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STUDIES IN AROMATIC AND AMINE NITRATION

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Final Report

By: David S. Ross and William Blucher

Covering the Period May 3, 1976--November 30, 1979

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Piperidene
Nitrogen Tetroxide
Nitric Oxide
Oxygen
Cerium (IV)
Electron Transfer
Hydrogen Peroxide
Phenol
Oxidation Potential

20. ABSTRACT (*Continue on reverse side if necessary and identify by block number*)

Toluene and benzene are nitrated very rapidly at 0°C in liquid nitrogen tetroxide when NO and O_2 are simultaneously bubbled into the stirred system. Without the added gases no nitration takes place. Dinitrobenzenes (DNBs) are formed when benzene is the substrate, while nitrobenzene itself is not nitrated under the reaction conditions. The DNB isomer ratio is consistently o:m:p = 11:62:27, in contrast to 7:91:2 observed commonly in mixed acid nitrations. Small quantities of trinitrobenzenes are observed as well. With toluene, only nitrotoluenes are observed, with an isomer ratio o:m:p = 55:4:41. Experiments

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20. Abstract (Concluded)

with piperidine in N_2O_4 show that nitrosoamine is the major product, but the small quantity of nitramine observed in runs without O_2/NO is increased substantially with the addition of the two gases. A shock sensitive substance consistent with the structure $CH_2=NNO_2$ is isolated in runs with hexamine and an NO-rich O_2/NO mixture. When the gas stream is O_2-rich, a mixture of stable products is isolated, the major component of which is 1,3-dinitro-5-nitroso-perhydrotriazine. For mixed acid nitration it can be shown that an estimate of the increase in oxidation potential of nitric acid with increasing acidity parallels the increase in nitration rate over the same acidity range about as well as does the H_R + log a_w function. In accord with an electron transfer nitration model, experimental results show that the addition of H_2O_2 to phenol nitrations in 56.2% sulfuric acid suppresses the nitration severely. However, the addition of both hydrogen peroxide and N(III) to the system substantially accelerates the nitration to rates greater than that in the absence of both hydrogen peroxide and N(III). In the nitration of phenol in 56.2% sulfuric acid both nitrous acid and Ce(IV) promote the process. The conventional explanation for nitrous acid catalysis cannot play a role since in this work o:p = 0:78, whereas phenol nitrosation yields virtually solely the p-isomer. The promotion by Ce(IV) suggests a radical cation role in aromatic nitration.

FOREWORD

SRI International has conducted a program of research for the Army Research Office under Contract No. DAAG29-76-C-0040. The research effort included laboratory studies of aromatic nitration and amine nitrolysis, and dealt with the question of the validity of the currently accepted mechanism of aromatic nitration in mineral acid media.

This final report is presented in four sections. Sections 1 and 2 describe our work dealing with nitrations and nitrolyses of aromatic and amine substrates, respectively, in liquid nitrogen tetroxide. Section 3 discusses our new view of mixed acid nitration in terms of electron transfer chemistry, and Section 4 deals with HONO and Ce(IV) catalysis of nitration.

In the Appendix is a copy of a manuscript on the HNO_2 catalysis to appear in Chemical Communications.

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AROMATIC NITRATIONS IN LIQUID NITROGEN TETROXIDE*

Introduction

Nitrogen tetroxide (N_2O_4) reacts very slowly with benzene. Only after reaction at high temperature or for very long periods is some nitration observed.[1] ~~We~~ IT IS reportED here on nitrations in N_2O_4 systems that are rapid at 0°C and IT appears to represent a new route to aromatic nitration. As ~~we shall~~ IT WILL BE discussED, the rates of aromatic nitration as well as the distribution of isomers in products from the system are significantly different from those obtained classically in mixed acid systems.

Results and Discussion

This work began with the attempted generation of N_2O_5 in the presence of an aromatic substrate, after the work of Ogg,[2] who suggested that N_2O_5 could be prepared in the gas phase in $NO/NO_2/O_2$ mixtures through the sequence

$$O_2 + NO \rightleftarrows O_2NO \qquad (1) \qquad\qquad O_2NO + NO_2 \rightleftarrows NO_2 + NO_3 \quad (2)$$

$$NO_2 + NO_3 \rightleftarrows N_2O_5 \qquad (3)$$

In our initial efforts, we were unable to repeat his results, and using the latest kinetic data available for equations (1)-(3),[3] we established that, contrary to his conclusions, Ogg could not have had significant quantities of N_2O_5 in his gas phase, ir experiments. We suggest that he may have mistaken the ir bands of HNO_3 for those of the pentoxide. The two species have very similar gas phase ir spectra, and in our experience HNO_3 was always seen despite rigorous drying of our gases and all equipment.

We then investigated the same chemistry in liquid N_2O_4, where work was performed in the apparatus shown in Figure 1. N_2O_4 is an easily handled liquid, boiling at about room temperature. Usually the experiments were performed under 2 atm of O_2, with NO bubbled into the rapidly stirred liquid at such a rate that the developing green color of N_2O_3

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*A portion of this work was presented at the 174th Meeting of the American Chemical Society, Chicago, Illinois, in August 1977. A full manuscript is in preparation.

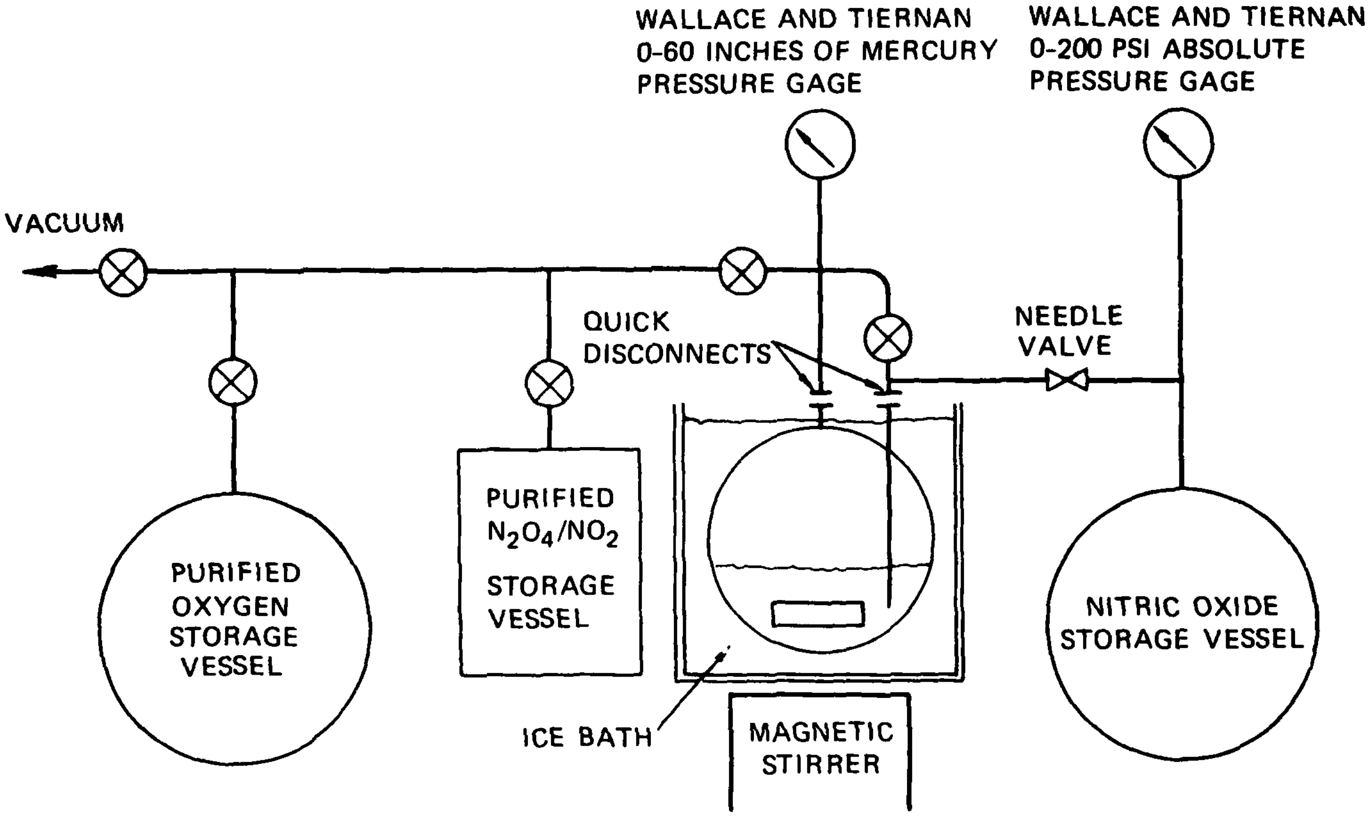

FIGURE 1 ASSEMBLY FOR N_2O_4 NITRATIONS

faded at the tip of the NO inlet tube. Benzene was the first substrate studied; the results for a series of runs with and without oxygen are shown in Figure 2. Results for other agents are shown as well for comparison.

In the absence of oxygen, only N_2O_5 provided significant nitration. In each case the reaction had ceased within 15 min, and surprisingly, nitric acid failed to yield more than about 15% nitration. Runs for longer periods and with increased nitric acid quantities yielded only little more nitration, with the nitration stopping long before all the HNO_3 (the limiting reagent) was consumed. This curious result is an area of research we propose to pursue in the next research period. It is also seen that neither NO nor water is very active.

The results with oxygen are strikingly different. Nitrogen pentoxide appears to be somewhat more active, and water itself yields about 40% nitration. However, nitric acid and NO are clearly very active, providing yields in excess of 100% of nitrated products. Again, the nitric acid results are curious, and we propose to direct our attention to that system. Their similarity to the NO results suggests a similar reaction path. Our interest, however, was focused on the NO results.

In the NO work, both di- and trinitrobenzenes are formed in addition to nitrobenzene, and the yield in excess of 100% suggests a chain process for the nitration. Also, the superior performance of the NO system relative to that of N_2O_5 clearly suggests that the pentoxide is not an intermediate in the nitration process.

The results for other substrates studied with the NO/O_2 system are shown in Table 1. Comparison of runs 1 and 2 shows that the addition of NO/O_2 to the system results in the nitration of 60% of the starting benzene. Based on the limiting quantities of NO used, the yield value is actually higher; the % yield column in the table displays yields for most of the runs calculated in this way.

