NITROUS ACID AND NITROSATION

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I. INTRODUCTION

In this review only brief mention is made of aspects of nitrous acid and nitrosation already covered elsewhere. Thus the chemistry of the nitrosonium ion (5, 57), the preparative aspects of aliphatic *C*-nitrosation (75), and the deamination of primary aliphatic amines (68) have not been considered.

This account begins with a discussion of the structure of molecular nitrous acid, the related chemical species that can be formed from it, and their thermodynamic stability in solution. These chemical species can be regarded as carriers of nitrosonium ion and are the active entities in nitrosation. The general pattern of nitrosation is demonstrated by consideration of *N*-nitrosation, *C*-nitrosation, *O*-nitrosation, and inorganic oxidation. Finally, mention is made of the decomposition of nitrous acid, which is of practical significance, and of the rates of formation of some active nitrosating species.

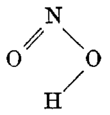
II. STRUCTURE OF NITROUS ACID

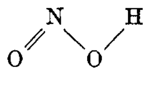
A. General theory

Observations as early as 1774 established the existence of nitrites and nitrous acid as separate chemical entities (44). In the gas phase nitrous acid exists in equilibrium with nitric oxide, nitrogen dioxide, and water vapor (80), the experimental equilibrium constant for the reaction

$$NO(g) + NO_2(g) + H_2O(g) \rightleftarrows 2HNO_2(g)$$

being 1.74 atm.$^{-1}$ at 25°C. The infrared absorption spectrum (28) suggests that the structure is a mixture of the two forms





with the trans form three times as abundant as the cis form. The data offer no evidence for a nitro form. The interatomic distances reported are

$$r(\text{O—H}) = 0.98\ \text{A.};\ r(\text{N—O}) = 1.46\ \text{A.};\ r(\text{N=O}) = 1.20\ \text{A.}$$

with the O—N=O bond angle 114° in the cis form and 118° in the trans form. An adequate theoretical treatment, using valence-bond methods, has been given for this structure (47).

B. Carriers of nitrosonium ion

In aqueous solution nitrous acid is best regarded as the hydroxylated form of the nitrosonium ion (NO^+). Other bases can replace OH^- to give compounds of the general type NOX.

Adequate spectroscopic evidence now exists for the species NO^+ [ultraviolet (12, 65) and Raman spectra (45)] and HONO [ultraviolet spectrum (12, 65)] in

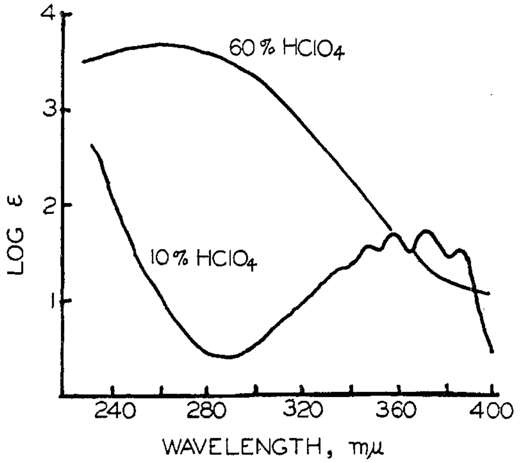


FIG. 1. Absorption spectra of nitrous acid in perchloric acid

TABLE 1

NO stretching frequency in nitrosyl compounds

Compound	NO Stretching Frequency	Reference	Compound	NO Stretching Frequency	Reference
	$cm.^{-1}$			$cm.^{-1}$	
$NO^+ClO_4^-$(c)	2329	(7)	N_2O_4(g)	1749	(69)
$NO^+HSO_4^-$(c)	2311	(8)	HONO(g), trans	1696	(28)
$Na_2[Fe(CN)_5NO]$(c)	1938	(23)	HONO(g), trans	1653–1681	(70)
NO(g)	1877	(5)	HONO(g), cis	1610–1625	(70)
NOCl(g)	1800	(5)			

aqueous solution (figure 1). In addition, some evidence exists for all the species $H_2NO_2^+$ (12), $NOHSO_4$ (12), N_2O_3 (17), N_2O_4 (39, 41), $N_2O_2^+$ (60), $N_2O_3^+$ (21), NOCl (53), NOBr (51), and NOCNS (31, 67) in aqueous solution.

An idea of the character of the nitroso group in some of these compounds can be gained from a consideration of the NO stretching frequency in a series of nitrosyl compounds, as in table 1. The interpretation of this table is that there is a gradation between the higher frequencies, corresponding to the triple bond of the nitrosonium ion, and the lower frequencies, typical of the double bond of the covalently bound nitroso group. The table provides independent evidence for part of a suggested order of decreasing electrophilic activity of nitrosating agents (26): NO^+, $NO\cdot OH_2^+$, NO·Hal, $NO\cdot NO_3$, $NO\cdot NO_2$, $NO\cdot OCOCH_3$, NO·OH.

