The hydrogen bond strength in bulk methanol, $\Delta H_1$, is about 5 kcal/mol.[32] The activation energy for reaction 3 is probably $\sim$2.5 kcal/mol to judge from the results with isopropyl nitroxide and *tert*-butyl nitroxide in benzene. Hence, the negative activation energy for the overall process can be accounted for if the strength of the hydrogen bond formed between the nitroxide and methanol, $\Delta H_2$, has a value of $\sim$2 kcal/mol, which seems not unreasonable.

The preexponential factor for the overall reaction is

$$A_{\mathrm{epr}} = 10^{3.1} = A_3 e^{2(\Delta S_2 - \Delta S_1)/R}\ M^{-1}\ \mathrm{sec}^{-1}$$

Since the preexponential factor for reaction 3 is probably $\sim 10^{8.8}\ M^{-1}\ \mathrm{sec}^{-1}$ (Table I), the experimental data indicate that $\Delta S_2 - \Delta S_1 \approx -13$ gibbs/mol. This result is consistent with the general observation that entropy differences accompanying reversible associations decrease as the size of the species involved increases.

While the above picture of the mechanism of the self-reaction of methyl nitroxide in methanol is undoubtedly oversimplified, it does account qualitatively for the observed results. Furthermore, it suggests that similar unusual Arrhenius parameters may be found for other bimolecular reactions in hydrogen bonding media, provided the reactants can form a weak hydrogen bond to give unreactive species, whereas the nonbonded reactants can react together rapidly and with little activation energy.

**Thermochemistry of Nitroxide Self-Reactions.** Benson and coworkers[33] have recently summarized the heats of formation of a number of compounds of interest in connection with the present work. Thus, $\Delta H_f°(\mathrm{CH_3NHOH}) = -12$ kcal/mol, $\Delta H_f°(\mathrm{CH_3NO}) = 16$ kcal/mol, and since $\Delta H_f°(\mathrm{H}\cdot) = 52$ kcal/mol the overall reaction

$$\mathrm{CH_3NHOH} \longrightarrow \mathrm{CH_3NO} + \mathrm{H}\cdot + \mathrm{H}\cdot$$

is endothermic by $12 + 16 + 2 \times 52 = 132$ kcal/mol in the gas phase. If we make the reasonable assumption that the O–H bond strength in methylhydroxylamine, D(CH$_3$NHO–H), is the same as that in 2,2,6,6-tetramethyl-*N*-hydroxypiperid-4-one (*viz.*, 71 kcal/mol[34]) it follows that $\mathrm{D(CH_3N(\dot{O})\text{–}H)} = 132 - 71 = 61$ kcal/mol. The reaction

$$2\mathrm{CH_3NHO}\cdot \longrightarrow \mathrm{CH_3NHOH} + \mathrm{CH_3NO}$$

is therefore exothermic by 10 kcal/2 mol of radical in the gas phase. This value is in excellent agreement with the value of 10 kcal/mol found in this work for the analogous self-reaction of *tert*-butyl nitroxide in benzene. The low activation energies found for these self-reactions are consistent with exothermic processes. The lower rate constant for phenyl nitroxide compared with the alkyl nitroxides is probably related to the lower exothermicity found for this reaction, *viz.*, 7.2 kcal.

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(32) G. C. Pimentel and A. L. McClellan, "The Hydrogen Bond," W. H. Freeman, San Francisco, Calif., 1960.

(33) S. W. Benson, F. R. Cruickshank, D. M. Golden, G. R. Haugen, H. E. O'Neal, A. S. Rodgers, R. Shaw, and R. Walsh, *Chem. Rev.*, **69**, 279 (1969).

(34) L. R. Mahoney, private communication.

# Kinetic Applications of Electron Paramagnetic Resonance Spectroscopy. III. Self-Reactions of Dialkyl Nitroxide Radicals[1,2]

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**Abstract:** Bimolecular rate constants and activation parameters have been measured for the self-reactions of several dialkyl nitroxides in a number of solvents. It has been shown that there is a large deuterium isotope effect on the reaction if the hydrogens on the $\alpha$-carbons are replaced by deuterium. The results are consistent with a mechanism involving the reversible formation of a diamagnetic nitroxide dimer and its slow decomposition to a hydroxylamine and a nitrone. The equilibrium constant for dimer formation has been determined with $\mathrm{(CH_3CD_2)_2NO}\cdot$ in isopentane and in $\mathrm{CF_2Cl_2}$. The analogy between the self-reactions of nitroxides and of peroxy radicals is pointed out.

The first paper in this series[5] described a study of the self-reaction of diethyl nitroxide in solution. The second-order rate constants for decay of the radicals were measured by epr spectroscopy over a range of temperature in a variety of solvents. Studies at low temperatures (−100 to −150°) indicated that the nitroxide radicals exist in equilibrium with a diamagnetic dimer

$$2\mathrm{Et_2NO}\cdot \rightleftharpoons \mathrm{(Et_2NO)_2}$$

Dimer was identified in isopentane and in $\mathrm{CF_2Cl_2}$ and, in the latter solvent, the irreversible decay of the radicals was just slow enough to obtain a very rough estimate of the equilibrium constant and the heat and en-

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(1) Issued as N.R.C.C. No. 12295.

(2) Part II: D. F. Bowman, J. L. Brokenshire, T. Gillan, and K. U. Ingold, *J. Amer. Chem. Soc.*, **93**, 6551 (1971).

(3) N.R.C.C. Fellow, 1967–1969.

(4) N.R.C.C. Fellow, 1969–1970.

(5) K. Adamic, D. F. Bowman, T. Gillan, and K. U. Ingold, *J. Amer. Chem. Soc.*, **93**, 902 (1971).

