458. *The Solvolysis of Acetic Anhydride in 50% Aqueous Acetone catalysed by Nitrite Ion.*

By E. B. Lees and B. Saville.

Kinetic studies have shown that the solvolysis of acetic anhydride is markedly catalysed by nitrite ions. The rate-determining step involves nucleophilic attack by nitrite ion on the anhydride molecule (specific rate constant $\sim$9 l. mole^{-1} min.$^{-1}$) to produce an intermediate, believed to be acetyl nitrite, which undergoes relatively instantaneous decomposition in the buffered solvent to regenerate nitrite ion. Evidence for the intermediate formation of acetyl nitrite is afforded by experiments where α-naphthylamine was included in the solvolysis systems: 4-amino-1 : 1′-azonaphthalene was then formed in amounts which could be calculated on the assumption of acetyl nitrite's functioning as a nitrosating agent. A new procedure for the determination of acetic anhydride in low concentrations has been developed.

Catalytic effects in the solvolysis of acetic anhydride have formed the subject of several studies during the last few decades. Carboxylate ions [1] and pyridine or quinoline derivatives [2] influence it and their effects have been quantitatively investigated. Catalysis by these bases can usually be ascribed to the formation, by nucleophilic attack of the catalyst on the anhydride carbonyl group, of a very easily hydrolysable intermediate which rapidly regenerates the catalyst. Catalysis by acetate ion [1] cannot be interpreted by a similar scheme and one may assume that in this case Brönsted-type classical base-catalysis is responsible. Catalysis by nitrite ions has however not been recorded, though Liddell and Saville [3] show that nitroso-compounds were formed when acetic anhydride was added to an acetate-buffered solution of sodium nitrite and an aromatic amine in aqueous acetone. The present paper deals with the action of nitrite ion upon acetic anhydride in aqueous acetone, in both the presence and the absence of α-naphthylamine as nitrosatable substrate.

Experimental

Materials.—" AnalaR " acetic anhydride was redistilled through a helix-packed column, and the fraction boiling at 139·1°/743 mm. collected. Acetone was purified by refluxing it over potassium permanganate for several days before final distillations. " AnalaR " α-naphthylamine was purified by several recrystallisations of its hydrochloride from water and liberation

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[1] Kilpatrick, *J. Amer. Chem. Soc.*, 1928, **50**, 2891.

[2] Bafna and Gold, *J.*, 1953, 1406; Gold and Jefferson, *J.*, 1953, 1409.

[3] Liddell and Saville, *Chem. and Ind.*, 1957, 493.

of the free base, which was distilled. It had m. p. 49° and contained no β-isomer as judged by chromatography on alumina with toluene as eluent, or by chromatography of the dye formed by coupling with diazotised *m*-chloroaniline: mixtures of the α- and β-isomers could be separated by these techniques.

" AnalaR " sodium acetate and sodium nitrite were used as supplied, after drying for several hours in a vacuum-oven.

Determination of Acetic Anhydride.—This was done colorimetrically by a new procedure based on the principle described by Liddell and Saville.[3] Calibration was by adding 1 ml. of a stock solution of acetic anhydride (0·0001—0·0013M) in acetone to about 20 ml. of reagent A contained in a 25 ml. volumetric flask. [The reagent A is made up as required by mixing 20 ml. of N-sodium nitrite and 80 ml. of N-sodium acetate (both aqueous) with 100 ml. of 0·006M-α-naphthylamine solution in acetone. It is important to keep this reagent from strong daylight. Determinations of acetic anhydride were performed in a dark corner of the laboratory; there the reagent could be used for at least 2—3 hr.] The contents of the flask were made up to the mark with further reagent A, mixed, and set aside for 15—20 min., the intensity of the resultant orange-coloured solution being measured on a Hilger " Spekker " absorptiometer with 601 filters and 1 cm. cells. Reagent A was used as the control.

There existed a linear relation between absorptiometer readings (S) and the anhydride concentrations in the stock solutions. Up to 100 times the concentration of acetic acid did not interfere. Typical results are presented in Table 1.

