cis- and *trans*-4-*t*-Butylcyclohexyl Acetates.—The commercially available mixture of *cis*- and *trans*-4-*t*-butylcyclohexanol (Matheson Chemical Co.) containing about 75% *trans*-alcohol was esterified by acetyl chloride in benzene–pyridine to give a mixture of *cis*- and *trans*-acetates, b.p. 104–105° (15 mm.).

Relative Rates of Pyrolysis.—The results for one run on *cis*- and *trans*-4-*t*-butylcyclohexyl acetate are typical of the method.

A mixture of approximately 70% of the acetates and 30% tetralin was made and analyzed by g.p.c. A peak area ratio of 3.13 was found for *trans* and *cis* esters and one of 2.27 for total esters and tetralin. One ml. of this mixture was pyrolyzed at 450° and the column cooled and washed with hexane. The products were similarly analyzed, giving ratios of 4.94 and 1.07, respectively. From the initial and final ester/tetralin ratios the conversion, 52.9%, could be calculated.

It may be shown that if the initial and final mole ratios of reactants, a_0/b_0 and a/b, respectively, and the conversion fraction, C, are known, then the ratio of the two first-order rate constants is given by

$$\frac{k_1}{k_2} = \frac{\log\left(\frac{1}{1-c} \times \frac{1+b_0/a_0}{1+b/a}\right)}{\log\left(\frac{1}{1-c} \times \frac{1+a_0/b_0}{1+a/b}\right)}$$

Substitution of the above values gives a value for k_{cis}/k_{trans} of 1.69.

Pyrolysis of *cis*- and *trans*-1-Methyl-4-*t*-butylcyclohexyl Acetates.—The *cis* ester (0.25 ml.) or melted *trans* ester (0.25 ml.) was passed at the rate of 0.3–0.4 ml./min. in a slow stream of nitrogen through the pyrolysis column at 400 ± 3°. The products were collected in a Dry Ice trap and the column was cooled and washed with 2 ml. of hexane. The crude pyrolysis mixture was analyzed by gas chromatography on a 3-meter column of Ucon LB 550X on Celite.

# Studies on the Mechanism of Peroxide Cleavage of Benzils[1a,b]

By Harold Kwart and Norman J. Wegemer

The kinetics of the cleavage of various substituted benzils were studied in 35% *t*-butyl alcohol–65% water at three different temperatures over a wide range of *p*H value. The reaction was found to be first order in each of the components, benzil and hydrogen peroxide, but its dependence on hydroxyl ion concentration was cleanly first order only in the *p*H range 7.4 to 10.7. Outside this range the rate of cleavage showed a much smaller variation with *p*H change. The cleavage of benzil with *t*-butyl hydroperoxide was also investigated. No upper *p*H limit to the range of hydroxyl ion dependence could be established in this case. Furthermore, the rate of cleavage with the alkyl hydroperoxide was only approximately one-ninetieth as great as with hydrogen peroxide. Peracetic acid, on the other hand, was almost two thousand times faster in the cleavage of benzil. Rate constants for the reaction of hydrogen peroxide with various *p*-substituted benzils could be correlated with Hammett σ-values but not with Brown σ+-values. The activation parameters of benzil cleavage and the substituent effects on the rates of cleavage were shown to be comparable to the related cases of cyanide ion-catalyzed cleavage and hydroxide ion-catalyzed rearrangement of benzils. The mechanism of cleavage under all these circumstances has been discussed and related to other observations in the literature.

Introduction

The reaction of simple ketones with hydrogen peroxide and hydroperoxide derivatives is the well known Baeyer–Villiger oxidation. The product of this oxidation is an ester (or lactone, in the case of cyclic ketones), the structure of which has been reported to depend on the relative migratory abilities of the substituent groups attached to the carbonyl group[2] or, when the substituent groups differ significantly in bulk, on conformational effects as well.[3] The over-all pattern of these and related reactions, where a peroxyester intermediate is clearly indicated, has been followed in a number of cases using labeled oxygen.[4-6] The results of these and other experiments reported[7-9]

strongly suggest a reaction transition state involving an ion pair intermediate in which the anion remains associated with the two cationic reaction centers (oxygen and carbon) during the entire course of rearrangement. Essentially, the rate-determining step in all these cases is the heterolytic breaking of the –O–O– bond in a peroxy intermediate[10] and the transition state,[3b] as illustrated for an aromatic ketone reaction, resembles a heterocyclic "phenonium ion."
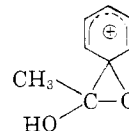
Transition State in Baeyer–Villiger and Related Reactions.—The transition state of the corresponding reaction where R is the oxygen of a carbonyl group, *viz.*, the cleavage of α-diketones with hydroperoxides, has received comparatively less attention. A number of authors have pointed out the analogy to the Baeyer–Villiger reaction and have on this basis postulated that the immediate reaction product is the anhydride, which then undergoes further solvolytic reaction in the medium to give the observed products.[11-14]

(1) (a) Presented in part before 138th A.C.S. Meeting, New York, N. Y., Sept., 1960. (b) Part of this work has been abstracted from the thesis of Norman J. Wegemer submitted in partial fulfillment of the requirements for the degree of Master of Science at the University of Delaware.

(2) W. von E. Doering and L. Speers, *J. Am. Chem. Soc.*, **72**, 5515 (1950).

(3) (a) S. L. Friess and N. Farnham, *ibid.*, **72**, 5518 (1950); (b) M. F. Hawthorne, W. D. Emmons and K. S. McCullum, *ibid.*, **80**, 6393 (1958).

(4) W. von E. Doering and E. Dorfman, *ibid.*, **75**, 5595 (1953).

(5) D. B. Denny and D. G. Denny, *ibid.*, **77**, 1706 (1955); **79**, 4806 (1957).

(6) C. A. Bunton, T. A. Lewis and D. R. Llewellyn, *J. Chem. Soc.*, 1226 (1956).

(7) R. Criegee, R. Kaspar and W. Dietrich, *Ann.*, **560**, 127, 135 (1948).

(8) P. D. Bartlett and J. L. Kice, *J. Am. Chem. Soc.*, **75**, 5591 (1953).

(9) H. L. Goering and A. C. Olson, *ibid.*, **75**, 5853 (1953).

(10) The corresponding homolytic reaction has been discussed in several review articles: (a) J. E. Leffler, *Chem. Revs.*, **45**, 385 (1949); (b) C. H. Hassall in R. Adams, *et al.*, *Organic Reactions*, Vol. IX, John Wiley and Sons, Inc., New York, 1957, p. 73.

Among earliest workers on the mechanism of peroxide reactions with electrophilic centers Weitz and Scheffer[15] had remarked on the similarity of α,β-unsaturated ketones and α-diketones in their behavior toward hydrogen peroxide. They proposed the intermediate formation of a cyclic peroxide via cis addition of the reagent peroxide oxygen atoms across the unsaturated carbon centers.

A more detailed analysis of the transition state for the hydroperoxidic cleavage of α-diketone and related ketones possessing an unsaturated α-carbon based on kinetic data has been undertaken by Bunton[16] and Bunton and Minkoff.[17] In the case of benzil cleavage by H_2O_2 the observation of base catalysis[16] has been inferred to suggest rate-determining attack by peroxide anion on the carbonyl, with little information to indicate the occurrence of intermediates of transient lifetime. Rate-determining attack by HOO^- has also been suggested[17] for the reaction with α,β-unsaturated ketones.

In addition to accounting for base catalysis, Bunton[16,18] justified the proposed rate-determining step by analogy to the addition of other nucleophilic anions, such as cyanide ion to carbonyl groups. We have noted, however, the contrary proposal of Kwart and Baevsky[19a] that the addition of cyanide ion was not the rate-determining step in the analogous cyanide ion-catalyzed cleavage of benzils and the driving force of reaction was not alone derived from the ability of this anion to add to the carbonyl. This was the incentive for a detailed kinetic study aimed at elucidating more of the mechanistic features of the peroxide cleavage of benzils, the results of which we are reporting here.

Experimental

A. Equipment.—Absorbancy measurements were made with a Beckman model B spectrophotometer at 375 mμ using matched Corex absorption cells of 5-cm. light path length.

