Oxidative Decarboxylation. On the Mechanism of the Potassium Persulfate Promoted Decarboxylation Reaction[1]

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The competitive rates for the potassium persulfate promoted decarboxylation of a series of aliphatic and aromatic carboxylic acid salts were determined. The rates for the ring-substituted phenylacetic acids showed a linear free energy correlation, $\rho = -0.44$, with σ^+ substituent constants. The aliphatic carboxylates, likewise, showed a structure–reactivity relationship; the carboxylates that produce the most stable radicals upon loss of carbon dioxide are decarboxylated at the fastest rate. No clear demarcation between the rates of decarboxylation of the aromatic and aliphatic carboxylates was found. The benzyl-substituted phenylacetic acids, both phenyl and methyl substituted, show only small increases in the rates of electron transfer from their anions, while ring substitution was more effective in increasing the rates. Electron transfer from the HOMO of the aromatic carboxylate is marginally easier than transfer from the aliphatic carboxylate anions that undergo concerted-electron-transfer fragmentation.

Introduction

A number of years ago we reported the preliminary results of a study of the oxidative decarboxylation of a series of salts of substituted phenylacetic acids with potassium persulfate.[2]

Competitive decarboxylation reactions between the salts of phenylacetic acid-*1*-^{14}C and nonradioactive ring-substituted phenylacetic acids were carried out in degassed systems (74.3 °C). The ratios of rate constants for the oxidative decarboxylation of various substituted phenylacetates relative to phenylacetate-*1*-^{14}C was found to decrease with the decreasing electron density at the benzylic carbon. This trend is opposite to that observed in the effects of substituents on the dissociation constants for substituted phenylacetic acids[3] and is consistent with substituent stabilization leading to the electron transfer from a carboxylate anion.

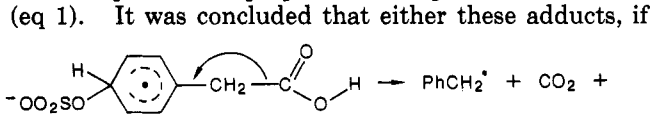
The relative rate data were fitted to the Hammett equation and the following linear free energy correlations were obtained.

$$\log (k/k_0) = -0.626\sigma + 0.12 \qquad s = 0.081 \qquad r = 0.939$$

$$\log (k/k_0) = -0.436\sigma^+ + 0.051 \qquad s = 0.035 \qquad r = 0.984$$

The better correlation with σ^+ values in comparison with that obtained with σ substituent constants was interpretated as an electron transfer from the carboxylate anion which was a process concerted, to some extent, with the loss of carbon dioxide during the rate-determining step.

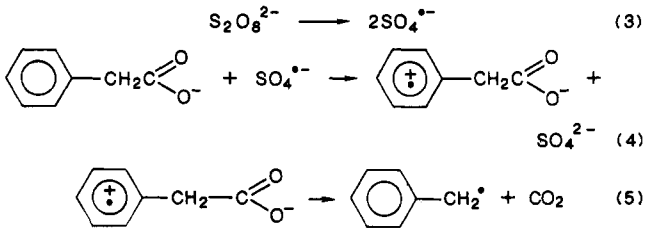
Subsequent to this paper the results of a study on the mechanism of the persulfate-promoted decarboxylation of phenylacetic acid was reported.[4] The reaction of phenylacetic acid with hydroxy radicals gave rise to ESR spectra assigned to a hydroxy radical adduct which, in acidic media, lost both water and CO_2 to yield a benzyl radical. By analogy, although no ESR signal was detected, the same process was proposed for the persulfate oxidation (eq 1). It was concluded that either these adducts, if

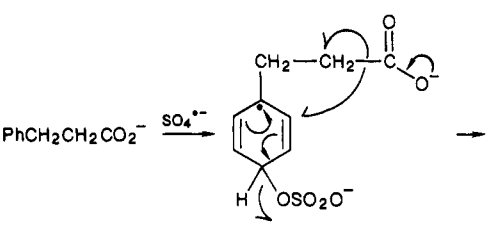


(1)

formed, fragment more readily than the corresponding hydroxyl adducts or the sulfate radical anion removes one electron from the aromatic ring before, or synchronously with, heterolytic fragmentation of the side chain. The addition–elimination fragmentation was favored by the authors since dihydrocinnamic acid underwent oxidation three times faster than valeric acid (eq 2). Most subsequent workers,[5-8] however, have considered the oxidative decarboxylation to be a one-step electron-transfer reaction followed by carboxylate fragmentation, Scheme I.

