doublet, $J \sim 9$ cps, shift varies with concentration, disappears on deuteration, NH), 4.75–5.20, 5.50–5.85, 6.28–6.60 (multiplets, six protons, H-2,3,4,5,6,6′), 7.14 (one proton, broad, shift varies with concentration, disappears on deuteration, OH), 7.90, 8.00, 8.14 (three-, six-, three-proton singlets, acetyls).

The product was homogeneous by tlc. A solution of the product (0.521 g) in dry pyridine (2 ml) containing chlorotriphenylmethane (0.418 g, 1 equiv) was shaken for 4 days at room temperature, and the solution was poured into ice and water (50 ml). The product that precipitated was washed with water and dried to yield 810 mg (92%), mp 155–156°, identical with authentic 2-acetamido-1,3,4-tri-*O*-acetyl-2-deoxy-6-*O*-trityl-α-D-glucopyranose[2] by mixture melting point, tlc, comparative infrared and nmr spectra, and by X-ray powder diffraction pattern.

2-Acetamido-1,3,4-tri-*O*-acetyl-2-deoxy-6-*O*-trideuterioacetyl-α-D-glucopyranose (3).—To 2-acetamido-1,3,4-tri-*O*-acetyl-2-deoxy-α-D-glucopyranose (1.84 g) were added, in rapid succession, dry pyridine (2 ml) and acetic anhydride-d_6, and the resultant solution was kept for 18 hr at room temperature. The mixture was poured into ice–water (20 ml) and the product was extracted with four 50-ml portions of chloroform. The extract was evaporated and the resultant syrup was codistilled with toluene and then with carbon tetrachloride. Crystallization of the syrup from chloroform–ether gave **3**, yield 1.6 g (76%), mp 138–139°. The product gave an X-ray powder diffraction pattern identical with that of 2-acetamido-1,3,4,6-tetra-*O*-acetyl-2-deoxy-α-D-glucopyranose (**1**); its nmr spectrum (chloroform-*d*) was identical with that of **1** except that the three-proton singlet (τ 7.93) present in the spectrum of **1** was absent in the spectrum of **3**.

Repetition of the experiment with nondeuterated acetic anhydride gave **1** in similar yield, identical with an authentic sample.

1,3,4,6-Tetra-*O*-acetyl-2-amino-2-deoxy-α-D-glucopyranose Hydrochloride (9).—A solution of 2-acetamido-3,4,6-tri-*O*-acetyl-2-deoxy-α-D-glucopyranosyl chloride[15] (**7**, 25 g) in acetone (500 ml) containing water (1.25 ml, 1 molar equiv) was refluxed gently for 38 hr. The resulting suspension was refrigerated for 1 hr, and the product was filtered and washed with two 50-ml portions of acetone: yield 19–20 g (73–75%); mp 181°; $[\alpha]^{23}D$ +144° (*c* 4.7, water) [lit.[14] mp 180°, $[\alpha]^{20}D$ +130° (*c* 1, water)]; λ_{max}^{KBr} ~3.5, 6.24, 6.35 (NH_3^+), 5.68, 5.74 μ (OAc); nmr (deuterium oxide) τ 3.60 (one-proton doublet, $J_{1,2}$ = 3.7 cps, H-1), 4.42 (one-proton triplet, $J_{2,3}$ = 10.2 cps, H-3), 4.88 (one-proton broadened triplet, $J_{3,4}$ = 9.1 cps, H-4), ~5.75 (three-proton multiplet, H-5,6,6′), 6.02 (one-proton quartet, H-2), 7.75, 7.86, 7.92 (three-, three-, and six-proton singlets, acetyl groups); X-ray powder diffraction 10.34 m, 8.89 s (2), 7.67 w, 6.27 vw, 5.56 m, 5.29 m, 4.93 m, 4.75 vw, 4.51 s (3), 3.90 vs (1), 3.74 w, 3.58 w, 3.38 w, 3.27 m.

Under the same conditions, 3,4,6-tri-*O*-acetyl-2-deoxy-2-trideuterioacetamido-α-D-glucopyranosyl chloride[2] (**8**) was converted into the corresponding 1-*O*-trideuterioacetyl analog (**10**) of **9**. The nmr spectra of **9** and **10** were identical, except that the three-proton singlet observed at τ 7.75 in the spectrum of **9** (deuteriumoxide) was absent in that of **10**.

