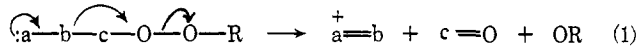
The Base-Catalyzed Fragmentation of a Peroxide, 2-*t*-Butylperoxy-2-methylpropanoic Acid

William H. Richardson and Ronald S. Smith

Contribution from the Department of Chemistry, San Diego State College, San Diego, California 92115. *Received February 6, 1969*

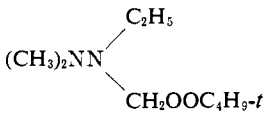
Abstract: A kinetic and product study of the amine-catalyzed fragmentation reaction of 2-*t*-butylperoxy-2-methylpropanoic acid is reported. The kinetic data are most simply explained by a mechanism which includes an equilibrium step between the acid and amine to give an ion pair. The ion pair then undergoes a fragmentation reaction in a subsequent step. The final products of the reaction are *t*-butyl alcohol, acetone, and carbon dioxide. Data are presented which show that the fragmentation reaction is concerted. Solvent effects are consistent with the proposed ionic mechanism and they are correlated with the χ_R solvent parameter. Activation parameters in some of these solvents are reported for the amine-catalyzed reaction. The parameters are consistent with the proposed mechanism and differ markedly from the thermal decomposition of the acid in the absence of amine. Comparisons are made with analogous fragmentation and elimination reactions.

A recent review[1] stresses the importance of heterolytic fragmentation reactions as a class of chemical transformations. Decompositions of several peroxidic compounds have been interpreted in terms of a heterolytic fragmentation mechanism. The general formulation of this mechanism is given by eq 1.[2] Such

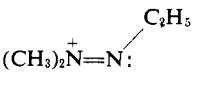


(1)

a mechanism is advanced for certain nitrogen-containing peroxides, where the nitrogen atom is represented by a or a and b in eq 1.[3] For example, N,N-dimethyl-N′-*t*-butylperoxymethyl-N′-ethylhydrazine (I) undergoes decomposition to give *t*-butyl alcohol, formaldehyde, and N,N-dimethyl-N′-ethylhydrazine, where II is a

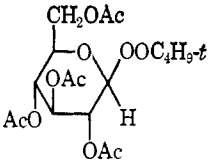


I

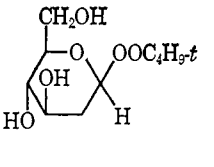


II

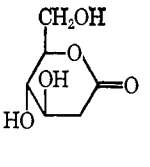
proposed intermediate in the reaction. A peroxide fragmentation mechanism is suggested for a series of reactions that are employed in the degradation of peroxy sugars.[4] Here, the oxygen atom of an alkoxide anion corresponds to a in eq 1. The intermediate alkoxide anion is generated from the acetyl derivative in these reactions. For example, when *t*-butylperoxy 2,3,4,6-tetraacetyl-β-D-glucoside (III) was treated with sodium methoxide in methanol, D-arabinose, *t*-butyl alcohol, and methyl formate were produced.[4e] In contrast, if the 2-acetyl group is absent, an elimination reaction occurs. For example, the basic reaction of *t*-butylperoxy 2-deoxy-D-glucoside (IV) gave lactone V and *t*-butyl alcohol. Several other reactions may be interpreted in terms of a peroxide fragmentation reaction according to eq 1, where a is an oxygen anion.[5]



III



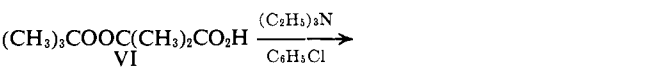
IV



V

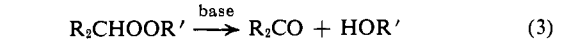
In many of the peroxide fragmentation reactions, the actual fragmentation step originates from an intermediate which is formed in a series of preliminary reactions. This is particularly true when a is an oxygen anion. Furthermore, little kinetic data are available to substantiate the mechanism of peroxide fragmentation. The half-lives for fragmentation of two nitrogen-containing peroxides are reported in ethanol and benzene.[3a] An ionic mechanism was suggested from the observation that the half-lives were decreased when the solvent was changed from benzene to ethanol. This appears to be the extent of the kinetic data.

In a preliminary communication[5i] we reported the fragmentation reaction of 2-*t*-butylperoxy-2-methylpropanoic acid (VI) with triethylamine in chlorobenzene solvent (eq 2). This reaction appeared to be ideally



(2)

suited to pursue the mechanism of peroxide fragmentation in detail for systems where a is an oxygen anion (eq 1). For example, direct comparisons may be made between this fragmentation reaction and the analogous peroxide elimination reaction (eq 3). The kinetics and



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(1) C. A. Grob and P. W. Schiess, *Angew. Chem. Intern. Ed. Engl.*, **6**, 1 (1967).

(2) A. Rieche, *ibid.*, **5**, 523 (1966).

(3) (a) E. Schmitz, A. Rieche, and A. Stark, *Ber.*, **101**, 1035 (1968); (b) L. A. Cohen and B. Witkop, *J. Am. Chem. Soc.*, **77**, 6595 (1955); (c) B. Witkop and J. B. Patrick, *ibid.*, **73**, 2196 (1951); (d) B. Witkop, *ibid.*, **72**, 1428 (1950).

(4) (a) M. Schulz and H. Steinmaus, *Angew. Chem. Intern. Ed. Engl.*, **2**, 623 (1963); (b) M. Schulz and H.-F. Boeden, *Tetrahedron Letters*, 2843 (1966); (c) M. Schulz and L. Somogyi, *Angew. Chem. Intern. Ed. Engl.*, **6**, 168 (1967); (d) M. Schulz, H.-F. Boeden, and E. Gründemann, *Z. Chem.*, **7**, 13 (1967); (e) M. Schulz, H.-F. Boeden, and P. Berlin, *Ann.*, **703**, 190 (1967).

