[CONTRIBUTION FROM THE CHEMICAL LABORATORY OF HARVARD UNIVERSITY]

The Influence of *n*-Alkyl Groups on the Rate of a Cyclization Reaction

BY ERNST BERLINER

A general inductive effect, which is described as manifesting itself in a feeble electron-release character, is attributed to alkyl groups.[1] This effect is considered to increase with the amount of branching in the order: methyl, ethyl, *i*-propyl, *t*-butyl, and with increasing length of a normal alkyl chain from methyl to about *n*-propyl or *n*-butyl, where a maximum is reached beyond which the effect is usually not transmitted. The nature and order of the effect are not very well defined or understood as yet, but it seems increasingly probable that in some cases (where the order appears reversed) no-bond resonance or hyperconjugation[2] (the Baker–Nathan effect[3]) may play an important, if not decisive, part. In addition, steric considerations cannot be neglected.[4]

Even though still imperfectly defined, the concept of an inductive effect has been used frequently and successfully to account for certain regularities in changes of physical constants (dipole moments, acid dissociation constants) and reaction rates, and considerable kinetic data derived from measurements in homologous series have been interpreted on this basis.[5]

In the present investigation measurements have been made of the rate of cyclization of ketones of the type I to the 9,10-disubstituted anthracenes (III), a reaction first described by Bradsher.[6] The reaction was carried out in a mixture of boiling acetic and hydrobromic acid, and the group R was varied from methyl to *n*-hexyl. The first step in the cyclization reaction involves the addition of a proton to the carbonyl group with the rapid establishment of a mobile equilibrium between the ketone and the conjugate acid II.[7] The latter exists as a resonance hybrid that attacks the *ortho* position of the opposite benzene ring in a normal aromatic substitution reaction.[8,9] As expected, the rate was found to be first order with respect to the ketone and is determined presumably by the attack of the carbonium ion on the benzene ring. The first order rate constants for the reactions of the different ketones as well as the half-life time of the cyclization reactions are listed in Table I.

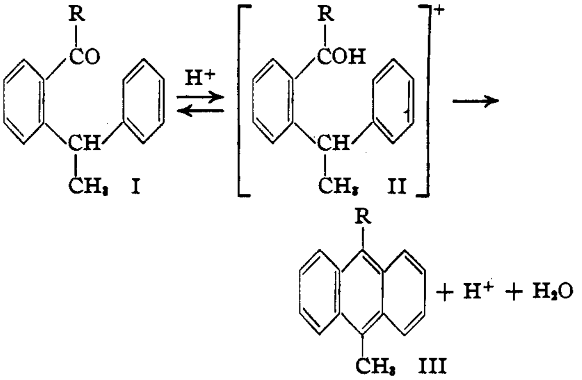

TABLE I

Ketone R	K (min.$^{-1}$) $\times 10^{-2}$	$t_{1/2}$, min.
Methyl	4.6	15
Ethyl	1.8	38.4
n-Propyl	0.99	70
n-Butyl	.35	198
n-Pentyl	.36	192
n-Hexyl	.36	192
Phenyl	.16	433[a]
Benzyl	.91	76[a]

The results obtained with the phenyl and benzyl ketones are not strictly comparable to those of the other ketones since some of the material remained undissolved when the reaction started. The values may, however, be taken as fair approximation.

The rate was found to decrease steadily from methyl to *n*-butyl and then to remain constant within the limits of the experimental error. The rate is also dependent upon the acid concentration; when the concentration of acid was lowered to three-fourths of the original, the observed rate constant for the methyl ketone was 0.76×10^{-2}.[10]

The decrease in rate with increasing length of the alkyl chain can be explained in terms of the inductive effect: an increase of electron density on the positive central carbon atom, caused by the electron release of the alkyl groups, results in a decrease of the effective positive character. Due to an increase in the energy of activation of the substitution reaction, the attack on the benzene ring becomes more difficult, and a decrease in rate results. This decrease reaches a limit with the limit of the inductive effect.

