March, 1987] © 1987 The Chemical Society of Japan Bull. Chem. Soc. Jpn., 60, 851—855 (1987) 851

Fast Reactions of Organic Anion Radicals with Organic Halides in Hexamethylphosphoric Triamide Studied by Pulse Radiolysis

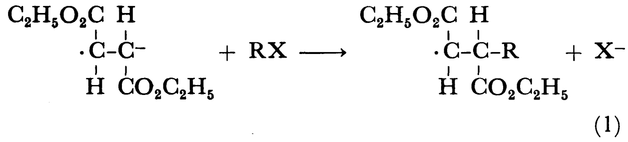
Eiji Honda, Masao Tokuda, Hiroshi Yoshida, and Masaaki Ogasawara*
Faculty of Engineering, Hokkaido University, Sapporo 060

(Received August 25, 1986)

Fast reactions of diethyl fumarate anion radical (DEF$^-$) and fluorenone anion radical (Fl$^-$) with various organic halides (RX) in hexamethylphosphoric triamide have been studied by means of ns pulse radiolysis at room temperature. Reactions of acetophenone anion radical were also studied for comparison. It was found that the reaction rate of Fl$^-$ was subject to the steric and resonance effects of R groups of RX in accord with the classical concept of S_N2 reactions: the rate constant was reduced by 2 orders of magnitude by the steric effect when R was changed from ethyl to bulky isopropyl or *t*-butyl, and it was still large by the resonance effect of R even if R was changed from ethyl to an allyl or a benzyl group. While the reaction rate of DEF$^-$ was not much affected when R was changed to more bulky groups, the rate constant was correlated to the reduction potential of RX. The results were interpreted in terms of a VB correlation diagram approach or rate-equilibrium relationship within a framework of S_N2 reactions.

Most organic anion radicals react efficiently with organic halides, the mechanism being dependent on the kind of anion radical and organic halide. An electron-transfer reaction from radical anions to organic halides accompanied by a dissociation of the electron acceptor into two fragments, organic radical and halide anion, is most commonly observed.[1] Nucleophilic substitution reaction at a saturated carbon atom of the organic halides may also take place. The possibility of an $S_{RN}1$-type substitution reaction,[2] in which the dissociative electron transfer from anion radical to organic halide is followed by addition reaction of the organic radicals to the product neutral molecules, should also be considered.

A few years ago, Satoh et al. studied electrochemical reactions of diethyl fumarate with alkyl halides and confirmed the carbon–carbon bond formation at a saturated carbon atom of the halide.[3] Following nucleophilic substitution reaction of diethyl fumarate anion radical with alkyl halides, RX, was proposed to explain the results.



$$\begin{array}{c}\mathrm{C_2H_5O_2C\ \ H}\\ |\quad\ |\\ \cdot\mathrm{C{-}C^-}\\ |\quad\ |\\ \mathrm{H\ \ CO_2C_2H_5}\end{array} + \mathrm{RX} \longrightarrow \begin{array}{c}\mathrm{C_2H_5O_2C\ \ H}\\ |\quad\ |\\ \cdot\mathrm{C{-}C{-}R}\\ |\quad\ |\\ \mathrm{H\ \ CO_2C_2H_5}\end{array} + \mathrm{X^-} \quad (1)$$

Similar nucleophilic substitution reactions were found in electrochemical reactions of the fluorenone anion radical.[4]

We report here the pulse radiolysis kinetics of the reactions of the diethyl fumarate anion radical and the fluorenone anion radical with several organic halides, in special reference to free-energy relationship of the reactions, in hexamethylphosphoric triamide (HMPA) at room temperature. Our aim is to attempt to understand the factors which determine the reactivity of these anion radicals toward organic halides.