III. EQUILIBRIA IN NITROUS ACID SOLUTIONS

A. General theory

In dilute acidic solution nitrous acid exists mainly as the molecule HONO. The equation for the formation of the nitrosonium ion is

$$\mathrm{HONO} + \mathrm{H^+} \rightleftarrows \mathrm{NO^+} + \mathrm{H_2O} \qquad K_1 = \frac{[\mathrm{NO^+}]a_w}{[\mathrm{HONO}]a_{\mathrm{H^+}}} \tag{1}$$

The formation of nitrosonium ion is favored by high acidity and low water activity; it has been detected (7, 65) in concentrated acids which provide both these conditions.

The existence of the nitrous acidium ion, $H_2NO_2^+$, in acidic media has been suggested by Bayliss and Watts (12). However, Seel (57), on the basis of a consideration of isoelectronic molecules, claims that its stability would be low. Kinetic evidence (26) suggests that $H_2NO_2^+$ may be an important nitrosating species in aqueous diazotization. It seems reasonable to suppose that $H_2NO_2^+$ is formed rapidly from molecular nitrous acid by a proton transfer, but its equilibrium concentration may be small.

Dinitrogen trioxide is the anhydride of nitrous acid.

$$2\mathrm{HONO} \rightleftarrows \mathrm{N_2O_3} + \mathrm{H_2O} \qquad K_2 = \frac{[\mathrm{N_2O_3}]a_w}{[\mathrm{HONO}]^2} \tag{2}$$

TABLE 2

Predicted and experimental equilibrium constants for the reaction

$$HONO + H^+ + X^- \rightleftarrows NOX + H_2O$$

in aqueous solution (25°C.)

X^-	NOX	K (calculated)	K (experimental)	References
—	NO^+	7×10^{-5}	2×10^{-7} (20°C.)	(78)
HSO_4^-	$NOHSO_4$	9×10^{-6}*	3×10^{-5}	(43)
NO_3^-	N_2O_4	4×10^{-5}*	3×10^{-3}	(39)
Cl^-	NOCl	6×10^{-4}*	1×10^{-3}	(42, 53)
Br^-	NOBr	9×10^{-2}*	5×10^{-2}	(51)
NO_2^-	N_2O_3	18	—	(78)

* These values have not previously been published.

The presence of N_2O_3 is favored by low water activity. The equilibrium concentration of dinitrogen trioxide depends on the square of the concentration of nitrous acid. Thus on diluting a blue (N_2O_3) acidic solution of nitrous acid, the color disappears to a greater extent than would be expected from normal proportional dilution.

Molecules of the type NOX, where X is Cl, Br, HSO_4, NO_3, etc., are formed in aqueous solution by the reaction:

$$HONO + H^+ + X^- \rightleftarrows NOX + H_2O \qquad K = \frac{[NOX]a_w}{[HONO][X^-]a_{H^+}}$$

Hence the concentration of the nitrosyl compound depends on acidity and water activity as well as on the concentration of X^-.

The nitric oxide-nitrosyl cation is formed from nitric oxide and the nitrosonium ion (60).

$$NO + NO^+ \rightleftarrows N_2O_2^+ \qquad K = \frac{[N_2O_2^+]}{[NO^+]P_{NO}}$$

$$K = 6 \times 10^{-3} \text{ atm.}^{-1} \text{ (20°C.)}$$

This ion does not seem to be important in nitrosation. $N_2O_3^+$ has been suggested as an intermediate in the oxidation of bromide ion by nitric acid (37).

In the following sections, thermodynamic calculations and experimental measurements of some of the important equilibrium constants are reviewed; a summary of the results is presented in table 2.

B. Thermodynamic calculations

Seel and Schwaebel (61) have used a cycle of hypothetical reactions to make a thermodynamic prediction of the equilibrium constant (K_1) for the formation of the nitrosonium ion. This cycle involved the chloride ion and yielded a value of K_1 impossibly small.

Recently (78), reaction cycles using the gas phase have been employed:

$$HONO(aq) + H^+(aq) + X^-(aq) \rightarrow NOX(g) + H_2O(l)$$

$$NOX(g) \rightarrow NOX(aq)$$

where X^- is a base such as nitrite ion. The free energies of formation of the substances in the first reaction are usually known, and the free-energy changes are readily found. In the case of the nitrosonium ion there is no X^-, and the free energy of the nitrosonium ion in the gas phase is found from the ionization potential of nitric oxide (5). The free energies of hydration required for the second step can be estimated by comparison with similar molecules. The difficulty of finding this solubility term is the limitation of the method. The overall free-energy changes, found by the addition of the two steps, are used to find the equilibrium constants in table 2. From the equilibrium constant for dinitrogen trioxide in the table, together with the acidity constant of nitrous acid (52) (5.1×10^{-4} at 25°C.), the value of K_2 is found to be 9×10^{-3}.

C. Experimental determinations of the equilibrium constants

Photometric or spectrophotometric measurement of the concentration of a nitrosyl compound in solution, whilst varying the acidity or anion concentration, is a convenient method of measuring the equilibrium constant. Kinetic and electrode methods have also been attempted.

