boiled with methyl alcohol. It crystallizes in colorless plates. Like the open chained esters it was readily reduced catalytically but the reduction product showed little tendency to crystallize and we were able to isolate only the saturated ketonic acid.

Anal. Calcd. for $C_{17}H_{13}O_3Cl$: C, 67.9; H, 4.4. Found: C, 67.8; H, 4.4.

Relative Stability and Isomerization.—The experiments by which the relative stability of the geometrical isomers was ascertained with sufficient accuracy for the present purpose have been described adequately in the introduction. The conditions that promote isomerization have not been studied with sufficient care for discussion at this time, but it is certain that in the case of the esters isomerization invariably leads to an equilibrium in which the cis ester preponderates. This has been found true when the isomerization was effected by increase of temperature, by exposure to sunlight, by treatment with a weak base like sodium acetate in glacial acetic acid, and by treatment with hydrochloric acid in glacial acetic acid.

Summary

This paper contains: a description and a comparison of the methods that are available for determining the configuration of unsaturated ketonic acids and their esters; a proof that each of two diastereomeric bromo esters loses hydrogen bromide in but one way and forms but one unsaturated ester; and the reason why a mixture of unsaturated esters is generally obtained, nevertheless, from each of the bromo esters.

CAMBRIDGE, MASS. RECEIVED AUGUST 8, 1934

NOTE

Note on a New Technique for the Preparation of Amino Nitriles[1]

BY G. A. MENGE

At the time, now rather remote, that this investigation was initiated, repeated attempts by the writer to duplicate the preparation of aminoacetonitrile, and other similar compounds, as described by Klages,[2] yielded results which, while not complete failures, were highly unsatisfactory, both with reference to the preliminary preparation of the cyanhydrin and to its conversion into the aminonitrile. Finally, the use of anhydrous hydrogen cyanide, freshly prepared by the method of Wade and Panting,[3] in connection with Ultee's[4] method for the preparation of the cyanhydrin gave excellent results. For the conversion of the cyanhydrin into the corresponding aminonitrile a new, and remarkably successful, technique was finally developed, involving the substitution of liquid ammonia,[5] to serve both as reagent and solvent, for the alcoholic or aqueous ammonia generally used. Application of the new technique is indicated in the following brief description of the preparation of aminoacetonitrile.

(1) Developed by the writer several years ago while a member of the Division of Pharmacology, Dr. Reid Hunt, Chief, of the then Hygienic Laboratory, U. S. P. H. S. Publication delayed in the vain hope that opportunity would develop for refining and extending the investigation.

(2) Klages, *J. prakt. Chem.*, [2] **65**, 189 (1902).

(3) Wade and Panting, *J. Chem. Soc.*, **73**, 256 (1898).

(4) Ultee, *Rec. trav. chim.*, **28**, 1, 248, 260 (1909).

(5) Grateful acknowledgment is due Dr. E. C. Franklin, for his kindly assistance and instruction in the proper and safe use of liquid ammonia.

To about 5 g. of glycolic nitrile, contained in a suitable Carius tube and cooled to the temperature of liquid ammonia, a large excess (about 20 cc.) of liquid ammonia was added. Keeping the reagents cooled, by immersion in liquid ammonia, the tube was carefully sealed and after the seal had cooled the contents were gently mixed (with the tube wrapped in a towel), forming a clear, colorless solution. After standing at room temperature for twenty-four hours a slight tinge of yellow color had developed in the clear solution. The pressure in the tube was then released by fusing the capillary seal and the ammonia allowed to evaporate spontaneously until boiling had ceased. The residue, transferred to a flask, was treated with successive portions of absolute alcohol, each portion being evaporated off under vacuum, until all free ammonia was removed, leaving a final product of very clear pale yellow liquid—presumably aminoacetonitrile. No attempt was made to distil it. The hydrogen chloride salt was prepared by dissolving the free base in dry ether and adding this solution carefully to hydrogen chloride dissolved in a 1:1 mixture of absolute alcohol and dry ether. Promptly, there resulted an abundant separation, with evolution of considerable heat, of a clean, white, crystalline solid which proved to be the pure hydrochloride salt, easily recrystallized from absolute alcohol.

Melting point: darkens gradually above 135° and melts slowly to a viscous brown mass at 165.5–166.5°; behavior identical with the pure salt obtained from the free base prepared by other methods and in close agreement with the m. p. of 165° reported by Klages.