(5) (a) J. Hoffman, *J. Am. Chem. Soc.*, **79**, 503 (1957); (b) D. D. Jones and D. C. Johnson, *J. Org. Chem.*, **32**, 1402 (1967); (c) D. H. R. Barton and E. Seoane, *J. Chem. Soc.*, 4150 (1956); (d) H. H. Wasserman and M. B. Floyd, *Tetrahedron Letters*, 2009 (1963); (e) W. Reusch and R. LeMahieu, *J. Am. Chem. Soc.*, **85**, 1669 (1963); (f) L. J. Bollyky, R. H. Whitman, R. A. Clarke, and M. M. Rauhut, *J. Org. Chem.*, **32**, 1663 (1967); (g) M. Avramoff and Y. Sprinzak, *J. Am. Chem. Soc.*, **85**, 1655 (1963); (h) M. M. Rauhut, D. Sheehan, R. A. Clarke, B. G. Roberts, and A. M. Semsel, *J. Org. Chem.*, 30, 3587 (1965); (i) W. H. Richardson and R. S. Smith, *J. Am. Chem. Soc.*, **89**, 2230 (1967).

products of eq 3 have been reported in chlorobenzene with amines serving as the base.[6,7] A comparable peroxide elimination reaction of *t*-butylperoxy formate was studied using amines as the base.[9] Again kinetics were obtained in a number of nonpolar solvents including chlorobenzene.

Results

Products. The over-all reaction for the base-catalyzed decomposition of 2-methyl-2-*t*-butylperoxypropanoic acid (VI) is given by eq 2, where the yields of acetone and *t*-butyl alcohol are both 100 ± 1%. In addition, the infrared spectrum of triethylamine and VI in chlorobenzene was measured periodically during the course of reaction. A strong absorption band develops at 2310 cm^{-1} and then gradually disappears after completion of the reaction. The position of absorption of this band coincides with carbon dioxide.[10]

Order in Reactants. Acceptable first-order rate constants were obtained for individual runs as indicated by the probable error. Ideally we would have desired to maintain a high constant concentration of the amine and vary the peroxy acid concentration to verify the order in the latter reactant. However, precision in the titrimetric data usually suffered when the amine to peroxy acid concentration ratio exceeded about 2:1. When the peroxy acid concentration was varied by a factor of 2.5 with the amine concentration constant at 20.0 × 10^{-2} *M* (Table I), the first-order rate constants were the same within experimental error.

Table I. Effect of Concentration and Temperature on the Rate of Reaction of 2-Methyl-2-*t*-butylperoxypropanoic Acid (VI) in Chlorobenzene[a]

Run	Temp, °C	[VI] × 10^2, *M*	[$(C_2H_5)_3N$] × 10^2, *M*	$k \times 10^4$, sec^{-1}
1	0.20	4.00	8.00	0.386 ± 0.003
2	0.20	4.00	8.00	0.390 ± 0.007
3	15.00	4.00	8.00	1.68 ± 0.03
4	25.00	1.00	2.00	3.52 ± 0.04
5	25.00	1.00	2.00	3.55 ± 0.08
6	25.00	2.00	4.00	3.98 ± 0.18
7	25.00	2.00	4.00	3.96 ± 0.06
8	25.00	4.00	4.00	3.77 ± 0.06
9	25.00	4.00	8.00	4.55 ± 0.04
10	25.00	4.00	8.00	4.73 ± 0.08
11	25.00	7.50	15.0	5.61 ± 0.26
12	25.00	4.00	20.0	5.96 ± 0.19
13	25.00	10.0	20.0	5.91 ± 0.08
14	35.00	4.00	8.00	10.9 ± 0.1
15	35.00	4.00	8.00	11.0 ± 0.3
16	85.00	7.00	0.00	8.92 ± 0.14
17	85.00	7.00	0.00	8.94 ± 0.14

[a] Rate constants are given with probable error.

The reaction is clearly base catalyzed. At 25° in the absence of base, the reaction is immeasurably slow. A temperature of 85° was required in the absence of amine for rates comparable to those at 35° with an amine concentration of 8.00 × 10^{-2} *M*. When the amine to peroxy acid concentration ratios were maintained at 2:1, the first-order rate constant increased with increasing reactant concentration. Since the reaction is apparently first order in the peroxy acid, the increase in rate is associated with increasing amine concentration. This is further supported by comparing the runs where the peroxy acid concentration is maintained constant at 4.00 × 10^{-2} *M* and the amine concentration is increased from 4.00 × 10^{-2} to 8.00 × 10^{-2} M. A noticeable increase in rate is observed under these conditions. Thus, the reaction has not reached zero-order dependence in the amine with amine to peroxy acid concentration ratios of 2:1. In fact, a log–log plot of the rate constant *vs.* amine concentration for these data at 25° gives a slope, corresponding to the order in the amine of 0.23.

Table II. Rate Comparisons of Model Compounds with 2-Methyl-2-*t*-butylperoxypropanoic Acid (VI)

Temp, °C	Compd, C	[C] × 10^2, *M*	[$(C_2H_5)_3N$] × 10^2, *M*	$k \times 10^4$, sec^{-1}
25.00	VI	10.0	20.0	5.91 ± 0.08
25.00	VII	10.0	20.0	Inert[a]
100.00	VIII	10.0	0.00	0.0069 ± 0.0007
100.00	VIII	10.0	20.0	0.0155 ± 0.0005

[a] No change in the titer after 240 hr.

The presence of a phenolic free-radical trapping agent increases the rate somewhat. Under the conditions of run 11, but with 2 mol % (based on VI) 2,6-di-*t*-butyl-*p*-cresol, the first-order rate constant is 6.70 ± 0.06 × 10^{-4} $sec.^{-1}$ If a free-radical chain reaction was operative, a decrease in the rate constant would be expected.

Comparative Rate Data. To distinguish between a concerted fragmentation reaction of the peroxy acid VI and a stepwise process, rate comparisons of VI with a nonperoxy acid of similar structure were desired. For this purpose 2-methyl-2-neopentoxypropanoic acid (VII) was selected as the model compound. The rate data are presented in Table II.

$(CH_3)_3CCH_2OC(CH_3)_2CO_2H$ VII

$(CH_3)_3COOCH(CH_3)_2$ VIII

Furthermore a kinetic comparison between the fragmentation reaction of VI and the corresponding elimination reaction of a peroxide was desired. The most appropriate model compound for the peroxide elimination reaction is *t*-butyl isopropyl peroxide (VIII). Since the rate of base-catalyzed elimination of VIII was quite slow, it was necessary to verify that the reaction was truly base catalyzed. Although the magnitude of base catalysis is not large, a significant rate difference is observed in the runs where base is present and omitted. These data are presented in Table II.

