884. *Nitration at Nitrogen and Oxygen Centres. Part I. Kinetics and Mechanism of the Conversion of Secondary Amines into Nitroamines.*

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The *N*-nitration of *N*-methyl-2 : 4 : 6-trinitroaniline by nitric acid in constant excess in nitromethane of not-too-high water-content is of zeroth order and, at not-too-high concentrations of nitric acid, the absolute rate of the nitration is identical with the common rate of nitration of aromatic hydrocarbons under the same conditions. Small additions of sulphuric acid increase, of nitrate ions reduce, and of water do not change, the zeroth-order rate, but larger additions of water increase rate, with maintenance of the zeroth-order form, and still larger ones change the kinetics to first-order form, with reduction in the absolute rate during the kinetic transition, and in the first-order rate-constant after it is completed. These effects fully establish the nitronium-ion mechanism for this *N*-nitration.

Two successive steps of nitration of *N*-methyl-2 : 4-dinitroaniline occur in similar conditions. The first step is an *N*-nitration, the kinetic characteristics of which, as far as studied, are practically identical with those of the nitration just described. The second kinetic step is stoicheiometrically a *C*-nitration, and is observed only in first-order form. Almost certainly the rate-controlling process here is an intramolecular rearrangement, in which the *N*-nitro-group moves into the aromatic ring, this slow reaction then being very quickly succeeded by another *N*-nitration.

Electrophilic substitution has hitherto been studied largely in the form of aromatic substitution, where the attack by the substituting agent is on the unsaturation electrons, *i.e.*, conjugated carbon $2p$ electrons. Since non-bonding $2p$ electrons of nitrogen and oxygen can participate deeply in such conjugation, we should expect them to share many properties with the unsaturation $2p$ electrons of carbon, including general vulnerability to electrophilic substituting agents. Indeed, for a given substitution of this class, *e.g.*, nitration, nitrosation, or chlorination, we should expect the pattern of mechanisms to be much the same for *C*-, *N*-, and *O*-substitutions.

Nitration has proved to be one of the simpler forms of electrophilic aromatic substitution, because it is dominated by one outstanding mechanism, that of the nitronium ion. Aromatic nitration through attack of the nitronium ion was kinetically demonstrated, at first in non-hydroxylic, and in hydroxylic but non-aqueous solvents,[1, 2] and, more recently, even in solvent water.[3]

$$\left.\begin{aligned}
HNO_3 + HNO_3 &\underset{-1}{\overset{+1}{\rightleftharpoons}} H_2NO_3^+ + NO_3^- \text{ (fast)} \quad \cdots\cdots \\
H_2NO_3^+ &\underset{-2}{\overset{+2}{\rightleftharpoons}} NO_2^+ + H_2O \text{ (slow)} \quad \cdots\cdots \\
NO_2^+ + XH &\xrightarrow[3]{} XH{\cdot}NO_2^+ \text{ (slow)} \quad \cdots\cdots \\
XH{\cdot}NO_2^+ + NO_3^- &\xrightarrow[4]{} X{\cdot}NO_2 + HNO_3 \text{ (fast)} \quad \cdots\cdots
\end{aligned}\right\} \text{(I)}$$

We here apply the general method of these demonstrations to *N*-nitration, which again proves to be an affair of the nitronium ion. The principle of the method can be most easily appreciated by recalling the conclusion to which it led concerning the mechanism of aromatic nitration. This conclusion is expressed in scheme (I), XH replacing ArH, for convenience. There are four stages. The nitronium ion is formed in two steps, of which the former, a protonation, is always fast. Then the substitution in the substrate goes in two steps, of which the latter, a deprotonation, is always fast. The intermediate steps, 2 and 3, are in principle slow, and, in any particular case, the kinetic form depends on which of them is rate-controlling. With a sufficient concentration of a sufficiently reactive substrate, and in a not-too-aqueous medium, step 3 prevails over step -2 as a means of consumption of formed nitronium ion, so that the latter reacts, as fast as it is formed, with the substrate, and the measured nitration rate is the rate of its formation in step $+2$. This rate in turn is proportional to the stationary concentration of nitric acidium ion, and is therefore increased by added small amounts of a very strong acid, such as sulphuric acid, and is decreased by added nitrate ions. But as long as the nitronium ions are trapped almost wholly by the substrate, the measured rate must be independent of the concentration and nature of the latter: it will be of zeroth-order in the substrate.