2-Acetamido-3,4,6-tri-*O*-acetyl-2-deoxy-1-*O*-trideuterioacetyl-α-D-glucopyranose (2).—Substance **10** was acetylated with acetic anhydride in pyridine by the procedure[2] used for preparation of 1,3,4,6-tetra-*O*-acetyl-2-deoxy-2-trideuterioacetamido-α-D-glucopyranose and substance **3**. The nmr spectrum of the product was identical with that of 2-acetamido-1,3,4,6-tetra-*O*-acetyl-2-deoxy-α-D-glucopyranose (**1**), except that the three-proton singlet at τ 7.81, present in the spectrum of **1** (chloroform-*d*) was absent in the spectrum of **2**.

Repetition of the experiment, with acetic anhydride-d_6, gave 3,4,6-tri-*O*-acetyl-2-deoxy-2-trideuterioacetamido-1-*O*-trideuterioacetyl-α-D-glucopyranose (**4**). The nmr spectrum of **4** was identical with that of **2**, except that the three-proton singlet at τ 8.09 in the spectrum of **2** (chloroform-*d*) was absent in the spectrum of **4**.

1,3,4,6-Tetra-*O*-acetyl-2-amino-2-deoxy-β-D-glucopyranose Hydrochloride (6).—This product, prepared by the procedure of Bergmann and Zervas,[16] darkened at 210°; $[\alpha]^{21}D$ +29.5 ± 0.5° (*c* 2.8, water) (lit.[16] darkens at 230°, $[\alpha]^{21}D$ +29.7° in water); λ_{max}^{KBr} ~3.6, 6.30 (NH_3^+), 5.70, 5.75 μ (OAc); nmr (deuterium oxide) τ 4.00 (one-proton doublet, $J_{1,2}$ = 8.7 cps, H-1), 4.48 (one-proton triplet, $J_{2,3}$ = 10.0 cps, H-3), 4.88 (one-proton multiplet, $J_{3,4}$ = 9.0 cps, H-4), ~5.70 (three-proton multiplet, H-5,6,6′), 6.23 (one-proton triplet, H-2), 7.77, 7.86, 7.90 (three-, three-, and 6-proton singlets, acetyl groups); X-ray powder diffraction 12.90 w, 8.98 vs (1), 6.01 vw, 5.35 s (2), 4.62 m, 4.24 m, 3.98 vw, 3.76 s (3), 3.57 w, 3.37 w, 3.19 w, 3.02 w, 2.88 w.

Registry No.—2-Acetamido-1,3,4-tri-*O*-acetyl-2-deoxy-α-D-glucopyranose, 10034-17-0; **3**, 10034-18-1; **9**, 10034-19-2; **6**, 10034-20-5; **1**, 7784-54-5; **5**, 7772-79-4.

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Electrochemical Oxidation of Primary Aliphatic Amines

KAREN K. BARNES AND CHARLES K. MANN

Department of Chemistry, Florida State University, Tallahassee, Florida 32306

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Ammonia and a series of primary aliphatic amines were oxidized at a platinum electrode in acetonitrile. A dual mechanism is proposed, one sequence of reactions predominating at low potentials, the other at high potentials. The low-potential mechanism involves a one-electron transfer to form a cation radical which decomposes for the most part to a carbonium ion and amidogen radical. At higher potentials, two electrons and a proton are lost to form the iminium salt, which can hydrolyze to aldehyde and ammonia.

We have undertaken an examination of the electrochemical oxidation of primary aliphatic amines as part of a larger study of anodic reactions of nitrogen compounds in which we hope to examine similarities and differences between electrochemical, chemical, and biological oxidations. Previous reports on this work involved anodic oxidation of triethylamine in diemethyl sulfoxide,[1] and oxidation of primary, secondary, and tertiary aliphatic amides.[2] In a survey of aliphatic amine oxidations by cyclic voltammetry,[3] it was noted that while substituent inductive effects could be correlated with voltammetric peak potentials for secondary and tertiary aliphatic amines, this did not apply to primary amines. Further, the voltammetry peaks for primary amines showed characteristically small slopes with the peaks drawn out over an appreciable potential range, as compared with secondary and tertiary amines. No other references to anodic oxidation of primary aliphatic amines appear, although a coulometric titration of amines was described under conditions which would seem, from this work, to involve direct amine oxidation.[4]

(1) R. F. Dapo and C. K. Mann, *Anal. Chem.*, **35**, 677 (1963).

(2) J. F. O'Donnell and C. K. Mann, *J. Electroanal. Chem.*, **13**, 157 (1967).

(3) C. K. Mann, *Anal. Chem.*, **36**, 2424 (1964).

(4) C. A. Streuli, *ibid.*, **28**, 130 (1956).

Results

Product Analyses.—Results are collected in Table I. All oxidations were carried out at 1.2 v *vs.* the silver–silver ion electrode except in the cases of benzylamine (1.5 v) allylamine (2.0 v), cyclohexylamine (1.4 v), and hexylamine (1.4 v).