Anal. The Pt salt was not easily prepared but was finally obtained by mixing alcoholic solutions of the free base and H_2PtCl_6 and adding dry ether to the mixture. On igniting for Pt: Calcd. for $C_4H_{10}N_4Cl_6Pt$: Pt, 37.35. Found: Pt, 37.20, 37.30.

Yield.—The entire product from the pressure-tube reaction was carefully converted into the hydrochloride salt, as described above, and the yield measured in that form. The total pure dry salt recovered weighed 7.82 g., or close

to 95% of the theoretical yield (about 8.25 g.). It follows that the reaction between glycolic nitrile and liquid ammonia, under the conditions noted, is smooth, clean, and, allowing for unavoidable loss through indirect measure of yield, practically quantitative.

The same method, with modified technique in individual cases, was successfully applied to the preparation of other known aminonitriles, desired for pharmacological investigation. Further study of the method and its application was interrupted by the author's change in affiliation and field of activity and has not been resumed because of lack of opportunity.

LAFAYETTE COLLEGE
EASTON, PENNA.

RECEIVED JULY 23, 1934

COMMUNICATIONS TO THE EDITOR

OXIDATION OF CYSTINE WITH PERMONOSULFURIC ACID

Sir:

Sulfur oxides of cystine [Toennies and Lavine, *J. Biol. Chem.*, **100,** 463 (1933)] should be obtainable in presence of water if oxygenation is rapid compared with speed of hydrolysis of the —S—S— bond. As cystine disulfoxide [Lavine, Toennies and Wagner, THIS JOURNAL, **56,** 242 (1934)] proved to be relatively stable toward aqueous acids the action on cystine of a potent oxygen donor such as permonosulfuric acid (prepared from $K_2S_2O_8$ and H_2SO_4 [Gleu, *Z. anorg. allgem. Chem.*, **195,** 70 (1931)]) seemed interesting. The following is a preliminary report on this investigation. In presence of 10.7 mole equivalents of H_2SO_5 oxygen consumption stopped after two hours at 5.0 atoms, the theoretical amount for oxidation to cysteic acid, suggesting that absence of free electron pairs in the acid form ($-NH_3^+$) protects the amino group against oxygen addition. The speed of oxygen consumption with different H_2SO_5 : cystine ratios (cystine 0.025 *M*, H_2SO_4 3.5 *M*, 0°) was as follows:

		Atoms O consumed per mol. of cystine, after min.					
	H_2SO_5:cystine	2	9	13	20	30	70
I	1.01:1	0.75	1.00	1.01			
II	2.10:1	1.2	1.7	1.9	2.0	2.10	
III	5.37:1	2.6	3.8	4.0	4.2	4.3	4.5

After oxidation I the direct cyanide–nitroprusside test indicated 15% of unchanged cystine while by a preceding reduction with iodide [Toennies and Lavine, *J. Biol. Chem.*, **105,** 119 (1934)] the test was increased to 93% of the cystine used. After oxidation II the corresponding figures were 5 and 95%. Thus the chief reaction products seem to be the monosulfoxide in oxidation I and the disulfoxide in oxidation II.

By diluting the H_2SO_5 solution (Ref. 3) with methanol the potassium sulfate present was nearly completely precipitated and the filtrate (0.1 *M* H_2SO_5, 0.9 *M* H_2SO_4, 75% CH_3OH), in which H_2SO_5 is as stable as in aqueous solution, was used to oxidize cystine, dissolved as perchlorate in CH_3CN (Ref. 1), with 1 mole equivalent of H_2SO_5. Precipitates obtained—after oxidation—by neutralization with pyridine, contained at least twice as much cystine as was indicated by direct test on the oxidized solution. Further evidence of a dismutative change of the primary oxidation product was obtained by fractionated neutralization of the precipitate, inasmuch as it resulted in further formation of cystine together with a decrease of total precipitate, and by polarimetric observation. On oxidation of cystine with 1 mole of H_2SO_5 the initial high negative rotation decreases in less than one hour to a slightly positive value, only to slowly turn negative again during two to three days, passing through a maximum of about one-third of the initial value and slowly decreasing again during the next two weeks. This last decrease presumably represents esterification (Ref. 1) by methanol of the cystine formed by dismutation during the second stage of change. Addition of ethanol and ether immediately after the first brief reaction stage yielded a white precipitate which, according to analysis and properties, consists of sulfates of cystine (10%) and its monosulfoxide (90%) and which in *N* sulfuric acid gives an initial specific rotation of about +5° for the total content of unoxidized and oxidized cystine. This value however, changes during the next five days to a constant negative level which would correspond to an amount of free cystine equivalent to 70% of the organic sulfur present.

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