The other limiting situation is that in which the substrate is so unreactive, or so dilute, or the medium is so aqueous, that step -2 prevails over step 3 in the consumption of nitronium ion. The latter is then established in a stationary concentration, not sensibly depleted by the slow withdrawal of material by the substrate; the rate of such withdrawal becomes proportional to the concentration, and dependent on the nature, of the substrate. The rate still depends on a pre-equilibrium protonation, and hence will be affected as before by an added strong acid, and by added nitrate. But it will now be of first-order in the substrate, and the rate-constants applying to different substrates will reflect their intrinsic reactivities.

A firm proof of mechanism requires the observation of both limiting kinetic forms, and of the transition between them, effected, as this may be, by changing the trapping power for nitronium ion either of the substrate, or of the water, *i.e.*, by changing the balance of competition between these two consumers of formed nitronium ion.

These principles are here applied to the conversion of secondary amines into nitroamines by means of nitric acid in constant excess in solvent nitromethane. In this polar,

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[1] Benford and Ingold, *J.*, 1938, 929.

[2] Hughes, Ingold, and Reed, *J.*, 1950, 2400.

[3] Bunton, Halevi, and Llewellyn, *J.*, 1952, 4913; Bunton and Halevi, *ibid.*, p. 4917; Bunton and Stedman, *J.*, 1958, 2420.

but initially non-aqueous, solvent, the concentration of water during nitration is of the same order of magnitude as that of the material being nitrated; and so the water sets a relatively easy standard of trapping efficiency for the substrate to overcome when zeroth-order kinetics are required to be established. Our secondary amines were, in fact, reactive enough to prevail as trapping agents at this competitive level. We could then raise the competitive standard by adding water, until the amines lost their predominating position as trapping agents. Thus we could observe the transition to first-order kinetic form.

Kinetics were followed by dilatometry, subject to check by chemical examination of the products. In both principle and technique, methods were modelled on those formerly used to establish mechanism in aromatic nitration, although some refinement of apparatus was introduced. A preliminary note on the present work has appeared.[4]

(*A*) *Conversion of* N-*Methyl-*2 : 4 : 6-*trinitroaniline into* N-*Methyl-*N : 2 : 4 : 6-*tetranitroaniline.*—This *N*-nitration was chosen first for its expected simplicity, inasmuch as the secondary amine is so weakly basic that heavy protonation in nitration conditions seemed likely to be avoidable. Moreover, the nitro-amine could not undergo rearrangement, and the aromatic ring was incapable of being itself nitrated.

(1) *Kinetic form. Effect of substrate concentration.* The nitration in nitromethane is associated with a contraction. It is only about one-quarter of that accompanying the nitration of toluene, but is large enough for easy measurement.

A typical reaction–time curve, obtained dilatometrically, for nitration in initially dry nitromethane, is shown in Fig. 1. Similar curves were obtained in many runs covering initial substrate concentrations from about 0·04 to 0·16M, sometimes also with other added substances. The reaction is clearly of zeroth order in substrate.

Since each run has zeroth-order form, the zeroth-order rate constant, as deduced from the constant slope of the curve, should be independent of the initial concentration of the substrate at concentrations over which the zeroth-order law prevails. This is so, as is illustrated in Table 1, the mean value of 10^5k_0 being 2·89 moles l.$^{-1}$ sec.$^{-1}$.

After a number of runs, the contents of the dilatometer were examined chemically. A high yield of *N*-methyl-*N* : 2 : 4 : 6-tetranitroaniline was invariably isolated.

(2) *Comparison with rate of nitration of toluene.* The rate of nitration of *N*-methyl-2 : 4 : 6-trinitroaniline increases steeply with increasing concentration of nitric acid. A quite similar behaviour has been repeatedly observed in aromatic nitration, for example, of toluene. In 2M-nitric acid the absolute nitration rates of the secondary amine and of toluene are identical (Table 2).

TABLE 1. *Zeroth-order rate-constants* (k_0 *in moles l.*$^{-1}$ *sec.*$^{-1}$) *for nitration of* N-*methyl-*2 : 4 : 6-*trinitroaniline by* 3·0M-*nitric acid in nitromethane at* 25°.

