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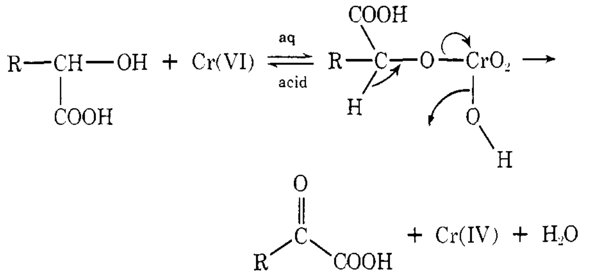

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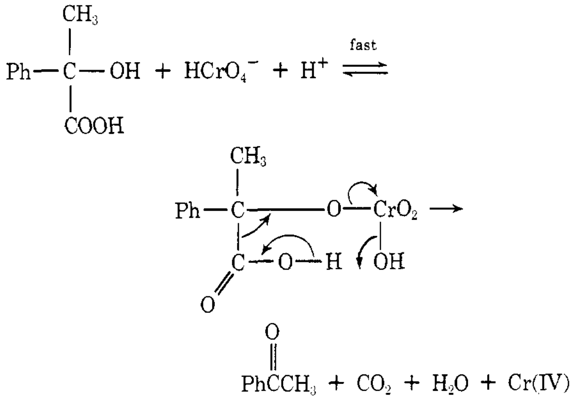

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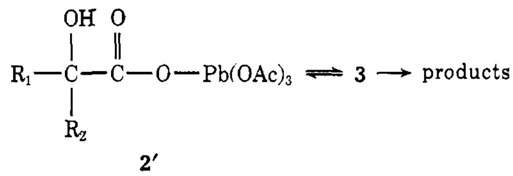
The effects of para substitutents on the rate of mandelic acid oxidation by $Pb(OAc)_4$ are found to be very minor. However, with greater structural changes in the 2-hydroxy acids, somewhat larger differences in oxidation rate are observed. These structural effects may reflect differences in the stabilization of developing sp^2 character or may indicate that such cleavages, though concerted, are not strictly synchronous.

All the evidence so far cited strongly indicates that 2-hydroxycarboxylic acid oxidations occur *via* a more or less concerted mechanism in which the rate-determining step involves the decomposition of a Pb(IV)–hydroxy acid intermediate. On the basis of the present study it seems more probable that the cyclic intermediate **3** is involved in the rate-determining step. However, none of the facts so far reported would preclude the formation of **3** from **2′**, Scheme III. The importance of the hydroxyl group is made apparent in Table III. Thus, whereas 1,2-diols, 2-hydroxy acids, hydrobenzoin, and 2-hydroxy esters are oxidized by LTA, 2-keto acids, mandelic acetate, and benzil are not. At present, Scheme II is preferred, although Scheme III is certainly not excluded.

Scheme III



The Production of Singlet Oxygen in Electrogenerated Radical Ion Electron Transfer Reactions

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Abstract: The electrochemical reduction of oxygen generated superoxide ion and the electrochemical oxidation of ferrocene yielded the ferricenium ion. Both of these ion radicals were produced alternately in the same solution by pulsing the electrode potential between the reduction potential of oxygen and the oxidation potential of ferrocene. The homogeneous electron transfer reaction between superoxide ion and the ferricenium ion then produced ferrocene and singlet oxygen. The intermediacy of singlet oxygen was demonstrated by chemical quenching with 1,3-diphenylisobenzofuran (DPBF) which produced *o*-dibenzoylbenzene (DBB). The electron transfer reaction between superoxide and DPBF cation radical also produced singlet oxygen.