Experimental

Pulse Radiolysis Apparatus. The source of the electron pulse was the Hokkaido University 45-MeV linear accelerator. The pulse was reducible in its duration down to as short as 10 ns, but only 100 ns pulse with a dose of 140 Gy/pulse was used in the present study. The cells and optical system have been described elsewhere.[5] The electric signals from a Hamamatsu Photonics R446 photomultiplier were digitized by an Iwatsu TS-8123 Storage-scope and the digitized data were processed by a NEC PC-9801 computer.

Materials. HMPA was dried over calcium hydride for 24 h, refluxed with fresh calcium hydride for 2 h, and distilled at reduced pressure before use. 9-Fluorenone (Nakarai Guaranteed Reagent) was used without purification. Ethyl bromide, iso-propyl bromide, *t*-butyl bromide, allyl bromide and ethyl iodide were distilled over calcium chloride before use. Benzyl bromide was distilled over calcium chloride at a reduced pressure before use. The solutions were degassed in vacuum line and sealed into Suprasil cells of the size 1×1×4.5 cm^3.

Cyclic Voltammetry. Cyclic voltammetry was made with Nichia HP-E$_{500}$ potentiostat at a sweep rate of 0.2 V s^{-1} using a platinum disc electrode (1 mm diameter) or hanging mercury drop electrode (HMDE) in solutions of 0.1 mol dm^{-3} n-Bu$_4$NClO$_4$ in HMPA. The potential was measured in V vs. Ag/AgI.

Results

Diethyl Fumarate Anion Radical. Figure 1 shows the transient absorption spectra observed in the pulse radiolysis of solutions of 4×10^{-3} mol dm^{-3} diethyl fumarate (DEF) in HMPA. The spectrum with an absorption maximum at 340 nm which was observed immediately after a pulse is assigned to the diethyl fumarate anion radical (DEF$^-$) formed by electron capture by DEF. For comparison, the absorption spectrum of *dimethyl* fumarate anion radical obtained at 77 K in γ-irradiated 2-methyltetrahydrofuran matrix is also shown.[6] Spectra observed with the samples containing a small amount (1×10^{-2} mol dm^{-3}) of ethyl bromide as the second solute are the same as those observed without ethyl bromide except for the band in the UV region (<320 nm) which is discernible at 250 μs after the pulse. This UV band may be due to the reaction product of DEF^- with the halide.

The time-profile of the absorption monitored at 340 nm consists of a fast-decaying component and a slow-decaying one, although the contribution of the latter component is only one tenth of the initial absorption. It is assumed that only the fast-decaying component which is separable by subtracting the slow-decaying one from the original curve is due to the DEF^-. In the absence of ethyl bromide the decay curve of DEF^- followed a second-order kinetics, suggesting that the DEF^- disappeared via recombination reactions with solvent cations or radicals produced by ionizing radiation. In the presence of ethyl bromide the decays were analyzed as being mixed first- and second-order processes. Evidently, the first-order process is due to a pseudo- first-order reaction of the DEF^- with ethyl bromide. The rate equation therefore is expressed as:

$$-d[DEF^-]/dt = k_1[DEF^-]^2 + k_2[DEF^-]. \tag{2}$$

Where k_1 is the second-order rate constant of the decay reaction of DEF^- in the absence of RX. This was determined empirically as $(1.1\pm0.3)\times10^9$ $OD^{-1}s^{-1}$ under the same experimental conditions. The k_2 infers the rate constant of the pseudo-first-order process which can be rewritten as $k_2=k[RX]$ with the k being the second-order rate constant between DEF^- and RX. With the k as a parameter the decay curves of the DEF^- were simulated, as exemplified in Fig. 2, by the method described elsewhere.5) The k thus obtained increases linearly with increase of the RX concentration and the second order rate constants, k, are determined from the slope of the straight lines as summarized in Table 1.

