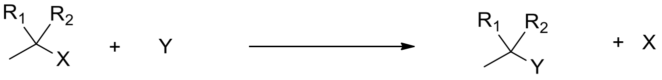
Solutions to Problems – Set 2 : Kinetics and Thermodynamics

Problem 1

Consider the substitution reaction of an alkyl halide



Give two possible reaction profiles corresponding to an SN1 and SN2 reaction, clearly labeling the important parts of the diagram.

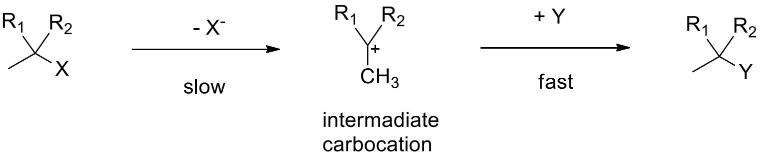
Problem 2

For the above reaction in the SN2 case, show the mechanism with the transition state including the breaking and forming bonds. Also draw the mechanism for SN1 reaction showing the intermediates.

Solution for Problem 1 + 2

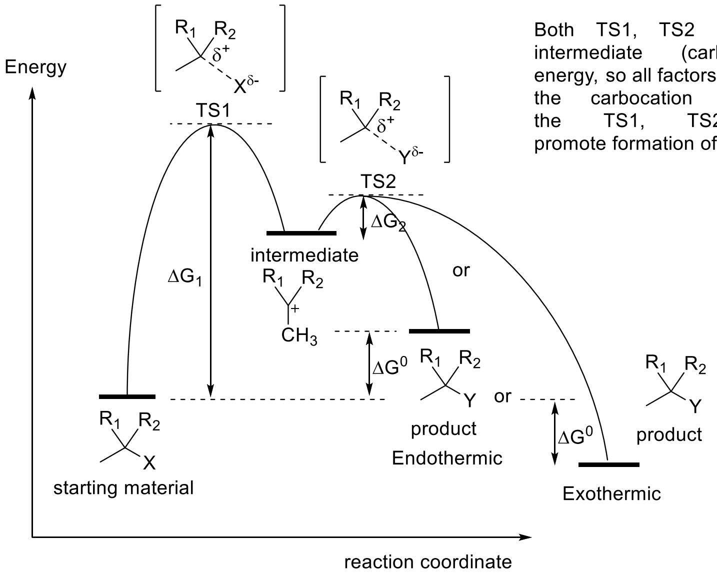
SN1 reaction:

Mechanism



Rate = k[starting material]

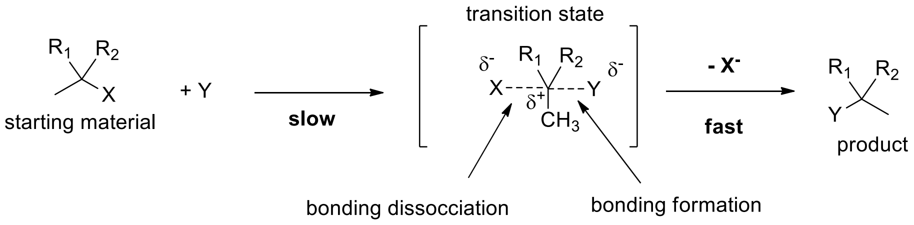
Energy profile of *SN1* reaction



Both TS1, TS2 are close to intermediate (carbocation) in energy, so all factors which stabilize the carbocation will stabilize the TS1, TS2, therefore promote formation of product.