The yields in excess of 100% for benzene suggest the operation of a short chain process in which dinitrophenol is a common side product. For runs 3, 4, and 5, the yields with respect to NO are high, but the total quantities of benzene nitrated are small. Although no attempt was made

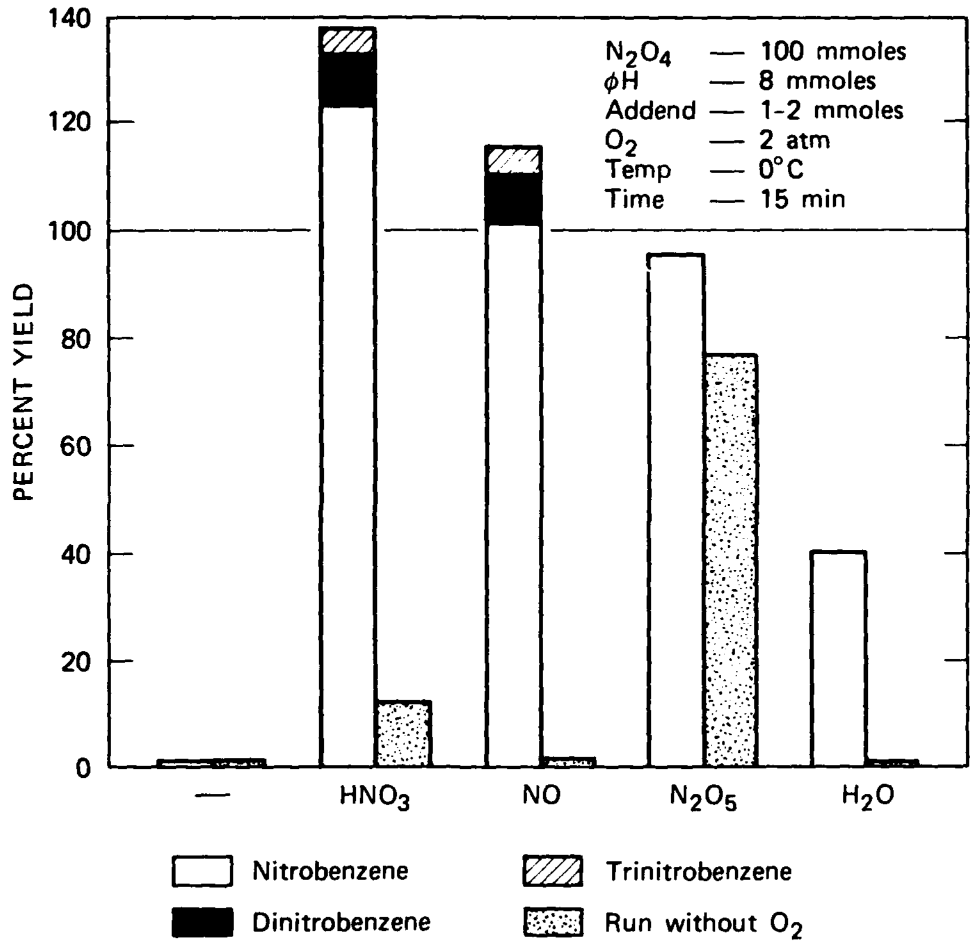

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FIGURE 2 RESULT FOR A SERIES OF NITRATION IN LIQUID N_2O_4

The percent yields are calculated in terms of the limiting reagent.

Table 1

SUMMARY OF *RESULTS FOR NITRATION OF SOME AROMATIC* SUBSTRATES IN LIQUID N_2O_4 AT 0°C FOR 15 MIN[a]

Run No.	Aromatic (mmoles)	Aromatic/N_2O_4[b]	Oxidant	% Yield[c]	mono:di:tri Nitroaromatic Products	Remarks	
1	PhH (2)	0.02	--	< 1	--	--	
2	PhH (3)	0.03	NO/O_2	90	89:11:+2	Ratio of o:m:p *for DNB* ≈ 11:62:27 for most runs	
3	PhH (102)	1.0	NO/O_2	118	88:12:+2		
4	PhH (202)	10.1	NO/O_2	106	87:11:1		
5	PhH (160)	> 160	NO/O_2	30	63:34:3		
6	PhH (11)	0.1	Co(III)[d]	20	90:8:2		
7	$PhCH_3$ (2)	0.02	--	< 1	--	--	
8	$PhCH_3$ (2)	0.02	NO/O_2	99+	No di- or tri- nitrotoluenes *observed*	55:4:41	Ratio of o:m:p *for* nitro-toluenes
9	$PhCH_3$ (100)	1.0	NO/O_2	98		58:4:38	
10	$PhCH_3$ (200)	10	NO/O_2	97		51:10:41	
11	$PhCH_3$ (100)	>100	NO/O_2	20		53:20:26	
12	$PhCH_3$ (10)	0.1	Co(III)[d]	$\sim$ 2		44:32:26	
13-15	$PhNO_2/PhCF_3$	- No nitration observed[e] -					

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[a] The quantities of NO and oxygen used were 2 and 20 mmoles, respectively.

[b] Molar ratio.

[c] Based on the limiting reagent in each case. The limiting reagent for runs 1 and 7 were benzene and toluenes respectively, Co(III) for the runs where it was used, and NO for all *the others.*

[d] As the acetate. For runs 6 and 12, 3.2 and 5.0 mmoles were used, respectively; however, it was not fully soluble in the reaction medium for both reactions.

[e] Starting substrates recovered quantitatively.

to maximize the process, in some runs with larger quantities of NO and limiting quantities of oxygen, 20%-30% of the benzene was nitrated. In these cases the dinitrobenzene yields exceeded those of nitrobenzene. The TNB yields were always small and highly variable. The system is complicated, however, and no full study was made of the effects of all the variables.

In summary,

- The system works for benzene and toluene, but not for substrates with electron-withdrawing groups.
- Benzene yields not only nitrobenzene, but also di- and trinitrobenzene under conditions where nitrobenzene itself is not nitrated.
- The relative yield of dinitrobenzene increases with increasing quantities of benzene in the reaction mixture.
- The o:m:p ratio for the dinitrobenzene formed is consistently ∿11:62:27, in contrast to the reported ratio for the nitric acid nitration of nitrobenzene of 7:91:2.[4]
- All three trinitrobenzenes are observed in almost equal quantities.
- Toluene in the system yields consistently only mononitrotoluenes under conditions where benzene is converted to polynitrobenzenes.
- The nitrotoluenes formed at low initial toluene concentrations are in the classical o:m:p ratio,[4] and with increasing toluene concentration, the ratio shifts to one with relatively large quantities of meta isomer.
- Both benzene and toluene are oxidized by Co(III) to their radical cations,[5] and the substitution of Co(III) for NO/O_2 provides products similar to those obtained in the NO/O_2 runs.

It seems clear that the familiar nitronium ion nitration is not operative in these systems and that there must be some radical character to the process. However, a straight free radical nitration cannot be operative, since we observe for toluene primarily o/p nitration, and the free radical nitration of toluene yields mainly m-substitution.[6]

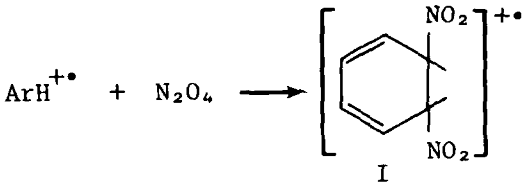
The fact that benzene yields increasing quantities of polynitrobenzens when its concentration in the reaction medium is increased is distinctly contrary to what would be expected for a simple substitution reaction. Also since toluene is more reactive than benzene toward electrophilic substitution,[7] it is notable that the benzene, not toluene, is polysubstituted. Finally, it would appear that like Co(III), the NO/O_2 system can act as a one-electron oxidant.

The results are complicated, and no clear-cut reaction path has currently been developed consistent with all the observations. Clearly, however, we are dealing with a nitration process significantly different from those commonly studied, and we can propose the following scheme to explain most of the data. The scheme remains tentative,pending confirmation through further research.

It is recalled that equations (1) and (2) yield NO_3. We suggest that NO_3 in these systems can act as an effective one-electron oxidant.*

$$NO_3 + ArH \rightarrow NO_3^- + ArH^{+\bullet} \qquad (4)$$

Then the radical cation adds molecular N_2O_4, yielding a cyclohexadienyl intermediate I



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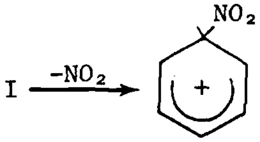
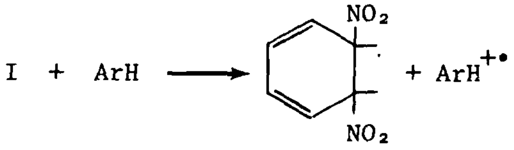
(5)

which can either lose NO_2 to go to the common Wheland intermediate, and thence to $ArNO_2$



(6)

or undergo *electron transfer with a second molecule of aromatic* substrate to yield dinitrocyclohexadiene



(7)

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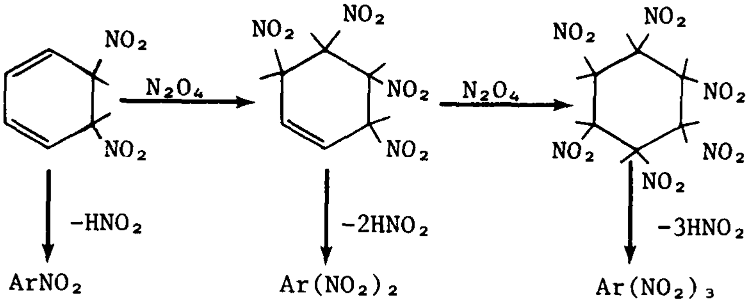
[*] Although we write equation (4) as a simple electron transfer, it is likely that free ions are not formed. Because of the low dielectric constant for N_2O_4, ion pairs or clusters are probably formed.

The diene can then go through a series of N_2O_4 additions and HONO eliminations to provide mono-, di-, and trinitrobenzenes



(8)

The competition between equations (6) and (7) explains why benzene is polynitrated, while toluene is not. The Wheland intermediate formed from toluene is expected to be relatively more stable than that formed from benzene, and the ratio of rates for steps (6) and (7), $k_6/k_7(ArH)$, must be significantly greater for toluene than for benzene. Thus for a given starting concentration of substrate for the two cases, a major portion of the toluene follows path (6), while the benzene is more equally partitioned. For this scheme the product ratio depends on the relation

$$\text{polynitro/mononitro} = k_7(ArH)/k_6$$

and it is therefore clear that an increase in the concentration of benzene increases the relative quantity of polynitrobenzenes in the product mixture.

Consistent with this discussion, the shift in nitrotoluene isomer distribution with change in toluene concentration can be explained by the fact that at the extreme condition where the toluene/N_2O_4 ratio is > 100, reaction (7) begins to compete with (6), and the sequence following them predominates. Apparently for the case of toluene, the formation of polynitro products is less favorable than loss of HONO from the methyl dinitrocyclohexadiene intermediate.

A one-electron transfer scheme is particularly attractive in application to the results we obtained with substrates classically considered highly reactive to electrophilic substitution.[7] We found that in accord with reported work, phenol,[8a] anisole,[8b] and mesitylene[8c] react with N_2O_4 at 0°C without the necessity of added NO/O_2. Results of runs both with and without NO/O_2 are presented in Table 2.

