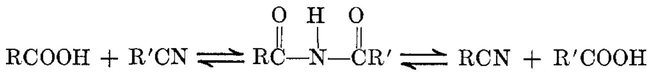
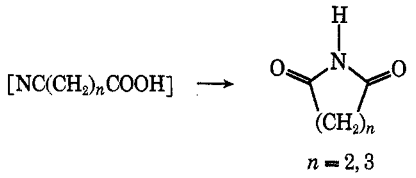
an equilibrium exists which proceeds through an imide addition product, the equilibrium being displaced toward the side of the weaker carboxylic acid.[3]



$$RCOOH + R'CN \rightleftharpoons RC(=O){-}N(H){-}C(=O)R' \rightleftharpoons RCN + R'COOH$$

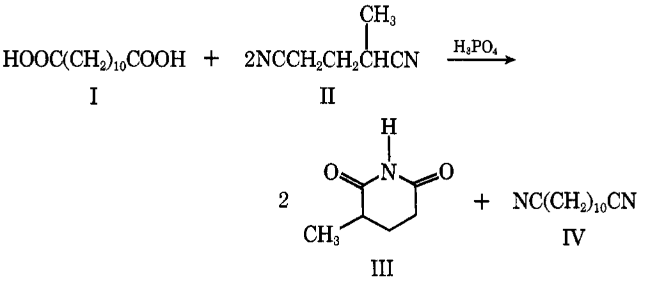
We have found that replacing acetonitrile with short-chain dinitriles (specifically succinonitrile, glutaronitrile, and α-methylglutaronitrile) provides two distinct advantages in the use of this reaction as a synthetic method. First, the use of pressure equipment to reach the required high temperatures (150–300°) is avoided. Second, distinct improvements in yield of product nitrile, especially for aliphatic systems, have been realized due to the fact that the once-exchanged short-chain cyano acid undergoes an internal cyclic imide formation, removing it from the equilibrium and thus driving the reaction to completion.

$$RCOOH + NC(CH_2)_nCN \rightleftharpoons RCN + [NC(CH_2)_nCOOH]$$



Although the reaction proceeds uncatalyzed, there are apparent advantages in adding 0.5–1.0 wt % of compounds such as sulfonic, sulfuric, or phosphoric acids or their various salts.

We have used this method for the synthesis of aliphatic dinitriles. Reaction of 1,12-dodecanedioic acid (I) with 2 molar equiv of α-methylglutaronitrile (II) afforded a 98% recovered yield of α-methylglutarimide (III) and a 97% recovered yield of 1,12-dodecanedinitrile (IV). Azelaonitrile similarly was prepared in 87% recovered yield from azelaic acid.



Conversion of I into IV also was accomplished in high yield by refluxing with 2 molar equiv of glutaronitrile or succinonitrile. The products glutarimide, mp 158–159°, and succinimide, mp 126–127°, were purified by recrystallization from chloroform and characterized by infrared, proton magnetic resonance, and elemental analyses.

Experimental Section

A typical reaction involved refluxing (*ca.* 285°) a mixture of 1150 g (5.0 mol) of dodecanedioic acid (I) and 1150 g (10.65 mol) of α-methylglutaronitrile (II) containing 11.6 g of 85% H_3PO_4 for 18 hr. The resulting black solution was cooled and vacuum distilled to give 1246 g (98% of theoretical) of tan solid α-methylglutarimide (III), bp 137° (7 mm), and 929.7 g (97% of theoretical) of yellow liquid dodecanedinitrile (IV), bp 193° (8 mm). Aqueous NaOH washing and redistillation afforded colorless liquid IV, mp 21–22°.

Anal. Calcd for $C_{12}H_{20}N_2$: C, 75.00; H, 10.42; N, 14.58. Found: C, 74.95; H, 10.52; N, 14.54.

Redistillation and recrystallization from 1:1 benzene–cyclohexane afforded white, crystalline III, mp 98.5–99°.

Anal. Calcd for $C_6H_9NO_2$: C, 56.69; H, 7.09; N, 11.02. Found: C, 56.44; H, 7.14; N, 11.00.

Proton magnetic resonance and infrared spectra of III were consistent with the proposed cyclic imide structure.

Registry No.—I, 693-23-2; III, 29553-51-3; IV, 4543-66-2.

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[3] F. Becke and T. F. Burger, *Justus Liebigs Ann. Chem.*, **716**, 78 (1968).