Most pH measurements were made with a Beckman model G pH meter using a type 42 glass electrode and a fiber type reference electrode. At high pH values a Leeds and Northrup model 7664 pH indicator was used with a Std. 1199–30 glass electrode and a Std. 1199–31 reference electrode. A KinTel electronic galvanometer was used to expand the scale of this meter, as shown in Fig. 1. At pH 8.16 the L. and N. pH meter shows zero signal at the recorder terminals. The change is 71.4 microamps per pH unit. Considering the 20 ohm resistor in the terminals, this converts to 1.43 millivolts per pH unit. At any pH reading, this output is canceled by adjusting the output of the mercury cell so that the KinTel galvanometer reads zero. Any change in pH will then be observed as a deviation from zero on the galvanometer, the magnitude of the deviation depending on the sensitivity setting. The 10^{-1} setting gives
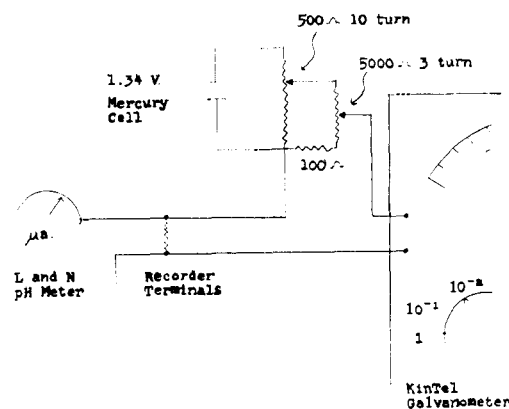
(11) M. A. F. Holleman, *Rec. trav. chim.*, **23**, 169 (1904).

(12) P. Karrer and F. Haab, *Helv. Chim. Acta*, **32**, 950 (1949).

(13) P. Karrer and L. Schneider, *ibid.*, **30**, 859 (1947).

(14) J. E. Leffler, *J. Org. Chem.*, **16**, 1785 (1951).

(15) E. Weitz and A. Scheffer, *Ber.*, **54**, 2327 (1921).

(16) C. A. Bunton, *Nature*, **163**, 444 (1949).

(17) C. A. Bunton and G. J. Minkoff, *J. Chem. Soc.*, 665 (1949).

(18) See also base catalysis of formation of hydroxyperoxides from HOOH and RCHO; C. E. Bunton, E. S. Halberstadt, A. J. Everett and G. L. Minkoff, *Nature*, **161**, 172 (1948).

(19) (a) H. Kwart and M. M. Baevsky, *J. Am. Chem. Soc.*, **80**, 580 (1958); (b) A. W. de R. v. Steveninck and E. C. Kooyman, *Rec. trav. chim.*, **79**, 414 (1960); (c) H. K. Hall, Jr., *J. Am. Chem. Soc.*, **77**, 5993 (1955).

Fig. 1.—Circuit diagram for expanding pH scale.

a full scale deflection for 0.07 pH unit, but requires precise grounding of all sources of interference. Most runs were made at the 10^{-2} setting, where full scale deflection is 0.7 pH units.

In the one instance where a correction for sodium ion error was necessary it was made according to Leeds and Northrup Nomograph 77–11–0–9.

B. Materials.—Benzil, Eastman Kodak Co. white label grade, was recrystallized from carbon tetrachloride as recommended by Gilman.[20] It was dried in vacuum overnight at 70°.

Substituted benzils were prepared according to reference 19a.

A sample of potassium 9,10-phenanthroquinone-3-sulfonic acid, supplied by Mr. Donald Bretzger, was used without further purification.

t-Butyl hydroperoxide, Lucidol 70, was distilled at 24° and 0.5 mm.[21] Analysis of the distillate showed 73.6% hydroperoxide.

Baker and Adamson reagent grade 30% hydrogen peroxide was used without further purification.

Peracetic acid (40%) from Buffalo Electrochem. Corp. was used, analyzed by iodometry.

C. Methods.—Water was used as the solvent for the anionic benzil, but all other studies were made in 35% t-butyl alcohol. This was found to be inert to the oxidizing agents used, although it will form the hydroperoxide under certain conditions of concentration and acid catalysis.[22] All reactions were run at 45°, except those necessary to determine the activation parameters.

The disappearance of benzil was followed spectrophotometrically below pH 5.5. At higher pH values the reaction was followed by titrating the benzoic acid formed with concentrated sodium hydroxide using a microburet. In either case this titration was performed to maintain constant pH without buffering the system.[23] Only those runs made at low pH and low concentration of peroxide were too slow to permit manual monitoring of the pH, and then sodium acetate–acetic acid or sodium acetate–hydrochloric acid buffers were used, prepared according to the concentrations ratios tabulated by Britton.[24]

All reactions were made using an excess of peroxide. A stock solution of benzil in spectro grade acetone was prepared such that a dilution of 1 to 100 would give the desired concentration. The standard procedure was to place 70 ml. of t-butyl alcohol, the necessary peroxide solution, and any additive desired in a 200-ml. volumetric flask. After dilution to volume in a 45° bath, the contents were emptied into a jacketed beaker which was maintained at the desired temperature by water circulating from a 45.7° bath. A magnetic stirrer was used for agitation. The pH meter

(20) H. Gilman, *Organic Syntheses*, Coll. Vol. I, John Wiley and Sons, Inc., New York, N. Y., 1932, p. 80.

(21) I. M. Kolthoff, E. J. Meehan, S. Bruckenstein and H. Minato, *Microchem. J.*, **4**, 33 (1960).

(22) A. G. Davies, R. V. Foster and A. M. White, *J. Chem. Soc.*, 1541 (1953).

(23) J. N. Brönsted, M. Kilpatrick and M. Kilpatrick, *J. Am. Chem. Soc.*, **51**, 428 (1929).

(24) H. T. S. Britton, "Hydrogen Ions," Vol. I, Chapman, Hall Ltd., London, 1942, pp. 301, 305.
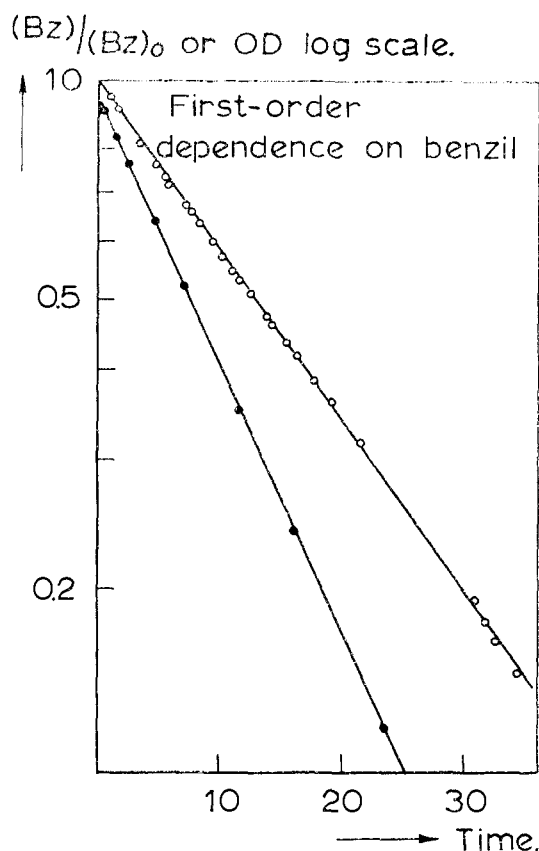
Fig. 2.—0.0024 M benzil in 35% t-BuOH at 45°: O, [Bz]/[Bz]$_0$ vs. min., pH 8.96, [HOOH]$_0$ 0.120 M, k' = 26.2 l.mole^{-1}hr.$^{-1}$; ●, optical density vs. hr., pH 2.1 [HOOH]$_0$ 0.757 M, k' = 0.113 l.mole^{-1}hr.$^{-1}$.

electrodes were inserted, and the pH adjusted to the desired value by careful addition of sodium hydroxide, acetic acid or hydrochloric acid. A sample was then withdrawn for peroxide analysis, and the benzil stock solution added to start the run.