The annihilation reaction between ion radicals has recently been an area of intense study. It has been well documented that the mechanism of these reactions is an electron transfer reaction between a cation radical and an anion radical producing an electronically excited species and a neutral ground state molecule.[1-3] The effect of these reactions is the ultimate emission of radiation characteristic of the fluorescence of either R_1 or R_2. When the ion radicals are generated electrochemically, this phenomenon is called electrogenerated chemiluminescence (ecl). In ecl, both ion radicals are generated in the same solution by alternately pulsing the electrode potential from that where the reduction of R_1 occurs to that for the oxidation of R_2. In this manner a very high concentration of ion radicals can be produced in a small volume near the electrode.

$$\left.\begin{array}{l} R_1 \xrightarrow{-e} R_1\cdot^+ \\ \\ R_2 \xrightarrow{+e} R_2\cdot^- \end{array}\right] \longrightarrow \begin{array}{lr} R_1 + R_2^* \longrightarrow R_2 + h\nu & (1) \\ \text{or} & \\ R_2 + R_1^* \longrightarrow R_1 + h\nu & (2) \end{array}$$

Luminescence was not, however, the desired product in this investigation. Since electronically excited species are generated in ecl when the enthalpy of the

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(1) D. M. Hercules, *Accounts Chem. Res.*, 2, 301 (1969); A. Wessberger and B. Rossiter, Ed., "Physical Methods of Organic Chemistry," 4th ed, Part IIA, Academic Press, New York, N. Y., 1971.

(2) A. J. Bard, K. S. V. Santhanam, S. A. Cruser, and L. R. Faulkner, "Fluorescence," G. G. Guilbault, Ed., Marcel Dekker, New York, N. Y., 1967, Chapter 14.

(3) T. Kuwana, "Electroanalytical Chemistry," Vol. 1, A. J. Bard, Ed., Marcel Dekker, New York, N. Y., 1966, Chapter 3.

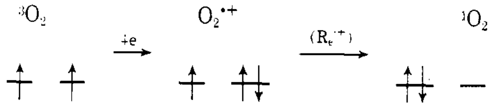
annihilation reaction is greater than the energy of the excited state, we proposed to produce singlet oxygen, the electronically excited state of oxygen, by using ecl techniques. The desired reaction is

$$R \xrightarrow{-e} R\cdot^{+} \quad (3)$$

$$^{3}O_2 \xrightarrow{+e} O_2\cdot^{-} \quad (4)$$

$$O_2\cdot^{-} + R\cdot^{+} \longrightarrow R + {}^{1}O_2 \quad (5)$$

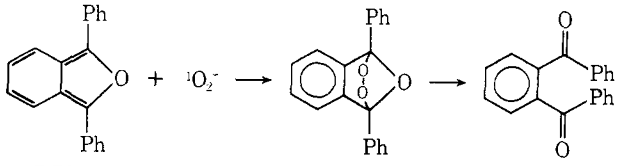
The one-electron reduction of oxygen produces superoxide ion and, in the presence of an appropriate cation radical, the annihilation reaction could produce singlet oxygen. A molecular orbital diagram of the process would be



The production of singlet oxygen requires that the enthalpy of reaction 5 be greater than the energy of the $^{1}\Delta$ state of O_2 (0.97 eV) above the ground state. The only previous observation of such a reaction was reported by Khan.[4] He observed that when a fluorescent dye was added to a saturated solution of potassium superoxide in dimethyl sulfoxide, fluorescence emission characteristic of the dye was observed. He also found that the addition of 2,5-dimethylfuran to the superoxide solution gave a positive peroxide test. Foote[5] had previously utilized this technique as a qualitative test for the presence of 2-methoxy-5-hydroperoxy-2,5-dihydrofuran, a known product of 2,5-dimethylfuran and singlet oxygen. These data were taken asev idence for the intermediacy of singlet oxygen and that the 2(Σ) state was sensitizing the dye fluorescence. The mechanism by which superoxide ion was oxidized was not clearly defined, since no chemical oxidants were added. Recently, Kearns[6] has suggested that the peroxide test was somewhat unreliable, since hydrogen peroxide can be generated from the reaction of superoxide and 2,5-dimethylfuran and this could explain the positive peroxide test.

We now wish to present evidence which indicates that singlet oxygen can be produced by oxidation of O_2^{-} under appropriate conditions.

The chemical method for detecting singlet oxygen was used since observation of emission from either the $^{1}\Sigma$, the $^{1}\Delta$ state, or dimeric states in solution is very improbable. The chemical quencher utilized was 1,3-diphenylisobenzofuran (DPBF) which has the largest rate constant, 8×10^{8} M^{-1} sec^{-1}, for quenching singlet oxygen.[7] The mechanism for quenching is outlined in the following reaction.