Scheme I





It is now clear that aromatic alkyl side-chain oxidation proceeds by a radical cation intermediate and numerous examples of these processes are present in the literature for oxidations utilizing a variety of reagents,[9] including the sulfate anion radical.[5,8,10]

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Scheme II

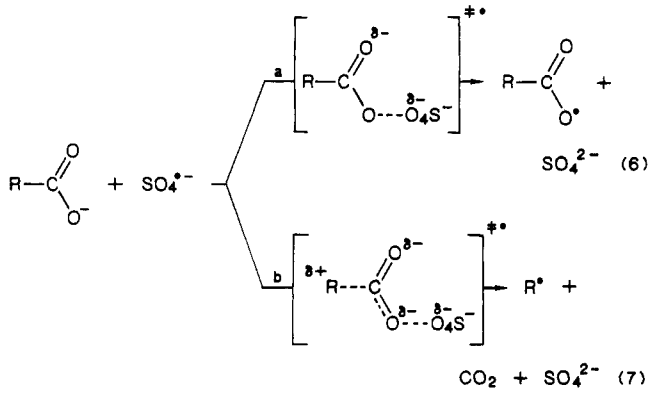


Table I. Relative Rate Constants for the Potassium Persulfate Promoted Decarboxylation of the Salts of Substituted Phenylacetic Acids at 74.3 °C

substit	rel rate[a]	substit	rel rate[a]
p-methoxy	2.51 ± 0.09	*p*-chloro	0.919 ± 0.007
p-phenoxy	2.18 ± 0.02	*m*-fluoro	0.875 ± 0.010
p-methyl	1.46 ± 0.03	*m*-bromo	0.815 ± 0.006
H	1.00	*m*-chloro	0.784 ± 0.020
p-bromo	0.958 ± 0.008	*o*-methyl	1.31 ± 0.02

[a] The values given are averages of four or five independent determinations and the errors reported are average deviations from the mean value given.

If the radical cation fragmentation process is the pathway responsible for the decarboxylation of phenylacetate (see Scheme I), then a second competitive mechanism is responsible for the decarboxylation of the aliphatic substrates (see Scheme II).

Both processes have been proposed for the anodic oxidation of the cesium salts of aromatic carboxylic acids,[11] Scheme I (eq 4 and 5) for the salt of 9-methylanthracene-10-acetic acid and Scheme II for the salts of mono-, di-, and triphenylacetic acids.

It appears necessary to compare the two competitive processes and to obtain further evidence for the proposed process responsible for the persulfate-promoted oxidation of aromatic carboxylate anions.

Results

The competitive decarboxylation rate ratios between the salts of phenylacetic acid-*1*-^{14}C and a series of substituted phenylacetic acids are listed in Table I.

The values listed were obtained by the measurement, upon acidification, of the total amounts of carbon dioxide produced from the reactions of aqueous mixtures of phenylacetic acid-*1*-^{14}C (0.0132–0.0148 M), substituted phenylacetic acids (0.0132–0.0148 M), potassium hydroxide (0.26 M), and potassium persulfate (0.057 M). The reaction mixtures gave quantitative yields of labeled and unlabeled carbon dioxide at 20 kinetic half-lives for the decarboxylation reaction (14 days). Quantitative determination of the amount of carbon dioxide formed during the course of the competitive decarboxylation reactions was carried out on acidified reaction mixtures by using standard vacuum-line procedures.[12] The measured carbon dioxide was absorbed quantitatively in a mixture of ethanolamine–ethylene glycol monomethyl ether on the vacuum line, and the amount of radioactive carbon dioxide was determined by using liquid scintillation counting techniques.[13]