Solvent Effects. The effect of solvent on the rate of base-catalyzed decomposition of VI was investigated in number of nonpolar solvents. Previous linear free energy relationships between rate data and solvent parameters were restricted primarily to polar solvents.[11]

(6) R. P. Bell and A. O. McDougall, *J. Chem. Soc.*, 1697 (1958).
(7) Product studies have been made with a variety of bases.[8]
(8) N. Kornblum and H. E. DeLaMare, *J. Am. Chem. Soc.*, 73, 880 (1951).
(9) (a) R. E. Pincock, *ibid.*, 86, 1820 (1964); (b) R. E. Pincock, *ibid.*, 87, 1274 (1965); (c) R. E. Pincock and T. E. Kiosky, *ibid.*, 87, 2072 (1965); (d) R. E. Pincock and T. E. Kiosky, *ibid.*, 87, 4100 (1965).
(10) "The Sadtler Standard Spectra," Sadtler Research Laboratories, Philadelphia, Pa. 19104, infrared spectrum 1924.
(11) (a) E. M. Kosower, *J. Am. Chem. Soc.*, 80, 3253 (1958); (b) K. Dimroth, C. Reichardt, T. Siepmann, and F. Bohlmann, *Ann.*, 661, 1 (1963); (c) E. Grunwald and S. Winstein, *J. Am. Chem. Soc.*, 70, 846 (1948); (d) S. Winstein, E. Grunwald, and H. W. Jones, *ibid.*, 73, 2700 (1951); (e) A. H. Fainberg and S. Winstein, *ibid.*, 78, 2770 (1956); 79, 1597, 1602, 1608 (1957); (f) S. Winstein, A. H. Fainberg, and E. Grunwald, *ibid.*, 79, 4146 (1957); (g) S. Brownstein, *Can. J. Chem.*, 38, 1590 (1960).

Table III. Rates of Fragmentation of 2-Methyl-2-*t*-butylperoxypropanoic Acid (VI) in Various Solvents[a,b]

Run	Solvent	Temp, °C	$k \times 10^4$, sec^{-1}
18	*n*-Heptane	25.00	1.39 ± 0.04
19	*n*-Heptane	25.00	1.42 ± 0.04
20	*n*-Heptane	35.00	3.36 ± 0.07
21	*n*-Heptane	45.00	7.17 ± 0.09
22	*n*-Heptane	55.00	16.1 ± 0.3
23	Isooctane	25.00	1.43 ± 0.04
24	Cyclohexane	25.00	1.43 ± 0.04
25	*n*-Butyl ether	25.00	0.390 ± 0.006
26	*n*-Butyl ether	35.00	1.48 ± 0.04
27	*n*-Butyl ether	45.00	3.76 ± 0.12
28	*n*-Butyl ether	55.00	7.07 ± 0.14
29	Benzene	25.00	1.88 ± 0.06
9, 10	Chlorobenzene	25.00	4.64 ± 0.09[c]
30	Methylene chloride	25.00	10.7 ± 0.2
31	Ethanol	25.00	23.3 ± 0.4
32	Nitrobenzene	25.00	117.0 ± 6.2

[a] Rate constants are given with probable error. [b] Reactant concentrations are [VI] = 4.00×10^{-2} *M* and [$(C_2H_5)_3N$] = 8.00×10^{-2} *M*. [c] Average of two runs with average error.

with triethylamine are given in Table V. For comparison the rate of fragmentation of VI with triethylamine is also included. Since the peroxy acid was 86% deuterated, the corrected average rate constant for VI-d_1 is 6.87×10^{-4} sec^{-1} at 25°. The magnitude of the deuterium isotope effect is then $k_D/k_H = 1.48$. The change in base from triethylamine to N,N-dimethylbenzylamine decreases the rate by about 50%.

Discussion

The mechanism which most simply explains the data is given below, where the base B: is an amine. The usual steady-state treatment of this mechanism results

$$\underset{\text{VI}}{(CH_3)_3COOC(CH_3)_2CO_2H} + B: \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \underset{\text{VIa}}{(CH_3)_3COOC(CH_3)_2CO_2^-BH^+} \quad (4)$$

$$\text{VIa} \xrightarrow{k_2} (CH_3)_3CO^- + O{=}C(CH_3)_2 + CO_2 + BH^+ \quad (5)$$

Table IV. Activation Parameters for the Decomposition of 2-Methyl-2-*t*-butylperoxypropanoic Acid (VI) in Various Solvents[a]

Solvent	E_a, kcal/mol	Log *A*	$\Delta H^{\pm}$, kcal/mol	$\Delta S^{\pm}$, eu
n-Heptane[b]	15.7 ± 0.2	7.64	15.0 ± 0.2	−25.6 ± 0.7
Chlorobenzene[b]	16.2 ± 0.2	8.50	15.6 ± 0.2	−21.5 ± 0.7
n-Butyl ether[b]	18.8 ± 1.2	9.43	18.2 ± 1.2	−17.5 ± 4.7
Chlorobenzene[c]	36.3 ± 0.3	16.1	35.5 ± 0.3	12.7 ± 1.0

[a] Data are given with probable error. [b] With triethylamine and peroxy acid concentrations of 8.00×10^{-2} and 4.00×10^{-2} *M*. [c] Without triethylamine and with peroxy acid concentration of 7.00×10^{-2} *M*.

However, a measure of the solvent property of typically nonpolar solvents has been recently formulated which is based on the spectral properties of certain merocyanines in these solvents.[12] The transition energy (in kcal/mol) for the particular merocyanine in a given solvent is designated as χ_R and is used as a solvent property indicator.[12] Table III presents the kinetic data for the fragmentation of VI in a number of solvents where the χ_R value is reported. A plot of the kinetic data *vs.* χ_R is given in Figure 1.