TABLE I
ELECTROLYSIS PRODUCTS

Amine reacted	Electrolysis potential (*vs.* Ag/Ag^+), v	*n* value	% starting amine reacting by path A	Amine salt recovered, %[a] (calcd, %)
n-Butylamine	1.2	0.75	90	59 (61)
Hexylamine	1.4	0.80	90	66 (61)
Isobutylamine	1.2	0.75	99	65 (60)
Benzylamine	1.5	0.78	Trace	80 (79)
Ammonia	1.2	0.75	...	...
n-Propylamine	1.2	0.73	90	52 (61)
Cyclohexylamine	1.4	0.71	90	66 (61)
Allylamine	2.0	0.79	50	70 (66)
Methylamine	1.2	0.72	90	61 (61)
t-Butylamine	1.2	0.71	100	56 (57)

NH_4 salt, %[a] (calcd, %)	% NH_3 oxidized	% aldehyde (calcd, %)	% N_2[a] (calcd, %)	Hydrocarbons found
11 (16)	25	3 (3)	+[b] (+)	C_1–C_4
19 (16)	25	3 (3)	25 (23)	C_1–C_6, ethylene predominates
14 (15)	50	−[b] (trace)	20 (25)	C_1–C_4, isobutylene and propylene predominate
5 (5)	75	15 (18)	+ (+)	None found
76 (75)	...	...	24 (25)	...
14 (16)	25	4 (3)	22 (23)	C_1–C_3, methane predominates
15 (16)	25	3 (3)	17 (23)	Trace cyclohexene
18 (17)	33	14 (17)	+ (+)	C_1–C_3
22 (16)	25	− (3)	25 (23)	Methane
25 (25)	17	4[c] (2)	+ (+)	C_1–C_4, isobutylene predominates

[a] Mole per cent based on initial amount of amine. [b] + = present in products; − = absent in products. [c] Alcohol.

At least half of the starting amine was recovered in the unreactive protonated form. Varying amounts of ammonium ion, aldehyde, and nitrogen were recovered. These yields, shown in Table I, indicate that a substantial fraction of the starting amine nitrogen has been accounted for. In addition, for all amines except cyclohexylamine and benzylamine, small amounts of hydrocarbons were identified. In general, the hydrocarbons ranged from parent olefin to methane, with distribution depending upon the structure of the starting material.

Controlled-Potential Coulometry.—Electron counts for the amines studied varied from 0.71 to 0.87 electron per molecule of amine. Since currents decayed to approximately 40 μa, in contrast to values in the range of 20–50 ma in the early parts of experiments, and since no starting material remained, this indicates that starting material reacts with an intermediate product to form an unreactive compound. In controlled-potential electrolysis without complications, the current decays exponentially with time as in a first-order reaction. If however, there are subsequent electron-transfer steps, then exponential decay may not be observed. The current–time curve may have more than one exponential segment, may have a sigmoid shape, or may pass through a minimum, rise, and then decay. This last behavior would result from delayed production of electroreactive material in a complex sequence of reaction steps. A typical current–time curve is illustrated in Figure 1. This is representative of those obtained in the low-potential electrolyses of *n*-butylamine, *n*-hexylamine, isobutylamine, cyclohexylamine, *n*-propylamine, methylamine, amylamine, and *n*-nonylamine. Benzylamine and allylamine showed a similar type of curve, but with very much smaller peaks. Ammonia and *t*-butylamine showed no peaks, but instead showed continuous current decay to background. A plot of log (current) *vs.* time in these cases did not yield a straight line, but a curve with two linear segments. *n*-Propylamine, which gave a curve typified by Figure 1 at 1.2 v, showed only a continuous decay to background at 1.9 v.

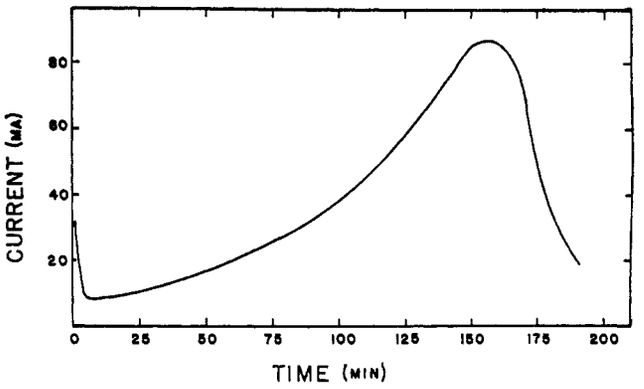


Figure 1.