[Amine]$_0$ ...	0·0314	0·0458	0·0688	0·0752	0·0786	0·0832	0·0887	0·127	0·150	0·163
10^5k_0	2·75	2·75	2·80	2·85	2·92	2·95	3·02	2·95	3·01	2·89

TABLE 2. *Zeroth-order rate-constants* (k_0 *in moles l.*$^{-1}$ *sec.*$^{-1}$) *for nitration of* N-*methyl-*2 : 4 : 6-*trinitroaniline and of toluene (initially about* 0·1M) *by nitric acid in nitromethane at* 25°.

[HNO_3]	2·0	2·5	3·0	3·5	4·0	4·5
10^5k_0 (amine)	0·675	1·41	2·89	6·04	9·64	14·9
,, (toluene) ...	0·688	1·69	3·34	6·80 *	12·7	—

* Value for ethylbenzene.

At the higher nitric acid concentrations the rate for nitration of the amine drops behind the rate for benzenoid hydrocarbons. The cause of this is not established, but it may be that the amine becomes partly protected by protonation at the higher acid concentrations. There is some evidence[4] that the deviation becomes diminished in the presence of so-called nitrous acid. This is really dinitrogen tetroxide, partly in ionised

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[4] Blackall and Hughes, *Nature*, 1952, **170**, 972.

form,[5] and hence leads to the presence of nitrate ion, which, as the lyate ion of the system, must reduce its acidity.

(3) *Kinetic effects of added solutes.* When sulphuric acid is added to the system, nitration is accelerated without change to its zeroth-order form (Table 3). Above a certain small concentration of sulphuric acid, the rate increase is linear with the concentration (Fig. 2). We interpret the change in rate as reflecting the change in the stationary concentration of nitric acidium ion, from which is derived the nitronium ion, the rate of production of which is measured in the observed nitration. Hence, we should expect the linear law to hold only after enough of the strong acid has been added fully to suppress formation of nitric acidium ion by autoprotolysis of the nitric acid, *i.e.*, to secure practically complete protonation of the nitrate ions thus formed. Accordingly, the accelerating effect of progressive addition of the strong acid should be expressed by a curve, such as is in Fig. 2, the slope of which is only somewhat slowly built up to that of its linear portion.

FIG. 1. *Zeroth-order nitration of* N-*methyl*-2 : 4 : 6-*trinitroaniline* (0·16M) *by nitric acid* (4·0M) *in nitromethane at* 25°.

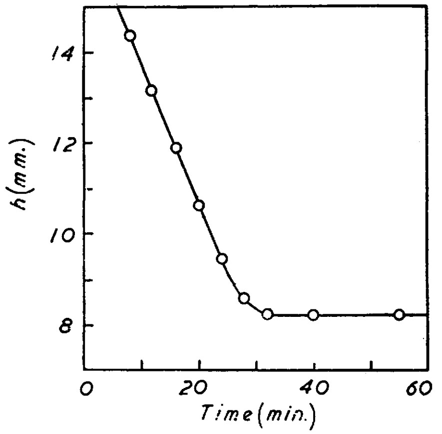
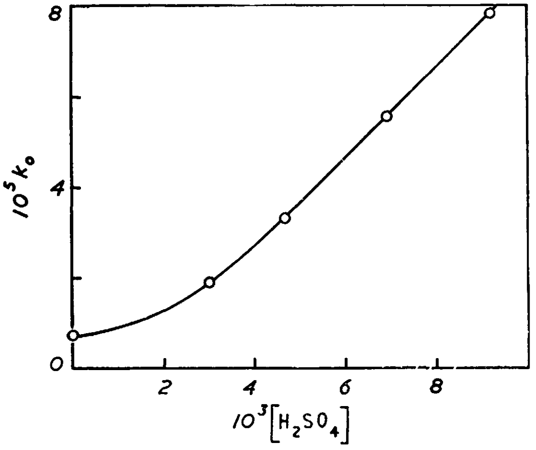


FIG. 2. *Effect of sulphuric acid on zeroth-order rate of nitration of* N-*methyl*-2 : 4 : 6-*trinitroaniline* (0·05M) *by nitric acid* (2·0M) *in nitromethane at* 25°.