Table I. Bimolecular Rate Constants and Activation Parameters for Dialkyl Nitroxide Radical Decays

| Nitroxide | Solvent | $[NO\cdot]_0 \times 10^6$,[a] $M$ | $(k_{epr})_{25°} \times 10^{-3}$, $M^{-1}\ sec^{-1}$ | Temp range, °C | Log $A_{epr}$, $M^{-1}\ sec^{-1}$ | $E_{epr}$, kcal/mol |
|---|---|---|---|---|---|---|
| $\overline{(CH_2)_5\text{—}N}O\cdot$ | Isopentane | 1.0–4.0 | 700 ± 100 | 25 to −72 | 5.3 | −0.75 ± 0.4 |
| $\overline{(CH_2)_5\text{—}N}O\cdot$ | $CF_2Cl_2$ | 1.0–6.0 | 370 ± 100 | 25 to −133 | 6.0 | 0.60 ± 0.4 |
| $\overline{(CH_2)_5\text{—}N}O\cdot$ | Benzene | 3.0–50 | 120 ± 20 | | | |
| $\overline{(CH_2)_4\text{—}N}O\cdot$ | Isopentane | 3.0–20 | 180 ± 20 | 25 to −54 | 5.2 | −0.1 ± 0.3 |
| $\overline{(CH_2)_4\text{—}N}O\cdot$ | $CF_2Cl_2$ | 2.0–9.0 | 110 ± 20 | 25 to −49 | 5.0 | 0.0 ± 0.3 |
| $\overline{(CH_2)_4\text{—}N}O\cdot$ | Benzene | 1.0–90 | 39 ± 5 | | | |
| $\overline{CD_2CH_2CH_2CD_2\text{—}N}O\cdot$ | Benzene | 9.0–50 | 2.9 ± 0.2 | | | |
| $(PhCH_2)_2NO\cdot$ | Isopentane | 7.0–20 | 41 ± 10 | 25 to −51 | 3.9 | −1.0 ± 0.5 |
| $(PhCH_2)_2NO\cdot$ | Benzene | 4.0–100 | 9.1 ± 1.0 | | | |
| $(i\text{-}Pr)_2NO\cdot$ | $CF_2Cl_2$ | 50 | 0.012 ± 0.002 | 25 to −22 | 7.9 | 9.4 ± 2.0 |
| $(i\text{-}Pr)_2NO\cdot$ | Benzene | 100–1000 | 0.0041 ± 0.0008 | 25 to 71 | 4.7 | 5.5 ± 1.0 |
| $Me_2NO\cdot$ | $CF_2Cl_2$ | 1.0–4.0 | 47 ± 10 | 25 to −115 | 5.1 | 0.6 ± 0.5 |
| $Me_2NO\cdot$ | Benzene | 8.0–230 | 21 ± 3 | | | |
| $(CH_3CD_2)_2NO\cdot$ | Isopentane | 50 | 5.5 ± 1 | 25 to −114 | 4.2 | 0.6 ± 0.5 |
| $(CH_3CD_2)_2NO\cdot$ | $CF_2Cl_2$ | 45 | 1.1 ± 0.2 | 25 to −119 | 4.9 | 2.5 ± 0.8 |
| $Et_2NO\cdot$[b] | Isopentane | | 33 ± 10 | | 3.8 | −0.93 ± 0.5 |
| $Et_2NO\cdot$[b] | $CF_2Cl_2$ | | 15.5 ± 5 | | 5.0 | 1.0 ± 0.5 |
| $Et_2NO\cdot$[b] | Benzene | | 13.6 ± 3 | | 5.2 | 1.4 ± 1.0 |
| $Et_2NO\cdot$[b] | *n*-Propanol | | 7.2 ± 2 | | 5.4 | 2.1 ± 0.5 |
| $Et_2NO\cdot$[b] | Methanol | | 1.5 ± 0.3 | | 4.6 | 1.9 ± 0.5 |
| $Et_2NO\cdot$[b] | Water | | 0.8 | | | |

[a] Range in initial nitroxide concentration at 25°. [b] Reference 5.

tropy of formation of the dimer. The activation parameters for the overall reaction suggested that the decay of the nitroxide radicals was due to an irreversible decomposition of the dimer to diethylhydroxylamine and a nitrone

$$(Et_2NO)_2 \longrightarrow Et_2NOH + CH_3CH{=}N({\rightarrow}O)Et$$

It was suggested that this reaction involved a cyclic transition state.

The present work describes an analogous, but less detailed, epr study of the self-reactions of dimethyl nitroxide, diisopropyl nitroxide, dibenzyl nitroxide, piperidine nitroxide, and pyrrolidine nitroxide. In addition, the effects of deuteration at the $\alpha$-carbon on the rates of the self-reactions of diethyl nitroxide and pyrrolidine nitroxide have been measured. The isotope effect was large enough, *i.e.*, radical decay was slow enough, that it was possible to measure the equilibrium constant for dimerization of $(CH_3CD_2)_2NO\cdot$ with considerably more accuracy than our earlier measurement[5] of the equilibrium constant for dimerization of $(CH_3CH_2)_2NO\cdot$.

## Experimental Section

The nitroxide radicals were generated from the corresponding *N*,*N*-dialkylhydroxylamines either by direct photolysis or by photolysis in the presence of di-*tert*-butyl peroxide. The experimental procedures for measuring radical concentrations and decay rate constants have been described previously.[2,5]

**Materials.** *N*-**Hydroxypiperidine** was obtained from Aldrich. *N*,*N*-Dimethylhydroxylamine hydrochloride was prepared by the method of Hepworth.[6] The free base was displaced from the hydrochloride with liquid ammonia.[7] *N*,*N*-**Diisopropylhydroxylamine** and *N*,*N*-**dibenzylhydroxylamine** were prepared by Hepworth's method except that the free bases were recovered by ether extraction rather than by steam distillation.

*N*-**Hydroxypyrrolidine** was prepared from *N*-**hydroxysuccinimid** by reduction with lithium aluminum hydride. The *N*-**hydroxy succinimide** was prepared from succinic anhydride according to the procedure of Wang, *et al.*,[8] and was reduced according to the procedure of Zinner and Moll.[9] Reduction with lithium aluminum deuteride yielded **2,2,5,5-tetradeuterio-*N*-hydroxypyrrolidine** with a deuterium content in the 2 and 5 positions of >98%.

**$\alpha,\alpha,\alpha',\alpha'$-Tetradeuteriodiethylhydroxylamine** was prepared by the lithium aluminum deuteride reduction of both triacetylhydroxylamine and *N*,*N*-diacetylhydroxylamine. Triacetylhydroxylamine was prepared from hydroxylamine hydrochloride and acetic anhydride by the procedure of Urbanski.[10] The reduction of this compound was carried out in anhydrous ether under reflux for 7 days. The excess lithium aluminum deuteride was destroyed with an aqueous solution of sodium hydroxide and, after removing the water layer, the ethereal solution of the free base was dried and treated with HCl gas and the ether was then removed under vacuum. The tetradeuteriodiethylhydroxylamine was displaced from the hydrochloride with liquid ammonia.[7] *N*,*N*-Diacetyl-*O*-benzylhydroxylamine was prepared from benzyloxyamine hydrochloride and acetic anhydride according to the procedure of Ames and Gray.[11] An ethereal solution of this compound was hydrogenated over 5% Pd/C at room temperature and pressure to yield *N*,*N*-diacetylhydroxylamine. This compound was reduced with lithium aluminum deuteride following the same procedure as for the triacetylhydroxylamine. Both samples of tetradeuteriodiethylhydroxylamine had a deuterium content in the $\alpha$ positions >98%.

