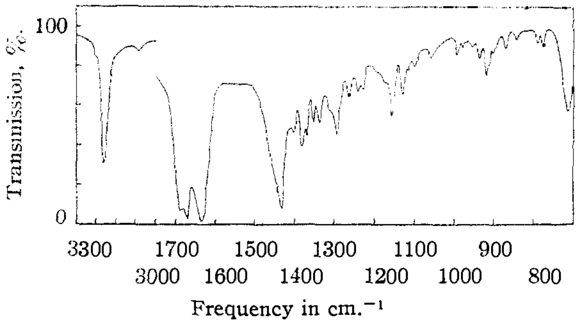

Fig. 1.

reacted with ninhydrin to give the deep purple color characteristic of α-amino acids. Paper strip chromatograms[6] showed the hydrolysis mixtures to contain leucine and proline. These data indicated the unknown to be one of the leucylproline anhydrides, and a search of the literature revealed it to be identical with the L,L-isomer described by Fischer and Reif.[2] Hydrolysis of this material with 33% aqueous sulfuric acid has been found to yield L-leucine and L-proline of good optical purity. Fischer and Reif reported hydrolysis with 20% aqueous hydrochloric acid to cause appreciable racemization of the amino acids.

Experimental

Isolation of L-Leucyl-L-proline Anhydride from the Culture Filtrates.—The culture fluid from a 100-gallon fermenter of the *Streptomyces* sp. (180 l.) was clarified by filtration. The filtrate was extracted with six 3.5-liter portions of chloroform, and the chloroform extract was concentrated *in vacuo* to about 750 ml. On standing overnight at 5°, a brown crystalline precipitate formed. It was collected by filtration and recrystallized from amyl acetate. The yield was 3.61 g. of white hexagonal plates. The sample for analysis was recrystallized from methanol; m.p. 158–161° after softening at about 145°; $[\alpha]^{21}\text{D}$ −142.4 ± 0.5° (*c*, 3.33 in ethanol). Fischer and Reif[2] reported $[\alpha]^{20}\text{D}$ −143.4°.

Anal. Calcd. for $C_{11}H_{18}O_2N_2$: C, 62.82; H, 8.62; N, 13.33; mol. wt., 210. Found: C, 62.91; H, 8.36; N, 13.48; mol. wt. (Rast), 235.

Acid Hydrolysis and Isolation of the Amino Acids.—A suspension of 3 g. of the fermentation metabolite in a mixture of 16 ml. of concd. sulfuric acid and 32 ml. of water was heated overnight on a steam-bath, then under reflux in an oil-bath for five hours. The reaction mixture was diluted to 250 ml., and solid barium hydroxide octahydrate was added to the hot solution until it was only weakly acidic. The precipitated barium sulfate was removed by filtration, and the filtrate was evaporated to dryness yielding 3.82 g. of a white crystalline residue. The residue was extracted with four 25-ml. portions of hot absolute alcohol.

L-Proline.—The alcohol extract was concentrated *in vacuo* to 10 ml. and 60 ml. of ether was added to precipitate the L-proline. The precipitate was collected by filtration, and the precipitation from absolute alcohol was twice repeated. The yield of L-proline was 1.11 g., or 70% of the theoretical amount; $[\alpha]^{22}\text{D}$ −77.7 ± 0.5° (*c*, 4.36 in water). Fischer and Zemplén[7] reported $[\alpha]^{20}\text{D}$ −79.8°.

Anal. Calcd. for $C_5H_9NO_2$: C, 52.16; H, 7.88. Found: C, 51.97; H, 7.74.

L-Leucine.—The alcohol-insoluble residue was recrystallized three times from water yielding 1.47 g. of L-leucine, 78% of the theoretical amount; $[\alpha]^{22}\text{D}$ +16.83 ± 0.4° (*c*, 5.02 in 20% HCl). Fischer[8] reported $[\alpha]^{20}\text{D}$ +17.5°.

Anal. Calcd. for $C_6H_{13}NO_2$: C, 54.94; H, 9.99. Found: C, 54.98; H, 9.80.

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[6] Run according to the procedure of R. J. Williams and H. Kirby, *Science*, **107**, 481 (1948).

[7] E. Fischer and G. Zemplén, *Ber.*, **42**, 2989 (1909).

[8] E. Fischer, *ibid.*, **34**, 433 (1901).