Care was taken to ensure that the potentiostat was in control throughout. To assure that resistance drop between anode and reference probe was not responsible for the unusual curve shapes, amine concentrations were varied from 0.002 to 0.08 *M* in order to change the current, hence the resistance drop error. No change in curve shape was noted. The fact that *t*-butylamine and ammonia, as well as many other compounds not reported in this work, gave continuously dropping curves with the equipment used here indicates that the unusual curve shapes are not generated by the potentiostat.

Reactions of Intermediates.—The anodic reactions of possible intermediates and products were examined by cyclic voltammetry. Peak potentials are quoted relative to the silver–silver ion reference electrode. Simple olefins, 1-butene, and 1-hexene, are very difficult to oxidize. Their reactions were observable only as shoulders on the limiting current in sodium tetrafluoroborate solutions in acetonitrile at 3.4 to 3.7 v. Other compounds oxidized with $NaBF_4$ supporting electrolyte are cyclohexene (2.16 v) and isobutene (2.8 v). The following were oxidized in $NaClO_4$–acetonitrile solutions: bibenzyl (2.15 v), stilbene (1.36 v), N-propylpropanaldimine (1.8 v), and 2-methyl-2-pentenalpropylamine (1.8 v).

The reaction of ammonia was examined using controlled-potential coulometry, cyclic voltammetry, and product analysis. The reaction produces ammonium ion, nitrogen gas, and 0.75 electron per molecule as shown in Table I. The cyclic voltammetry peak at

1.02 v was very similar in shape to those of primary amines.

Amines having hydrogen on the carbon α to nitrogen generally produce aldehydes having the same number of carbons as the amine. *t*-Butylamine, lacking these hydrogens, formed *t*-butyl alcohol instead. Propylamine, unlike the others examined, did not produce aldehyde in the free form. Instead, 2-methyl-2-pentenalpropylamine (MPP) was formed. Cyclic voltammetry showed that this compound could be expected to oxidize only very slowly at 1.2 v, the value used for low-potential oxidation of propylamine. Consequently, it would be a possible product in low-potential runs, but not in high-potential runs. The presumed precursor, N-propylpropanaldimine (NPP), had the same voltammetry peak potential as MPP. However, the curve for NPP was very much more drawn out than that for MPP, making it seem unlikely that NPP is a possible product of either low or high potential runs.

When the course of an electrolysis is followed by running ultraviolet absorption spectra of successive aliquots, one finds, in addition to any peaks owing to amine, *e.g.*, benzylamine, or to expected products, *e.g.*, MPP, a peak at 256 mμ. This peak was observed on oxidation of *n*-propylamine, *n*-hexylamine, isobutylamine, cyclohexylamine, and *n*-butylamine but not of *t*-butylamine or methylamine. It is significant that this peak was not formed continuously from the start of the electrolysis, but only appeared after a period of time to coincide with the rising portion of the peak in the current–time curve. Hammond reported formation of a peak at this wavelength in the spectra of acetonitrile solutions. It was attributed to attack by a carbonium ion on the nitrogen of the nitrile, followed by condensation of the resulting ion to a polymer. Reactions of electrochemically generated carbonium ions with acetonitrile have been reported by Eberson and Nyberg.[5]

Discussion

On the basis of the electrochemical evidence, from cyclic voltammetry and controlled-potential electrolysis, a generalized mechanism will be proposed. Data for individual amines, including product analyses, will then be considered in light of the proposed mechanism.

Voltammetry curves for primary, secondary, and tertiary amines have been published previously.[3] Primary amines produce curves which differ from those of secondary and tertiary amines in that they are more drawn out and do not show changes in peak potentials which reflect changes in polar substituent effects. With secondary or tertiary amines, as small an alteration as the change from methyl to ethyl causes the oxidation to take place at a noticeably lower potential. With primary amines a change from methyl to *t*-butyl has no effect. It seems unreasonable to attribute this insensitivity simply to change from two to one alkyl and one to two hydrogens attached to nitrogen when no such change is observed when tertiary and secondary amines are compared.

(5) G. S. Hammond, M. F. Hawthorne, J. H. Waters, and B. M. Graybill, *J. Am. Chem. Soc.*, **82**, 704 (1960); L. Eberson and K. Nyberg, *Acta Chem. Scand.*, **18**, 1567 (1964); *Tetrahedron Letters*, 2389 (1966).