Fluorenone Anion Radical. In the pulse radiolysis of solutions of fluorenone in HMPA, transient absorption spectra with a single peak at 570 nm were observed. The spectra are assigned to the fluorenone

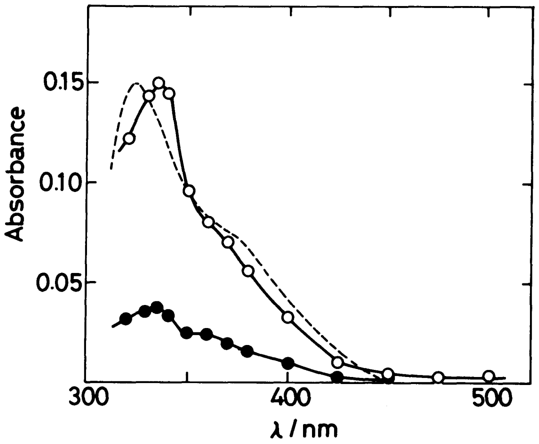


Fig. 1. Transient absorption spectra obtained in the pulse radiolysis of solutions of 4×10^{-3} mol dm^{-3} diethyl fumarate in HMPA: —○— immediately after the pulse, —●— 250 μs after the pulse. The dashed curve indicates the steady state absorption spectrum of dimethyl fumarate anion radical in 2-methyltetrahydrofuran matrix at 77 K.6)

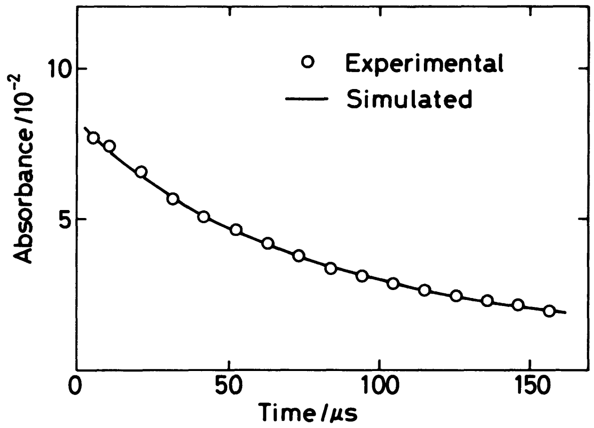


Fig. 2. Decay curve of diethyl fumarate anion radical observed at 340 nm in HMPA solution containing 4×10^{-3} mol dm^{-3} diethyl fumarate and 9×10^{-3} mol dm^{-3} ethyl bromide. The solid line indicates the decay curve simulated on the basis of superimposed first- and second- order kinetics (see the text).

Table 1. Rate Constants of the Reactions between Organic Anion Radicals and Organic Halides

Halogen Compound	Reduction Potential $-E/V$	DEF^- k/mol^{-1} dm^3 s^{-1}	Fl^- k/mol^{-1} dm^3 s^{-1}	$Acet^-$ c) k/mol^{-1} dm^3 s^{-1}	Solvent
C_2H_5Br	2.30a)	$(5.0\pm1.1)\times10^5$	$(3.5\pm0.9)\times10^8$	—	HMPA
$(CH_3)_2CHBr$	2.24a)	$(8.1\pm1.4)\times10^5$	$(3.1\pm0.9)\times10^6$	$(7.5\pm2.5)\times10^6$	HMPA
ibid.	ibid.	—	$(2.0\pm0.3)\times10^6$	—	CH_3CN
$(CH_3)_3CBr$	2.02b)	—	$(1.6\pm1.0)\times10^6$	—	HMPA
$CH_2=CHCH_2Br$	1.62a)	$(1.2\pm0.4)\times10^6$	$(6.3\pm2.8)\times10^8$	—	HMPA
$C_6H_5CH_2Br$	1.47a)	$(2.6\pm0.9)\times10^6$	$(1.1\pm0.3)\times10^8$	—	HMPA
C_2H_5I	1.84a)	$(1.6\pm0.4)\times10^6$	—	—	HMPA

a) Determined by CV with Pt disc electrode; V vs. Ag/AgI. b) Determined by CV with a hanging mercury drop electrode; V vs. Ag/AgI. c) Acetophenone anion radical.

anion radical (Fl^-) because of the similarity to those obtained in the pulse radiolysis of ethanol solutions at low temperature and of *N,N*-dimethylformamide solutions at room temperature.[7] In the presence of ethyl bromide essentially the same spectra were observed immediately after the pulse, but a small absorption at ca. 530 nm, probably due to the reaction products of Fl^- with halides, survived for a long time.