Table II. Relative Rates of Decarboxylation of Salts of Secondary, Tertiary, Allylic, and Benzylic Acids

radical	acid	rel rate
benzylic	phenylacetic	1.000
allylic	cyclohexene-3-carboxylic	0.359 ± 0.014
tertiary	pivalic	0.161 ± 0.005
secondary	cyclohexanecarboxylic	0.0024 ± 0.009
secondary	isobutyric	0.0008 ± 0.0003

Table III. Relative Rates of Peroxydisulfate-Promoted Decarboxylation of the Salts of α- and Ortho-Substituted Phenylacetic Acids at 74.3 ± 0.10 °C

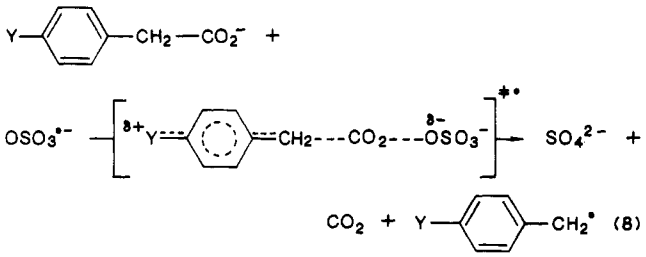
acid	rel rate	acid	rel rate
phenylacetic	1.00	α,α-dimethyl-phenylacetic	0.743 ± 0.017
diphenylacetic	1.19 ± 0.06		
triphenylacetic	1.44 ± 0.03	*o*-tolylacetic	1.31 ± 0.02
mandelic	0.509 ± 0.004	α-methyl-*o*-tolylacetic	1.05 ± 0.02
α-methylphenylacetic	0.651 ± 0.007	α-methyl-*p*-tolylacetic	1.06 ± 0.02

Competitive decarboxylations between aromatic and nonaromatic carboxylates give some indication of the importance of aromatic electron-transfer–decarboxylation and decarboxylation by oxidation of the carboxylate anion. Concerted-electron-transfer–fragmentation, Scheme II, path a, will show a dependence upon the rate of transfer and the structure of the radical formed upon loss of CO_2. The results obtained from a representative number of structurally different carboxylic acids are listed in Table II.

Structurally different benzylic substituted aromatic carboxylate anions were subjected to the persulfate-promoted decarboxylation (see Table III) since benzylic substitution will affect the magnitude of the relative rates if the oxidation proceeds by either Scheme I or Scheme II.

Discussion

The relative rates of decarboxylation for a series of substituted phenylacetic acids and phenylacetic acid itself show an excellent correlation with σ^+ substituent constants. The original interpretation for this observation was that the mechanism for the oxidative decarboxylation proceeded by an electron transfer from the carboxylate anion to the sulfate anion radical, through a polar transition state, which was concerted with the breaking of the benzylic carbon–carbon dioxide bond (eq 8). This conclusion



was fortified by the observation that, under these conditions, the relative rates of decarboxylation of several aliphatic acids also appeared to be dependent upon the structure of the radical produced upon oxidative decarboxylation.[14]

Persulfate-promoted oxidations of a series of substituted benzoate anions initiated by pulse radiolysis have been reported to proceed by an electron-transfer process.[15] The

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authors reported ρ values, using σ substituent constants, of −2.4 for electron transfer from the carboxylates as well as from a homologus series of substituted benzenes. With use of their kinetic data for the reactivity of the arenes, a recalculation of their Hammett plot, using σ_p substituent constants, gives a ρ of −1.59 (r = 0.94, s = 0.24). Considering only the para-substituted benzoates (9 points, omitting o-Br, o-CH_3, and m-CH_3),[16] a ρ value of −1.91 (r = 0.84, s = 0.38) was obtained. It appears that the arenes and the benzoates do not give the same ρ values nor does the kinetics of the benzoates oxidations represent a satisfactory Hamett correlation. Prior to the report of the pulse radiolysis investigation[15] both the charge-transfer spectra and by implication the ionization potentials of substituted arenes were reported to show a better correlation with σ^+ substituent constants than with σ constants.[17] More recently the correlation of ionization potentials with σ^+ constituent constants has been substantiated in a more extensive study by Kochi.[18]

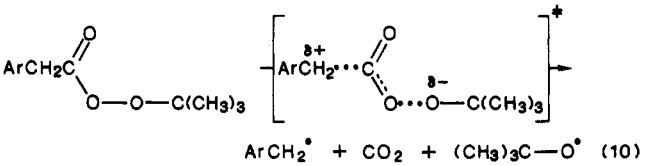
The suggestion that the decarboxylation of phenylacetic acid involved electron-transfer–side-chain fragmentation was in accord with the initially observed correlation[2] with σ^+, although it was not in agreement with the results obtained in the pulse radiolysis studies which show, for the arenes only, a correlation with σ and not with σ^+.[15] The observations that the rates of aliphatic decarboxylations werrr comparable to those found for the reaction of the aromatic carboxylates[2,14] are also in accord with the differences reported for the rates of reactions determined by pulse radiolysis, which measured differences of only 10^1–10^2 between the oxidation of aromatic and aliphatic carboxylates.[15,19–21] Carboxylates that on decarboxylation yield secondary radicals are found to undergo oxidation at rates 10^2–10^3 times slower than that of aromatic acids (see Table II). Since the pseudo-first-order rate constant for the reaction of the sulfate radical anion with hydroxyl (eq 9)[22]

$$SO_4^{\bullet-} + {}^-OH \rightarrow SO_4^{2-} + HO^{\bullet} \qquad (9)$$

is 5×10^7, the rate under our conditions is (5×10^7) $[HO^-] = 7.5 \times 10^6$ s^{-1} M^{-1}. Carboxylates whose rate constant for electron transfer to the sulfate radical anion are 10^7–10^8 will show a rate of decarboxylation $[RCO_2^-] \times 10^7$–10^8 = 1.5×10^6 to 10^7 which is competitive with the formation of hydroxyl radicals. Under these conditions the reactions of the slower carboxylates (last two entries in Table II) cannot be compared to those of the phenylacetic acids (k = 10^9 M^{-1} s^{-1})[15] since they no doubt react to a large extent by transfer to hydroxyl radical formed from reaction of the sulfate radical anion.

A reasonable possibility exists that two different decarboxylation mechanisms, one for aliphatic carboxylic acids and one for aromatic acids, could both show a simular structure–reactivity relationship. An analysis of the relative rates of decarboxylation of the salts of a series of aliphatic and aromatic acids sheds some light on this proposal.

The rates of oxidation of the anions of the series of substituted phenylacetic acids, Table I, showed a ρ value, −0.44, which correlated with σ^+. Both electron transfer from the carboxylate anion with concerted loss of carbon dioxide (Scheme II, eq 6) and the two-step oxidation (Scheme I, eq 4 and 5) can accommodate these observations. The magnitude of the selectivity for decarboxylation found in this work is consistent with the relative rates of decarboxylation reported for the thermolysis of a series of aryl-substituted *tert*-butyl phenylperacetates, ρ = −1.09 for a correlation with σ^+ substituent constants.[23] As in the perester deompositions, a 100% yield of CO_2 was formed as were products from the benzyl radical (bibenzyl, benzaldehyde, and toluene). Under the reaction conditions the major fraction of the benzyl radicals produced resulted in the formation of polymeric material (average M_r 1400) whose elemental analysis corresponded to the formula, $C_{72}H_{60}O_{17}S_1$. The rate correlation obtained for the perester decomposition was rationalized on the basis of a concerted decomposition (eq 10), not unlike the concerted persulfate decarboxylation transition state originally proposed[2] (Scheme II, eq 6).



$$ArCH_2^{\bullet} + CO_2 + (CH_3)_3C{-}O^{\bullet} \qquad (10)$$

As is seen in Table II, no clean division between the rates of decarboxylation of aromatic and nonaromatic substrates is observed. The largest difference is between aliphatic substrates that produce tertiary or secondary radicals. By analogy to the perester decompositions,[24,25] the acids that produce primary and secondary radicals are predicted to proceed by a two-step, nonconcerted, elecron transfer followed by a loss of CO_2. In the case of the persulfate-promoted decarboxylations these carboxylates most likely are oxidized by transfer to the hydroxy radical.

More substantive evidence concerning the electron-transfer mechanism is obtained by an examination of the relative rates of decarboxylation of a series of benzyl-substituted phenylacetic acids (see Table III).