TABLE 1. *Determination of acetic anhydride with reagent A.*

[Acetic anhydride] (10^{-4} mole/l.)	2·5	5·0	7·5	10	12·5
S	0·151 ±0·001	0·302 ±0·001	0·454 ±0·002	0·608 ±0·002	0·761 ±0·003

Rate of Decomposition of Acetic Anhydride in Aqueous Acetone.—To check the validity of the above procedure for anhydride determinations when applied to kinetic measurements, the rates of hydrolysis of acetic anhydride in several aqueous-acetone solvents were measured as follows.

To the solvent (100 ml.) in a stoppered flask, which had been brought to 25° in a thermostat, was added 0·1M acetic anhydride (in acetone) (1 ml.). Aliquot parts (1·00 ml.) were withdrawn with a calibrated tuberculin syringe and added at recorded times to about 20 ml. of reagent A and made up to 25 ml. with further reagent A, further procedure being as described above.

First-order plots of excellent linearity were obtained by plotting $\log_{10} S$ against time, from which the first-order rate constants k_0 were calculated in the conventional manner.

In Table 2 these rate constants are compared with Gold's results,[4] the good agreement indicating that the new procedure is satisfactory for anhydride determinations in kinetic runs.

TABLE 2. *First-order rate constants (k_0) for hydrolysis of acetic anhydride in aqueous acetone at 25°.*

Acetone (wt. %) in solvent	0	20	30	50
k_0 (min.$^{-1}$), this work	0·1495	0·0510	0·0252	0·00584
Gold's work	0·1575	0·0520	0·0265	0·00587

Effect of Nitrite-ion Concentration on the Rate of Hydrolysis of Acetic Anhydride in 50% (v/v) Aqueous Acetone.—Stock 1·00M-solutions of sodium nitrite and sodium acetate in 50% (v/v) aqueous acetone were mixed with extra solvent to produce several solutions ranging from 0·1 to 0·5M in sodium acetate and 0·005 to 0·040M in sodium nitrite. The acetate ion was present to buffer the effect of acetic acid liberated in the anhydride hydrolysis which might otherwise reduce the effective nitrite-ion concentration during reaction.

100 ml. of these acetate–nitrite solutions at 25° were used and the rate of decomposition of added acetic anhydride (initial concentration 0·001M) was followed as just described. The rates were found to be of strictly first order with respect to the anhydride, and the values of the constants were calculated graphically in the usual manner.

Reaction of Acetic Anhydride with Acetate–Nitrite Solutions in the Presence of α-Naphthylamine.—The almost quantitative formation of 4-amino-1 : 1′-azonaphthalene from acetic anhydride, excess of sodium acetate, sodium nitrite, and α-naphthylamine in aqueous acetone

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[4] Gold, *Trans. Faraday Soc.*, 1948, **44**, 506.

was observed as follows: To sodium acetate (41 g., 0·5 mole) and sodium nitrite (6·9 g., 0·1 mole) in water (500 ml.) was added α-naphthylamine (14·3 g., 0·1 mole) in acetone (450 ml.). The solution was cooled, and acetic anhydride (1·02 g., 0·01 mole) in acetone (50 ml.) added with stirring which was continued for 30 min., before extraction with chloroform (3 × 150 ml.). The combined extracts were washed with three equal volumes of water, dried, and evaporated under reduced pressure. The residue was taken up in toluene (100 ml.), and concentrated hydrochloric acid (10 ml.) added. The mixture was shaken vigorously and the solid hydrochlorides were filtered off, washed with a little acetone, and suspended in N-hydrochloric acid (300 ml.) at 50°. The naphthylamine hydrochloride dissolved, leaving the crude dye hydrochloride which was filtered off and shaken with aqueous ammonia and chloroform to extract the free base. The extracts were dried and evaporated, yielding the crude dye which was recrystallised from benzene–light petroleum (b. p. 60—80°). 4-Amino-1 : 1′-azonaphthalene (2·77 g., 93·6% based on anhydride) was obtained, with m. p. and mixed m. p. 184—185° (Found: C, 80·6; H, 5·2. Calc. for $C_{20}H_{15}N_3$: C, 80·8; H, 5·1%).