Nitrous acid reacts with ammonium chloride in hydrochloric acid solution (55). The rate of the reaction is dependent on the hydrochloric acid activity and shows a maximum at 5 *M* hydrochloric acid. The decrease in rate beyond this maximum was stated (42) to be due to the formation of nitrosyl chloride from molecular nitrous acid. Thus the proportion of nitrosyl chloride was found, and the equilibrium constant was obtained (table 2). The method has also been used for nitrosyl bisulfate (43). However, in these experiments the effects of acidity and anion concentration have not been treated separately.

The equilibrium constant for the formation of nitrosyl chloride has also been determined (53) by a photometric method, using the yellow color of nitrosyl chloride. Nitrous acid in solutions of hydrochloric acid of increasing concentration showed increasing extinction with a maximum at 12 *M* hydrochloric acid. This was taken as complete conversion to nitrosyl chloride. Hence it was possible to estimate the concentration of nitrosyl chloride at intermediate concentrations of hydrochloric acid. The equilibrium constant obtained agreed well with that from the kinetic method above, but again acidity and chloride-ion concentration were altered simultaneously. It is possible that an appreciable amount of nitrosonium ion may form at the high acidities. The corresponding equilibrium constant for nitrosyl bromide has been determined in a similar fashion (51).

Longstaff and Singer (39) have measured the absorption spectra of nitrous acid in various concentrations of nitric acid. From their data for the proportion of dinitrogen tetroxide present it is possible to make a rough estimate of the equilibrium constant for the formation of this compound from molecular nitrous acid (table 2). There is some uncertainty due to lack of knowledge of the extent of the equilibrium $N_2O_4 \rightleftharpoons 2NO_2$ in these solutions. In 50 per cent nitric acid the concentrations of dinitrogen tetroxide and nitrous acid are approximately equal.

The absorption spectra of solutions of nitrous acid in sulfuric, hydrochloric, and phosphoric acids have been measured (12). Molecular nitrous acid exists at low acidities, and a strong absorption band at high acidities was assigned to the nitrosonium ion. In the middle range of acid concentrations, the existence in quantity of the nonabsorbing nitrous acidium ion was inferred because all of the nitrous acid could not be accounted for as HONO and NO^+.

The effect of coördinating anions can be eliminated by using perchloric acid (65). The absorption spectra shown in figure 1 are obtained. In the region of 50 per cent perchloric acid the spectra correspond to a mixture of nitrous acid and nitrosonium ion. By measuring the absorption band of nitrosonium ion (2600 A.) in this region, the ratio $[NO^+]/[HONO]$ has been found. Measurements on the absorption bands of nitrous acid gave similar results (78). The ratio $[NO^+]/[HONO]$ was independent of the total concentration of nitrous acid ($< 5 \times 10^{-3} M$), and there was no evidence for the presence of any other species. Calculation of the equilibrium constant from this ratio is difficult, because the necessary activity coefficients are unknown and there is no satisfactory method of estimating them. An approximate acidity function J_0 was used to find K_1 as given in table 2. An attempt to estimate the activity coefficient of nitrosonium ion by use of an NO/NO^+ electrode system was not successful (76).

D. Relation of equilibria to reactivity

The rate of reaction of nitrous acid is given generally by the expression

$$\text{Rate} = k(\text{HONO})(\text{H}^+)(\text{X}^-)(\text{B})$$

where (B) is the base being attacked.

The term $(\text{HONO})(\text{H}^+)(\text{X}^-)$ represents the concentration of NOX, which in turn depends on K_{NOX} and is the equilibrium concentration of the particular species dominating the kinetics.

The rate will vary according to the reactivity of NOX. This reactivity will depend on the polarity of NOX and its degree of hydration. From the experimental values of the rates of reaction in the presence of differing species X^-, combined with a knowledge of the equilibrium concentrations NOX, it is possible to estimate the relative reactivities of the NOX species.

The rate is also dependent on the relative basicity of the center being attacked. If a strong base in sufficiently high concentration is taken, it is possible to get conditions under which an equilibrium concentration of NOX is no longer achieved and the rate of formation of this species [rate $= k'(\text{HONO})(\text{H}^+)(\text{X}^-)$] then becomes rate determining.

In the subsequent sections this reaction scheme will be considered in relation to *N*-, *C*-, and *O*-nitrosations and to oxidations by nitrous acid.

IV. *N*-NITROSATION

N-Nitrosation is the term used to include the nitrosation of secondary amines, the deamination of primary aliphatic amines, and the diazotization of primary aromatic amines. These reactions are characterized by an initial attack of nitrous acid on the unshared pair of electrons on the nitrogen atom of the free amine.

TABLE 3

Comparison of specific rates of nitrosation (mole^{-2} l. min.$^{-1}$) and the base strength of amines at 25°C.