As is illustrated in Table 3, a substantial accelerating effect is experienced when sodium perchlorate is added to the original system. We assume that, at the concentrations used, the salt would be largely present as dissociated ions, and we interpret the acceleration as a normal salt effect on a reaction, the kinetically significant part of which, *i.e.*, the part up to and including the slow step, is producing ions (nitronium and nitrate) from initially neutral species (nitric acid molecules).

The addition of sodium nitrate to the original system produces a retardation, as is also shown in Table 3. This must be the residue of a larger specific retardation which outweighs the accelerating normal effect of the salt. Although we have not enough data to allow a quantitative separation of these two effects, the specific retardation is qualitatively consistent with its interpretation as a suppression by added nitrate ions of the production of nitric acidium ion by the autoprotolysis of nitric acid.

(4) *Kinetic transition from zeroth to first order.* When, to the original system, water was added in successively increasing quantities, the following kinetic changes were observed. In the millimolar and centimolar ranges of concentration, the ranges in which substantial kinetic effects were produced by the added solutes considered in the preceding sub-section, added water produced no observable kinetic effect. As water does not appear

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[5] Goulden and Millen, *J.*, 1950, 2620.

in the equation for the autoprotolysis equilibrium, the only pre-equilibrium in zeroth-order nitration, this is very reasonable. Near the concentration 0·1M-water, a definite acceleration appeared, which we view as an ordinary medium effect; but there was still no detectable disturbance to the zeroth-order form of the reaction. Since the kinetically significant part of a zeroth-order nitration produces ions from originally neutral species, a highly polar co-solvent, such as water, should be expected to accelerate the reaction.

TABLE 3. *Effects of added sulphuric acid, sodium perchlorate, and sodium nitrate on the zeroth-order rate-constants* (k_0 *in moles l.*$^{-1}$ *sec.*$^{-1}$) *of nitration of* N-*methyl*-2 : 4 : 6-*trinitroaniline* (*initially about* 0·05M) *by nitric acid in nitromethane at* 25°.

$[HNO_3]$ = 2·0M		$[HNO_3]$ = 3·0M		$[HNO_3]$ = 3·0M	
$[H_2SO_4]$	10^5k_0	$[NaClO_4]$	10^5k_0	$[NaNO_3]$	10^5k_0
0	0·680	0	2·75	0	2·75
0·00301	1·89	0·00315	3·16	0·00320	2·03
0·00469	3·29	0·00643	3·47	0·00784	1·53
0·00690	5·54	0·00719	3·82	0·0120	1·42
0·00920	7·81			0·0177	1·31

And since water is still not a pre-equilibrium product, it has no specific kinetic effect which could counter this general co-solvent effect. The idea that added water can, in certain circumstances, accelerate nitration, though unconventional, is thus rational. However, when water was added above a concentration of about 0·2M, the absolute rate began to fall, and the reaction simultaneously began to lose its zeroth-order form. As the water concentration was further increased the rate fell sharply, and the kinetic order rose from zero towards unity, at first near the " tail " of reaction, and later throughout its course. At 0·8M-water, a first-order kinetic form had been established throughout reaction with good approximation. On still further increase of the water concentration, the rate, now specifiable by means of a first-order rate-constant, continued to fall, though the first-order law remained in force to the highest water-concentration investigated, 1·0M. This first-order kinetic form of nitration is illustrated in Fig. 3. We can understand why water retards the first-order reaction, although it mildy accelerates the zeroth-order reaction, by noting that, in the first-order reaction, water will have not only its general co-solvent effect, but also a specific retarding effect, because not only the autoprotolysis, but also the cationic dehydration, are now included in the pre-equilibrium, and, in this enlarged pre-equilibrium system, water is a product. Some rate-constants illustrating the main steps in these kinetic effects are assembled in Table 4.

TABLE 4. *Zeroth- and first-order rate-constants* (k_0 *in moles l.*$^{-1}$ *sec.*$^{-1}$ *and* k_1 *in sec.*$^{-1}$) *of nitration of* N-*methyl*-2 : 4 : 6-*trinitroaniline* (*about* 0·05M) *by nitric acid* (3·0M) *at* 25° *in nitromethane containing various concentrations of initially added water.*

$[H_2O]$ *	10^5k_0	Order	$[H_2O]$ *	10^4k_1	Order
—	2·75	0	0·794	~2·61	impure
0·149	3·29	0	0·825	1·96	1
0·228	3·23	0	0·972	1·40	1
0·469	~1·78	impure	1·02	1·34	1
0·610	~0·99	impure			

* *Added* water.