Activation Parameters. From the data in Tables I and III, activation parameters were calculated for the fragmentation of VI with triethylamine. These results are presented in Table IV. For comparison the activation parameters for thermal decomposition of VI in the absence of base are included.

Variation in Base and the Isotope Effect. The effect of variation in the basic and acidic sites of the reaction system was investigated. In the previous reactions triethylamine was employed as the base with a pK_a of 10.67 at 25° in water.[13] To determine the effect of varying the pK_a of the amine, the rate of fragmentation was measured with N,N-dimethylbenzylamine (see Table V). The pK_a of this amine is reported to be 8.9 at 25° in water.[14]

The acidic site of the peroxy acid (VI) was modified by deuterium substitution. The rates of fragmentation of 2-methyl-2-*t*-butylperoxypropanoic acid-d_1 (VI-d_1) in rate law 6 where $[VI]_T$ is the total peroxy acid concentration. Since the rate constants were calculated on the basis of a first-order rate of disappearance of VI, $k_{obsd} = k_1k_2[B]/(k_{-1} + k_2 + k_1[B])$. The observed rate constant (k_{obsd}) may be conveniently approximated

$$-d[VI]_T/dt = \frac{k_1k_2[VI]_T[B]}{(k_{-1} + k_2 + k_1[B])} \quad (6)$$

Table V. The Effect of Base Strength and the Deuterium Isotope Effect[a,b]

Acid, concn × 10², *M*	Base, concn × 10², *M*	$k \times 10^4$, sec^{-1}
VI, 4.00	$C_6H_5CH_2N(CH_3)_2$, 8.00	2.35 ± 0.11
VI, 4.00	$(C_2H_5)_3N$, 8.00	4.64 ± 0.09[c]
VI-d_1, 4.00	$(C_2H_5)_3N$, 8.00	6.45 ± 0.10[d]
VI-d_1, 4.00	$(C_2H_5)_3N$, 8.00	6.66 ± 0.07[d]
		Av 6.56 ± 0.11

[a] At 25.00° in chlorobenzene solvent. [b] Rate constants are given with probable error. [c] Average of two measurements, with average error. [d] Rate constants are uncorrected for the per cent deuterium content of VI-d_1.

under two conditions. If $(k_{-1} + k_2) \gg k_1[B]$, the observed rate constant is given by $k_{obsd} \cong k_1[B]/[k_{-1}/(k_2 + 1)]$. In contrast, if $(k_{-1} + k_2) \ll k_1[B]$, then one obtains $k_{obsd} \cong k_2$. The order in amine may then vary from zero to first order according to the above mechanism. With an amine to peroxy acid concentration ratio of 2:1, our data do not neatly fit into either of these

(12) L. G. Brooker, A. C. Craig, D. W. Heseltine, P. W. Jenkins, and L. L. Lincoln, *J. Am. Chem. Soc.*, **87**, 2443 (1965).
(13) W. S. Fyfe, *J. Chem. Soc.*, 1347 (1955).
(14) H. Goldschmidt and R. M. Saleher, *Z. Physik. Chem.*, **29**, 89 (1899).

extremities. Instead, the order in amine is approximately 0.23 at 25° in chlorobenzene.

The question of a free-radical pathway in the reaction of the peroxy acid VI with triethylamine needs to be pursued. A radical chain reaction is eliminated by the observation that the reaction rate is not depressed in the presence of a radical trapping agent, 2,6-di-*t*-butyl-*p*-cresol. The yield of products presents a strong argument against a free-radical reaction as well. Radical decomposition of the peroxy acid VI would no doubt occur by rupture of the peroxide bond to give two alkoxy radicals. The *t*-butoxy radicals will undergo further reaction to give *t*-butyl alcohol by hydrogen atom abstraction and acetone plus methyl radicals by radical fragmentation.[15] This would conflict with the observed quantitative yields of acetone and *t*-butyl alcohol in a 1.00:1.00 ratio. The product studies require that the *t*-butoxy portion of VI be converted quantitatively to *t*-butyl alcohol. Finally, the rate of free-radical decomposition of dialkyl peroxides is relatively insensitive to solvent changes.[16] In contrast, the decomposition of the peroxy acid VI in the presence of triethylamine is quite sensitive to solvent changes (*vide infra*).

Both kinetic and product studies indicate that fragmentation of the carboxylate ion VIa is concerted rather than stepwise. Kinetic arguments were presented previously,[5i] which were based in part on the facile decomposition of the peroxy acid VI compared to the unreactive alkoxy acid VII (see Table II). It was suggested that both acids VI and VII should undergo stepwise decarboxylation to the carbanion at about the same rate. The large difference in the rate of decomposition between VI and VII indicates that VI decomposes by a concerted fragmentation reaction which is not available to VII. Furthermore, some *t*-butyl isopropyl peroxide might be expected to arise from the carbanion intermediate if a stepwise decarboxylation of VIa occurred. Yet *t*-butyl isopropyl peroxide was not detected by glpc analysis. The data in Table II indicate that this peroxide should be reasonably stable under the reaction conditions.

We have presented a mechanism which most simply explains our data. It should be noted that more complex schemes could be formulated which would be accommodated equally well with our data. For example, hydrogen-bonded species of the acid and amine[17] could be interdisposed in equilibrium 4. Also VIa is written as an ion pair[18] for simplicity; however, it could equally well exist as an ion aggregate.

