

THE NITRATION OF PYRIDINE

Joan and John Jones

Department of Chemistry, The University, SHEFFIELD 10, U.K.

(Received 18 June 1964)

Recent work has shown that the nitration of quinoline, imidazole⁽¹⁾ and isoquinoline⁽²⁾ in nitric/sulphuric acid mixtures is a slow reaction involving electrophilic attack by nitronium ion on protonated heterocyclic base. So it would not appear unreasonable to assume that the nitration of pyridine itself in acidic media involves attack by nitronium ion on protonated pyridine.

To date there appears to be but one report on the nitration of pyridine in an aprotic medium⁽³⁾.

Now solid nitronium fluoborate, NO_2BF_4 is stable to 170°C and is known from infra-red studies^(4,5) to consist of nitronium and fluoborate ions. Its solutions in the aprotic solvent sulfolane (tetrahydrothiophen 1:1 dioxide) are stable, have strong nitrating powers, and have been shown by conductance studies to contain free ions, ion-pairs and higher ionic aggregates⁽⁶⁾. Accordingly we chose to use this stable nitrating system in our study of pyridine nitration under aprotic conditions.

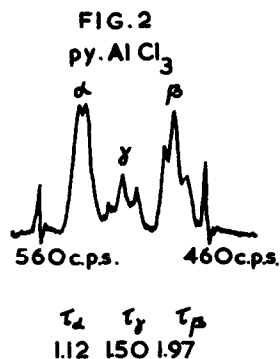
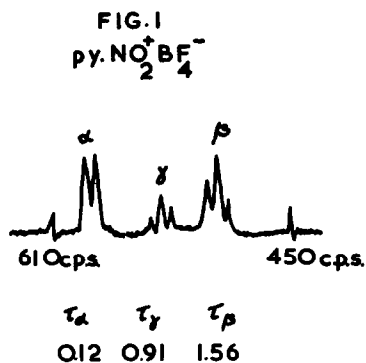
In our first series of experiments solutions of pyridine and nitronium fluoborate in sulfolane were prepared and mixed in a heated (30°C) dry box under dry nitrogen, and their u.v. spectra were taken in thin (0.1 mm, 0.3 mm and 0.6 mm) cells. Pyridine in sulfolane displays twin maxima at 252 m μ ($\epsilon_M = 2300$) and 256 m μ ($\epsilon_M = 2200$) with shoulders at 247 m μ ($\epsilon_M = 2000$) and 263 m μ ($\epsilon_M = 1700$), and a much more intense absorption at 199 m μ , ($\epsilon_M = 8700$). When equimolar amounts of 0.01M pyridine and nitronium fluoborate solutions are mixed an immediate change in the spectrum of pyridine occurs, resulting in a single maximum at 256 m μ ($\epsilon_M = 4800$) with shoulders at 252 m μ ($\epsilon_M = 4300$) and 263 m μ ($\epsilon_M = 3500$). There is also a change in the spectrum at shorter wavelengths, the maximum absorption lying at 202 m μ ($\epsilon_M = 6900$). Such results were reproducible. We compared this spectrum with that of 2-nitro-pyridine, 3-nitro-pyridine and pyridine-N-oxide, all of which it did not resemble. However, there was a close resemblance to the spectrum of pyridinium fluoborate dissolved in sulfolane, which displays maxima at 257 m μ ($\epsilon_M = 5700$) and 202 m μ ($\epsilon_M = 3400$) and shoulders at 252 m μ and 262 m μ . This was our first indication that we might be observing N-nitro-pyridinium salt formation.

A series of experiments was then performed with 0.04M solutions of our reagents, which solutions were mixed together in differing proportions, with the proviso that the total volume of solution was always the same (10 ml.). The spectra of these solutions were taken and the combining ratio of nitronium fluoborate to pyridine could be calculated as 1:(1 - 1.3). Since there is n.m.r. evidence that an

excess of pyridine causes further reaction to occur, total success of this application of Job's method⁽⁷⁾ cannot be expected. However, the results do lend support to the proposal that a 1:1 complex of pyridine with nitronium fluoborate has been formed.

More concentrated solutions of our reagents were used for n.m.r. spectroscopy. All chemical shifts were measured relative to tetramethyl-silane as internal standard, using the side band (field modulation) technique with our A.E.I. "R.S.-2" 60 mc. high resolution instrument. Teflon caps were used to protect the contents of the n.m.r. sample tubes from moisture.

When 1.0M solutions of pyridine and nitronium fluoborate were mixed in equal proportions much heat was evolved and the reaction mixture had the n.m.r. spectrum shown below (Fig. 1). Identical spectra could be obtained by adding an equivalent, or half-equivalent amount of pyridine alone to a measured volume of 1.0M nitronium fluoborate solution, so the spectrum would appear to be concentration invariant. The solutions were very, very pale yellow.