Kinetics of Azo-dye Formation.—Aqueous-acetone solutions containing α-naphthylamine, sodium nitrite, sodium acetate, and acetic acid were prepared from stock solutions of each component. 100 ml. portions of these solutions (at 25°) were treated with M/100-acetic anhydride (1·00 ml.) in acetone, and 5 ml. portions then added to N-ammonia in 50% aqueous acetone (20 ml.) to stop the reaction (excess of anhydride is decomposed instantaneously) and the intensity of colour was measured as before except that 3·0 cm. cells were used.

S_∞ being the reading at completion of the reaction (taken at ten half-lives), and S_t a reading at time t, plots of $\log_{10} (S_\infty - S_t)$ against time gave very good straight lines from which first-order constants were obtained. It had been previously demonstrated that the relation between azo-dye concentration and absorptiometer reading was linear.

Results and Discussion

The first-order rate constants for decomposition of acetic anhydride bore a linear relation to both acetate- and nitrite-ion concentration. Thus the relation $k_0 = k_H + k_1[AcO^-] + k_2[NO_2^-]$ is obeyed where k_H is the first-order rate constant for the uncatalysed reaction and k_1 and k_2 are the respective catalytic coefficients (in l. mole^{-1} min.$^{-1}$) for acetate and nitrite ions.

The value of k_H was found by studying the hydrolysis of the anhydride in 50% aqueous acetone in the absence of other substances. The effect of added acetate ion on the decomposition rate in the absence of nitrite ion allows k_1 to be calculated, and the value of k_2 was found by plotting k_0 against nitrite-ion concentration at constant acetate concentration. Results are summarised in Table 3.

Table 3. *The effect of nitrite- and acetate-ion concentration on the first-order rate constants* (k_0) *for the decomposition of acetic anhydride* (*initially* 0·001M) *in* 50% *aqueous acetone at* 25°.

	NaOAc 0·100		NaOAc 0·200		NaOAc 0·500 mole l.$^{-1}$	
[$NaNO_2$] (mole l.$^{-1}$)	k_0 (min.$^{-1}$)	Mean k_2 *	k_0 (min.$^{-1}$)	Mean k_2 *	k_0 (min.$^{-1}$)	Mean k_2 *
0·000	0·0203	8·85	0·0305	9·25	0·0603	9·35
0·005	0·0645		0·0770		0·106	
0·010	0·109		0·124		0·150	
0·020	0·198		0·216		0·247	
0·030	0·286		0·308		0·340	

In absence of $NaNO_2$ and NaOAc, k_0 (= k_H) = 0·0110 min.$^{-1}$.
Mean calc. value of k_1 = 0·0986 l. mole^{-1} min.$^{-1}$.
* l. mole^{-1} min.$^{-1}$.

The fact that nitrite ion functions purely as a catalyst in these reactions was established in two ways. Some runs were performed with 0·01M-nitrite and 0·2M-acetate with initial concentrations of 0·01M- and 0·02M-acetic anhydride. When the concentration of acetic anhydride in the system had fallen to about 0·001M measurements on the rate of decomposition were commenced. The rate constants thus obtained were only very slightly lower than that recorded in Table 3, indicating very little loss of nitrite from the systems. Table 4 gives this comparison.

TABLE 4. *Rate data for decomposition of acetic anhydride in* 0·2M-*acetate plus* 0·01M-*nitrite in* 50% *aqueous acetone at* 25°.

Initial [Ac_2O] (mole $l.^{-1}$)	0·001	0·010	0·020
Final decompn. rate, k_0 ($min.^{-1}$)	0·124	0·118	0·111

Determination of the residual nitrite concentration after reaction with acetic anhydride indicated that even in the system where the initial anhydride concentration was double the nitrite concentration, the latter was reduced to only some 92% of its original value. In a purely aqueous system there was no overall utilisation of nitrite at all, even after the reaction with a ten-fold excess of anhydride.

The rate of decomposition of acetic anhydride in these systems can be accurately defined by the equations

$$-d[Ac_2O]/dt = k_0[Ac_2O]$$

and

$$k_0 = k_H + k_1[AcO^-] + k_2[NO_2^-]$$

where the last term refers to the nitrite-catalysed hydrolysis.