All hydroxylamines were carefully purified by distillation and/or crystallization. The solvents employed were deoxygenated before use. Radical decays were monitored only on "clean" spectra identical with, or at least similar to, those reported by other workers.

## Results

**Kinetics.** All the nitroxides studied in the present work decayed with clean second-order kinetics provided the usual precautions were taken.[2,5] Table I lists the

$$2R_2NO\cdot \xrightarrow{k_{epr}} \text{nonradical products}$$

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(6) H. Hepworth, *J. Chem. Soc.*, **119**, 255 (1921).

(7) T. C. Bissot, R. W. Parry, and D. H. Campbell, *J. Amer. Chem. Soc.*, **79**, 796 (1957).

(8) K.-T. Wang, D. N. Brattesani, and B. Weinstein, *J. Heterocycl. Chem.*, **3**, 98 (1966).

(9) G. Zinner and R. Moll, *Chem. Ber.*, **99**, 1292 (1966).

(10) T. Urbanski, *J. Chem. Soc.*, 3374 (1949).

(11) D. E. Ames and T. F. Gray, *ibid.*, 631 (1955).

rate constants for decay, $k_{\mathrm{epr}}$, at 25° together with the range in initial concentration of nitroxide over which measurements were made. This table also lists the overall activation energies, $E_{\mathrm{epr}}$, and preexponential factors, $A_{\mathrm{epr}}$, for decay, together with the temperature range studied. Our earlier data on diethyl nitroxide[5] have also been included.

In most cases there was little change in the epr spectrum of the nitroxide as the temperature was changed. However, the spectrum of piperidine nitroxide is highly dependent on the temperature as has been previously reported.[12,13] The variation with temperature is due to changes in the hyperfine interaction of the protons at the 2 and 6 position which result from conformational changes in the ring.

Irradiation of dibenzylhydroxylamine in isopentane containing di-*tert*-butyl peroxide at temperatures below ~−70° gave a rather broad single line instead of the usual nitroxide signal. The position of this line ($g \approx 2.013$) suggests that it is a peroxy radical. Apparently the hydroxylamine is almost completely intermolecularly hydrogen bonded at these low temperatures. The hydroxylic hydrogen can then no longer be abstracted by a *tert*-butoxy radical which instead attacks a benzylic hydrogen. The resultant carbon radical picks up residual traces of oxygen to form a peroxy radical, $\mathrm{PhCH_2(PhCHOO\cdot)NOH}$.

We have shown previously that diethyl nitroxide exists in an equilibrium with a diamagnetic dimer at low temperatures in isopentane and in $\mathrm{CF_2Cl_2}$[14]

$$2(\mathrm{CH_3CH_2})_2\mathrm{NO}\cdot \underset{-1}{\overset{1}{\rightleftharpoons}} ((\mathrm{CH_3CH_2})_2\mathrm{NO})_2$$

The evidence for dimer formation was that the nitroxide signal could be reversibly increased or decreased by raising or lowering the temperature in the range −100 to −145°. Even at these low temperatures the irreversible decay of diethyl nitroxide was extremely rapid and in isopentane the equilibrium constant could not be measured. However, in $\mathrm{CF_2Cl_2}$ the equilibrium constant could be roughly measured by correcting for the change in nitroxide concentration due to irreversible decay that occurred in the time required to make measurements at more than one temperature. The variation in the equilibrium constant with temperature could be represented by the van't Hoff relation

$$K_1 = \frac{k_{-1}}{k_1} = e^{\Delta S_1/R}e^{-\Delta H_1/RT}$$

The data were not very good because of the rapid irreversible decay of nitroxide. The best fit to all the data in $\mathrm{CF_2Cl_2}$ was obtained with $\Delta H_1 = 6 \pm 3$ kcal/mol and $\Delta S_1 = 18 \pm 20$ gibbs/mol.[5]

The mechanism we have proposed for irreversible decay of diethyl nitroxide involves a hydrogen atom transfer *via* a five-centered cyclic transition state of the dimer.[5] Replacement of the hydrogen atoms that can undergo this transfer by deuterium slowed down the rate of irreversible decomposition (Table I)[16] and thus allowed $\Delta H_1$ and $\Delta S_1$ to be measured for $(\mathrm{CH_3CD_2})_2\mathrm{NO}\cdot$ with much more accuracy than for the undeuterated radical (see Figure 1). In $\mathrm{CF_2Cl_2}$, $(\Delta H_1)_\mathrm{D} = 2.6$ kcal/mol and $(\Delta S_1)_\mathrm{D} = 0.1$ gibb/mol while in isopentane $(\Delta H_1)_\mathrm{D} = 7.6$ kcal/mol and $(\Delta S_1)_\mathrm{D} = 24.8$ gibbs/mol. These values are compared with our previous data[5] in Table II. For the majority of the other nitroxides studied in this work irreversible decay was too rapid for dimer formation to be detected. In the case of diisopropyl nitroxide, dimer was not detected although the decay was very slow. Presumably in this case steric effects reduce $\Delta H$ to such a degree that dimer is not formed in significant quantities and, as a result, irreversible decomposition is exceedingly slow. Dimer formation was detected