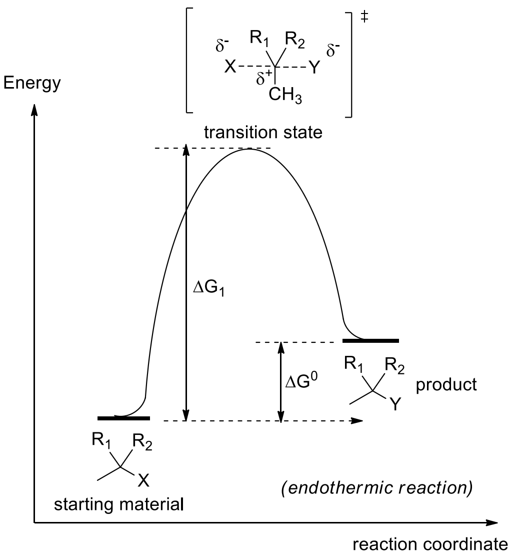
S_N2 reaction:

Mechanism

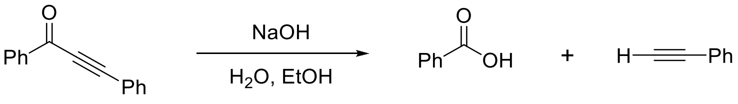


Rate = k [starting material][Y]

Energy profile of *S_N2* reaction



Problem 3

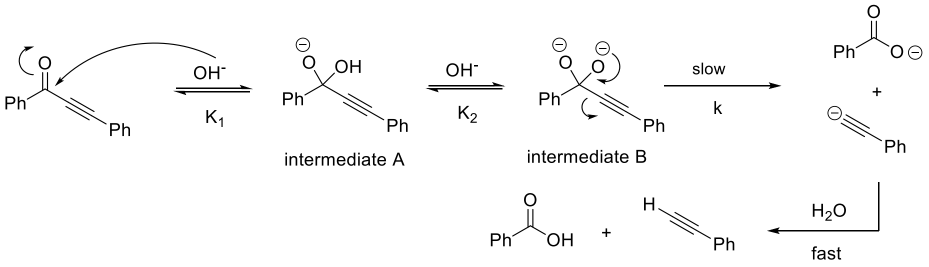


This reaction shows third order kinetics as the rate expression is rate=**k**[ketone]$[OH^-]^2$. Suggest a mechanism that explain these observations and draw the energy profile. (k: constant)

Solution

The hydroxide group must attack the ketone to form a tetrahedral intermediate. The best leaving group from this intermediate is the hydroxide ion that has just come in (pK_{aH} 15) rather than the alkyne anion (pK_{aH} ~25). If we use the second hydroxide ion to deprotonate the intermediate, we have only one possible leaving group, though it is a bad one, and the decomposition of the dianion must be the rate-determining step. This mechanism is usually found for nucleophilic substitution at a carbonyl group with a very bad leaving group such as amide hydrolysis. As the reaction is performed in protic solvents the generated alkynyl anion is immediately protonated, pushing though the reaction.

Possible mechanism:



Explain the rate expression according above proposed mechanism:

$$\text{rate} = k[B]$$

For 2 equilibrium, we have:

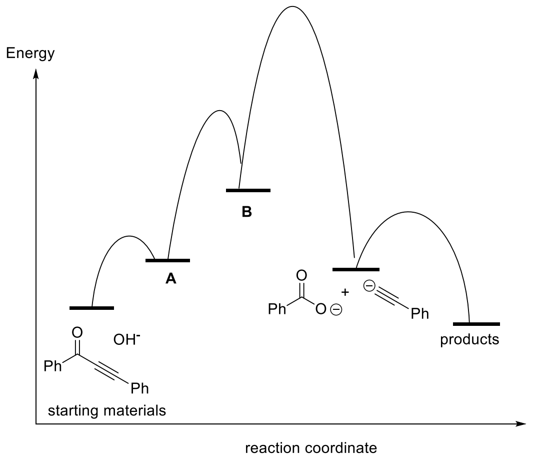
$$[A] = K_1[OH^-][\text{ketone}]$$

$$[B] = K_2[OH^-][A]$$

$$\Rightarrow [B] = K_1K_2[OH^-]^2[\text{ketone}]$$

$$\Rightarrow \text{rate} = kK_1K_2[OH^-]^2[\text{ketone}] \sim k'[OH^-]^2[\text{ketone}]$$