Since the rate of the latter process, in which we are primarily interested, is directly proportional to the nitrite-ion concentration it seems that one could set up a crude mechanistic interpretation of this reaction by postulating that nitrite ion attacks acetic anhydride in a bimolecular rate-controlling step to yield some intermediate which decomposes relatively instantaneously producing acetic acid and nitrite ion which is then available for participation in further reactions:

$$O{\cdot}NO^- + Ac_2O \xrightarrow[\text{Slow}]{k_2} \text{Intermediate} \xrightarrow[\text{Fast}]{H_2O} O{\cdot}NO^- + 2AcOH$$

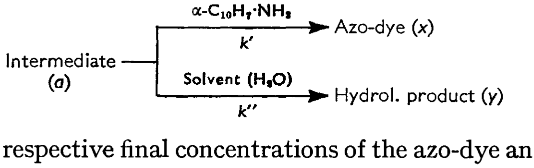
k_2 then becomes the second-order rate constant for the slow step.

The existence of the intermediate would be substantiated if some substance included in the system competed with the solvent molecules (water) for the intermediate to produce an easily recognisable product. Now, addition of α-naphthylamine to the nitrite–acetic anhydride reaction led to the formation of 4-amino-1 : 1′-azonaphthalene (referred to below as " azo-dye "). Correspondence between the amounts of this dye produced and the calculated fraction of intermediate which could be formed if its hydrolysis were avoided, and the kinetics of azo-dye formation, were in fact consistent with existence of the intermediate.

TABLE 5. *Effect of* α-*naphthylamine concentration on azo-dye formation from* 0·1M-*nitrite and* 0·4M-*acetate in* 50% *aqueous acetone after reaction of* 0·00004M-*acetic anhydride at* 25° *for* 20 *minutes.*

[α-Naphthylamine] (10^{-4} mole $l.^{-1}$)	1·25	2·5	5·0	12·5	25	50	100
Mole ratio of azo-dye formed to Ac_2O used	0·420	0·561	0·710	0·909	0·924	0·924	0·911

In solutions where the acetate and the nitrite ion were in effectively constant excess over the initial anhydride, the final ratio of dye formed to anhydride used varied with initial α-naphthylamine concentration as shown, for example, in Table 5. If the azo-dye is produced by competition of the α-naphthylamine with solvent for the intermediate, the theory of competing first-order reactions can be applied to the scheme:



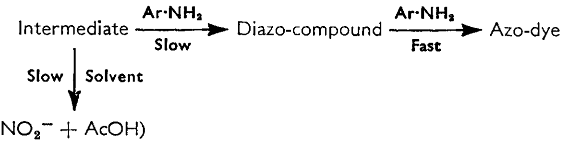
where *x* and *y* are the respective final concentrations of the azo-dye and hydrolysis products, *a* is the initial concentration of intermediate, and k' and k'' are the first-order rate constants for the two competing reactions; this leads to the expression:

$$a/x = (k''/k')[\alpha\text{-}C_{10}H_7{\cdot}NH_2] + 1 \qquad . \; . \; . \; . \; . \; . \quad (1)$$

where $$k[\alpha\text{-}C_{10}H_7{\cdot}NH_2] = k'$$

This has been tested graphically by plotting a/x (inverse of row 2 in Table 5) against the reciprocal of the amine concentration. A straight line results, from the slope of which one obtains a value of $k''/k = 1{\cdot}54 \times 10^{-4}$. α-Naphthylamine at this concentration ($1{\cdot}54 \times 10^{-4}$M) competes on equal terms with the solvent for the intermediate. Since the azo-dye must be formed from two molecules of α-naphthylamine but the kinetics of its formation from the intermediate are of the first order in amine concentration, it appears that a subsequent even faster step is involved whereby the primary reaction product of intermediate and α-naphthylamine reacts with another molecule of α-naphthylamine.

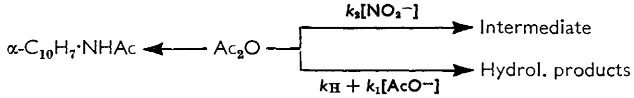
Very probably the complete decomposition of the intermediate can be represented, thus: *



This scheme requires that the rate of azo-dye formation from the intermediate is of the first order in α-naphthylamine, as observed experimentally.