Amine	$10^{-6}k_1$	pK_a	Reference	Amine	$10^{-6}k_1$	pK_a	Reference
NH_3	0.39	9.2	(71)	$C_2H_5NH_2$	11	10.5	(72)
CH_3NH_2	2.9	9.7	(71)	$(CH_3)_2NH$	20	10.8	(72)

The kinetics and mechanism of diazotization have been reviewed by Schmid (51) and by Hughes, Ingold, and Ridd (26).

A. Second-order N-nitrosation by nitrous acid

The classical example of *N*-nitrosation is the deamination reaction

$$RNH_2 + HONO \rightarrow ROH + H_2O + N_2$$

and the rate measurements (71), variously expressed, reduce to:

$$\text{Rate} = k_1[RNH_2][HONO]^2$$

The aromatic amines parallel the aliphatic amines (54), the difference in product being due to the fact that the diazonium ion is stabilized by conjugation with the aromatic ring. At low acidity, where the proportion of free amine in the system is increased, the rate $= k_2[HONO]^2$, indicating that the nitrosating agent is dinitrogen trioxide, which, at low acidity, is consumed as rapidly as it is formed (26). The interpretation of these rate laws is:

$$2HONO \rightleftarrows N_2O_3 + H_2O$$

$$RNH_2 + N_2O_3 \rightarrow RNHNO + HONO$$

$$RNHNO + H^+ \rightarrow RN_2^+ + H_2O \quad \text{(rapid)}$$

The above scheme is modified in the presence of a base X^-, when the system is:

$$HONO + H^+ + X^- \rightleftarrows NOX + H_2O$$

$$RNH_2 + NOX \rightarrow RNHNO + H^+ + X^-$$

The first step of this scheme may under certain circumstances become rate determining (26).

In accord with the above mechanisms a marked correlation of the rate of diazotization with base strength of the amine is observed (table 3).

B. First-order N-nitrosation by nitrous acid

From the equilibria quoted in Section III is obtained

$$\frac{[N_2O_3]}{[NO^+]} = 130\,\frac{[HONO]}{a_{H^+}}$$

The relative concentrations of dinitrogen trioxide and nitrosonium ion in equilibrium with the bulk of the nitrous acid in solution thus vary with the conditions. The ratio of the reactivities of dinitrogen trioxide and nitrosonium ion will be somewhat less than this concentration ratio, but how much less is not known in the absence of any information on their specific reactivities. The expression, however, indicates the conditions under which a change in attacking species from dinitrogen trioxide to nitrosonium ion might be favored, viz.: low concentration of nitrous acid, high acidity, and, since nitrosonium ion is more reactive than dinitrogen trioxide, lowering of the base strength of the base that is attacked.

Such a change in mechanism has been realized with *o*-chloroaniline in dilute perchloric acid, where the rate law is found to be (26)

$$\text{Rate} = k_3[o\text{-ClC}_6\text{H}_4\text{NH}_2][\text{HONO}][\text{H}^+]$$

Since it was not possible to obtain a limiting rate $= k_4[\text{HONO}][\text{H}^+]$, which would be expected for a rate-determining step,

$$\text{H}_2\text{NO}_2^+ \rightarrow \text{NO}^+ + \text{H}_2\text{O}$$

the nitrous acidium ion is regarded as the active nitrosating species.

$$\text{H}^+ + \text{HONO} \rightleftarrows \text{H}_2\text{NO}_2^+ \quad \text{(rapid)}$$

$$\text{RNH}_2 + \text{H}_2\text{NO}_2^+ \rightarrow \text{products} \quad \text{(slow)}$$

Other arguments are also advanced in support of this view, but nitrosonium ion has not been excluded as the attacking electrophilic species.

Meta-directing groups may so lower the basicity of an amine that it cannot be diazotized in the usual concentrations of mineral acids (20). Increasing the concentration of mineral acid increases the concentration of the diazotizing species and reduces the concentration of the free base. If, however, glacial phosphoric acid is used, the dehydrating power favors the formation of dinitrogen trioxide and strong diazotizing conditions may be produced at low acidities.

The complexity of the products obtained following the initial formation of the diazonium ion will not be the concern of this review. The fact that the proportions of products obtained in diazotization reactions are different from others in which carbonium ions are formed suggests that many of the reactions take place by direct action with the RN_2^+ ion and are not preceded by a separation to R^+ and nitrogen (75).

C. Direct diazotization

An interesting reaction is that in which phenols are diazotized directly by nitrous acid, perhaps through an intermediate nitrosophenol. Philpot and Small (49) showed that *p*-cresol gave the diazonium compound on treatment with nitrous acid. They showed that *o*-nitroso-*p*-cresol was an intermediate by isolating it as its copper complex. In a separate experiment the intermediate reacted with nitric oxide to give the diazonium compound. Recently (74) the method has been extended, and excess nitrous acid has been used to effect a preparative diazotization of activated aromatic compounds. The mechanism is obscure.

V. *C*-NITROSATION

*C-N*itrosation is the attack on a carbon atom by a nitrosating species. It includes the nitrosation of reactive aromatic compounds—for example, phenols in the Liebermann reaction—and the formation of oximino derivatives from active methylene groups.