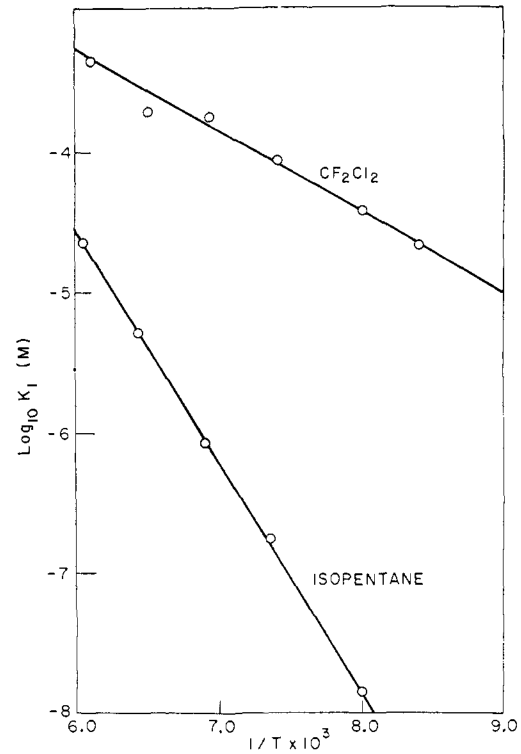
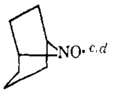
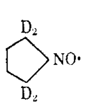


Figure 1. The $\alpha,\alpha,\alpha',\alpha'$-tetradeuteriodiethyl nitroxide radical–dimer equilibrium constant as a function of temperature.

**Table II.** Nitroxide–Dimer Equilibria

| Radical | Isopentane $(K_1)_{25°}$, $M$ | Isopentane $\Delta S_1$[a] | Isopentane $\Delta H_1$[b] | $\mathrm{CF_2Cl_2}$ $(K_1)_{25°}$, $M$ | $\mathrm{CF_2Cl_2}$ $\Delta S_1$[a] | $\mathrm{CF_2Cl_2}$ $\Delta H_1$[b] |
|---|---|---|---|---|---|---|
| $(\mathrm{CH_3CH_2})_2\mathrm{NO}\cdot$[c] | | | | 0.34 | 18 ± 20 | 6 ± 3 |
| $(\mathrm{CH_3CD_2})_2\mathrm{NO}\cdot$ | 0.63 | 24.8 | 7.6 | 0.018 | 0.1 | 2.6 |
|  [c,d] | 0.009 | 14.8 | 7.2 | 0.007 | −0.2 | 2.9 |
|  | | | | | | 1–2 |

[a] Gibbs per mole. [b] Kilocalories per mole. [c] Reference 5. [d] Nortropane-*N*-oxyl.

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(12) J. J. Windle, J. A. Kuhnle, and B. H. Beck, *J. Chem. Phys.*, 50, 2630 (1969).

(13) R. E. Rolfe, K. D. Sales, and J. H. P. Utley, *Chem. Commun.*, 540 (1970).

(14) Benson[15] has pointed out that appreciable dimerization in such solvents is not unexpected because nitroxides are probably fairly polar. The heat of dimerization could be ~3 kcal/mol.

(15) S. W. Benson, private communication.

(16) At 25°, $(k_{\mathrm{epr}})_\mathrm{H}/(k_{\mathrm{epr}})_\mathrm{D}$ is 6.0 in isopentane and 14 in $\mathrm{CF_2Cl_2}$ for diethyl nitroxide. For pyrrolidine nitroxide at 25° in benzene the isotope effect is 13.

in $CF_2Cl_2$ with dibenzyl nitroxide and with 2,2,5,5-tetradeuteriopyrrolidine nitroxide. In the latter case decay was just slow enough for $\Delta H$ to be estimated as ~1–2 kcal/mol, but $\Delta S$ could not be determined.

**Thermochemistry.**[17] This work was undertaken to determine a rough value for the heat of the self-reaction of $Et_3NO$. The radicals were generated from the hydroxylamine by oxidation with ceric ammonium nitrate in methanol as solvent in a homemade calorimeter. The reaction was carried out at −40° to prevent hydrolysis of the product nitrone.[5] The necessity for this low temperature created a number of problems in temperature control not all of which were completely overcome.

An accurately weighed quantity ($3 \times 10^{-3}$–$1 \times 10^{-2}$ mol) of the compound to be oxidized was sealed under vacuum into a thin-walled glass vessel. This vessel was attached to the bottom of a plunger and immersed in the calorimeter in 100 ml of methanol containing a 10% excess of the stoichiometrically required amount of ceric ammonium nitrate, the whole being cooled to ~−40°. The temperature of the calorimeter was monitored with a thermocouple, the output of which was fed through a dc amplifier to a strip chart recorder. After thermal equilibrium was achieved the sample vessel was crushed and the temperature was recorded until the reaction—as judged from the chart trace—was complete. The calorimeter and contents were then calibrated electrically in the usual way, *i.e.*, the heat capacity of the system was determined with an electric heater. Blank experiments were carried out using the same procedure but without the ceric ammonium nitrate in order to correct for the heat of solution of the compound to be oxidized. The measured heats of reaction are probably reliable to within about ±1.5 kcal/mol.

The heat of the $Ce^{IV} \rightarrow Ce^{III}$ conversion under the required experimental conditions was determined using the ceric ion–hydrogen peroxide reaction

$$2Ce^{IV}_{soln} + H_2O_{2liq} \longrightarrow 2Ce^{III}_{soln} + 2H^+_{soln} + O_{2gas}$$

Experimentally this reaction was found to be exothermic by 37 kcal/mol (~2.5 kcal/mol of this being the heat of solution of the $H_2O_2$). The known heats of formation are, $\Delta H_f°(H_2O_2)_{liq} = -44.9$, $\Delta H_f°(O_2)_{gas} = 0$, and $\Delta H_f°(H\cdot)_{gas} = 52.1$ kcal/mol.[18] Hence, for the reaction

$$2Ce^{IV}_{soln} + 2H\cdot_{soln} \longrightarrow 2Ce^{III}_{soln} + 2H^+_{soln}$$

$\Delta H_{233} = -37 - 45 - 104 + 2a = 2a - 186$ kcal/mol, where $2a$ is the heat of the reaction $2H\cdot_{gas} \rightarrow 2H\cdot_{soln}$.

Hydroquinone and *N*-methylhydroxylamine were also tried as calibrating compounds, but without success. Hydroquinone oxidized only as far as the quinhydrone rather than all the way to quinone and the quinhydrone precipitated from solution. *N*-Methylhydroxylamine reacted too slowly with ceric ion under the required experimental conditions for the heat of this reaction to be determined.