Results and Discussion

A. Dependence of Rate on Peroxide and Benzil Concentration.—Initially, rate studies were made in an acid and basic pH range, respectively, to determine the order of reaction with respect to benzil and to ascertain at the same time whether this dependence on benzil could vary with acidity of the medium. Using an excess of hydrogen peroxide of 50- to 300-fold, we obtained good linear adherence to a first-order rate law in benzil, as typically illustrated by Fig. 2 for the low and high pH runs (e.g., 2.1, 8.96).

Rate studies were then carried out on the cleavage of benzil in the same medium as above (35% t-butyl alcohol, standard for almost all the rate studies being reported here) using five different concentrations of hydrogen peroxide. The data for a series of five typical runs are summarized in Table I. A plot of log k_{obs} vs. log (HOOH)$_0$ shows a least square slope of 1.09. This clearly establishes a first-order dependence of the rate on hydrogen peroxide.

Similar dependence of the rate on the first power of the concentration of t-butyl hydroperoxide and peracetic acid was demonstrated. From all these

RATE SUMMARY AS A FUNCTION OF INITIAL HYDROGEN PEROXIDE

(Bz)$_0$ = 0.0023 M, 45°, pH 2.1 ± 0.1

k_{obs}, hr.$^{-1}$	0.0237	.0507	.0675	.0854	.110
(HOOH)$_0$	0.236	.459	.624	.757	.945

measurements we infer the general rate law of the cleavage reaction: velocity of cleavage = k_2 (benzil)(peroxide).

B. Dependence of Rate on pH.—The rate of reaction of benzil with hydrogen peroxide was studied in unbuffered systems but at controlled pH over the entire range of pH values available for measurement by the methods used. Accurate measurements could not be made above pH 12.3 in this system because of the speed of the cleavage reaction and the increased rate of unimolecular decomposition of HOOH at high pH at the 45° reaction temperature.

The reaction was followed spectrophotometrically at pH values up to 5.5. Above this pH the studies were carried out by pH-Stat measurement (see Experimental). The values of k_{obs} were converted to values of k' (according to k_{obs} = k'-(HOOH)$_0$) by dividing by the average concentration of (HOOH) as determined by intermittent titration with standard permanganate of samples withdrawn during the course of a run. This average differed from (HOOH)$_0$ by a significant amount only at the very high pH values where the unimolecular decomposition of hydrogen peroxide becomes important at the reaction temperature.

The data listed in Table II show an almost 20,000-fold k' variation in the pH range of ca. 1 to 12.

TABLE II

DEPENDENCE OF HYDROGEN PEROXIDE CLEAVAGE RATE ON pH

(Bz)$_0$ = 0.0024 M, 45°, 35% t-BuOH

pH	k', l. mole^{-1} hr.$^{-1}$	k'calc[a]	pH	k', l. mole^{-1} hr.$^{-1}$	k'calc[a]
1.04	0.181	0.143	9.89	330	280
1.93	.188	.189	10.56	1210	1349
2.95	.238	.250	10.78	1690	..[b]
4.00	.341	.347	11.00	1890	..
5.02	.453	.463	11.29	1630	..
5.46	.525	.521	11.33	1950	..
6.09	.608	.656	11.65	2390	..
7.03	1.18	1.20	11.90	2080	..
7.93	4.10	4.47	12.06	2510	..
8.93	36.2	31.9	12.28	2000	..
8.96	26.6	33.7	12.30	3600	..
9.68	132	173			

[a] Calculated on the basis of empirical eq. A. [b] These points depart from line; see Fig. 4.

These data can be fitted quite well to the empirical equation

$$k'_{calc} = 5.10(OH)^{1/4} + 8.98 \times 10^5(OH) \quad (A)$$

considering the enormous range of hydroxide ion concentration encompassed (ca. 10^{13}-fold variation). Apparently there are two rate-determining factors, one of which is pH dependent and the other almost independent of the pH. This can readily be perceived on a plot of log k' vs. pH as given by Fig. 3 which shows a central region with a slope of exactly

unity where the familiar catalytic equations 1 and 2 can be said to apply precisely.

$$k_0 = k'_0 + k_c(\text{cat}) \tag{1}$$

where k_c is the specific catalytic rate coefficient of any catalyst of concentration (cat). Thus, for hydroxide ion catalysis we have

$$\log (k' - k'_0) = \log k_{OH} - pOH \tag{2}$$

The value of k_{OH} (8.98 × 10⁵ l.² mole⁻² hr.⁻¹) was computed by determining first a value for k'_0, assuming this was a pH independent term, by extrapolation of the linear portions of eq. 1 plots. Clearly, however, these k'_0 values do not apply except as a first approximation in the low pH region of the rate $-p$H profile, since here, as stated before (eq. A), this term is slightly dependent on the hydroxide ion concentration, $(OH)^{1/8}$.

The corresponding dependence of the rate of cleavage of benzil with t-butyl hydroperoxide on the pH could be studied over a pH range of $ca.$ 1 to 13, without encountering any interference from significant unimolecular decomposition of the peroxide at high pH values. These data are listed in Table III and plotted for the same variables, log

TABLE III

DEPENDENCE OF t-BUTYL HYDROPEROXIDE CLEAVAGE RATE ON pH

$(Bz)_0 = 0.0024M,\ 45°,\ 35\%\ t\text{-BuOH}$

pH	k', l. mole⁻¹ hr.⁻¹	pH	k', l. mole⁻¹ hr.⁻¹
1.71	0.0136	10.99	58.4
1.91	.0093	11.51	143
4.46	.0199	12.09	420
5.54	.0346	12.55	1360
6.18	.0395	12.58	1530
8.80	.327	12.60	1420
9.79	2.87	13.0–13.2[a]	10⁴

[a] Corrected for sodium ion error; computed: k_{OH} = 0.109 × 10⁵ l.² mole⁻² hr.⁻¹.

k' vs. pH, in Fig. 3. Again, it will be noted that this plot possesses a linear region with a slope of unity starting from pH values a little beyond 7.5 and extending as far into the pH range as it was possible to measure with our instrumentation. The value of k_{OH} determined in the same manner as for the corresponding value for hydrogen peroxide (according to eq. 2) thus permits a comparison of the relative effectiveness of the peroxidic cleavage reagents. Apparently hydrogen peroxide is almost ninety times as reactive as t-butyl hydroperoxide in the cleavage of α-diketones.

Some experimental difficulty was encountered in measurement of the rate dependence on pH for the benzil cleavage with peracetic acid due to the extremely rapid reaction at 45°. Nonetheless, the curve that could be obtained, described by fairly reliable data, for the plot of log k' vs. pH showed a linear portion of unit slope (as with the other peroxide cleavage reagents) in the region of pH 7–8 (beyond pH 8 it was not possible to obtain reliable rates). Again, the value of log k_{OH} was determined according to eq. 2 and the value of k_{OH} deduced in this fashion was 1.7 × 10⁹ l.² mole⁻² hr.⁻¹. Evidently per acetic acid is about two thousand times as effective as hydrogen peroxide and nearly two hundred thousand times more ef-
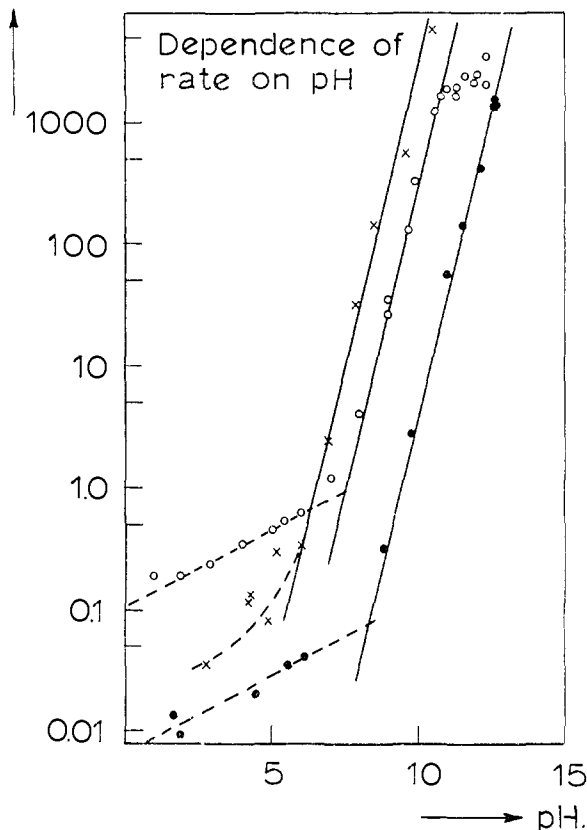
k' (log scale).