Voltammetry curve slopes reflect rates of electron transfer when they result from a single, irreversible step. Thus, assuming a single step, all tertiary and secondary amines examined show rates which are similar to each other. Primary amines, on this basis, show drastically lower rates. The initial step is, in all cases, the removal of an electron from the nitrogen. This is indicated by the completely unreactive nature of the remainder of the molecule in the case of saturated alkyl groups or the much higher reaction potential for the unsaturated substituents. Accordingly, it is difficult to understand why primary amines differ so completely from secondary and tertiary amines unless the curves for primary amines result from superposition of two or more electron-transfer steps, separated by a chemical step as indicated in eq 1.

$$\mathrm{A} \xrightarrow[1]{-e} \mathrm{B} \xrightarrow[2]{} \mathrm{C} \xrightarrow[3]{-e} \mathrm{D} \qquad (1)$$

If steps 1 and 3 are electron transfers occurring about the same potential, intervention of rate-controlling chemical step 2 could cause the observed drawing out of the curve. The experiment involves sweeping the potential over the range of interest in about 0.2 sec. Sequential electron-transfer reactions occurring at the same potential would simply show one peak. Sequential electron transfers at different potentials generally show two peaks, providing the second reaction occurs at the higher potential. Two electron-transfer steps, at the same or slightly different potentials, separated by a rate-controlling chemical or electrochemical step would show a drawn-out curve. The intervening step must be slow enough to affect the overall kinetics but fast enough to produce significant concentrations of C within 0.1 sec.

The observed differences in curve shape for the various classes of amines do not necessarily imply drastic differences in reaction mechanism. They could result from differences in the rates of step 2. With the suggested intervention of step 2 in primary amine oxidation, the observed failure of peak potentials to reflect ease of removal of an electron is understandable.

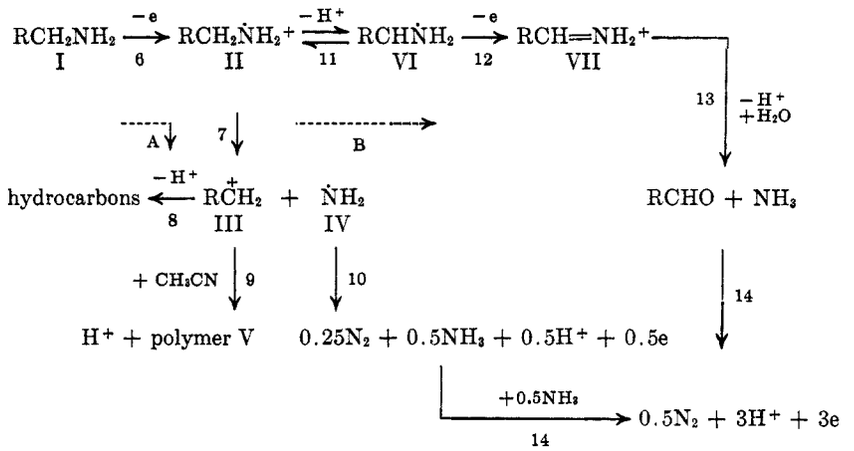
In order to generate the type of current–time curve illustrated in Figure 1, a sequence of at least three steps is necessary. These are outlined as eq 2. This be-

$$\mathrm{A} \xrightarrow[1]{-e} \mathrm{B} \xrightarrow[4]{} \mathrm{E} \xrightarrow[5]{-e} \mathrm{F} \qquad (2)$$

havior would be caused by occurrence of an initial electrochemical step or steps symbolized by step 1 followed by a chemical step or steps (reaction 4) which do not contribute to the current and which are quite slow. These are followed by an electrochemical step or steps which do contribute to the current. Because step 4 is slow, there is a delay in the build-up of component E, hence a delay in the appearance of current from step 5 which shows as a peak in the current–time curve.

The reaction sequence (1 to 4 to 5) responsible for the current–time behavior on exhaustive electrolysis cannot contain the same steps observed by voltammetry (1 to 2 to 3) because of the necessity for a great difference in rate between steps 2 and 4. Step 2 must be fast enough to produce significant concentrations of

SCHEME I



product in a fraction of a second. Step 4 produces significant concentrations of product only after many minutes.

It is suggested, therefore, that a mechanism involving two separate paths is warranted by the evidence. Each of these involves more than one electron transfer, may involve one or more chemical steps, and may have the initial electron-transfer step in common. The reaction is outlined in Scheme I.

In this scheme, the steps of eq 1 are designated as path B, consisting of steps 6, 11, 12, 13, and 14. Path A corresponds to eq 2 and consists of steps 6, 7, 8, 9, 10, and 14.

It has been acertained that ammonia does react at the potential in question and that the products are nitrogen gas and protons. Accordingly, it is suggested that ammonia serves as the reactive intermediate responsible for the peak in the low-potential current–time curve. It may be noted that if the reaction is carried out while sparging with an inert gas, the characteristic peak is not observed; this would be expected if ammonia, responsible for the peak, were being removed as rapidly as it was being formed. In addition, the 256-mμ peak, attributed to reaction of carbonium ion with solvent, appears as the current starts to rise. This indicates that reaction 7 is slow enough to cause the delayed formation of ammonia. When secondary and tertiary amines are allowed to react under comparable conditions, no peak is obtained.[6] This is reasonable, since although all amines could undergo reactions analogous to steps 6, 7, 8, and 9, only primaries would produce a precursor of ammonia on carbon–nitrogen cleavage.