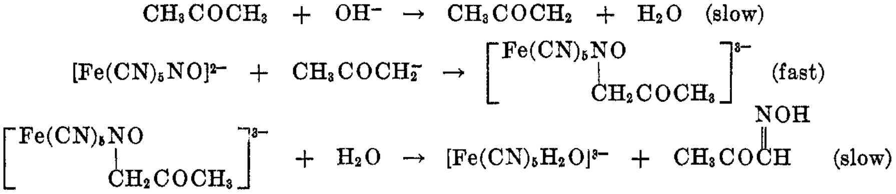
A. General theory

Nitrous acid in solution yields the electrophilic species dinitrogen trioxide and nitrosonium ion. The particular species undergoing reaction will depend on relative concentration effects and the basicity of the center attacked. For *C*-nitrosation in the aromatic nucleus the basicity of the carbon is low and attack can only be made by the more strongly electrophilic nitrosonium ion. With amines where the basicity of the nitrogen is higher, attack may be made by the weaker, electrophilic dinitrogen trioxide. When nitrous acid reacts with compounds containing an activated methylene group we might expect the reaction to proceed by way of the ion ($R\bar{C}HR'$). Since this ion will have a basicity comparable with that of an amine, we might expect such a compound to show attack by dinitrogen trioxide as the predominant reaction. This has not been experimentally realized. In the case of a methylene group activated by a carbonyl group it is found that an intervening reaction occurs and that the rate-determining step appears to be the attack of nitrosonium ion on the carbonyl oxygen (66). Work on the mechanism of *C*-nitrosation is limited.

B. C-Nitrosation by nitroprusside

The NO stretching frequencies in table 1 suggest that sodium nitrosopentacyano iron(II) (sodium nitroprusside) can be regarded as a carrier of the nitrosonium ion.

In alkaline solution reaction takes place with bases stronger than hydroxyl, and many of these are compounds which normally react with nitrous acid. The reaction with acetone appears to be as follows:



$$\mathrm{CH_3COCH_3} + \mathrm{OH^-} \rightarrow \mathrm{CH_3COCH_2} + \mathrm{H_2O} \quad \text{(slow)}$$

$$[\mathrm{Fe(CN)_5NO}]^{2-} + \mathrm{CH_3COCH_2^-} \rightarrow \begin{bmatrix} \mathrm{Fe(CN)_5NO} & \\ | & \\ \mathrm{CH_2COCH_3} & \end{bmatrix}^{3-} \quad \text{(fast)}$$

$$\begin{bmatrix} \mathrm{Fe(CN)_5NO} & \\ | & \\ \mathrm{CH_2COCH_3} & \end{bmatrix}^{3-} + \mathrm{H_2O} \rightarrow [\mathrm{Fe(CN)_5H_2O}]^{3-} + \mathrm{CH_3CO}\overset{\mathrm{NOH}}{\overset{\|}{\mathrm{C}}}\mathrm{H} \quad \text{(slow)}$$

The rate-determining step in this scheme is the enolization reaction. This is followed by the rapid formation of a colored compound, which then slowly decomposes to aquopentacyano iron(III) and pyruvic aldoxime. The total extent to which the reaction occurs is small, and it would appear to have no preparative value (36).

C. Catalyzed nitration

Ingold (27) has recognized two modes by which nitrous acid catalyzes nitration. With nitric acid, nitromethane, or acetic acid as solvent, the aromatic nitration of compounds not readily attacked by the nitrosonium ion is retarded by nitrous acid. The retardation is caused by the ionization of the nitrogen oxides dinitrogen trioxide and dinitrogen tetroxide to give nitrite ion and nitrate ion. These anions deprotonate the nitric acidium ion ($H_2NO_3^+$) and so reduce the production of the nitronium ion.

In aqueous solvents the nitration of compounds that react with the nitrosonium ion is accelerated in the presence of nitrous acid. This arises from the fact that nitrosonium ion is more stable in aqueous solvents than NO_2^+. For nitrations in the presence of nitrous acid the steps are considered to be nitrosation followed by oxidation.

$$\mathrm{ArH} + \mathrm{NO}^+ \rightarrow \mathrm{ArHNO}^+ \qquad \text{(slow)}$$

$$\mathrm{ArHNO}^+ \rightarrow \mathrm{ArNO} + \mathrm{H}^+ \qquad \text{(fast)}$$

$$\mathrm{ArNO} + \mathrm{HNO_3} \rightarrow \mathrm{ArNO_2} + \mathrm{HNO_2} \qquad \text{(fast)}$$

Evidence for this scheme has been obtained from consideration of the products formed and from kinetic data. The ratio of *o*- to *p*-nitrophenol is 1:9 in the nitrous acid-catalyzed nitration and this is the same as the ratio of *o*- to *p*-nitrosophenol in the simple nitrosation reaction. In the absence of nitrous acid the ratio of *o*- to *p*-nitrophenol is 7:3.