Oxidation of Nitric Oxide at High Pressures of Reactants

BY HAROLD S. JOHNSTON AND LOREN W. SLENTZ

The reaction between nitric oxide and oxygen was studied in considerable detail at pressures of reactants up to one-half atmosphere at the one temperature, 25°. The method and apparatus are identical with the intermediate pressure equipment of Mills and Johnston.[1] Three series of runs were made: (1) with equal pressures of reactants, (2) with nitric oxide in tenfold excess and oxygen diluted with 9 parts of nitrogen, and (3) with oxygen in tenfold excess and nitric oxide diluted with 9 parts of nitrogen. At high pressures of reactants the rate was followed on the oscilloscope since half-times were of the order of magnitude of one-tenth second. At lower pressures of reactants, the rate was followed visually on an electronic voltmeter. The third-order rate law was observed throughout this range of conditions, and no trend in the values of the rate constants was observed in going from high to low pressures. The results are summarized in Table I.

TABLE I

THIRD ORDER RATE CONSTANTS FOR THE REACTION BETWEEN NITRIC OXIDE AND OXYGEN AT 25°

Ratio $(NO)/(O_2)$	Range of initial pressures of NO, mm.	Number of runs	Average rate constant, sec.$^{-1}$ mm.$^{-2}$ × 10^5	Standard error of mean, sec.$^{-1}$ mm.$^{-2}$ × 10^5
10	10.6 to 340	28	1.98	0.06
1	18.4 to 130	15	2.11	.05
0.1	8.1 to 43.0	20	2.02	.05
	Average of all	63	2.02	.03

The average rate constant is 7.10×10^9 cc.^{2} mole^{-2} sec.$^{-1}$, and the standard error of the mean is 0.12×10^9. This value agrees exactly with Bodenstein's[2] 7.06×10^9 cc.^{2} mole^{-2} sec.$^{-1}$ at 30°, and it is slightly, though perhaps not significantly, higher than Smith's[3] 6.00×10^9 cc.^{2} mole^{-2} sec.$^{-1}$ at 25°.

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(1) Robert L. Mills and Harold S. Johnston, THIS JOURNAL, **73**, 938 (1951).

(2) M. Bodenstein, *Z. Elektrochem.*, **24**, 183 (1918); *Z. physik. Chem.*, **100**, 87 (1922).

(3) J. H. Smith, THIS JOURNAL, **65**, 74 (1943).

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Thermal Decomposition of Nitrogen Pentoxide at High Temperature

BY HAROLD S. JOHNSTON AND YU-SHENG TAO

The thermal decomposition of nitrogen pentoxide is a very extensively investigated reaction. The earliest kinetic examination by Daniels and Johnston[1] was a careful study over the temperature range 25 to 65°, and over this range a good estimate was made of the energy of activation. The primary object of this extension of the temperature range was to see how and if the energy of activation itself is a function of temperature.

Experimental

The apparatus used in this study was very nearly the same as that described by Johnston and Yost.[2] It consisted of a calibrated flowmeter system, U-shaped nitrogen pentoxide saturator, Hartridge and Roughton type of mixing chamber, stainless-steel stop gate, 30-cm. coil of glass tubing through heating bath to bring gases up to desired temperature, reaction cell 10 cm. long and 6 mm. internal diameter, bath of paraffin oil stirred and thermally regulated to ±0.3°, H-4 mercury arc filtered to give line at 436 mμ, electron multiplying photoelectric tube RCA 1P28, DuMont 208 oscilloscope, and 35 mm. camera. By means of thermocouples it was established that the temperature inside the reaction cell was identical with that of the bath at the flow rates used. The bath temperature was measured by a thermometer calibrated by the Bureau of Standards. Nitrogen pentoxide was prepared by the method and apparatus described by Mills and Johnston.[3] All studies were made with approximately 2 to 5 mm. of nitrogen pentoxide and at a total pressure of one atmosphere with nitrogen making up the difference. Reaction was followed colorimetrically by the appearance of nitrogen dioxide. The absorption coefficient expressed in units of length and concentration (not pressure) was constant from 25 to 125° and agreed with previously found and published values.[4]

Results

The reaction is

$$2N_2O_5 = 4NO_2 + O_2$$

and the first-order rate constant is defined as k in the expression

$$-(^1/_2)\ d(N_2O_5)/dt = k(N_2O_5)$$

At the concentrations of products encountered and in the temperature range 65 to 123° the correction for nitrogen tetroxide, N_2O_4, was negligible, greatly simplifying the computation of rate constants. By letting the reaction run for 20 or more half-lives, the initial concentration of nitrogen pentoxide could be computed as one-half the final concentration of nitrogen dioxide. The concentration of nitrogen pentoxide at any time was considered to be this initial value minus one-half the concentration of nitrogen dioxide. The plot of the logarithm of nitrogen pentoxide concentration against time gave a straight line (at least for the first half-life) whose slope times 2.303/2 gave the first-order rate constant. A summary of values of rate constants is given in Table I.