Fig. 3.—Log k' vs. pH. Solid lines have unit slope, dashed straight lines have a slope of one-eighth: O, HOOH cleavage, 0.0024 M benzil in 35% t-BuOH, 45°, k_{OH}- = 8.98 × 10⁴ l.² mole⁻²hr.⁻¹; ●, t-BuOOH cleavage, 0.0024 M benzil in 35% t-BuOH, 45°, k_{OH}- = 1.09 × 10⁴ l.² mole⁻² hr.⁻¹; +, HOOH cleavage of phenanthroquinone-3-sulfonate anion in water at 45° and constant ionic strength, k_{OH}- = 6.8 × 10⁶ l.² mole⁻²hr.⁻¹.

fective than t-butyl hydroperoxide in the mechanism of cleavage of α-diketones in the central pH region characterized above for all three peroxidic cleavage reagents.

The curves in Figs. 3 also possess other features that merit comparison. For one, as stated previously, the unit slope portion of the t-butyl hydroperoxide graph is linear as far out on the pH axis as it was possible to measure. A break from linearity and a tailing off of the rate with increasing pH beyond 10.3 is quite noticeable in the hydrogen peroxide plot. The significance of this observation will be discussed at greater length in a later section of this report. The possible origin of this difference in behavior of the alkyl hydroperoxide and hydrogen peroxide, however, may be correlated with the fact that the latter is potentially a dibasic acid whereas the former has only a single proton available for transfer to base. This would also mean that any complex anion formed between hydroperoxy anion and benzil is an acid by virtue of possessing an additional proton, but the corresponding complex formed from t-butylperoxy-anion does not have this potential. Furthermore,
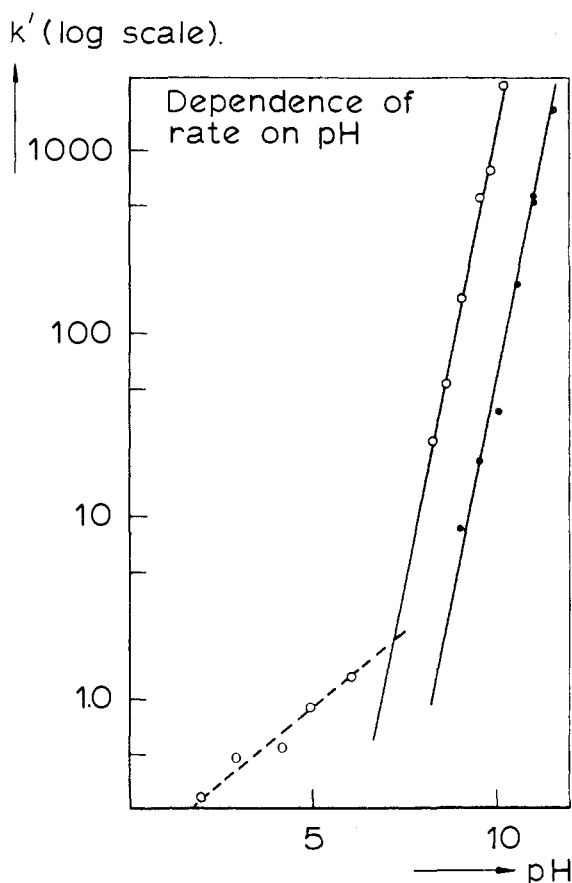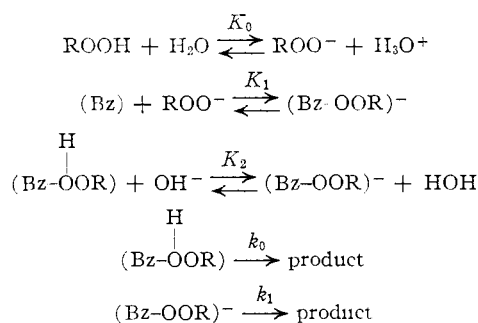
k' (log scale).

Fig. 4.—log k' vs. pH Solid lines have unit slope, dashed line has a slope of one-sixth; O, HOOH cleavage of 4-chlorobenzil in 35% t-BuOH at 45°, $k_{OH^-} = 3.50 \times 10^5$ l.² mole⁻² hr. ⁻¹; ●, HOOH cleavage of 4-methoxybenzil in 35% t-BuOH at 45°, $k_{OH^-} = 1.41 \times 10^5$ l.² mole⁻² hr.⁻¹.
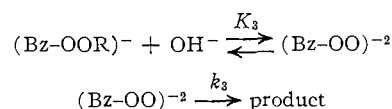
the acidity of the complex anion formed with hydroperoxy anion would be strongly dependent on substitution on the nuclei of the (original) benzil. It is for this reason that we observe the high pH breakpoint in the case of benzil occurs at pH 10.8, in the case of 4-chloro-4'-methoxy-(approx.) at pH 11.4 and for 4-methoxybenzil (Fig. 4), where the single substituent is strongly electron releasing, there is no indication of a high pH break-point even at 11.5.

A second detail in the rate–pH profiles in Fig. 3 that might be compared comprises that portion of the curves prior to the attainment of unit slope. The slope of the line along which the rate rises with increasing pH is slightly greater for hydrogen peroxide. Furthermore, the pH at which the break in the curve occurs is more acid by almost one pH unit in the case of hydrogen peroxide (ca. pH 7.4 vs. pH 8.2). It is interesting to note, too, that the (low) pH break-point in the curve for peracetic acid also occurs at a lower pH (ca. 6.9) than for t-butyl hydroperoxide.

It is possible to account for much of the detail of the rate–pH profile of these typical peroxide cleavage reactions, as well as the intimate points of difference discussed above, by the following gross mechanistic course of the reaction.

$$ROOH + H_2O \underset{}{\overset{K_0}{\rightleftharpoons}} ROO^- + H_3O^+$$

$$(Bz) + ROO^- \underset{}{\overset{K_1}{\rightleftharpoons}} (Bz\text{-}OOR)^-$$

$$\underset{(Bz\text{-}OOR)}{\overset{H}{|}} + OH^- \underset{}{\overset{K_2}{\rightleftharpoons}} (Bz\text{-}OOR)^- + HOH$$

$$\underset{(Bz\text{-}OOR)}{\overset{H}{|}} \xrightarrow{k_0} product$$

$$(Bz\text{-}OOR)^- \xrightarrow{k_1} product$$

where R = H

$$(Bz\text{-}OOR)^- + OH^- \underset{}{\overset{K_3}{\rightleftharpoons}} (Bz\text{-}OO)^{-2}$$

$$(Bz\text{-}OO)^{-2} \xrightarrow{k_3} product$$

where k_3 is smaller than k_1 and sometimes may be neglected. The rate expression derivable from this reaction course can be obtained readily

$$-d(Bz)/dt = k_0(Bz\overset{H}{\underset{|}{-}}OOR) + k_1(Bz\text{-}OOR)^- \quad (3)$$

Since

$$(Bz\text{-}OOR)^- = \frac{K_0K_1 (Bz)(ROOH)}{(H_3O^+)}, \text{ and}$$

$$\underset{(Bz\text{-}OOR)}{\overset{H}{|}} = \frac{(Bz\text{-}OOR)^-}{K_2(OH^-)} = \frac{K_0K_1 (Bz)(ROOH)}{K_2K_w}, \text{ then}$$

$$\frac{-d(Bz)}{dt} = \frac{k_0K_0K_1(Bz)(ROOH)}{K_wK_2} + \frac{k_1K_0K_1(Bz)(ROOH)}{(H_3O^+)} \quad (4)$$