D. Nitrosation of phenols

The reactions of nitrous acid with phenols are complicated because of the variety of products that can be formed.

The reaction with "nitrous acid" in concentrated sulfuric acid is the well-known Liebermann color test for phenols (35). If a phenol gives a positive reaction (i.e., blue or green in concentrated sulfuric acid, red in dilute sulfuric acid, blue or green in alkali), then the position para to the hydroxyl group is unoccupied. Phenols containing activating groups stronger than alkyl groups and deactivating groups stronger than halogens do not give the above reaction. The detailed procedure is discussed elsewhere (77).

At lower acid concentrations difficulties arise because of the fact that nitrosophenols are attacked by nitric oxide. Nesmeyanov and Ioffe (46) have shown that *p*-nitrosophenol reacts with nitric oxide in acetic acid to give the diazonium compound. Bamberger (11) had earlier shown that nitrosobenzene in acetic acid reacted with nitric oxide to give a diazonium compound. The diazonium compounds formed in the above reactions may react with unchanged phenol, giving a series of colored by-products. Detailed methods for the preparation of nitrosophenols have been given by Hodgson (24).

Nitrous acid alone is known to produce nitro derivatives of phenols under some conditions. Whether this is due to nitrous acid acting as an oxidizing agent after initial nitrosation, homolytic attack by nitrogen dioxide, or nucleophilic attack by nitrite ion is not clear (25, 50).

With *o*-chloromercuryphenol, nitrous acid forms the red *p*-nitroso-*o*-chloromercuryphenol (19). Such colors are the basis of Millon's reaction, used as a test for phenols. Though the effect of the mercury chloride group would be expected to bring about deactivation of the phenol, it is found, surprisingly, that the compound is nitrosated faster than the unsubstituted phenol (76).

VI. *O*-NITROSATION

A. Organic O-nitrosation

Despite the ease with which alkyl nitrites are formed and hydrolyzed (64), comparatively little seems to be known concerning the mechanism of organic *O*-nitrosation. The nitro compounds, RNO_2, are prepared by nucleophilic displacement with nitrite ions,

$$RX + NO_2^- \rightarrow RNO_2 + X^-$$

while the nitrites, RONO, are prepared by the action of nitrous acid on the corresponding alcohol.

$$ROH + HONO \rightarrow RONO + H_2O$$

The nitrites can be readily converted to the nitro compounds by heating (14).

B. Inorganic O-nitrosation

The above reactions are paralleled in the formation of inorganic nitro and nitrito compounds. Bailar (10) cites the following reaction:

$$[Co(NH_3)_5{}^*OH]^{2+} + HONO \rightarrow [Co(NH_3)_5{}^*ONO]^{2+} + H_2O$$

where *O is labelled oxygen. All of the isotopic oxygen is retained in the nitrite pentammine cobalt(III) ion, indicating no fission of the cobalt–oxygen bond in the transformation. Kinetically (48) the rate of formation of the nitrito compound is given by

$$\text{Rate} = k[\text{hydroxyl complex}][\text{HONO}]^2$$

This form of kinetics indicates attack by dinitrogen trioxide, a result which would be expected, since the species $[Co(NH_3)_5OH]^{2+}$ is a relatively strong base ($K_a = 10^{-6}$) (15).

When the pink nitrito salt was heated either in the solid state or in solution, the yellow nitro isomer was formed. This, when treated with excess sodium hydroxide to re-form the hydroxypentammine cobalt(III) salt, released all the heavy oxygen as nitrite ion. Since the nitro compound can be made by direct nucleophilic displacement of chloride,

$$[Co(NH_3)_5Cl]^{2+} + NO_2^- \rightarrow [Co(NH_3)_5NO_2]^{2+} + Cl^-$$

a direct analogy with the organic case is established (29).

C. Oxygen-exchange reactions

Experiments have been carried out on the rate of exchange of ^{18}O between labelled water and nitrous acid. The reaction can be regarded as the nitrosation of water. Anbar and Taube (6) found that

$$\text{Rate} = k_4[\text{HONO}][H^+]$$

using [HONO] $< 10^{-2}M$ in phosphate buffer solutions. The postulated mechanism is

$$\mathrm{HONO + H^+ \rightleftarrows NO^+ + H_2{}^{18}O}$$

Bunton, Llewellyn, and Stedman (17) have measured the same exchange rate and have found

$$\text{Rate} = k_2[\mathrm{HONO}]^2$$

with [HONO] $> 10^{-2}M$ in mixtures of excess nitrite and perchloric acid. This rate law indicates the mechanism

$$\mathrm{2HONO \rightleftarrows N_2O_3 + H_2{}^{18}O}$$

Seel (58) has suggested that the conflicting results may be due to the method of interpretation used in the second experiment. Bunton, Llewellyn, and Stedman (17) state that the use of buffers in the first experiment makes exchange through nitrosyl compounds likely.

$$\mathrm{HONO + H^+ + X^- \rightleftarrows NOX + H_2{}^{18}O}$$

Acceleration of the exchange rate has been demonstrated for the addition of chloride and acetate. In the presence of phosphate buffer the rate is somewhat dependent on the buffer concentration, but the precise effect is not known. Another possible explanation of these experiments is that there is a change of attacking species from nitrosonium ion to dinitrogen trioxide as the concentration of molecular nitrous acid is increased. This is discussed in Section IV,B.