Energy profile:



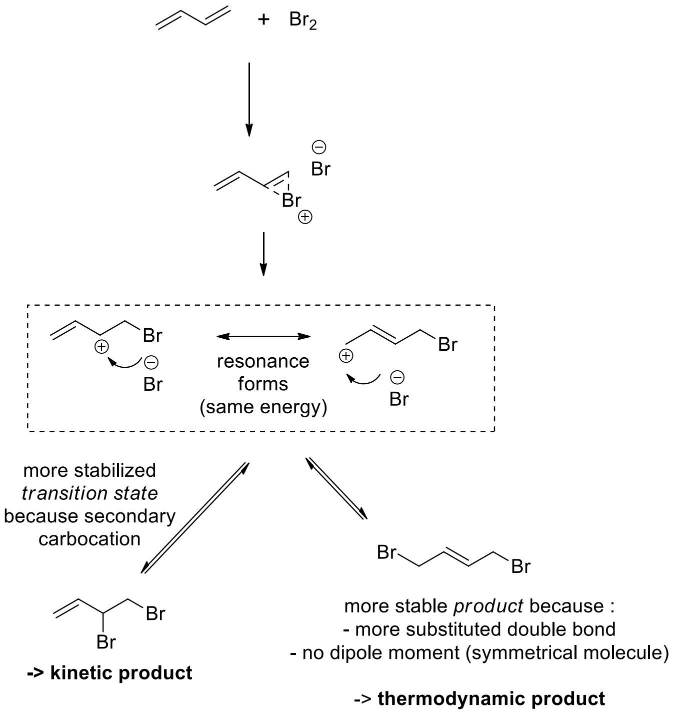
Problem 4

Consider the reaction between Br_2 and 1,3-butadiene

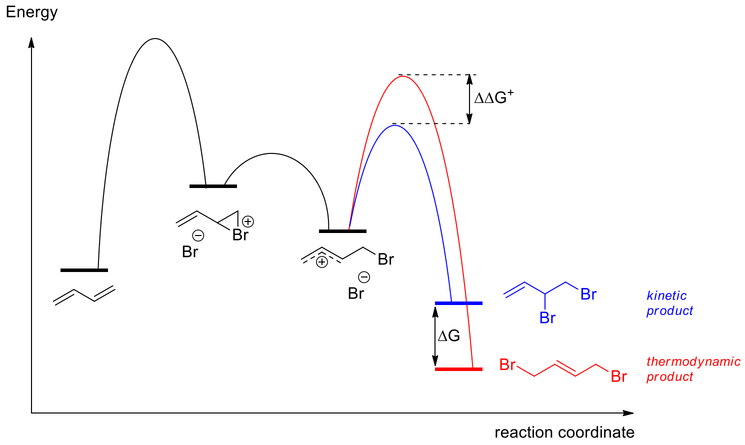
Draw the energy profile diagram and explain the difference between the formation of these 2 products.

Solution

Mechanism:

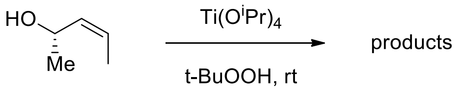


Energy profile:



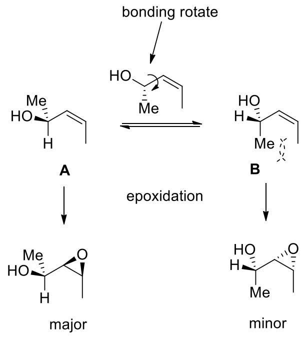
Problem 5

Draw the major/minor products expected from this reaction:



The theoretical transition state energy $\Delta\Delta G^{++}$ = -1.43 kcal/mol, and experimentally the product ratio = 91:09. Determine whether this reaction obeys the Curtin-Hammett principle by calculating the expected ratio of products based on the theoretical $\Delta\Delta G^{++}$ value.