Table 2

REACTION OF REACTIVE SUBSTRATE
IN LIQUID N_2O_4 AT 0°C FOR 15 MIN

	Products, % Yield	
Substrates	N_2O_4 Alone	With NO/O_2
Phenol	98^+% 2,4-dinitrophenol	98^+% 2,4-dinitrophenol
Mesitylene	56% nitromesitylene	85% nitromesitylene
Anisole	60% nitroanisole (o/p = 0.66)	98^+% nitroanisole (o/p = 0.66)

The phenol is fully converted under both sets of conditions to 2,4-dinitrophenol. Traces of mononitro product were seen by both gas chromatography and high pressure liquid chromatography. The anisole and mesitylene are converted partially to their mononitro derivatives, and the addition of NO/O_2 substantially increases the conversion for both cases. Notable for anisole, the o:p ratio does not change, and thus similar chemistry must be operative in both cases. Also significant is the quantitative formtion of nitroanisole, as opposed to the common observation in mineral acid anisole nitrations of major conversion to nitrophenols.[9a]

It has been suggested that the mechanism of nitration in these reactive substrates in N_2O_4 includes initial nitrosation, followed by oxidation of the nitroso derivative to the nitro compound.[8b] This mode of reaction is suspect, however, since the purposeful initial nitrosation of anisole in sulfuric acid yields a product with an o:p = 0.06,[9b] in strict contrast to both our N_2O_4 results and those of Underwood et al.[8c][*] On the other hand the substitution cannot be through a nitronium ion scheme, since Bonner and Hancock[8b] found that N_2O_4 was far superior to nitric acid in anisole nitration in CCl_4.

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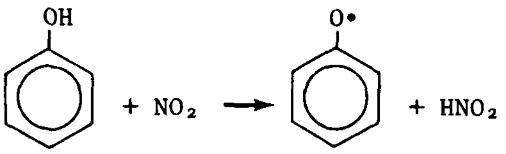
[*] It is expected that a change in medium could affect the o/p ratio in the product mixture. However, the overwhelming differences discussed here can only be due to different mechanisms.

It might be suggested that at least the mononitration of phenol proceeds via the phenoxyl radical.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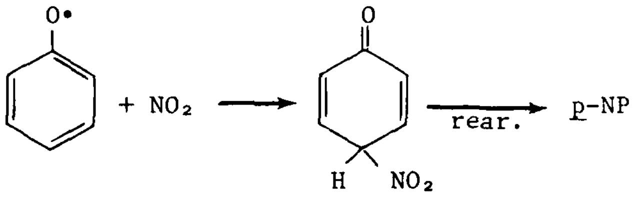
In contrast to the nitration of toluene in a free radical process, this intermediate should readily be substituted in the o- and p- positions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This scheme is not possible at 0°C, however, since the abstraction in equation (9) is endothermic by about 7 kcal/mol; i.e., D(H-ONO) = 78 kcal/mole and D(H-OPh) = 85 kcal/mol.[10a]

Since NO/O_2 addition in these systems increases the yield but does not change the product isomer distribution, and since we have shown that NO/O_2 systems can act as one-electron oxidants, perhaps via NO_3, we can conclude that a radical cation scheme similar to that discussed above can be operative for these reactive substrates.

An overall view of the process then is one in which aromatic substrates with increasingly greater ionization potentials in turn require increasingly more potent oxidants for successful nitration. The mononitration of mesitylene and anisole under conditions where benzene is dinitrated is consistent with the similar finding for toluene and is in line with the expected highly stabilized Wheland intermediates for these compounds.

A puzzling aspect of the result is the full dinitration of phenol. In a preliminary experiment we established that p-nitrophenol was converted quantitatively to 2,4-dinitrophenol in liquid N_2O_4 under the above conditions in the absence of NO/O_2. Possibly the acidity of the substrate might be promoting the process in some way. It is known, for example, that the oxidizing potential of N_2O_4 is promoted by acid[10b]

$$N_2O_4 + 2H^+ + 2e^- \rightarrow 2HNO_2$$

and investigations along these lines are continuing.

Section 2

REACTIONS OF AMINES IN $N_2O_4/NO/O_2$ SYSTEMS

Our favorable results with aromatic substrates (Section 1) suggested that it would be feasible to synthesize nitramines in the same system. It has been reported that secondary amines react readily with N_2O_4, present as the limiting reagent, in various solvents to form either equimolar quantities of nitrosamine and the nitrate salt of the starting amine

$$2R_2NH + N_2O_4 \xrightarrow{0°C} R_2N\text{-}NO + R_2\overset{+}{N}H_2NO_3^- \qquad (11)$$

or, at low temperatures, the nitramine and nitrite salt[11]

$$2R_2NH + N_2O_4 \xrightarrow{-80°C} R_2N\text{-}NO_2 + R_2\overset{+}{N}H_2NO_2^- \qquad (12)$$

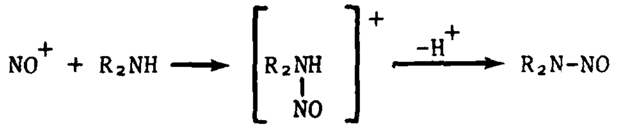
When the N_2O_4 is present in excess, however, the nitrosamine is formed fully in runs at 25°C, with traces of nitramine observed.[12] Apparently the nitrate salt in equation (11) can react with N_2O_4 to yield more nitrosamine.

In work related to that described above, we performed experiments with piperidine as substrate in liquid N_2O_4 at 0°C. The results are presented in Table 3.

In N_2O_4 alone, piperidine is converted to nitrosamine with small quantities of nitramine in the product mixture. The addition of NO/O_2 increases the quantity of nitramine somewhat, but the nitrosamine itself is not oxidized by the system. Piperidine and NO are unreactive under these conditions; however, piperidine reacts with NO/O_2 to form only nitrosamine.

A simple scheme for the nitrosation of the amine through NO^+ explains most of the data

$$N_2O_4 \rightleftarrows NO^+ + NO_3^- \qquad (13)$$



$$NO^+ + R_2NH \longrightarrow \left[\begin{array}{c} R_2NH \\ | \\ NO \end{array}\right]^+ \xrightarrow{-H^+} R_2N\text{-}NO \qquad (14)$$

Table 3

REACTIONS OF PIPERIDINE AND NITROSOPIPERIDINE IN VARIOUS NO_x SYSTEMS AT 0°C/15 MIN

System	Substrate	Product[a]
N_2O_4	Piperidine	94% nitrosamine 6% nitramine
$N_2O_4/NO/O_2$	Piperidine	85% nitrosamine 15% nitramine
$N_2O_4/NO/O_2$	Nitroso-piperidine	No reaction
NO/O_2[b]	Piperidine	100% nitrosamine[c]
NO	Piperidine	Very slow reaction[d]

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[a] The overall product recoveries in all cases are > 95%.

[b] No N_2O_4 was used. NO was bubbled into stirred piperidine under 2 atm O_2. Some N_2O_4 was seen near the end of the reaction.

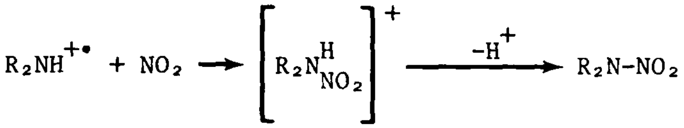
[c] Quantitative recovery of nitrosamine.

[d] See reference 12.

However, the formation of some nitramine and the increase in nitramine production with the addition of NO/O_2 suggest that possibly electron transfer reactions compete with reactions (13) and (14.

Thus, since simple amines have ionization potentials in the range 8.2-8.3 eV,[13] similar to that for anisole, they should be very active in these systems. We can therefore propose

$$NO_2 + R_2NH \longrightarrow NO_2^- + R_2NH^{+\bullet} \tag{15}$$



$$R_2NH^{+\bullet} + NO_2 \longrightarrow \left[R_2N\overset{H}{\underset{NO_2}{}} \right]^+ \xrightarrow{-H^+} R_2N\text{-}NO_2 \tag{16}$$

An alternative route for the formation of nitramine has been proposed.[12]

$$NO_2 + R_2NH \longrightarrow HONO + R_2N^\bullet \tag{17}$$

$$R_2N^\bullet + NO_2 \longrightarrow R_2N\text{-}NO_2 \tag{18}$$

This process, however, cannot be operative in these systems, at least at 0°C, because equation (17) is endothermic by about 17 kcal/mol.* Moreover, the chemistry in equation (15) cannot continue along the line

$$R_2NH^{+\bullet} + NO_2^- \longrightarrow R_2N^\bullet + HONO \tag{19}$$

since the sum equations (15) and (19) is equation (17). The radical cation is therefore a stable component of the reaction.[14]

In other work still in progress, we studied the reactions of hexamine (hexamethylenetetramine) in these systems. Hexamine is the precursor to HMX and RDX in treatment with nitric acid, ammonium nitrate, acetic acid, and acetic anhydride.[16]

Our current results in runs for 2 hr at 0°C include the observation that hexamine in N_2O_4 alone or under 2 atm O_2 yields a small quantity of a solid that quickly decomposes upon standing. The same solid is obtained as a major product when NO is bubbled into the N_2O_4/O_2 system at a relatively

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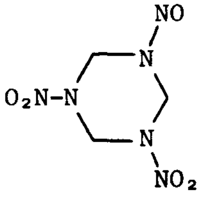
*The appropriate thermochemistry for the calculation is available: $D(H\text{-}NR_2) = 95$ kcal/mol;[15] $D(H\text{-}ONO) = 78$ kcal/mol.[10a]

rapid rate such that the green color from the initial formation of N_2O_3 fades only slowly with stirring. The solid is very shock sensitive and, with DMSO-d_6 as solvent, it yields a single NMR peak at 4.2 ppm. Its elemental analysis is consistent with that for both HMX and RDX, but since its NMR spectrum is different from that for either of those compounds, which provide singlets at 6.10 ppm and 6.17 ppm, respectively, we have tentatively considered the product to be methylenenitrimine, $H_2C{=}N{-}NO_2$. Methylenenitrimine has never before been prepared although many attempts have been made.[17] The firm identification of our product is in progress.

In further experimental work, hexamine was treated with N_2O_4/O_2 into which NO was bubbled at a very slow rate, so that no green color developed. In this case, all the hexamine was converted to a mixture of stable products, the major component of which has been identified as I, the nitroso-dinitro analog of RDX.



I

Compound I was prepared by Brockman et al.[18] from the known trinitroso-RDX analog by treatment with H_2O_2/HNO_3. We have currently not identified any of the other mixture components, but have speculated, on the basis of the high pressure liquid chromatography profiles, that similar nitroso-nitro analogs of HMX may be present. This work is continuing.

The chemistry of these systems needs to be developed. In the hexamine work, the observation of Compound I demonstrates that significant nitrolysis (nitramine formation) takes place, in some contrast to the results for our piperidine work. The formation of Compound I under conditions considerably less vigorous than those used by others suggests that the chemistry of this system has potential in practical application.

The chemistry itself is novel and of interest. The hexamine system provides access to both 6- and 8-membered ring systems, and it would be of interest to establish the effects of both one-electron oxidants such as Ce(IV) and oxidation via anodic electrochemical procedures on the ultimate 6/8 product distribution. This area of research will be pursued in our current program.

Section 3

AROMATIC NITRATION IN MINERAL ACID MEDIA

Current View

The question of the intermediacy of nitronium ion (NO_2^+) in aromatic nitration in mineral acids has been considered for more than two decades, and it is generally agreed that the nitroaromatic products arise from electrophilic substitution by the nitronium ion on the aromatic substrate.

$$HONO_2 + H^+ \rightleftarrows H_2O + NO_2^+ \qquad (20)$$

$$NO_2^+ + ArH \longrightarrow ArHNO_2^+ \xrightarrow{-H^+} ArNO_2 \qquad (21)$$

The cornerstone in the collective evidence in favor of this scheme is the work of Westheimer and Kharasch, who showed that the rate of nitration of nitrobenzene in nitric acid/sulfuric acid media peaked at about 90% sulfuric acid and ran parallel in the lower acid region to the equilibrium ionization of trinitrotriphenylcarbinol.[19] This work led to the development of the H_R (or J_o) acidity function,[†] which defines the proton-donating power of a medium in terms of the equilibrium,[20]

$$ROH + H^+ \rightleftharpoons R^+ + H_2O$$

so that

$$-H_R = \log \frac{a_{H^+}\,\gamma_{ROH}}{a_w\,\gamma_{R^+}} \qquad (22)$$

[†] As a point of interest historically, we note that Westheimer and Kharasch stated that their work, presumably because of its impact on the war effort, "...although sufficiently precise to support the statements above, is not of the highest accuracy; in general the speed with which the work could be accomplished was considered of first importance."

Later nitration work with other substrates confirmed the original finding.[21] The correlation between the nitration rate and the H_R function is stated to establish nitronium ion as the reactive intermediate because of the similarity in equations (20) and (22). More recent work has shown that, in fact, a better correlation exists between the log of the nitration rate and the function (H_R + log a_w).[21] At 25°C, slopes of the straight lines of near unity are found; however, since with an increase in temperature the nitration rate dependence on acidity becomes less steep while the acidity function itself becomes more steep, the value of unity is more coincidental than meaningful.[4,22] Nonetheless, based on the linear correlation, nitronium ion is said to be well established in the mechanism of aromatic nitration.

The case is not without its difficulties, however. Encounter rate kinetics are required to deal with the nitration of reactive substrates, such as anisole, because the nitronium ion has been labeled as the active species at acidities as low as 55% sulfuric acid, where only vanishingly small quantities are present in equilibrium with HNO_3. The question then arises of how so reactive an electrophile, successfully reacting at every encounter, can at the same time be so highly selective, yielding almost exclusively o/p substituted products. The problem has been dealt with by introducing an encounter pair into the mechanism, developed after the rate limiting step, but before the positional selectivity step.[22] However, no direct evidence for the formation of such an intermediate has been offered.

New Approach

Perrin has recently suggested that, for nitration of activated substrates, electron transfer to nitronium ion is the initial step, followed by reaction of the radical cation of the substrate with NO_2.[23]

$$ArH + NO_2^{+} \rightarrow ArH^{+\bullet} + NO_2 \qquad (23)$$

$$ArH^{+\bullet} + NO_2 \rightarrow ArHNO_2^{+} \xrightarrow{-H+} ArNO_2 \qquad (24)$$

Perrin's rationale involved his finding through electrochemical techniques that the anodic half-wave potential for NO_2 in acetonitrile was greater than that for a number of aromatics, and he concluded that equations (23) and (24) were applicable for aromatics more reactive than toluene.

Perrin's observation that NO_2^+ can oxidize reactive aromatic compounds seems to explain the general observation that nitrations of substrates such as phenol, anisole, naphthalene, and similar compounds in mixed acid media lead to the production of nitrous acid.[22] Thus, in terms of Perrin's scheme, some leakage of the NO_2 must occur, and hydrolysis takes place

$$2NO_2 + H_2O \rightarrow HNO_3 + HONO$$

The nitrous acid produced leads, in turn, to catalysis of the nitration, considered to be via prior nitrosation,[4] and the net reaction is an autocatalytic process. (The addition of urea to these systems is said to eliminate the accumulation of the HONO, and no catalysis is observed.)

In some experimental work to be discussed in Section 4, we have recently shown that in HNO_3/H_2SO_4 media, the presence of the 1-electron oxidant Ce(IV) significantly promotes the rate of nitration of phenol, the products being solely o- and p-nitrophenol. Since 1-electron oxidants generate the radical cation of phenol,[24] thus it would seem that the radical cation plays a significant role in the nitration. In that same work, we found that nitrous acid catalyzes phenol nitration through a scheme <u>not</u> involving prior nitrosation. We speculate that, in accord with the oxidation work discussed above, the oxidizing ability of N(III) in the acid system was sufficient to oxidize the aromatic substrate as Ce(IV) did.

It seems clear therefore that a one-electron oxidation, either like that proposed by Perrin or one similar, can be an element in the process of aromatic nitration in mineral acids. There is some similarity here to our discussion of aromatic nitration in liquid N_2O_4. Consistent with these observations is a statement by Latimer[25] that the oxidizing ability of nitric acid increases with increasing acidity, in line with the observation that those substrates difficult to nitrate and requiring higher acidities for nitration are those with electron withdrawing groups, which in turn make the aromatic more difficult to oxidize.

It thus becomes attractive to consider the possibility that the oxidizing power of nitric acid-containing system contributes to the observed nitration kinetic behavior of the system. In electrochemical work on the oxidizing potential of nitric acid systems, Vetter[26] found that the cathodically active component in aqueous nitric acid systems (that is, the electron acceptor) was N(IV) rather than N(V). His scheme for nitric acid acting as an oxidant is

$$HNO_3 + HNO_2 \rightleftarrows H_2O + 2NO_2 \qquad (25)$$

$$2H^+ + 2NO_2 + 2e^- \rightarrow 2HNO_2 \qquad (26)$$

with the net reaction being the two-electron reduction of HNO_2

$$HNO_3 + 2e^- + 2H^+ \rightarrow H_2O + HNO_2 \qquad (27)$$

This scheme is consistent with the earlier data of Monk and Ellingham[27] and later work of Julien and Comtat[28] who established the electromotive potential of an aqueous nitric acid cell containing N(III) as a function of acidity. The Monk and Ellingham measurements were made in media ranging from 1 to 12 M nitric acid. These authors found their observations to be in compliance with the Nernst equation

$$E = E_o + \frac{0.0591}{2} \log \frac{a_{HNO_3} a^2_{H^+}}{a_w a_{HNO_2}} \qquad (28)$$

so that, for example, a plot of E versus the log of the reciprocal of the measured HNO_2 concentration had a slope of 0.0296 for conditions where the other variables were kept constant. E_o for equation (28) is 0.94 V.[25]

Also in good agreement with theory, we have found that if the Monk and Ellingham's acidities are viewed in terms of the Hammett acidity function H_o, then for values of E at various acidities of constant HNO_2, a plot of E versus $-H_o$ (Figure 3) yields a straight line with a slope of 0.060 (in contrast to a theoretical slope of 0.059).

Although there is no reason to expect a linear relationship to exist, much less a slope in accord with the Nernst equation, the fit of the $-H_o$ function

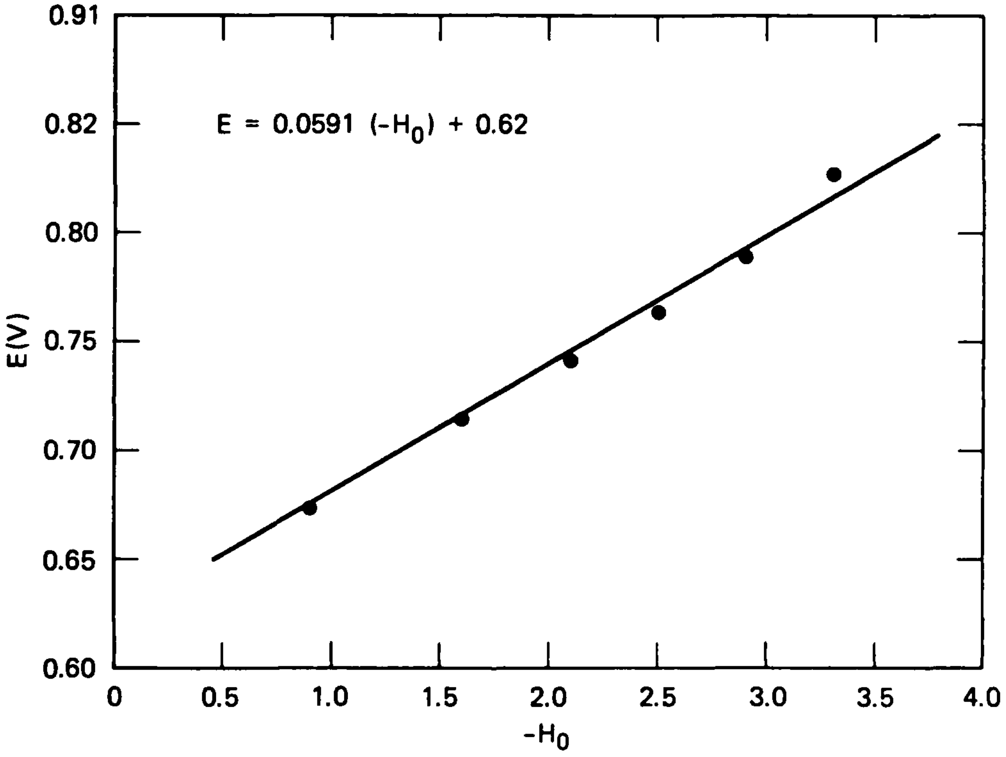


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FIGURE 3 PLOT OF E FROM THE WORK OF MONK AND ELLINGHAM (REF. 27) FOR A NITRIC ACID CELL vs -H_o

in the Nernst expression could suggest that NO_2 behaves like a Hammett base and that NO_2 is protonated before the electron transfer. Thus

$$NO_2 + H^+ \rightleftarrows HNO_2^+ \xrightarrow[e^-]{} HNO_2$$

This suggestion needs to be confirmed.

From the plot, the relationship between the two parameters is

$$E = 0.0591(-H_o) + 0.62 \tag{29}$$

If we equate equations (28) and (29) we get a simple expression for the activity ratio in terms of H_o, that ratio being a function of the properties of the medium in terms of the electron transfer chemistry

$$\log \frac{a_{HNO_3} a^2_{H^+}}{a_w a_{HNO_2}} = \frac{E \times 2}{0.0591} - 31.81 = 2(-H_o) - 10.73 \tag{30}$$

We can now generate values of the log term for acidities greater than those covered by Monk and Ellingham and for other acid systems such as H_2SO_4. We recognize that such an extension must be made with caution since this expression is purely empirical and is developed from limited electrochemical data for a nitric acid cell. Yet we can show that, even with these limitations, there is a surprisingly good correlation when applied to aromatic nitration kinetics in sulfuric acid media.