Oxygen exchange between nitric acid ($<$ 60 per cent) and water occurs only by way of "nitrous acid" in the solution (16). The mechanism depends on the particular nitrosyl compound present and on the concentration of nitric acid.

VII. OXIDATIONS BY NITROUS ACID

Table 4 lists kinetic data for a number of oxidation reactions by nitrous acid. These are regarded as being initial nitrosations followed by rapid rearrangements and transformations to give the oxidized products. The table gives the experimental rate law, the active nitrosating species indicated by this law, and the reactant base strength pK, where

$$K = \frac{[\mathrm{B}][\mathrm{H^+}]}{[\mathrm{BH^+}]} \quad \text{and} \quad \mathrm{B} = \text{reactant base}$$

It can be seen from table 4 that there is a correlation between base strength and mechanistic type, as would be expected for an electrophilic reaction. For the stronger bases the attacking species is dinitrogen trioxide, while for the weaker bases the more electrophilic nitrosonium ion is required. Sulfite and azide are the two strongest bases that normally react with nitrosonium ion, and a limiting rate, independent of the reactant concentration, can be achieved in these cases. The base is sufficiently strong to be able to use up the nitrosonium ions as rapidly as they are formed. However, it has been pointed out (17) that these experiments were conducted in solutions containing buffer anions (usually acetate); hence the reactive entities may be nitrosyl compounds.

TABLE 4

Oxidations by nitrous acid

Reactant	pK	Rate Law	Active Species	References
$(SO_3)_2NO^{2-}$	—	$k[HNO_2]^2$	N_2O_3	(33)
$H_2AsO_3^-$	7.7	$k[H_2AsO_3^-][HNO_2]^2$	N_2O_3	(4)
SO_3^{2-}	5.3	$k_1[HNO_2][H^+] +$ $k_2[SO_3^{2-}][HNO_2][H^+]$	NO^+	(59)
N_3^-	4.9	$k[HNO_2][H^+]$	NO^+	(61)
$S_2O_3^{2-}$	4.6	$k[S_2O_3^{2-}]^{0.2}[HNO_2][H^+]$	NO^+	(18, 63)
$C_2O_4^{2-}$	4.3	$k[C_2O_4^{2-}][HNO_2][H^+]$	NO^+	(79)
$HCOO^-$	3.6	$k[HCOO^-][HNO_2][H^+]$	NO^+	(38)
$SO_3NH_2^-$	0.6	$k[SO_3NH_2^-][HNO_2][H^+]$	NO^+	(34)
I^-	—	$k[I^-][HNO_2][H^+]$	NO^+	(13)
ClO_3^-, BrO_3^-	—	$k[\text{Hal } O_3^-][HNO_2][H^+]$	NO^+	(40)
H_2O_2	− 3	$k[H_2O_2][HNO_2][H^+]$	NO^+	(6, 22, 62)
Fe^{2+}	—	$k_1[Fe^{2+}][HNO_2]$	HONO	(3)
		$+ k_2[Fe^{2+}][HNO_2][H^+]$	NO^+	
		$+ \frac{k_3[Fe^{2+}][HNO_2]^2}{P_{NO}}$	N_2O_3	
Cu	—	$k[HNO_2][H^+]$	NO^+	(30)

The results for arsenious, oxalic, and formic acids have been stated here in a form different from that of the original report but still consistent with the experimental data.

VIII. DECOMPOSITION OF NITROUS ACID

Nitrous acid in solution undergoes disproportionation according to the equilibrium:

$$3HONO \rightleftarrows HNO_3 + 2NO + H_2O$$

giving nitric acid and nitric oxide as decomposition products.

The equilibrium constant has been measured (2, 32).

$$K = \frac{[H^+][NO_3^-]P^2_{NO}}{[HNO]^3} = 29 \text{ (25°C.)}$$

Abel and Schmid (2) have also determined the kinetics of decomposition.

$$\text{Rate} = k_5 \frac{[HONO]^4}{P^2_{NO}} - k_6[HONO][H^+][NO_3^-]$$

$$\log_{10} k_5 = 22.65 - \frac{6250}{T} \text{ (atm.}^2 \text{ mole}^{-3} \text{ l.}^3 \text{ min.}^{-1}\text{)}$$

The rate law can be accounted for by the following mechanism:

$$4HONO \rightleftarrows 2N_2O_3 + 2H_2O$$

$$2N_2O_3 \rightleftarrows N_2O_4 + 2NO$$

$$N_2O_4 + H_2O \underset{k_6}{\overset{k_5}{\rightleftarrows}} HONO + HNO_3$$

The last step is rate determining. The reaction involves evolution of nitric oxide gas; hence it is influenced by agitation (73) and the presence of certain catalysts in solution (9). It is concluded from these data that provided the reactions of nitrous acid with reactive substances are sufficiently rapid, its decomposition is not inconvenient. This conclusion has been confirmed experimentally (26).

It has been observed (65, 78) that solutions of nitrous acid in concentrated perchloric acid decompose rapidly in the region where the principal species present are the nitrosonium ion and molecular nitrous acid. It can be shown qualitatively that the maximum rate of decomposition corresponds to that perchloric acid concentration at which nitrosonium ion and nitrous acid are present in approximately equal amounts. This can be explained (12) by assuming that in these solutions dinitrogen trioxide is formed by the reaction:

$$NO^+ + HONO \rightleftarrows N_2O_3 + H^+$$

The maximum concentration of dinitrogen trioxide (and hence the maximum decomposition rate) occurs when the concentrations of nitrosonium ion and nitrous acid are approximately equal.

IX. RATES OF FORMATION OF ACTIVE SPECIES

Table 5 records the specific rates of formation of some active nitrosating species together with the Arrhenius activation energies *E*, where they are known. In each case the rate law has been expressed so as to correspond with the reaction of the nitrous acidium ion to form the carrier concerned.

$$HONO + H^+ \rightleftarrows H_2NO_2^+ + X^- \xrightarrow{k} NOX + H_2O$$

These rates of formation have been found from experiments in which a limiting maximum rate is observed, independent of the reactant concentration, but dependent on [HONO], [H$^+$], and [X$^-$]. Two oxygen-exchange results are given in

TABLE 5

Specific rates of formation of nitrosating species

Reactant	Rate Law	Species	k*	T	E	Reference
				°C.	kcal. mole^{-1}	
N_3^-	$k[HNO_2][H^+]$	NO^+	1.9×10^4	0	12.8	(61)
			1.9×10^5	25		
SO_3^{2-}	$k[HNO_2][H^+]$	NO^+	1.7×10^5	25	10	(59)
$H_2{}^{18}O$	$k[HNO_2][H^+]$	NO^+	1.3×10^5	25	—	(6)
H_2O_2	$k[HNO_2][H^+]$	NO^+	1.8×10^5	25	—	(6)
$C_6H_5NH_2$	$k[HNO_2][H^+][NO_2^-]$	N_2O_3	1.6×10^5	0	—	(56)
$C_6H_5NH_2$	$k[HNO_2][H^+][NO_2^-]$	N_2O_3	1.1×10^5	0	—	(26)
$H_2{}^{18}O$	$k[HNO_2][H^+][NO_2^-]$	N_2O_3	0.9×10^5	0	13.5	(17)
			6.9×10^5	25		
$NO(SO_3)_2^{2-}$	$k[HNO_2][H^+][NO_2^-]$	N_2O_3	8.6×10^5	0	20.1	(33)
			1.4×10^7	25		
$C_6H_5NH_2$	$k[HNO_2][H^+][I^-]$	NOI	8.2×10^4	0	—	(26)
$C_6H_5NH_2$	$k[HNO_2][H^+][Br^-]$	NOBr	7.0×10^4	0	—	(26)
NO_3^-	$k[HNO_2][H^+][NO_3^-]$	N_2O_4	1.6	25	—	(2)
I^-	$k[HNO_2][H^+][NO_3^-]$	N_2O_4	1.6	25	—	(1)
$C_6H_5NH_2$	k[acetate]$[HNO_2][H^+][NO_2^-]$	N_2O_3	5.1×10^6	0	—	(26)
$C_6H_5NH_2$	k[phthalate]$[HNO_2][H^+][NO_2^-]$	N_2O_3	2.2×10^7	0	—	(26)

* Units of k are mole l.$^{-1}$ and min.$^{-1}$

table 5, because in the absence of further evidence it is possible that both are valid under the appropriate conditions (Section VI,C).

Since all the measurements of the rate of formation of nitrosonium ion have been carried out in buffer solutions, it has been suggested (17) that the buffer anions take part in the reaction and that the observed rates are the rates of formation of nitrosyl compounds. The constancy of the specific rates quoted, however, indicates that they may have some validity.

The values of the uncatalyzed specific rate of formation of dinitrogen trioxide agree well with the exception of the determination made by using the nitrosyl disulfonate ion (33). In this case acetate buffer was used, and it is shown conclusively in the last two entries in table 5 that carboxylate ions have a marked catalytic effect (26) on the formation of dinitrogen trioxide.

$$H_2NO_2^+ + {}^-OCOCH_3 \rightleftarrows NO\cdot OCOCH_3 + H_2O$$

$$NO\cdot OCOCH_3 + NO_2^- \rightarrow N_2O_3 + {}^-OCOCH_3$$

Hence the discrepancy of k and E for this case are explained.

From table 5 it can be seen that the rate of substitution in the nitrous acidium ion follows the order $NO_2^- > I^- > Br^- > NO_3^-$.

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