TABLE I

FIRST-ORDER RATE CONSTANTS FOR THE DECOMPOSITION OF NITROGEN PENTOXIDE AT HIGH TEMPERATURES

Temperature, °K.	Number of runs	Av. rate constant sec.⁻¹	Standard error of the mean sec.⁻¹
337.6	10	0.00256	0.00007
348.6	11	.00683	.00013
357.6	15	.0161	.0002
367.6	10	.0458	.0013
378.0	9	.122	.009
388.2	7	.258	.004
396.2	13	.525	.009

Discussion

The energy of activation is defined as E in the relation

$$d \ln k/dT = E/RT^2 \qquad (1)$$

By first assuming the energy of activation to be independent of temperature and by combining these data with those of Daniels and Johnston,[1] the energy of activation was computed by the method of least squares to give

$$k = 2.05 \times 10^{13}\ e^{-24650/RT}\ \text{sec.}^{-1} \qquad (2)$$

These results are in almost exact agreement with those obtained in the low temperature range alone.[1] To test the temperature dependence of the energy of activation itself, the same data were fitted to the usual three-parameter relation to give

$$k = 1.97 \times 10^{18}\ T^{-1.68}\ e^{-25790/RT}\ \text{sec.}^{-1} \qquad (3)$$

It has been said[5] with reference to the low temperature data alone that the results could equally well be expressed by Eq. (2) or by expressions like Eq. (3) wherein the absolute temperature is given a negative exponential value up to 14. These results cover a wide enough range of temperature to make that statement no longer true. The least-squares exponent of the absolute temperature in Eq. (2) is very nearly zero, and the energy of activation of this reaction is essentially constant from 25 to 123°.

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[1] F. Daniels and E. H. Johnston, THIS JOURNAL, **43**, 53 (1921).

[2] H. S. Johnston and D. M. Yost, *J. Chem. Phys.*, **17**, 386 (1949).

[3] R. L. Mills and H. S. Johnston, THIS JOURNAL, **73**, 938 (1951).

[4] H. H. Holmes and F. Daniels, *ibid.*, **56**, 630 (1934).

[5] R. Fowler and E. A. Guggenheim, "Statistical Thermodynamics," University Press, Cambridge, 1949, p. 526.

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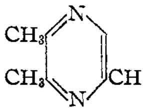
The Action of Organolithium Compounds on 2,5-Dimethylpyrazine. II[1]

BY BERNARD KLEIN[2] AND PAUL E. SPOERRI

In an earlier communication,[3] it was reported that treatment of 2,5-dimethylpyrazine with methyllithium produced 2,5,6-trimethylpyrazine (I)



I

Similarly, with phenyllithium, the corresponding 2,5-dimethyl-6-phenylpyrazine was formed.

We wish to present additional evidence that this reaction is a general one for the direct alkylation in the pyrazine series. Thus, when 2,5-dimethylpyrazine was treated, in turn, with one equivalent of a series of alkyllithium compounds, the corresponding 2,5-dimethyl-6-alkylpyrazine was obtained (II), presumably by the 1,2 addition across the azomethine linkage. Evidently, the high electron

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[1] From the Ph.D. thesis of Bernard Klein, Polytechnic Institute of Brooklyn, May, 1950. A major portion of the work was done in the laboratory of the Veterans Administration Hospital, Bronx, N. Y. Reviewed in the Veterans Administration and published with the approval of the Chief Medical Director. The statements and conclusions published by the authors are the result of their own study and do not necessarily reflect the opinion or policy of the Veterans Administration.

[2] Veterans Administration Hospital, Bronx 63, N. Y.

[3] B. Klein and P. E. Spoerri, THIS JOURNAL, **72**, 1844 (1950).