Figure 4a summarizes the kinetic data for aromatic nitration in HNO_3/H_2SO_4 mixtures over a broad range of acidity, from 40% to 100% sulfuric acid. The data for three substrates covering that acidity range fall well on one curve. At the highest acidities, the rate of nitration actually decreases with increasing acidity.

Figure 4b shows a plot of $-(H_R + \log a_w)$, and the agreement with the nitration rate data is clear. A plot for the activity ratio from equation (30) is shown in Figure 4c, and below 90% H_2SO_4, the electrochemically based function models the reaction about as well as the alcohol indicator function, although it deviates from the experimental curve at high acidities.*

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*Commonly, log-log plots are prepared to make these comparisons. We prefer this type of examination since we feel a better overall view of the fit is provided.

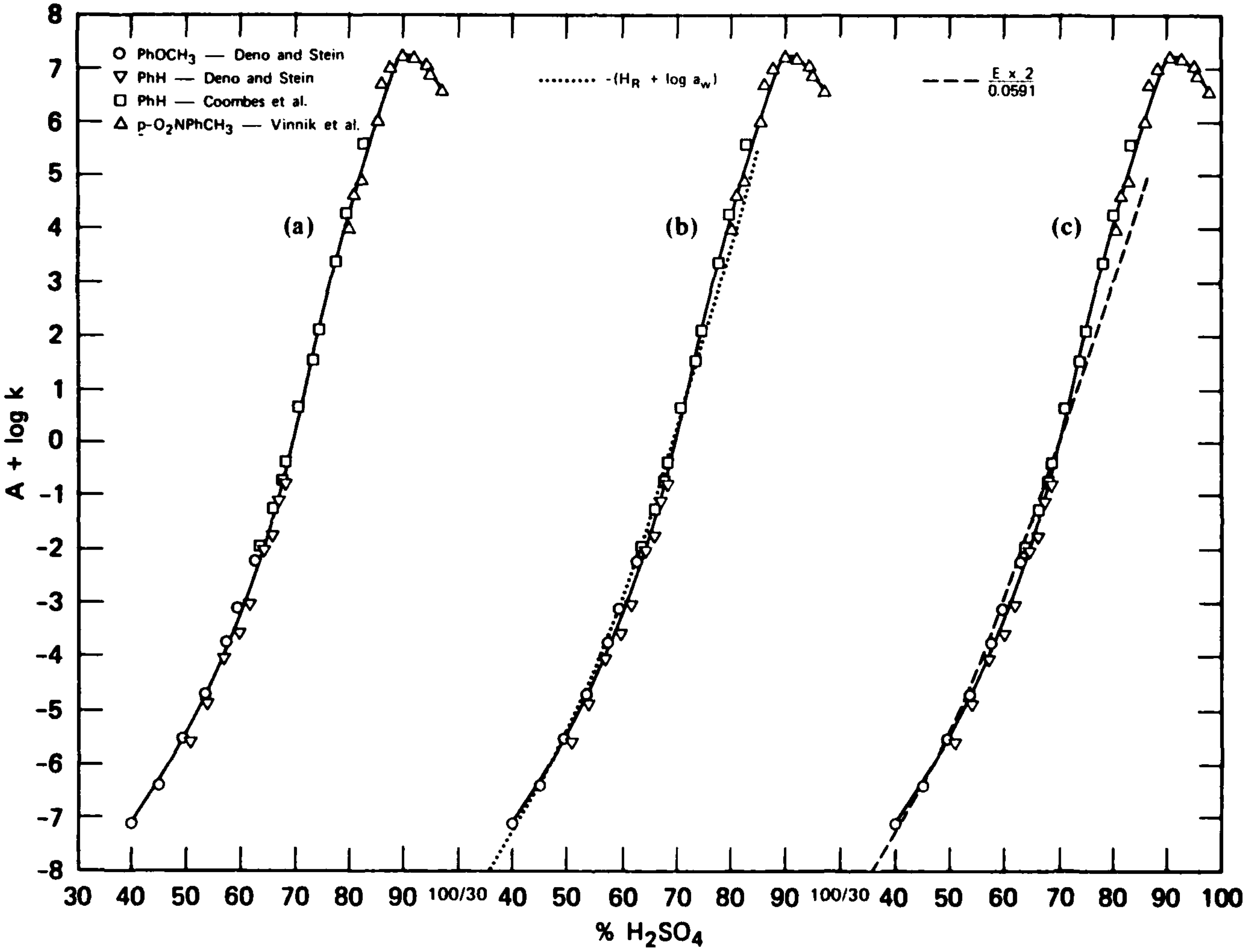

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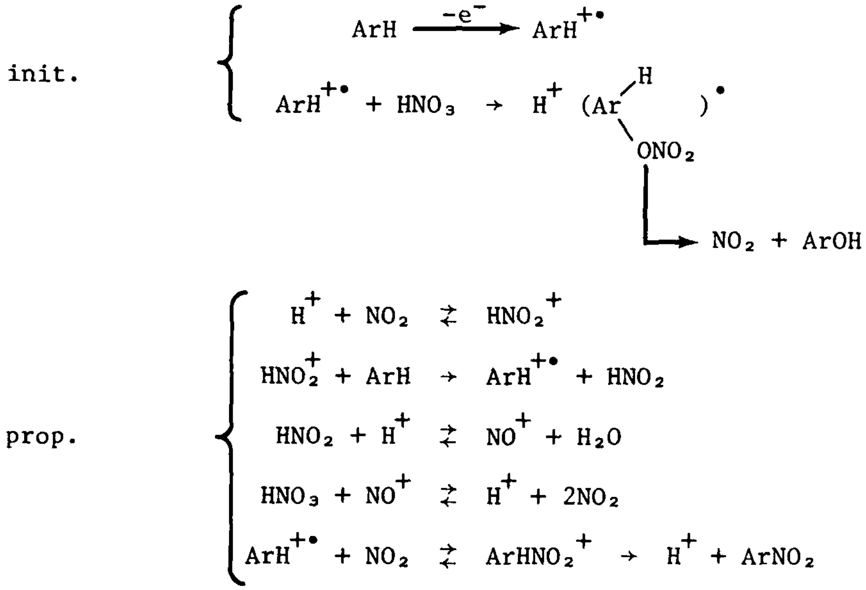
FIGURE 4 a) NITRATION RATE vs %-H_3SO_4; b) COMPARISON WITH -(H_R + LOG a_w); c) COMPARISON WITH ACTIVITY COEFFICIENT RATIO FROM EQ. (30)

This treatment comes down to the fact that the potential of a nitric acid cell is proportional to $2(-H_o)$. In fact, for the range 10% to 95% sulfuric acid, $H_R \approx 2H_o$,[20] and so $E \approx (-H_R + \text{const})$. Thus in retrospect the observed fit is not surprising. Of course, the chemistries involved in the H_R function and in the potential of a nitric acid cell are very different, and the agreement must be coincidental.

It can be suggested therefore as a working hypothesis and on the basis of Figure 4c that the correlation between nitration kinetic data and the H_R acidity function is fortuitous and that the actual factors governing nitration rates are more closely related to redox potentials than to the acid-promoted formation of nitronium ion. Using Vetter's reaction scheme, Perrin's results, and our own, we suggest the following chain process for the nitration of aromatic substrates



init.

$$ArH \xrightarrow{-e^-} ArH^{+\bullet}$$

$$ArH^{+\bullet} + HNO_3 \rightarrow H^+ \left(Ar\begin{smallmatrix} H \\ ONO_2 \end{smallmatrix}\right)^\bullet \rightarrow NO_2 + ArOH$$

prop.

$$H^+ + NO_2 \rightleftarrows HNO_2^+$$

$$HNO_2^+ + ArH \rightarrow ArH^{+\bullet} + HNO_2$$

$$HNO_2 + H^+ \rightleftarrows NO^+ + H_2O$$

$$HNO_3 + NO^+ \rightleftarrows H^+ + 2NO_2$$

$$ArH^{+\bullet} + NO_2 \rightleftarrows ArHNO_2^+ \rightarrow H^+ + ArNO_2$$

The overall stoichiometry is the common

$$HNO_3 + ArH \rightarrow ArNO_2 + H_2O \qquad (31)$$

but the detailed chemistry is decidedly different from the nitronium ion-based scheme.

As written, the above sequence accounts for the acid catalysis observed below 90% H_2SO_4, and a critical aspect for initiation is the formation of small amounts of NO_2. The initiation steps shown apply to the presence of some oxidant to generate the radical cation of the aromatic substrate. However, nitric acid in the condensed phase is relatively unstable, both thermally and with respect to trace quantities of catalytic impurities.[29] It is therefore likely that even the most carefully purified samples of HNO_3 have some trace of NO_2 present. An additional initiation possibility is a direct electron transfer reaction of the substrate with the acid

$$H^+ + HNO_3 + ArH \rightarrow ArH^{+\bullet} + NO_2 + H_2O$$

The above scheme is consistent with observations noted to this point, including the material to be discussed in the next Section. We present here some very recent data in Table 4, dealing with the addition of H_2O_2 to phenol/H_2SO_4/N(V) systems. As shown for run 4, the addition of H_2O_2 to a sulfuric acid solution of phenol and N(V) suppresses the nitration severely. Thus virtually no nitration takes place in 28 min. The second-order rate constant for phenol nitration (without nitrous acid intervention) at 25°C can be estimated to be $1.2 \times 10^{-3}\ M^{-1}\ sec^{-1}$ in 56.2% sulfuric acid from the data of Coombes et al.[22] Based on that value, almost 20% of the phenol should have been converted in the 28-min period.

Further, this finding can be compared with runs 2 and 3, without and with urea, respectively, where substantial nitration has occurred.* Clearly the peroxide is acting in a significant way to inhibit the reaction, but it is not clear how such action can be explained in terms of a nitronium ion mechanism. This finding can be anticipated, however, in terms of the proposed scheme, since the peroxide would be expected to oxidize any NO_2 present back to HNO_3, and thus nitration could not proceed.

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*The chemistry for runs 2 and 3 is like that discussed in the next section where nitrous acid promotion of the nitration takes place, but not entirely via prior nitrosation. The observation that urea does not affect the result is surprising, and should be investigated further.

