THE MECHANISM OF AROMATIC MERCURATION. II

REACTION KINETICS[1]

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INTRODUCTION

The rapid rate of mercuration of benzene in nitric acid solutions was discovered during a study of the oxynitration of benzene (28).

$$Hg^{++} + C_6H_6 \rightarrow C_6H_5Hg^+ + H^+ \qquad (1)$$

It was found that the reaction rate in 60 per cent nitric acid is more than a hundred times that in 20 per cent acid. It has now been shown that the reaction occurs homogeneously in the aqueous phase, that the rate is proportional to the product of the concentrations of benzene and of mercuric ion, and that the reaction is strongly accelerated both by acids *and by the corresponding salts.*[4] Furthermore, the reaction is most strongly accelerated by salts of those anions (e.g., nitrates and perchlorates) which are generally considered to be the least effective as complex-forming agents. The data are presented in the following paragraphs; then the mechanism of mercuration and of similar reactions will be discussed.

KINETIC ORDER OF THE REACTION

In these studies the mercuration was usually carried out with an excess of benzene, stirred rapidly with an aqueous solution of acid and mercuric salt. Under these conditions the concentration of benzene in the aqueous layer was constant throughout the experiment, and the rate proved essentially first order with respect to mercuric salt. However, only the initial reaction rate (first 10–20 per cent of the reaction) was measured, since polymercuration (13) interfered with the rate determinations at high conversions. That the reaction takes place in the aqueous layer is shown by the fact that the rate is independent of stirrer speed and of the volume of the benzene solution (see table 1). The fact that the reaction is first order with respect to mercuric ion is shown by the data of table 2.

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[4] In a previous publication (28) it was erroneously concluded that salts were ineffective in accelerating the reaction. This error was caused by neglect of the effect of added salts on the solubility of benzene in the solutions used (see table 5).

The dependence of the rate on the benzene concentration was determined at 25°C. instead of at 40°C. (where most measurements were made). The lower temperature was chosen to minimize the volatilization of benzene from the extremely dilute aqueous solutions of benzene, but the slight decrease in the second-order rate constant at lower benzene concentrations is probably due to volatility losses.

TABLE 1

The rate of mercuration of benzene in 300 ml. of 3.82 M sulfuric acid at 40°C.

CONCENTRATION OF $HgSO_4$	STIRRER SPEED	VOLUME OF BENZENE LAYER	k
moles/liter	R.P.M.	*ml.*	*min.*$^{-1}$
0.092	130	25	0.00051
0.092	600*	25	0.00052
0.091	600*	100	0.00053

TABLE 2

The order of mercuration with respect to mercuric ion at 40°C.

ACID	CONCENTRATION OF ACID	CONCENTRATION OF Hg^{++}	$10^5\ k$
	moles/liter	*moles/liter*	*min.*$^{-1}$
Nitric*	3.54	0.0597	14.
Nitric*	3.54	0.297	16.
Nitric*	5.62	0.0600	45.
Nitric*	5.62	0.298	37.
Nitric*	7.91	0.0598	110.
Nitric*	7.91	0.295	100.
Nitric*	10.40	0.0598	350.
Nitric*	10.40	0.290	360.
Nitric*	10.40	1.22	560.†
Perchloric	0.53‡	0.025	6.5
Perchloric	0.53	0.092	6.2

* With a small amount of added urea.

† The reason for this large rate constant is the large increase in nitrate ion (*vide infra*).

‡ Interpolated from two nearby values.

TABLE 3

The order of mercuration in aqueous solution with respect to benzene at 25°C.

SOLVENT	INITIAL Hg^{++} CONCENTRATION	INITIAL BENZENE CONCENTRATION	SECOND-ORDER RATE CONSTANT
	moles/liter	*moles/liter*	*min.*$^{-1}$ (*moles/liter*)$^{-1}$
3.4 M $HClO_4$	0.030	0.029*	4.5×10^{-3}
3.4 M $HClO_4$	0.030	0.010	4.1×10^{-3}
8.0 M HNO_3	0.040	0.040	3.0×10^{-3}
8.0 M HNO_3	0.040	0.010	2.5×10^{-3}

* The solubility of benzene in 3.4 M perchloric acid. This reaction was carried out in the presence of excess benzene.

Since the reaction rate is proportional to the concentration of benzene, and since there is no apparent irregularity at saturation, most of the experiments were carried out, in accordance with earlier practice (28), in the presence of excess benzene. The second-order rate constants, k_2, were determined by dividing the observed first-order constants, k_1, by the solubility of benzene in the particular solvent used.

"SALT EFFECT"

The effect of acids and salts on the mercuration rate is shown in figures 1 and 2 and in tables 4, 5, and 6.

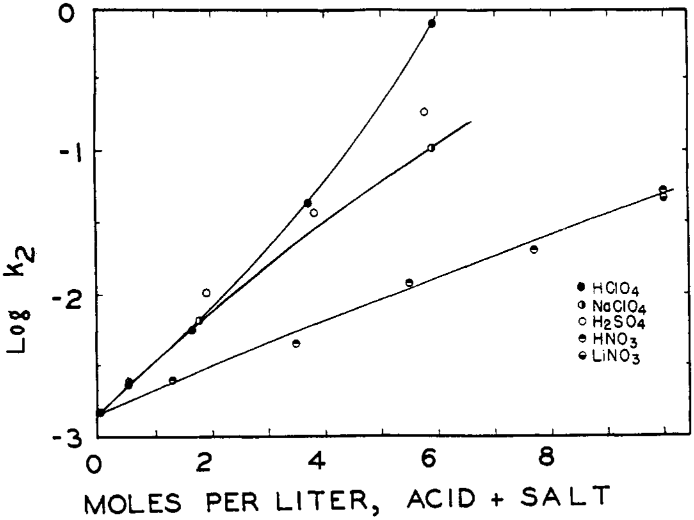

FIG. 1. The effect of added acids and salts on the rate of mercuration of benzene

In the data here recorded, the incomplete ionization (21) of nitric acid has not been taken into account. Since nitric acid is a fairly strong acid, this omission is inconsequential for solutions less concentrated than 3 *M*. However, if the incomplete ionization is taken into account, 6 *M* nitric acid solution is a better catalyst than is the corresponding lithium nitrate solution. This question is further considered in the discussion.