The measured heat of the reaction

$$2Ce^{IV}_{soln} + Et_2NOH_{solid} \longrightarrow 2Ce^{III}_{soln} + 2H^+_{soln} + EtN(\rightarrow O){=}CHCH_{3soln}$$

was −72 kcal/mol based on hydroxylamine with no appreciable heat of solution for the hydroxylamine. Hence, for the reaction

$$Et_2NOH_{soln} \longrightarrow EtN(\rightarrow O){=}CHCH_{3soln} + 2H\cdot_{soln}$$

$\Delta H_{233} = -72 - (2a - 186) = 114 - 2a$ kcal/mol. It seems reasonable to assume that the O–H bond strength in diethylhydroxylamine should be the same as that found for 2,2,6,6-tetramethyl-*N*-hydroxypiperid-4-one, *viz.* 71 kcal/mol.[19] If we assume $2a$ is small and ignore any differences arising from the fact that our data were obtained in methanol at −40° whereas the O–H bond strength was determined in benzene at room temperature, then

$$D[EtN(O\cdot)C(CH_3)H{-}H] = 114 - 71 = 43 \text{ kcal/mol}$$

Therefore, for the self-reaction of diethyl nitroxide

$$Et_2NO\cdot_{soln} \longrightarrow 0.5Et_2NOH_{soln} + 0.5EtN(\rightarrow O){=}CHCH_{3soln}$$

$\Delta H_{233} = 0.5(43 - 71) = -14$ kcal/mol with an error that probably does not exceed ±3 kcal/mol.

As a rough check on the validity of the approximations used in this method of estimating the exothermicity of the diethyl nitroxide self-reaction a similar experiment was carried out with *n*-butyl hydroperoxide.[20] The overall oxidation of this compound with ceric ion can be represented by

$$Ce^{IV}_{soln} + n\text{-}BuOOH_{soln} \longrightarrow Ce^{III}_{soln} + H^+_{soln} + 0.5n\text{-}BuOH_{soln} + 0.5C_3H_7CHO_{soln} + 0.5O_{2gas}$$

In methanol at −40° this reaction is exothermic by 54 kcal/mol (based on hydroperoxide and corrected for the heat of solution of the hydroperoxide). Hence, for the reaction

$$n\text{-}BuOOH_{soln} \longrightarrow H\cdot_{soln} + 0.5n\text{-}BuOH_{soln} + 0.5C_3H_7CHO_{soln} + 0.5O_{2gas}$$

$\Delta H_{233} = 0.5(186 - 2a) - 54 = 39 - a = 39$ kcal/mol with an error that probably does not exceed ±3 kcal/mol.

The O–H bond strength in hydroperoxides, D(ROO–H), has been estimated to be 90 kcal/mol in the gas phase[21a] and an experimental value of 88.0 ± 1.7 kcal/mol has recently been measured in solution.[21b] Assuming that $a$ is small and that the heat of solution of the alcohol plus aldehyde is similar to that of 2 mol of hydroperoxide we obtain $\Delta H_{233} = -88 + 39 = -49 \pm 5$ kcal/mol for the reaction

$$n\text{-}BuOO\cdot \longrightarrow 0.5n\text{-}BuOH + 0.5C_3H_7CHO + 0.5O_2$$

The exothermicity in the gas phase of a reaction exactly analogous to this last reaction can be calculated from the heats of formation of the corresponding ethyl compounds which have been given by Benson,[22] *viz.*, $\Delta H_f°(EtOO\cdot) = -1.8$, $\Delta H_f°(EtOH) = -56.2$,

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[17] The work described in this section was done with the aid of L. E. Iton.

[18] D. D. Wagman, *et al.*, *Nat. Bur. Stand.* (*U. S.*), *Tech. Note*, 270 (1968).

[19] L. R. Mahoney, private communication.

[20] Prepared according to the procedure of H. R. Williams and H. S. Mosher (*J. Amer. Chem. Soc.*, **76**, 2984 (1954)).

[21] (a) S. W. Benson and R. Shaw, *Advan. Chem. Ser.*, **75**, 288 (1968); (b) L. R. Mahoney and M. A. DaRooge, *J. Amer. Chem. Soc.*, **92**, 4063 (1970).

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

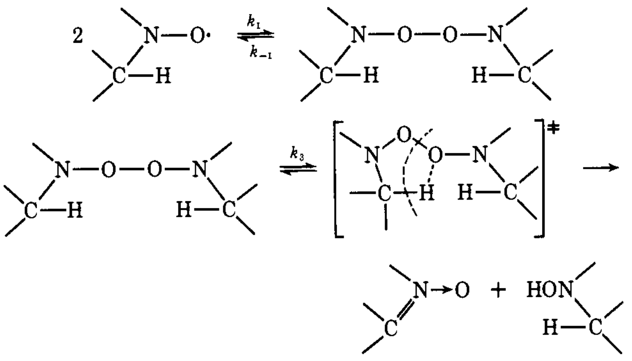
and $\Delta H_f°(CH_3CHO) = -39.7$ kcal/mol. That is, for the reaction

$$EtOO\cdot \longrightarrow 0.5EtOH + 0.5CH_3CHO + 0.5O_2$$

$\Delta H = -0.5(39.7 + 56.2) + 1.8 = -46.2$ kcal/mol which is in reasonable agreement with the measured value of $-49 \pm 5$ kcal/mol for *n*-butylperoxy in methanol. This implies that the value of $-14 \pm 3$ kcal/mol estimated for the diethyl nitroxide self-reaction is probably reasonably accurate.

## Discussion

All the data obtained in the present work lend support to the previously proposed mechanism for the irreversible bimolecular decay of dialkyl nitroxides having a hydrogen atom on one of the $\alpha$-carbon atoms, *viz.*[23]



$$k_{epr} = A_{epr}e^{-E_{epr}/RT} = k_3/K_1 = A_3e^{-\Delta S_1/R}e^{-(E_3-\Delta H_1)/RT}(M^{-1}\ sec^{-1})$$

This two-step mechanism is supported by the small values of the preexponential factors for these radical–radical reactions. The generally small activation energies indicate that $\Delta H_1$ and $E_3$ are of comparable magnitude. In most solvents $E_{epr}$ is positive implying that $E_3$ is slightly greater than $\Delta H_1$ but in isopentane $E_{epr}$ is negative (except for the $(CH_3CD_2)_2NO\cdot$ radical) implying that $E_3 < \Delta H_1$. For diisopropyl nitroxide in benzene $A_{epr}$ is in the usual range but in $CF_2Cl_2$ $A_{epr}$ is larger than usual. The activation energies in both solvents are considerably larger than for any of the other nitroxides. Our failure to detect any dimerization of this nitroxide in spite of its very slow decay at low temperatures implies that for this hindered nitroxide $\Delta H_1$ can be no more than 1 kcal/mol at best. In fact the activation parameters in $CF_2Cl_2$ suggest that the reaction may occur by a straightforward disproportionation not involving the formation of an intermediate dimer.