(5) *Mechanism of* N-*nitration.* The preceding four sections cover all the important points of evidence on which the mechanism of scheme I [with X = 2 : 4 : 6-$(NO_2)_3C_6H_2{\cdot}NMe$] may be based: thus *N*-nitration is certainly proceeding by attack of pre-formed nitronium ion on the unshared electrons of the amino-nitrogen atom.

(6) *Comparison of reactivity of amine and hydrocarbons.* It is interesting to compare the kinetic reactivity towards nitronium ion of these unshared nitrogen electrons with that of carbon unsaturation electrons in an aromatic hydrocarbon. This we can do by comparing the first-order rate constants for nitration of the amine, and of one or more aromatic

hydrocarbons, in nitromethane containing approximately 1M-added water. For $[H_2O] = 0{\cdot}972$, some first-order rate constants, expressed as 10^4k_1, with k_1 in sec.$^{-1}$, are: Benzene 1·0; *N*-methyl-2 : 4 : 6-trinitroaniline, 1·4; toluene, 23·4. The amine thus lies between benzene and toluene, though it is closer to benzene in kinetic affinity for the nitronium ion. But it must be much more basic than either hydrocarbon.

(*B*) *Conversion of* N-*Methyl-2* : *4-dinitroaniline successively into* N-*Methyl*-N : 2 : *4-trinitroaniline and* N-*Methyl*-N : 2 : 4 : *6-tetranitroaniline*.—(1) *The first step of nitration.* When *N*-methyl-2 : 4-dinitroaniline (0·05—0·15M) is nitrated by nitric acid (3·0—4·0M) in nitromethane at 25°, two kinetically distinct steps are seen, each marked by a contraction. In a typical experiment, the volume–time curve fell linearly for the first half-hour, with a slope similar to that which would have applied to the *N*-nitration described in Section (*A*). Then the slope changed abruptly, but not, as in the reaction of Section (*A*), to horizontality: the new direction had a small finite slope, and actually marked the start of an extended further curve, which continued to fall slowly, with exponentially diminishing slope, for many hours. We thus had a fast initial reaction of zeroth-order, succeeded by a slow continuing reaction of first order.

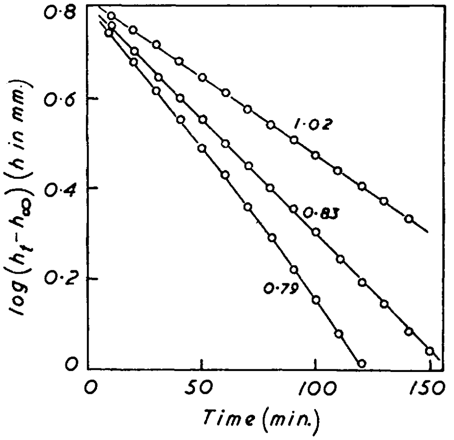

FIG. 3. *First-order nitration of* N-*methyl-2* : 4 : 6-*trinitroaniline* (~0·05M) *by nitric acid* (3·0M) *in nitromethane at* 25°. *The figures by the curves are concentrations of added water. One curve is not straight* (*cf. Table* 4).

On stopping experiments at the break in the curve, we found that the product at that point was *N*-methyl-*N* : 2 : 4-trinitroaniline. Thus the initial reaction is an *N*-nitration. And, like the *N*-nitration described in Section (*A*), it is of zeroth-order in the substrate, when conducted in initially dry nitromethane.

This being so, we might have expected that the absolute rate of *N*-nitration of the dinitro-amine and of the trinitro-amine would be the same in the same conditions. However, as shown in Table 5, this is not the case for the conditions specified; the dinitro-amine is *N*-nitrated somewhat more slowly than is the trinitro-amine. This may be a matter of protection by protonation of some of the amine from attack by the nitronium ion. At a given concentration of nitric acid, a fixed fraction of the stoicheiometric total of either amine will be present in free-basic form, and this fraction will remain constant as the stoicheiometric total of amine becomes consumed in reaction. With, for example, 3M-nitric acid, the fraction of trinitro-amine thus left free is probably high, though we saw signs of partial protection against nitration at this acidity in Section (*A*2). The dinitro-amine is, however, a considerably stronger base, and so it seems likely that more of it will be protonated, and less left in free-basic form in the conditions specified in Table 5. This might explain the rate differences.