Not all salts, however, are effective in accelerating the reaction (4): chlorides are strong inhibitors, as shown in table 7. There is too much uncertainty in the equilibrium constant (18, 24) at high ionic strength, for the reaction

$$Hg^{++} + HgCl_2 \rightleftarrows 2HgCl^+ \qquad (2)$$

to permit a calculation of whether the $HgCl^+$ ion is or is not an effective mercurating agent.

The data for both the nitrate and the perchlorate catalysis extrapolate at

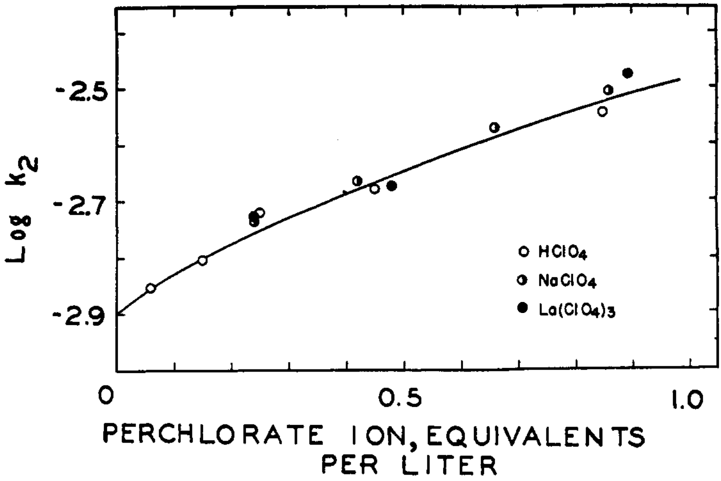

Fig. 2. The effect of perchloric acid and of perchlorates on the rate of mercuration of benzene.

TABLE 4

The rate of mercuration of benzene in perchlorate solutions at 40°C.

Hg^{++}	$HClO_4$	$NaClO_4$	$La(ClO_4)_3$	ClO_4^-	μ	$10^5 k_1$	SOLUBILITY OF BENZENE	$10^3 k_2$
moles/liter	*moles/liter*	*moles/liter*	*moles/liter*	*moles/liter*		*min.*$^{-1}$	*moles/liter*	*min.*$^{-1}$ (*moles/liter*)$^{-1}$
0.025	0.010			0.060	0.085	3.6	0.026	1.40
0.025	0.100			0.150	0.175	4.1	0.026	1.57
0.090	0.036			0.212	0.306	3.9*	0.026	1.50
0.025	0.010	0.180		0.240	0.265	4.8	0.026	1.84
0.025	0.010		0.060	0.240	0.445	4.9	0.026	1.88
0.025	0.200			0.250	0.275	4.9	0.026	1.90
0.025	0.010	0.360		0.420	0.445	5.4	0.025	2.2
0.025	0.400			0.450	0.475	5.7	0.027	2.1
0.025	0.010		0.140	0.480	0.925	5.5	0.026	2.1
0.025	0.010	0.600		0.660	0.685	6.1	0.023	2.7
0.092	0.034	0.500		0.718	0.810	5.8*	0.024	2.4
0.092	0.53			0.712	0.806	6.2*	0.027	2.3
0.025	0.800			0.850	0.875	8.0	0.028	2.8
0.025	0.010	0.800		0.860	0.885	6.9	0.022	3.1
0.025	0.010		0.278	0.894	1.753	8.4	0.025	3.4
0.092	1.66			1.84	1.93	18.5	0.033	5.6
0.091	0.035	1.75		1.97	2.06	11.7	0.018	6.5
0.092	3.71			3.89	3.98	150.	0.036	42.
0.088	0.037	5.84		6.06	6.15	102.	0.010	102.
0.093	5.92			6.10	6.19	2700.	0.036	750.
0.226	11.2			11.7	11.9	*ca.* 1,400,000		*ca.* 400,000

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* There is a slight discrepancy here between experiments conducted by R. M. Schramm and those by W. Klapproth; the data marked by the asterisks were provisional and are presumed to be of less accuracy than the other measurements.