Benzylic substitution by one or two phenyl groups (see Table III) increases the rate of oxidative decarboxylation, but only marginally. The relative rates of decomposition, by a concerted one-step fragmentation, of *tert*-butyl phenylperacetate/*tert*-butyl diphenylperacetate is reported[24,25] as 1:65. By analogy, if the persulfate-promoted oxidative decarboxylations of diphenylacetic and triphenylacetic acid, Table III, proceeded by an electron-transfer–fragmentation mechanism, a much larger difference in rates would have been expected. The marginal increase in rate most likely reflects the dependence upon the ionization potential of the aromatic anion radical and not the stability of the benzylic radical. In accord with this assumption, methyl substitution on the ring is more effective in stabilizing the transition state for electron transfer than benzylic methylation (see Tables II and III).

Electron transfer from the carboxylate anion to the sulfate radical anion takes place from the HOMO of the

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(16) The symmetry of the HOMO is discriptive of a substituted aromatic ring, making the evaluation of the effect of electron donation or withdrawal ambiguous for both ortho and meta substituents.[17] The σ and σ^+ values used are taken from Exner, O. *Advances in Linear Free Energy Relationships*; Chapman, N. B.; J., Shorter, Ed.; Plenium Press, New York; 1972.
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anion, be it aliphatic or aromatic, the ionization potential for the aromatic substrates being lower than those for the aliphatic anions. The ease of transfer for the aliphatic anions must also depend upon the degree of concerted carbon–CO_2 bond breaking, which in turn is determined by the structure of R•.

Experimental Section

Materials. The carboxylic acids used were either commercially available or were synthesized or purified by the methods given below. The commercially available carboxylic acids (the solids) were purified to constant melting points by repeated recrystallization from the appropriate solvents. Their purity was checked by titration and by comparison of their melting points with those given in the literature.

Isobutyric Acid (Aldrich). The acid was purified by distillation, bp 147 °C (690 mm) [lit.[26] bp 152–155 °C (760 mm)].

Cyclohexanecarboxylic Acid (Eastman Organic Chemicals). The acid was distilled under reduced pressure, bp 121–122 °C (15 mm) [lit.[27] bp 232 °C (760 mm)]. The acid solidified on cooling, mp 29–30 °C (lit.[27] mp 29 °C).

Pivalic Acid (Aldrich). The acid was distilled and gave a colorless liquid, bp 158 °C (694 mm) [lit.[28] bp 162–165 °C (760 mm)].

α-Methylphenylacetic Acid (Aldrich). The acid was purified by distillation using a Teflon spinning band. The colorless acid distilled at 132 °C (17 mm) [lit.[29] bp 155 °C (21 mm)]; n^{25}_D 1.5209 (lit.[26] n^{25}_D 1.5204).

Phenylacetic Acid-*1*-^{14}C. The carboxyl-labeled acid was obtained from the Tracerlab Division of International Chemical and Nuclear Corp. Waltham, MA. Purified phenylacetic acid (150 g) was dissolved in ethyl alcohol. To this solution was added 0.1 mg of phenylacetic acid-*1*-^{14}C. The alcohol was evaporated and the diluted radioactive acid was dried in a desiccator over phosphorus pentoxide. A weighed amount of the acid (0.0123 g) was dissolved in a scintillation solution and its radioactivity determined by using a scintillation counter. The scintillation solution was a mixture of 2-ethanolamine, ethylene glycol monomethyl ether, toluene, and the scintillator, diphenyloxazole (PPO). The acid had an activity of 859 cpm/mg.

***p*-Phenoxyphenylacetic Acid.** The acid was prepared from diphenyl ether by the following sequence of reactions.

Acetylation of Diphenyl Ether. The acetylation of diphenyl ether (45 g, 0.265 mol) by acetyl bromide (30 g, 0.322 mol) was performed by following a procedure by Kipper.[30] The crude *p*-phenoxyacetophenone was obtained in a yield of 30 g (33%) and was crystallized from ethyl alcohol, mp 45 °C (lit.[30] mp 45 °C).