Although the electronic interpretation seems more attractive at the present time and fits well into the general picture of reaction mechanisms,

---

(1) (a) Watson, "Modern Theories of Organic Chemistry," Oxford, 1941, pages 92 ff.; (b) Ingold, *Chem. Rev.*, **15**, 238 (1934).

(2) Mulliken, Rieke and Brown, THIS JOURNAL, **63**, 41 (1941); Hughes, Ingold and Taher, *J. Chem. Soc.*, 949 (1941).

(3) Baker and Nathan, *ibid.*, 1844 (1935).

(4) Bartlett and Rosen, THIS JOURNAL, **65**, 543 (1943).

(5) Ref. 1a, pages 92 ff. and 135 ff.

(6) Bradsher, THIS JOURNAL, **62**, 486, 1077 (1940).

(7) Hammett, "Physical Organic Chemistry," McGraw-Hill Book Co., Inc., New York, N. Y., 1940, p. 237.

(8) Berliner, THIS JOURNAL, **64**, 2894 (1942).

(9) Bradsher and Smith, *ibid.*, **65**, 854 (1943).

(10) The lack of exact relationship between the rate and acid concentration may be due to the possibility that the actual concentration of acid in the refluxing mixture of acetic acid, hydrobromic acid, and water was not exactly three-fourths of the original. Because of this fact and a possible loss of solvent the rate constants presumably do not constitute absolute values. The stress is laid on the difference in rate of the different ketones.

it should be emphasized that equally plausible explanations possibly might be obtained from purely stereochemical considerations. The slow reaction of the phenyl ketone can be ascribed to steric effects as well as to additional resonance of the conjugate acid with the benzene ring, which distributes the charge over two nuclei and makes it less available for the attack. This additional effect is not present with the benzyl group, which has a rate comparable to that of the *n*-propyl group, but on the other hand has steric requirements quite different from those of the phenyl group.

The results obtained do not confirm the statement of Bradsher and Smith[9] that the rate of cyclization is roughly independent of the nature of the group R.

Acknowledgment.—I wish to express my sincere thanks to Professor L. F. Fieser, Harvard University, for his help and kind interest in the work.

Experimental

Rate Measurements.—All determinations were carried out under the same conditions: approximately 0.002 mole of the ketone (about 500 mg.) was dissolved in 10 cc. of glacial acetic acid in a 50-cc. round-bottomed flask equipped with a ground glass reflux condenser. Four cc. of hydrobromic acid (48%) and one cc. of water were added, and the flask was immersed in a Wood's metal-bath which was preheated to 133–134° and maintained at this temperature. The first three ketones were soluble in the cold mixture, while the other three dissolved in the boiling mixture. The hydrocarbons resulting from the cyclization separated from the boiling solution. After a given time the flask was removed from the bath and cooled in an ice-bath. A crystal of the pure hydrocarbon was added if crystallization was slow. The solid material was filtered and washed with two 25-cc. portions of distilled water. It was dried in a vacuum oven overnight and weighed. The amount of hydrocarbon left in the mother liquor was negligible and never exceeded a few mg. In a typical case the following results were obtained

R = methyl

Ketone, g.	Hydrocarbon, g.	Time, min.	K (min.$^{-1}$) × 10^{-2}
0.4464	0.1877	13	4.70
.4470	.2492	20	4.66
.4412	.3103	30	4.82
.4471	.3451	40	4.57
.4485	.3612	50	4.17
			Av. 4.58

Preparation of Materials.[11]—The preparation of the intermediate ketones was similar to the procedure used by Bradsher.[6]