Ionization Scheme for the *N*,*N*-Di(carboxymethyl)anilines

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It is well known that aliphatic amino acids exist in aqueous media in the form of dipolar molecules (zwitterions). In fact, the equilibrium between the neutral form and the dipolar form favors the dipolar form by a factor of several tens of thousands.[1] However, if the basicity of the nitrogen is reduced by interaction with an aromatic ring, then the neutral form and dipolar form can have comparable stabilities in aqueous media. For example, the *o*-, *m*-, and *p*-aminobenzoic acids have dipolar molecule–neutral molecule ratios of 0.2, 2.5, and 0.17.[2] Infrared work in D_2O and D_2O–dioxane mixtures[3] demonstrated that the three pyridinecarboxylic acids existed largely as dipolar molecules in aqueous media, but the neutral form was present to a nonnegligible extent.

The purpose of this work was to decide whether or not dipolar molecules or ions were important species in the ionization scheme of the *N*,*N*-di(carboxymethyl)anilines. Two complementary studies on five of these acids (unsubstituted and the para-substituted chloro, fluoro, methyl, and methoxy acids) led to the conclusion that dipolar molecules or ions are not involved to any appreciable extent in the ionization scheme of these acids.

The ionization scheme is presented in Scheme I. It can be seen that H_2A and HA^- can both exist in either a dipolar form or a neutral form, *i.e.*, protonated or nonprotonated on the nitrogen. H_3A^+ represents the fully protonated species. Since thermodynamics does not discriminate between the possible forms, the three experimental pK's which completely describe the ionization scheme for any one acid are actually composite pK's. However, using nonthermodynamic experiments and reasoning, it has been possible to demonstrate

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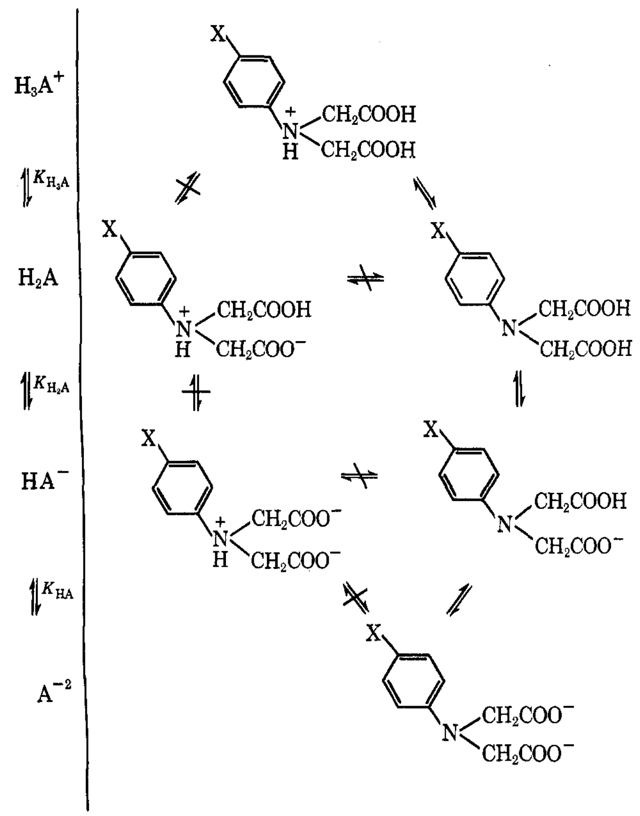
(1) J. P. Greenstein and M. Winitz, "Chemistry of the Amino Acids," Vol. 1, Wiley, New York, N. Y., 1961, p 447.

(2) J. J. Christensen, D. P. Wrathall, R. M. Izatt, and D. O. Tolman, *J. Phys. Chem.*, **71**, 3001 (1967).

(3) J. F. Wojcik and T. H. Stock, *ibid.*, **73**, 2153 (1969).

SCHEME I



that the pathway connected by solid arrows is the major pathway for these acids.

