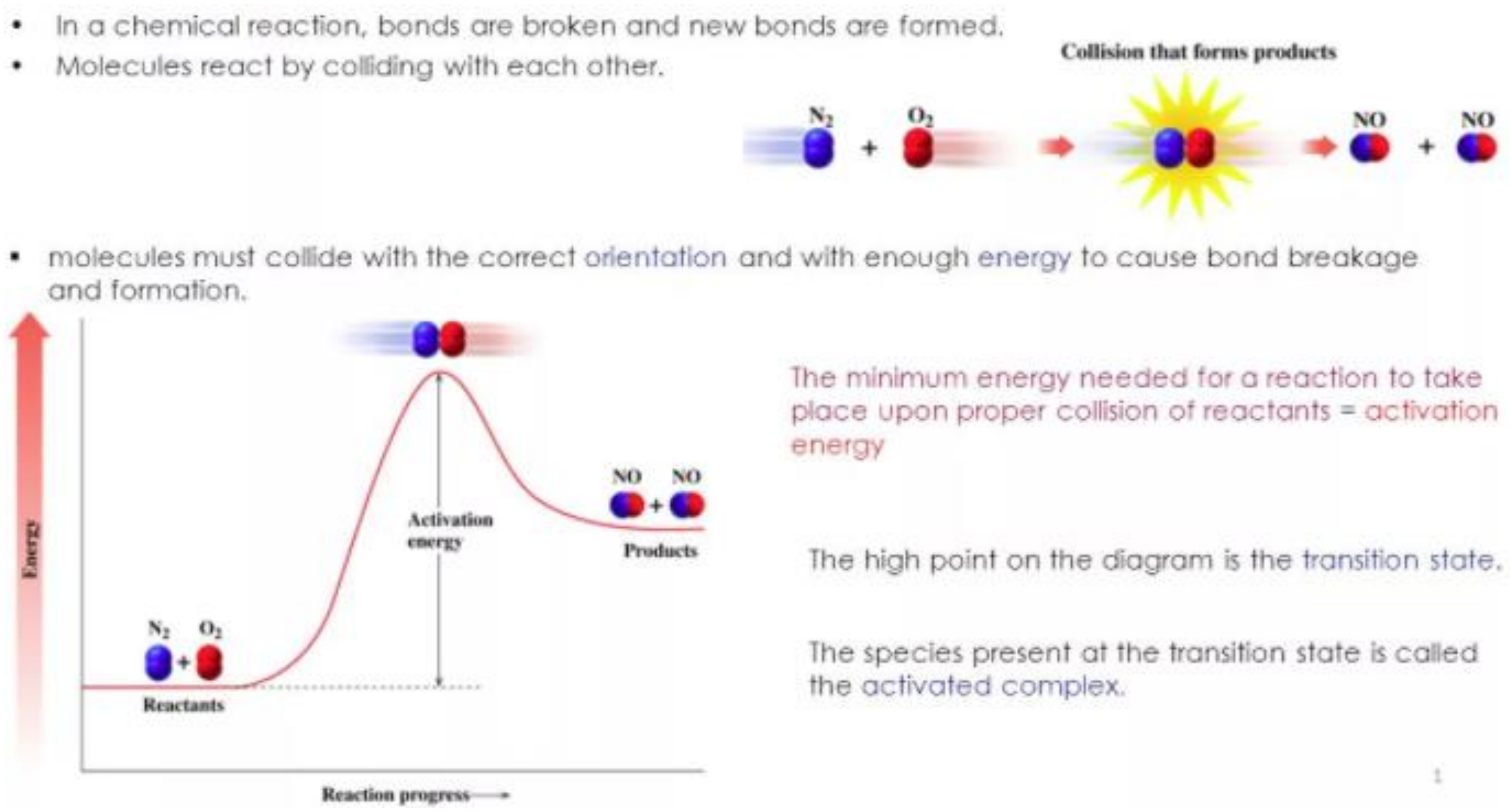
- Higher T ⇒ Larger *k* ⇒ Increased/faster rate
- Smaller E_a ⇒ Larger *k* ⇒ Increased/faster rate
- Lower E_a (or T) ⇒ Smaller *k* ⇒ Decreased/slower rate
- A is related to both the collision frequency an orientation probability factor (dependent on structural complexity)