The appearance of lower hydrocarbon fragments is suggestive of a radical reaction, possibly indicating that 7 may also yield some $R\dot{C}H_2$ and NH_2^+.

When the anode potential is increased, the rate of electrochemical step 12 is increased while that of chemical step 7 is unaffected. Accordingly, path B becomes more important at high potentials. We suggest that this path is responsible for the voltammetry results. If so, all steps must be fast compared with the rate of reaction 7. As an indication that this is the case, when the electrolysis of *n*-propylamine is run at 1.2 v the ultraviolet absorption peak of MPP, the product of step 13, appears and increases continuously from the start of the experiment, while that for the carbonium ion condensation product (V) appears only at the time of the current–time peak. Furthermore, when the controlled-potential electrolysis of *n*-propylamine is carried out at 1.9 v, making path B dominant, no current–time peak is formed. It was noted that at the high potential, the yields of hydrocarbon and of V were considerably smaller. Yields of MPP could not be compared, since it is destroyed at the high potential.

The structure of *t*-butylamine prevents its fitting the general mechanism; it is impossible for it to go by path B since it cannot form a double bond between carbon and nitrogen. Accordingly, one would predict a steep voltammetry peak and a current–time curve similar to that of Figure 1. Actually, however, the voltammetry peak is very similar to those of other amines. If the explanation advanced above for voltammetry curve shapes is valid, a rate-controlling step between two electron-transfer steps, this must occur for *t*-butylamine in path A of the general mechanism. This is illustrated in Scheme II.

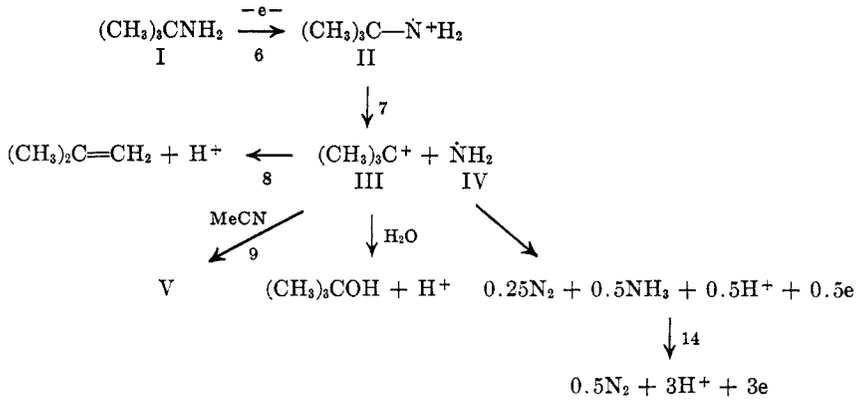
The stability of *t*-butylcarbonium ion could account for step 7 being much faster than is normally the case. If so, the voltammetry curve could be due to the sequence 6 to 7 to 10 to 14.

If it is true that reaction 14 is important in cyclic voltammetry of *t*-butylamine, analogous to reaction 12 for other amines, because of the greater rate of step 7, then step 7 for *t*-butylamine would not be slow enough to produce a peaked current–time curve. This is quite in agreement with the observation that of the amines studied, *t*-butylamine was the only primary amine which produced a low-potential current–time which decayed continuously to background.

We therefore suggest a dual mechanism with the relative importance of the two parts determined by the applied potential and molecular structure. To determine, for a particular compound and potential, the extent of one path relative to the other, one looks to quantitative determination of reaction products and finds that there are significant difficulties. Both paths give rise, in part, to the same products, nitrogen and proton. In acetonitrile, the yield of olefin is not indicative of the extent of path A, since a significant amount of the carbonium intermediate reacts with solvent. There is competition between ammonia and unreacted amine for protons formed in the reaction and competition between protonation and oxidation for the

(6) P. J. Smith, unpublished results.

SCHEME II



ammonia formed in the reaction. Finally, although aldehyde is formed only by reaction sequence B, it is not possible to choose relative amounts of paths A and B on the basis of aldehyde found because the aldehyde may self-condense in the presence of amine or react to form the azomethine which is slowly oxidized at electrolysis potentials. Thus, various combinations of A and B and various amounts of ammonia oxidation were assumed. The expected product yields were calculated and compared with experimental results. By successive approximations the best fit was found. This information is given in Table I for each compound studied by indicating the fraction of reaction going by path A and the fraction of ammonia oxidized. Product yields on the basis of assumed mechanisms are shown in parentheses beside the corresponding experimental values.

