

# Lecture 11

## Theories of Reaction Rates

There are two basic theories: Collision theory and activated complex theory (transition state theory).

Simplest is the collision theory – accounts for gas phase reactions

### Reactions in solution

1. Diffusion controlled - Diffusion equation can account for rates
2. Activation controlled – Activated complex and thermodynamics are involved

Rates in terms of potential energy surfaces is more advanced

Both (collision and activated complex) theories say that:

Molecular encounters make them excited making reactions. Excitation can be through collisions (collision theory) or the energy can be taken from surroundings (activated complex, in solutions).

## Collision Theory



$$\text{Rate} = k_2[A][B]$$

$k_2$  to involve 1. Rate of encounter 2. Energy 3. Orientation

$k_2 \propto$  rate of collisions  $\times$  energy requirement  $\times$  steric requirement

### Assumptions

- Molecules are hard spheres
- For reactions to occur, molecules must collide
- Reactions occur only if molecules are energetic
- Collisions should transfer certain minimum energy

# Number of Collisions

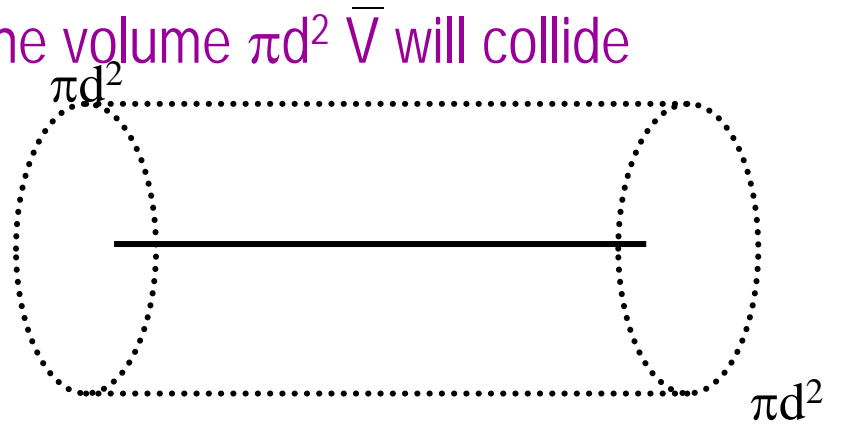
Number of molecules  $N$  in volume  $V$

Number of collisions of one molecule per second

is the collision frequency,  $z$

Average velocity of the molecule =  $\bar{V}$  m/s, the molecule collides every molecule within the distance,  $d$

Therefore, all the molecules within the volume  $\pi d^2 \bar{V}$  will collide with the test molecule per second.