DPBF

DBB

Two types of experiments were undertaken. In one, ferrocene was employed as the cation radical oxidant and DPBF as the quencher and, in the other experiment, DPBF was used as both the cation radical oxidant and quencher.

Experimental Section

The 1,3-diphenylisobenzofuran (DPBF) from K & K Laboratories was used without further purification, since the ir spectrum (Nujol mull) showed no carbonyl absorption at 1650 cm^{-1}. Ferrocene from Aldrich Chemical Co. was purified by sublimation. Polarographic grade tetra-*n*-butylammonium perchlorate (TBAP), from Southwestern Analytical Chemical Co., was dried *in vacuo* at 100° for 6 hr and then stored in a desiccator over calcium chloride. Reagent grade acetonitrile from Matheson Coleman and Bell was purified by refluxing with sodium carbonate and potassium permanganate for 1 hr, then distilling. A few drops of concentrated sulfuric acid were added to the distillate. The acetonitrile was decanted from the white precipitate and then distilled. The fraction between 81–82° was collected and then eluted twice through a 1 m × 3 cm silica gel column and stored in a flask under He. Matheson Coleman and Bell 60–200 mesh silica gel was used.

The electrochemical cell used has been previously described.[8] A gas inlet was added to the working compartment. The auxiliary electrode was a platinum wire and the working electrode was either a 4.2-cm^2 platinum foil or a 9.60-cm^2 mercury pool. The reference electrode was a saturated calomel electrode (sce). For cyclic voltammetry studies a 0.78-mm^2 platinum disk was utilized. In a typical experiment TBAP (7 mmol) was added to the freeze–pump–thaw chamber of the cell and then the cell was evacuated. Acetonitrile (70 ml) was then transferred into the cell under vacuum. A positive pressure of He was then applied and then the appropriate substrates and electrodes were added. The cell was then placed inside a box which could be closed to exclude room light. Dry oxygen was then continuously bubbled through the solution.

For the pulse experiments the potentials were controlled by a Princeton Applied Research Model 176 electrochemistry system. A Wavetek Model 114 function generator operating in the square mode controlled the pulse frequency. For cyclic voltammetry studies a Princeton Applied Research Model 170 electrochemistry system was used.

After completion of the electrolysis the solution in the working compartment was transferred to a round bottom flask. The solvent was then evaporated by means of a Rotovap. The residue was extracted three times with ether. The ether layer was dried with magnesium sulfate and then evaporated. The resulting residue was placed on a 30 cm × 12 mm silica gel (Matheson Coleman and Bell, 60–200 mesh) chromatography column. The column was eluted with benzene and fractions 1 and 2 (50-ml aliquots) contained, depending upon the particular experiment, either DPBF or DPBF and ferrocene. The column was then eluted with acetone and fractions 3 and 4 contained DBB. Recrystallization from ether yielded colorless crystals. The ir spectrum was identical with the known spectrum for DBB (Sadtler No. 20976). The mp was 145–146° (uncorrected) compared with the literature value, 146–147°.[9]

Results

Pulse Electrolysis of DPBF and Oxygen. A continuous flow of oxygen was passed through a solution of DPBF (0.55 mmol) in 0.10 M TBAP in acetonitrile by means of a bubbler. Because DPBF has been reported to react strongly with oxygen, even in the absence of light,[10] the experiments were performed in a closed box. The potential of the Pt electrode was pulsed from +0.90 to −1.3 V at a frequency of 0.75 cps for 4 hr. Cyclic voltammetry indicated that the oxidation of DPBF was irreversible with $E_p = +0.8$ V while the reduction of DPBF exhibited Nernstian behavior with $E_p = -2.0$ V. Zweig and coworkers[11] found similar

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[6] D. R. Kearns, *Chem. Rev.*, **71**, 395 (1971).