Solvent and Entropy Effects. The effect of solvent on the rate of reaction provides a further test of the proposed mechanism. Exclusive of the most nonpolar solvents, namely cyclohexane, isooctane, and *n*-heptane, the rate increases uniformly with increasing polarity of the solvent as measured by χ_R[12] (see Figure 1). The solvent effect data (Table III) were determined with an

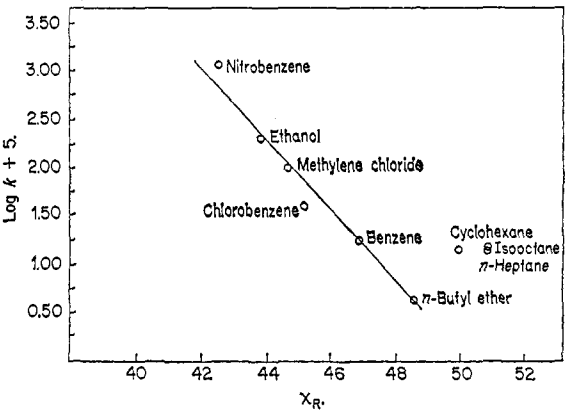
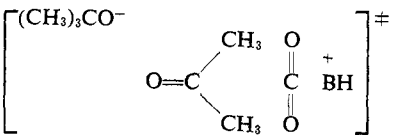


Figure 1. The dependence of the rate of fragmentation of VI with triethylamine at 25° on the χ_R value of the solvent. The slope, excluding the saturated hydrocarbon solvents, is −0.36.

amine to peroxy acid ratio of 2:1. Under these conditions the observed rate constant is still somewhat dependent on the amine concentration so that the approximation $k_{obsd} \cong k_2$ cannot be made. Instead the observed rate constant is dependent on both the equilibrium (4) and fragmentation (5) step. One structure that is expected to contribute to the activated complex of the latter step is given by IX. It is reasonable that VIa is an ion pair (or possibly an ion aggregate) considering the low dielectric constant of all of the solvents employed.[18] In the activated complex for fragmentation, isolated charge development is anticipated as seen by IX. Thereby, the solvent is called upon for charge



IX

stabilization beyond that required for stabilization of the two charged species in close proximity. In other words, an increase in solvation is predicted in traveling along the potential energy surface from the ion pair VIa to the activated complex. From this it is expected that the rate of this step will increase with increasing solvent polarity.[19] Turning now to the question of solvent effects on equilibrium 4, it is evident that increasing solvent polarity should displace equilibrium 4 to the right. Qualitatively, it is reasonable that increasing solvent polarity should increase the rate of reaction due to both the equilibrium and fragmentation step.

As yet few reactions have been correlated with the χ_R constants of Brooker and coworkers.[12] This could prove to be an extremely useful solvent polarity parameter, since few such parameters are applicable to low polarity solvents.[12,21] As seen from Figure 1, an acceptable correlation with the χ_R constant is obtained for all of the solvents employed, except the saturated hydrocarbons. In contrast, the rate data in *n*-heptane solvent were well correlated by χ_R for the amine-catalyzed

(15) For recent reviews see: (a) P. Gray, R. Shaw, and J. C. J. Thynne, *Progr. Reaction Kinetics*, 4, 63 (1967); (b) J. Heicklen, International Oxidation Symposium Abstracts, Aug 28–Sept 1, 1967, San Francisco, Calif., p I-343.

(16) J. H. Raley, F. F. Rust, and W. E. Vaughan, *J. Am. Chem. Soc.*, 70, 88, 1336 (1948).

(17) M. Oki, M. Hirota, and Y. Morimoto, *Bull. Chem. Soc. Japan*, 39, 1620 (1966).

(18) See: (a) H. Van Loog and L. P. Hammett, *J. Am. Chem. Soc.*, 81, 3872 (1959); (b) E. J. Corey, *ibid.*, 75, 1172 (1953).

(19) Although the rate typically increases when solvent polarity is increased for reactions where charge is developed in the activated complex, this is in many instances an entropy rather than a potential energy effect.[20]

(20) A. A. Frost and R. G. Pearson, "Kinetics and Mechanism," 2nd ed, John Wiley & Sons, Inc., New York, N. Y., 1961, pp 137–142.

(21) A. Allerhand and P. von R. Schleyer, *J. Am. Chem. Soc.*, 85, 371 (1963).

decomposition of *t*-butylperoxy formate.[9,12] With 16 solvents, the correlation coefficient was 0.958.[12] We believe that the deviation of our data from the χ_R plot when saturated hydrocarbons are employed is not an artifact.[22] A possible explanation for this deviation is seen by noting that the rate law may be approximated in the two extremities by zero-order dependence on the amine and by first-order dependence. Although a constant amine to peroxy acid concentration ratio of 2:1 was maintained for all solvents, the observed rate may be more or less sensitive to amine concentration as the solvent is varied. The rates would be faster than expected as the reaction approached first-order dependence on the amine. In one instance, the data support this interpretation. The observed rate constant for the decomposition of VI (4.00×10^{-2} *M*) in *n*-heptane with a triethylamine concentration of 4.00×10^{-2} *M* is $0.831 \pm 0.019 \times 10^{-4}$ sec^{-1} at 25°. Doubling the amine concentration (average of runs 18 and 19) gives a rate constant of 1.41×10^{-4} sec^{-1} or increases the rate by a factor of 1.71. In chlorobenzene, a comparable doubling of the amine concentration increases the rate by a factor of only 1.23 (Table I, run 8 and the average of runs 9 and 10). Thus, it appears that the reaction in *n*-heptane more nearly approaches first-order dependence on amine than in chlorobenzene. The result is that the rate constant in *n*-heptane is larger than would be expected if the order in amine remained constant for all solvents. This is consistent with the data in Figure 1.

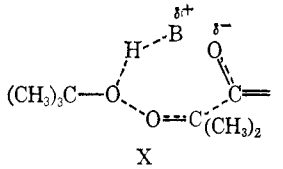
As was indicated previously, few correlations have been made with the χ_R solvent parameter. For this reason it is of interest to compare the response of the rate constant to solvent polarity as measured by χ_R for the reaction reported here and for the decomposition of *t*-butylperoxy formate with amines. The latter reaction was shown to proceed in a concerted manner to give the protonated amine, carbon dioxide, and the *t*-butoxide anion.[9] The slope from the χ_R plot for this reaction, which indicates the sensitivity of the reaction to solvent polarity, is -0.358.[12] Although the over-all reaction of an amine with the peroxy acid VI is proposed to involve a change from two uncharged reactant molecules to a dipolar transition state, similar to the amine–*t*-butylperoxy formate reaction, a prior equilibrium step is interdisposed in this sequence. Yet, the slope in the χ_R plot for the amine–VI reaction is essentially the same (-0.36) as found for the amine–*t*-butylperoxy formate reaction.