The theory of active collisions

- In order for a reaction to occur, the particles of the reactant must have enough energy, and must collide at the correct angles (proper orientation).
- The collision theory explains the factors that affect the rate of a reaction.
- The greater the rate of *effective collisions*, the greater the reaction rate is
- The collision theory explains the factors that affect the rate of reaction



The theory of active collisions



- In a chemical reaction, bonds are broken and new bonds are formed.
- Molecules react by colliding with each other.

- molecules must collide with the correct orientation and with enough energy to cause bond breakage and formation.

The minimum energy needed for a reaction to take place upon proper collision of reactants = activation energy

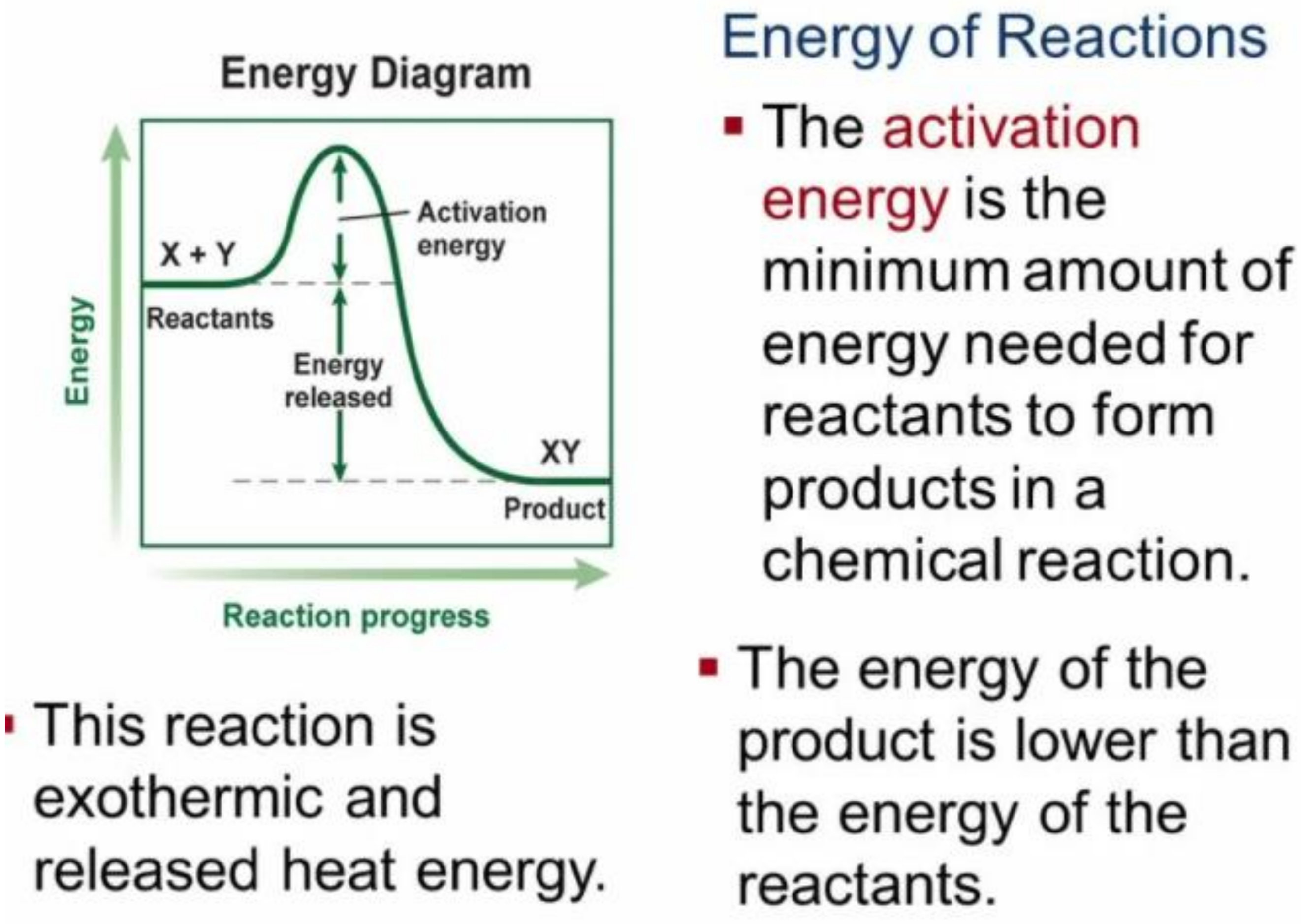
The high point on the diagram is the transition state.