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Table I. Summary of Experimental Results

Reactants	Species generated	Potential, V	Yield mmol	DBB %[a]
O_2, DPBF	None	None	0.07	12
O_2, DPBF	$O_2\cdot^-$, DPBF$\cdot^+$	+0.9 to −1.3	0.35	64[b]
O_2, DPBF	DPBF$\cdot^+$	+0.9 to −0.8	0.07	12
O_2, DPBF	O_2^-	+0.7 to −1.3	0.07	12
O_2, ferrocene, DPBF	None	None	0.06	12
O_2, ferrocene, DPBF	$O_2\cdot^+$, ferr$\cdot^+$	+0.4 to −0.9	0.30	55[b]

[a] Based upon theoretical yield. [b] Average value for three experiments.

behavior in ecl studies of DPBF. The reduction of oxygen exhibited quasi-reversible behavior with E_p = −1.4 V in agreement with results obtained by Peover and coworkers.[12] Isolation of products yielded DPBF (0.1 mmol) and DBB (0.35 mmol).

Control Experiments with DPBF and Oxygen. Three control experiments were performed utilizing identical reaction conditions as those employed in the pulse experiment, except for the potential applied. In the first experiment, the potential was pulsed from +0.9 to −0.8 V. Under these conditions, the DPBF cation radical is formed but no superoxide is generated. Product analysis indicated that DPBF (0.05 mmol) and DBB (0.07 mmol) were formed.

In the second experiment, the potential was pulsed from +0.7 to −1.3 V. Under these conditions, superoxide was generated but no DPBF cation radicals were formed. DPBF (0.43 mmol) and DBB (0.07 mmol) were isolated from the reaction mixture.

In the third experiment neither the DPBF cation radical nor superoxide was produced. The products isolated were DPBF (0.44 mmol) and DBB (0.07 mmol).

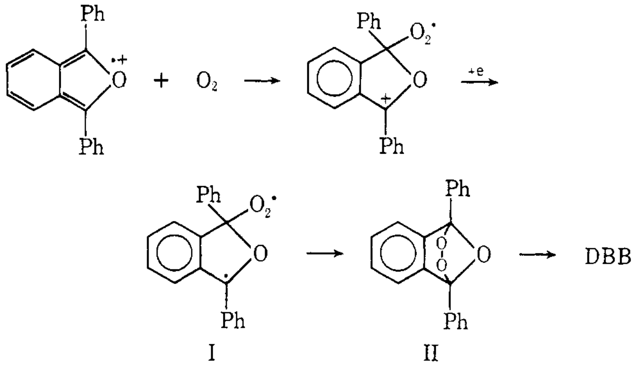
Pulse Electrolysis of Ferrocene, Oxygen, and DPBF. A solution of DPBF (0.55 mmol) and ferrocene (0.55 mmol) was saturated with oxygen in the usual manner. A mercury pool was utilized as the working electrode, because filming occurred during pulsing at the platinum electrode, so that a continuous current could not be maintained. This filming was also observed with a mercury pool but slow stirring with a magnetic stirring bar kept the electrode surface renewed. The potential was pulsed from +0.40 to −0.9 V at 1.0 cps for 4 hr. Cyclic voltammetry of ferrocene exhibited the expected Nernstian behavior for its oxidation with $E_{1/2}$ = +0.30 V.[13] The reduction of ferrocene was not observed up to −2.0 V. The reduction of oxygen is reversible at mercury with $E_{1/2}$ = −0.80 V.[14] DBB (0.30 mmol) was isolated from the reaction mixture.

Control Experiments with DPBF, Oxygen, and Ferrocene. The identical conditions employed in the pulsing experiment were utilized except a potential was not applied. Product analysis indicated that DB (0.06 mmol) was formed.

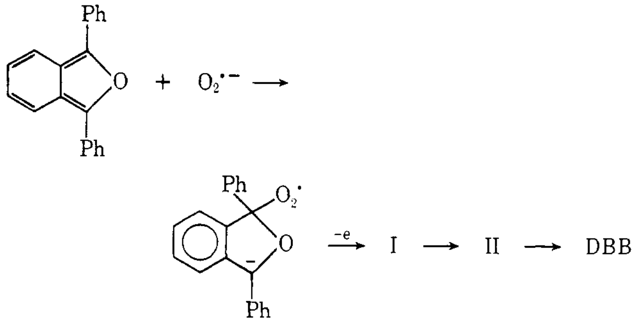
Discussion

The experimental data show that the direct reaction between DPBF and oxygen occurs at a rather slow rate but the yield of DBB is definitely increased when a potential pulse program is applied. In the case in which DPBF was utilized as both the cation radical and quencher, there are several mechanisms which could explain the results obtained. The following shows three such mechanisms.

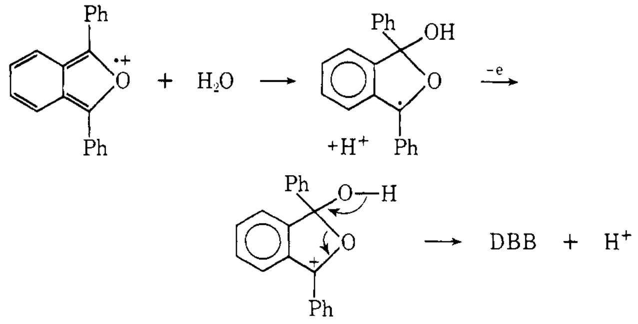
Mechanism a



Mechanism b

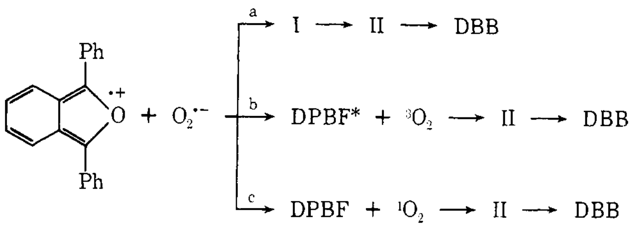


Mechanism c



The common feature of the preceding mechanisms is that they only utilize one of the electrogenerated species, either DPBF cation radical or superoxide. The two control experiments were performed in such a manner that the potential pulse could only generate the DPBF cation radical or superoxide but not both. It can be seen that only the control amount of DBB was formed. This indicates that in order to increase the yield of DBB, both DPBF cation radical and superoxide must be present in solution, thus eliminating the preceding mechanisms.

Again, there are several ways in which DPBF cation radical and superoxide may react to form DBB. These mechanisms are outlined below.



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(12) R. Dietz, M. E. Peover, and P. Rothbaum, *Chem.-Ing.-Tech.*, **4**, 185 (1970).

(13) T. Kuwana, D. E. Bublitz, and G. Hok, *J. Amer. Chem. Soc.*, **82**, 5811 (1960).

(14) M. E. Peover and B. S. White, *Electrochim. Acta*, **11**, 1061 (1966).

Direct elimination of any of these three possibilities is difficult. For pathway a, the production of the intermediate moloxide by an alternate method is rather difficult, and for b and c, it has been shown that these intermediates do in fact yield DBB. What may be tested, however, is the relative contributions of any of the three processes to the overall yield. It is apparent that for processes a and b, the DPBF cation radical must be the oxidant, since its presence is required to form either the moloxide or DPBF*. In process c, the DPBF cation radical is participating only as a chemical oxidant and quencher. If pathways a and b contribute heavily to the overall process, then their absence or presence will have a great effect upon the yield. This was tested by using ferrocene cation radical as the oxidant.

$$\text{Ferr}\cdot^{+} + \text{O}_2\cdot^{-} \longrightarrow \text{Ferr} + {}^{1}\text{O}_2 \xrightarrow{\text{DPBF}} \text{DBB}$$

The oxidation potential of ferrocene is approximately 350 mV less positive than DPBF's potential so that DPBF cation radical could not be generated and the moloxide and DPBF*, therefore, could not be formed. The experimental data show essentially the same yield of DBB was isolated. This indicates that contributions by pathways a and b are minimal.

This, of course, assumes that the ferrocene system proceeds *via* the singlet oxygen mechanism. There are two possible mechanisms which had to be considered for this system. These are outlined below.

$$\text{(a)}\quad \text{Ferr}^{+} + \text{O}_2\cdot^{-} \longrightarrow \text{Ferr}^{*} + {}^{3}\text{O}_2$$

$$\text{Ferr}^{*} + \text{DPBF} \longrightarrow \text{Ferr} + \text{DPBF}^{*}$$

$$\text{DPBF}^{*} + {}^{3}\text{O}_2 \longrightarrow \text{DPBFO}_2 \longrightarrow \text{DBB}$$

$$\text{(b)}\quad \text{Ferr}\cdot^{+} + \text{O}_2\cdot^{-} \longrightarrow \text{Ferr} + {}^{1}\text{O}_2$$

$${}^{1}\text{O}_2 + \text{DPBF} \longrightarrow \text{DPBFO}_2 \longrightarrow \text{DBB}$$