Typically reactions which proceed from neutral reactants to a dipolar transition state show an increase in activation energy and entropy of activation with increasing solvent polarity.[9,20] This trend is also observed for the decomposition of the peroxy acid VI with triethylamine in three solvents (Table IV). The activation parameters then appear to reflect the over-all reaction of the neutral amine and acid reactants to a dipolar transition state. The same conclusion was reached from the χ_R plot of the data. The polar character of the reaction of amine with the peroxy acid VI is further emphasized by comparing the activation parameters of this reaction with the thermal and presumably free-radical decomposition of VI in the absence of amine. Both the enthalpy and entropy of activation for the latter reaction are considerably greater in the positive sense than the corresponding values for the amine-catalyzed decomposition.

The large negative entropy values (see Table IV) support the concept that solvent orientation is an important feature in determining the rate of decomposition of VI with amines. The rotation of two bonds will be frozen in the activated complex of reaction 5, namely, the O–O and C_1–C_2 bonds. In the absense of solvent effects this would amount to about -8 eu,[23] which is considerably less than the observed value of -21.5 eu in chlorobenzene. The difference between these two values reflects the increase in solvent orientation in proceeding from the reactant amine and peroxy acid molecules to the transition state.

Base Strength and the Isotope Effect. It is seen from Table V that the rate constant is increased by a factor of about 2 when the weaker base N,N-dimethylbenzylamine is replaced by the stronger base triethylamine. The sensitivity of the rate constant to the basicity of the amine indicates that equilibrium 4 is not displaced completely to the right. The nonzero-order dependence of the reaction on amine concentration is consistent with the influence of base strength on the rate constant. It appears unlikely that the difference in rate with these two bases is due to a solvent effect. N,N-Dimethylbenzylamine is expected to be more polar than triethylamine due to the replacement of an alkyl group with an aromatic group. Considering the χ_R plot, a more polar solvent should increase rather than decrease the rate.

The observed isotope effect of $k_D/k_H = 1.48$ for peroxy acids VI and VI-d_1 clearly rules out a one-step reaction between the amine and the acid to give the products directly. The isotope effect requires an equilibrium step prior to the rate-determining step, which is consistent with our proposed mechanism. As the fragmentation step (5) is written, no isotope effect should be associated with it. An alternative step could be formulated whereby the ion pair VIa directly yields *t*-butyl alcohol, acetone, carbon dioxide, and the amine. The activated complex for this process can be described by X. However, the observed isotope effect is inconsistent with this alternative step, since for this step k_D/k_H should be less than one. The isotope effect must then arise from equilibrium 4.



X

Comparison to Analogous Reactions. The fragmentation reaction of VI with base may be compared to the corresponding decarboxylative fragmentation where an olefinic rather than a carbonyl π system is generated. Since little kinetic data are available, the comparison is crude, but nonetheless it demonstrates that the frag-

(22) The deviation of the rate data in *n*-heptane, isooctane, and cyclohexane solvents from the χ_R plot cannot be explained in terms of a solvent change due to high concentrations of triethylamine. The χ_R value for triethylamine is 49.3. It can be seen from Figure 1 that the use of this value for these solvents will not bring the points onto the line.

(23) H. E. O'Neal and S. W. Benson, *J. Phys. Chem.*, 71, 2903 (1967), and references therein.

mentation of VI is more facile. The rate constant for the decomposition of the sodium salt of *trans*-cinnamic acid dibromide (XI) in refluxing ethanol (78.3°) is reported to be 4.2×10^{-4} sec^{-1} where decarboxylation constitutes 58% of the reaction.[24] In contrast, the rate constant for the triethylamine-catalyzed decomposition of VI in ethanol at 25° is 23.3×10^{-4} sec^{-1}. The rate of fragmentation of VI is clearly more facile than for XI even though the leaving group is bromide in the latter instance *vs.* *t*-butoxide with VI. The difference in rate is most likely due to the development of the π-carbonyl bond of acetone in the decomposition of VI in contrast to the π-olefinic bond in the reaction of XI. The π-bond energies in a carbonyl group and in an olefinic group are approximately 75[25] and 58[26] kcal, respectively. This energy difference is reflected in the activated complexes of the two fragmentation reactions.

Previously the base-catalyzed elimination reaction of peroxides bearing an α-hydrogen atom was reported to proceed according to reaction 3.[6,8] All of the peroxides that were studied possessed an α-phenyl group as well. We now find that an amine-catalyzed elimination reaction occurs in the absence of an α-phenyl substituent. Thus, elimination of *t*-butyl isopropyl peroxide is affected by triethylamine. However, under the conditions employed (see Table II), the amine-catalyzed reaction is only about twice as fast as the thermal decomposition in the absence of amine. The rate data do clearly indicate that the analogous fragmentation reaction of VI is considerably more facile than the elimination reaction.

Experimental Section[27]

Materials. Triethylamine (Matheson Coleman and Bell) was distilled from sodium hydroxide[6] through a 2-ft glass helices column to give a heart cut, bp 88.0–88.5°. Gas–liquid partition chromatography (glpc) showed the fraction to be pure. N,N-Dimethylbenzylamine (Matheson Coleman and Bell) was used without further purification.

Solvents. Chlorobenzene was dried over phosphorus pentoxide and then distilled through a 2-ft glass helices column. A heart cut, bp 130.2–130.8°, was collected and was found to be pure by glpc. The solvent was stored over Drierite and under nitrogen. Purification of *n*-heptane was accomplished by washing with sulfuric acid followed by sodium bicarbonate solution and water. After drying over magnesium sulfate, the *n*-heptane was distilled from calcium hydride, bp 98.0–98.5°. Spectral grade isooctane (Matheson Coleman and Bell) and cyclohexane (Spectrar) were used without further purification. Reagent grade benzene and *n*-butyl ether (Matheson Coleman and Bell) were both distilled from calcium hydride. The cuts boiling at 80.0–80.5° and 142–143°, respectively, were collected. Reagent grade absolute ethanol (U. S. Industrial Chemicals) was used without further purification. Nitrobenzene (Matheson Coleman and Bell) was dried over magnesium sulfate and then distilled, bp 210–212°.