Therefore

$$\frac{-d(Bz)}{(Bz)(ROOH)dt} = k' = \frac{k_0K_0K_1}{K_wK_2} + \frac{k_1K_0K_1(OH^-)}{K_w} \quad (5)$$

The last equation, (5), expresses a linear relationship between k' and (OH⁻). Most of our data (above) has been presented on log scaled plots since both k' and (OH⁻) varied over many orders of magnitude. The unit slope region of the plots in Fig. 3 corresponds to the condition where the first term on the right in eq. 5 is negligible. This should occur when the predominant fraction of the total rate is attributable to the decomposition of the anionic peroxy complex (Bz–OOR)⁻ and comparatively little of the uncharged complex is present or decomposed in solution: hence, eq. 5 reduces to (in the unit slope region)

$$k' = \frac{k_1K_0K_1}{K_w}(OH^-)$$

A comparison with the empirical eq. 1, therefore, justifies the relationships

$$k_{OH} = \frac{k_1K_0K_1}{K_w} \text{ and } k_0' = \frac{k_0K_0K_1}{K_wK_2} \quad (6)$$

The parameter k'_0, evaluated from the data and plots in Fig. 3 for hydrogen peroxide ($k'_0 = 108 \times 10^{-3}$ and t-butyl hydroperoxide ($k'_0 = 6.5 \times 10^{-3}$), is therefore a composite measure of the acidity of the hydroperoxide (K_0), the parallel acidity of its complex with benzil (K_2), and the product of the nucleophilicity of the peroxy anion (K_1), and the ease of decomposition of the neutral complex with benzil (k_0). Since the difference in acidity of

t-butyl hydroperoxide and hydrogen peroxide is almost a factor of ten and the differences in the k'_0 values is also almost a factor of the same magnitude, it appears that differences in the portion of the rate–pH profile curves preceding the low pH break-point *for any given benzil* are largely determined by differences in acid dissociation of the hydroperoxide.

However, by changing the structure and nature of the benzil for a given hydroperoxide cleavage agent some change in k'_0 and in the portion of the curve preceding the low pH break-point can be expected. This can be viewed for the cleavage of 4-chlorobenzil ($k'_0 = 131 \times 10^{-3}$) presented in Fig. 4 for comparison with the rate –pH profile of the unsubstituted compound in Fig. 3.

It is possible to perceive in another way the factors which control the kinetics of reaction. Thus, examination of eq. (5) shows that the low pH break-point (unit slope rate) is attained when k_0/K_2 approximately equals $k_1(OH^-)$, occurring at pH 7.4 for hydrogen peroxide and 8.2 for t-butyl hydroperoxide. If we make the reasonable assumption that the k_0/K_2 term, expressing a ratio of rate and equilibrium factors in the neutral peroxy complex, will remain relatively unchanged with changes in the structure of the peroxidic cleavage reagent, then we see the low pH break-point is a reflection of the magnitude of k_1, a kinetic measure of the instability of the particular anionic peroxy complex. In agreement with expectation we find that k_1 for the hydrogen peroxide derived complex is much larger than the corresponding t-butyl hydroperoxide and smaller than for the peracetic.

We may observe, also, from Fig. 3 that the low pH break-point varies with the structure of the benzil. For the phenanthraquinone sulfonate, an α-diketone substituted by strongly electron-attracting groups, this pH value is about 6, implying that electronegative substituents increase the magnitude of k_1 and the ease of decomposition of the peroxy complex. Conversely, electron-releasing substituents decrease k_1, as witness Fig. 4 where the p-OCH$_3$ substituent has raised the low pH break-point. Further correlation of the cleavage rate and peroxide structure will be undertaken in a later section of this report.

C. The Effect of Added Neutral Salts.—To test the response of the cleavage rate to neutral electrolyte, the ionic strength of the solution in several runs was adjusted by means of added sodium perchlorate to a value equal to about twenty times that expected on the basis of sodium benzoate formed at completion of reaction. It was noted that such runs did not deviate from the line drawn through points determined at minimum ionic strength. Other data of rates taken at two different ionic strength levels, $\mu = 0.05$ and 0.10, also bear out the conclusion that the rate of cleavage of benzil with hydrogen peroxide in acidic or basic solution shows little or no response to added neutral salts.[25,27]

D. The Effect of Temperature on Rate of Cleavage. The Activation Parameters.—Measurements of the rates of cleavage of benzil with hydrogen peroxide were also made, in 35% t-BuOH medium, at 25° and 1°. These results are listed in Table IV. The energy (ΔH^*) and entropy (ΔS^*)

TABLE IV

DATA FOR ACTIVATION PARAMETER DETERMINATIONS[a]

	25°		1°
pH	k', l. mole^{-1} hr.$^{-1}$	pH	k', l. mole^{-1} hr.$^{-1}$
7.01	1.36	7.01	0.417
7.53	2.33	8.15	0.828
8.18	2.10	9.00	4.19
9.21	9.34	9.00	6.03
9.76	24.4	9.93	6.88
10.21	64.7	11.15	28.7
10.60	178	11.29	58.5
11.01	566	11.38	65.4
11.61	947	11.98	290
12.07	1120		
12.10	1490		
12.48	2000		
12.75	872		

$k_{OH} = 4.73 \times 10^5$ l.^{2} mole^{-2} hr.$^{-1}$ using $pK_w = 14.00$　　　$k_{OH} = 2.29 \times 10^5$ l.^{2} mole^{-2} hr.$^{-1}$ using $pK_w = 14.902$

[a] The initial concentration of benzil used in all runs at both temperatures was 0.0024 M.

of activation were computed from these data and from the value of $k_{OH} = 8.98 \times 10^5$ obtained previously by measurements at 45°, using the relation

$$\ln \left(\frac{h k_{OH^-}}{kT} \right) = \frac{\Delta S^*}{R} - \frac{\Delta H^*}{RT} \qquad (7)$$

where the various constants have the magnitudes and significance designated in Laidler.[28] The slope and intercepts of the straight line plotted on the scales $\ln (h\, k_{OH}^-/kT)$ vs. $1/T$ according to eq. 7 were evaluated by the method of least squares and these results are listed in Table V.

The values of the activation parameters are quite close to those found by Kwart and Baevsky[19a] for the analogous reaction of benzil cleavage by cyanide ion: see Table V. These observations, therefore, tend to strengthen the analogy in mechanism of the two reactions, first suspected by Bunton.[16] On the basis of the earlier work, again, we

actly the response to neutral salts predicted by the Brönsted "kinetic" salt equation,[26,27] log $k^0 + A'\, Z_A Z_B \mu^{1/2}$ where $A' = 1.06$ in water solution at 45° and Z_A and Z_B are the usual symbols designating the charge on the reacting species entering the reaction transition state. Some of the data we obtained on the cleavage of potassium 9,10-phenanthraquinone-3-sulfonate with HOOH in the presence of varying amounts of sodium perchlorate to control ionic strength were plotted accordingly. It seems apparent, despite some scatter due to difficulties in measurements of very fast reaction rates, that the adherence to the Brönsted kinetic salt relation approached expectation (slope = 1.2 with a correlation coefficient of 0.95). This is to say, while the transition state of the unsubstituted benzil comprises a complex derived from the reaction of a neutral molecule with a negative ion, the corresponding transition state for the anionic α-diketone reagent is clearly the result of a complex formed from interaction of two negatively charged reagents.

(25) A somewhat more indirect test of the occurrence of a kinetic salt effect in the cleavage reaction of neutral, unsubstituted benzil is to determine the exact magnitude of the effect in the case of an anionic α-diketone. It would be predicted in this case that if the neutral benzil showed zero salt effect, the anionic reagent would show ex-

(26) For a discussion of this and other pertinent relationships see A. A. Frost and R. G. Pearson, "Kinetic and Mechanism," John Wiley and Sons, Inc., New York, N. Y., 1953, pp. 138–139, and ref. 27.

(27) J. E. Leffler and E. Grunwald in "Technique of Organic Chemistry," Vol. VIII, A. Weissberger, Ed., Interscience Publishers, Inc., New York, N. Y., 1953, p. 333.