The spectra of Fl^- depended markedly on the solvent and the counterion. In acetonitrile solutions the λ_{max} was shifted to ca. 530 nm and a long tail extending on the longer wavelength side was observed. In 2-methyltetrahydrofuran (MTHF) solutions containing sodium tetraphenylborate as *the third solute,* transient spectrum with λ_{max} at 540 nm which was ascribed to ion pair between Fl^- and Na^+ was observed. The shift of λ_{max} in acetonitrile and in MTHF should be due to solvation and coulombic interaction of the counterion, respectively.

The decay kinetics of Fl^- also depended on the solvent. In HMPA, the decay curve monitored at 570 nm arose from a first-order process with a rate constant of $7.3 \times 10^4\ s^{-1}$. The reaction mechanism of the decay reaction is not known, but the anions may be consumed by reactions with impurities in HMPA. Addition of the second solutes, RX, apparently accelerated the decay rate, indicating efficient reaction of Fl^- with RX. The pseudo-first-order rate constant increased linearly with increase of the RX concentration. Second-order rate constants of Fl^- were determined from the slope of the straight lines for various RX. The reactions of the Fl^- in acetonitrile, which follows a second-order kinetics in the absence of RX, are analyzed by the same method as the case of DEF^-. The results are summarized in Table 1. They showed that the ion pairs, Fl^-Na^+, in MTHF did not react with organic halides in a μs time-domain.

Acetophenone Anion Radical. The absorption spectrum of the acetophenone anion radical in HMPA has a broad band in the visible with double maxima at 460 and 480—490 nm, which disappear following a first-order kinetics. The second-order rate constants *k* of the reaction between acetophenone anion radical and ethyl bromide are listed in Table 1.

Reduction Potentials of Organic Halides. Reduction potentials (E_{red}) of organic halides in HMPA were determined by cyclic voltammetry. The type of electrode employed in the measurements, Pt or HMDE, did not have much effect, though in general the HMDE electrode gives a little more positive potentials for E_{red}. In Table 1 only the results obtained with the Pt electrode are shown except the one for *t*-butyl bromide.

Figure 3 shows cyclic voltammograms of HMPA solutions containing allyl bromide and $n\text{-}Bu_4NClO_4$. The single peak observed in the cyclic voltammetry with Pt electrode (Fig. 3(a)) may be ascribed to the following reactions.

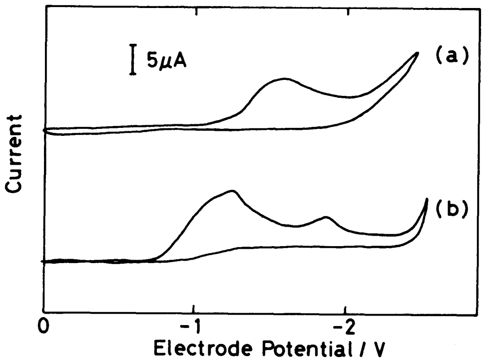


Fig. 3. Cyclic voltammograms of HMPA solutions containing 5×10^{-3} mol dm^{-3} allyl bromide and 0.1 mol dm^{-3} $n\text{-}Bu_4NClO_4$ recorded with Pt electrode (a) and HMDE electrode (b). The electrode potential was measured vs Ag/AgI.

$$CH_2{=}CHCH_2Br + e^- \longrightarrow (CH_2{=}CHCH_2Br)^{\cdot -} \longrightarrow CH_2{=}CHCH_2\cdot + Br^- \qquad (3)$$

$$CH_2{=}CHCH_2\cdot + e^- \longrightarrow CH_2{=}CHCH_2^- \qquad (4)$$

The second peak at −1.87 V, which was observed exclusively with HMDE electrode (Fig. 3(b)), can be ascribed to the following reactions.