Synthesis of *p*-Phenoxyphenylacetamide. The compound was prepared from *p*-phenoxyacetophenone (12 g; 0.057 mol) by the Willgerodt reaction following a procedure by Tomita and Hashimoto.[31] The crude *p*-phenoxyphenylacetamide (9 g) was purified by a silica gel chromatography (acetone, ethyl acetate) and recrystallized from ethyl alcohol, mp 172–173 °C (lit.[31] mp 172 °C).

Hydrolysis of *p*-Phenoxyphenylacetamide. The amide (2.5 g) was heated under reflux with 200 mL of 10% aqueous potassium hydroxide solution for 8 h and acidified with concentrated hydrochloric acid, and the precipitated acid was crystallized from *n*-pentane to yield *p*-phenoxyphenylacetic acid, mp 78 °C (lit.[31] mp 78 °C): NMR ($CDCl_3$) δ 2.90 (m, 9 H), 6.25 (s, 2 H), 11.9 (s, 1 H). The equivalent weight of the acid obtained by titration was 228.

Cyclohexene-3-carboxylic Acid. The acid was prepared by following a procedure of Boorman and Linstead.[32] The crude acid was purified by distillation, bp 122–123 °C (11 mm) [lit.[32] bp 120 °C (10 mm)]; n^{25}_D 1.4847 (lit.[32] n^{25}_D 1.4814). The equivalent weight, obtained by titration, was 236.

α-Methyl-*o*-tolylacetic Acid. The acid was prepared by the methylation of the methylene group of *o*-tolylacetic acid as described in the following sequence of reactions.

Esterification of *o*-Tolylacetic Acid. *o*-Tolylacetic acid (15 g, 0.1 mol) was mixed with 98% ethyl alcohol (50 mL) and concentrated hydrochloric acid (4 mL) was slowly added. The mixture was heated under reflux for 2 h, cooled, and poured into water, and the ester was extracted with ether. The ethereal solution was washed with water and then with 5% sodium carbonate solution and dried over anhydrous sodium sulfate, and the ether was removed. The yield of the ethyl ester was 16 g (80%).

Methylation of the Ethyl Ester of *o*-Tolylacetic Acid. A general procedure for the alkylation of the methylene group of ethyl phenylacetate reported by Kenyon, Kaiser, and Hauser[33] was followed. Hydrolysis of the methylated ester yielded the free acid: NMR ($CDCl_3$) δ 7.77 (s, 3 H), 6.45 (s, 2 H), 2.90 (s, 4 H), 11.9 (s, 1 H). The α-methyl-*o*-tolylacetic acid was purified by repeated crystallization from water–ethyl alcohol and then *n*-heptane, mp 93–94 °C (lit.[34] mp 92–94 °C). The equivalent weight obtained by titration was 164.

Anal. Calcd for $C_{10}H_{12}O_2$: C, 73.15; H, 7.36. Found: C, 73.14; H, 7.26.

α-Methyl-*p*-tolylacetic Acid. The acid was prepared from *p*-tolylacetic acid in the reaction sequence used for the synthesis of α-methyl-*o*-tolylacetic acid. The α-methyl-*p*-tolylacetic acid was crystallized from *n*-pentane. The purified acid, mp 40 °C (lit.[35] mp 39–40 °C); NMR (CCl_4) δ 12.40 (s, 1 H), 2.20 (d, 4 H), 8.62 (d, 3 H), 7.76 (s, 3 H), 6.54 (q, 1 H), had an equivalent weight, obtained by titration of 164.

Anal. Calcd for $C_{10}H_{12}O_2$: C, 73.15; H, 7.36. Found: C, 73.45; H, 6.99.

Potassium Peroxydisulfate. The commercially available potassium peroxydisulfate (Fisher, Certified, >99.6%) was used without any further purification.

Competitive Decarboxylation of the Labeled and Unlabeled Carboxylic Acids. Aliquots of aqueous solutions 1.32×10^{-2}–1.48×10^{-2} M in phenylacetic acid-*l*-^{14}C, 1.32×10^{-2}–1.48×10^{-2} M in the nonradioactive acids, about 5.8×10^{-2} M in potassium peroxydisulfate and about 0.16–0.18 M in potassium hydroxide, were placed in break seals cooled in ice, and the break seals were degassed and sealed.