(28) K. J. Laidler, "Chemical Kinetics," McGraw-Hill Book Co., Inc., New York, N. Y., 1950, p. 75.

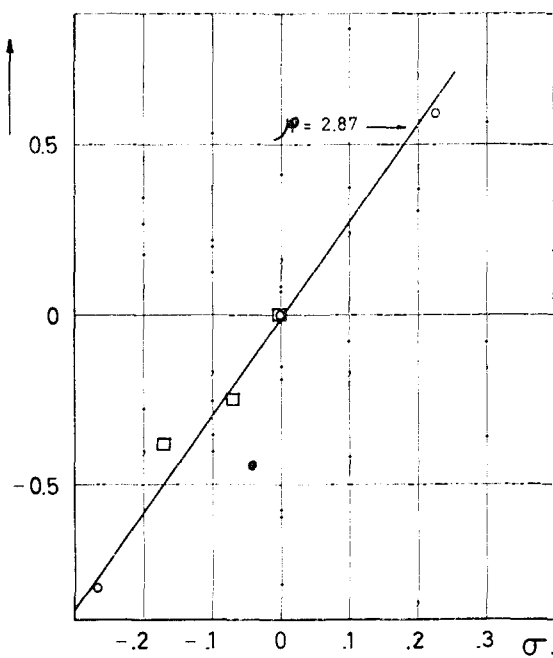$$\log{}^k/_{k_0} = \sigma_1\rho_1 + \sigma_2\rho_2 = \rho_{av}\,(\sigma_1 + \sigma_2).$$

Fig. 5.—Hammett plot: O, monosubstituted benzils–HOO⁻ cleavage; □, monosubstituted benzilic rearrangement; ●, disubstituted benzil–HOO⁻ cleavage.

may interpret our results to signify the formation of a relatively stable complex between benzil and hydrogen peroxide in a rapid equilibrium. The transition state attained during the decomposition of this complex apparently involves a high degree of order and the various bond cleavages must be de-

TABLE V

THE ACTIVATION PARAMETERS COMPARED

	Benzil + HOO⁻	Benzil + CN⁻	Acid cat. decompn. $\overset{\text{R}}{\underset{\text{XC}_6\text{H}_4\text{COOH}}{|}}$	Nucl. disp.[a] $\overset{\text{R}}{\underset{\text{O}}{\overset{|}{\underset{||}{\text{OH}^- + \text{CCl}}}}}$
ΔH^*	4.4	4.7	22	14 → 19
ΔS^*	−34.0	−37.6	−15	−7 → −14
Ref.	Present work	19a	19b	19c

[a] For evidence of intermediate addition complex formation here see C. A. Bunton, T. A. Lewis and D. R. Llewellyn, *Chemistry & Industry*, **73**, 1154 (1954), and V. Gold, J. Hilton and E. G. Jefferson, *J. Chem. Soc.*, 2756 (1954).

picted as occurring with a good deal of concertedness, judged from the magnitude and sign of ΔS^*. Other reactions analogous to the Baeyer–Villiger peroxidic oxidations have also been shown to involve the almost simultaneous migration of several bonding pairs in the transition state[5,7,9] and the large, negative activation entropy exhibited by the peroxide cleavage of benzil is consistent with this interpretation. Table V illustrates data for a typical (and possibly analogous) bond-breaking mechanism of peroxide decomposition[19b] ($\rho = -4.6$) and a typical bond-making mechanism[19c] ($\rho = 2$ to 3). The disparities in the activation parameters of these latter reactions when compared

with both CN⁻ and HOO⁻ cleavage of benzils emphasize, once again, that the cleavage mechanisms do not occur by either purely bond-making or bond-breaking reaction paths.

E. **The Substituent Effect.**—It can be readily deduced from eq. 6 that for a series of cleavage reactions with hydrogen peroxide where the zero superscript refers to the parameter characteristic

$$\log\left(\frac{k_{OH}}{k_{OH}{}^0}\right) = \log\left(\frac{k_1}{k_1{}^0}\right) + \log\left(\frac{K_1}{K_1{}^0}\right) = \sigma\rho \quad (8)$$

of the unsubstituted benzil, the other designates the corresponding value for a *m*- or *p*-substituted benzil, and the term on the extreme right is the familiar product of the Hammett linear free energy relationship.[29] On this basis, therefore, the ρ-value of the over-all reaction is composed of terms representing the sum of the equilibrium formation of complex (K_1) and its rate-determining decomposition (k_1) to product. Of course, this analysis does not take into consideration the fact that the complex is formed at one of the carbonyl groups (K_1) whereas bond migration involves the second carbonyl group as well. Consequently, a substituent may regulate the ease and position of complex formation but may have a different effect on the complex decomposition than the substituent in the second benzene ring and *vice versa*. Nonetheless, there appear to be two factors that afford some basis for correlating mechanistic features of the reaction with the magnitude and sign of the reaction constant ρ even under these somewhat complicated circumstances. For one, the substituents effects have been studied for the analogously complicated cleavage of benzils with cyanide ion so that comparison with the presently reported results on the peroxide cleavage could be informative. Our objective here was to evaluate if possible the relative importance of the substituent effect in complex formation and decomposition.

Table VI contains a listing of the data plotted as a linear free energy function in Fig. 5. Consider-

TABLE VI

CORRELATION OF k_{OH} WITH HAMMETT'S VALUES

Compound	k_{OH} (10⁻⁵)	$\log (k_{OH}/k_{OH}{}^0)$	$(\sigma_1 + \sigma_1)$
4-Chlorobenzil	35.0	0.591	0.227
Benzil	8.98	0	0
4-Methoxybenzil	1.41	−0.804	−0.268

$$\rho = 2.87 \pm 0.13$$

| 4-Methoxyl-4'-chlorobenzil | 3.98 | −0.443 | −0.041 |

TABLE VII

SOME SUBSTITUENT EFFECTS ON THE RATES OF BENZILIC ACID REARRANGEMENT[a]

Benzil	k_{OH}	$\log (k_{OH}/k_{OH}{}^0)$	$(\sigma_1 + \sigma_2)$
4-Methyl-	0.206	−0.375	−0.170
Unsubstd.	.488	0	0
3-Methyl-	.273	−0.252	−0.069
4,4'-Diamino-	.023	−1.326	−1.32

[a] Computed from data reported in ref. 33.

ing only monosubstituted benzils, a good straight line intersects the three points observed correspond-

(29) For a general review of this subject see H. H. Jaffé, *Chem. Revs.*, **53**, 191 (1953).

ing to a reaction constant ($\rho = 2.87$) very much the same as that observed for the cyanide cleavage ($\rho = 3.45$). Also contributing to the similarity of the two reactions is the fact that using Brown's σ^+-values[30] the correlation is entirely destroyed; that is, the relationship doesn't begin to approach linearity. This latter result can probably be interpreted for the peroxide cleavage as it was for cyanide: namely, the absence of a strong electromeric requirement to assist the breaking and migration of the central $-C-C-$ bond in the transition state.

We may reasonably assume that a m- or p-substituent on the ring can have several effects on the rate of peroxide cleavage of benzil. For example, if it is an electron-attracting substituent it will both foster complex formation on the adjacent carbonyl and enhance decomposition of the complex by exerting a "pull" on the central bonding pair.[31] If it is electron releasing it will retard complex formation (relative to the unsubstituted ring) at the adjacent carbonyl; but in the resulting complex, occurring presumably at the remote carbonyl (with lesser ease), it can aid delocalization of the bonding pair or, in other words, exert a "push" contributing to the breaking of the central $-C-C-$ bond. If the situation is exactly analogous to the cyanide ion-catalyzed cleavage,[17] then we should observe the over-all ρ-constant to be an average value of the almost equal contributions of both the equilibrium complex-forming step: hence we should expect to observe good linear adherence to the Hammett additivity relation.[17,31]

$$\log \left(\frac{k_{OH}}{k_{OH}{}^0} \right) = \rho_{av} (\sigma_1 + \sigma_2) \qquad (9)$$

We tested this by determining the rates in a disubstituted case where one ring is substituted by the electron-releasing p-OCH$_3$ and the other by the electron-attracting p-chloro. In the cyanide ion-catalyzed cleavage this case fell exactly on the line[19a] of the plot of eq. 9. However, the plot in Fig. 5 demonstrates conclusively that the additivity relation eq. 9 does not apply to the peroxide cleavage and most probably the two reaction steps do not make nearly the same contributions to the determination of the over-all ρ-value, as in the cyanide cleavage. As can be perceived in Fig. 5 the rate for 4-methoxy-4'-chlorobenzil cleavage with peroxide should have been almost twice as great as observed in order for eq. 9 to apply.