$$2\text{-}i\text{-}Pr_2NO\cdot \longrightarrow i\text{-}Pr_2NOH + i\text{-}PrN(\rightarrow O){=}C(CH_3)_2$$

The transfer of an $\alpha$-hydrogen atom in a rate-controlling step is indicated by the large deuterium isotope effect, $(k_{epr})_H/(k_{epr})_D \sim 6\text{–}14$, at 25°.[25a] As we might expect, this isotope effect is due primarily to a difference in the overall activation energies for the deuterated and undeuterated radicals. The activation energy difference, *i.e.*, $(E_{epr})_D - (E_{epr})_H$, is 1.5 kcal/mol for diethyl nitroxide in both isopentane and $CF_2Cl_2$ which is of the same order of magnitude as the difference in zero-point energy of the C–D and C–H bonds.[26] This suggests that most of the zero-point energy of the transferred hydrogen is lost in the transition state and hence that the isotope effect arises primarily from step 3. This is in accord with our expectation that there should be little or no isotope effect on the equilibrium constant.[27] Unfortunately, our measurements of $K_1$, $\Delta H_1$, and $\Delta S_1$ for diethyl nitroxide[5] were too inaccurate to allow this supposition to be checked by comparison with the far more accurate data now obtained with tetradeuteriodiethyl nitroxide. However, the data for this deuterated nitroxide in isopentane and in $CF_2Cl_2$ are in excellent agreement with our earlier data[5] for dimerization of the stable, unhindered nortropane-*N*-oxyl (see Table II). We therefore believe that the measured entropy and enthalpy changes accompanying dimerization of tetradeuteriodiethyl nitroxide in the two solvents, will be generally applicable to other relatively nonhindered nitroxides.

The loss in entropy for the dimer of a nitroxide containing four transferable hydrogen atoms which accompanies the formation of the cyclic transition state, $\Delta S_3{}^{\ddagger}$, was calculated previously to be $-8.4 \pm 1.4$ gibbs/mol.[5] In this calculation any solvent effect on reaction 3 was ignored and hence this value can only be correct if the reaction is carried out in the gas phase or in a totally inert (*i.e.*, nonsolvating) solvent. The overall preexponential factor at the mean temperature of most experiments (220°K) can be calculated from the relation

$$A_{epr} = \frac{ekT}{h} \times \frac{e^{\Delta S_3^{\ddagger}/R}}{e^{\Delta S_1/R}} = \frac{10^{13.02} \times 10^{-1.84}}{10^{\Delta S_1/2.3R}} = 10^{11.18-\Delta S_1/4.6}\ M^{-1}\ sec^{-1}$$

Our original data with diethyl nitroxide[5] suggested that $CF_2Cl_2$ was inert. That is, the calculated value of $A_{epr}$ was $10^{7.3}\ M^{-1}\ sec^{-1}$ in reasonable agreement with the measured value of $10^{5.0\pm1.0}\ M^{-1}\ sec^{-1}$. However, the revised value of $\Delta S_1 = 0.1$ gibbs/mol (based on tetradeuteriodiethyl nitroxide) gives a calculated $A_{epr} =$

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[23] The reasons for formulating the dimer as >NOON< in the absence of experimental evidence were formulated previously.[5] Additional support for this structure comes from a recent formulation[24] of the $NO_3$ radical dimer as $O_2NOONO_2$. On the other hand, there are strong thermochemical arguments against oxygen–oxygen coupling of nitroxides. These radicals are stabilized by internal dative bonding of nitrogen to oxygen; thus $D(R_2NO\text{–}H) \approx 71$ kcal/mol[19] compared with $D(RO\text{–}H) \approx 104$ kcal/mol[21a] which indicates a stabilization energy of 33 kcal/mol. Since $D(RO\text{–}OR) \approx 38$ kcal/mol[21a] we can calculate that $D(R_2NO\text{–}ONR_2) \approx -28$ kcal/mol. The validity of this type of calculation is attested by the results for the isoelectronic peroxy radical. That is, $D(ROO\text{–}H) \approx 90$ kcal/mol[21a] corresponding to a stabilization of 14 kcal/mol and a calculated value of $D(ROO\text{–}OOR) \approx 10$ kcal/mol in excellent agreement with the experimental value of ~9 kcal/mol.[25]

[24] R. W. Glass and T. W. Martin, *J. Amer. Chem. Soc.*, **92**, 5084 (1970).

[25] K. Adamic, J. A. Howard, and K. U. Ingold, *Can. J. Chem.*, 47, 3803 (1969); J. E. Bennett, D. M. Brown, and B. Mile, *Trans. Faraday Soc.*, **66**, 397 (1970).

[25a] NOTE ADDED IN PROOF. The bimolecular self-reaction of isopropyl *tert*-butyl nitroxide exhibits an isotope effect, $k_H/k_D = 5.5$ at 25° in cyclohexane, when the tertiary H is replaced by deuterium. The activation energy for reaction of the undeuterated radical is *ca.* 9 kcal/mol in cyclohexane (R. Briere, Thesis, Centre d'Etudes Nucléaires de Grenoble, 1969).

[26] L. Melander, "Isotope Effects on Reaction Rates," Ronald Press, New York, N. Y., 1960.

[27] Deuteration of the nitroxide would most probably increase the concentration of dimer because the characteristically smaller amplitude of vibration of deuterium relative to hydrogen would reduce nonbonded repulsions.[28] This would contribute an inverse isotope effect to the overall reaction.

[28] L. S. Bartell, *Tetrahedron Lett.*, 13 (1960); *J. Amer. Chem. Soc.*, **83**, 3567 (1961).