$$CH_2{=}CHCH_2\cdot + Hg \longrightarrow CH_2{=}CHCH_2Hg\cdot \xrightarrow{CH_2=CHCH_2\cdot} (CH_2{=}CHCH_2)_2Hg \qquad (5)$$

$$CH_2{=}CHCH_2Hg\cdot + e \longrightarrow CH_2{=}CHCH_2^- + Hg \qquad (6)$$

We adopted the value evaluated from the first peak of the cyclic voltammogram shown in Fig. 3(a) as the E_{red} for allyl bromide.

Discussion

The effect of RX on the decay curve of DEF^- clearly indicates that DEF^- reacts with RX very efficiently. In the study of electrochemical reaction of DEF, Satoh et al. concluded that the electron transfer from the electrode to DEF was followed by an S_N2 reaction of the produced DEF^- with RX.[3] Analogously, the most plausible reaction of DEF^- with RX would be an S_N2-type substitution reaction. The resulting radical should be the R-adduct radical of DEF as shown by Eq. 1, although its formation has not yet been confirmed optically. It should be noted that the reaction rate is not much affected when R is changed from ethyl to a bulky isopropyl group. This result is not fully consistent with the classical concept of an S_N2 reaction which is characterized by effective hindrance by the presence of bulky group. Instead, the rate constant is correlated to E_{red} of RX as shown in Fig. 4.

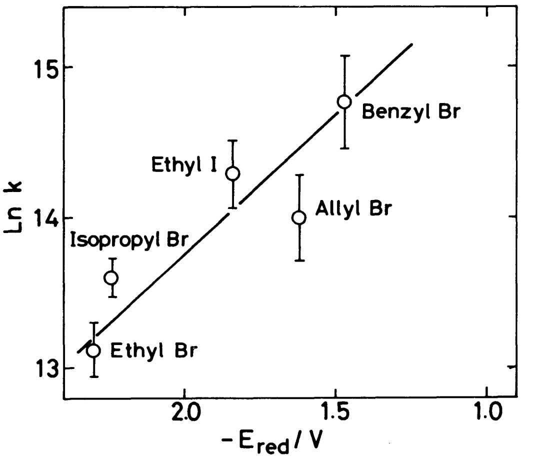


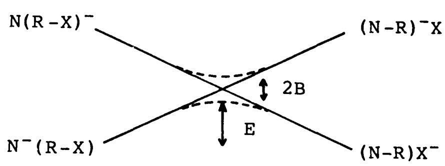
Fig. 4. Second-order rate constant for the reactions at room temperature between diethyl fumarate anion radical and organic halides as a function of the reduction potential of the organic halides.

In S_N2 reactions a nucleophilic reagent forms a new bond to the carbon atom under attack, the old bond breaking. In the transition state the incoming group and the leaving group are both "half-bonded" to the carbon atom. Therefore it had been thought that "crowdedness" in the formation of the activated complex was the most important factor in determining the reaction rate.[8] Actually, in the classical bromide-exchange reaction of alkyl bromide in acetone, the reaction rate changes drastically, by a factor of about 100, when alkyl group of the organic halides is replaced from an ethyl to an isopropyl group.[9] However, recently, some of the theoretical and experimental studies suggested that the difference in the shape or size of the reactant molecules played no obvious part in determining the reaction rate. In some cases those effects can be incorporated into parameters such as proton affinity, ionization potential, or reduction potential of the reactants.[10] The reaction of DEF$^-$ with RX observed in the present study is one of the examples of such cases.