The reaction of benzils with hydroxide ion, known as the benzilic acid rearrangement, has also been suggested to occur *via* a rapidly established equilibrium complex. The rate-determining decomposition of the complex does not involve the breaking of the central $-C-C-$ bond but rather the migration of the group from the tetrahedral, complexed carbon to the adjacent, unsaturated center.[32] The peroxide cleavage and benzilic

rearrangement of α-diketones are formally quite similar in that electron-attracting substituents enhance the rate, although in the former the migration or delocalization occurs to an electron-deficient oxygen center rather than to an unsaturated carbon center, as in the latter.[33] Some quantitative data on substituent effects in the benzilic acid rearrangement are listed in Table VII and also plotted in Fig. 5 for comparison to the peroxide cleavage. Evidently the ρ-values are similar in both sign and magnitude, but once again the additivity relation eq. 9 does not apply. The p,p'-diamino point is so far removed from the line of the monosubstituted point it couldn't be readily plotted on the same graph no matter what σ-value is used (among several available) for the dimethylamino group. Apparently, the equilibrium complex-forming step and the complex-decomposition step also do not contribute equally to determining the over-all reaction constant.

On the other hand, the Baeyer–Villiger reaction of ketones with peracids, which has been considered[34] to be formally related to the α-diketone peroxy cleavage, is clearly different in the nature of its driving force. The conclusion that the rate-determining step in the Baeyer–Villiger is involved with the acid-catalyzed heterolysis of the $-O-O-$ bond[35] is supported also by the observation[36] that electron-releasing substituents enhance this process ($\rho \cong -1.4$). An analogous transition state for peroxy cleavage of α-diketones would be
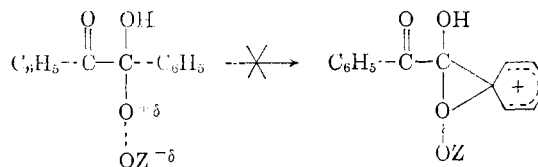
quite untenable since it would predict incorrectly a preference for phenyl rather than the observed benzoyl migration. (See reference 39 for evidence of the formation of anhydride, the product of benzoyl migration in the cleavage reaction.)

The Mechanism of Cleavage.—In all the preceding mechanistic discussion in connection with derivation of the rate law governing the cleavage of benzils we have assumed the equilibrium formation of a complex, which then underwent rate-determining decomposition to product. However, Bunton,[16] the most recent author to attempt formulation of the mechanism of this reaction, has advanced the idea that base catalysis of cleavage must be accounted for by assuming a rate-determining attack by the extremely strong nucleophile on the benzil carbonyl. Decomposition of the resulting short-lived complex was assumed to occur *via* a four-membered transition state
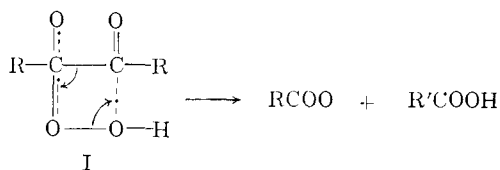
(30) H. C. Brown and Y. Okamoto, *J. Am. Chem. Soc.*, **79**, 1913 (1957).

(31) Compare with the effects of substituents in C. G. Swain, W. H. Stockmayer and J. T. Clarke, *ibid.*, **72**, 5426 (1950), and in ref. (19a).

(32) For a discussion of this mechanism of benzilic acid rearrangement see F. H. Westheimer, *ibid.*, **58**, 2209 (1936), and I. Roberts and H. C. Urey, *ibid.*, **60**, 880 (1938), as well as H. Kwart and I. M. Sarasohn, *ibid.*, **83**, 909 (1961).

(33) M. T. Clark, E. C. Hendley and O. K. Neville, *ibid.*, **77**, 3280 (1955).

(34) Y. K. Syrkin and J. I. Moiseev, *Russ. Chem. Revs. (Uspekhi Khim.)*, **29**, 193 (1960).

(35) M. F. Hawthorne and W. D. Emmons, *J. Am. Chem. Soc.*, **80**, 6398 (1958); see, however, Y. Yukawa and T. Yokayama, *Mem. Inst. Sci. Ind. Res. Osaka Univ.*, **13**, 171 (1956).

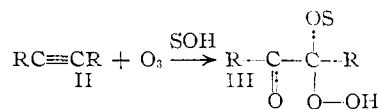(36) S. L. Friess and A. H. Soloway, *J. Am. Chem. Soc.*, **73**, 3968 (1951).

$$R-\underset{\underset{O}{\overset{O}{\|}}}{\overset{\overset{O}{\|}}{C}}-\underset{\overset{O-O-H}{}}{\overset{O}{\|}}{C}-R \longrightarrow RCOO + R'COOH$$

I

A somewhat similar suggestion involving a 6-membered transition state for decomposition was preferred by Syrkin and Moiseev.[34]

II
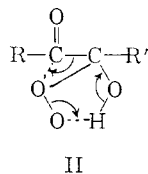
There are, however, several observations that we have discussed above which do not accord with Bunton's suggestions or those of the Russian authors regarding the rate-determining step. For one, the relative rate order observed in which peracetate is several thousand times more effective in the cleavage reaction than hydroperoxy anion and the latter, in turn, affords a much faster (ca. 100 x) cleavage than t-butylperoxy anion, demonstrates that the driving force of reaction is not alone derived from the bond making ability of the peranion. Peracetate ion is a much weaker nucleophile than either hydroperoxy or t-butylperoxy, although the acetate ion is a much better leaving group than hydroxide or t-butoxide. Another observation we have made that is in disagreement with both these proposed transition states (I and II) is the decrease in rate at high pH only in the case of hydrogen peroxide cleavage, which is discussed below.

Apparently, the over-all rate constant is best correlated with both the properties, the leaving group characteristic of the anion and the nucleophilicity of the peranion. The mechanism most consistent with this experimental conclusion requires the rapid, equilibrium formation of the complex, where the position of equilibrium is directly related to the nucleophilic tendency of the peroxyanion. There is considerable evidence that such complexes lie at a potential energy minimum, contrary to the expectation of Bunton.[16]

Thus, Criegee[37] and Bailey[38] and their co-workers have established the existence of quite stable intermediates formed during the ozonolysis of acetylene derivatives in polar solvents which can be identified as possessing the general structure

$$RC{\equiv}CR + O_3 \xrightarrow{\text{SOH}} R-\underset{\underset{O}{\overset{|}{}}}{\overset{}{C}}-\underset{\underset{O-OH}{\overset{|}{}}}{\overset{\overset{OS}{|}}{C}}-R$$

II III

These intermediates have sufficient stability for isolation and accurate characterization of the nature of their decomposition products in a variety of solvolytic media, SOH.

The remaining data which our mechanism must account for are also in accord with prior equilibrium

(37) R. Criegee and M. Lederer, Ann., 583, 29 (1953); R. Criegee Record Chem. Progr., 18, 111 (1957).

