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Solvent Effects on the Monomer-Dimer Equilibrium and Dimerization of the 1-Methyl-2-methoxycarbonylpyridinyl Radical

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Solvent effects on the hyperfine splitting constants, monomer-dimer equilibrium, and kinetic parameters of the dimerization of the 1-methyl-2-methoxycarbonylpyridinyl radical have been examined in detail, using methylcyclohexane, toluene, 2-methyltetrahydrofuran, acetone, and acetonitrile as the solvents. A comparatively large activation energy for the fast step in the two-step dimerization mechanism suggests strongly that the dimeric intermediate is the 2,2′-dimer of the radical, which subsequently rearranges to the stable 4,4′-dimer.

Equilibrium between the 1-methyl-2-methoxycarbonylpyridinyl radical (R) and the covalently bonded dimer (D) was discussed in our previous papers.[1,2] Since the equilibrium lies largely in favor of D at room temperature, and since D decomposes reversibly to generate R upon irradiation with light in the near-UV region, the dimerization mechanism showing a two-step decay of R in 2-methyltetrahydrofuran (MTHF) was successfully elucidated by kinetic ESR spectroscopy.[3] In this study, however, the structure of the dimeric intermediate (DI) in the two-step decay process of R was left for further investigation, though the π-complex dimer was suggested for it. Concerning the structure of DI, a kinetic study of the dimerization of the 1-alkyl-4-phenylpyridinyl radicals suggests the 4,4′-dimer as DI; it subsequently rearranges to the stable 2,2′-dimer through the [3,3]sigmatropic mechanism.[4] The activation energy for the DI formation was too large (>27 kJ mol^{-1}) for us to suggest the π-complex dimer as DI. Because of the short lifetime of DI, the attempted spectroscopic analysis of DI has been unsuccessful.[3]

We then examined the thermodynamic parameters of the equilibrium and kinetic parameters for the dimerization in various solvents with a view to establishing the magnitude of solvent dependence for these parameters and also whether or not the two-step mechanism can be used in solvents with different polarities. We are also interested in solvent effects on neutral radicals, as have been reported previously for the stable 1-methyl-4-methoxycarbonyl- and 1-methyl-4-acetyl-pyridinyl radicals.[5] The present examination provides valuable information about solvent effects on the equilibrium system formed by a pyridinyl radical with an intermediate stability.

Methylcyclohexane, toluene, MTHF, acetone, and acetonitrile were used as solvents. The Block-Walker reaction field parameters were adopted in discussing the experimental data. This paper will end by proposing the 2,2′-dimer of R as a favorable structure for DI.

Experimental

Materials. The pyridinyl radical R was prepared by the reduction of 1-methyl-2-methoxycarbonylpyridinium iodide with 3% sodium amalgam in degassed acetonitrile, purified by distillation as previously reported,[1] and then dissolved in the desired solvent. Commercial methylcyclohexane (spectral grade) was dried over sodium metal, degassed, and then distilled under a vacuum. The purification of the other solvents has been described in a previous paper.[6]

Concentration of the Species. The addition of 1,1-dimethyl-4,4′-bipyridinium dichloride (methylviologen) to a solution of R and D caused a quantitative electron-transfer reaction, thus generating the methylviologen cation radical. The total concentration of the species was determined from the absorption measurement of the cation radical (ε 13000 at λ_{max} 605 nm). An electronic double-integrator of ESR signals was used to determine the concentration of R. A benzene solution of 2,2-diphenyl-1-picrylhydrazyl was used as the standard sample, and the concentration was determined from its optical absorbance (ε 14150 at 519 nm). The ESR sensitivity-variation in different solvents was calibrated with an internal standard of strong pitch in a capillary tube.