Table 4

PRELIMINARY DATA FROM NITRATION EXPERIMENTS WITH ADDED HYDROGEN PEROXIDE AND UREA AT 25°C AND IN 56.2% H_2SO_4[a]

Run No.	Added (mmol)	N(V) (mmol)	Reaction Time	o/p	Yield Nitrophenol (%)	Conversion (%)
1	nothing	0.32	30 sec		98% phenol recovered	
2	nothing	0.32	7 min	0.8	87	99+
	Urea					
3	0.1	0.32	7 min	0.5	76	99+
	H_2O_2/N(III)					
4	0.40/--[b]	0.20	28 min		98% phenol recovered	
5	0.40/0.20[c]	0.20	30 sec	1.3	39[d]	19
6	0.40/0.20[c]	0.20	7 min	1.1	36[d]	16
7	0.40/0.20[e]	0.20	7 min	2.1	98	15

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[a] The volume of the solutions was 3.3 ml. In each case 0.02 mmol phenol was present. All runs were run at least in duplicate.

[b] No N(III) was added.

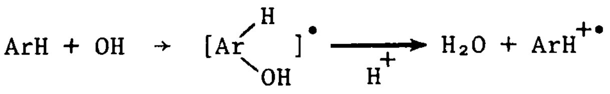
[c] H_2O_2 and N(III) mixed together just before addition to the H_2SO_4/N(III)/phenol mixture.

[d] The major product in these cases may be some oxidatively coupled product. Analyses are currently under way.

[e] The H_2SO_4/N(V)/phenol/H_2O_2 mixture was prepared first; then the N(III) was added.

When both hydrogen peroxide and N(III) are added, however, we find that nitration proceeds very rapidly as in run 5, with results similar to those discussed in Section 4 with Ce(IV). The added material significantly catalyzes the nitration, and again an explanation in terms of a nitronium ion mechanism is difficult. Comparison of runs 5 and 6 shows that the reaction is completed within 30 sec; the differences brought about by addition order changes are shown in runs 7 and 6, where 98% of the converted phenol is nitrated when N(III) is added last, in contrast to a large degree of oxidative side produce when H_2O_2 is added last. The accelerated rate and yet clean conversion in run 7 is of particular interest, and we are currently investigating this feature in detail.

The likely mode of reaction here is initial, rapid formation of HOONO, followed by its equally rapid conversion to NO_2 and HO.[30a] The OH then, in accord with recent findings by Walling et al.[30b] and O'Neil et al.,[30c] generates the aromatic radical cation OH through addition, followed by protonation and loss of water



$$ArH + OH \rightarrow \left[Ar\genfrac{}{}{0pt}{}{H}{OH}\right]^{\bullet} \xrightarrow[H^{+}]{} H_2O + ArH^{+\bullet}$$

and the nitration sequence proceeds. Alternatively, we might propose a simple acid-catalyzed electron transfer

$$H^{+} + ArH + HO \rightarrow H_2O + ArH^{+\bullet}$$

These observations support our proposition, but require the performance of a full kinetics study for confirmation.

It is tempting to extend these ideas to acidities above 90% sulfuric acid, where the nitration rate begins to decline with increasing acidity. A nitronium ion approach to nitration leads to the conclusion that the nitration rate should become independent of acidity above 90% H_2SO_4, where the conversion to nitronium ion is complete. The rate profile in this region should therefore be flat. The observed decline in rate seems to be inconsistent with the nitronium ion mechanism, but an explanation has been attributed recently by Vinnik et al.[31] to the effects of the medium on the activity coefficient ratio in the Brønsted rate equation for aromatic nitration

$$\text{rate} = k(NO_2^+)(ArH)(f_{NO_2}+f_{ArH}/f_{tr}+)$$

However, in terms of the radical cation proposition, the decline in rate at high acidities is directly evident from the chemistry. The process depends on the production of NO_2, and from the chain process, the source of NO_2 is N(III) + N(V). But above 90% H_2SO_4, where all of the N(V) is NO_2^+

$$NO_2^+ + NO^+ \not\longrightarrow 2NO_2$$

In other words, molecular nitric acid must be present for the generation of NO_2, but it is present in declining quantities at > 90% H_2SO_4. We suggest therefore that

$$H_2O + NO_2^+ \rightarrow H^+ + HNO_3 \qquad \text{(slow)}$$

$$HNO_3 + NO^+ \rightleftarrows H^+ + 2NO_2 \rightarrow \text{etc.} \qquad \text{(fast)}$$

and the nitration proceeds with the observed rate directly proportional to the activity of water.

Alternatively, NO_2 could be extensively protonated and the same conclusion drawn, with

$$HNO_2^+ + H_2O \rightarrow H_3O^+ + NO_2$$

as participating in the rate control.* For either case, the rate follows the activity of water.

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*Topol et al.[32] conducted linear sweep, cyclic voltammetric studies in 96% sulfuric acid solutions of N(III), N(IV), and N(V). They found both cathodic and anodic waves for

$$NO^+ + e^- \rightleftarrows NO$$

but only a cathodic wave for

$$NO_2^+ + e^- \rightarrow NO_2$$

Even when NO_2 was bubbled into the solution, there was no wave for its oxidation to NO_2^+. In line with the above discussion, this result suggests that at these acidities either NO_2 is rapidly and essentially completely converted to NO_2^+ and NO^+ hydrolytically, or it is fully protonated and resistant to oxidation.

Figure 5 shows that this limitation in rate by the availability of free water is in accord with experiment. The figure shows an enlarged view of nitration kinetics behavior at the high acid end, along with a curve for the activity of water to acidities as high as reliable data are available. It is clear from the figure that, if the decrease in log a_w continues smoothly above 90% H_2SO_4, the decrease in reaction rate will parallel the decline in the activity of water.

The oxidizing power of nitric acid and the kinetics of aromatic nitration therefore seem to coincide. In turn, the oxidizing power of the system may be due to increasing protonation of NO_2 yielding HNO_2^+, likely a very strong oxidant. Other potential oxidants are present in the system, including NO^+, NO_3 from the NO^+-HNO_3 reaction discussed earlier, and at higher acidities, NO_2^+. And for each of these, a scheme similar to the one above can be written, each scheme a chain process involving one-electron transfer, and each with the ultimate stoichiometry of equation (31). It would seem reasonable to suppose that each of these processes operates individually over portions of the entire acidity range, and the ensemble is the result in Figure (4a). This possibility needs to be evaluated.

In summary, we have presented an alternative to the nitronium ion-based explanation for aromatic nitration, and the proposed mechanism includes electron transfer reactions aligned with the known electrochemistry of nitric acid systems. Further, within the limits of available data discussed here, the acidity profile for proposed reaction sequence closely matches the observed profile for mixed acid nitration over acidities not only up to 90% H_2SO_4,but beyond that value through the region where the rate declines.

The preliminary and speculative nature of some of this discussion notwithstanding, it appears that the current view of the nitration chemistry in mixed acid media can be significantly challenged. Moreover, the possible route to aromatic nitration via radical cationic intermediates suggests, in addition to a new theoretical understanding, some interesting and potentially useful variations for our present nitration technology. In particular, the possibility of electrochemically controlled aromatic nitration seems attractive.

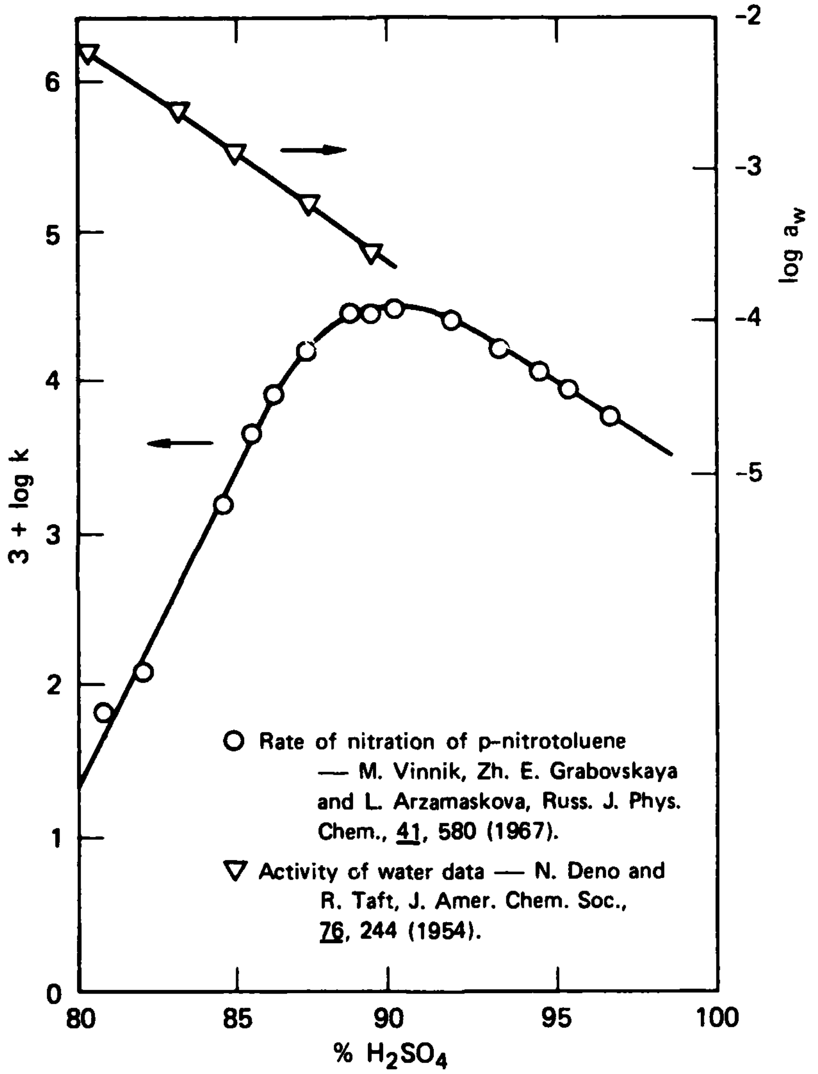

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FIGURE 5 NITRATION RATE AND WATER ACTIVITY IN CONCENTRATED SULFURIC ACID

Section 4

PROMOTION OF NITRATION BY NITROUS ACID AND Ce(IV)*

Nitrous acid can play contrasting roles in aromatic nitration. On the one hand, it acts as an anticatalyst in the nitration of most aromatic substrates in nitromethane/nitric acid mixtures, and in nitric acid alone. On the other hand, in nitric acid/sulfuric acid mixtures, it catalyzes the nitration of phenol, aniline, and related compounds activated toward electrophilic substitution.[4]

The catalytic action of nitrous acid in the latter case is explained through initial nitrosation of the aromatic substrate, followed by oxidation to the corresponding nitroaromatic product by nitric acid. Thus, the HONO-catalyzed nitration of phenol or anisol, for example, yields mainly p-substitution.[33,9b] When steps are taken to guard against the accumulation of HONO by the addition of urea, however, the reaction rate is substantially reduced, and the o:p ratio is reported to become ∿ 1.9.[9b,22]

Our initial study of the system by uv spectroscopy led to the curious observation that the uv profiles of the products of the nitrous acid catalyzed nitration of phenol were different, depending on the order in which the HONO and HNO_3 were added to the system in otherwise identical experiments. This finding led to further study using high-pressure liquid chromatography for product analysis. The data for several runs in 56.2% sulfuric acid at 25°C are shown in Table 5 and summarized in the following chart. (The oxidation states of the oxy-acids of nitrogen are used

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*This portion of the final report was prepared initially as a manuscript dealing with both nitrous acid and Ce(IV) promotion of nitration in mixed acid media. A shortened version of it has been accepted for publication in Chemical Communications, and is presented in the Appendix.

here, since there is more than one specific form present for both acids in 50-60% sulfuric acid. Thus, N(III) in these solutions is both NO^+ and HONO; N(V) is both HNO_3 and NO_3^-.[34] In all of the work, aliquots of aqueous sodium nitrate and sodium nitrite solutions were added to the acid medium, and thus served as fresh sources of the respective acids.)

It is seen that when the N(III) is present initially, nitrosation is complete within seconds. With the subsequent addition of N(V) after 30 sec, p-nitrophenol is virtually the only product after 7 min, with only traces of the o-isomer seen in the product mixture. This result is in accord with reported nitrosation studies.[35a]

However, when the N(V) is present initially, followed in 30 sec by N(III) addition, while the total quantities are small, there are significant and consistent increases in the amounts of o-nitrophenol formed, and some benzoquinone is seen as well.

NITRATION OF PHENOL IN 56.2% SULFURIC ACID
AT 25°C

$$\text{PhOH} + \text{N(III)} \longrightarrow \underset{\text{full conversion}}{p\text{-ONPhOH}} \xrightarrow{\text{N(V)}} \underset{\text{98\% yield}}{\text{O}_2\text{NPhOH (o:p} \approx 0.03)}$$

(see text)

$$\text{PhOH} + \text{N(V)} \longrightarrow \underset{\text{98\% recovered}}{\text{PhOH}} \xrightarrow{\text{N(III)}} \underset{\text{98\% yield}}{\text{O}_2\text{NPhOH (o:p} = 0.11)}$$

if quenched here

$$\left.\begin{array}{c}\text{PhOH} + \text{N(V)} \\ + \\ \text{Ce(IV)}\end{array}\right\} \longrightarrow \underset{\text{20\% yield}}{\text{O}_2\text{NPhOH (o:p} \approx 1.2)}$$

——— 30 sec ———→ ——— 6.5 min ———→

Table 5

NITRATION OF PHENOL IN 56.2% SULFURIC
ACID AT 25°C[a]

Mode	Reaction Period	NP[b] (%)	o:p	Benzoquinone[c] (%)	Material Balance[d] (%)	Remarks
N(V) alone	30 sec	< 1	--	--	98	98% phenol recovered
	7 min	87	0.78	tr	87	No phenol observed in product
N(III) first[e]	7 min	98^+	0.03±0.02	0	98^+	No phenol observed in product
N(V) first[e]	7 min	98^+	0.11±0.01	tr	98^+	No phenol observed in product
Ce(IV)/N(V)	30 sec[f]	19-23	1.2-1.3	2-3	75-90	50-70% phenol recovered

---

[a] The volume of the solutions was 3.3 ml. Initial concentrations: N(V), 6.1 - 9.7 x 10^{-2} M; phenol, 6.1 x 10^{-3} M; N(III), 6.1 x 10^{-2} M; Ce(IV), 6.1 x 10^{-3} M.

[b] Fraction of starting phenol converted to nitrophenols.

[c] Fraction of starting phenol converted to benzoquinone.

[d] Fraciton of starting phenol accounted for.

[e] Second component added 30 sec after first.

[f] Three separate experiments yielded somewhat scattered results.

The observation of significant o-substitution for the latter case might be simply explained through the suggestion that in the 30 sec period before the N(III) addition standard nitronium ion [N(V)] nitration took place. However, as shown, in control experiments quenched after 30 sec when only N(V) was present, there was no measurable nitration. Alternatively, it might be suggested that in the period after the N(III) was added, nitronium ion nitration competed with nitrosation kinetically. However, phenol nitrosation is very rapid.[35a] In some preliminary kinetic studies with these systems, we have found that at 25°C, the ratio of rates of initial nitration to nitrosation under our conditions is about 3×10^{-5}. On the other hand, the generation of a product mixture with an o:p ratio of 0.11 from an appropriate combination of straight nitration and prior nitrosation requires a rate ratio of about 10^{-2}. Thus, nitration is about three orders of magnitude too slow under our conditions, and cannot compete with nitrosation. Therefore, when nitrous acid is present in a mixed acid nitration system, at least with phenol as substrate, there appears to be a reaction route competitive with nitrosation, requiring nitrous acid, and leading to ortho-rich nitration.

Support for a significant role of N(III) in phenol nitration comes from the brief kinetic studies of these solutions. The nitration of phenol in N(V)/H_2SO_4 media is smoothly second order for an initial period. But then the process suddenly accelerates, presumably promoted by the accumulation of nitrous acid. At 25°C, the second-order rate constant of the initial reaction in 56.2% sulfuric acid was found to be $3.9 \times 10^{-3}\ M^{-1}\ s^{-1}$, and the point of incursion of the acceleration was a strong function of the starting phenol and N(V) concentrations. For the concentrations used in the runs shown in Table 5, the acceleration began after about 2 min or after only about 2% of phenol was nitrated. In the accelerated regime the nitration was complete in 5 min.

The 7-min data in Table 5 for N(V) alone are therefore for a system in which catalysis of some sort takes place. However, if the catalysis were the result of prior nitrosation, since the large majority of phenol remained to be nitrated, the nitration product would be expected to be very rich in p-nitrophenol. We should, in fact, expect a product mixture similar to that obtained for the runs in which N(III) was added. It is seen in the runs with no N(III) added that the o:p ratio of 0.78 is inconsistent with a scheme in which nitrosation occurs initially, and we conclude again that N(III) is involved in the nitration process, but not totally via nitrosation. It can be estimated that if the catalytic route not involving nitrosation characteristically yields product with an o:p ratio of 1.2 (see below), then that route must represent about 60% of the entire process. This observation is qualitatively similar to that of Ridd and coworkers,[36] who found nitration promotion by N(III) in 85% H_2SO_4, at which acidity nitrosation is very slow.

In other work we have found nitrous acid in sulfuric acid to be a strong oxidizing system, at least toward ethers, alcohols[37] and SO_2,[38] and the possibility therefore existed that an oxidative step was operative in the scheme.* Accordingly a number of experiments were run in which Ce(IV) was added to sulfuric acid solutions of phenol and nitric acid. Ce(IV) readily removes an electron from simple aromatic systems to produce a radical cation,[39] and our initial expectation was that the major products of reaction would be phenol oxidation products such as benzoquinone and diphenoquinone.[24] Instead, as shown in Table 5, significant nitration took place within just 30 sec. Indeed, 20% of the phenol was converted to an ortho-rich mixture of nitrophenols, with an o:p ratio of 1.2-1.3. No m-isomer was observed. Also the product mixture contained about 3% benzoquinone. These results thus support our suggestions about the oxidative chemistry of systems that contain N(III) and represent a new route to nitration.

The 7-min runs with cerium yielded a nitrophenol mixture similar to that obtained in the runs with N(V) alone with an o:p ratio of 0.7-0.8.

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*We suggest that the active oxidizer in the system is NO_3, formed via the protonated form of O-bridged N_2O_4

$$NO^+ + HNO_3 \rightleftarrows O{=}N{-}O{-}\overset{+}{N}(O){-}OH$$

$$(HN_2O_4)^+ \rightarrow NO + H^+ + NO_3$$

$$NO_3 + e^- \rightarrow NO_3^-$$

This result is difficult to explain. Since the first 30-sec period of the reaction yields product with an o:p of ∿ 1.3, to provide the observed 7-min product, the nitration proceeding over the following 6.5 min must generate nitrophenols with an o:p ratio of about 0.4. Detailed kinetic studies are required to understand this anomaly. Nonetheless, it appears that the initial rate of aromatic nitration of phenol can be substantially increased with one-electron oxidants and that nitrous acid can play a significant oxidizer role in nitration, in addition to that of prior nitrosation.

We believe this report to be the first of a radical-based nitration in which ortho and para products are formed in predominance over the meta isomer. Kurz et al. have reported on the free radical initiated nitration of toluene in which meta substitution predominates.[40]

REFERENCES

1. T. Urbanski, Chemistry and Technology of Explosives, Vol. 1 (The Macmillan Co., New York, 1964).

2. R. Ogg, J. Phys. Chem., 18, 770 (1950).

3. R. F. Hampsen, Jr. and D. Garvin, "Reaction Rate and Photochemical Data for Atmospheric Chemistry, 1977", National Bureau of Standards Special Publication 513, U.S. Dept. of Commerce, May 1978.

4. J. Hoggett, R. Moodie, F. Penton, and K. Schofield, Nitration and Aromatic Reactivity (Cambridge University Press, 1971).

5. J. Kochi, R. Tangand, and T. Bernath, J. Amer. Chem. Soc. 95, 7114 (1973).

6. T. Suehiro, M. Hirai and T. Kaneko, Bull. Chem. Soc. Jap., 44, 1402 (1971).

7. E. Gould, Mechanism and Structure in Organic Chemistry (H. Holt and Co., 1959), p. 428.

8. (a) H. Wieland, Ber., 54B, 1776 (1921); (b) T. Bonner and R. Hancock, Chem. Comm., 780 (1967); (c) G. Underwood, R. Silverman, and A. Vanderwalde, J. Chem. Soc. Perkin II, 1177 (1973).

9. (a) R. Moodie, K. Schofield, and G. Tobin, J. Chem. Soc., Chem. Comm., 180 (1978); (b) J. Hoggett, R. Moodie, and K. Schofield, Chem. Comm. 605 (1969).

10. (a) S. W. Benson, Thermochemical Kinetics (John Wiley and Sons, Inc., New York, 1968); (b) G. Charlot, D. Bezier, and J. Courtot, Selected Constants--Oxydo-Reduction Potentials (Pergamon Press, New York, 1958).

11. E. White and W. Feldman, J. Amer. Chem. Soc., 79, 5832 (1959).

12. B. Challis and S. Kyrtopoulos, Chem. Comm., 877 (1976).

13. J. Franklin, J. Dillard, H. Rosenstock, J. Herron and K. Draxl, "Ionization Potentials, Appearance Potentials and Heats of Formation of Gaseous Positive Ions," Nat. Stand. Ref. Data Ser., Nat. Bur. Stand., NSRDS-NGS 26, U.S. Govt. Printing Office, Wash. DC (1969).

14. B. Challis and J. Ostram, Chem. Comm., 707 (1978).

15. D. Golden, R. Solly, N. Gac, and S. Benson, J. Amer. Chem. Soc., 94, 363 (1972).

16. G. Wright in The Chemistry of the Nitro and Nitroso Groups, H. Feuer, Ed. (Interscience Publishers, New York, 1969), p. 651.