Of the mechanisms proposed, path B bears a resemblance to that suggested by Wei and Stewart[7] for the oxidation of primary amines by alkaline permanganate. This involves removal of a hydrogen atom by permanganate to leave a radical with the unshared electron α to the amine group. This radical is further oxidized to the imine conjugate acid which is hydrolyzed to ammonia and aldehyde. Removal of hydrogen atom is accomplished by our steps 6 and 11. Loss of the second electron followed by hydrolysis is reaction 13. Path B is also similar to the mechanism proposed by Taborsky[8] for amino acid oxidase deamination of α-amino acids. Here the oxidation involves the formation of the imino acid, analogous to our imine conjugate acid (VII), followed by hydrolysis to the keto acid.

Path A is somewhat similar to the deamination of primary amines by nitrous acid in that decomposition of the diazonium salt leads to formation of a carbonium ion having the same carbon skelton as the original amine.

Experimental Section

All boiling points are uncorrected. Near-infrared work was done on a Cary Model 14 spectrophotometer as was the ultraviolet study. The gas chromatographic work was done on an F & M Model 500 programmed high-temperature gas chromatograph equipped with a thermal conductivity detector and an Aerograph Hy-Fi Model 600-C with a flame-ionization detector.

Cyclic voltammetry curves were obtained on the apparatus previously described,[9] with sweep rates of 10 v/sec. Oxidation was accomplished at a platinum electrode. Typical solutions were 0.005 M in amine.

Controlled-Potential Coulometry.—Electrolyses were performed with a conventional potentiostat and the coulombs were measured with a hydrogen–nitrogen coulometer of the type described by Lingane.[10]

Reagents.—Reagent grade liquid amines were purchased from Eastman Organic Chemicals or from Matheson Coleman and Bell, purified by distillation, and checked for purity by titration with perchloric acid in acetic acid and by gas chromatography. They were then stored in the dark over sodium hydroxide pellets. Gaseous amines were obtained from Matheson lecture bottles and bubbled into MeCN; solution concentrations were determined by titration in acetic acid.

n-Propylpropanaldimine was prepared by the method of Campbell, bp 99–101°.[11] NPP was refluxed for 12–24 hr and the 2-methyl-2-pentenalpropylamine was distilled rapidly from sodium hydroxide, bp 173–78°.

Electrolytically produced MPP was compared by gas chromatography and by ultraviolet spectroscopy to the sample described above. The ultraviolet peak was used for quantitative determination of this compound.

Acetonitrile was purified as previously described.[12]

Sodium perchlorate was recrystallized from 95% ethanol, dried in a vacuum oven, and stored in a vacuum desiccator over P_2O_5.

Controlled-Potential Electrolysis.—Runs were made in a two-compartment cell using a Pt mesh anode, a 0.1 M Ag/Ag^+ in MeCN reference electrode, and a platinum cathode, dipped in a mercury pool to prevent production of gas. The cell was fitted with ground joints to permit operation with exclusion of air. Dried, degassed, supporting electrolyte–solvent mixtures could be introduced without being contaminated.

In a typical experiment, the cell was loaded with 0.10 M $NaClO_4$–MeCN solution and a pre-electrolysis was carried out at a potential slightly higher than that to be used for amine oxidation. When the current had become constant, the potential was readjusted and 50 μl of amine was added to the anode compartment through a septum. The reaction was then permitted to proceed until current had decayed to a constant value. Current was recorded as a function of time and was integrated with respect to time by the coulometer.

Product Analysis.—Approximately one-half of the initial amine is found as amine perchlorate and another 10–20% as ammonium perchlorate. Thus the electrolysis solution could be pipetted directly into dimethyl sulfoxide and the salts were determined as acids by titration with KOMe.[13] The amines were then freed by addition of NaOH flakes to the electrolysis solution, and identified quantitatively and qualitatively by injecting the solution directly into the gas chromatograph[14] or by running its near-infrared spectrum.

Volatile hyrocarbons were identified in the headspace gas by withdrawing 300-μl samples directly from the cell and analyzing

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(8) G. Taborsky, *Yale J. Biol.*, **27**, 267 (1955).
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them by glpc on a Porapak Q column. Nitrogen was determined in the same way by thermal conductivity chromatography on a 4A molecular sieve column.

Aldehydes were identified and determined in the electrolysis solution itself by ultraviolet spectroscopy and by chromatography on 9N9, THFP, and Porapak columns.