The reaction vessels were thermostated at 74.3 ± 0.1°. At this temperature the half-life of potassium peroxydisulfate is 7 h.[36] The flasks were removed from the bath at intervals and quenched in ice water. The reaction was followed to about 50% completion, which was usually a period of approximately 1 h. One flask was kept as a blank, while two others were run to infinity.

The carbon dioxide was liberated by addition of sulfuric acid, added via a break seal. The reaction vessel was connected to a high vacuum line via a second break seal and the carbon dioxide was collected through a series of cold traps (two *n*-pentane (−130 °C) traps and a liquid nitrogen (−198 °C) trap). The carbon dioxide was distilled to a calibrated Toepler pump where its volume and pressure were measured. After it was measured, the carbon dioxide was absorbed in a scintillation solution and its radioactivity measured following a procedure by Jaffay and Alvarez.[13]

The trapping solution was a 1:2 mixture of 2-aminoethanol in ethylene glycol monomethyl ether. The scintillation solution was a 1:2 mixture of ethylene glycol monomethyl ether and toluene, to which was added the scintillator, 2,5-diphenyloxazole (PPO) (0.60 g/100 mL).

Upon the decarboxylation of phenylacetic acid-*l*-^{14}C alone, a linear relation was obtained between the amount of carbon dioxide measured at intervals of the decarboxylation reaction and the

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Table IV. Decarboxylation of Pivalic Acid (1.46×10^{-2} M) and Phenylacetic Acid-*1*-^{14}C (1.46×10^{-2} M) at 74.3 ± 0.10 °C[a]

time (min)	total CO_2 (mol × 10^4)	cpm	active CO_2 (mol × 10^4)	inactive CO_2 (mol × 10^4)	k^I/k^{II}
22.0	0.253	25173	0.214	0.039	0.165
28.0	0.328	32687	0.278	0.050	0.157
31.1	0.363	35916	0.306	0.057	0.160
35.5	0.427	42064	0.358	0.069	0.163
41.8	0.501	48742	0.414	0.087	0.169
70.0	0.757	72520	0.614	0.137	0.150
14 days	1.892	110500	0.965	0.927	
		average	k^I/k^{II} =	0.161 ± 0.005	

[a] Pivalic acid (0.985×10^{-4} mol); phenylacetic acid-*1*-^{14}C (0.953×10^{-4} mol); potassium peroxydisulfate (3.70×10^{-4} mol).

Table V. Products from the Reaction of Phenylacetic Acid (2.94×10^{-2} M) and Potassium Peroxydisulfate (5.70×10^{-2} M) at 74.3 × 0.10 °C[a]

product	mol × 10^4	% of phenylacetic acid
CO_2	5.73	100
toluene	0.057	1.49
benzaldehyde	0.178	4.42
bibenzyl	0.412	10.86
polymer (0.51 g)		

[a] Phenylacetic acid (5.73×10^{-4} mol, 0.078 g); potassium peroxydisulfate (11.13×10^{-4} mol).

number of counts/min (cpm) recorded on the scintillation counter. This linear relation was used to relate the activity observed to the amount of radioactive gas evolved in the competitive decarboxylation reactions.

The reactions when carried out to 20 kinetic half-lives gave quantitative yields of labeled and unlabeled carbon dioxide. The purity of the gas was confirmed by mass spectroscopic analysis.

A typical kinetic run is given in Table IV.

Product Analysis from the Decarboxylation of Phenylacetic Acid. Aqueous solutions which were 2.94×10^{-2} M in phenylacetic acid, 5.8×10^{-2} M in potassium hydroxide, and 3.80 × 10^{-2} in potassium peroxydisulfate were placed in break seals, degassed, sealed, and thermostated at 74.3 °C. The decarboxylation reactions were carried to infinity (20 h). After the completion of the reaction, the CO_2 was measured after acidification. The reaction mixture was saturated with potassium bromide, and the organic material was extracted repeatedly with ether. The ethereal solution was separated, leaving a yellow solid suspended in the aqueous layer.

Analysis of the ethereal solution was carried on by GLPC using a 10 ft × $^1/_8$ in. SE-30, 5% on 60/80 Chromosorb W column on a Varian Aerograph Model 600-D with a flame ionization detector. Freon-112 was added as an external standard.