(38) P. S. Bailey, private communication, July, 1960.

complex formation followed by rate-determining heterolysis of its -O-O- bond. (i) Exceedingly high base concentrations diminish the rate of cleavage in the cases of hydrogen peroxide, but not in the case of t-butyl hydroperoxide. This inhibition at very high pH's, mentioned earlier, is clearly due to the fact that in sufficiently strong base a change has occurred with proton transfer resulting in a dianion of lesser leaving group tendency; $i.e.$, $O^=$ is a poorer leaving group than OH^-, but the t-BuO$^-$ cannot be converted to a dianion. (ii) The large positive reaction constant ($\rho \cong +2.9$) is strongly suggestive of the analogous cleavage of benzils with cyanide ion ($\rho \cong +3.4$) as well as the benzilic acid rearrangement (Fig. 5). In both of these analogous cases, it was established that cleavage is the result of a fast equilibrium complex-forming step succeeded by a much slower decomposition step. It was pointed[17] out also that the ability of cyanide ion to effect the cleavage derives not so much from the nucleophilic character (10^{-2} as great as hydroxide ion and only 10^{-4} that of hydroperoxyanion in bonding to the carbonyl group) but rather from the strong electron attraction exerted by the -CN moiety bonded to the (former carbonyl) carbon center in the complex as illustrated in IV. The analogy is clear on the basis of the complex we can write for the peroxide cleavage as illustrated in V. The electrophilic power

IV   V
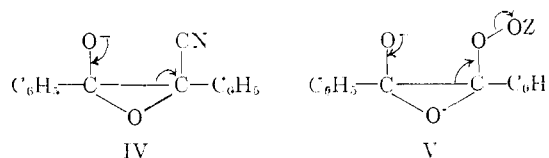
generated by potential heterolysis of the oxygen-oxygen bond in V is the probable driving force for complex decomposition, just as the inductive effect of the -CN group in IV contributes to the destruction of this complex.

The complexes of peroxyanions and α-diketones are apparently capable of some variation in this mode of decomposition other than shown in V. The nature of the products formed can be readily understood on this basis. For example, Karrer, Schwyzer and Neuwirth[39] have shown that in ether solution the product formed from the reaction of monoperphthalic acid and 4-methyl-O-benzoquinone is the cis,cis-β-methylmuconic anhydride.

However, an entirely different type of product is reported when alkyl hydroperoxide in alcoholic solution is used to cleave benzil. Here the original product is one mole of alkyl ester and one mole of benzoic acid[40] and there is no evidence that anhydride occurred during the course of reaction. On the other hand, when substituents exist on the aromatic rings and the substitution is asymmetric, some change occurs in the nature of the product composition. Thus, 4,4'-dimethoxybenzil still gives about equal parts of anisic acid and its ethyl ester, but 4-methoxybenzil gives more than four times

(39) P. Karrer, R. Schwyzer and A. Neuwirth, Helv. Chim. Acta, 31, 1210 (1948).

(40) R. P. Barnes and R. E. Lewis, J. Am. Chem. Soc., 58, 947 (1936).

as much mixed acid as ethyl ester product. Unfortunately, the composition of the acid and ester products, with respect to benzoic and anisic moieties was not reported,[40] and, until we have had the opportunity to investigate this matter further, little can be said.

Furthermore it seems possible that two other factors could alter the structure of the transition state; (i) the nature of the solvent medium, and (ii) the nature of the leaving group Z. For example, where $Z = RCO^-$, an alternative to V is

VI   VII
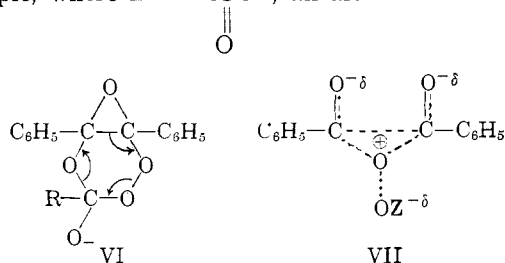
the transition state VI in which a neighboring group effect is operating. Apparently the solvent medium also plays a role in choosing between the alternatives. A poorly ionizing solvent like ether would tend to resist separation of the carboxylic acid anion and thereby enhance the neighboring group relationship. A solvent of low dielectric such as ether or chloroform may favor the occurrence of cleavage in concert with bond making and the transition state for such a mechanism might possibly be that represented by VII. Edwards and co-workers,[41] most recently, and Bartlett[42] have accepted as reasonable an analogous transition state for epoxidations by means of perbenzoic acid (Z = benzoyl).

(41) D. R. Campbell, J. O. Edwards, J. Maclachlan and K. Polgar, J. Am. Chem. Soc., 80, 5308 (1959).

(42) P. D. Bartlett, Record Chem. Prog., 11, 47 (1950).

[CONTRIBUTION FROM THE CHEMISTRY DEPARTMENT OF THE UNIVERSITY OF MARYLAND, COLLEGE PARK, MD.]

The Synthesis of α-Methoxyarylacetic Acids from the Base-catalyzed Condensation of Arylaldehydes with Bromoform and Methanol. Studies of Aldehyde Reactivities

BY WILKINS REEVE AND EDWARD L. COMPERE, JR.[1]

Certain arylaldehydes undergo a base-catalyzed reaction with bromoform in methanol solution to form α-methoxyarylacetates. This reaction has been studied with 23 aldehydes and one ketone, and it is concluded that the effect of ring substituents in this reaction is much like their effect in the Perkin reaction. Yields of approximately 75% of theory are obtained with the chlorobenzaldehydes, but p-hydroxybenzaldehyde and p-dimethylaminobenzaldehyde do not react. Many of the α-methoxyarylacetic acids form insoluble sodium acid salts, and the structural features necessary for this to occur are discussed. Many of the acids are potent plant growth regulators.

In an earlier publication,[2] a one-step synthetic method of preparing α-methoxyarylacetate esters and salts by the base-catalyzed condensation of an aryl aldehyde with chloroform or bromoform in methanol solution was described, and data were presented for seven substituted benzaldehydes to illustrate the utility of the method. The object of the present work was to study an extensive series of aromatic aldehydes to determine the relationship between structure and reactivity of the aldehyde in this reaction. The α-methoxyarylacetic acids are especially interesting because of two unusual characteristics. First, many of them have the ability to form relatively insoluble, crystalline, sodium acid salts, containing equimolecular amounts of the acid combined with its sodium salt, and at least one of these acid salts is useful in analytical chemistry[2,3]; second, they are potent plant growth regulators and, unlike most arylacetic acids, they are translocated readily throughout the plant and some of them have the property of even being exuded through the roots into the soil.[4]

(1) Bakelite Fellow, 1957–1958.

(2) W. Reeve and C. W. Woods, J. Am. Chem. Soc., 82, 4062 (1960).

(3) W. Reeve and I. Christoffel, Anal. Chem., 29, 102 (1957); W. Reeve, ibid., 31, 1066 (1959).

(4) W. H. Preston, J. W. Mitchell and W. Reeve, Science, 119, 437 (1954); P. J. Linder, J. C. Craig and T. R. Walton, Plant Physiol., 32, 572 (1957); J. W. Mitchell, B. C. Smale and W. H. Preston, J. Agr. Food Chem., 7, 841 (1959).

The α-methoxyarylacetate salts synthesized in this work were all prepared by what amounted to a standard procedure; 0.5 mole of potassium hydroxide in methanol was slowly added over a three-hour period to a mixture of 0.1 mole of the aldehyde and 0.12 mole of bromoform maintained at a temperature near 5°. The reaction consisted of the bromoform condensing with the carbonyl group to form first a tribromomethylcarbinol (I), which then was converted successively into a postulated epoxide and an ester (II) of the desired acid.[2]
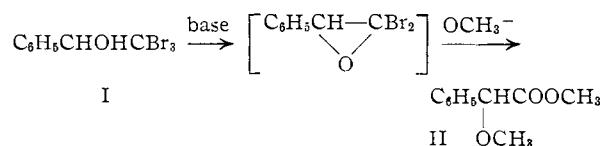
The tribromomethylphenylcarbinol has been prepared in 15% yield by Howard from benzaldehyde and bromoform with potassium hydroxide,[5] and the ester II has been demonstrated to be an intermediate when the haloform employed was chloroform.[2] Since the over-all yield of the methoxy acid, resulting from the hydrolysis of the ester and its isolation in the form of its sodium acid salt, is 40% it is obviously advantageous to carry out the reaction as described here so that the tribromocarbinol is not isolated. m-Chlorobenzal-

(5) J. W. Howard, J. Am. Chem. Soc., 52, 5059 (1930).