Comparison of the spectrum with that of pyridinium fluoborate in sulfolane, or better, that of the pyridine. AlCl_3 complex in sulfolane (Fig. 2) indicates clearly that the spectrum of Fig. 1. is the result of attaching a strongly electron-withdrawing group to the nitrogen atom of the pyridine ring. Reflection will indicate that the spectrum in Fig. 1. cannot be generated by substitution at a carbon atom in the pyridine ring. Thus we are led again to the conclusion that N-nitropyridinium fluoborate has been formed. This conclusion is reinforced by the fact that the areas under the peaks labelled α , γ and β are in the ratio 2:1:2, corresponding to the number of protons in the α , γ and β positions on the pyridine ring. Furthermore, N-nitropyridinium fluoborate should be to pyridinium fluoborate as nitrobenzene is to benzene. Using this analogy and well known results⁽⁸⁾, one can calculate what the chemical shifts should be for the α , γ and β protons in N-nitropyridinium fluoborate, knowing the shifts in pyridinium fluoborate. The close agreement between this simple calculation and experiment is shown in Table I.

The solutions prepared for n.m.r. work show a very slow subsequent reaction to substances as yet unidentified. Any excess of pyridine over nitronium salt also produces substances as yet unidentified.

All the investigations so far recorded relate to a product in a freshly prepared solution. When dry chloroform was added to the solution which resulted from mixing equal portions of 1.0M reagents, a white salt-like material was precipitated. This solid was filtered in a dry box using a sintered glass funnel. Its solutions in sulfolane

TABLE I
Chemical Shifts (τ)

Experimental shifts in $\text{py.H}^+ \text{BF}_4^-$ (sulfolane)	α	γ	β
	1.08	1.23	1.80
Calculated shifts in $\text{py.NO}_2^+ \text{BF}_4^-$	0.11	0.81	1.50
Observed shifts in $\text{py.NO}_2^+ \text{BF}_4^-$	0.12	0.91	1.56

appeared from n.m.r. and u.v. spectra to contain N-nitropyridinium fluoborate and some impurity. Analysis would seem to show that it was mainly N-nitro-pyridinium fluoborate. (Calculated for $\text{py.NO}_2^+ \text{BF}_4^-$: C = 28.2%, H = 2.4%, N = 13.2%. Found :- C = 28.2%, H = 2.6%, N = 13.5%). An infra-red spectrum of the solid in mulls and KCl discs showed clearly that it was a fluoborate, though not nitronium fluoborate itself. Comparison with the i.r. spectrum of pyridinium fluoborate revealed extra bands at 650 cm^{-1} , 725 cm^{-1} , 810 cm^{-1} , 1000 cm^{-1} , 1200 cm^{-1} , 1290 cm^{-1} and 1720 cm^{-1} . Without using N^{15} reagents firm assignments cannot be made, but there is a case for assigning the band at 1720 cm^{-1} to $-\text{NO}_2$ asymmetric stretching, that at 1000 cm^{-1} to $\text{N}-\text{N}$ stretching, and that at 650 cm^{-1} to $\text{N}-\text{N}$ deformation. The bands around 800 cm^{-1} are characteristic of nitro compounds, too⁽⁹⁾.

Thus two independent lines of evidence have indicated that pyridine and nitronium fluoborate combine in a 1:1 mole ratio, and two, one conclusive, have indicated that the nitronium ion is bonded to the nitrogen atom in the pyridine ring. In view of the high electron density due to the lone pair on the nitrogen⁽¹⁰⁾ this is hardly surprising, and is in line with other electrophilic additions at this position⁽¹¹⁾.

These findings also explain why nitration proceeds through the pyridinium ion in nitric/sulphuric media. Presumably there is no free base here, but a mixture of conjugate acid and N-nitropyridinium ions, which should be even more de-activated than pyridinium ions.

We have also found u.v. and n.m.r. evidence for N-nitro-pyridinium ion formation with 2:6 lutidine, 2:6 dibromopyridine, 3:5 lutidine, γ -picoline, 4-amino pyridine and 4-cyanopyridine, but none whatsoever in the case of 2:6 dicyanopyridine.

It is gratifying that other workers appear to have reached some of our conclusions quite independently⁽¹²⁾.

We are grateful to the University of Sheffield for an I.C.I. Fellowship and a Junior Research Fellowship, to our colleagues and to Dr. J. A. Elvidge for helpful discussions, to Shell Chemicals for a gift of sulfolane, and to Dr. G. A. Olah for making available some results prior to publication.

References

- (1) M. W. Austin, M. Brickmann, J. H. Ridd & B. V. Smith, Chem. & Ind. p 1057 (1962).
- (2) R. B. Moodie, K. Scholfield & M. J. Williamson Chem. & Ind. p 1283 (1963).
- (3) R. W. Foster, Ph.D. Thesis, U. of London 1954.
- (4) D. Cook, S. J. Kuhn & G. A. Olah, J. Chem. Phys. 33 p 1669 (1960).
- (5) M. Kilpatrick, M. L. Kilpatrick & J. G. Jones, unpublished observations.
- (6) S. J. Kuhn and G. A. Olah, J. Amer. Chem. Soc. 83 p 4564 (1961).
- (7) cf. F. J. C. Rossotti and H. Rossotti, "The Determination of Stability Constants" pp. 46 - 56. McGraw Hill (1961).
- (8) L. M. Jackman, "Nuclear Magnetic Resonance" Table 4.14, p 63. Pergamon, (1959).
- (9) L. J. Bellamy, "Infra-Red Spectra of Complex Molecules", Chap. 17. John Wiley (1958).
- (10) R. D. Brown and M. L. Heffernan, Australian J. Chem. 12 p 554 (1959).
- (11) O. Hassel, Proc. Chem. Soc. p 250 (1957).
- (12) G. A. Olah, private communication.