In order to interpret the correlation between the k of the S_N2 reaction and the E_{red} of the reactants, the VB correlation diagram approach can be invoked. The S_N2 reaction is a transformation that involves a synchronous electron shift and a bond interchange as expressed below.

$$N^- + R\text{-}X \longrightarrow N\text{-}R + X^- \qquad (7)$$

In the VB correlation diagram model which was invented and developed by Shaik,[11] four anchor states, i.e. the ground state reactants ($N^-(R-X)$) and products ($(N-R)X^-$) and the corresponding valence-charge-transfer forms of the reactants ($N(R-X)^-$) and products ($(N-R)^-X$), are correlated with each other.



In between these four anchor states a barrier arises owing to avoided crossing of the intersecting curves. The height of the barrier (E) is a fraction (f) of the energy gap separating the anchor states $N(R-X)^-$ and $N^-(R-X)$, less the avoided crossing, B, i.e.

$$E = f(I_{N^-} - A_{RX}) - B \qquad (8)$$

I_{N^-} is the vertical ionization potential of the nucleophile, N^-, while A_{RX} is the vertical valence electron affinity of the substrate, RX, that can be related to E_{red}. If constancy of I_{N^-}, f and B for a reaction series is maintained, variation in the activation energy depends only on E_{red}.

Alternatively within a framework of rate-equilibrium relationship originally developed by Marcus,[12] the following elementary steps are assumed.

$$Y^- + RX \rightarrow \underset{\mathrm{I}}{Y^-\ RX} \rightarrow \underset{\mathrm{II}}{YR\ X^-} \rightarrow RY + X^- \qquad (9)$$

Brauman and his co-workers proposed a double minimum potential surface at the first (I) and the second (II) intermediates in Eq. 9.[13] As the reactants approach one another, long-range ion-induced dipole and ion-dipole interactions produce a drop in potential energy. There is then a potential energy increase up to the central transition state between the I and II states which becomes the intrinsic energy barrier of the S_N2 reaction. The difference in the E_{red} of Y and RX may be related to the free energy difference, ΔG, between the I and II states. The ΔG is related, in turn, to the intrinsic energy barrier of the S_N2 reaction by use of the Marcus equation.[10b] Thus, the rate of S_N2 becomes a function of E_{red}. It should be emphasized that the Marcus approach is a very general one and is not limited to electron-transfer reactions, as the parabolic free energy profile arises from summing displacement along many different coodinates.[14]

Another possible reaction mechanism is a dissociative electron-transfer reaction from anion radicals to organic halides. The Marcus quadratic equation for reversible electron-transfer reaction predicts the slope α of the acitivation energy vs. ΔG plot to be >0.5 for positive ΔG. In some systems where the electron-transfer is part of an overall irreversible transformation, such as chemical-bond formation or dissociation into two fragments, the experiments give values for α usually in the ranges 0.2 to 0.5.[15] Actually, α in the present study, which is deduced from the gradient of the straight line in Fig. 4, is as small as 0.1. This value is not consistent with the prediction or the empirical low of the electron-transfer reaction. Besides, free energy differences of the electron-transfer from DEF$^-$ or FI$^-$ to RX evaluated from the E_{red} of the reactants are very endothermic. The electron-transfer reaction should, even if it occurred, be too slow to be observed by pulse radiolysis in ns or μs time-domains. However, the reaction with benzyl bromide can proceed via an electron-transfer mechanism, since the E_{red} of this halide is only slightly more negative than that of DEF. As it is known that C–C bond formation is accomplished by the reaction of DEF$^-$ with benzyl bromide, so called $S_{RN}1$ mechanism in which the electron transfer occurs prior to C–C bond formation may be applicable.[2]

In the case of the reactions of Fl$^-$, it is obvious that there is no correlation between the *k* and the E_{red} of RX. Rather, *k* is determined by the steric and resonance effect of the R groups as is commonly seen in typical S_N2 reaction. The *k* is reduced by 2 orders of magnitude by the steric effect when R changes from ethyl to the bulky isopropyl or *t*-butyl groups. However *k* still remains large because of the resonance effect of R if R changes from ethyl to a bulky allyl or benzyl group. When a benzene ring of the Fl$^-$ is replaced by methyl group, the reaction is accelerated as demonstrated by the rate constant for acetophenone anion radical shown in Table 1. Thus, an S_N2-type substitution reaction between Fl$^-$ and RX is strongly indicated. This conclusion is in accord with the electrochemical study by Lund and Simonet which indicated that Fl$^-$ reacted with RX to form R-adducts.[4] In these substitution reactions,the reaction site of the nucleophilic reagent is supposed to be a negatively charged oxygen atom. The difference between Fl$^-$ and DEF$^-$, in the hindrance of S_N2 reactions by the presence of a bulky group, may be interpreted as due to the bond length in the transition state between the incoming group and the carbon atom under attacked being shorter as compared with the case of DEF$^-$, with Fl$^-$ as the nucleophile. The shorter bond length may be caused by the difference in the reaction site, O or C atom, of the substitutions.

The authors would like to express their thanks to Messrs. Hiroaki Tanida and Kohichi Sato for operating the 45-MeV Electron Linear Accelerator of Hokkaido University.

References

1) a) J. F. Garst, *Acc. Chem. Res.*, **4**, 400 (1971); b) B. Bockrath and L. M. Dorfman, *J. Phys. Chem.*, **77**, 2618 (1973); c) S. Takamuku, H. Kigawa, S. Toki, K. Tsumori, and H. Sakurai, *Bull. Chem. Soc. Jpn.*, **54**, 3688 (1981).

2) R. A. Rossi, *Acc. Chem. Res.*, **15**, 164 (1982).

3) a) S. Satoh, T. Taguchi, M. Itoh, and M. Tokuda, *Bull. Chem. Soc. Jpn.*, **52**, 951 (1979); b) S. Satoh, H. Suginome, and M. Tokuda, *ibid.*, **54**, 3456 (1981).

4) H. Lund and J. Simonet, *Bull. Soc. Chim. Fr.*, 1843 (1973).

5) M. Ogasawara, N. Kajimoto, T. Izumida, K. Kotani, and H. Yoshida, *J. Phys. Chem.*, **89**, 1403 (1985).

6) T. Shida, S. Iwata, and M. Imamura, *J. Phys. Chem.*, **78**, 741 (1974).

7) M. Hoshino, S. Arai, and M. Imamura, *J. Phys. Chem.*, **78**, 1473 (1974).

8) E. S. Gould, "Mechanism and Structure in Organic Chemistry," Henry Holt and Co., New York, N. Y. (1959), p. 251.

9) P. B. D. de La Mare, L. Foeden, E. D. Hughes, C. K. Ingold, and J. D. Mackie, *J. Chem. Soc.*, 3200 (1955).

10) a) F. G. Bordwell and D. H. Hughes, *J. Org. Chem.*, **48**, 2206 (1983); b) C. C-. Han, J. A Dodd, and J. I. Brauman, *J. Phys. Chem.*, **90**, 471 (1986).

11) a) S. S. Shaik, *J. Am. Chem. Soc.*, **103**, 3692 (1981); b) S. S. Shaik, *J. Am. Chem. Soc.*, **105**, 4359 (1983); c) S. S. Shaik, *J. Am. Chem. Soc.*, **106**, 1227 (1984).

12) R. A. Marcus, *J. Chem. Phys.*, **24**, 966 (1947).

13) a) M. J. Pellerite and J. I. Brauman, *J. Am. Chem Soc.*, **102**, 5993 (1980); b) M. J. Pellerite and J. I. Brauman, *J. Am. Chem. Soc.*, **105**, 2672 (1983).

14) W. Albery and M. M. Kreevoy, *Adv. Phys. Org. Chem.*, **16**, 87 (1978).

15) F. Scandola, V. Balzani, and G. B. Schuster, *J. Am. Chem. Soc.*, **103**, 2519 (1981).