TABLE 5

The rate of mercuration of benzene in nitrate solutions at 40°C.*

Hg^{++}	HNO_3	$NaNO_3$	$LiNO_3$	$N(C_2H_5)_4NO_3$	GROSS NO_3^-	μ	$10^5 k_1$	SOLUBILITY OF BENZENE	$10^3 k_2$
moles/liter	*moles/liter*	*moles/liter*	*moles/liter*	*moles/liter*	*moles/liter*		*min.*$^{-1}$	*moles/liter*	*min.*$^{-1}$ *(moles/liter)*$^{-1}$
0.025	0.050				0.100†	0.125	3.6	0.026	1.40
0.025	0.150				0.200	0.225	3.9	0.026	1.50
0.025	0.100	0.250			0.400	0.425	4.1	0.025	1.64
0.025	0.150			0.300	0.500	0.525	4.0	0.033	1.20
0.025	0.600				0.650	0.675	4.7	0.027	1.73
0.025	0.150	1.000			1.200	1.225	4.1	0.021	1.97
0.090	1.32				1.50	1.59	7.0	0.028	2.5
0.025	0.150			1.914	2.114	2.139	10.2	0.110	0.93
0.025	2.450				2.500	2.525	9.7	0.029	3.3
0.025	0.100	3.000			3.150	3.175	3.6	0.014	2.6
0.060	3.50				3.62	3.68	14.	0.031	4.5
0.025	0.200	4.167			4.417	4.442	4.5	0.011	4.1
0.060	5.50				5.62	5.68	45.	0.036	12.
0.025	0.098	5.910			6.057	6.082	4.1	0.007	5.6
0.049	7.7				7.8	7.85	100.	0.049	20.
0.061	10.0				10.1	10.2	370.	0.071	52.
0.090	1.32		8.75		10.3	10.3	23.	0.0050	46.

* 0.005 mole/liter (or more) of urea was added to every solution to destroy oxides of nitrogen.

† Without taking into account the incomplete ionization of nitric acid (see text).

TABLE 6

The rate of mercuration of benzene in sulfuric acid solutions at 40°C.

Hg^{++}	H_2SO_4	$10^5 k_1$	SOLUBILITY OF BENZENE	$10^3 k_2$
moles/liter	*moles/liter*	*min.*$^{-1}$	*moles/liter*	*min.*$^{-1}$ *(moles/liter)*$^{-1}$
0.091	1.86	18.5	0.018	10.3
0.092	3.82	52.	0.014	37.
0.061	5.77	240.	0.013	185.

TABLE 7

The rate of mercuration of benzene in the presence of chloride ion at 40°C.

Hg^{++}	Cl^-	$HClO_4$	$10^5 k_1$
moles/liter	*moles/liter*	*moles/liter*	*min* $^{-1}$
0.092	0.19	4.00	1
0.092		3.75	150
0.0454	0.0478	3.75	62

zero salt concentration to the same rate constant. This rate constant, k_0, for the reaction of a hydrated mercuric ion with a benzene molecule has the value of 1.27×10^{-3} min.$^{-1}$ (moles/liter)$^{-1}$ at 40°C.

In solutions containing less than 1 mole/liter of perchlorate ion, and in all the nitrate solutions, it is an empirical fact that the rate can be approximated by equation 3:

$$v = k_{obs}(C_6H_6)(Hg^{++}) = k_0(C_6H_6)(Hg^{++}) + k_{3A}(C_6H_6)(Hg^{++})(A^-) \qquad (3)$$

where k_{obs} is the bimolecular rate constant found for mercuration in any particular solution, k_{3A} is the termolecular rate constant appropriate to the particular anion present at a concentration (A), and k_0 has the value previously assigned. The third-order rate constants, determined according to equation 3, are presented in tables 8 and 9.

TABLE 8

Third-order rate constant for perchlorate ion-catalyzed mercuration of benzene

PERCHLORATE	$10^3 (k_{obs} - k_0)$	$10^3 k_3 = 10^3 \frac{(k_{obs} - k_0)}{(ClO_4^-)}$
equiv./liter	*min.*$^{-1}$ *(moles/liter)*$^{-1}$	*min.*$^{-1}$ *(moles/liter)*$^{-2}$
0.060	0.13	2.2
0.150	0.30	2.0
0.240	0.57	2.4
0.240	0.61	2.5
0.250	0.63	2.5
0.420	0.90	2.1
0.450	0.83	1.8
0.480	0.85	1.8
0.660	1.36	2.1
0.850	1.59	1.9
0.860	1.85	2.1
0.894	2.08	2.3
Average..........		2.1

TABLE 9

Third-order rate constant for nitrate ion-catalyzed mercuration of benzene

NITRATE	$10^3 (k_{obs} - k_0)$	$10^3 k_3 = 10^3 \frac{(k_{obs} - k_0)}{(NO_3^-)}$
equiv./liter	*min.*$^{-1}$ *(moles/liter)*$^{-1}$	*min.*$^{-1}$ *(moles/liter)*$^{-2}$
0.100	0.13	1.3
0.200	0.23	1.2
0.400	0.35	0.9
0.650	0.46	0.7
1.200	0.70	0.6
2.500	2.07	0.8
3.150	2.44	0.8
4.417	2.59	0.6
6.057	4.29	0.7
Average..........		0.8

Of course, that k_3 should be so constant as appears to be the case in the perchlorate series is entirely fortuitous. Actually, for a reaction of the type shown by the last term of equation 3, one involving ions of opposite charge, a negative primary salt effect (2) should be observed. Such a decrease in k_3 is quite apparent in the nitrate series.

The only exception to the general rule that the second-order reaction rate constant is linear with the anion concentration arises from the experiments with tetraethylammonium nitrate. Here both the solubility of benzene and the first-order rate constant are markedly increased; however, the second-order rate constant, $k_1/(C_6H_6)$, is decreased by the addition of the salt. This effect becomes understandable if the reason for the increased solubility of benzene is considered. The hydrocarbon must certainly be closely associated with the large organic cations and correspondingly close to a positive charge. It follows therefore that the rate of reaction with a positively charged mercuric ion must be strongly decreased.

TABLE 10

Mercuration of benzene in 97 per cent acetic acid solution at 40°C.