The insoluble polymeric material suspended in the aqueous layer was separated by centrifugation and repeatedly washed with water, and the solid was dried over P_2O_5 at reduced pressure. The polymer was insoluble in the common organic solvents (*n*-pentane, benzene, toluene, carbon tetrachloride, methylene chloride, chloroform, methyl alcohol, and ethyl alcohol). It dissolved in dimethyl formamide and dimethyl sulfoxide. The average molecular weight of the polymer was 1400 and the elemental analysis showed, C, 63.26; H, 4.31; O, 20.19; S, 2.33. An approximate molecular formula for the polymer was calculated to be $C_{72}H_{60}O_{17}S_1$. The IR spectrum showed a strong hydroxyl band (3340 cm^{-1}) and a weak carboxyl band (1700 cm^{-1}). A typical analysis is shown in Table V.

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Registry No. *o*-$MeC_6H_4CH_2CO_2H$, 644-36-0; *p*-$MeC_6H_4CH_2CO_2H$, 622-47-9; *p*-$MeOC_6H_4CH_2CO_2H$, 104-01-8; *p*-$PhOC_6H_4CH_2CO_2H$, 6328-74-1; $PhCH_2CO_2H$, 103-82-2; *p*-$BrC_6H_4CH_2CO_2H$, 1878-68-8; *m*-$BrC_6H_4CH_2CO_2H$, 1878-67-7; *p*-$ClC_6H_4CH_2CO_2H$, 1878-66-6; *m*-$ClC_6H_4CH_2CO_2H$, 1878-65-5; *m*-$FC_6H_4CH_2CO_2H$, 4771-80-6; 1-cyclohexene-3-carboxylic acid, 4771-80-6; pivalic acid, 75-98-9; cyclohexanecarboxylic acid, 98-89-5; isobutyric acid, 79-31-2; ethyl *o*-tolylacetate, 40291-39-2; α-methyl-*o*-tolylacetic acid, 62835-95-4; α-methyl-*p*-tolylacetic acid, 938-94-3; potassium peroxydisulfate, 7727-21-1; diphenylacetic acid, 117-34-0; triphenylacetic acid, 595-91-5; mandelic acid, 90-64-2; α,α-dimethylphenylacetic acid, 826-55-1.

β-Lactam Annulation Using (Phenylthio)nitromethane

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2-[(Phenylthio)carbonyl]-1-azabicyclo[4.2.0]octan-8-one (**8a**), 2-[(phenylthio)carbonyl]-1-azabicyclo[3.2.0]-heptan-7-one (**10a**), 2-[(phenylthio)carbonyl]-4-oxa-1-azabicyclo[3.2.0]heptan-7-one (**19a**), 3,3-dimethyl-2-[(phenylthio)carbonyl]-4-oxa-1-azabicyclo[3.2.0]heptan-7-one (**19c**), and 3-[[(*tert*-butyldimethylsilyl)oxy]-methyl]-2-[(phenylthio)carbonyl]-5-oxa-1-azabicyclo[4.2.0]oct-2-en-8-one (**26**) were prepared in good overall yields from the monocyclic β-lactam aldehydes **6c**, **6g**, **18b**, **18f**, and **24d**. The key process in this novel annulation was the condensation reaction of the aldehydes **6c**, **6g**, **18b**, **18f**, and **24d** with (phenylthio)nitromethane (1) followed by cyclization of the resultant (*Z*)-nitroalkenes **6e**, **6i**, **18d**, **18h**, and **24f** with tetrabutylammonium fluoride followed by ozone. These studies unequivocally establish (phenylthio)nitromethane (1) as a versatile reagent for the construction of the carbapenam, carbacepham, oxapenam, and oxacephem frameworks. These units occur in diverse β-lactam antibiotics and β-lactamase inhibitors.

Recently we had occasion to study (phenylthio)nitromethane (**1**)[1,2] as a convenient reagent for the homologation of aldehydes to produce α-substituted phenylthio esters.[3] Thus, for example, acetaldehyde was reacted with **1**, catalyzed by potassium *tert*-butoxide in THF and *tert*-butyl alcohol, followed by dehydration with methanesulfonyl

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