**Table III.** Comparison of Nitroxide and Peroxy Radical Self-Reactions

| Reaction | $k_{25°}$, $M^{-1}$ sec$^{-1}$ | $\Delta H$, kcal for 2 mol of radical | D(X–H) in radical |
|---|---|---|---|
| $2H_2NO\cdot \rightarrow H_2NOH + HNO$ | | $-5^{a}$ | $66^{a}$ |
| $2MeNHO\cdot \rightarrow MeNHOH + MeNO$ | $3.6 \times 10^{7\,b}$ | $-10^{b}$ | $61^{b}$ |
| $2HOO\cdot \rightarrow HOOH + O_2$ | $1.3 \times 10^{9\,c}$ | $-43^{d}$ | $47^{d}$ |
| $2Et_2NO\cdot \rightarrow Et_2NOH + EtN(\rightarrow O){=}CHCH_3$ | $1.4 \times 10^{3\,e}$ | $-28^{f}$ | $43^{f}$ |
| $2EtOO\cdot \rightarrow EtOH + CH_3CHO + O_2$ | | $-92^{f}$ | $63^{g}$ |
| $2n\text{-}BuOO\cdot \rightarrow n\text{-}BuOH + C_3H_7CHO + O_2$ | $4 \times 10^{7\,h}$ | $-102^{f}$ | $63^{g}$ |

[a] $\Delta H_f°(H_2NOH) = -9$, $\Delta H_f°(HNO) = 23.8$ (ref 21b), D($H_2NO$–H) taken as 71 kcal/mol (see text). [b] In benzene (ref 2). In methanol, $k_{25°} = 1.5 \times 10^6$ $M^{-1}$ sec$^{-1}$. [c] In chlorobenzene at 30° (J. A. Howard and K. U. Ingold, *Can. J. Chem.*, **45**, 785 (1967)). In water at room temperature $k = 2 \times 10^6$–$6 \times 10^6$ $M^{-1}$ sec$^{-1}$. [d] D(HOO–H) = 90, D(H–OO·) = 47 kcal/mol (ref 21a). [e] In benzene (see Table I). [f] See text. [g] Reference 15. [h] In tetralin and α-methylstyrene (J. A. Howard and K. U. Ingold, *J. Amer. Chem. Soc.*, **90**, 1058 (1968)).

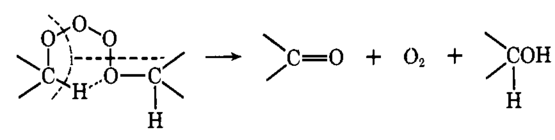
$10^{11}$ $M^{-1}$ sec$^{-1}$ in very poor agreement with experiment. In contrast, the agreement between experiment and theory in isopentane is fairly good. In this solvent, $\Delta S_1 = 24.8$ gibbs/mol and hence $A_{epr}$ is calculated to be $10^{5.8}$ $M^{-1}$ sec$^{-1}$. The mean of the measured values in this solvent is $10^{4.5}$ $M^{-1}$ sec$^{-1}$.

It is clear from the foregoing that the effect of solvents on the overall rate of the self-reaction of nitroxides cannot be explained solely on the basis of their effect on nitroxide dimerization. The mechanism outlined above gives a satisfactory explanation of the results obtained in isopentane which is a "poor" solvent for nitroxide and dimer since it is both nonpolar and nonpolarizable. That is, isopentane approximates the gas phase because it is a poorly solvating solvent and, as a result, theory and experiment are in reasonable agreement. In contrast, in a "good" solvent such as $CF_2Cl_2$ the nitroxide radicals are strongly solvated because of their dipolar character. The dimer is expected to be less strongly solvated than the radicals. In a good solvent the desolvation accompanying step 1 will make $\Delta S$ more positive than in a poor solvent, and the heat of dimerization will be reduced by an amount equal to the difference in the heats of solvation of radicals and dimer. Furthermore, the polarity of the final products is expected to be greater than the polarity of the dimer. Since the transition state in step 3 must resemble the products to some extent it will be more polar, and hence more strongly solvated, than the dimer. As a consequence, $\Delta S_3^{\ddagger}$ should be more negative than is the case for a poor solvent and the activation energy for step 3 will be reduced. (Thus, for diethyl nitroxide $E_3 = 7.6 - 0.9 = 6.7$ kcal/mol in isopentane and $2.6 + 1.0 = 3.6$ kcal/mol in $CF_2Cl_2$.) The effects of changes in solvent on $k_{epr}$ and $E_{epr}$ are therefore minimized since changes in $\Delta S_1$ and $\Delta H_1$ are offset by changes of comparable magnitude (but opposite sign) in $\Delta S_3^{\ddagger}$ and $E_3$.

A fairly strong interaction between diethyl nitroxide radicals and solvent $CF_2Cl_2$ is suggested by our results. Significant solvent effects on free-radical reactions were first reported by Russell[29] who found that chlorine atoms were less reactive in hydrogen atom abstractions in certain aromatic solvents than in aliphatic solvents and in $CCl_4$. Since that time, solvent effects on many free-radical reactions have been observed.[30] Of special note in connection with the present work is Bentrude and MacKnight's[31] study of the partial molal heats of solution of di-*tert*-butyl nitroxide and di-*tert*-butylhydroxylamine. These workers found that in transfer from cyclohexane to an aromatic solvent or carbon tetrachloride the radical undergoes a more favorable enthalpy of solvation change than does the hydroxylamine but the reverse is the case for transfer to another alkane, to cyclohexene, or to *tert*-amyl alcohol. For benzene, toluene, and $CCl_4$ the solvation enthalpy of transfer favors the radical by ~900 cal/mol. It is concluded that the aromatics (and $CCl_4$) are able to solvate the nitroxide in some specific manner. If the highly hindered di-*tert*-butyl nitroxide radical has an "extra" solvation enthalpy of 0.9 kcal/mol in $CCl_4$ it seems not unreasonable that the unhindered diethyl nitroxide and nortropane-*N*-oxyl radicals should have an "extra" solvation enthalpy of several kilocalories per mole in $CF_2Cl_2$.

**Comparison of Nitroxide and Peroxy Radical Self-Reactions.** The overall mechanism for the self-reactions of nitroxides and of structurally analogous peroxy radicals appears to have several features in common which is not entirely unexpected since the two classes of radicals are isoelectronic. The self-reactions of monoalkyl nitroxides and of the HOO· radical involve transfer of the hydrogen atom attached to the atom adjacent to the formal site of the unpaired electron. These reactions are very rapid (see Table III). The nitroxide reactions are considerably less exothermic than the HOO· reaction, in part because the N–H bond strength in these readicals is ~12–14 kcal/mol greater than D(H–OO·).