Effect of perchloric acid and sodium perchlorate

Hg^{++} SALT	Hg^{++} SALT	$HClO_4$	$NaClO_4$	TOTAL (ClO_4^-)	*T*	C_6H_6	$10^3 k_2$
moles/liter		*moles/liter*	*moles/liter*	*moles/liter*	°C.	*moles/liter*	*min.*$^{-1}$ (*moles/liter*)$^{-1}$
0.100	$Hg(OAc)_2$				25	1.00	0.0025
0.100	$Hg(OAc)_2$	0.05		0.05	25	1.00	0.21
0.100	$Hg(OAc)_2$	0.10		0.10	25	1.00	0.60
0.100	$Hg(OAc)_2$	0.20		0.20	25	1.00	1.60
0.100	$Hg(OAc)_2$	0.40		0.40	25	1.00	4.96
0.100	$Hg(OAc)_2$	0.60		0.60	25	1.00	10.7
0.100	$Hg(ClO_4)_2$			0.20	25	1.00	1.95
0.100	$Hg(ClO_4)_2$		0.20	0.40	25	1.00	2.04
0.100	$Hg(ClO_4)_2$		0.40	0.60	25	1.00	2.60
0.100	$Hg(ClO_4)_2$		0.60	0.80	25	1.00	3.21
0.089*	$Hg(OAc)_2$	0.32		0.32	40	0.071	51.
0.089*	$Hg(OAc)_2$	0.32		0.32	40	0.35	48.

* Glacial acetic acid.

MERCURATION IN ACETIC ACID SOLUTION

The mercuration of benzene was also carried out in homogeneous solution in acetic acid. The rate constants so obtained are presented in table 10 and figure 3.

The reaction is strongly accelerated by perchloric acid and slightly accelerated by sodium perchlorate. Although these results are at first glance at variance with those obtained in the aqueous system (where salts are about as effective as acids), the discrepancy is more apparent than real. It has previously (13) been shown that an ionic mercurating agent is very much more effective than one in which the mercury is strongly coördinated with acetate or chloride. But in

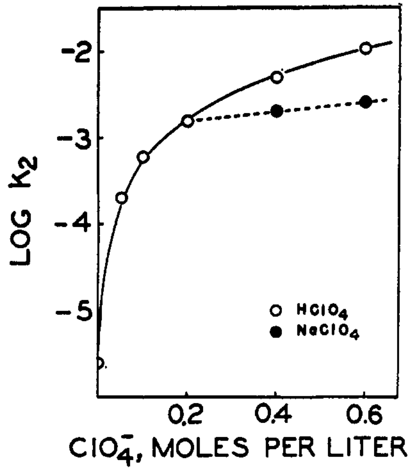

FIG. 3. Catalysis of the mercuration reaction in 97 per cent acetic acid

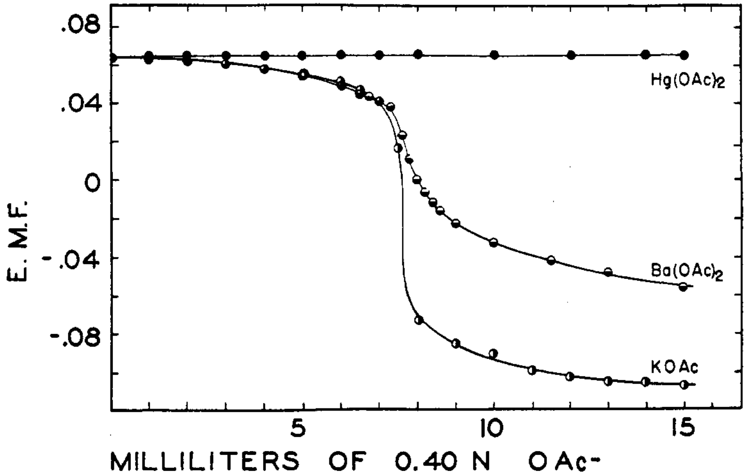

FIG. 4. Potentiometric titration in 97 per cent acetic acid. The 0.40 N potassium acetate, barium acetate, or mercuric acetate was added to 75 ml. of 0.04 N perchloric acid in 97 per cent acetic acid.

acetic acid solution the mercury is certainly held primarily as mercuric acetate. Even when perchloric acid is added, the equilibrium

$$Hg(OCOCH_3)_2 + HClO_4 \rightleftarrows Hg(OCOCH_3)^+ + CH_3COOH \qquad (4)$$

is displaced far to the left. The primary effect of acid, then, is to liberate a small amount of ionized mercury compound; the effect of the perchlorate ion, although real, is secondary.

The fact that the position of equilibrium 4 lies far to the left was established by electrometric titration. The curves for the titration of perchloric acid with barium, sodium, and mercuric acetates are shown in figure 4. Since mercuric acetate does not neutralize perchloric acid, the extent to which reaction 4 proceeds to the right must be small.

EXPERIMENTAL

Reagents

Except in the following instances, analytical reagent grade chemicals were used directly. Anhydrous sodium perchlorate was prepared from the commercial monohydrate by recrystallization at temperatures above 50°C.; it was then dried 24 hr. at 135°C. "Thiophene-free" benzene was allowed to stand overnight over mercuric acetate and Drierite and then distilled through an 18-in. Podbielniak column. Mercuric perchlorate hexahydrate was prepared by the method of Geffcken (7), except that the product was recrystallized from 20 per cent perchloric acid to prevent the formation of a basic salt. Analyses: Hg, 39.8 per cent (by titration with standard potassium thiocyanate solution); ClO_4^-, 39.1 per cent (by precipitating the mercury with hydrogen sulfide, then titrating the resulting perchloric acid with standard alkali); calculated for $Hg(ClO_4)_2 \cdot 6H_2O$:Hg, 39.5 per cent; ClO_4^-, 39.2 per cent.