***o*-Chlorobenzophenone.**—This compound has been prepared previously but no exact procedure is given.[12] The acid chloride was prepared from *o*-chlorobenzoic acid (147 g.) and phosphorus pentachloride (100 g.). The phosphorus oxychloride was removed *in vacuo* and the acid chloride dissolved in 300 cc. of dry benzene. This solution was dropped on a suspension of 170 g. of aluminum chloride and 400 cc. of benzene in a three-necked flask equipped with a reflux condenser, a dropping funnel, and a hydrogen chloride trap. The addition took two hours, after which time the mixture was allowed to stand at room temperature for five hours. The reaction was completed by refluxing the mixture gently for thirty minutes. Hydrochloric acid and small pieces of ice were added carefully; the layers were separated, and the aqueous layer extracted with some fresh benzene. The combined layers were dried, the solvent evaporated, and the product distilled. It distilled at 181° at 18–19 mm. (redistilled at 178–180° (14–15 mm.)) and solidified to small crystals melting at 43–44°.[13] The yield was 175 g. (85.7%).

1-(*o*-Chlorophenyl)-1-phenylethylene was prepared according to Bergmann and Bondi[14] from magnesium (8 g.), methyl chloride gas, and *o*-chlorobenzophenone (39 g.). The mixture was refluxed overnight and decomposed with ice-cold 30% sulfuric acid. The yield was 29.7 g. (77%) of a colorless oil that distilled at 164–166° (17 mm.). Bergmann and Bondi report the value of 162–163° (18 mm.). The same compound could also be prepared from *o*-chloroacetophenone and phenylmagnesium bromide, but the yield was not higher than 50%.

1-(*o*-Chlorophenyl)-1-phenylethane.—Twelve grams of the above olefin was dissolved in 20 cc. of absolute alcohol and hydrogenated in the presence of Raney nickel catalyst at atmospheric pressure. The product distilled at 156–157° (12 mm.) as a colorless oil with a pleasant, fruit-like odor. The yield was almost quantitative. A sample redistilled at 158° (12 mm.) had the following composition:

Anal.[15] Calcd. for $C_{14}H_{13}Cl$: C, 77.59; H, 6.05. Found: C, 77.10; H, 6.03.

1-(*o*-Cyanophenyl)-1-phenylethane.—A mixture of the above chloro compound (27.5 g.), 14 g. of cuprous cyanide (dried at 100° for ten hours before use), and 12.5 cc. of dry pyridine was refluxed for thirty-six hours in a Wood's metal-bath maintained at 250–260°. The product after isolation in the usual way distilled at 183–188° (16 mm.) as an amber colored oil.[16] The yield was 20 g. (75.7%). Material redistilled at 190–191° (17–18 mm.) showed the composition:

Anal. Calcd. for $C_{15}H_{13}N$: C, 86.92; H, 6.32. Found: C, 86.76; H, 6.11.

All the ketones and hydrocarbons were prepared in the same way. In a typical procedure the following method was used:

1-(*o*-Acetylphenyl)-1-phenylethane. (I, R = methyl). —A Grignard solution was prepared from 2.7 g. of magnesium turnings and methyl chloride gas. The cyano compound (10.3 g.) in benzene solution was added to the solution of methylmagnesium chloride, and the mixture was refluxed overnight after which time a crystalline, white solid had separated. (In no other case was a precipitate formed.) The Grignard solution was decomposed with a 20% solution of ammonium chloride, the organic layer was separated, and the solvent evaporated without further drying. Thirty cc. of concentrated hydrochloric acid, 40 cc. of water, and 40 cc. of acetone were added to the residue, and the mixture was refluxed for thirty minutes. During this time most of the organic solvent distilled from a side arm of the condenser. The remaining oil, which showed a slight fluorescence, was taken up in benzene; the benzene layer was dried, and the solvent evaporated. The product distilled at 184–186° (16–17 mm.) as a light, almost colorless oil. The yield was 8.4 g. (75.5%).

Anal. Calcd. for $C_{16}H_{16}O$: C, 85.67; H, 7.19. Found: C, 85.52; H, 6.81.

9,10-Dimethylanthracene.—A mixture of 1.2 g. of the above ketone, 15 cc. of glacial acetic acid, 10 cc. of hydrobromic acid, and 2 cc. of water was refluxed in an all-glass apparatus. After fifteen minutes a crystalline product

---

(11) All melting points are corrected.