(2) *The second step of nitration.* By isolating the material present when the kinetic form changes, we found that the second step starts from *N*-methyl-*N* : 2 : 4-trinitroaniline. By isolation of the material present at the end of the second step, it was found that the

final product is *N*-methyl-*N* : 2 : 4 : 6-tetranitroaniline. Stoicheiometrically the second kinetic step is a *C*-nitration. It will suffice thus to consider it at first, though we shall see later that appearances may be deceptive.

For greater convenience, two changes were now made in the conditions. First, to provide a sharper start, we introduced *N*-methyl-*N* : 2 : 4-trinitroaniline as such, instead of waiting for its concentration to be built up by the original *N*-nitration. Then, since, in the conditions for the observation of this *N*-nitration, namely, by the use of **3—4**M-nitric acid in nitromethane at 25°, the ensuing *C*-nitration is inconveniently slow, the *C*-nitration of introduced *N*-methyl-*N* : 2 : 4-trinitroaniline was studied with concentrations of nitric acid raised to 5—8M.

TABLE 5. *Zeroth-order rate-constants* (k_0 *in moles* $l.^{-1}$ *sec.*$^{-1}$) *for the* N-*nitration of* N-*methyl-2 : 4-dinitro- and* N-*methyl-2 : 4 : 6-trinitro-aniline in nitromethane at 25°.*

$[\text{Amine}]_0$	$[HNO_3]$	10^5k_0 (dinitro-amine)	10^5k_0 (trinitro-amine)
0·05	3·0	2·13	2·75
0·10, 0·14	4·0	7·2	9·6

TABLE 6. *First-order rate-constants* (k_1 *in sec.*$^{-1}$) *for the* C-*nitration of* N-*methyl*-N : 2 : 4-*trinitroaniline* (*initially* 0·1M *throughout*) *by nitric acid in nitromethane at 25°, and the kinetic effects of added sulphuric acid, sodium perchlorate, potassium nitrate, and water.*

$[HNO_3]$	5·0	5·5	6·0	6·5	7·0	7·5	8·0
10^4k_1	2·17	2·77	3·85	5·45	7·52	10·6	13·7

$[HNO_3]$ = 5·5		$[HNO_3]$ = 7·0		$[HNO_3]$ = 7·5		$[HNO_3]$ = 7·5	
$[H_2SO_4]$	10^4k_1	$[NaClO_4]$	10^4k_1	$[KNO_3]$	10^4k_1	$[H_2O]$	10^4k_1
—	2·77	—	7·52	—	10·6	—	10·6
0·00146	3·60	0·00263	7·87	0·00168	9·79	0·100	9·7
0·00245	4·05	0·00485	8·45	0·00323	7·20	0·150	10·0
0·00410	6·43	0·00918	9·41	0·00623	6·40	0·292	9·0
0·00572	8·02	0·0115	9·88	0·0104	5·34	0·453	5·95
0·00580	8·25	0·0171	10·5	0·0154	4·42	0·594	5·18
0·00874	16·3			0·0205	3·52	0·748	4·08

The *C*-nitration is approximately of first order in the substrate, though this kinetic law loses accuracy at late times. As shown in Table 6, the rates increase steeply with increasing concentration of nitric acid.

Some observations on the kinetic effect of added substances are recorded in Table 6. The reaction is strongly accelerated by added small amounts of sulphuric acid. It is moderately accelerated by added sodium perchlorate. Nevertheless, it is strongly retarded by added potassium nitrate. Relative to these additives, water has only a trivial kinetic effect, though, when added in excess of 0·2M, it produces a retardation. All these rate effects occur without change to the essentially first-order form of the reaction.