Solution



$\Delta\Delta G^{++}$ is difference in the free energies of activation (which is also mentioned in next problem)

Do not forget to use the good unit: R= 8.314 $J.K^{-1}.mol^{-1}$= 1.987 cal $.K^{-1}.mol^{-1}$=

-1.987*10^{-3} kcal K^{-1} mol^{-1}

If this reaction obeys Curtin-Hammett principle, the distribution of products will be dependent on $\Delta\Delta G^{++}$ by following equation:

$\Delta\Delta G^{++}$= -RTln$[P_a]/[P_b]$

How we get that equation?

- First, ratio $[P_a]/[P_b]$ ~ ratio of v_a/v_b (rate of formation A and B) ~ ratio of k_a/k_b (because A, B are products from the same starting material)
- From Arrhenius equation, we have

  $\Delta\Delta G^{++}$= -RT ln$[k_a]/[k_b]$ = -RT ln$[P_a]/[P_b]$ = -1.987*10^{-3} kcal K^{-1} mol^{-1} * 293 K * ln(91/9) = -1.347 kcal/mol

The theoretical value of $\Delta\Delta G^{++}$ = -1.43 kcal/mol is close to the experimental value of

-1.347 kcal/mol and therefore the reaction follows the Curtin-Hammett principle.

To present the problem from another point of view, you can determine the ratio of product you should obtain starting from the theoretical $\Delta\Delta G^{++}$:

$\Delta\Delta G^{++}$= -RT ln$[k_a]/[k_b]$ = -RT ln$[P_a]/[P_b]$

$\Delta\Delta G^{++}/(-RT)$= ln$[P_a]/[P_b]$ → e$(\Delta\Delta G^{++}/(-RT))$= $[P_a]/[P_b]$= e(-1.43 kcal.mol^{-1}/(-1.987*10^{-3} kcal K^{-1} mol^{-1} * 293 K))= 11.66

$[P_a]$= 11.66$[P_b]$

To obtain the % of each products:

100% = %$[P_a]$ + %$[P_b]$ = 11.66*%$[P_b]$ + %$[P_b]$ =12.66*%$[P_b]$

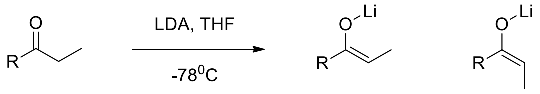
%$[P_b]$=100/12.66= 7.9% ≈ 8%

%$[P_a]$= 100 - 8 ≈ 92%

The theoretical ratio of products 92/8 is close to the experimental value of 91/9 and therefore the reaction follows the Curtin-Hammett scenario. This means that the interconversion of the two conformers is faster than the rate of reactions to the products. Thereby the product ratio depends on the difference in the free energy of the transition states going to each product (and not solely on the equilibrium constant between the conformers).

Problem 6

Use the proportion of enolates shown for the two cases (R=t-Bu and R=Et) to determine the difference in the free energies of activation at -78°C, assuming that enolate formation is kinetically controlled?



R=t-Bu	98%	2%
R=Et	30%	70%

Solution

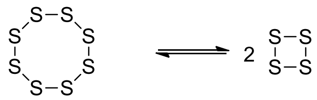
Similar to problem 5, we apply the same equation $\Delta\Delta G^{++}$= -RTln$[k_a]/[k_b]$ = -RTln$[P_a]/[P_b]$

R = t-Bu: $[P_a]/[P_b]$ = 98/2 so $\Delta\Delta G^{++}$= -RTln$[P_a]/[P_b]$ = -1.98*10^{-3} kcal K^{-1} mol^{-1} * 195 K * ln(98/2) = -1.5 kcal/mol (this mean activation energy of A is lower than B's)

R = Et: $[P_a]/[P_b]$ = 30/70 so $\Delta\Delta G^{++}$= -1.98*10^{-3} kcal K^{-1} mol^{-1} * 195 K * ln(30/70) = 0.33 kcal/mol

Problem 7

Elemental sulfur normally exists as an eight-membered ring (S_8), but it can also be found in a number of other states. How would entropy and enthalpy affect the equilibrium between sulfur in these two forms?