The double salt $La(NO_3)_3 \cdot 2NH_4NO_3 \cdot 4H_2O$ was kindly furnished by Dr. Harold Friedman. From this salt the chalk-white hydroxide was precipitated and thoroughly washed. A lanthanum perchlorate solution of known concentration was prepared by adding to a suspension of lanthanum hydroxide an accurately known quantity of perchloric acid, insufficient to dissolve all the precipitate. Since the solution was neutral[5] (23), the lanthanum concentration was equivalent to that of the perchlorate.

Tetraethylammonium bromide was prepared from triethylamine and ethyl bromide. The nitrate was prepared from the bromide by reaction in alcoholic solution with an equivalent of silver nitrate. The product was twice recrystallized from alcohol–ether and dried at 115°C. Tests for both silver and bromide ions were negative.

A standard solution of mercuric nitrate was prepared by dissolving a weighed amount of pure mercury in concentrated nitric acid. The solution was heated until it became colorless; it was used to standardize the potassium thiocyanate solutions.

Rate determinations

The method used was similar to that previously employed (28). A 500-ml. two-necked ground-glass reaction flask fitted with stirrer and a paraffin oil seal

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[5] In a private communication to Professor T. F. Young, Dr. T. Shedlovsky states that a recalculation of the data (23) shows that the activity coefficients of lanthanum chloride solution are normal.

was immersed in a thermostat maintained at 39.90°C. ± 0.03°. When the catalyst for mercuration was a salt, rather than an acid, a small amount of acid was added to suppress hydrolysis (6, 19) of the mercuric salt; this acid was kept at 0.01 mole per liter of perchloric acid or 0.1–0.2 mole per liter of nitric acid. A small amount of urea (usually 0.005 mole/liter) was added to all the nitric acid and nitrate solutions in order to prevent the accumulation of oxides of nitrogen, which would otherwise destroy the mercurated product (28).

The reaction was started (except where otherwise noted) by adding an excess of benzene; the mixture was stirred vigorously throughout each experiment. The first aliquots, from which the initial concentration of mercuric ions was calculated, were taken 3–5 min. after the addition of the benzene. The reaction was followed by titrating at 10–15°C. the mercuric and phenylmercuric ions with standard potassium thiocyanate (14).

Since the reactions were allowed to proceed only about 10 per cent to completion, it was not necessary to correct the rate constants for back-reaction. It was, however, independently established that the equilibrium, with perchloric acid and sulfuric acid solutions, lies much farther in the direction of complete mercuration than is the case with nitric acid solution (28). Because of the small amount of reaction, the accuracy of the rate constants here determined cannot be greater than about ±5 per cent, despite the fact that the analytical procedure is reproducible to 0.3 per cent. Furthermore, the fact that there is considerable polymercuration even at 10 per cent reaction introduces additional uncertainty into the calculated rate constants.

In the reaction mixtures containing concentrated solutions of sodium nitrate, titration of the benzene layer showed that a small amount of phenylmercuric nitrate was being extracted into the hydrocarbon. This amounted to about 1 per cent in 1 M, 3 per cent in 3 M, and 6 per cent in 6 M sodium nitrate, in experiments where the volume of the benzene layer was about 15 ml. and that of the aqueous layer averaged 200 ml.; the salting out was almost complete with 8.75 M lithium nitrate. Since this loss of mercurated product from the aqueous layer increases the apparent rate of reaction, the rate constants were corrected accordingly. The "salting-out" phenomenon here reported may be useful in the preparation of phenylmercuric nitrate. Tests for nitrophenols (28) in the nitrate solutions were all negative.

Test of the kinetic order with respect to benzene

For those experiments where the concentration of benzene was below saturation, a weighed amount of benzene, in a small glass cup, was added to the reaction mixture while the latter was stirred in a creased flask. Glass-capped ampoules were rapidly filled with aliquots of this solution.

Determination of the solubility of benzene

A section of calibrated 10-mm. tubing was sealed onto a bulb of approximately 500-ml. capacity. About 450 ml. of the acid or salt solution was introduced into this flask, which was then thermostatted. A measured volume of benzene was added, the flask stoppered, and shaken vigorously. The shaking was repeated at intervals for several hours, and then the two phases were allowed to separate. Special attention was given to avoiding small droplets of benzene on the sides of the flask. The measurement was completed by introducing an additional amount of acid or salt solution sufficient to force the undissolved benzene into the calibrated 10-mm. tubing. A measurement of the height of the layer of benzene, together with the measurement of the volume initially introduced, allowed a calculation of the solubility of benzene in the solution.

Rate determination in acetic acid

Most of these experiments were carried out in a mixture of 97 per cent acetic acid and 3 per cent water. The analytical procedures were the same as those previously outlined.

Potentiometric titrations in 97 per cent acetic acid

The cell used was that developed by Watanabe (27). It was in general similar to that described by Conant and Hall (9), except that (*a*) a glass electrode was introduced into the titration compartment and (*b*) the entire system contained only a single solvent. The cell was the following:

Glass electrode		*Titration compartment*		
chloranil		$HClO_4$		chloranil
chloranil hydroquinone				chloranil hydroquinone
platinum	glass	titrating solution	salt bridge	platinum
buffer				buffer

The buffer in the glass electrode and the final cell compartment was a 2:1 sulfuric acid–urea mixture. The salt bridge was an inverted U-tube, stoppered at both ends with glass plugs, and filled with a supersaturated solution of potassium methanesulfonate in 97 per cent acetic acid. Gelatin was added to prevent the salt from crystallizing out.

The titration assembly consisted of the cell described above and of an amplification system. This was a thermionic amplifier built by Mr. C. Bard and Mr. J. Bjorkland from a design by Pennick (20); the amplifier was made available through the courtesy of Dr. George Shaeffer. The glass electrodes were Leeds & Northrup's Student No. 1199 of low resistance and low asymmetry.