These kinetic results are closely in line with what is to be expected for *C*-nitration by the nitronium-ion mechanism. In zeroth-order nitration the rate of formation of nitronium ion, and in first-order nitration the stationary concentration of nitronium ion, are both proportional to the stationary concentration of its precursor, the nitric acidium ion, formed, as this is, by the preliminary protonation of nitric acid. The addition of a very strong acid, such as sulphuric acid, must raise the stationary concentration of nitric acidium ion. The addition of the strongest admissible base, the lyate ion of the system, namely, nitrate ion, must reduce it. Apart from such specific effects, neutral salts will raise the stationary concentration of the cation by the electrostatic effect of their ionic strength. Water will not directly affect the concentration of nitric acidium ion until it is in sufficient concentration to dominate autoprotolytic nitrate ions as a base, but it will then cause retardation.

Thus, our kinetic results are entirely consistent with the view that what is being measured is the rate of a *C*-nitration, which is what the reaction is stoicheiometrically, a nitration proceeding with first-order kinetics, as is to be expected for *C*-nitration in an aromatic nucleus already deactivated with two nuclear nitro-groups.

(3) *An alternative interpretation of the second kinetic step.* This is that the reaction measured is a rearrangement of *N*-methyl-*N* : 2 : 4-trinitroaniline to *N*-methyl-2 : 4 : 6-trinitroaniline, and that this change is followed by a relatively rapid *N*-nitration of its product to give *N*-methyl-*N* : 2 : 4 : 6-tetranitroaniline. The final *N*-nitration would certainly be relatively rapid, according to the specific investigation of it described in Section (*A*). It is necessary to try to decide whether the rate-controlling process is indeed a *C*-nitration, or is such a rearrangement of a nitro-amine.

Let us first consider the bearing of the kinetic results on this question. Since the original demonstration of the mechanism of aromatic nitration by nitronium ion, including the mechanism of formation of nitronium ion, was given largely on the basis of a pattern of kinetic results similar to those described in the preceding sub-section, it has been all too easy to accept the observation of such a kinetic pattern as pointing to a nitronium-ion process, supposing that the reaction being studied is stoicheiometrically a nitration. However, what the kinetic pattern, which has been encountered several times, discloses is that the sequence of reactions under observation is based on a pre-equilibrium protonation of something—so far as this evidence alone goes, of anything. Some collateral consideration must be imported in order to identify what is being protonated, and in the prototype investigation of aromatic nitration it was indeed thus that the protonated material was identified as molecular nitric acid. For example, when the measured reaction was a zeroth-order nitration of *any* substrate, the protonated substance has to be nitric acid because the substrate is not kinetically involved. When the measured reaction is a first-order nitration of a non-basic substrate, such as an aromatic hydrocarbon or halogenohydrocarbon, the protonated substance still has to be nitric acid, because (though for a different reason) there is still no chemically sensible alternative. When the measured reaction is a first-order nitration of a base, such as an amine, the protonated substance even yet has to be nitric acid, because protonation of the substrate, though possible, would be (and probably is) inhibitory. However, if we should be observing a first-order rearrangement catalysed by nitric acid, then a kinetic disclosure of pre-equilibrium protonation would have to be interpreted as protonation of the substrate itself. Should the rate-controlling process be indeed a rearrangement, proceeding through a conjugate acid formed from the substrate with nitric acid as the proton-donor, then the kinetic findings would be *exactly the same* as for a rate-controlling first-order nitration by nitronium ion formed in consequence of the autoprotolysis of nitric acid. Equilibria of this form, $AH^+ + B \rightleftharpoons A + BH^+$ generally preclude purely kinetic information as to whether A or B receives a proton, and thus our kinetic results do not solve the question.

The chemical considerations to be set beside the kinetic are those derived from past investigations concerning the rearrangement of aromatic nitro-amines,[6] and the order of the successive steps in the polynitration of *N*-methylaniline.[7]

As to the former, the rearrangements of aromatic nitro-amines, which are acid-catalysed, are intramolecular. If the catalyst is nitric acid, any concomitant hydrolysis of the *N*-nitro-group, *i.e.*, any reversibility in *N*-nitration, is likely to be inappreciable. With other catalysing acids it may be appreciable, but this does not provide a route of any importance for rearrangement; it simply protects, temporarily, a certain fraction of the original nitro-amine from intramolecular rearrangement, reducing the rate, but not the ultimate extent, of that process; it does not change its intramolecular character.