ESR Spectroscopy. A Varian E-109E EPR spectrometer was used for the ESR measurements. The system of kinetic ESR spectroscopy has been described previously.[3]

Results and Discussion

Solvent Effects on the Hyperfine Splitting Constants of R. In a previous paper,[5] an expression for the solvent effects on the hyperfine splitting (hfs) constants of a free radical was derived by obtaining the π-electron spin densities of the radical in solution by means of the perturbation theory; it was demonstrated that, for some neutral radicals, the hfs constants are correlated with the Block and Walker reaction-field parameter of $\theta(\varepsilon_r)=3\varepsilon_r(\ln \varepsilon_r)/(\varepsilon_r\ln \varepsilon_r-\varepsilon_r+1)-6/(\ln \varepsilon_r)-2$, where ε_r is the relative permittivity of the solvent.[7] We then used this parameter to express the solvent dependence of the hfs constants for the present radical. The equilibrium ($2\,R \rightleftharpoons D$) system showed well-resolved hfs of the ESR spectrum in all solvents, and each spectrum was analyzed with six splitting constants, to which the assignments were given by comparing them with the spin densities calculated by the McLachlan procedure. Plots of the hydrogen and nitrogen hfs constants (a_{obsd}) *vs.* $\theta(\varepsilon_r)$ are shown in Fig. 1. In this figure, the sign of each a_{obsd} value was chosen according to these relations; $a^{H}_{obsd}=Q^{H}_{CH}\,\rho_{calcd}$, with $Q^{H}_{CH}<0$,[8] and $a^{N}_{obsd}=Q^{N}\rho^{\pi}_{N}+Q^{N}_{CN}\times\sum\rho^{\pi}_{CN}$, with $|Q^{N}|>|Q^{N}_{CN}|$ and positive Q^{N},[9,10] nega-

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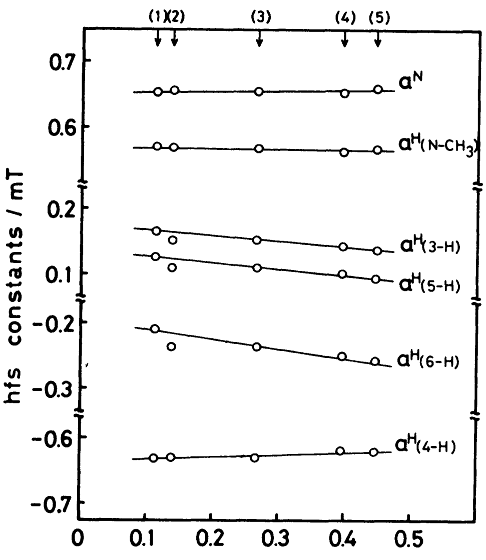

Fig. 1. Relationship between the hfs constants of R and the Block-Walker's parameters, $\theta(\varepsilon_r)$. The numbers in parentheses correspond to solvents. (1) Methylcyclohexane, (2): toluene, (3): MTHF, (4): acetone, and (5): acetonitrile.

tive spin densities appearing at the 3 and 5 positions.

The correlation between the values of a_{obsd} and $\theta(\varepsilon_r)$ is good, though the plots for the toluene solvent deviate significantly from the lines obtained by the least-squares method. The deviation is probably due to a specific interaction, such as a charge-transfer mechanism between the radical and the solvent toluene. The variations in the a_{obsd} values with the change in the solvent are small, the largest change being observed at the 6 position from 0.208 mT in methylcyclohexane to 0.256 mT in acetonitrile. Accordingly, the solvent effects on the spin distribution are not very large, not even for the density at the 2 position with the largest spin density in the pyridine ring; therefore, the dimerization mechanism is not influenced by the perturbation in the shape of the singly occupied molecular orbital due to the change in solvent. Therefore, the solvent effects described hereinafter will mainly be ascribed to the stabilization of the energy level by solvation.

Solvent Effects on the Monomer-Dimer Equilibrium. The monomer-dimer equilibrium constant is expressed as K=[D]/[R]^2=([Total]−[R])/2[R]^2, where [Total] is the total concentration of D and R, which can be determined by titration with 1,1′-dimethyl-4,4′-bipyridinium dichloride. [R] is determined by the use of an electronic double-integrator of ESR signals. The measurements were carried out at temperatures between 3 °C and 45 °C. Plots of log K *vs.* 1/T afforded a good straight line for each solvent, as is shown in Fig. 2. The thermodynamic parameters obtained from these linear relationships are listed in Table 1. The equilibrium tends to shift towards the D side with an increase in the solvent polarity. Plots of the free energy change (ΔG) against $\theta(\varepsilon_r)$ gave a linear relationship, with the exception of the value in toluene, as is shown in Fig. 3. The enthalpy change (ΔH) also shows a solvent dependence, the absolute value ($|\Delta H|$) becoming smaller with an increase in the solvent polarity.

Because R has a dipole moment, while that of D is equal to zero or very small, the stabilization of R by solvation with polar solvents would be larger than that of D. The solvation enthalpy, $\Delta\Delta H$, is assumed to be the difference between the dimer-formation enthalpy in a solution (ΔH_{solv}) and that *in vacuo* (ΔH_{vac});[11] therefore, $\Delta\Delta H_{solv}=\Delta H_{solv}-\Delta H_{vac}=\Delta H_{D,solv}-\Delta H_{2R,solv}$, where

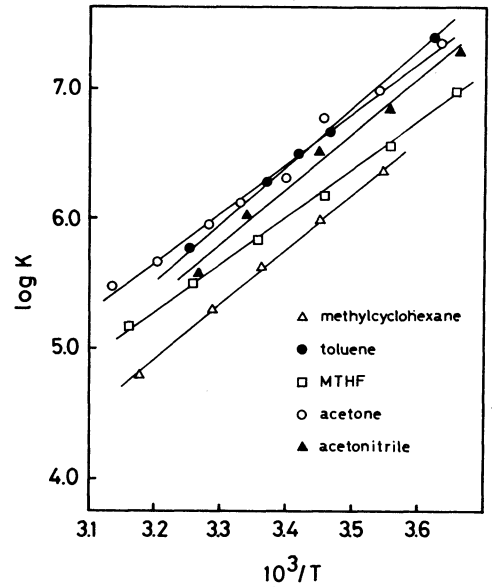

Fig. 2. Monomer-dimer equilibrium constants (K) *vs.* 1/T.

TABLE 1. THERMODYNAMIC PARAMETERS FOR THE MONOMER-DIMER EQUILIBRIUM OF R IN VARIOUS SOLVENTS

Solvent	ΔH	ΔS	$\Delta G^{a)}$	$K^{a)}$
	kJ mol^{-1}	J K^{-1} mol^{-1}	kJ mol^{-1}	mol^{-1} dm^3
Methyl-cyclohexane	−82.5	−170	−31.8	3.86×10^5
Toluene	−86.7	−172	−35.5	1.66×10^6
MTHF	−69.0	−121±4	−32.9	7.44×10^5
Acetone	−74.1	−129	−35.6	1.89×10^6
Acetonitrile	−76.5	−141	−34.5	1.10×10^6

a) At 25 °C.

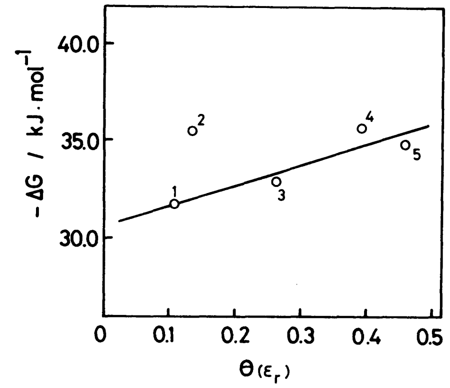

Fig. 3. Free energy change (ΔG) at 25 °C *vs.* the Block-Walker's parameters ($\theta(\varepsilon_r)$).

$\Delta H_{D,solv}$ and $\Delta H_{2R,solv}$ are the solvation enthalpies of D and R respectively. As the dipole moment of methylcyclohexane is approximately zero, the enthalpy change in this solvent is assumed to be that under a vacuum. Then, $\Delta H_{vac}=\Delta H_{methylcyclohexane}=-82.5$ kJ mol^{-1}. For the case of the acetone solvent, $\Delta\Delta H_{acetone}$ becomes 8.4 kJ mol^{-1}; this positive value indicates the large stabilization of the R side. In a similar way, the solvation free energy is obtained: $\Delta\Delta G_{acetone}=-3.8$ kJ mol^{-1}; this indicates that D is stabilized by the solvation at this temperature. The difference in signs of $\Delta\Delta G$ and $\Delta\Delta H$ is caused by the entropy term: $\Delta\Delta S_{acetone}=\Delta S_{acetone}-\Delta S_{vac}=41$ JK^{-1} mol$^{-1}=\Delta S_{D,acetone}-\Delta S_{2R,acetone}$. Since $\Delta S_{methylcyclohexane}$ ($\simeq\Delta S_{vac}$) is largely negative (Table 1), the positive $\Delta\Delta S_{acetone}$ value can be explained if there is a stronger orientation of acetone molecules to R than to D. It is thus assumed that the ΔS value of -170 JK^{-1}mol^{-1} in methylcyclohexane is mostly the entropy change for D formation and that, in acetone, the $\Delta\Delta S_{acetone}$ value contributes positively to ΔS, suppressing it to the observed value (-129 JK^{-1}mol^{-1}). Consequently, the importance of the entropy term in polar solvents is established for the present equilibrium system.

Kinetic Parameters of Dimerization. In our previous paper,[3] a kinetic ESR study aimed at the dimerization of R generated by the flash photolysis of D in MTHF was described. Similar examinations for R in methylcyclohexane exhibited the growth-and-decay curves shown in Fig. 4. The following features were recognized in the same manner as in MTHF: (1) a two-step decay of R is observed around −65 °C, as shown in Fig. 4b; (2) the decay of the fast step observed at tempratures below −40 °C is represented as a reversible second- and first-order for the range of −95—−70 °C; (3) the slow step is a first-order decay in the −40—40 °C range, and (4) no change in the hfs pattern is exhibited. The same features were observed in toluene and acetone. In acetonitrile, the measurements were carried out at temperatures above −40 °C, and a first-order decay was observed.

Although the fast step could be treated as a second-order reaction, some deviations were observed because of the influence of the reverse reaction. Therfore, the step was treated as a reversible second- and first-order reaction, k_1 one which can be represented by the following equation:

$$2k_1t = -\frac{1}{2[\bar{R}]}\ln\left|\frac{[R]-[\bar{R}]}{[R]+[\bar{R}]}\Big/\frac{[R]_0-[\bar{R}]}{[R]_0+[\bar{R}]}\right|$$

where [R], $[\bar{R}]$, and $[R]_0$ are the concentration, the pre-equilibrium concentration, and the initial concentration of the radical respectively. The plots based on this equation gave good linear relationships, as is shown in Fig. 5a. The Arrhenius plots yielded an activation energy (E_{a1}) of 28.9 kJ mol^{-1} and a frequency factor (A) of 1.5×10^{14} mol^{-1} dm^3 s^{-1}. These values and the rate constant obtained by extrapolation to 25 °C are

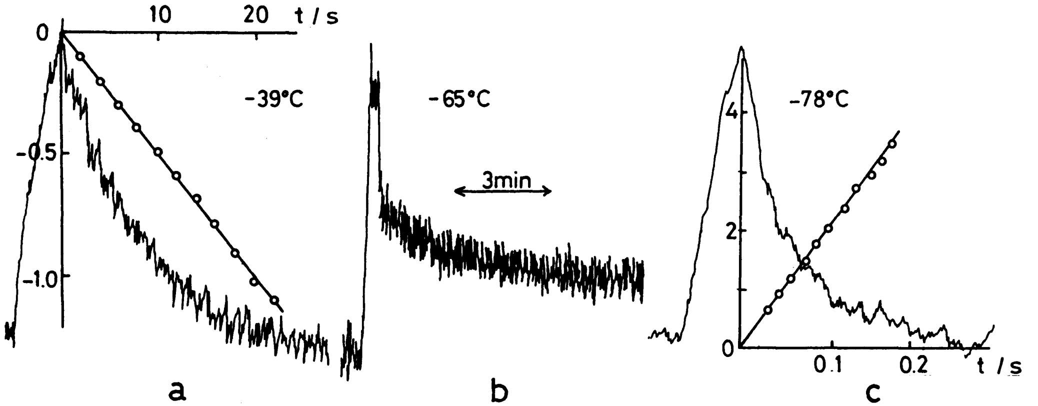

Fig. 4. Growth-and-decay curves of R in methylcyclohexane.
a: The first-order decay at −39 °C; the unit of the vertical axis is $\ln([R]-[\bar{R}])/([R]_0-[\bar{R}])$.
b: Two-step decay at −65 °C. c: The reversible second- and first-order decay at −78 °C; the unit of the vertical axis is

$$-\frac{1}{2[\bar{R}]}\ln\left|\frac{[R]-[\bar{R}]}{[R]+[\bar{R}]}\Big/\frac{[R]_0-[\bar{R}]}{[R]_0+[\bar{R}]}\right|.$$

TABLE 2. ACTIVATION PARAMETERS FOR THE DIMERIZATION OF R IN VARIOUS SOLVENTS

Solvent	Fast step				Slow step			
	$\log A$	E_{a1}	ΔG_1[a]	$2k_1$[a]	$\log A$	E_{a2}	ΔG_2[a]	k_2[a]
	mol^{-1}dm^{3}s^{-1}	kJ mol^{-1}	kJ mol^{-1}	mol^{-1}dm^{3}s^{-1}	s^{-1}	kJ mol^{-1}	kJ mol^{-1}	s^{-1}
Methyl-cyclohexane	14.2	28.9	20.9	1.4×10^9	11.8	56.9	62.8	63.8
Toluene	9.5	15.9	34.7	5.1×10^6	10.9	54.8	65.3	21.3
MTHF[b]	10.88	20.9	31.8	1.6×10^7	10.98	54.4	64.4	28.0
Acetone	10.7	19.2	31.4	2.0×10^7	8.32	42.7	68.6	6.15
Acetonitrile					12.9	65.3	64.4	31.8

a) At 25 °C, b) From Ref. 3.

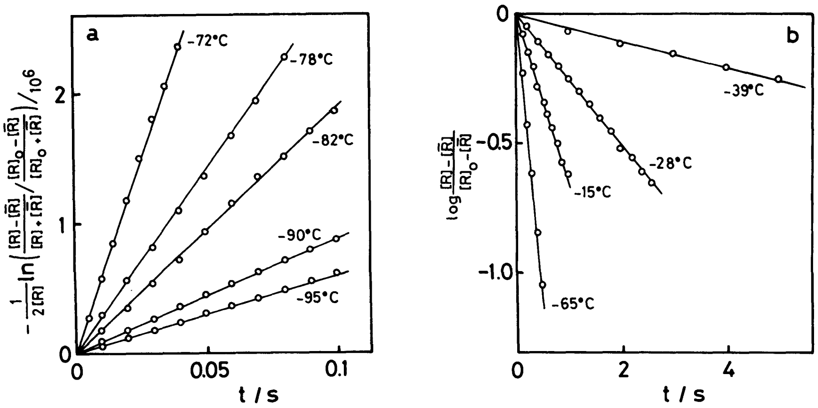

Fig. 5. (a) Reversible second- and first-order plots for the fast step and (b) first-order plots for the slow step in the decay of R in methylcyclohexane.

shown in Table 2, together with those obtained using toluene, MTHF, and acetone as solvents. The rate constant in methylcyclohexane is comparable to that in a diffusion-controlled reaction.[12]

The E_{a1} value of 28.9 kJ mol^{-1} in methylcyclohexane tends to decrease with an increase in the solvent polarity, showing 19.2 kJ mol^{-1} in acetone, while the ΔG value at 25 °C increases and k_1 decreases with an increase in the solvent polarity. The relationship between the ΔG values and the solvent parameters $\theta(\varepsilon_r)$ was poor, though a tendency could be recognized. The activation parameters for the fast step are also indicative of a great influence of the frequency factor upon k_1. The ΔS values, as estimated from the A values, are 4.5, −17.1, −10.7, and −11.5 JK^{-1} mol^{-1} for R in methylcyclohexane, toluene, MTHF, and acetone respectively. It is noteworthy that a positive value is obtained for methylcyclohexane. Since the fast step passes through a transition state in which two radicals are oppositely oriented to each other by the dipole-dipole interaction, the positive ΔS value presumably originates from the strong interaction between two radicals, regardless of the nonpolar solvent molecules. This orientation would be reduced in its strength by solvation to R in polar solvents, resulting in the negative ΔS values, because the solvation corresponds to the reduction of the dipole moment of the radical.

The rate constant, k_2, of the slow step is represented by this equation:

$$k_{\text{obsd}} = \frac{k_2}{2}t = -\ln\frac{[R]-[\bar{R}]}{[R]_0-[\bar{R}]}$$

The good relationships shown in Fig. 5b were obtained for the measurements in methylcyclohexane. The activation parameters obtained from the Arrhenius plots are summarized in Table 2 for the results in five solvents.

The E_{a2} value has a tendency to decrease with an increase in the solvent polarity. The large value for acetonitrile is abnormal. As a long irradiation of the acetonitrile solution caused significant damage to the sample, there may be a specific interaction between the sample and the solvent. The values of ΔG_2 and k_2 show that the reaction becomes slow with an increase in the ΔG_2. Concerning the solvent dependence of ΔG_2, the plots *vs.* $\theta(\varepsilon_r)$ showed a good relationship among the solvents except for a slight deviation for toluene.

Preequilibrium between R and DI. The above results clearly prove that the radicals generated by the photolysis of D rapidly establish a preliminary equilibrium. In the irradiation with a strong, short light pulse, the initial concentrations of R and DI immediately after the irradiation are represented as $[R]_0$ and $[DI]_0$ respectively. At the preequilibrium, these concentrations become $[\bar{R}]$ and $[\overline{DI}]$ respectively. The constant (K_1) can be determined from the following equation:

$$K_1 = \frac{[\overline{DI}]}{[\bar{R}]^2} = \frac{[R]_0-[\bar{R}]}{2[\bar{R}]^2}$$

since $[DI]_0/([R]_0-[\bar{R}]) \ll 1$.

The measurements were performed at the shutter speed of 1/15 s, using a UV-29 glass filter at temperatures between −70 °C and −95 °C. Under these conditions, however, accurate values for K_1 could not be obtained. For instance, the measurement in methylcyclohexane gave no fine signal because the formation of DI was very rapid. The reliable results obtained in toluene, MTHF, and acetone were in these ranges: $K_1=(1.0—2.0)\times10^7$ mol dm^{-3} (at −75 °C) and $\Delta H_1=$ −23—−33 kJ mol^{-1}.

On the Structure of DI. The above results can enable us to draw a free energy *vs.* reaction coordinate

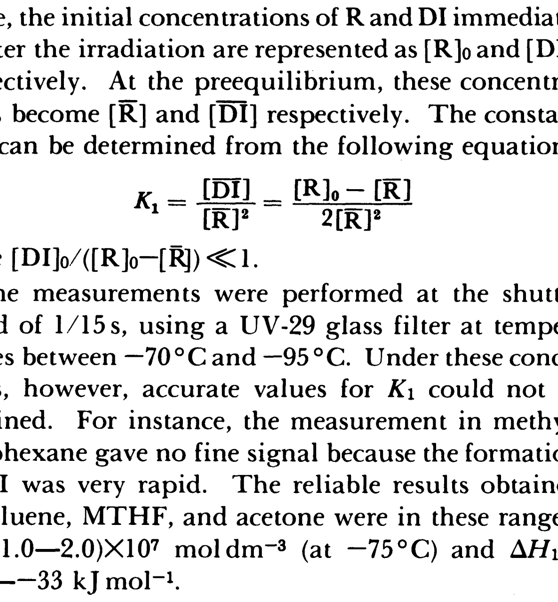

Fig. 6. A free energy at 25 °C *vs.* reaction coordinate diagram for the dimerization of R proceeding *via* DI. ——: in methylcyclohexane. -----: in acetone.

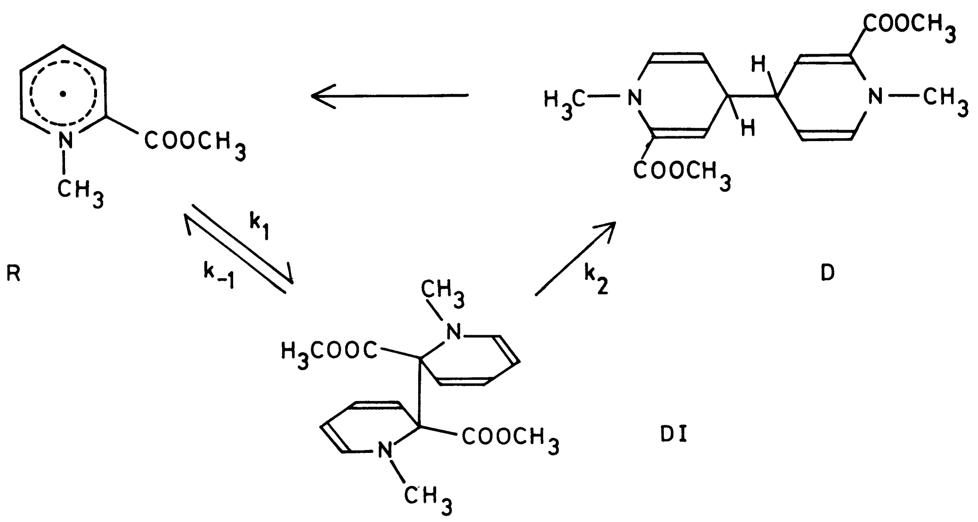

diagram for the present dimerization, as depicted for the methycyclohexane and acetone solvents in Fig. 6.[13] The structure of D reported previously[1] was recently confirmed by NMR measurements at a temperature below 0 °C, the spectrum disclosing that the dimer consists of two diastereoisomers.[14]

Concerning the structure of DI, the previous paper[3] suggested that DI was a singlet π-complex with a sandwich-like structure and expected that a charge-transfer absorption due to the complex would be observed. Further examinations, however, provided no evidence supporting the π-complex structure.

The observation of the two-step decay of R in all five solvents means that the dimerization mechanism does not differ with the solvent, and the relatively large $E_{\mathrm{a}1}$ values strongly imply that DI is a σ-bonding dimer rather than a π-complex. The value of 28.9 kJ mol^{-1} in methylcyclohexane is not conceivable for the $E_{\mathrm{a}1}$ of π-complex formation. Since free radicals are apt to react at positions with a high spin density, it is simply assumed that the coupling point is the 2 position, where the highest spin density is indicated by the McLachlan calculation.[2] The transformation of the intermediate 2,2′-dimer into the stable 4,4′-dimer can be interpreted as a [3,3]sigmatropic rearrangement, by analogy with the dimerization of the 1-alkyl-4-phenylpyridinyl radicals.[4]

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13) In this figure, the ΔG value for the $2\mathrm{R}\rightleftharpoons 2,2'$-dimer equilibrium is assumed to be about 30 kJ mol^{-1}, which was estimated from the K_1 value at 25 °C obtained by the extrapolation of the values measured in the range of −70—−95 °C.

14) Y. Ikegami and others, unpublished results.