Let it now be supposed that the intermediate does not decompose in the solvent, and consider what hypothetical relation exists between the final intermediate concentration and the initial anhydride concentration. (This is one way, in fact, of calculating what fraction of the anhydride decomposes *via* the intermediate.)

In terms of the scheme:



the expression for the final intermediate concentration becomes

$$[\text{Intermediate}] = [Ac_2O]k_2[NO_2^-]/k_d \qquad . \; . \; . \; . \; . \; . \quad (2)$$

where k_d is the total decomposition rate $= k_H + k_1[AcO^-] + k_2[NO_2^-] + k_N$ and k_N is the first-order rate constant for α-naphthylamine acetylation, which has the value 0·023 min.$^{-1}$ for 0·005M-α-naphthylamine. The very good agreement between calculated intermediate and azo-dye concentrations shown in Table 6 again confirms the reaction mechanism postulated.

TABLE 6. *Formation of azo-dye in reaction of* 0·005M-α-*naphthylamine, varying nitrite concentrations,* 0·0001M-*acetic anhydride,* 0·02M-*acetate, and* 0·005M-*acetic acid in* 50% *aqueous acetone, and the concentration of intermediate calculated from equation* 2.

[Nitrite] (mole l.$^{-1}$)	0·005	0·010	0·020	0·030	0·040
[Intermediate] (10^{-6} mole l.$^{-1}$)	44·5	61·4	76·1	82·2	86·4
[Azo-dye] (mole l.$^{-1}$) (±1%)	43·2	59·3	75·7	82·7	86·4

Finally, if the azo-dye is formed in a very fast step from the intermediate, the kinetics of dye formation are really the kinetics which would be observed for the formation of the

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* All processes are assumed to be very fast compared with the rate of formation of the intermediate.

intermediate if the latter could be preserved from subsequent reaction. Kinetic analysis of the system:

$$\underset{(\beta)}{\text{Intermediate}} \xleftarrow{k_2[NO_2^-]} \underset{(\alpha - \beta - \gamma)}{Ac_2O} \xrightarrow{k_a} \underset{(\gamma)}{\text{Other products}}$$

where k_a is the sum of the first-order constants for alternative decomposition routes not involving nitrite reaction, α = initial[Ac_2O], and β and γ the time-variable concentrations of the respective product species, yields the expression:

$$-\log(\beta_\infty - \beta) = (k_2[NO_2^-] + k_a)t/2{\cdot}303 + \text{constant} \quad . \quad . \quad . \quad (3)$$

Since β is proportional to the dye concentration which is in turn proportional to the absorptiometer readings S it follows that the constants obtained by plotting $\log_{10}(S_\infty - S_t)$ against time (see Experimental section, rates of azo-dye formation) should equal the total rate of decomposition of the acetic anhydride, *i.e.*, $k_a + k_2[NO_2^-]$. We can also arrive at the value of $(k_a + k_2[NO_2^-])$ in another way, not presupposing an intermediate. This is by direct measurement of the rate of decomposition of the anhydride in nitrite–acetate solutions in the absence of α-naphthylamine, and adding an independently determined term for the rate of acetylation of α-naphthylamine. Table 7 indicates the excellent correspondence between acetic anhydride decomposition rates obtained (*a*) by measurement of azo-dye production and (*b*) by direct measurement of residual anhydride. The figures in the difference column are seen to be approximately constant and equal to the independently determined rate constant for the reaction between α-naphthylamine and acetic anhydride. Both sets of measurements are linear with nitrite-ion concentration and yield 8·73 l. mole^{-1} min.$^{-1}$ for the catalytic coefficient (k_2) of nitrite ions, in satisfactory agreement with the values in Table 3.

TABLE 7. *Rates of decomposition of acetic anhydride* (0·0001M) *in the absence and the presence of α-naphthylamine at* 25°.