## Encounter of A with B

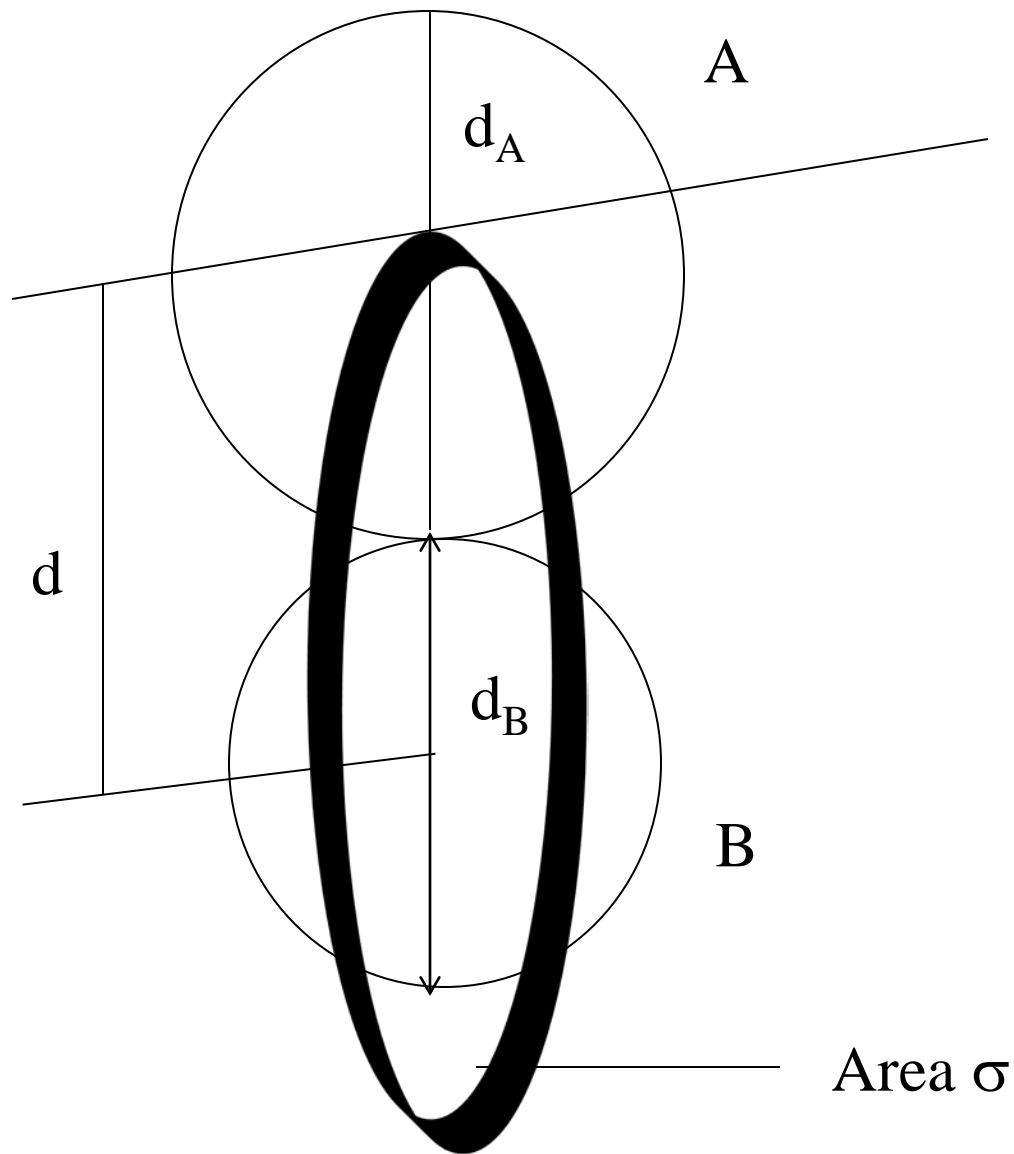
$\pi d^2$  - collision cross section

This is the area within which centre of A has to enter around the target molecule B for collisions to occur.

This is the volume swept by the molecule/sec

The solid line represents the average velocity of the molecule.

Area  $\pi d^2$ , is the area covered by the dotted lines.



# Rate of collisions

$N/V \rightarrow$  Number of molecules in unit volume

Number of collisions of one molecule per second,

$$z = (N/V) (\pi d^2) (\bar{V} s^{-1})$$

Relative speed  $\bar{V} = (8kT/\pi m)^{1/2}$

$m$  is replaced by  $\mu = (m_1 m_2 / m_1 + m_2)$ , reduced mass  
identical particles,  $\mu = m/2$

$$\bar{V} = (8kT/\pi m/2)^{1/2}$$

$$z_{11} = \sqrt{2} N/V \pi d^2 (8kT/\pi m)^{1/2} s^{-1}$$

$$z_{11} = \sqrt{2} N/V \pi d^2 (8RT/\pi M)^{1/2} s^{-1}$$

$$z_{11} = \sqrt{2} N/V \sigma (8kT/\pi m)^{1/2} s^{-1}$$

$\pi d^2 = \sigma \leftarrow$  Collision cross – section

## Collision number

Number of bimolecular collisions in 1 m<sup>3</sup> per sec,

$$Z_{11} \text{ (s}^{-1}\text{m}^{-3}\text{)} = \frac{1}{2} (N/V) z_{11}$$

$\frac{1}{2}$  not to count collisions twice

$$= \frac{1}{\sqrt{2}} (N/V)^2 \pi d^2 \bar{V}$$

for different molecules,

$$Z_{12} = (N_A/V) (N_B/V) \pi (r_1 + r_2)^2 (8kT/\pi \mu)^{1/2}$$

Collision Number is very large  $\sim 10^{34}$  for N<sub>2</sub> at STP

# Rate constant



$$d[P]/dt = k[A][B]$$

if all collisions lead to reactions,  $d[P]/dt = Z_{AB}/N_{av}$

**$N_{av}$  → Avogadro number**

(Since rate is in moles/unit volume and not in molecules/unit volume)

We know,

$$d[P]/dt = k (N_A/V.N_{av}) (N_B/V.N_{av}) \quad \dots 1$$

$$d[P]/dt = Z_{AB}/N_{av} = [1/N_{av}] (N_A/V) (N_B/V) \pi d^2 \bar{V} \quad \dots 2$$

Compare 1 and 2

$$k/N_{av} = \pi d^2 \bar{V}$$

$$k = N_{av} \pi d^2 \bar{V} \text{ mol}^{-1} \text{m}^3 \text{s}^{-1}$$

The actual value is very very small !!  $\text{H}_2 + \text{I}_2 \rightarrow \text{HI}$ , Calculated

$$k = 7.5 \times 10^8 \text{ m}^3 \text{ mol}^{-1} \text{ s}^{-1} = 7.5 \times 10^{11} \text{ L mol}^{-1} \text{ s}^{-1}$$

Observed →  $1.15 \times 10^{-4} \text{ L mol}^{-1} \text{ s}^{-1}$  !!!



# Energy Considerations

We expect the rate constant for a bimolecular reaction to depend on both frequency and energy of the collisions.

The above considerations do not take into account the role of energy:

The reaction of molecular pairs with energy above  $E_a$  is given by  $\exp^{-E_a/RT}$

These are the fruitful collisions leading to products.

$$Z_{AB}(\text{effective}) = Z_{AB} \times \exp^{-E_a/RT}$$

The rate constant for a bimolecular reactions is

$$k' = N_{av} \pi d_{AB}^2 (8kT/\pi m)^{1/2} \exp^{-E_a/RT}$$

This equation is similar to the Arrhenius law except for  $\sqrt{T}$  term in  $\bar{V}_{AB}$  of pre-experimental factor.

Kinetic data over a wide range of temperature range show that the pre-experimental factor depends on  $\sqrt{T}$  (by about 2%) as predicted.

$$\therefore k = A' \sqrt{T} \exp^{-E_a/RT}$$

$$A' = N_{AV} \pi d_{AB}^2 (8k/\pi m)^{1/2} \text{ mol}^{-1} \text{ m}^3 \text{ sec}^{-1}$$

# Steric requirement

Collision energy is not the only criterion.

Relative orientation is important.

Disagreement in  $A$  values between theory and experiment

Explained in terms of a steric factor,  $P$

$$k_{11} = P Z_{11} e^{-E_a/RT}$$

$$k_{12} = P Z_{12} e^{-E_a/RT}$$

$$P = (A_{\text{expt}}/A_{\text{theory}})$$

The steric factor is difficult to evaluate

$$k = P\sigma(8kT/\pi\mu)^{1/2} N_{av}e^{-Ea/RT}$$

The effective  $\sigma$  can be considered as  $P\sigma$  which is taken as the reactive cross section  $\sigma^*$ .

$$k = P\sigma(8kT/\pi\mu)^{1/2} N_{av}e^{-Ea/RT}$$

$$k = Ae^{-Ea/RT}$$

$$A = P\sigma(8kT/\pi\mu)^{1/2} N_{av}$$

This kind of evaluation shows that  $P$  is very small, of the order of  $10^{-6}$

# Reactions in solution

Cage effect

Activation energy, Encounter pair



$$d[AB]/dt = k_d[A][B] - k_d'[AB] - k_a[AB] \sim 0$$

$$[AB] \sim k_d[A][B]/[k_a + k_d']$$

$$d[P]/dt \sim k_a[AB] = k_a k_d[A][B]/[k_a + k_d'] = k_2[A][B]$$

$$k_d' \ll k_a, \quad d[P]/dt \sim k_d[A][B] \quad \text{diffusion controlled}$$

$$k_a \ll k_d', \quad d[P]/dt \sim k_a k_d/k_d'[A][B] = k_a K[A][B] \quad \text{activation controlled}$$

# Transition State Theory

Other names: Activated complex theory and Absolute rate theory

Drawbacks of collision theory:

Difficult to calculate the steric factor from molecular geometry for complex molecules.

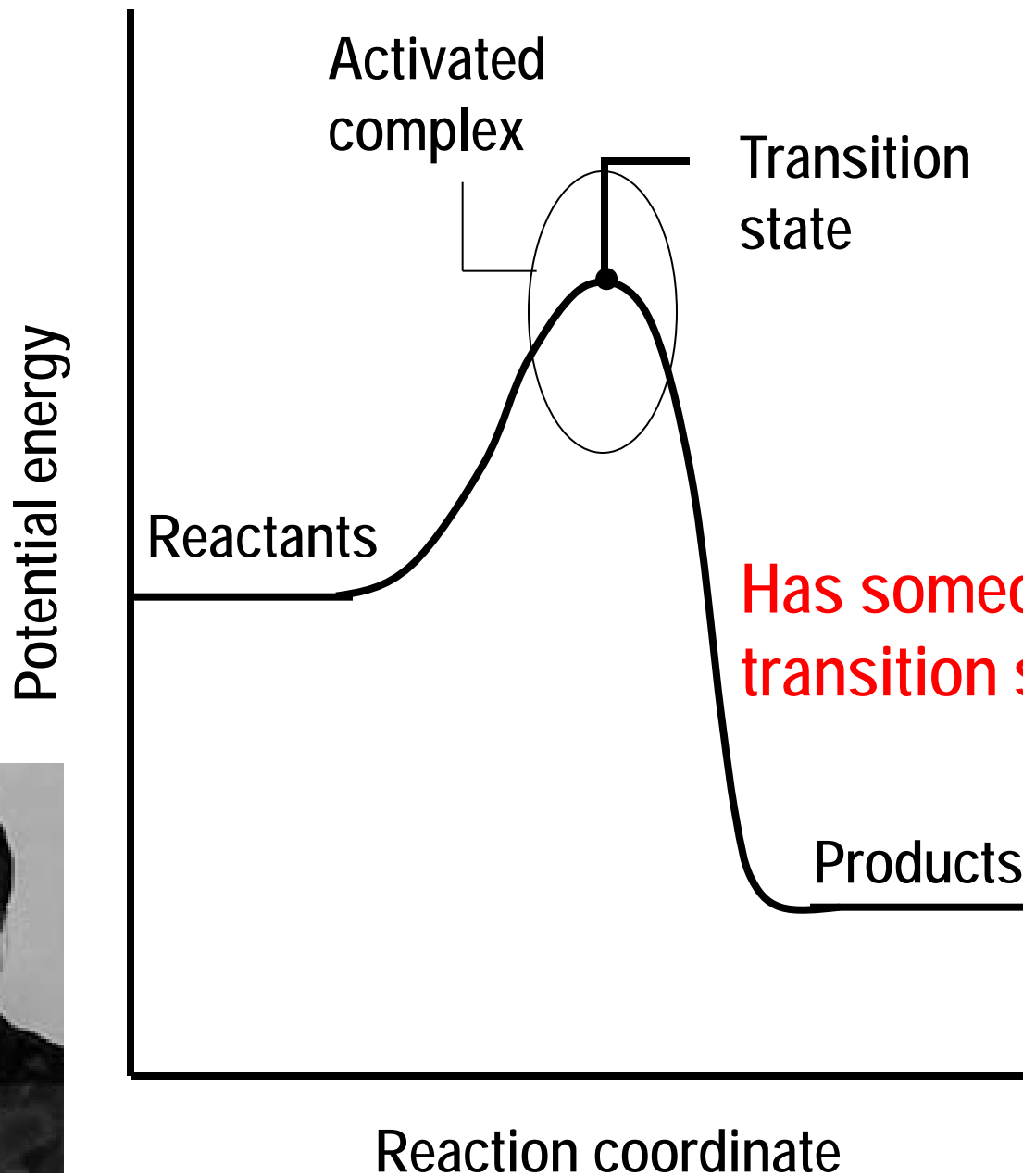
The theory is applicable essentially to gaseous reactions

Consider  $A + B \rightarrow P$  or  $A + BC = AB + C$



A and B form an activated complex and are in equilibrium with it.

The reactions proceed through an activated or transition state which has energy higher than the reactions or the products.



The rate depends on two factors,

(i). Concentration of  $[AB^\ddagger]$

(ii). The rate at which activated complex is decomposed.

$\therefore$  Rate of reaction =  $[AB^\ddagger]$  x frequency of decomposition of  $AB^\ddagger$

$$K_{eq}^\ddagger = [AB^\ddagger] / [A] [B]$$

$$[AB^\ddagger] = K_{eq}^\ddagger [A] [B]$$

The activated complex is an aggregate of atoms and assumed to be an ordinary molecule. It breaks up into products on a special vibration, along which it is unstable.

The frequency of such a vibration is equal to the rate at which activated complex decompose.

$$-d[A]/dt = -d[B]/dt = k_2[A][B]$$

$$\begin{aligned} \text{Rate of reaction} &= [AB^\ddagger] \nu \\ &= K_{eq}^\ddagger \nu [A] [B] \end{aligned}$$

Activated complex is an unstable species and is held together by loose bonds.



∴ The vibration of low frequency will decompose the activated complex.

∴ Energy of vibration  $E = h\nu$

$$\nu = E/h = k_B T/h$$

In a classical sense, the average energy of such a vibrational degree of freedom,  $k_B T$  (where  $k_B = R/N$ ). Remember equipartition principle.

If we know the energy distribution of molecules clearly (statistical mechanics) we can make a better description.

## Eyring equation

$$k = k_B T/h K_{eq}^\ddagger (C^\circ)^{1-m}$$

$C^\circ$  is the standard – state concentration of  $1 \text{ mol L}^{-1}$ ,  $m$  is the order. The factor  $(C^\circ)^{1-m}$  is introduced because equilibrium constant is dimensionless, but the units for rate constant depend on the order of the reaction.

$$\therefore k = k_B (\text{J K}^{-1}) T (\text{K})/h (\text{J s}) \underset{\uparrow}{K_{eq}^\ddagger} (C^\circ)^{(1-m)}$$

Dimensionless

$k$  is in  $\text{s}^{-1}$  for 1<sup>st</sup> order

for second order,  $\text{L mol}^{-1}\text{s}^{-1} = \text{s}^{-1}(C^\circ)^{-1}$

$$\therefore k_2 = K_{eq}^\ddagger k_B T/h$$

$K_{eq}^\ddagger$  for the formation of the transition state can be written in terms of the molecular partition functions for the reactants and the transition state.

# Thermodynamic Aspects

The equilibrium constant  $K_{eq}^\ddagger$  can be interpreted in terms of  $\Delta G^\ddagger$ ,  $\Delta S^\ddagger$  and  $\Delta H^\ddagger$ .

These quantities represent the difference between the values of the respective thermodynamic functions between activated complex and reactants.

All substances being in their standard states, is activity of unity.

$$\Delta G^\ddagger = -RT \ln K_{eq}^\ddagger$$

$$\Delta H^\ddagger - T\Delta S^\ddagger = -RT \ln K_{eq}^\ddagger$$

$$K_{eq}^\ddagger = \exp(-\Delta H^\ddagger/RT) \exp \Delta S^\ddagger/R$$

In the form of Arrhenius equation  $\rightarrow \therefore k_2 = k_B T/h \exp^{-\Delta H^\ddagger/RT} \exp \Delta S^\ddagger/R$   
This can be expressed in a form which involves  $E_a$  (experimental) instead of  $\Delta H^\ddagger$ .

For that we need to find relation between  $\Delta H^\ddagger$  and  $E_a$ .

The variation of  $K_{eq}^\ddagger$  with temperature is given by,

$$d(\ln K_{eq}^\ddagger)/dT = \Delta E^\ddagger/RT^2 \text{ (one form of Arrhenius law).}$$

The assumption here is that  $K$  is actually,  $K_c$ .

$\Delta E^\ddagger$  represents the increase in internal energy when reactants go from the initial state to the activated state.

$$k_2 = K_{eq}^\ddagger k_B T/h$$

$$\ln k_2 = \ln T + \ln K_{eq}^\ddagger + \ln k_B/h$$

Differentiate with respect to temperature and rewrite.

$$d \ln k_2/dT = 1/T + d \ln K_{eq}^\ddagger/dT$$

$$\therefore d \ln k_2/dT = 1/T + \Delta E^\ddagger/RT^2$$

$$\therefore d \ln k_2/dT = (RT + \Delta E^\ddagger)/RT^2$$

Compare this with Arrhenius equation

$$d \ln k/dT = E_a/RT^2$$

We get  $E_a = \Delta E^\ddagger + RT$

$$[\Delta E^\ddagger \equiv \Delta U^\ddagger]$$

$$H = E + PV$$

Or  $E_a = \Delta H^\ddagger - P\Delta V^\ddagger + RT$

$$\Delta E^\ddagger = \Delta H^\ddagger - P\Delta V^\ddagger \text{ or } \Delta U^\ddagger = \Delta H^\ddagger - P\Delta V^\ddagger$$

At constant P,  $P\Delta V^\ddagger = \Delta n^\ddagger RT$

For ideal gases, it becomes

$$E_a = \Delta H^\ddagger + RT - \Delta n^\ddagger RT$$

$\Delta V^\ddagger$  is volume change and  $\Delta n^\ddagger$  is the change in number of moles from initial state to the activated state.

For bimolecular reaction

$$\Delta n^\ddagger = -1$$

$$E_a = \Delta H^\ddagger + RT - (-1)RT = \Delta H^\ddagger + 2RT$$

$$\Delta H^\ddagger = E_a - 2RT$$

$$\therefore k_2 = k_B T/h \exp^{\Delta S^\ddagger/R} \exp^{-(E_a - 2RT)/RT}$$

$$k_2 = k_B T/h e^2 \exp^{\Delta S^\ddagger/R} \exp^{-E_a/RT}$$

**Note this  $e^2$**

$\Delta S^\ddagger$  can be calculated from  $k_2$  &  $E_a$ .

This term is the A in the Arrhenius expression

Can we show that A is the same by both the theories?

For reactions in solution and for unimolecular reactions,  $\Delta H^\ddagger \approx E_a$   
(almost equal)

$\Delta S^\ddagger$  negative means activated complex is more ordered than the reactants and  $A$  will have smaller value.

$$A = k_B T/h \exp^{\Delta S^\ddagger/R} e \text{ (unimolecular)}$$

$$A = k_B T/h \exp^{\Delta S^\ddagger/R} e^2 \text{ (bimolecular)}$$

Positive  $\Delta S^\ddagger$  means that entropy of the complex is greater than the entropy of the reactants.

A loosely bound complex has a higher entropy of activation than a tightly bound one.

Multiply rate constants with a factor  $(C^\circ)^{1-m}$  to take care of the units of the rate constants (which depend on order).

Example: For the reaction

$\text{CH}_3 + \text{D}_2 \rightarrow \text{CH}_3\text{D} + \text{D}$ , the rate constant is given by

$k = 2 \times 10^9 \exp^{-6389/T} \text{ M}^{-1}\text{s}^{-1}$  over the temperature range 300 to 626 K  
calculate enthalpy and entropy of activation for this reactions at 300K.

Since  $n = -1$ .

$$\Delta H^\ddagger = E_a - 2RT$$

$$= 6389 \times 8.314 - 2 \times 8.314 \times 300$$

$$= 48.1 \text{ kJ/mol}$$

$$\exp \Delta S^\ddagger/R = AhC_0/k_B T e^2$$

$$= ((2 \times 10^9 \text{ M}^{-1}\text{s}^{-1}) (6.63 \times 10^{-34} \text{ Js}) (1\text{M})) / (1.38 \times 10^{-23} \text{ JK}^{-1}) (300\text{K}) e^2)$$

$$= 4.35 \times 10^{-15}$$

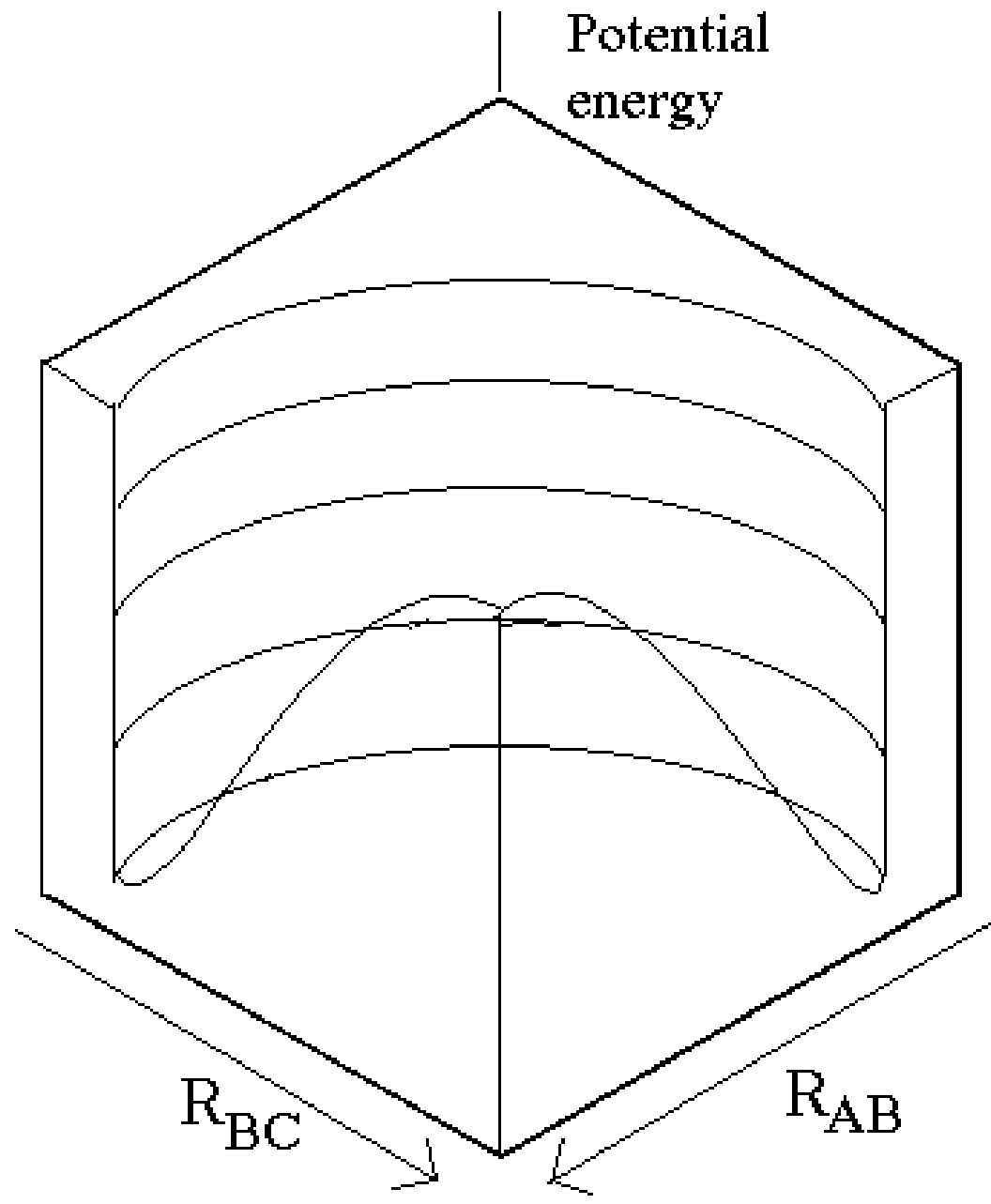
$$\therefore \Delta S^\ddagger = -83.50 \text{ JK}^{-1} \text{ mol}^{-1}$$

$\Delta S$  is negative; which means that the activated complex is more restricted than the reactants. This suggests the formation of a complex between  $\text{CH}_3$  and  $\text{D}_2$  with a bond energy  $\sim 48$  kJ. Evidently, the structure of the complex is rigid enough (vibrational frequencies are high).



## Dynamics of molecular collisions

Reactions between species are better understood by undertaking studies in the gas phase. Here molecular beams can be used to study reactive collisions. Reactants can be prepared in selected energy states and the products can be studied as a function of their energy states. In the case of a typical kinetics experiment, there is a distribution of energy states. The study of factors governing collisions of molecules constitute the area called reaction dynamics. The results can be better understood in terms of potential energy surfaces. For example, the collision between a hydrogen molecule ( $H_2$ ) and a hydrogen atom ( $H$ ) can be represented in terms of a potential energy surface. In this system, it is known that the lowest energy of interaction is when the collision is along a line. The variation in the potential energy of the system as a function of two distances give the potential energy surface. This is depicted below. The interaction can also be depicted in terms of a contour diagram. There are also other approaches. A study of this collision tells how the molecule  $H_A-H_B$  results in the formation of  $H_B-H_C$  upon collision with an  $H$  atom and the various factors which govern this reactive encounter. Such a procedure can be extended to more complicated systems.



## Different equations used

For like molecules

$$\text{Rate} = Z_{AA}/N_{AV} = 1/\sqrt{2} (N_A/V)^2 \pi d^2 \bar{V} (1/N_{AV})$$

$$\text{Rate} = k (N_A/V N_{AV}) (N_A/V N_{AV}) = 1/\sqrt{2} (N_A/V)^2 \sigma \bar{V} (1/N_{AV})$$

$$k/N_{AV} = 1/\sqrt{2} \pi d^2 \bar{V}$$

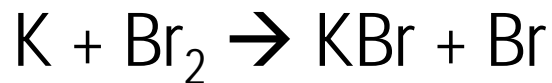
$$k = 1/\sqrt{2} N_{av} \pi d^2 \bar{V}$$

$$= 1/\sqrt{2} N_{av} \pi d^2 (8kT/\pi m)^{1/2}$$

$$= 2 N_{AV} d^2 (\pi kT/m) \text{ mol}^{-1} \text{ m}^3 \text{ s}^{-1}$$

$$k = 2 N_{av} d^2 (\pi RT/M)^{1/2} e^{-E_a/RT}$$

One case where P is large is for the reaction,



Here  $P = 4.8$  (experimentally)

Suggested to follow Harpoon Mechanism

Harpoon extends the cross-section of reactive encounter

Reaction need not occur as a result of simple mechanical encounter

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Calculate P assuming harpoon mechanism

Net change in energy upon transfer of an electron from K to  $\text{Br}_2$  to form  $\text{K}^+$  and  $\text{Br}_2^-$  is

$E = I - E_a - e^2/(4\pi\epsilon_0 R)$ , this will go to zero when harpoon will transfer.

$$e^2/4\pi\epsilon_0 R = I - E_a$$

$$\text{So, } P = \sigma^*/\sigma = R^2/d^2 = \{e^2/4\pi\epsilon_0 (I - E_a)\}^2 d^2$$

Value determined is 4.2, in agreement with experiment (4.8).