The species present at the transition state is called the activated complex.

The activation energy

The activation energy is the excess energy that a molecule must have at the time of collision to be able to enter into the given chemical reaction. The role of activation energy for reaction rate is very important. We can think of activation energy as a barrier that prevents less energetic molecules from reacting. If elements of a physicochemical system have sufficient energy to overcome the barrier, reaction will occur. Normally, only a small fraction of the colliding molecules have enough kinetic energy to exceed the activation energy. These molecules can therefore take part in the reaction. The increase in the rate with temperature can now be explained: since more high-energy molecules are present at the higher temperature, the rate of product formation is also greater at the higher temperature

The activation energy



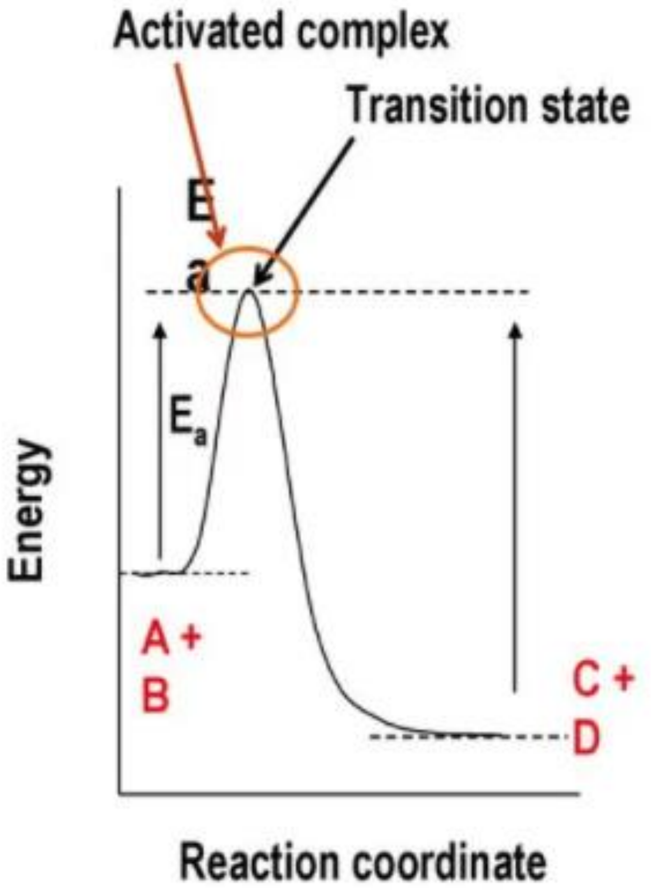
Energy of Reactions

- The activation energy is the minimum amount of energy needed for reactants to form products in a chemical reaction.
- The energy of the product is lower than the energy of the reactants.

- This reaction is exothermic and released heat energy.

Activated complex theory

- In a reaction step as the reactant molecules A and B come together they distort and begin to share, exchange or discard atoms.
- They form a loose structure $AB^{\ddagger}$ of high potential energy called the activated complex that is poised to pass on to products or collapse back to reactants C + D.
- The peak energy occurs at the transition state. The energy difference from the ground state is the activation energy E_a of the reaction step.
- The potential energy falls as the atoms rearrange in the cluster and finally reaches the value for the products



Pharmacokinetics

- Clearance
- Half-lives and Residence Times
- Distribution Volumes
- Absorption & Bioavailability Measures
- Compartmental Modeling

Pharmacokinetics