As to the latter, introduction of the third nuclear nitro-group into the vacant *ortho*-position in the polynitration of *N*-methylaniline is slow and difficult, hardly to be accomplished except by " mixed acid," and certainly not by nitric acid alone under

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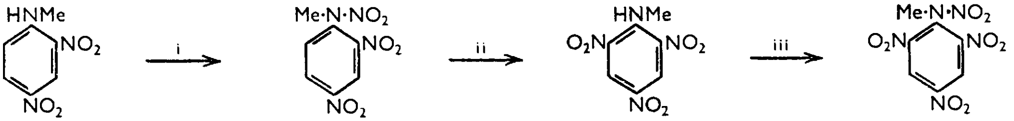
[6] Hughes and Jones, *J.*, 1950, 2678; Brownstein, Bunton, and Hughes, *Chem. and Ind.*, 1956, 981; *idem*, preceding paper.

[7] Glazer, Hughes, Ingold, James, Jones, and Roberts, *J.*, 1950, 2657.

conditions which suffice for the *N*-nitration of either *N*-methyl-2 : 4-dinitro- or *N*-methyl-2 : 4 : 6-trinitro-aniline. Apart from complications which may arise if nitrous acid is allowed into the system, the later stages of polynitration of *N*-methylaniline were shown to take the following course: *N*-methyl-2 : 4-dinitroaniline is first *N*-nitrated; then the formed *N*-methyl-*N* : 2 : 4-trinitroaniline rearranges to *N*-methyl-2 : 4 : 6-trinitroaniline; and then this is further *N*-nitrated to the final product, *N*-methyl-*N* : 2 : 4 : 6-tetranitroaniline.

In the light of this evidence, there seems no doubt that the second kinetic step of the double nitration of *N*-methyl-2 : 4-dinitroaniline is actually the nitro-amine rearrangement, which kinetically carries with it a relatively fast subsequent *N*-nitration. Our favoured interpretation of the two kinetic steps may thus be summarised as follows:



Kinetic first step: (i) N-Nitration by nitronium ion (somewhat fast; zeroth order).
Kinetic second step: (ii) Rearrangement (acid-catalysed; slow; first order). (iii) As (i).

(*C*) N-*Nitration of Other* N-*Alkylanilines*.—Measurements have been made on the zeroth-order *N*-nitrations of *N*-ethyl-, *N*-*n*-propyl-, and *N*-*n*-butyl-trinitroaniline by 3M-nitric acid in nitromethane at 25°, but the results are so similar to those described in Section *A* that we do not record them.

Measurements have also been made on the *N*-nitration in similar conditions of *N*-*iso*-propyl-2 : 4 : 6-trinitroaniline. In this case the zeroth-order law is not accurately obeyed, the order in substrate rising above zero especially towards the end of reaction. Also, the absolute rate is now only about 0·6 times as large as the nearly common rate applying to the other *N*-alkyl analogues. The beginnings of a transition to first-order kinetics, perhaps on account of steric shielding of the unshared nitrogen electrons, may here be under observation.

EXPERIMENTAL

Materials.—The amines nitrated and the products formed in the work described in Sections *A* and *B* were as previously described.[6, 7] Nitric acid was purified as usual by distillation twice with sulphuric acid in all-glass apparatus under an oil-pump vacuum, and stored (white powder) in the dark at −80°: it had its maximum freezing-point, and no nitrous acid was detected. Nitromethane was carefully purified, with gas-chromatography as a preliminary criterion and the nitration rates as a final criterion of purity. The most satisfactory material was obtained with omission of drying with phosphoric oxide and chromatography on alumina, which have sometimes been introduced. After passing nitrogen through the nitromethane under reflux for a day, the sample was dried ($CaSO_4$), filtered, distilled, and refractionated. The water content (Karl Fischer) was about 0·01% and the nitrous acid content [8] about 10^{-6}M. Gas-chromatography revealed no impurity except water, and it was verified by deliberate additions that the common impurities, formaldehyde and formic acid, were not being obscured in the chromatogram by the large peak due to nitromethane. In the few cases in which the presence of nitrous acid was required, it was added as solid dinitrogen tetroxide, prepared and purified in the customary way.

Kinetics.—These were followed essentially by our standard dilatometry.[1] However, the investigation in Part II (following paper) demanded a dilatometric precision at least an order of magnitude higher than previously needed, and most of the results herein recorded were obtained with this improved apparatus.

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[8] Mellon, " Organic Reagents in Organic Analysis," Blakiston Co., Philadelphia, 1941.