DISCUSSION

Anion catalysis

The data illustrated in figure 2 show that, at least up to a concentration of 0.9 equiv./liter, the accelerating effect of perchlorates is as large as that of perchloric acid. It should be noted (table 4) that, for all these solutions, the solubility of benzene is essentially constant at approximately 0.026 mole/liter. The changes in rate are therefore (as a first approximation) independent of the effect of the environment on the activity coefficient of benzene. Furthermore, the effect of lanthanum perchlorate is identical with that of sodium perchlorate at the same perchlorate-ion concentration, and not at the same ionic strength. Finally, the data of table 8 show that, in the concentration range from 0 to 1 M, the rate follows equation 3, i.e., the rate is linear in the perchlorate-ion concentration. It therefore follows that the reaction is catalyzed by perchlorate ion. A consideration of the results with nitric acid, sodium nitrate, and lithium nitrate suggests that the reaction is likewise catalyzed by nitrates.

In solutions of perchlorates more concentrated than 1 M, perchloric acid is more effective in accelerating the mercuration than is sodium perchlorate. But the effect of the salt is nevertheless large; 6 M (saturated) sodium perchlorate increases the observed rate almost thirtyfold, and increases the second-order rate constant almost a hundredfold, over the rate at infinite dilution. Although the effect of 6 M perchloric acid is greater by an order of magnitude than that of the corresponding salt, it seems probable that this difference represents simply a specific ionic effect. Certainly a large part of the effect of perchloric acid, even in concentrated solutions, is due to the perchlorate ion; if there is any effect of acidity as such, it operates only at the highest acid concentrations. Similarly, the effect of 9 M lithium nitrate is comparable to that of a similar concentration of nitric acid. If the incomplete dissociation (21) of nitric acid is taken into account, the effect of the acid would, it is true, considerably exceed that of the salt of the same nitrate-ion concentration. Furthermore, the effect of acids and salts, in such concentrated solutions, on the kinetic activity coefficient (2) ratio,

$$\frac{\gamma_{\text{benzene}}\,\gamma_{\text{Hg}^{++}}\,\gamma_{\text{NO}_3^-}}{\gamma_{\text{complex}}}$$

may be far from trivial. Nevertheless, the salts have a strong accelerating effect, and as a first approximation a nitrate-ion catalysis accounts for the effect of both acid and salts.

"Salt effect"

Since the effect of anion concentration upon the rate of the reaction is large, it is desirable to explain it. The first possibility to be considered is that the reaction represents a large, nonspecific salt effect (2) upon the reaction between benzene and mercuric ion. This possibility cannot be excluded, even though the reaction rate depends upon the anion concentration and not upon the ionic strength. The work of Brønsted and Delbanco (3), and more especially the recent work of Olson and Simonson (19), has shown that the rates of reaction of polyvalent ions (like Hg^{++}) depend not upon ionic strength but upon the concentration of ions of charge opposite to that of the polyvalent reactant. Thus the rate of mercuration in perchlorate solution might be expected to depend upon the perchlorate-ion concentration, and equation 3 might be expected to hold, even though the effect of salt is merely that of a nonspecific salt effect. Furthermore, the differences in rate between acid and salt in concentrated solutions may be interpreted in terms of specific ion effects. This explanation, however, seems inadequate. The Brønsted–Bjerrum salt effects (2) cannot explain any large acceleration of a reaction between an ion and a neutral molecule; in any event a "salt effect" of such magnitude as that here described requires further discussion.

The possibility that mercuration proceeds by way of a chain involving mercurous ion can be effectively disposed of, since in nitric acid solution the rate of mercuration of benzene is essentially identical in the presence and in the absence of oxides of nitrogen (28). Furthermore, the possibility that mercuration proceeds by way of a "π-complex" between mercuric ion and benzene in no way helps in explaining the effect upon the reaction of neutral salts.

Possible mechanisms

It has already been shown that the activated complex probably consists of a benzene molecule, a mercuric ion, and an anion (e.g., perchlorate); the details of the mechanism depend upon the geometric arrangement and function of these entities. The authors have been unable to arrive at a totally satisfactory explanation of the experimental facts. However, an attractive possibility, here presented for purposes of discussion, is a mechanism in which a complex of mercuric ion and an anion reacts directly with benzene. Presumably mercuric ion in aqueous solution is tetrahedrally coördinated with water molecules. In the absence of a high concentration of anions, this hydrated ion may react with benzene; a water molecule is displaced when the carbon-to-mercury bond is formed. It is here tentatively assumed that, in high concentrations of neutral salts, a small amount of a complex[6] is formed in which one water molecule is replaced by an anion.[7]

The hypothesis under consideration suggests that those complexes of mercuric ion (such as the chloride and acetate) which contain anions more tightly bound than water necessarily react slowly; the coördinating anion cannot easily be displaced, and the reagent is less strongly electrophilic than the doubly charged aquo ion. Those complexes which contain anions less tightly bound than water are present only in low concentration; however, they react very rapidly. It should be especially easy to displace a perchlorate ion, for example, from its coördination with mercury. It is not possible to predict, *a priori*, that the overall reaction rate of an unstable perchlorate complex will exceed that of the $Hg(H_2O)_4^{++}$ ion; however, there is no reason why such should not be the fact. Further, it is interesting to speculate upon the possibility that the displacement reactions, discussed above, occur with stereochemical inversion about the tetrahedrally coördinated mercuric ion.

[6] Although nitrate and perchlorate ions do not generally coördinate with cations, evidence for $Hg(ClO_4)^+$ has been obtained (15). However, the favorable equilibrium constant for $Hg(NO_3)^+$ which has been reported (11) appears improbable; a careful analysis of the data shows that the constant can be quantitatively correct only if the ionic strength principle holds in 3 M salt solutions.

[7] The ion $(H_2O)_3Hg(ClO_4)^+$ may be regarded as an ion pair or as a trihydrated mercuric ion stabilized by an anion; at the present time, the distinction between these concepts and that of a true complex is vague.

The hypothesis for the mercuration advanced above explains why those anions which are least tightly bound accelerate the reaction most, and why the rate constant, in dilute solutions of perchlorates, is linear in the anion concentration. The hypothesis, further, is consistent with the orientation effects observed for mercuration (13). Specifically, mercuration by way of a complex of an anion with mercuric ion explains the ortho mercuration of benzoic acid (5).

Analogs

There are at least two other reactions between neutral molecules and ions which show "salt effects" comparable to those found in mercuration. The reactions in question are the hydrolysis of acetal (22) and of cyanamide (8). Not only is there a strong positive "salt effect" for the latter reaction, but dichloroacetate and acetate ions are catalysts for the hydrolysis (26). Mrs. Kilpatrick (12) has thoroughly discussed the mechanism of this and related reactions; although her mechanisms may not be correct in all details, they parallel to some extent that suggested here for mercuration. Specifically, she postulated that the activated complex for cyanamide hydrolysis may contain a dichloroacetate or trichloroacetate ion. The decrease in reaction rate found at high concentrations of halogen acids was explained (12, 26) by postulating stable complexes between cyanamide and halogen acids. The work of Michael (16) suggests that $NH_2CCl{=}NH$ hydrolyzes rapidly; if the compound $NH_2CCl{=}NH_2^+$ hydrolyzes slowly (12), all the results with hydrochloric acid can be explained.

A parallel mechanism for acetal hydrolysis, involving catalysis by undissociated acids or the transient formation of hemiacetal chlorides (1, 17) etc., can similarly be formulated.

For each case, then, for which a large neutral salt effect has been found for a reaction between a neutral molecule and an ion, an explanation of the salt effect can be advanced in terms of an activated complex which includes an anion. It is, however, rather surprising that the magnitude of the anion effect, in all three cases here under discussion, should be so nearly the same.

Additional problems

Even if the tentative hypothesis here advanced for mercuration is correct, there are several unanswered questions. Perhaps the most important of these concerns the fact that, in concentrated solutions of nitric and especially of perchloric acid, the expression $(k_{obs} - k_0)$ increases more than linearly with the concentration of the anion. Perhaps this will eventually prove to be true acid catalysis. Or perhaps it may prove to be a general law of concentrated solutions (25), comparable to those effects on the acidity of concentrated aqueous solutions of acids which find expression in the H_0 function (10). It is because of uncertainties in the physical chemistry of concentrated solutions that the rate of mercuration in concentrated nitric acid solution has been plotted against the gross concentration of nitrate ions rather than against the true concentration of nitrate ions.

SUMMARY

1. The rate of mercuration of benzene in aqueous solution is first order in mercuric ion and first order in benzene.

2. The second-order rate constant is increased many hundred fold by concentrated solutions of perchloric, sulfuric, or nitric acid; in the concentration range 0–3 *M* the second-order rate constant is increased as much by the addition of solutions of sodium perchlorate or nitrate as by the corresponding acid. Chlorides strongly inhibit the reaction.

3. In 97 per cent acetic acid as solvent, mercuration is strongly catalyzed by perchloric acid and weakly catalyzed by sodium perchlorate.

4. The reaction thus appears to be catalyzed by anions.

5. These facts are tentatively correlated with a mechanism in which mercuration takes place by way of a mercuric ion complexed with the catalyzing anion.

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DISCUSSION

F. A. **Long** (Cornell University): Concerning the data on cyanamide and acetal discussed in this paper, it is worth pointing out that the Brønsted primary salt effect for hydrolysis reactions can actually be relatively large, if the reactant nonelectrolyte is strongly salted out. An example of this is the hydrolysis of acetal in sodium chloride in which Riesch and Kilpatrick (J. Phys. Chem. **39,** 561 (1935)) report a k/k_0 value of 2.49 in 1 *M* salt. According to data of Olson and Tong (J. Am. Chem. Soc. **66,** 1555 (1944)), the activity coefficient of acetal in 1 *M* sodium chloride is about 1.8. Thus the strong salting out of the acetal accounts for most of the observed large salt effect.

One may frequently expect large salt effects for reactions where the rate is proportional to the Hammett acidity function. An example of this is the hydrolysis of β-propiolactone, for which addition of 4 *M* sodium perchlorate increases the rate by a factor of 15 and yet this large increase is consistent with the measured increase in H_0 caused by this same amount of salt. It is worth noting that Hammett and Paul (J. Am. Chem. Soc. **56,** 827 (1934)) showed that the rate of hydrolysis of cyanamide varies with H_0 and that the data for solutions of nitric acid, and for nitric acid and sodium nitrate mixtures, fit the H_0 function rather well.

The acid-catalyzed hydrolysis of β-lactones is independent of acid concentration up to acid concentrations of over 1 *M*. With still higher acid concentrations, the rate increases very rapidly and turns out to be proportional to H_0. It occurs to me that a similar explanation might hold for the mercuration of benzene. Under this postulate the effect shown by salts and acids up to about 2 *M* concentration would be due simply to a salt effect on the bimolecular reaction, and beyond perhaps 2 *M* acid, an acid-catalyzed reaction enters for which the rate is proportional to H_0, i.e., $k_{\text{obs}} = k_0 + k_1H_0$. To test this crudely, I have taken the data given in this paper for the three pure acids, perchloric, sulfuric, and nitric, and have applied the above equation, using a value of $k_0 = 4$. The data for sulfuric and perchloric acids fit the H_0 plot rather nicely and give a slope of unity. The data for nitric acid fall somewhat below this line but appear to give about the same slope. It might be interesting to put this hypothesis to further test by running further experiments in other concentrated acid solutions and also in 6 *M* mixtures of perchloric acid and sodium perchlorate for which the H_0 data are known from the work of Dodson and Harbottle.

F. H. **Westheimer** (University of Chicago), in reply to Dr. Long: Professor Long has suggested that the mercuration reaction proceeds by two mechanisms, one of which is independent of acidity but has a strong positive salt effect, whereas the other depends upon acidity as expressed by the H_0 function, and shows the strong salt effects associated with that function. His explanation may, of course, turn out to be correct. But it has been rejected here because we have been unable to assign any important role to the hydrogen ion in a detailed mechanism for the mercuration reaction.

M. **Kilpatrick** (Illinois Institute of Technology): When the logarithm of the first-order velocity constant for the hydrolysis of cyanamide in hydrochloric acid solution is plotted *vs.* the H_0 function of Hammett (Chem. Revs. **16,** 67 (1935))

$$H_0 = -\log a_{H^+} f_B / f_{BH^+}$$

it is found that up to 1 M acid the points lie on a line of slope -1, but at higher acid concentrations they rapidly drop below it.

For the acetal reaction it should be pointed out that the H_0 relation is not obeyed, but the suggestion that acids stronger than hydronium ion might catalyze the reaction was examined in 1935 (Riesch and Kilpatrick: J. Phys. Chem. **39**, 561 (1935)). For Westheimer anion catalysis we would need intermediates with ClO_4^-, Cl^-, NO_3^-, $C_6H_5SO_3^-$, etc., and I am particularly unhappy in regard to the perchlorate complex. I fail to see how perchlorate ion can replace water and make a weaker bond.

R. P. **Bell** (Oxford University, England): In my opinion there is not necessarily a sharp dividing line between an interpretation in terms of a specific salt effect and one in terms of anion catalysis. As Brønsted pointed out many years ago, a reaction involving cations will be specifically affected by anions in the solution as soon as the salt concentration rises above the very dilute range in which the limiting laws hold. The modern interpretation of this is that the anions concerned will be sufficiently close either to the initial reactant or to the activated complex (or to both) for their individual properties to be important, and it is clearly difficult to distinguish this from a mechanism involving complex formation at some stage. On the other hand it seems unsatisfactory to use a picture involving covalent or coördinate binding when the maximum effect is shown by the perchlorate ion, an ion which shows great reluctance to enter into such binding. It seems more likely that the effects observed can be related to some physical characteristics of the anions, such as their polarizability. Physical interactions could become very large at high salt concentrations with a multiply-charged cation such as Hg^{++}.

Westheimer has considered primarily the interaction of the mercuric ion with anions in the absence of the other reactant (benzene) and concludes that anions having weak interaction will be most effective. This picture cannot be complete in itself, since an anion which interacts with mercuric ion much less strongly than a water molecule will be unable to replace the water molecule, and thus could not cause any large effect. It seems possible that the important interaction is between the anion and the activated complex, where a strong interaction will favor reaction. A similar effect occurs in the accelerating power of cations (e.g., cupric ion) in the bromination of some β-keto esters (K. J. Pedersen: Acta Chem. Scand. **2,** 252 (1948)). Pedersen has explained this in terms of interaction between the cations and the keto estèr, but it is more likely that the important interaction is between the cation and the activated complex, which will resemble the enol form of the ester. In the mercuration reaction we do not yet know enough about the activated complex to picture the effect of anions, but it should be possible by optical or electrochemical methods to detect whether mercuric ion alone interacts with perchlorate. I believe that no marked interaction would be found in this system. On the other hand, the retarding effect of chloride ion and acetate ion fits in well with their complex formation with mercuric ion, as pointed out by Westheimer.

THE RACEMIZATION OF THE DIMETHYL ESTER OF *l*-BROMOSUCCINIC ACID BY LITHIUM BROMIDE IN ACETONE[1]

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The addition of anhydrous lithium bromide to a solution of the optically active dimethyl ester of bromosuccinic acid in dry acetone causes racemization of the ester. If the logarithm of the rotation is plotted against the time in the usual way, as shown in figure 1, a straight line is obtained, showing that the reaction is pseudo-unimolecular. The specific rate of racemization, however, depends upon the concentrations of the lithium bromide and of the ester. The addition of other substances, such as lithium perchlorate and water, has pronounced effects upon the rate.

PREPARATION OF MATERIALS

The preparation of *l*-bromosuccinic acid from *l*-aspartic acid has been described previously (2). The dimethyl ester was prepared by refluxing for 2 hr. a solution of the acid in methyl alcohol to which 2 ml. of concentrated sulfuric acid had been added. Water was then added to the solution, and the ester was extracted with ether. After removal of the ether by distillation, the pale yellow liquid was dried over calcium sulfate and distilled under high vacuum to produce a colorless liquid. Microanalysis for carbon and hydrogen gave 32.02 ± 0.14 per cent and 4.20 ± 0.13 per cent, respectively. The theoretical values are 32.02 per cent and 4.03 per cent. In acetone $[\alpha]^{25°}_{5461} = -69.5°$.

The acetone was treated with potassium permanganate and potassium hydroxide and then distilled. Subsequently it was kept in contact with calcium sulfate in an evacuated system for 2 hr., with shaking at frequent intervals,

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