Registry No.—Butylamine, 109-73-9; hexylamine, 111-26-2; isobutylamine, 78-81-9; benzylamine 100-46-9; ammonia, 7664-41-7; propylamine, 107-10-8; cyclohexylamine, 108-91-8; allylamine, 107-11-9; methylamine, 74-89-5; *t*-butylamine, 75-64-9.

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ortho Metalations of Ring-Substituted Benzyldimethylamines by *n*-Butyllithium and Condensations with Benzophenone. Nucleophilic Mechanism. Cyclizations to Phthalans[1]

KARL P. KLEIN AND CHARLES R. HAUSER

Department of Chemistry, Duke University, Durham, North Carolina

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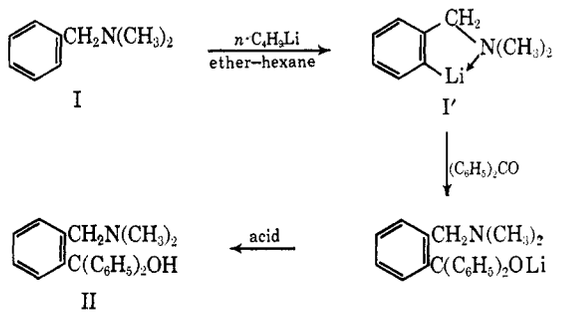
Several 2-, 3-, and 4-substituted and one 3,5-disubstituted benzyldimethylamines were lithiated with *n*-butyllithium in ether–hexane for 1 and 24 hr, and the resulting lithioamines were condensed with benzophenone to form the corresponding carbinolamines. The two 3-substituted benzyldimethylamines afforded only one of the two possible isomers: that which arose through lithiation at the position *ortho* to both the dimethylaminomethyl group, and the ring substituent. The yields were generally good to excellent after the longer lithiation period. Also, deuterations of certain of the lithioamines were effected after the 1-hr period. The results indicated that a nucleophilic mechanism operated. Six of the carbinolamines were cyclized through their methiodides to form substituted phthalans, generally in good yields. Both the lithiation–condensation and the cyclization methods should be useful in synthesis.

Recently,[2] benzyldimethylamine (I) was lithiated with *n*-butyllithium in ether–hexane at 25–30° to form *o*-lithioamine I′ which was condensed with benzophenone to give the *o*-carbinolamine (II, Scheme I).

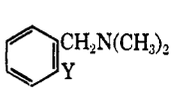
SCHEME I



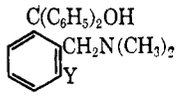
I I′ II

In the present investigation, nine ring-substituted benzyldimethylamines were similarly *ortho* lithiated and condensed with benzophenone to produce the corresponding carbinolamines, all of which were new. The lithiations were effected for both 1- and 24-hr periods before adding the ketone; the condensation period after adding the ketone was 4 hr. The results are summarized in Tables I and II; also in Table I are included, for comparison, the yields of the known carbinolamine (II) after the two lithiation periods.

The three 2-substituted benzyldimethylamines (IIIa–c) afforded the 2,6-disubstituted benzyldimethylamines (IVa–c), respectively. The structures of these products were supported by their infrared spectra which showed peaks in the region 790–800 cm^{-1}, ascribable to the three adjacent aromatic hydrogens.[3]

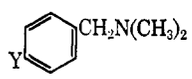


IIIa, Y = Cl
b, Y = F
c, Y = CH=C(C6H5)2

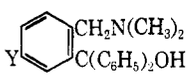


IVa, Y = Cl
b, Y = F
c, Y = CH=C(C6H5)2

The three 4-substituted benzyldimethylamines (Va–c) produced the 2,5-disubstituted benzyldimethylamines (VIa–c), respectively. The structures of these products were supported by infrared spectra which showed peaks in the region 800–820 and 870–893 cm^{-1}, indicating the two adjacent and one free aromatic hydrogen, respectively.[3]

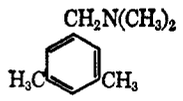


Va, Y = CH_3
b, Y = OCH_3
c, Y = Cl

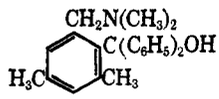


VIa, Y = CH_3
b, Y = OCH_3
c, Y = Cl

The 3,5-dimethylbenzyldimethylamine (VII) afforded the 2,3,5-trisubstituted benzyldimethylamine (VIII). The structure of VIII was indicated by its infrared spectrum which exhibited a peak at 863 cm^{-1}, consistent with a structure having one free aromatic hydrogen.[3]



VII



VIII

The two 3-substituted benzyldimethylamines (IXa, b), each of which might conceivably form isomers, nevertheless afforded only a single product. Evidently, lithiation occurred preferentially at the

(1) Supported by Army Research Office (Durham) and the National Science Foundation.

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