Solution

Gibbs energy of this reaction: $\Delta G_r = \Delta H_r - T\Delta S_r$

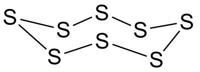
Reminder: If $\Delta G_r < 0$, reaction will favour to right, $\Delta G_r > 0$ will favour to left.

The equilibrium position is determined both by enthalpy and entropy.

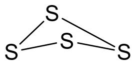
For this reaction: $\Delta H_r > 0$, $\Delta S_r > 0$

Why?

Both S_8 and $2S_4$ contain 8 S-S bonds. But S_8 bonds are more stable than S_4 because of ring strain in S_4. So from energy view, transformation from S_8 to 2 S_4 is endothermic process or $\Delta H_r > 0$.



no ring strain



ring strain

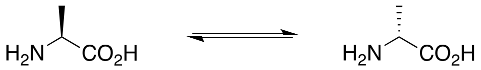
Moreover, transformation from S_8 to 2 S_4 increases number of molecules (1 to 2) or increases the randomness, then $\Delta S_r > 0$.

So **heat** (high temperature) will drive the equilibrium to the right because $T\Delta S_r$ has more effect than ΔH_r. In contrast, Enthalpy favoring the eight-membered ring will be more important at **lower temperature** and favor the left-hand side.

Problem 8

Enantiomerically pure α-amino acids are isomerized into their racemic mixtures. What is the Gibbs Energy of this reaction at 25°C?

Solution



(S)-P

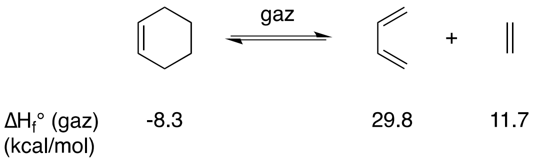
(R)-P

Before isomerization	100% (S)-**P**
At equilibrium	50% (S)-**P**, 50% (R)-**P**
At equilibrium	$\Delta G_r = -RT\ln K = -RT\ln 1 = 0$

$\Delta G_r = 0$ → Both enantiomers have the same energy level, the temperature does not affect the difference in energy. This racemization is possible in presence of a catalyst or by adding enough energy to promote the isomerization (to overpass the energy barrier).

Problem 9

The following reaction has a standard entropy of 35cal·mol^{-1}·K^{-1}. At which temperature will you obtain a full conversion (conversion < 99%) of the starting material?



Solution

The enthalpy of reaction is the sum of the enthalpies of formation of the products, minus the sum of the enthalpies of formations of the reagents.

$\Delta H_r = 29.8 + 11.7 - (-8.3) = 49.8$ kcal/mol.

By using the basic relations linking the Gibbs energy and the equilibrium constant, we can find the temperature needed.

Do not forget to use the good unit:

R= 8.314 J.K^{-1}.mol^{-1}= 1.987 cal .K^{-1}.mol^{-1}= -1.987*10^{-3} kcal K^{-1} mol^{-1}

ΔSr = 35cal·mol-1·K-1 =0.035 kcal K-1 mol-1

$\Delta G_r = \Delta H_r - T\Delta S_r = -RT\ln Keq$

Keq = [products]/[starting material] = (0.99*0.99)/0.01

$T = \Delta H_r/(\Delta S_r - R\ln Keq) = 49.8/(0.035-0.00198*\ln[(0.99*0.99)/0.01]) = 1921K = 1648°C$

We can see that this retro-Diels-alder is really difficult to obtain. The entropic factor is not big enough to compensate easily the endothermicity.