2-Methyl-2-*t*-butylperoxypropanoic Acid (VI). A solution of chromium trioxide (62 g, 0.62 mol), acetic acid (500 ml), and water (59 ml) was stirred and maintained at room temperature with a water bath during the dropwise addition of 2-methyl-2-*t*-butylperoxy-1-propanol[28] (25 g, 0.15 mol). Further stirring was continued for 12 hr at room temperature. The reaction mixture was then diluted with 300 ml of water, extracted with five 35-ml portions of carbon disulfide (reagent grade), and dried over sodium sulfate. Evaporation of the carbon disulfide extract gave 10.9 g (41% yield) of a white solid, mp 61.5–62.8°. The product was sublimed three times at 1 mm (bath 69°), and then dried in a vacuum desiccator over phosphorus pentoxide for 2 days giving 9.9 g (32% yield), mp 62.0–63.5°, of VI. The structure of VI was established by its infrared and nmr spectra. The latter spectrum showed the following absorptions: *t*-butyl protons 1.22 ppm, singlet, area = 9; *gem*-dimethyl protons 1.43 ppm, singlet, area = 6; and carboxylic acid proton 11.6 ppm, singlet, area = 1.

Anal. Calcd for $C_9H_{20}O_2$: C, 67.45; H, 12.57. Found: C, 67.52; H, 12.44.

2-Methyl-2-*t*-butylperoxypropanoic Acid-d_1 (VI-d_1). Deuterium exchange was carried out in three repetitive steps starting with 1.0 g of VI and using 1.0 g of deuterium oxide (99.5% minimum deuterium, Matheson Coleman and Bell) for each exchange reaction. After each step, the water was removed under vacuum. The final product was dried in a vacuum desiccator over phosphorus pentoxide for 2 days. The per cent deuterium content was determined by nmr analysis. The acid and methyl proton areas were determined for the nondeuterated acid VI and compared with these absorptions in the deuterated acid. The per cent deuterium content was found to be 86 ± 1.

2-Methyl-2-neopentoxypropanoic Acid (VII). Oxidation of 2-methyl-2-neopentoxy-1-propanol[28] was carried out by the procedure described for VI, except the reaction mixture was not diluted with water before the carbon disulfide extraction. Acid VII was obtained as white crystals, mp 59.5–62.4°, in 30% yield. After sublimation at 1 mm (bath 63°) and drying in a vacuum desiccator over phosphorus pentoxide for 2 days, the product was obtained in 26% yield, mp 60.0–62.4°. The structure of VII was confirmed by infrared and nmr spectral analysis. The nmr spectrum showed the following absorptions: tertiary methyl protons 0.88 ppm, singlet, area = 9; *gem*-dimethyl protons 1.40 ppm, singlet, area = 6; methylene protons 3.05 ppm, singlet, area = 2; and carboxylic acid proton 12.0 ppm, singlet, area = 1.

Anal. Calcd for $C_9H_{18}O_3$: C, 62.4; H, 10.41. Found: C, 61.71; H, 10.31.

***t*-Butyl Isopropyl Peroxide (VIII).** This peroxide was prepared from potassium *t*-butyl peroxide and isopropyl bromide according to a previously reported procedure,[29] bp 40° (70 mm) (lit.[29] 52° (125 mm)). The structure of VII was confirmed by the infrared and nmr spectrum. The latter spectrum showed the following absorptions: *t*-butyl protons 1.20 ppm, singlet, area = 9; dimethyl protons 1.08 ppm, doublet, area = 6; and α-proton 4.06 ppm, multiplet, area = 1. In addition the parent ion was detected in the mass spectrometer at *m/e* 132.

Product Studies. Quantitative product analyses were performed by glpc using thermal conductivity detection on a 6 ft × 0.25 in., 20% polypropylene glycol on Firebrick (60–80 mesh) column. Four analyses were made after 15 or more half-lives on solutions with initial concentrations of triethylamine and VI at 0.100 *M* in chlorobenzene. The reaction mixture contain 0.100 *M* cyclohexane as an internal standard. Peak areas were integrated with a planimeter. Area ratios of product to internal standard were compared to area ratios of known standards. The condensable products were collected upon elution from a 6 ft × 3/8 in. column packed with silicone fluid XF1150 on Chromosorb W. Infrared spectra of the collected products were compared with authentic spectra.

Kinetic Methods. For the fragmentation reaction, a weighed quantity of the peroxy acid VI was placed in a volumetric flask, and the reaction was begun by adding a thermally equilibrated solution of the amine to the mark. The flask was vigorously shaken for a few seconds to affect a homogeneous solution and placed in a constant temperature bath. The bath temperature was controlled to ±0.01°. Prior to withdrawing an aliquot, the loosely stoppered flask was jarred several times to release any dissolved carbon dioxide. Aliquots were quenched at recorded times by addition to a uniform amount of dilute hydrochloric acid to neutralize the amine. The resulting mixtures were titrated with standardized sodium hydroxide solution to a cresol red end point. Infinity titers were measured after 15 half-lives. From these data, first-order constants

(24) E. Grovenstein, Jr., and D. E. Lee, *J. Am. Chem. Soc.*, **75**, 2639 (1953).

(25) R. Walsh and S. W. Benson, *ibid.*, **88**, 3480 (1966).

(26) Calculation from data given in T. L. Cottrell, "The Strength of Chemical Bonds," Butterworth & Co. (Publishers), Ltd., London, 1954.

(27) All melting points and temperatures of kinetic measurements are corrected. All boiling points are uncorrected. Elemental microanalyses were performed by C. F. Geiger, Ontario, Calif. The nmr spectra were determined in 15% (w/v) carbon tetrachloride solution with tetramethylsilane as an internal standard on a Varian A-60 spectrometer. Infrared spectra were measured with a Perkin-Elmer 337 grating spectrometer. The mass spectrum was determined with a Hitachi RMU-6E spectrometer.

(28) W. H. Richardson and R. S. Smith, *J. Org. Chem.*, **33**, 3882 (1968).

(29) F. F. Rust, F. H. Seubold, Jr., and W. E. Vaughan, *J. Am. Chem. Soc.*, **72**, 338 (1950).

for the disappearance of VI were calculated with a computer program which included a least-squares subroutine. The program for the calculation of activation parameters included a least-squares subroutine as well. A similar titrimetric method was used for the thermal decomposition of VI in chlorobenzene in the absence of amine. In this case the aliquots were quenched by addition to boiled distilled water. The rate of decomposition of *t*-butyl isopropyl peroxide (VIII) was determined by glpc. The reactants were sealed in capillary tubes with the internal standard, cyclohexane. Tubes were periodically removed from the thermostated bath and the glpc area ratios of VIII to cyclohexane were determined. First-order rate constants were calculated from these data.

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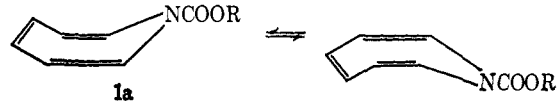
Thermochemical Reactions of 1H-Azepine Derivatives. I. Dimerization[1]

Leo A. Paquette, James H. Barrett, and Donald E. Kuhla[2]

Contribution from the Department of Chemistry, The Ohio State University, Columbus, Ohio 43210. *Received December 23, 1968*

Abstract: Heating of several 1-substituted 1H-azepines at 120–130° for short periods of time has been found to afford crystalline dimers. Classification of these dimerizations as (6 + 4) π cycloadditions was based upon ultimate degradation of the unsubstituted dimers to ethylcyclodecane. The stereochemistry of the dimerization was established as *exo* by X-ray crystallographic structure determination of a derived methiodide. It is suggested that the transition states of such reactions are controlled to a significant extent by orbital symmetry factors since the *exo*-bonding process is favored on the basis of secondary orbital interactions. Thermolysis of the monomeric 1H-azepines or their low-temperature dimers at 200° for several minutes was observed to give rise to different dimers. These high-temperature dimers were shown to be derivatives of the highly symmetrical 13,14-diazatricyclo[6.4.1.1^{2,7}]-tetradeca-3,5,9,11-tetraene system. The mechanistic interrelationship of the two types of dimers is discussed.

From the purely structural viewpoint, the 1H-azepine nucleus (**1**) can be considered to be closely related to cycloheptatriene (**2**) since both molecules contain a fully conjugated triene unit in a seven-membered ring. From the electronic standpoint, however, 1H-azepines with their cyclic array of 8 π electrons are isoelectronic with cyclooctatetraene (**3**).[3] Since the effects of a wide range of temperatures (from −150 to *ca.* 600°) on both of these hydrocarbons have been examined by a variety of techniques, we were interested in the possibility that 1H-azepines may respond in equally interesting fashion to varied thermal energies. We have already commented on the fact that the nmr spectra of a diverse number of 1H-azepine derivatives remain invariant over a substantial temperature range (−90 to +130°).[6a] Therefore, the energy of activation for the conformational ring inversion of nonplanar structures **1a** and **1b**[6] appears to be less than that required in the case of cycloheptatriene (E_a = 5.7–6.3 kcal/mol)[7] or cyclooctatetraene.[8] These nmr studies also provide no evidence for the existence of benzenimine tautomers such as **4**, especially in the



1a **1b**

higher temperature (90–130°) regions. This is particularly significant since it is suggestive that **4** is relatively

1 **4**

less stable than benzene oxide,[9] a number of norcaradienes,[10] and bicyclo[4.2.0]octatetraene[11] relative to their respective monocyclic counterparts.

At somewhat more elevated temperatures, cycloheptatriene and 7-monosubstituted cycloheptatrienes are subject to a series of successive suprafacial 1,5-sigmatropic hydrogen shifts;[12] under the most drastic

(1) Unsaturated Heterocyclic Systems. LI. For previous paper, see I. C. Paul, S. M. Johnson, J. H. Barrett, and L. A. Paquette, *Chem. Commun.*, 6 (1969).

(2) National Institutes of Health Predoctoral Fellow, 1965–1968.

(3) According to a current convention,[4] these heterocycles would be termed π-excessive analogs of cyclooctatetraene.[5]

(4) A. G. Anderson, Jr., W. F. Harrison, and R. G. Anderson, *J. Am. Chem. Soc.*, **85**, 3448 (1963).

(5) For a recent synthesis of the first monocyclic π-equivalent heterocyclic congeners of cyclooctatetraene, see L. A. Paquette and T. Kakihana, *ibid.*, **90**, 3897 (1968); L. A. Paquette and J. C. Philips, *ibid.*, **90**, 3898 (1968).

(6) (a) L. A. Paquette, D. E. Kuhla, J. H. Barrett, and R. J. Haluska, *J. Org. Chem.*, *in* press. (b) Recent X-ray crystallographic studies have provided evidence that 1H-azepines preferably adopt the boat conformation (at least in the crystalline state): I. C. Paul, S. M. Johnson, L. A. Paquette, J. H. Barrett, and R. J. Haluska, *J. Am. Chem. Soc.*,, **90**, 5023 (1968).

(7) (a) F. A. L. Anet, *ibid.*, **86**, 458 (1964); (b) F. R. Jensen and L. A. Smith, *ibid.*, **86**, 956 (1964).

(8) (a) F. A. L. Anet, *ibid.*, **84**, 671 (1962); (b) F. A. L. Anet, A. J. R. Bourn, and Y. S. Lin, *ibid.*, **86**, 3576 (1964).

(9) E. Vogel and H. Günther, *Angew. Chem. Intern. Ed. Engl.*, 6, 385 (1967).

(10) E. Ciganek, *J. Am. Chem. Soc.*, **89**, 1454, 1458 (1967), and references therein for earlier literature.

(11) (a) A. C. Cope, A. C. Haven, Jr., F. L. Ramp, and E. R. Trumbull, *J. Am. Chem. Soc.*, **74**, 4867 (1952); (b) R. Huisgen and F. Mietzsch, *Angew. Chem. Intern. Ed. Engl.*, **3**, 83 (1964); (c) E. Vogel, H. Kiefer, and W. R. Roth, *ibid.*, **3**, 442 (1964).

(12) (a) A. P. ter Borg, H. Kloosterziel, and N. van Meurs, *Proc. Chem. Soc.*, 359 (1962); *Rec. Trav. Chim.*, **82**, 717 (1963), and related work from this group; (b) G. Büchi and E. M. Burgess, *J. Am. Chem.*