17. D. Millen, J. Chem. Soc., 2600 (1950).

18. J. Brockman, D. Downing, and G. Wright, Can. J. Res., B, 27, 469 (1949).

19. F. Westheimer and M. Kharasch, J. Amer. Chem. Soc., 68, 1871 (1946).

20. C. Rochester, Acidity Functions (Academic Press, New York, 1970).

21. R. Moodie, K. Schofield, and M. Williamson, in Nitro Compounds, Proceedings of International Symposium, Warsaw (Pergamon Press, London, 1963), p. 89.

22. R. Coombes, R. Moodie, and K. Schofield, J. Chem. Soc., 800 (1968).

23. C. Perrin, J. Amer. Chem. Soc., 99 (16), 5516 (1966).

24. R. Cecil and J. S. Littler, J. Chem. Soc. B, 1420 (1968).

25. W. Latimer, The Oxidation States of the Elements and Their Potentials in Aqueous Solutions, Second Edition (Prenctice-Hall, Inc., New York, 1952), p. 93.

26. (a) K. Vetter, Z. Physik, Chemie, 194, 199 (1950); (b) K. Vetter, Electrochemical Kinetics (Academic Press, New York, 1967), p. 490.

27. R. Monk and H. Ellingham, J. Chem. Soc., 127 (1935).

28. E. Julien and M. Comtat, Revue de Chemie Minerule, 6, 885 (1969).

29. P. Gray and A. Yoffee, Chem. Rev., 55, 1069 (1955).

30. (a) E. Halfpenny and P. Robinson, J. Chem. Soc., 928, 939 (1952); (b) C. Walling, D. Cambioni, and S.S. Kim, J. Amer. Chem. Soc., 100 (15) (1978); (c) P. O'Neil, D. Schulte-Frohlinde, and S. Steenken, Faraday Disc. Chem. Soc., 63, 1418 (1977).

31. M. Vinnik, Zh. E. Grabovskaya, and L. Arazamaskova, Russ. J. Phys. Chem., 41, 580 (1967).

32. L. Topol, R. Osteryoung, and J. Christie, J. Electrochem. Soc., 112 (8), 861 (1965).

33. C. Bunton, E. Hughes, C. Ingold, D. Jacobs, M. Jones, J. Minkoff, and R. Reed, J. Chem. Soc., 2628 (1950).

34. N. Deno, H. Peterson and E. Sacher, J. Phys. Chem., 65, 199 (1961); N. Bayliss, R. Dingle, D. Watts and R. Wilkie, Aust. J. Chem., 16, 933 (1963); D. S. Ross and N. A. Kirshen, "Industrial and Laboratory Nitrations," Am. Chem. Soc., Symposium Series 22, Amer. Chem. Soc., Washington, D.C., 1975, pp. 114-131.

35. (a) B. Challis and A. Lawson, J. Chem. Soc. (B), 770, (1971).
(b) A. Granzow and A. Wilson, J. Org. Chem., 37 (8), 1189 (1972).

36. J. Giffney, D. Mills, and J. Ridd, J. Chem. Soc., Comm., 19 (1976).

37. D. S. Ross and G. Hum,Abstracts, 175th National Meeting of the American Chemical Society, Anaheim, CA, March, 1978, ORGN 168.

38. D. S. Ross, Environ. Sci. Tech., 12 (6), 726 (1978).

39. W. Trahanovsky and D. Brixius, J. Am. Chem. Soc., 95 (20), 6778 (1973).

40. M. Kurz, R. Foxdar, and S. Schultz, J. Org. Chem., 39, 3336 (1974).

APPENDIX

CATALYSIS OF AROMATIC NITRATION BY THE LOWER OXIDES OF NITROGEN

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Catalysis of Aromatic Nitration by the Lower Oxides of Nitrogen

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Summary In the absence of nitrous acid traps, the nitration of phenol in 56·2% sulphuric acid displays autocatalytic behaviour; on the other hand, the isomer ratio of the products is inconsistent with the commonly accepted prior nitrosation scheme, and some other route for the promotion of nitration must be operative.

We report here on anomalous nitration patterns of phenol in mineral acid nitration media which suggest an unusual catalysis of the nitration. Such nitrations of phenol and other substrates activated toward electrophilic substitution are recognized to be subject to the catalytic effects of nitrous acid and perhaps other lower oxides of nitrogen, and the mode of action is assumed to be prior nitrosation followed by oxidation of the nitrosoaromatic.[1] We have found that this reaction scheme is not necessarily operative, or, at the very least, is not alone in promoting nitration.

We used phenol at 25 °C in 56·2% sulphuric acid with no added nitrous acid trap. Aliquots of sodium nitrate and nitrite solutions served as fresh sources of the N^{V} and N^{III} species, respectively, in each run.† We added the two oxyacids to phenol–sulphuric acid solutions either alone or sequentially, with 30 s between additions. The reactions were quenched and worked up after either 30 s or 7 min total reaction time. The analyses were carried out by h.p.l.c., and the data are presented in the Table.

Table. Nitration of phenol in 56·2% sulphuric acid at 25 °C[a]

Run	Mode	Reaction time	Nitrophenol[b] (%)	Ratio $o:p$	Material balance[c] (%)	Remarks
1	N^{V} alone	30 s	< 1	—	98	98% phenol recovered
2	N^{III} alone[d]	30 s	quantitative (nitrosophenol)	—	—	—
3	N^{V} alone	7 min	87	0·78	87	No phenol observed in product
4	N^{V} first[e]	7 min	98^{+}	0·11 ± 0·01	98^{+}	No phenol observed in product
5	N^{III} first[e]	7 min	98^{+}	0·03 ± 0·02	98^{+}	No phenol observed in product

[a] The volume of the solutions was 3·3 ml. Initial concentrations: N^{V}, 6·1—9·7 × 10^{-2} M; phenol, 6·1 × 10^{-3} M; N^{III}, 6·1 × 10^{-2} M. Analyses on quenched samples by h.p.l.c. [b] Fraction of starting phenol converted into nitrophenols. [c] Fraction of starting phenol accounted for. [d] Performed in a spectrophotometer; no product isolated. [e] Both N^{III} and N^{V} added. Second component added 30 s after first.

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† The oxidation states of the oxyacids of nitrogen are used here because more than one specific form is present for both nitrous and nitric acids in 50–60% sulphuric acid. Thus, N^{III} in these solutions is both NO^{+} and HONO; N^{V} is both HNO_3 and NO_2

First, we see that whereas the intervention of NO_x catalysis of phenol nitration can bring about autocatalysis,[1] in the brief 30 s period employed in run 1 almost no nitration took place.

We then established that, in accord with work of Challis and Lawson, phenol nitrosation is rapid and yields almost entirely the *para*-isomer.[2] In our work the conversion was complete and quantitative within 30 s (run 2), with a measured second-order rate constant of $k_{NO} = 92$ dm^3 mol^{-1} s^{-1}. A measure of the very small degree of substitution was obtained from runs, shown collectively as run 5, in which the nitrophenol formed from the subsequent oxidation of the nitroso compound by addition of N^V was isolated and found to have an *o*:*p* ratio of (0·03 ± 0·02) : 1.

Next, in run 3 we see that with only N^V, after 7 min the nitration is effectively complete, in contrast with the 30-s result of run 1. Clearly some sort of catalysis is operative in the 7 min run, since from our measured initial rate constant for phenol nitration in these media ($k_{NO_2} = 3·9 \times 10^{-3}$ dm^3 mol^{-1} s^{-1}) only about 7% of the phenol should have been nitrated over that period. This value is in fair agreement with the value for nitration at that acidity in studies with added HONO traps.[1b] However, the product *o*:*p* ratio is 0·78 : 1, a value clearly inconsistent with a prior nitrosation route, which should yield an *o*:*p* ratio of about 0·03 : 1.

Finally, in accord with this assessment is the result of run 4 in which N^V was added 30 s before N^{III}. While small in an absolute sense, the product *o*:*p* ratio of (0·11 ± 0·01) : 1 shows that 3—4 times more *ortho* nitration is taking place in this mixture than would be expected from the essentially pure prior nitrosation route operative in run 5.

If we assume that in run 3 both nitrosation and nitration are proceeding in parallel, then given the nitrosation *o*:*p* ratio of 0·03 : 1, and recognizing that the *o*:*p* ratio for straight nitration is about 2·0 : 1,[1b] it can be shown that the observed product mixture can be obtained only if the nitration rate is about 1·9 times greater than the rate of nitration *via* nitrosation and oxidation. In addition, since the reaction is complete in 7 min, a lower limit can be calculated for the apparent second-order rate constant for the catalysed component of the aromatic nitration, $k^{cat}_{NO_2} > 5·0 \times 10^{-2}$ dm^3 mol^{-1} s^{-1}.

This value is an order of magnitude greater than that for the conventional nitration, and clearly a potent catalyst not involving nitrosation is operative. It is not clear if in fact the reaction is truly catalysed or merely promoted, and we use the former term here with reservation.

The chemistry in effect here must be elucidated. It can be suggested that phenol may be a special case, since at least in acetic anhydride its nitration is more sensitive to promotion by the lower nitrogen oxides than that of anisole.[4] On the other hand, our results can be compared with those recently presented by Giffney and Ridd,[3] who suggest that the lower oxides of nitrogen play a non-nitrosative role in the reaction of *N,N*-dimethylaniline in nitration media. Their proposed scheme includes electron transfer from the protonated substrate to NO^+, followed by radical cation reactions. The possibility of radical cation intermediates in aromatic nitration more generally has been recently raised,[5] and it seems the entire question of nitration mechanism is yet to be resolved.[6]

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[1] (a) J. Hoggett, R. B. Moodie, J. R. Penton, and K. Schofield, in 'Nitration and Aromatic Reactivity,' Cambridge University Press, Cambridge, 1971, p. 57; (b) R. G. Coombes, J. G. Golding, and P. Hadjigeorgiou, *J. Chem. Soc., Perkin Trans.* 2, 1979, 1451.
[2] B. C. Challis and A. J. Lawson, *J. Chem. Soc. B*, 1971, 770.
[3] J. C. Giffney and J. A. Ridd, *J. Chem. Soc., Perkin Trans.* 2, 1979, 618.
[4] J. G. Hoggett, R. B. Moodie, and K. Schofield, *Chem. Commun.*, 1969, 605.
[5] C. L. Perrin, *J. Am. Chem. Soc.*, 1977, **99**, 5516; L. Eberson, L. Johnson, and F. Radner, *Acta Chem. Scand., Ser.* B, 1978, **32**, 749.
[6] R. G. Coombes in 'Comprehensive Organic Chemistry,' Vol. 2, ed. I. O. Sutherland, Pergamon, Oxford, 1979, p. 309.