	First-order rate constants (min.$^{-1}$)		
[NO_2^-] (mole l.$^{-1}$)	By measurement of residual anhydride *	By measurement of azo-dye produced †	Difference
0·00	0·0434	—	—
0·005	0·0743	0·0980	0·025
0·010	0·118	0·140	0·022
0·020	0·206	0·230	0·024
0·030	0·295	0·319	0·024
0·040	0·381	0·400	0·019

* Initially, 0·20M-AcO^-, 0·005M-AcOH, 50% aqueous acetone.

† Initially, 0·20M-AcO^-, 0·005M-AcOH, 50% aqueous acetone, +0·005M-α-naphthylamine.

We next discuss the chemical identity of the intermediate. The most obvious initial product from nitrite ion and acetic anhydride under non-solvolytic conditions is acetyl nitrite,[5] whole potentialities as a nitrosating agent have not been studied. One may suppose (by analogy with the nitrating properties of acetyl nitrate towards benzene [6]) that it would be able to nitrosate effectively any convenient nucleophilic centre, Z:, by the general scheme, $Z{:} + O{=}N{\cdot}OAc \longrightarrow Z{\cdot}N{=}O + {:}OAc$. Presumably both water and α-naphthylamine react in this way in the systems discussed, so that the azo-dye formation can be explained rationally in terms of the equations:

$$\begin{tikzcd} ON{\cdot}OAc + Ar{\cdot}NH_2 \arrow[r] & Ar{\cdot}NH{\cdot}N{=}O + HOAc \arrow[d, shift left] & \\ & Ar{\cdot}N{=}N{\cdot}OH \arrow[u, shift left] \arrow[r, "Ar{\cdot}NH_2"] & Dye \end{tikzcd}$$

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[5] Ferrario, *Gazzetta*, 1910, **40**, II, 95; Francesconi and Cialdea, *ibid.*, 1904, **34**, I, 435.

[6] Cohen and Wibaut, *Rec. Trav. chim.*, 1935, **54**, 409.

Further to this argument we have noticed that methyl nitrite (inferred from the nitrosating properties of an organic extract) may be formed when sodium nitrite and acetic anhydride react in methanol.

It is not thought likely that heterolysis of the carbonyl–oxygen bond occurs during acetyl nitrite solvolysis. Thus, acetate ion did not significantly affect the nitrite-ion catalytic coefficient, k_2, as Table 3 shows. If acetate ion had attacked the acetyl nitrite to yield acetic anhydride, this would have led to considerable decrease of k_2, with increase of acetate-ion concentration.

There is also some indirect evidence that the acetyl nitrite may attack the acetone in our solvent. In certain cases there was a very small loss of nitrite ion from the system after reaction with excess of acetic anhydride and in these instances an extremely small quantity of cyanide was found in the final solution. The cyanide could arise from the reaction of the anhydride with hydroxyiminoacetone [7, 8] which could be formed by nitrosation of the acetone by the acetyl nitrite.

The high reactivity of nitrite ions towards other reactive carbonyl centres has been checked by demonstrating that *p*-substituted benzoyl chlorides, ethyl chloroformate, and succinic and butyric anhydride also give approximately equimolar amounts of azo-dye with nitrite–acetate–α-naphthylamine solutions. This high order of nitrite-ion reactivity was expected because it had been shown that the anions of oximes [8] and hydroxamic acids,[9] which like nitrite ion contain the $^{-}\dot{\text{N}}\text{–O}^{-}$ arrangement, are powerful nucleophilic reagents towards certain organophosphorylating agents, and because phosphoryl derivatives behave similarly to carbonyl compounds.[10]

These nucleophilic reagents sometimes exhibit about 1000 times the reactivity of simple buffer anions of the same basicity. The decomposition of the acylated or phosphorylated products which often occurs extremely rapidly usually involves nitrogen–oxygen bond fission in a similar manner to that postulated for decomposition of acetyl nitrite.

No satisfactory explanation has been advanced for the extremely high nucleophilicity : basicity ratio exhibited by anions of this type towards carbonyl or phosphoryl derivatives.

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[7] Saville, *Analyst*, 1957, **82**, 269.
[8] Green and Saville, *J.*, 1956, 3887.
[9] Green, Sainsbury, Saville, and Stansfield, *J.*, 1958, 1583.
[10] Dostrovsky and Halmann, *J.*, 1953, 508 *et seq.*