The first mechanism, a, may be eliminated by energy considerations. The source of the excitation energy can only come from the exothermic annihilation reaction. An empirical formula used to obtain an estimate of the energy available from such an annihilation reaction is shown below.[15]

$$-\Delta H^\circ = E_\text{p}(\text{R}/\text{R}\cdot^{+}) - E_\text{p}(\text{R}/\text{R}\cdot^{-}) - 0.16\ \text{eV}$$

Here ΔH° represents the standard enthalpy, E_p represents the peak potentials of two half-reactions, and the final term is an entropy contribution. The lowest excited state of ferrocene is the triplet state which is 2.4 eV above the ground state.[16] The enthalpy of the ferrocene cation radical–superoxide ion redox couple is 1.0 eV and, therefore, is not energetically sufficient to produce an excited ferrocene molecule. Therefore, the weight of the evidence presented strongly indicates that singlet oxygen is the reactive intermediate and that it is generated from an electron transfer reaction between a cation radical and superoxide ion.

The efficiency for the production of singlet oxygen was roughly estimated by first calculating the coulombs for a single pulse and then multiplying by the total number of pulses utilized. It was assumed that the current decayed as $t^{-1/2}$ so that the integrated form of the Cottrell equation could be used. The coulometric efficiency was then the number of faradays required to produce the amount of DBB isolated divided by the total number of faradays passed. For the ferrocene and oxygen system, the efficiency was found to be about 4%. Although this is probably not a very accurate measurement since there is some error in the measured current because of double layer charging, background current, and loss of starting material, it does fall in the range of efficiencies for other ecl processes.[17]

At the beginning of this investigation, other cation radical sources were tried, but no singlet oxygen was chemically detected. Tetraanisylethylene (TAE), 1,1′,-2,2′-tetramethylvinylene-3,3′-diindolizine (TMI), 10-methylphenothiazine (10-MP), and 9,10-diphenylanthracene (9,10-DPA) were all tried as sources of oxidants. The first two yielded stable cation radicals and dications[18,19] upon electrochemical oxidation while the latter two gave stable cations only.[20,21] For TMI and 10-MP, a different chemical.quencher, 2,3-dimethyl-2-butene (TME), was used and in both cases no 2,3-dimethyl-3-hydroperoxy-1-butene, the known product of singlet oxygen and TME,[22] was isolated. There are several possible explanations for this. First, TME has been reported to quench 20 times more slowly than DPBF,[7] so that with equal reaction times, the yield of the hydroperoxide should be very low. Also, in the case of 10-MP the singlet oxygen was probably quenched by 10-MP, since this behavior has been previously observed for amines.[23] In the case of 9,10-DPA and TAE, they were utilized as both cation radical source and quencher. Again, no products were observed and a possible explanation is that 9,10-DPA and TAE are poor singlet oxygen quenchers. For 9,10-DPA β (decay rate of oxygen/reaction rate) in benzene was found to be 0.05 M[24] while for DPBF in methanol β was 10^{-4} M.[25] The lifetime of singlet oxygen is approximately three times greater in benzene than in methanol, while the reaction rate constant for DPBF was not found to be dependent upon solvent.[7] Therefore, from the ratio of β's, DPBF is approximately 1000 times more reactive than 9,10-DPA. Moreover, 9,10-DPA has a limited solubility in acetonitrile, approximately 1 mM. Therefore, to use 9,10-DPA as a singlet oxygen trap one would have to pulse at least 1000 times longer than when DPBF is used to obtain the same yield of product. The oxidation potentials of TAE, 10-MP, and 9,10-DPA are positive enough that DPBF would be oxidized concurrently, thereby eliminating the possibility of using DPBF as the quencher with these cation radical sources.

Acknowledgment. The support of this research by the U. S. Army Research Office, Durham, and the National Science Foundation (GP-31414X) is gratefully acknowledged.

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