(12) Overton, *Ber.*, **26**, 29 (1893); Graebe and Keller, *ibid.*, **32**, 1687 (1899).

(13) Auwers, *Ber.*, **58**, 50 (1925), reports a boiling point of 185–188° (13 mm.).

(14) Bergmann and Bondi, *ibid.*, **64**, 1473 (1931).

(15) Microanalyses by Miss E. Werble.

(16) Bergmann, *J. Org. Chem.*, **4**, 1 (1939), reports a boiling point of 166–168° (5 mm.) and 151° (1.8 mm.).

separated from the boiling solution.[17] The flask was cooled after a total refluxing time of three hours and the hydrocarbon was filtered. Recrystallized from alcohol and little benzene, it forms yellow crystals, often described in the literature.[18] The melting point was 182.8–183.8°. The picrate forms dark brown needles of the melting point 179–180°. The physical constants of the other ketones and hydrocarbons are listed in Table II.

TABLE II

KETONE (I)

R	B. p. °C.	B. p. Mm.	Calcd., % C	Calcd., % H	Found, % C	Found, % H
Ethyl	189–190	16–17	85.67	7.61	85.24	7.32
n-Propyl	194–197	14–15	85.67	7.99	85.29	8.21
n-Butyl	205–206	17–18	85.67	8.33	85.34	8.04
n-Pentyl	209–212	15–16	85.67	8.63	85.46	8.31
n-Hexyl	217–219	14–15	85.66	8.9	85.31	8.61
Phenyl[23]	216–219	7–8	88.08	6.34	87.62	6.08
Benzyl	195–197	1–2	87.96	6.71	87.61	6.34

HYDROCARBON (III)

R	M. p., °C.	Solvent	Form
Ethyl	143.2–144	Alcohol	Yellow needles
n-Propyl	97.8–98.6	Alcohol	Orange plates
n-Butyl	78.2–78.8	Alcohol	Long yellow needles
n-Pentyl	71–71.8	Alcohol	Long yellow needles
n-Hexyl	65.8–66.5	Alcohol	Long yellow needles
Phenyl	113.5–114.5[19]	Benzene/ligr.	Yellow prisms
Benzyl	167.8–168.6	Alcohol/benz.	Long yellow needles

R	Calcd., % C	Calcd., % H	Found, % C	Found, % H	Picrates[20] M. p., °C.	Picrates[20] Nitrogen, % Calcd.	Picrates[20] Nitrogen, % Found
Ethyl	92.68	7.32	92.58	7.41	137.8–138.4	9.35	9.26
n-Propyl	92.11	8.04	92.26	7.74	125.5–126.2	9.06	8.57
n-Butyl	91.88	8.11	91.73	8.05	91.8–92.8	8.8	8.46
n-Pentyl	91.55	8.45	91.72	8.68	85.4–86.2	8.55	8.10
n-Hexyl	91.25	8.75	91.43	8.83	78.2–79.2[21]	5.38	5.13
Phenyl					125.2–126	8.64	8.46
Benzyl[22]	95.58	6.42	95.78	6.76			

Summary

The rate of the cyclization of different ketones to 9,10-disubstituted anthracenes has been measured and found to decrease as the number of carbon atoms on the alkyl groups increases. The decrease in rate can best be explained in terms of the inductive effect of alkyl groups. A number of 9-alkyl-10-methylanthracenes has been prepared and the physical constants are described.

---

[17] The cyclization of the ketone which leads to 9-methylanthracene is reported to take place in four days (Bradsher, ref. 6a).

[18] Bachman and Chemerda, *J. Org. Chem.*, **4**, 583 (1939); Barnett and Matthews, *Ber.*, **59**, 1437 (1926); J. E. Jones, Thesis, Harvard University, 1941.

[19] Barnett and Matthews, *Ber.*, **59**, 1437 (1926). report a melting point of 112°.

