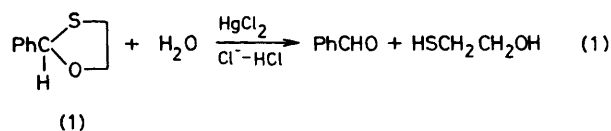


The Kinetics of the Mercury(II) Ion-promoted Hydrolysis of 2-Phenyl-1,3-oxathiolan and the Effects of pH and Chloride Ion Concentration

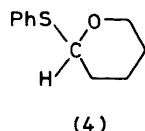
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The kinetics of the mercury(II) ion-promoted hydrolysis of 2-phenyl-1,3-oxathiolan (1) have been studied at various pH values and ambient chloride ion concentrations using dioxan–water and ethanol–water solvents. The products are always benzaldehyde and 2-mercaptoethanol. In the absence of mercury(II) ions the hydrolyses are comparatively slow. Hg^{2+} and (1) rapidly form a 1:1 adduct (2) in solution with a formation constant K ca. 1.8×10^3 at 25 °C. Depending upon the pH, either (2) or its protonated or deprotonated form reacts to give a hemiacetal whose subsequent hydrolysis is fast. The ring-opening process may involve intramolecular transfer of mercury-bond aquo species. In ethanol–water solvents (2) also reacts with ethanol to give a (small) quantity of an open-chain *OO*-acetal (3) whose subsequent hydrolysis is relatively slow. Added chloride ions reduce the competition from this ethanol reaction but also catalyse the hydrolysis of (3). Chloride ions have complex effects on the rate of the direct hydrolysis of (1) [via (2) and the hemiacetal] which suggest a reactivity sequence $\text{Hg}^{2+} \gg \text{HgCl}^+ < \text{HgCl}_2 > \text{HgCl}_3^- \gg \text{HgCl}_4^{2-}$. Mechanistic implications are discussed in the light of earlier work.

COMPARED with their *O*-analogues, *S*-acetals are hydrolysed relatively slowly in the presence of hydrogen acids.¹ Their hydrolyses are, however, promoted (often powerfully) by soft metal ions. There have been four kinetic studies of such promotion. Early studies^{2,3} of 2-phenyl-1,3-oxathiolan (1), 2-thiophenoxytetrahydropyran (4), and similar compounds promoted by mercury(II) chloride in the presence of a large excess of chloride ions and of hydrogen chloride [*e.g.*, equation (1)] showed that the



hydrolyses are of the first-order in the *S*-compound and in the total mercury(II) ion concentration. In these hydrolyses the slow step probably involves the reaction of an *S*-acetal– Hg^{II} ion adduct formed rapidly in low concentration. The effect of chloride and of hydrogen ion concentration on the promoted rate was not investigated. More recent studies^{1,4} show (i) that the hydrolysis of acetals like (4) proceed much faster at lower chloride concentrations, presumably owing to the presence of less HgCl_4^{2-} and HgCl_3^- and of more Hg^{2+} , HgCl^+ , and HgCl_2 species, and (ii) that the behaviour of 2,2-diphenyl-1,3-oxathiolan with thallium(III) ions implies that the use of an excess of free Hg^{2+} ions with (1)



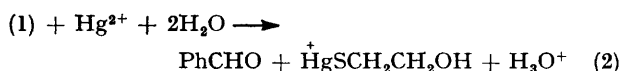
may lead to stoichiometric adduct formation followed by a rate-determining reaction of this adduct and to a consequent zero-order dependence of the hydrolysis on $[\text{Hg}^{2+}]$. We have now examined this possibility together with other features of the reaction.

EXPERIMENTAL

Materials.—Compound (1) was prepared by Djerassi and Gorman's method⁵ and purified by distillation. It had b.p. 111 °C at 6.3 Torr, analysed correctly, and had the expected i.r. and n.m.r. spectra. It was stored at 0 °C. AnalaR dioxan was refluxed with sodium for several hours and then distilled. It was stored at 0 °C under nitrogen. Benzaldehyde was the B.D.H. reagent, distilled before use, and 2-mercaptoethanol and mercury(II) chloride were Fisons reagents. Deuterium oxide (99.7 atom %) was a B.O.C. product and water was doubly distilled. All other chemicals were of AnalaR grade. Stock solutions of mercury(II) perchlorate were prepared by dissolution of mercury(II) oxide in the calculated quantity of aqueous perchloric acid. The final pH was normally <4.

Kinetic Arrangements.—Compound (1) is only slightly soluble in water; dissolution was facilitated by adding (1) to the reaction mixtures as a solution in dioxan, ethanol, or acetonitrile. Stock solutions of (1) in dioxan, *etc.* were made up by weight and most of the final reaction mixtures contained either 1% (v/v) dioxan or 20% (v/v) ethanol. Monitoring the course of the reaction normally required use of the stopped-flow technique (modified⁴ Durrum–Gibson D-110 instrument) but reactions at very high chloride ion concentrations and certain reactions in the ethanolic solvents could be followed using a u.v. recording spectrophotometer (Unicam SP 1750 fitted with SP 1802 and SP 1805 accessories and thermostatted cell housing). Reactions were studied (see Figures and Tables) mainly at 25 °C at fixed ionic strength (NaClO_4) and at different chloride (KCl), hydrogen ion (HClO_4), and mercury(II) perchlorate concentrations. The pH of every reaction mixture was measured (Radiometer PHM64 pH meter) after completion of the hydrolysis and was always found to be close to that expected for the quantities of stock perchloric acid and mercury perchlorate used. The hydrogen ion concentration in a reaction mixture was normally greatly in excess of any generated by the hydrolysis [*e.g.* equation (2)] since the initial concentration of (1) was usually ca. $4 \times 10^{-5}\text{M}$. Measurements were normally made at 250 nm where benzaldehyde absorbs strongly and (1) weakly. In reaction mixtures with $[\text{Cl}^-] \gtrsim 0.01\text{M}$ the complex Hg^{II} chlorides interfere at 250 nm and measurements were made at 258

nm. Reactions were studied under pseudo-first-order conditions (Tables) and k_{obs} , the observed first-order rate



constant for benzaldehyde formation, was computed from the absorbance measurements using an interactive Fortran program. Values of k_{obs} were generally reproducible to within $\pm 5\%$ and the data showed accurate first-order behaviour of three or more half-lives.

Reaction Products.—Comparison of the spectra of reaction mixtures after completion of the hydrolyses with those of synthetic product mixtures showed good agreement, indicating effectively quantitative formation of benzaldehyde and 2-mercaptoethanol [equation (2)]. Benzaldehyde undergoes little equilibrium hydration, acetal, or hemiacetal formation in the reaction mixtures used.^{6,7} Previous preparative-scale experiments have shown benzaldehyde to be formed under conditions of high chloride ion concentration.² We have demonstrated its formation in hydrolyses promoted by $\text{Hg}(\text{ClO}_4)_2$ only: ether extraction of reaction mixtures followed by g.l.c. analysis (on Apiezon L) indicated a $\gtrsim 95\%$ yield.

N.m.r. Measurements.—N.m.r. spectra (Bruker WM250) at 30 °C of solutions in a 10% (v/v) dioxan- D_2O solvent showed that hydrolysis of (1) in the presence of Hg^{2+} ions at $\text{pH} \approx 2.5$ does not involve the exchange of the formyl proton: the product remains PhCHO and is not PhCDO .

RESULTS AND DISCUSSION

The Hydrogen Ion-catalysed Reaction.—The hydrogen ion-catalysed hydrolysis of (1) was always relatively very slow compared to the various mercury ion-promoted processes, except for those promoted reactions conducted in the presence of very high chloride and hydrogen ion concentrations. However, even with $[\text{Cl}^-] 0.2\text{M}$, $[\text{Hg}^{2+}] 4 \times 10^{-4}\text{M}$, and $[\text{H}_3\text{O}^+] 0.05\text{M}$, the promoted reaction in 1% (v/v) dioxan-water is still *ca.* 200-fold faster than that in the absence of mercury. The contribution of the purely hydrogen ion-catalysed reaction to the observed rates can therefore usually be ignored. For the hydrogen ion-catalysed hydrolysis at 25 °C and ionic strength $I 1.0\text{M}$, we find $k_{\text{H}_3\text{O}^+} 4.7 \times 10^{-3} \text{ l mol}^{-1} \text{ s}^{-1}$ in 1%

Mercury(II) Ion Promotion.—(i) *In the absence of added chloride ions.* In ethanolic solvents, in the presence of mercury(II) ions, two reactions can, in appropriate circumstances, be observed; one is usually complete before the other has much progressed. The first reaction lies in the stopped-flow region and produces much of the benzaldehyde. The second, slower reaction leads to the remainder of the benzaldehyde. How much of the product is formed *via* the slower reaction depends upon the proportion of ethanol in the solvent and, to some extent, upon the method of mixing of the components in making up a reaction mixture. With rapid (stopped-flow) mixing and 20% (v/v) ethanol, $< 5\%$ of the benzaldehyde is formed *via* the slower reaction, so that, in these conditions, it forms a minor part of the overall reaction. The slow reaction is also absent from hydrolyses in dioxan-water mixtures and is reduced in importance in ethanol-water in the presence of added chloride ions (see below).

The slower reaction can conveniently be studied using

TABLE I

The hydrolysis of (3) in ethanol-water solvents at 25 °C.

$[\text{OS-acetal}]_{\text{init}} 4.20 \times 10^{-5}\text{M}$; $I 0.10\text{M}$

(a) Dependence on $[\text{H}_3\text{O}^+]$ $[\text{Hg}^{2+}] 4.20 \times 10^{-4}\text{M}$; 20% (v/v) ethanol}

$10^3 [\text{H}_3\text{O}^+]/\text{M}$	0.65	1.30	2.20	4.10	6.65	7.75
$10^3 k_{\text{obs}}/\text{s}^{-1}$	1.67	3.50	5.70	10.0	16.9	19.7

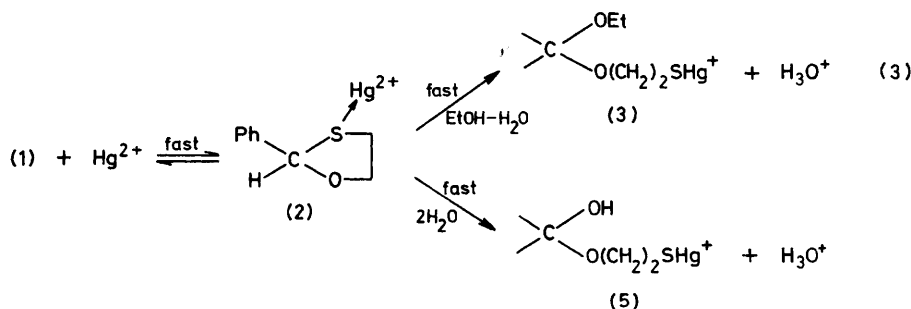
(b) Dependence on $[\text{Hg}^{2+}]$ $[\text{H}_3\text{O}^+] 7.75 \times 10^{-3}\text{M}$; 20% (v/v) ethanol}

$10^4 [\text{Hg}^{2+}]/\text{M}$	4.00	6.50	8.00	16.5	40.0
$10^3 k_{\text{obs}}/\text{s}^{-1}$	20.0	19.5	19.3	19.0	20.1

(c) Effect of ethanol content of solvent $[\text{Hg}^{2+}] 4.0 \times 10^{-4}\text{M}$; $[\text{H}_3\text{O}^+] 3 \times 10^{-3}\text{M}$

(v/v) Ethanol (%)	18	49	90
$[\text{PhCHO}]_{\text{total}}$ formed <i>via</i> (3) (%)	8	24	34
$10^3 k_{\text{obs}}/\text{s}^{-1}$	7.00	1.67	0.32

a recording spectrophotometer and initiating the hydrolysis by adding an aliquot portion of a stock solution of (1) in ethanol into the remainder of the reaction mixture in the thermostatted cell. Under such conditions the first reaction is complete before measurements



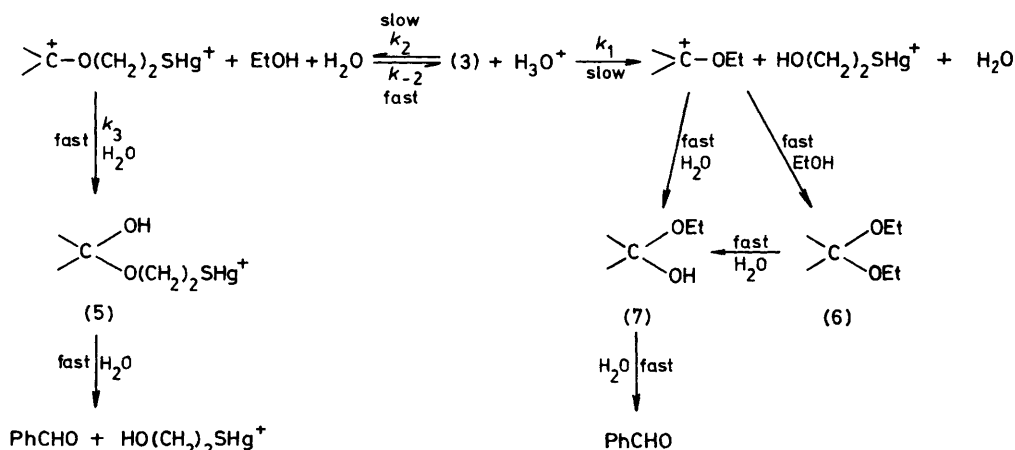
(v/v) dioxan-water. This value lies between the comparable values given by Fedor² ($1.1 \times 10^{-2} \text{ l mol}^{-1} \text{ s}^{-1}$ at 30 °C) and by Jensen⁸ ($2.6 \times 10^{-3} \text{ l mol}^{-1} \text{ s}^{-1}$ at 25 °C) and, in view of the temperature difference, is in fair agreement with the former.

begin and the slower reaction contributes *ca.* 10–20% of the overall benzaldehyde formation. At constant ionic strength the latter reaction is of zero order in $[\text{Hg}^{2+}]$ and of the first order in $[\text{H}_3\text{O}^+]$ (Table I).

We believe the slower reaction is the hydrolysis (Scheme

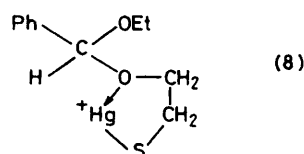
1) of (3), formed as shown in outline in equation (3). (Whatever the position of initial bond cleavage in the hydrogen ion-catalysed hydrolysis,⁸ in the mercury-promoted reaction the cleavage is reasonably assumed to involve the C-S bond.) Scheme 1 is based in outline on the mechanism recently proposed by Jensen *et al.*⁹ for

ethanol-water, to the major part of the aldehyde formed, has the dependencies on $[\text{Hg}^{2+}]$ and $[\text{H}_3\text{O}^+]$ shown in Figures 1 and 2. When Hg^{2+} is in sufficient excess over (1), k_{obs} becomes independent of $[\text{Hg}^{2+}]$. This is the type of result expected from our previous work on thallium(III) ion promotion.⁴ However, the pH profile, and especially



SCHEME 1

the Brønsted acid-catalysed hydrolysis of acetals of benzaldehyde and simple aliphatic alcohols. It is known¹⁰ that the rates of hydrolysis of the diethyl acetal (6) and its hemiacetal (7) are both fast compared with the observed rate of the slow reaction (Table 1) and it is assumed that the hydrolysis of (5) will also be fast by comparison. For simplicity not all the steps are included in Scheme 1, but it can be shown to lead to a pseudo-first-order rate constant, $k_{\text{obs}} = \{k_1 + k_2k_3[\text{H}_2\text{O}]/(k_{-2}[\text{EtOH}] + k_3[\text{H}_2\text{O}])\}[\text{H}_3\text{O}^+]$. This expression is in



general agreement with our results (Table 1). The effects on k_{obs} of altering the ethanol content of the medium must, of course, arise partly from solvent (*e.g.*, dielectric constant) changes as well as from concentration changes.¹¹ Which of the paths controlled by k_1 and k_2 is the more important is difficult to assess. Previous work suggests that the more acidic and bulkier alcohol will leave preferentially under proton catalysis, but the position is here complicated by the positive charge on (3) and, as will be argued below, by the possibility that (3) may exist in the cyclic form (8). That the hydrolysis of (3) should be slower (*ca.* 40-fold) than that of (6) is clearly a reasonable result. There is little evidence from the first-order plots for the hydrolysis of (3) for the build-up of significant amounts of hemiacetal in this system.¹⁰

The fast reaction which, in dioxan-water, leads to a quantitative yield of benzaldehyde and, in 20% (v/v)

the steep rise in k_{obs} , found at the lowest hydrogen ion concentrations used, is an unexpected result and different from the behaviour found with thallium. In this pH

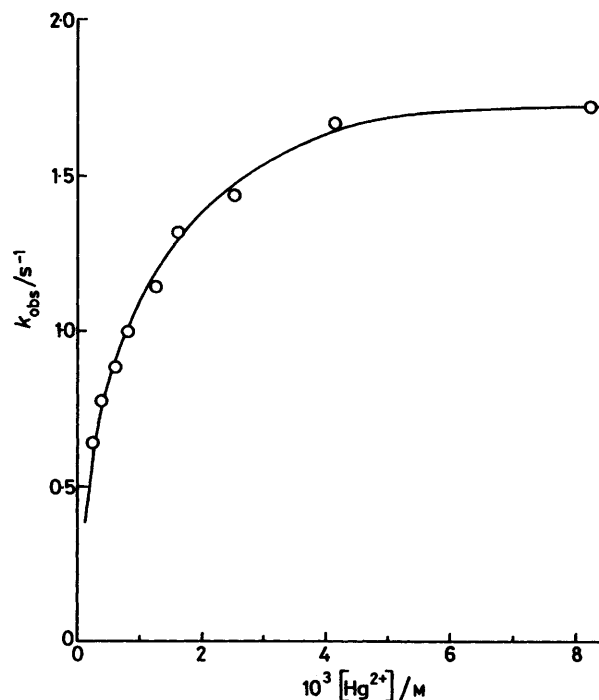


FIGURE 1 Dependence of the fast reaction on $[\text{Hg}^{2+}]$: $[\text{OS-acetal}]_{\text{init.}} 4.20 \times 10^{-5}\text{M}$; $[\text{H}_3\text{O}^+] 5.00 \times 10^{-2}\text{M}$; $I 1.00\text{M}$; temperature 25.0°C ; solvent 20% (v/v) ethanol-water

region (pH 2–3) equilibria (4) and (5) are known¹² to occur and analogous equilibria presumably obtain for Hg^{2+} co-ordinated to the OS-acetal or to other ligands.

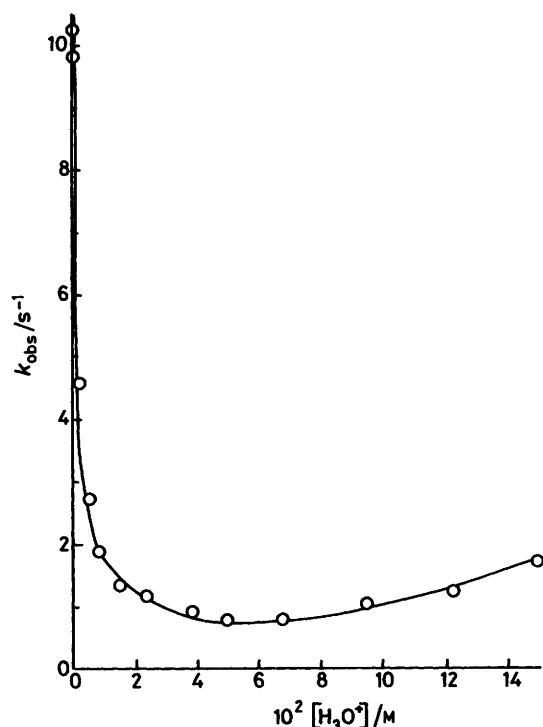
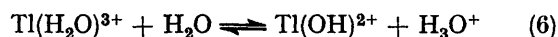
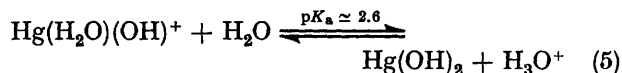
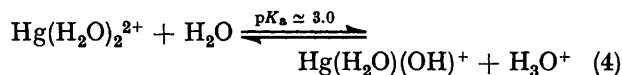


FIGURE 2 Dependence of the fast reaction on $[H_3O^+]$: $[OS\text{-}acetal]_{init.} 4.20 \times 10^{-5}M$; $[Hg^{2+}] 4.20 \times 10^{-4}M$; $I 1.00M$; temperature $25.0^\circ C$; solvent 20% (v/v) ethanol-water

Although this pH region has not been examined previously by us in any Hg^{2+} -promoted reaction, for Tl^{3+} the analogous ionisation (6) leads to a species which is a much less (rather than more) powerful promoter of *S*-acetal hydrolysis.⁴

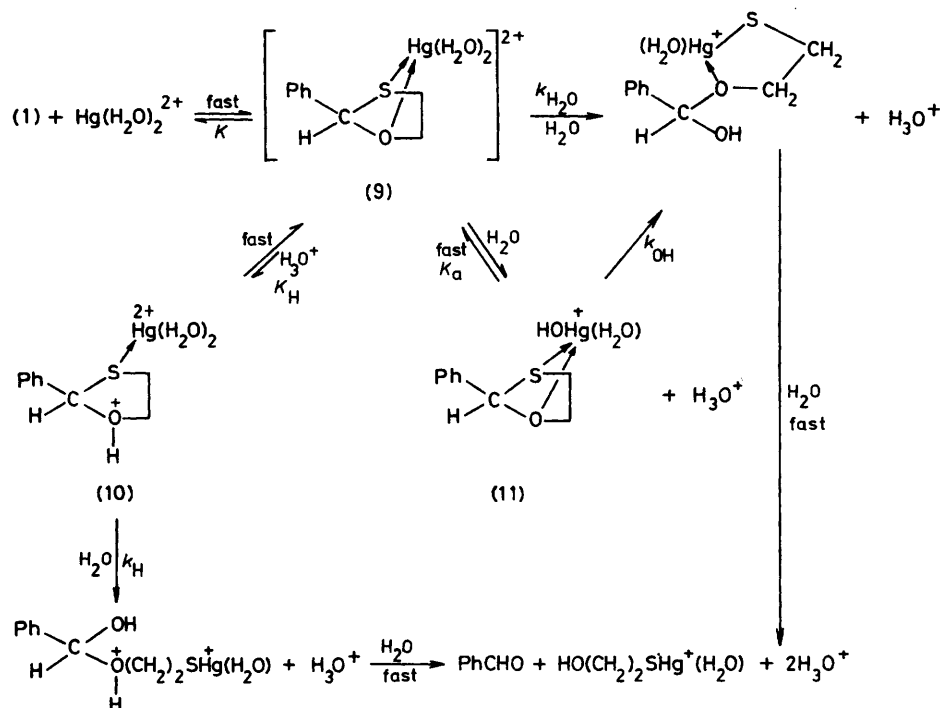
A superficially attractive explanation of the observed pH profile is that there occurs a very rapid (perhaps beyond the stopped flow region) formation of species (3) and (5) and that our observed fast reaction is the hydrolysis of the hemiacetal (5). Hemiacetal hydrolysis



is known¹⁰ to exhibit pH profiles similar to that in Figure 2, although the sharp upturn at high pH, due to the incursion of the hydroxide ion-catalysed path, has normally appeared at pH 5–6. This discrepancy could perhaps be explained by the positive charge on (5), although the minimum only moves to $pH \approx 3$ in the presence of chloride ions (Figure 3) which will neutralise the charge or even reverse it. The principal difficulty in accepting the foregoing explanation is that it is not easy to combine it with the observed equilibrium dependence of k_{obs} on $[Hg^{2+}]$. We believe therefore that the fast reaction is better explained by Scheme 2.

Scheme 2 envisages the rapid formation of Hg^{II} -OS-acetal adducts which can exist in three states of ionisation (9)–(11), each state providing a route to the hemiacetal. Subsequent hydrolysis of the hemiacetal is assumed to be fast. The possibility that the formyl proton in (9) undergoes dissociation is excluded by the product analysis in deuterium oxide.

If we assume as an approximation that $Hg(H_2O)_2^{2+}$



SCHEME 2

and $\text{Hg}(\text{H}_2\text{O})\text{OH}^+$ form adducts with (1) about equally readily (as they may in view of results for HgCl^+ given below) then Scheme 2 leads to equation (8) for k_{obs} , derived from the general equation (7).

$$d[\text{PhCHO}]/dt = k_{\text{H}_2\text{O}}[(9)] + k_{\text{OH}}[(11)] + k_{\text{H}}[(10)] \quad (7)$$

In the presence of a sufficient excess of Hg^{II} ions, all

$$k_{\text{obs}} = \left(k_{\text{H}_2\text{O}} \frac{[\text{H}_3\text{O}^+]}{[\text{H}_3\text{O}^+] + K_{\text{a}}} + k_{\text{OH}} \frac{K_{\text{a}}}{[\text{H}_3\text{O}^+] + K_{\text{a}}} + k_{\text{H}} K_{\text{H}} [\text{H}_3\text{O}^+] \right) \frac{K[\text{Hg}^{\text{II}}]}{1 + K[\text{Hg}^{\text{II}}]} \quad (8)$$

the acetal will be converted into adducts and the rate will become independent of $[\text{Hg}^{\text{II}}]$, as found. At $[\text{H}_3\text{O}^+] 5 \times 10^{-2}\text{M}$, where little (11) is probably formed and the adducts are therefore principally (9), our results (Figure 1) give $K = [(9)]/[(1)][\text{Hg}^{2+}] = \text{ca. } 1.8 \times 10^3 \text{ l mol}^{-1}$. Equation (8) is compatible with the three regions of promotion observed at fixed mercury but varying hydrogen ion concentration (Figure 2). Our results for 20% (v/v) ethanol-water at $I 1.0$ are well fitted with $k_{\text{H}_2\text{O}} 0.28 \text{ s}^{-1}$, $k_{\text{OH}} 9.6 \text{ s}^{-1}$, $K_{\text{H}} k_{\text{H}} 2.2 \text{ l mol}^{-1} \text{ s}^{-1}$, and $K_{\text{a}} 8 \times 10^{-4} \text{ mol l}^{-1}$. The k_{H} route is less prominent at lower ionic strengths. A similar behaviour pattern is found in 1% (v/v) dioxan-water with all the rate constants about two-

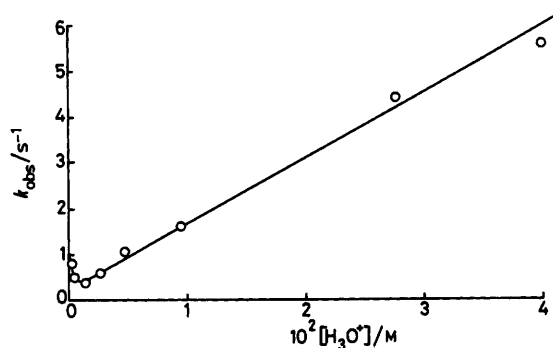


FIGURE 3 Dependence of the fast reaction on $[\text{H}_3\text{O}^+]$ in the presence of chloride ions: $[\text{OS-acetal}]_{\text{init.}} 4.20 \times 10^{-5}\text{M}$; $[\text{Hg}^{2+}] 4.20 \times 10^{-4}\text{M}$; $[\text{Cl}^-] 5.00 \times 10^{-3}\text{M}$; $I 0.20\text{M}$; temperature 25.0°C ; solvent 2% (v/v) ethanol-water

fold larger. The greater reactivity of (11) compared with (9) we attribute to a greater intramolecularity of the reaction of (11) owing to the availability of the OH^- group on the mercury atom. Considerable nucleophilicity for OH^- groups attached to Hg^{II} has been observed by other workers.¹ We suggest that in most of the adducts in Scheme 2 the mercury species is co-ordinated to both the O- and the S-atoms of the acetal (as shown). This will facilitate intramolecular transfer of Hg-bound aquo-species to carbon. The ΔS^\ddagger values for the pH regions where either the reaction of (11) or of (9) is dominant are -3.4 and $-42 \text{ J K}^{-1} \text{ mol}^{-1}$, respectively (Table 2). The former in particular suggests a unimolecular (and therefore possibly intramolecular) decomposition for (11). For the region where k_{obs} increases with $[\text{H}_3\text{O}^+]$ we consider it reasonable to assume that some (10) will form and be more reactive than (9).

(ii) *In the presence of chloride ions.* In their previous

study Fedor and De² used relatively very high ambient chloride and hydrogen ion concentrations in examining the effect of Hg^{II} ions. Their solvent was water containing a small amount of acetonitrile. We have tried to repeat their results using similar conditions but have been unable to do so: we have always obtained k_{obs} values significantly greater than theirs. For example

using 1% (v/v) acetonitrile-water at 25°C with $[\text{Cl}^-] 0.19\text{M}$, $I 0.19\text{M}$, $[\text{H}_3\text{O}^+] 0.149\text{M}$ (as HCl), and $[\text{Hg}^{\text{II}}]_{\text{total}} 4.0 \times 10^{-4}\text{M}$, we find $k_{\text{obs}} 4.3 \times 10^{-2} \text{ s}^{-1}$ compared with their $k_{\text{obs}} 3.6 \times 10^{-3} \text{ s}^{-1}$ at 30°C . It is, in fact, difficult

TABLE 2

Effects of temperature on the fast hydrolysis of (1). $[\text{OS-acetal}]_{\text{init.}} 4.2 \times 10^{-5}\text{M}$; $[\text{Hg}^{2+}] 4.2 \times 10^{-4}\text{M}$; $I 1.00\text{M}$; solvent 20% (v/v) ethanol-water

(i) $[\text{H}_3\text{O}^+] 1.00 \times 10^{-3}\text{M}$

$T/^\circ\text{C}$	8.00	25.0	39.1
$k_{\text{obs}}/\text{s}^{-1}$	1.86	9.88	34.8
$E_{\text{a}} 68 \pm 1 \text{ kJ mol}^{-1}$; $\Delta S^\ddagger -3.4 \pm 0.8 \text{ J K}^{-1} \text{ mol}^{-1}$			

(ii) $[\text{H}_3\text{O}^+] 6.60 \times 10^{-2}\text{M}$

$T/^\circ\text{C}$	6.40	25.0	39.1
$10k_{\text{obs}}/\text{s}^{-1}$	1.33	7.65	23.8
$E_{\text{a}} 65 \pm 1 \text{ kJ mol}^{-1}$; $\Delta S^\ddagger -42 \pm 5 \text{ J K}^{-1} \text{ mol}^{-1}$			

to be certain of the exact values of $[\text{Cl}^-]$ and $[\text{H}_3\text{O}^+]$ used by Fedor and De. It is clear, however, that at very high chloride ion concentrations the reaction is much slower than in the total absence of chloride. Using a 1% (v/v) dioxan-water solvent we have varied $[\text{Cl}^-]$ under conditions of fixed ionic strength, fixed total mercury concentration and at two fixed values of $[\text{H}_3\text{O}^+]$ (Figure

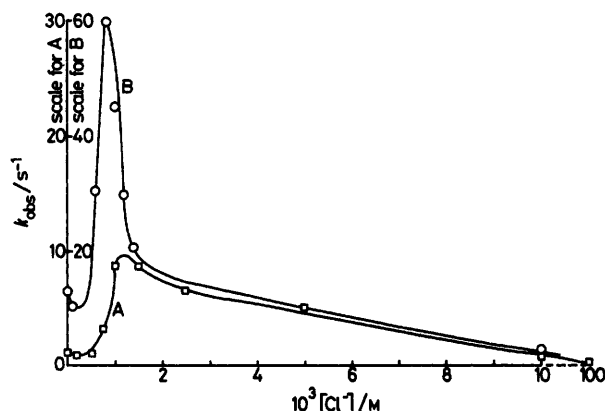
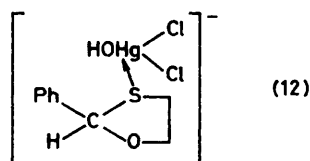


FIGURE 4 The effect of added chloride ions on the hydrolysis of (1): $[\text{OS-acetal}]_{\text{init.}} 4.20 \times 10^{-5}\text{M}$; $[\text{Hg}^{2+}] 4.20 \times 10^{-4}\text{M}$; $I 0.20\text{M}$; temperature 25.0°C ; solvent 1% (v/v) dioxan-water. A, $[\text{H}_3\text{O}^+] 5.00 \times 10^{-3}\text{M}$; B, $[\text{H}_3\text{O}^+] 5.00 \times 10^{-4}\text{M}$

4). The important features of our surprising results are (a) the very sharp increase in k_{obs} found, at both acidities, when $[\text{Cl}^-] \simeq 2[\text{Hg}^{\text{II}}]$, (b) the beneficial effects of chloride ions over a fairly wide concentration range, (c) the eventual fall in k_{obs} at very high chloride ion concen-

trations, and (d) the greater prominence of the hydrogen ion-catalysed route in the presence of chloride ions (Figure 3).

Fact (c) is in agreement with our previous work with compounds like (4) and doubtless arises because the species HgCl_3^- and HgCl_4^{2-} are less able than Hg^{2+} ,



HgCl^+ , and HgCl_2 to form adducts with OS-acetals. If K in Scheme 2 is reduced, k_{obs} will fall and the rate will eventually become first order in $[\text{Hg}^{\text{II}}]$, as found by Fedor.² The sharp and temporary increase in k_{obs} at intermediate values of $[\text{Cl}^-]$ is difficult to explain. We did not observe it¹ with compounds like (4) and it has

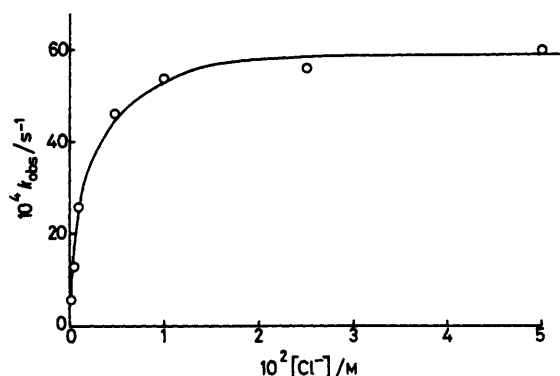


FIGURE 5 The effect of added chloride ions on the hydrolysis of (3): $[\text{OS-acetal}]$ $4.20 \times 10^{-5}\text{M}$; $[\text{Hg}^{2+}]$ $4.20 \times 10^{-4}\text{M}$; I 0.10M ; $[\text{H}_3\text{O}^+]$ $2.20 \times 10^{-4}\text{M}$; temperature 25.0°C ; solvent 20% (v/v) ethanol-water

been noted, in various contexts,¹³ that the addition of negatively charged ligands to soft metal ions (Hg^{2+} , Tl^{3+}) tends, if anything, to reduce their reactivity towards neutral substrates. Our present results, and the relevant association constants,¹² indicate a sequence of reactivity: $\text{Hg}^{2+} \simeq \text{HgCl}^+ < \text{HgCl}_2 > \text{HgCl}_3^- \gg \text{HgCl}_4^{2-}$. We attribute the great reactivity of HgCl_2 to a more favourable rate of transfer of aquo-species from mercury to carbon in species such as (12).

That more efficient reaction with water occurs in the presence of chloride ions is attested by a fall in the yield of (3) in ethanol-water solvents containing these ions. The enhancement by chloride ions of the hydrogen ion-catalysed hydrolysis of (3) has quite a different rate profile (Figure 5) from that found for the hydrolysis of (1) (Figure 4).

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REFERENCES

- D. P. N. Satchell, *Chem. Soc. Rev.*, 1977, **6**, 345.
- N. C. De and L. R. Fedor, *J. Am. Chem. Soc.*, 1968, **90**, 7266.
- L. R. Fedor and B. S. R. Murty, *J. Am. Chem. Soc.*, 1973, **95**, 8407.
- D. P. N. Satchell and T. J. Weil, *J. Chem. Soc., Perkin Trans. 2*, 1980, 1191.
- C. Djerassi and M. Gorman, *J. Am. Chem. Soc.*, 1953, **75**, 3704.
- P. Greenzaid, *J. Org. Chem.*, 1973, **38**, 3164.
- Y. Ogata and A. Kawasaki in 'The Chemistry of the Carbonyl Group,' ed. J. Zabricky, Interscience, New York, 1970, vol. 2, ch. 1.
- J. L. Jensen and W. P. Jencks, *J. Am. Chem. Soc.*, 1979, **101**, 1476.
- J. L. Jensen, L. R. Herold, P. A. Lenz, S. Trusty, V. Sergi, K. Bell, and P. Rogers, *J. Am. Chem. