Kinetic Applications of Electron Paramagnetic Resonance Spectroscopy. I. Self-Reactions of Diethyl Nitroxide Radicals[1,2]

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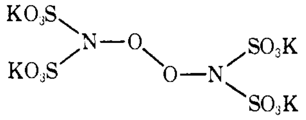
Abstract: The bimolecular rate constants for the self-reaction of diethyl nitroxide radicals have been measured in water, methanol, *n*-propyl alcohol, benzene, CF_2Cl_2, and isopentane by epr spectroscopy. Activation parameters for the reaction were obtained in the last five named solvents. A reaction mechanism involving the reversible formation of a diamagnetic dimer and its slow decomposition to diethylhydroxylamine and a nitrone is proposed. Strong evidence for dimer formation has been obtained at low temperatures in CF_2Cl_2 and in isopentane.

Although the inorganic nitroxide, Fremy's salt, $(KSO_3)_2NO\cdot$, has been known since 1845 and the first organic nitroxide was isolated in 1901, there have been very few quantitative measurements of the reactivity of nitroxides toward one another and toward other compounds.[5] While the electron paramagnetic resonance spectra of several hundred nitroxides have been recorded, the information on their reactivity is usually confined to qualitative statements that a particular radical is stable, moderately stable, or unstable.

Nitroxides have an inherently stable electronic arrangement around the oxygen and nitrogen atoms.[5] No known organic nitroxide shows any tendency to dimerize at the O–N center at room temperature. The stable free-radical bis(trifluoromethyl) nitroxide, which is a purple gas at room temperature, can be condensed to a brown liquid (bp −25°) and solidifies to a yellow diamagnetic solid (mp −70°).[6,7] The yellow solid is presumably a dimer of the radical. A study[6] of the epr spectrum of a 30% solution of this nitroxide in $CFCl_3$ at various temperatures gave a value of −2.5 kcal/mol for the heat of formation of the dimer.

$$2(CF_3)_2NO\cdot \rightleftharpoons \{(CF_3)_2NO\}_2$$

Solid potassium nitrosodisulfonate exists in two distinct crystalline forms.[8] Crystallization from aqueous solution at temperatures below 30° yields the "common" diamagnetic yellow-orange needles of Fremy's salt with the probable structure **1**. Above 30° orange-brown rhombs are obtained which possess approximately one unpaired electron per dimer unit. A structure containing two "nonbonded" monomer units should contain two free electrons per "dimer." There must therefore be a weak covalent bond between the monomer units. X-Ray analysis[9] of this form indicates the structure **2**. In solution, Fremy's salt exists as the free radical $(KSO_3)_2NO\cdot$.



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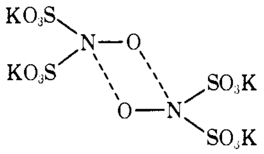
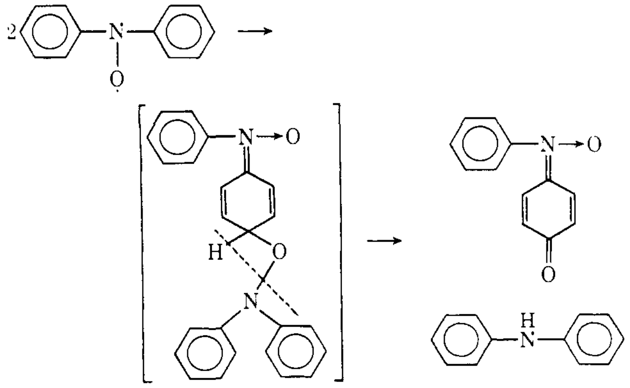


2

Structural factors are important in preventing an organic nitroxide from reacting with itself other than by dimerization at the O–N center. Thus, radicals such as di-*tert*-butyl nitroxide, 2,2,6,6-tetramethyl-4-pyridone nitroxide, and 4,4′-dimethoxydiphenyl nitroxide are stable in the sense that they can be isolated and kept for indefinite times. However, the vast majority of organic nitroxides are unstable because they react with each other through some other center of the molecule. For example, diphenyl nitroxide decomposes to diphenylamine and *N*-phenyl-*p*-benzoquinoneimine *N*-oxide.[10] A similar reaction occurs with *tert*-butyl phenyl nitroxide.[11] Decomposition can also occur by a simple hydrogen atom transfer between two nitroxide radicals, *i.e.*



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(1) Issued as NRCC No. 11804.

(2) A preliminary account of some of this work has appeared: K. Adamic, D. F. Bowman, and K. U. Ingold, *J. Amer. Chem. Soc.*, **92**, 1093 (1970).

(3) NRCC Postdoctorate Fellow, 1967–1969.

(4) NRCC Postdoctorate Fellow, 1969–1970.

(5) A. R. Forrester, J. M. Hay, and R. H. Thomson, "Organic Chemistry of Stable Free Radicals," Academic Press, New York, N. Y., 1968, Chapter 5.

(6) W. D. Blackley and R. R. Reinhard, *J. Amer. Chem. Soc.*, **87**, 802 (1965).

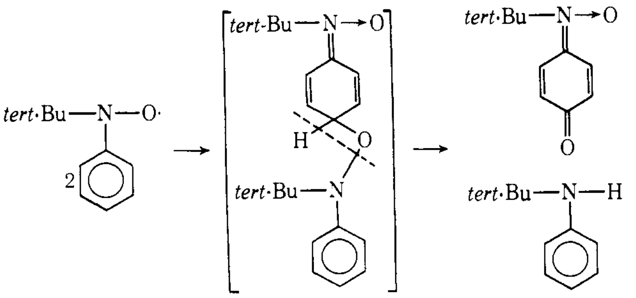
(7) S. P. Makarov, A. Ya. Yakubovich, S. S. Dubov, and A. N. Medvedev, *Dokl. Akad. Nauk SSSR*, **160**, 1319 (1965).

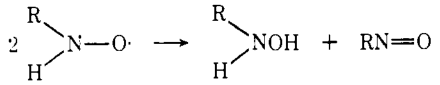
(8) W. Moser and R. A. Howie, *J. Chem. Soc. A*, 3039 (1968).

(9) R. A. Howie, L. S. D. Glasser, and W. Moser, *ibid.*, *A*, 3043 (1968).

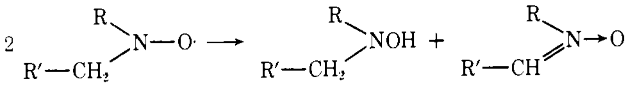
(10) H. Wieland and K. Roth, *Chem. Ber.*, **53**, 210 (1920).

(11) A. R. Forrester and R. H. Thomson, *Nature (London)*, **203**, 74 (1964).





$$2\,\mathrm{R(H)N{-}O\cdot} \longrightarrow \mathrm{R(H)NOH} + \mathrm{RN{=}O}$$



$$2\,\mathrm{R(R'CH_2)N{-}O\cdot} \longrightarrow \mathrm{R(R'CH_2)NOH} + \mathrm{R(R'CH){=}N{\rightarrow}O}$$

Our interest in these reactions was stimulated by our kinetic studies of the self-reactions of the isoelectronic peroxy radicals.[12–15] Since rate constants for the self-reactions of nitroxides had not been measured[16] we decided to undertake a systematic survey of these reactions using epr techniques in the same manner as for our studies of peroxy radicals. The present paper describes a quantitative study of the self-reaction of diethyl nitroxide radicals. Future papers will describe the self-reactions of monoalkyl nitroxides, of dialkyl nitroxides, and of iminoxy radicals (R'RC=NO·).

Experimental Section

Materials. Diethyl nitroxide radicals were prepared from recrystallized diethylhydroxylamine (Aldrich) by oxidation with ceric ammonium nitrate or with *tert*-butoxy radicals. The *tert*-butoxy radicals were generated by photolysis of di-*tert*-butyl peroxide. The commercial peroxide was passed through alumina before use to remove all traces of *tert*-butyl hydroperoxide. At low temperatures the nitroxide radicals were generated by direct photolysis of the hydroxylamine. All solvents were reagent grade materials that were deoxygenated and were then used without further purification.

Radical Formation and Decay. (i) Oxidation with Ceric Ion. In our initial work the hydroxylamine in methanol was oxidized with ceric ammonium nitrate in methanol in a flow system. Two streams containing the hydroxylamine and ceric ion were rapidly and vigorously mixed shortly before flowing through the cavity of a Varian E-3 EPR spectrometer. The two streams of liquid were cut off simultaneously by electrically operated valves and the decay of the radical was monitored on a Tektronix 564 storage oscilloscope which was triggered by the same impulse that operated the shut-off valves.

(ii) Photolytic Generation of Nitroxide Radicals. The above method for generating nitroxide radicals suffers from serious disadvantages since the range of solvents and temperatures that can be studied is very limited. In addition, the stop-flow method is not satisfactory for following very fast radical decays. These disadvantages can be overcome by generating the radicals photolytically. It was found that the nitroxide was best prepared by direct photolysis of $\sim 10^{-2}$ *M* hydroxylamine at temperatures below −60° and by photolysis in the presence of $\sim 10^{-1}$ *M* di-*tert*-butyl peroxide above −60°. The sample was placed in the V-4557 variable-temperature accessory of the spectrometer and the light from a 200-W Osram super-pressure mercury lamp (HBO 200 W) was focussed on the cavity. In some experiments the light was filtered with metal screens of known transmittance and/or with a Corning C.S. No. 7-51 filter (maximum transmittance 365 mμ).

When CF_2Cl_2, which is a gas at room temperature, was the solvent, the sample tubes were sealed *in vacuo* at −190°. However, some problems were encountered because repeated photolysis of the same sample produced a build-up of nitroxide radicals which were more stable than diethyl nitroxide. For this reason, with the other solvents (isopentane, benzene, *n*-propyl alcohol, methanol, and water) the sample tube was normally left open, and a fine quartz capillary led from a reservoir of reactant solution to the bottom of the tube. A Teflon needle valve between the reservoir and capillary and an overflow side arm attached to the sample tube allowed the sample to be replaced after every photolysis.

In the initial experiments the light was cut off by a shutter and a single radical decay was monitored on the oscilloscope. Subsequently, we followed the procedure pioneered by Weiner and Hammond[18] and rotated a sectored disk between the cavity and the light source. At the appropriate rotation speed a large number of identical radical decays were collected by a Fabri-Tek 1072 signal averager.

Radical Concentrations. The second integral of the first-derivative epr signal was normalized to the ~2-kG signal from a single crystal of synthetic ruby in order to correct for variations in cavity Q.[19] The normalized second integral was converted to a radical concentration by means of a calibration curve.[15] This curve gives the temperature dependence of the normalized second integrals for two stable radicals (2,2-diphenyl-1-picrylhydrazyl and 2,2,6,6-tetramethylpyridone nitroxide) over the entire temperature range and for all the solvents. It is only necessary to calibrate any particular sample tube with a known concentration of a stable radical at one temperature in order to be able to convert the normalized second integral for the diethyl nitroxide radical at any temperature to the appropriate radical concentration.

Products of the Oxidation of Diethylhydroxylamine. The initial product is the diethyl nitroxide radical the epr spectrum of which agrees with published spectra[20,21] and consists of 15 lines, three lines of equal intensity from the nitrogen nucleus, each of which is split into a quintet (1:4:6:4:1) by the four equivalent α-hydrogen atoms. In general only 13 of the lines are resolved because of overlap. The splitting constants at room temperature in the solvents employed in this work are listed in Table I. The splitting constants and line widths decrease slightly as the temperature is lowered.

Table I. Hyperfine Coupling Constants (Gauss) of Diethyl Nitroxide in Different Solvents at Room Temperature[a]

Solvent	a_N	a_H
Isopentane	14.9	10.1
CF_2Cl_2	15.0	10.1
Benzene	15.1	10.3
Propanol	15.8	10.9
Methanol	16.0	11.0
Water	16.9	11.7

[a] Calibrated against Fremy's salt with a_N taken to be 13.1 G in 1 *M* Na_2CO_3.

De La Mare and Coppinger[22] have shown that *N,N*-dibenzylhydroxylamine is oxidized with *tert*-butyl hydroperoxide *via* the nitroxide radical to the corresponding nitrone in high yield.[23] Diethyl nitroxide radicals were detected in a similar reaction with *N,N*-diethylhydroxylamine but it was not possible to isolate the nitrone. However, the nitrone was shown to be formed by hydrolyzing it and identifying acetaldehyde as its 2,4-dinitrophenyl hydrazone. In addition, the nitrone was trapped in 75% yield as

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(12) K. U. Ingold, *Accounts Chem. Res.*, **2**, 1 (1969).

(13) J. A. Howard and K. U. Ingold, *Can. J. Chem.*, 47, 3797 (1969).

(14) J. A. Howard, K. Adamic, and K. U. Ingold, *ibid.*, 47, 3793 (1969).

(15) K. Adamic, J. A. Howard, and K. U. Ingold, *Chem. Commun.*, 505 (1969); *Can. J. Chem.*, 47, 3803 (1969).

(16) Cowley and Waters[17] have since reported a bimolecular decay rate constant of $2 \times 10^3\ M^{-1}\,\mathrm{sec}^{-1}$ for dibenzyl nitroxide in 25% aqueous ethanol.

(17) D. J. Cowley and W. A. Waters, *J. Chem. Soc. B*, 96 (1970).

(18) S. Weiner and G. S. Hammond, *J. Amer. Chem. Soc.*, **90**, 1659 (1968).

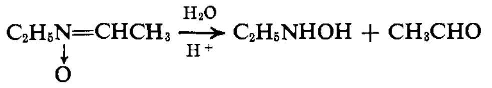
(19) L. S. Singer, *J. Appl. Phys.*, 30, 1463 (1959).

(20) G. M. Coppinger and J. D. Swalen, *J. Amer. Chem. Soc.*, 83, 4900 (1961).

(21) J. Q. Adams, S. N. Nicksic, and J. R. Thomas, *J. Chem. Phys.*, **45**, 654 (1966).

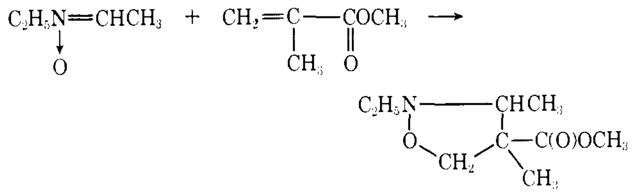
(22) H. E. De La Mare and G. M. Coppinger, *J. Org. Chem.*, 28, 1068 (1963).

(23) See also ref 17.

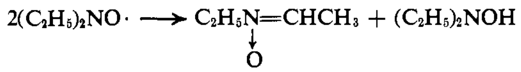


$$C_2H_5\underset{\substack{\downarrow\\ O}}{N}=CHCH_3 \xrightarrow[H^+]{H_2O} C_2H_5NHOH + CH_3CHO$$

an isoxazolidine with methyl methacrylate. It is clear that the



nitrone must be formed by the self-reaction of diethyl nitroxide radicals. The regeneration of diethylhydroxylamine in this re-



$$2(C_2H_5)_2NO\cdot \longrightarrow C_2H_5\underset{\substack{\downarrow\\ O}}{N}=CHCH_3 + (C_2H_5)_2NOH$$

action implies that 2 equiv of the oxidizing agent/mol of hydroxylamine will be required to oxidize the hydroxylamine completely. We have found that at −40° diethylhydroxylamine reacts with 1.98 mol of ceric ion. This oxidation was carried out at the low temperature to prevent hydrolysis of the nitrone and hence additional consumption of ceric by its oxidation of ethylhydroxylamine.