[20] All picrates form dark brown crystals and decompose when recrystallized from alcohol, benzene or ligroin. They are best recrystallized from methyl alcohol in the presence of some excess picric acid.

[21] 9-Hexyl-10-methylanthracene forms a picrate which contains two molecules of hydrocarbon. The same is reported for 9-butylanthracene (Stieglitz and Marx, *Ber.*, **56**, 1619 (1923)).

[22] This hydrocarbon did not form a picrate.

[23] This ketone solidified after long standing and had a crude melting point of 47–48°.

CAMBRIDGE, MASS. RECEIVED DECEMBER 18, 1943

[CONTRIBUTION FROM THE CENTRAL EXPERIMENT STATION, BUREAU OF MINES]

Hydrogenation of Anthracene by Tetralin[1]

BY MILTON ORCHIN[2]

Examination of some of the high-boiling liquid products obtained by hydrogenation of coal at the Bureau of Mines Central Experiment Station involved attempts at dehydrogenation. The physical properties of these oils indicated a large percentage of hydroaromatic compounds, yet catalytic liquid-phase dehydrogenation gave only small quantities of evolved hydrogen. To understand better the reactions that may occur when mixtures of aromatics and hydroaromatics such as are found in the products from the hydrogenation of coal are heated in the liquid phase in the presence of a catalyst, a study of the mixture anthracene and tetralin was undertaken.

The literature reveals that under the conditions necessary for liquid-phase dehydrogenation, the hydrogen is often utilized in what may be regarded as a sort of internal oxidation–reduction process. This transfer of hydrogen can be either intramolecular or intermolecular. Examples of the former are the conversion of methyl 2-(5,6,7,8-tetrahydro)-naphthyl ketone to 2-ethylnaphthalene[3]; and 1-hydroxy-1-allyl-1,2,3,4-tetrahydrophenanthrene to 1-*n*-propylphenanthrene.[4,5] Examples of the intermolecular transfer of hydrogen are the conversion of 2,3-naphthalic anhydride to 2,3-dimethylnaphthalene in small yield by means of *p*-cyclohexylphenol[6]; the disproportionation reactions in the terpene series[7]; the incomplete dehydrogenation of tetralin due to the equilibrium tetralin $\leftrightarrows$ naphthalene + $2H_2$[8]; and the work of Adkins and coworkers, who dehydrogenated many compounds by heating in a steel bomb with benzene as a hydrogen acceptor.[9]

---

[1] Published by permission of the Director, Bureau of Mines, U. S. Department of the Interior. Not copyrighted.

[2] Organic chemist, Central Experiment Station, Bureau of Mines, Pittsburgh, Pa.

[3] Newman and Zahm, THIS JOURNAL, **65**, 1097 (1943). Barbot, *Bull. soc. chim.*, [4] **47**, 1314 (1930), dehydrogenated the same ketone with sulfur at 220° and obtained a 71% yield of methyl 2-naphthyl ketone. Apparently, under these conditions the ketone group is not attacked.

[4] Bachmann and Wilds, THIS JOURNAL, **60**, 624 (1938).

[5] See also (a) Ruzicka, *Helv. Chim. Acta*, **19**, 419 (1936); (b) Fieser and Joshel, THIS JOURNAL, **61**, 2958 (1939); and (c) Mosettig and Duval, *ibid.*, **59**, 367 (1937); (d) for a recent complete review of the subject of dehydrogenation, see Plattner, *Die Chemie*, **55**, 131 (1942).

[6] Windaus and Thiele, *Ann.*, **521**, 163 (1935).

[7] Linstead, Michaelis and Thomas, *J. Chem. Soc.*, 1139 (1940).

[8] Linstead and Michaelis, *ibid.*, 1134 (1940).

[9] Adkins and Reid, THIS JOURNAL, **63**, 741 (1941), and Adkins, Richards and Davis, *ibid.*, **63**, 1320 (1941).