The self-reaction of dialkyl nitroxides appears to proceed by a mechanism analogous to that proposed by Russell for primary and secondary peroxy radicals.[32]



There is a considerable body of data which supports this mechanism for secondary peroxy radicals.[33] Evidence against the most frequently postulated alternative mechanism, which involves alkoxy radicals as intermediates, continues to grow.[34]

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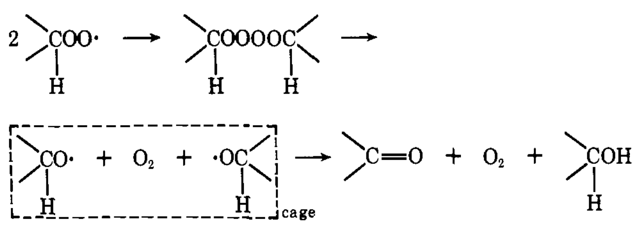
(29) G. A. Russell, *J. Amer. Chem. Soc.*, **79**, 2977 (1957); **80**, 4987, 4997, 5002 (1958).

(30) For a review see: E. S. Huyser, *Advan. Free Radical Chem.*, **1**, 77 (1965).

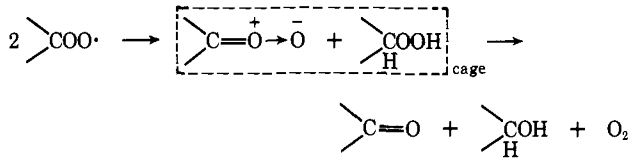
(31) W. G. Bentrude and A. K. MacKnight, *J. Amer. Chem. Soc.*, **92**, 5259 (1970).

(32) G. A. Russell, *ibid.*, **79**, 3871 (1957).

(33) For a summary see, K. U. Ingold, *Accounts Chem. Res.*, **2**, 1 (1969).



The rate constants at 30° for the self-reactions of a majority of secondary peroxy radicals are roughly two–three orders of magnitude less than the diffusion-controlled limit of $2 \times 10^9$ $M^{-1}$ sec$^{-1}$.[35-37] In the Russell mechanism this means that $k_2$ is generally in the range from $10^{-2}k_{-1}$ to $10^{-3}k_{-1}$. That is, when a di-*sec*-alkyl tetroxide is formed it has a 1–0.1% probability of yielding molecular products and a 99–99.9% probability of re-forming two secondary peroxy radicals. If we make the reasonable assumption that $\Delta S_{-1}^{\ddagger} = -20$ gibbs/mol and that $E_{-1} = 9.0$ kcal/mol (the enthalpy change for the decomposition of di-*tert*-alkyl tetroxides to two *tert*-alkyl peroxy radicals[25]) then $k_{-1} = (kT/h) \exp(\Delta S_{-1}^{\ddagger}/R) \exp(-E_{-1}/RT) = (kT/h) \exp(-4.8)$ sec$^{-1}$ at 30°. The minimum entropy loss in step 2 accompanying the loss of entropy of four hindered internal rotors is 3.6 gibbs/mol per rotor, for a total of $-14.4$ gibbs/mol.[15,38] Hence, if $k_2 = 10^{-2}k_{-1}$ then, at 30°, $e^{-4.8} \times 10^{-2} = e^{(14.4/2 - E_2/606)}$ from which $E_2$ is calculated to be 1.3 kcal/mol, and if $k_2 = 10^{-3}k_1$ then $E_2$ is calculated to be 2.7 kcal/mol.[39] Although this activation energy is very low for a six-center 1,5-hydrogen shift reaction (many of which are endothermic and have activation energies in the range 30–50 kcal/mol[22]) it does not, in our opinion, rule out the Russell mechanism because of the high exothermicity of alkyl peroxy radical self-reactions (see Table III). That is, in the dialkyl nitroxide self-reaction we believe we have a process analogous to the Russell mechanism. The activation energy for irreversible decomposition of the diethyl nitroxide dimer is 6.7 kcal/mol in isopentane and 3.6 kcal/mol in $CF_2Cl_2$ and yet the exothermicity for decomposition of the dimer is *not more than* $28 - \Delta H_1 = 20$ kcal in isopentane and 25 kcal in $CF_2Cl_2$. For the analogous decomposition of a *primary* or *secondary* alkyl tetroxide it seems not unreasonable that the activation energy should fall to ~1–2 kcal for a process which is exothermic by 60–70 kcal.[40] However, it has been pointed out[15] that such a low activation energy for irreversible tetroxide decomposition would lead to an apparent negative activation energy of several kilocalories per mole for the overall reaction. This is contrary to experimental observations (of low accuracy) of small positive activation energies (in the range 0–4 kcal/mol) with secondary peroxy radicals.[43] Some doubt therefore still remains as to the validity of the Russell mechanism. This difficulty has led Benson to propose the following mechanism.[15]



Some accurate activation energies for secondary peroxy radical self-reactions are needed at this point.

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(34) See in particular: R. Hiatt and S. Szilagyi, *Can. J. Chem.*, **48**, 615 (1970); R. Hiatt and L. Zigmund, *ibid.*, **48**, 3967 (1970).

(35) J. A. Howard and K. U. Ingold, *ibid.*, **45**, 793 (1967).

(36) J. A. Howard, *Advan. Free Radical Chem.*, 4, 49 (1971).

(37) The self-reaction rate constants for primary peroxy radicals are generally ~$2 \times 10^8$ $M^{-1}$ sec$^{-1}$.[35,36]

(38) H. E. O'Neal and S. W. Benson, *J. Chem. Eng. Data*, **15**, 266 (1970).

(39) This calculation was suggested to the authors by Professor Russell.

(40) The overall exothermicity of 92–102 kcal (Table III) must be corrected by 9 kcal for the enthalpy of formation of the tetroxide and by a further 22 kcal for the fact that the oxygen is formed in its first excited state,[41] ${}^1\Delta g$, which is 22 kcal above the ground state, ${}^3\Sigma\bar{g}$. If the oxygen is actually formed in its ground state and the carbonyl product in its excited triplet state, as has been proposed,[42] the overall exothermicity would be only ~10–20 kcal.

(41) J. A. Howard and K. U. Ingold, *J. Amer. Chem. Soc.*, **90**, 1056 (1968).

(42) R. E. Kellogg, *ibid.*, **91**, 5433 (1969).

(43) For a complete summary of this data see ref 36.

(44) N.R.C.C. summer student, 1970.