Results

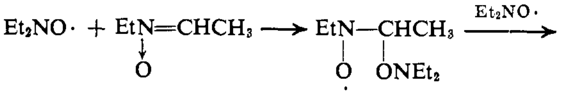
In solution diethylhydroxylamine reacts rather readily with atmospheric oxygen and unless oxygen is rigidly excluded the solution is likely to show the characteristic spectrum of diethyl nitroxide. At room temperature in a sealed tube the nitroxide signal disappears slowly as the oxygen is consumed. If the temperature is lowered the nitroxide signal decreases, disappearing at about −70°. The signal reappears when the temperature is raised and can be repeatedly increased and decreased by raising and lowering the temperature. Although this behavior is not inconsistent with the reversible formation of a diamagnetic nitroxide dimer we can rule out this explanation because the nitroxide signal does slowly disappear in sealed tubes. Moreover, if nitroxide is formed photochemically or with Ce^{IV} it decays very rapidly at room temperature.[24] The results suggest that this initial "impurity" of nitroxide in the hydroxylamine solution is formed by aerial oxidation and that this reaction has a significant temperature coefficient, the rate being very slow at −70°.

$$Et_2NOH \xrightarrow{O_2} Et_2NO\cdot$$

Hydroxylamine Oxidation in Methanol with Ceric Ion. Satisfactory stop-flow kinetic studies by epr of peroxy radicals generated by the Ce^{IV} oxidation of hydroperoxides required high concentrations of ceric ion relative to hydroperoxide in order to ensure that the hydroperoxide was completely converted to peroxy radicals before entering the spectrometer.[25] This procedure is not suitable for diethylhydroxylamine because the hydroxylamine is a product of the reaction of the nitroxide radicals, *i.e.*

$$2Et_2NOH \xrightarrow{Ce^{IV}} 2Et_2NO\cdot \longrightarrow Et_2NOH \xrightarrow{Ce^{IV}} Et_2NO\cdot \longrightarrow \text{etc.}$$

In addition, the nitrone appears to act as a trap for nitroxide radicals so that the epr signal during the decay of high concentrations of diethyl nitroxide changes continuously. A whole series of nitroxides could be formed, *i.e.*



$$Et_2NO\cdot + Et\underset{\substack{\downarrow\\ O}}{N}=CHCH_3 \longrightarrow Et\underset{\substack{|\\ \underset{\cdot}{O}}}{N}-\underset{\substack{|\\ ONEt_2}}{C}HCH_3 \xrightarrow{Et_2NO\cdot}$$

further nitrones and nitroxides

The rate constant for nitroxide decay gradually decreases during a decay presumably because of the increasing stability of the more hindered nitroxides that are formed by this reaction sequence.

These difficulties can be overcome by working with a large excess of diethylhydroxylamine relative to Ce^{IV} and with a low absolute concentration of Ce^{IV}. Excellent second-order decay curves were obtained. The calculated bimolecular rate constants, k_{epr}, were independent of the concentration of ceric ion and of the initial concentration of diethyl nitroxide radicals. The mean value of k_{epr} based on ten experiments was 1.9 ± 0.4 M^{-1} sec^{-1}.

Photoproduction of Diethyl Nitroxide. (i) Methanol. The photolysis of 0.2 *M* diethylhydroxylamine in methanol containing 1.0 *M* di-*tert*-butyl peroxide generated diethyl nitroxide radicals. At 25° the rate constant for the bimolecular decay of these radicals when the light was cut off was 1.5 ± 0.3 M^{-1} sec^{-1} which is in reasonably good agreement with the value found in the ceric ion stop-flow experiments. In addition to the excellent second-order decay curves, confirmation that the decay is bimolecular was obtained from the variation in the steady nitroxide concentration at different light intensities. The reactions taking place are

$$\textit{tert}\text{-BuOO-}\textit{tert}\text{-Bu} \xrightarrow{h\nu} 2\text{-}\textit{tert}\text{-BuO}\cdot$$

$$\textit{tert}\text{-BuO}\cdot + Et_2NOH \longrightarrow \textit{tert}\text{-BuOH} + Et_2NO\cdot$$

$$2Et_2NO\cdot \xrightarrow{k_{epr}} Et_2NOH + EtN(O){=}CHCH_3$$

The steady concentration of nitroxide radicals is given by

$$[Et_2NO\cdot]_0 = \left[\frac{cI[\textit{tert}\text{-BuOO-}\textit{tert}\text{-Bu}]}{k_{epr}}\right]^{1/2}$$

where *I* is the light intensity and *c* is a constant. Experimentally, the steady nitroxide concentration at 25° was found to be proportional to the square root of the light intensity and to the square root of the peroxide concentration. The rate constants decreased with temperature and over a range from 25 to −77° could be described by the expression $k_{epr} = 10^{(4.6\pm0.4)} \exp[(-1900 \pm 500)/RT]\ M^{-1}\ \text{sec}^{-1}$.

(ii) Dichlorodifluoromethane. In sealed tubes this solvent is suitable for studying free radicals in solution from room temperature down to −155°. Second-order rate constants for the decay of photochemically generated diethyl nitroxide radicals from +27 to −137° are given in Table II. The best results were obtained using 2 × 10^{-2} *M* [Et_2NOH] alone at temperatures below ∼−60° and with 2 × 10^{-1} *M* [*tert*-BuOO-*tert*-Bu]

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[24] Even in the presence of a residual nitroxide signal the true second-order rate constant, *k*, for decay of additional radicals generated photochemically can be calculated from the relation

$$\frac{1}{2[NO]_r}\log\frac{[NO]-[NO]_r}{[NO]+[NO]_r} = \frac{2kt}{2.303} + \text{constant}$$

where $[NO]_r$ and [NO] represent the residual and the total nitroxide concentrations, respectively.

[25] J. R. Thomas and K. U. Ingold, *Advan. Chem. Ser.*, 75, 258 (1968).

Table II. Rate Constants for the Decay of Diethyl Nitroxide in CF_2Cl_2

Temp, °C	$[Et_2NO\cdot]_0 \times 10^5$, M	$k_{epr} \times 10^{-3}$, M^{-1} sec^{-1}
−137	0.7	2.7
−137	0.9	3.2
−137	0.06	5.9
−135	0.07	4.6
−117	0.7	4.5
−117	0.07	2.0
−103	0.04	2.0
−102	1.2	2.1
−99	0.8	5.0
−78	1.1	8.6
−50	0.8	15
−39	3.8	12
−19	3.7	13
7	3.7	18
25	4.2	10
26	4.2	12
27	4.2	12
27	4.2	14
27	4.2	15

at temperatures above −60°. The radicals generally disappeared with reasonably clean second-order kinetics, occasional deviations being observed at the beginning or end of the main decay. These deviations are presumably due to the formation of some radicals having stabilities different from diethyl nitroxide. The bimolecular rate constants at any temperature show a fairly wide variation but the calculated values show no trend with the initial radical concentration. The rate constants decrease with temperature and can be approximately represented by the expression $k_{epr} = 10^{(5 \pm 1)} \exp[(-1000 \pm 500)/RT]\ M^{-1}\ \text{sec}^{-1}$.

Diethyl nitroxide radicals can be seen to exist in a reversible equilibrium with a diamagnetic dimer at temperatures in the range −100 to −145°. The existence of the dimer was deduced from the fact that the nitroxide concentration could be increased or decreased by raising or lowering the temperature in this range. The general behavior of the nitroxide radicals is similar to that of tertiary peroxy radicals which, in the same temperature range, exist in equilibrium with a tetroxide.[15,26,27]

It was not possible to form very much of the nitroxide radical-nitroxide dimer equilibrium mixture because the irreversible decomposition of the nitroxide is quite rapid even at these low temperatures (see Table II). That is, high concentrations of nitroxide generated by photolysis decrease by irreversible decay up to the point where this rate is small compared with the length of an equilibrium experiment. It is this "residual" nitroxide concentration, maximum $\sim 1 \times 10^{-6}\ M$, which is in an observable equilibrium with a "residual" concentration of dimer. The equilibrium constant is

$$2(CH_3CH_2)_2NO\cdot \underset{-1}{\overset{1}{\rightleftharpoons}} \{(CH_3CH_2)_2NO\}_2$$

given by

$$K_1 = \frac{k_{-1}}{k_1} = \frac{[(CH_3CH_2)_2NO\cdot]^2}{[\{(CH_3CH_2)_2NO\}_2]} = \frac{2[(CH_3CH_2)_2NO\cdot]^2}{[(CH_3CH_2)_2NO\cdot]_{max} - [(CH_3CH_2)_2NO\cdot]}$$

where $[(CH_3CH_2)_2NO\cdot]_{max}$ is the radical concentration at complete dissociation. The variation in the equilibrium constant with temperature can be represented by the van't Hoff relation

$$K_1 = e^{\Delta S_1/R} e^{-\Delta H_1/RT}$$

where ΔS is the change in entropy and ΔH the change in enthalpy.

The rapid irreversible decay of the radicals prevented the direct determination of K_1 and ΔH_1 and for this reason the following procedure was employed. The radical concentration was determined at a low temperature and the sample was then rapidly warmed or cooled by between 10 and 25°. As soon as the sample reached thermal equilibrium the radical concentration was remeasured and the decay of the radicals with time was then followed. The value that the radical concentration would have had at the second temperature had the temperature change been instantaneous was obtained by extrapolation. This extrapolation is reasonably accurate because the temperature coefficient for irreversible decay is so small that the decay rate constant does not alter significantly over a 10–25° range of temperature. A series of measurements of this kind were made in the temperature range −110 to −145°. For each pair of measurements equilibrium constants were calculated at the two temperatures for various assumed values of ΔH_1. The best fit to all the data over the entire temperature range was obtained with $\Delta H_1 = 6 \pm 3$ kcal/mol and, *via* the van't Hoff equation, $\Delta S_1 = 18 \pm 20$ gibbs/mol. These values are lower than those quoted in our preliminary communication,[2] *viz.*, 8.5 kcal/mol and 32 gibbs/mol.

(iii) Isopentane. At room temperature the value of k_{epr} in isopentane is about twice that found in CF_2Cl_2. However, as the temperature is *decreased* the value of k_{epr} *increases* (see Table III); that is, the overall reaction has a negative temperature coefficient in this solvent. The data given in Table III can be approximately represented by the expression $k_{epr} = 10^{3.3} \exp(930/RT)\ M^{-1}\ \text{sec}^{-1}$.

Table III. Rate Constants for the Decay of Diethyl Nitroxide in Isopentane

Temp, °C	$[Et_2NO\cdot]_0 \times 10^5$, M	$k_{epr} \times 10^{-3}$, M^{-1} sec^{-1}
−112	0.067	180
−112	0.091	95
−70	0.17	71
−70	0.40	55
−70	0.93	46
−59	0.76	57
−48	0.94	53
−20	1.2	45
−19	1.4	42
+22	0.25	36
+22	0.49	32

Although there was evidence for dimer formation in isopentane at low temperatures the decay rate constant

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(26) P. D. Bartlett and G. Guaraldi, *J. Amer. Chem. Soc.*, **89**, 4799 (1967).

(27) J. E. Bennett, D. M. Brown, and B. Mile, *Chem. Commun.*, 504 (1969).

Table IV. Rate Constants for the Decay of Diethyl Nitroxide in Benzene

Temp, °C	Rel light intensity (I)	[*tert*-BuOO-*tert*-Bu], M	[Et_2NOH], M	[$Et_2NO\cdot$]$_0$ × 10^5, M	$\frac{[Et_2NO\cdot]_0}{([tert\text{-BuOO-tert-Bu}]I)^{1/2}} \times 10^5$	$k_{epr} \times 10^{-3}$ M^{-1} sec^{-1}
25	100	1.0	0.1	28	2.8	10.1
25	100	1.0	0.1	28	2.8	9.5
25	30	1.0	0.1	13	2.4	11.5
25	10	1.0	0.1	8.9	2.8	11.6
25	3	1.0	0.1	4.8	2.8	12.8
25	10	0.1	0.01	2.9	2.9	11.7
25	3	0.1	0.01	1.4	2.6	18.0
25	10	0.01	0.001	0.92	2.9	15.2
25	3	0.01	0.001	0.47	2.7	18.7
25	3	0.01	0.001	0.47	2.7	17.1
						Mean = 13.6 ± 3.2
7	1	1.0	0.1	2.7	2.7	12.8
7	33	1.0	0.1	16	2.8	10.0
39	1	1.0	0.1	2.8	2.8	16.6
39	33	1.0	0.1	15	2.6	13.6
75	33	1.0	0.1	15	2.6	18.5

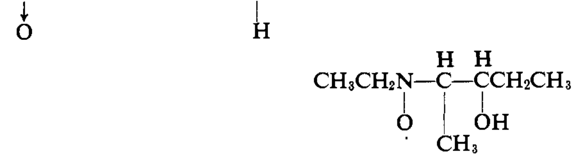
was much too rapid for the heat of formation of the dimer to be determined.

(iv) Benzene. In benzene containing di-*tert*-butyl peroxide and diethylhydroxylamine the steady nitroxide radical concentration is proportional to the square root of the peroxide concentration and to the square root of the light intensity (see Table IV). Once again this implies that the nitroxide radicals are generated *via* photolysis of the peroxide and disappear by reaction with one another. At room temperature the decay rate constant is (13.6 ± 3.2) × 10^3 M^{-1} sec^{-1}. From 7 to 75° the decay rate constant $k_{epr} = 10^{5.2} \exp(-1400/RT)$ M^{-1} sec^{-1}, with an uncertainty of at least 1 kcal/mol in the activation energy.

(v) *n*-Propyl Alcohol. Photolysis of 0.05 M diethylhydroxylamine with 0.5 M di-*tert*-butyl peroxide in a static system in this solvent gave diethyl nitroxide initially. However, this radical was quickly replaced by a 14-line spectrum consistent with an ethyl isopropyl nitroxide (**3**) ($a_N = 13$, $a_{CH_2} = 10$, $a_{CH} = 3$ G). A reasonable reaction sequence would be

$$tert\text{-BuO}\cdot + CH_3CH_2CH_2OH \longrightarrow tert\text{-BuOH} + CH_3CH_2\dot{C}HOH$$

$$2(CH_3CH_2)_2NO\cdot \longrightarrow (CH_3CH_2)_2NOH + CH_3CH_2N(O){=}CHCH_3$$



$$CH_3CH_2\underset{\downarrow \atop O}{N}{=}CHCH_3 + CH_3CH_2\underset{| \atop H}{\dot{C}}OH \longrightarrow CH_3CH_2\underset{| \atop \dot{O}}{N}-\overset{H}{\underset{| \atop CH_3}{C}}-\overset{H}{\underset{| \atop OH}{C}}CH_2CH_3$$

3

The new radical was more stable than diethyl nitroxide.

Photolysis in the absence of the peroxide yielded diethyl nitroxide as the principal radical which decayed with clean second-order kinetics for 70% or more reaction. A small residual signal due to radical **3** did, however, remain after decay of the diethyl nitroxide. Since the signal from radical **3** was narrower than that from diethyl nitroxide it was possible to follow the decay of the latter without interference from the former by monitoring the outermost line in the diethyl nitroxide spectrum. The amount of radical **3** is too small to affect the kinetics of decay or the measured rate constant.

Measurements over a temperature range from 97 to −85° suggest that there is a significant decrease in k_{epr} as the temperature is lowered (Table V). The rate constants can be correlated by $k_{epr} = 10^{5.4} \exp(-2100/RT)$ M^{-1} sec^{-1} with an estimated uncertainty of about 0.5 kcal/mol in the activation energy.

Table V. Rate Constants for the Decay of Diethyl Nitroxide in *n*-Propyl Alcohol (Et_2NOH = 0.1 M)

Temp, °C	[$Et_2NO\cdot$]$_0$ × 10^5, M	$k_{epr} \times 10^{-3}$, M^{-1} sec^{-1}
−85	1.3	1.03
−67	0.8	1.80
−65	1.2	1.40
−53	1.2	1.92
−42	1.2	1.90
−31	1.2	2.12
−21	1.2	2.48
0	1.1	2.30
23	8.2	7.70[a]
~80	10.7	13.2[a]
97 (bp)	8.1	12.8[a]

[a] Slow flow through cavity to prevent build-up of radical **3**.

In this solvent positive evidence for the reversible formation of a nitroxide dimer at low temperatures could not be obtained. On rapidly warming a nitroxide sample from about −95 to −40° there were small transient changes in the signal intensity (a sharp decrease, followed by a sharp increase, and then a slower decrease). However, within 2 min the signal had returned to its initial intensity, and since at least 2 min are required for thermal equilibration these changes are more probably due to changes in cavity Q than to changes in nitroxide concentration.

(vi) Water. A 0.5 M solution of diethylhydroxylamine in water was photolyzed directly at room temperature. Second-order rate constants for decay of the nitroxide radicals of 8.0 × 10^2 M^{-1} sec^{-1} and 7.9 × 10^2 M^{-1} sec^{-1} were obtained at initial nitroxide concentrations of 6.5 × 10^{-5} M and 4.9 × 10^{-5} M.

Discussion

The rate constants and activation parameters for the decay of diethyl nitroxide in the various solvents used are summarized in Table VI. The rate constants are

Table VI. Summary of Rate Constants and Activation Parameters for the Bimolecular Decay of Diethyl Nitroxide in Various Solvents

Solvent	$k_{epr} \times 10^{-3}$ at 25°, M^{-1} sec^{-1}	Log A_{epr}, M^{-1} sec^{-1}	E_{epr}, kcal/mol
Water	0.8		
Methanol (Ce^{IV})	1.9 ± 0.4		
Methanol ($h\nu$)	1.5 ± 0.3	4.6 ± 0.4	1.9 ± 0.5
n-Propyl alcohol	7.2 ± 2	5.4 ± 0.4	2.1 ± 0.5
Benzene	13.6 ± 3	5.2 ± 0.7	1.4 ± 1.0
CF_2Cl_2	15.5 ± 5	5.0 ± 1.0	1.0 ± 0.5
Isopentane	33 ± 10	3.8 ± 1.0	−0.93 ± 0.5

very much lower than the value of ~10^9 M^{-1} sec^{-1} which is found for the bimolecular self-reaction of most radicals in solution. The low rate constants arise from the small preexponential factor in the Arrhenius equations rather than from a large activation energy for the overall reaction.

The nitroxide dimer identified at low temperatures in CF_2Cl_2 and in isopentane might have no direct bearing on the overall reaction, *i.e.*

$$\{(CH_3CH_2)_2NO\}_2 \underset{-1}{\overset{1}{\rightleftharpoons}} 2(CH_3CH_2)_2NO\cdot \xrightarrow{2} (CH_3CH_2)_2NOH + CH_3CH{=}N(O)CH_2CH_3$$

However, it is difficult to believe that reaction 2 could occur almost without activation energy and with a preexponential factor as low as 10^4–10^5 M^{-1} sec^{-1}. We therefore prefer to represent the reaction as

$$2(CH_3CH_2)_2NO\cdot \underset{-1}{\overset{1}{\rightleftharpoons}} \{(CH_3CH_2)_2NO\} \xrightarrow{3} (CH_3CH_2)_2NOH + CH_3CH{=}N(O)CH_2CH_3$$

Support for a mechanism of this type comes from the negative temperature coefficient for the decay in isopentane. A negative activation energy for a one-step reaction in solution is not possible. A negative temperature coefficient indicates at least a two-step process, the first step of which is reversible.

The measured decay rate constant k_{epr} is equal to k_3/K_1. In CF_2Cl_2 the data gave

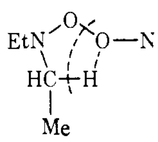
$$k_{epr} = 10^5 \exp(-1000/RT)\ M^{-1}\ \text{sec}^{-1}$$

$$K_1 = 10^4 \exp(-6000/RT)\ M$$

and hence

$$k_3 = 10^9 \exp(-7000/RT)\ \text{sec}^{-1}$$

The preexponential factor for reaction 3 is lower than is generally found for any type of unimolecular reaction in the gas phase. The lowest preexponential factors for gas-phase unimolecular decomposition are ~10^{11}–10^{12} sec^{-1} and are found for complex fissions proceeding by five- or six-center transition states.[28] Four possible cyclic transition states can be envisioned for reaction 3 depending on whether the dimer is formed by O–O coupling (**4**), or by O–N coupling (**5**,**6**), or by N–N coupling (**7**). Assuming, by analogy with Fremy's salt,[8,9] that the diethyl nitroxide dimer is formed by head–head coupling through the two oxygen atoms, then structure **4** represents the most probable transition state for reaction 3.



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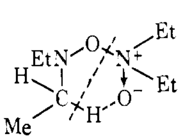

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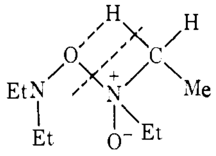

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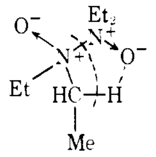

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The loss in entropy for the dimer which accompanies the formation of the cyclic transition state, **4** (*i.e.*, $\Delta S_3{}^{\ddagger}$) will be that due to the loss from three hindered internal rotors plus a contribution of $R \ln K_\sigma{}^{\ddagger}$ to allow for symmetry and optical isomerism. At 500°K or higher the entropy loss per rotor would be ~3.6 gibbs/mol but at ~200°K the ring torsions do not compensate the rotations as much as they do at higher temperatures and the entropy loss per rotor is about 4.2 gibbs/mol.[29] The value of $K_\sigma{}^{\ddagger}$ is somewhat uncertain because the exact nature of the transition state is unknown. Thus, if one assumes a flat, five-center transition state, as drawn above, then there are two optical centers, one at the ring nitrogen and the other at the ring carbon atom.[29] Of the four isomers, two are cis and the other two trans. If the cis–trans isomers had the same energy then $K_\sigma{}^{\ddagger} = 8$, but if one is favored by as little as 0.5 kcal/mol at 200°K, then $K_\sigma{}^{\ddagger}$ is closer to 4. On the other hand, if the ring hydrogen atom is out of the plane of the ring there are eight species since each of the four preceding species can be cis or trans with respect to the out-of-plane H atom. Depending on the energy differences between the possible cis–trans configurations $K_\sigma{}^{\ddagger}$ can be between 4 and 16.[29] Since we have no assurance that the transition state is represented by structure **4** rather than by structure **5**, **6**, or **7** it seems best to take $K_\sigma{}^{\ddagger} = 8$ with a possible uncertainty of a factor of 2. Hence, $\Delta S^{\ddagger}{}_3 = 3(-4.2) + R \ln 8 \pm R \ln 2 = -12.6 + 4.2 \pm 1.4 = -8.4 \pm 1.4$ gibbs/mol. The preexponential factor for the overall decay is given by

$$A_{epr} = \frac{ekT}{h}\frac{e^{\Delta S^{\ddagger}{}_3/R}}{e^{\Delta S_1/R}}$$

The entropy change for dimer formation ΔS_1 was found to be 18 ± 20 gibbs/mol in CF_2Cl_2. Therefore, in CF_2Cl_2 at the mean temperature of the experiments (220°K)

$$A_{epr} = \frac{ekT}{h}\frac{10^{-8.4/2.3R}}{10^{18/2.3R}} = 10^{7.3}\ M^{-1}\ \text{sec}^{-1}$$

Although this calculated preexponential factor is rather inaccurate because of the absence of any correction for solvent effects on $\Delta S^{\ddagger}{}_3$ and because of the uncertainties introduced by the large limits of error in ΔS_1, it is in reasonable agreement with the value of $10^{5.0 \pm 1.0}$ M^{-1} sec^{-1} found experimentally.

The results obtained in isopentane, CF_2Cl_2, benzene, *n*-propyl alcohol, and methanol suggest that the decrease in k_{epr} at 25° along this series is due to an increase in the overall activation energy, E_{epr}, rather than

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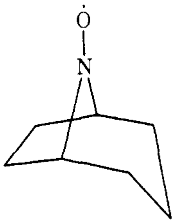
[28] S. W. Benson, "Thermochemical Kinetics," Wiley, New York, N. Y., 1968.

[29] S. W. Benson, private communication.

to any change in the overall preexponential factor, A_{epr}. Unfortunately, it proved impossible to measure the equilibrium constant for dimer formation in any solvent other than CF_2Cl_2 and therefore it was not possible to evaluate separately K_1 and k_3 in any other solvent. However, it seems likely that the *primary* cause of the decrease in k_{epr} in the more polar solvents is due to a decrease in the concentration of dimer rather than to any change in k_3. That is, in the more polar solvents (H_2O, alcohol, etc.) the heat of solvation of the nitroxide radicals will be greater than the heat of solvation of the dimer since the polar N–O group will be more exposed. This effect may be counterbalanced to some extent by an entropy factor if strong solvation of the nitroxide radicals results in an increase in the degree of ordering of the solvent molecules. Provided the entropy effect does not override the effect due to the heat of solvation, (*i.e.*, provided the temperature is not too high) the equilibrium constant K_1 should shift so as to favor the free radical as one goes along the solvent series from isopentane to water. Some support for this conclusion comes from measurements of the equilibrium constant for dimer formation for the unhindered nitroxide radical, nortropane-*N*-oxyl (**8**). This radical is stable at room temperature in neutral solution.[30] We have found that the epr signal due to this radical in solution can be reversibly decreased and increased by cooling and warming. This implies that **8** can form a diamagnetic dimer at low temperatures. Since **8** does not decompose, the equilibrium constant could be measured fairly accurately in some solvents. In isopentane, $\Delta H = 7.2$ kcal/mol and $\Delta S = 14.8$ gibbs/mol which give $K_{300°\mathrm{K}} = 10^{-2.0}$ and $K_{200°\mathrm{K}} = 10^{-4.6}$ *M*. In CF_2Cl_2, $\Delta H = 2.9$ kcal/mol and $\Delta S = -0.2$ gibbs/mol which give $K_{300°\mathrm{K}} = 10^{-2.2}$ and $K_{200°\mathrm{K}} = 10^{-3.2}$ *M*. Thus, the free-nitroxide radical is relatively more abundant below room temperature in CF_2Cl_2 than in isopentane. This is so in spite of the large entropy effect operating so as to favor radical formation in isopentane. We attribute the near zero entropy change for the equilibrium in CF_2Cl_2 to a relatively higher degree of ordering of the solvent molecules around the radicals than around the dimer. Unfortunately the equilibrium constant could not be measured for propanol becuase there was no significant dimerization down to $-100°$. The epr signal due to the radical did decrease at still lower temperatures but it became asymmetric which suggests that **8** may have crystallized out of the propanol. The equilibrium constant could not be measured in water because of the low solubility of **8**, nor in benzene because of the high freezing point of this solvent. In toluene ΔH was found to be ~3.8 kcal/mol but the relatively limited temperature range available precluded the determination of ΔS.



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Electron Spin Resonance Study on Intermediate Free Radicals in the Addition of Thiols to Unsaturated Compounds

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Abstract: Esr spectra of radicals of the type $RSCHX\dot{C}HY$ have been observed in a flow system by running a Ti^{3+}–H_2O_2 redox reaction in the presence of olefinic compounds or propargyl alcohol and thiols. The results show that the thiyl radical adducts observed in the present study have open-chain classical structures. Conformations of the thiyl radical adducts are discussed based on β-proton hyperfine coupling constants.

Thiols add to olefins and acetylenes *via* a free-radical chain mechanism to form sulfides.[1] The postulated reaction mechanism involves a free-radical species RS–A, in which A represents a unit of composition corresponding to that of the olefin or the acetylene employed. The stereospecificity of homolytic thiol addition reactions has been studied and, in some cases, preferential trans addition reactions have been reported.[2–7] To account for the observation[2] that methyl deuteriomercaptan undergoes photoinitiated addition in a stereospecific trans manner to *cis*- and *trans*-2-butenes in the presence of deuterium bromide, sulfur-bridged radicals are proposed as reaction intermediates.[4] The major product from radical addition of methyl mer-

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