Ultraviolet Spectroscopy.—The ultraviolet spectra of the five acids studied here were measured as a function of pH. A peak was observed for all of the acids around 300 nm in basic and weakly acid solution. This peak disappeared as the pH was lowered below about 2. Using these data, the pK data determined in this work, and an iterative computer program, the molar extinction coefficients were determined for all possible species for all of the acids at the proper maxima. These data are given in Table I. The extinction coefficient for H_3A^+ is far smaller than that for the other species for all five acids in the 300-nm region. The extinction coefficients for H_2A and HA^- represent composites made up from contributions from dipolar and neutral forms in the proportion that these make up H_2A and HA^-. Now aniline compounds are known to absorb in the ultraviolet region around 300 nm. However, the anilinium ions do not absorb here; protonation shifts the absorption to lower wavelengths. Thus, it is expected that the extinction coefficients of H_2A and HA^- represent a contribution from the neutral form only. The ratios of the extinction coefficient for A^{2-} to that for HA^- are 1.17, 0.99, 1.16. 1.09, and 1.17 for the chloro, fluoro, unsubstituted methyl, and methoxy compounds, respectively. The same ratios for A^{2-} and H_2A are 1.10, 0.98, 1.12, 1.11, and 1.06. It would be expected that either of these ratios would increase across the series if a dipolar form were important for either HA^- or H_2A. This would be so because the dipolar form in each case would make the extinction coefficient for that species smaller, the effect being greater as the dipolar form contributed more to the structures of H_2A and HA^-. The increase across the series would be expected because the basicity of the nitrogen increases across the series. This would make the dipolar form more important for the methoxy compound than the others. The absence of such an increase supports the ionization pathway proposed. Both sets of data can be used to estimate an upper limit to the concentration of dipolar species in the methoxy case. Assuming a 10% uncertainty in all the extinction coefficients, it is estimated that the amount of dipolar species in the methoxy case must be less than 15% in order to account for the approximate constancy of all of the ratios. The per cent dipolar species would be less than 15% for the other compounds.

TABLE I

MOLAR EXTINCTION COEFFICIENTS[a]

Substituent	Registry no.	H_2Y λ_{max}	ϵ	HY^- λ_{max}	ϵ	Y^{2-} λ_{max}	ϵ
p-Cl	30042-69-4	305	2000	308	2100	312	2300
p-F	31045-00-8	302	2500	308	2600	312	2500
p-H	1137-73-1	293	2200	297	2200	301	2500
p-CH_3	28444-51-1	298	2200	304	2100	308	2300
p-CH_3O	30042-67-2	309	2300	314	2600	318	2700

[a] λ values are in nm, ϵ values are in M^{-1} cm^{-1}.

Dissociation Constants.—The three different acid dissociation constants for each of the five acids are tabulated in Table II. Each of the three pK's was analyzed in terms of a Hammet $\sigma\rho$ plot.[4] It is expected that the pK (or pK's) which involves the ionization of the proton off the nitrogen will be most sensitive to substituent effects. The ρ value for the first ionization was 3.5, that for the second 0.0, and that for the third 0.8. It is seen that the first ionization is by far the most sensitive to substituent effects. This ρ value compares favorably with the value of 3.56 for the ionization of *N*,*N*-dimethylanilinium ions.[5]

TABLE II

p*K* VALUES OF THE *N*,*N*-DI(CARBOXYMETHYL)ANILINES[a]

Substituent	pK_{H_3Y}	pK_{H_2Y}[b]	pK_{HY}[b]
p-Cl	0.5 ± 0.2	2.38 ± 0.02	4.84 ± 0.02
p-F	0.2 ± 0.2	2.43 ± 0.02	4.90 ± 0.02
p-H	−0.4 ± 0.2	2.34 ± 0.03[c]	5.04 ± 0.02[c]
p-CH_3	−0.3 ± 0.2	2.21 ± 0.13	5.24 ± 0.02
p-CH_3O	−1.3 ± 0.2	2.41 ± 0.05	5.16 ± 0.02

[a] All averages based on five determinations. [b] Data for *m*-SO_3H and *p*-SO_3H agree reasonably well with these data depending on what σ values are used (L. G. Sillen and A. E. Martell, "Stability Constants," 2nd ed, The Chemical Society, London, 1964, p 648). [c] Compare with 2.40 ± 0.1 and 4.96 ± 0.1 and 20° [G. Schwarzenbach, G. Anderegg, W. Schneider, and H. Senn, *Helv. Chim. Acta*, **38**, 1147 (1955)].

Conclusions

The results presented here can be extended to other ring-substituted compounds. The σ value for the *p*-methoxy compound is one of the most negative of the σ values and suggests that the nitrogen in the methoxy compound would be the most basic of all aniline com-

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[4] J. E. Leffler and E. Grunwald, "Rates and Equilibria of Organic Reactions," Wiley, New York, N. Y., 1963, p 173.

[5] P. R. Wells, "Linear Free Energy Relationships," Academic Press, London, 1968, p 12.

pounds. This would then be the most likely possibility for the observation of a dipolar species. Since it is not conclusively observed in this case, no substituent of those commonly investigated would be expected to produce dipolar species.

Experimental Section

The acids were synthesized from the parent anilines using the procedure outlined by Pettit and Irving.[6] Neutralization equivalents were determined for all of the acids and found to agree with the theoretical value within 1%. The neutralization equivalent of the *p*-methoxy compound was that for a monohydrate. Microanalysis for this compound confirmed the result. Microanalyses were carried out by Galbraith Laboratories.

Anal. Calcd for *N*,*N*-di(carboxymethyl)aniline ($C_{10}H_{11}NO_4$): C, 57.41; H, 5.30; N, 6.70. Found: C, 57.49; H, 5.17; N, 6.57.

Anal. Calcd for *N*,*N*-di(carboxymethyl)-4-chloroaniline ($C_{10}H_{10}NO_4Cl$): C, 49.29; H, 4.14; N, 5.75. Found: C, 49.48; H, 4.08; N, 5.60.

Anal. Calcd for *N*,*N*-di(carboxymethyl)-4-fluoroaniline ($C_{10}H_{10}NO_4F$): C, 52.86; H, 4.44; N, 6.17. Found: C, 52.74; H, 4.25; N, 5.99.

Anal. Calcd for *N*,*N*-di(carboxymethyl)-4-methylaniline ($C_{11}H_{13}NO_4$): C, 59.18; H, 5.87; N, 6.28. Found: C, 59.96; H, 6.09; N, 6.21.

Anal. Calcd for *N*,*N*-di(carboxymethyl)-4-methoxyaniline ($C_{11}H_{13}NO_5 \cdot H_2O$): C, 51.36; H, 5.88; N, 5.45. Found: C, 51.51; H, 6.04; N, 5.44.

Nmr spectra of all of the acids were run in NaOD solutions on a Varian A-60. All of the compounds gave the expected aromatic chemical shifts around 6.72–6.92 relative to 3-(trimethylsilyl)-propanesulfonate. The methylene chemical shifts fell between 3.82 and 3.89. The methyl chemical shift in the *p*-methyl compound was 2.22. The methyl shift in the *p*-methoxy compound was at 3.70. Integrated results confirmed the molecular formulas.

The ultraviolet spectra were run on a Beckman DK-2 at 25°. A stock solution of the acid was prepared in dilute sodium hydroxide solution and then added to the final buffer solution. Buffers were chosen which did not absorb in the wavelength region of interest. The reference cell contained buffer. pH measurements were made on a Beckman Research pH meter. Buffer pH's were in steps of one between 1.00 and 10.00. In order to obtain the best set of extinction coefficients an iterative procedure was used. Based on the pH value and the p*K*'s of the acids the concentration of each species was measured. With these values, extinction coefficients were chosen which best fit the absorption change with pH.

The p*K* values for H_2A and HA^- were determined from buffer solutions of these acids at 25° and an ionic strength of 0.100 using KCl as the added electrolyte. A glass electrode–calomel electrode system was used to determine the pH of the solution. The pH meter was calibrated against an acetate buffer and 0.01 *M* HCl. The hydrogen ion concentrations of these buffer solutions are known and the pH reading for the unknown solutions was converted to hydrogen ion concentration. Thus the p*K* values are concentration constants. In some cases the pH of the solution drifted. This was attributed to an acid-catalyzed decarboxylation of the acids. All pH readings were measured as a function of time and extrapolated back to time of mixing. The correction was negligible for the chloro and fluoro compounds and was greatest for the methoxy. However, the pH drift was small and easily measured.

The p*K* of H_3A^+ was determined spectrophotometrically. The acidity of the solutions were adjusted with HCl and the ionic strength was kept constant at the value of 3.0 with KCl. Decomposition proceeded more rapidly at these acidities and all spectra were measured as a function of time. This constant is also a concentration constant.

Acknowledgments.—The authors wish to acknowledge support from the National Science Foundation under Research Grant GP-9034.

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[6] L. D. Pettit and H. M. N. J. Irving, *J. Chem. Soc.*, 5336 (1964).

The Reaction of Ionic Thiocyanates with Diacyl Peroxides. The Formation of Thiocyanogen

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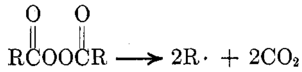
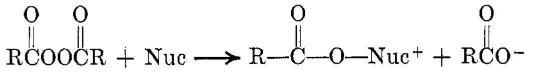
The thermolysis[1] and photolysis[2] of acyl peroxides represent convenient modes for the production of alkyl radicals, *e.g.*



$$\mathrm{R\overset{O}{\overset{\|}{C}}OO\overset{O}{\overset{\|}{C}}R \longrightarrow 2R\cdot + 2CO_2} \qquad (1)$$

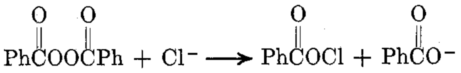
However, complications due to ionic processes can arise by rearrangement to carboxylic carbonic anhydrides involving carboxy inversion.[3] Acyl peroxides are also susceptible to a variety of nucleophiles (Nuc)[4] and the acyloxy intermediate **1** has been shown in a number of



$$\mathrm{R\overset{O}{\overset{\|}{C}}OO\overset{O}{\overset{\|}{C}}R + Nuc \longrightarrow \underset{\textbf{1}}{R{-}\overset{O}{\overset{\|}{C}}{-}O{-}Nuc^+} + R\overset{O}{\overset{\|}{C}}O^-} \qquad (2)$$

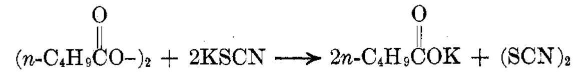
1

cases to be involved further in competing heterolytic and homolytic processes.[5] Thus, benzoyl hypochlorite, an intermediate in the reaction of benzoyl peroxide and ionic chloride, is capable of electrophilic and free-radical chlorinations.[6]



$$\mathrm{Ph\overset{O}{\overset{\|}{C}}OO\overset{O}{\overset{\|}{C}}Ph + Cl^- \longrightarrow Ph\overset{O}{\overset{\|}{C}}OCl + Ph\overset{O}{\overset{\|}{C}}O^-} \qquad (3)$$

We wish to report a similar facile reaction when diacyl peroxides are exposed to ionic thiocyanates. Thus, a solution of 0.040 *M* valeryl peroxide and 0.080 *M* potassium thiocyanate in acetonitrile afforded thiocyanogen and potassium valerate according to eq 4 after 5 hr at



$$(n\text{-}\mathrm{C_4H_9\overset{O}{\overset{\|}{C}}O\text{-})_2 + 2KSCN \longrightarrow 2}n\text{-}\mathrm{C_4H_9\overset{O}{\overset{\|}{C}}OK + (SCN)_2} \qquad (4)$$

room temperature. Less than 3% carbon dioxide and no butyl thiocyanate were detected.[7]

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[1] See A. G. Davies, "Organic Peroxides," Butterworths, London, 1961; E. Hawkins, "Organic Peroxides, Their Formation and Reactions," E. and F. Spon, London, 1961; C. Walling, "Free Radicals in Solution," Wiley, New York, N. Y., 1957; W. A. Pryor, "Free Radicals," McGraw-Hill, New York, N. Y., 1966.

[2] R. A. Sheldon and J. K. Kochi, *J. Amer. Chem. Soc.*, **92**, 4395 (1970).

[3] (a) F. D. Greene, H. S. Stein, C. C. Chu, and F. M. Vane, *ibid.*, **86**, 2081 (1964); (b) C. Walling, H. N. Moulden, J. H. Waters, and R. C. Neuman, *ibid.*, **87**, 518 (1965); (c) H. Hart and D. Wyman, *ibid.*, **81**, 4891 (1959); (d) J. K. Kochi, *ibid.*, **85**, 1958 (1963); (e) R. C. Lamb and J. R. Sanderson, *ibid.*, **91**, 5034 (1969); (f) S. Oae, T. Kashiwagi, and S. Kozuka, *Chem. Ind.* (*London*), 1964 (1965); (g) D. B. Denney and N. Sherman, *J. Org. Chem.*, **30**, 3760 (1965); (h) D. S. Tarbell, *Accounts Chem. Res.*, **2**, 296 (1969).

[4] J. O. Edwards, "Peroxide Reaction Mechanisms," Interscience, New York, N. Y., 1962, p 67 ff.

[5] (a) F. D. Greene, W. Adam, and J. E. Cantrill, *J. Amer. Chem. Soc.*, **83**, 3461 (1961); F. G. Greene and W. Adam, *J. Org. Chem.*, **28**, 3550 (1963); **29**, 136 (1964); (b) Y. Ogata and I. Tabushi, *Bull. Chem. Soc. Jap.*, **31**, 962 (1958); (c) L. Horner and E. Schwenk, *Angew. Chem.*, **61**, 411 (1949); L. Horner and B. Anders, *Chem. Ber.*, **95**, 2470 (1962); C. Walling and N. Indictor, *J. Amer. Chem. Soc.*, **80**, 5814 (1958); C. Walling and R. B. Hodgson, *ibid.*, **80**, 228 (1958); C. Sato and T. Otsu, *Chem. Ind.* (*London*), 125 (1970).

[6] (a) J. K. Kochi, B. M. Graybill, and M. Kurz, *J. Amer. Chem. Soc.*, **86**, 5257 (1964); (b) N. J. Bunce and D. D. Tanner, *ibid.*, **91**, 6096 (1969).

[7] C. L. Jenkins and J. K. Kochi, *J. Org. Chem.*, in press.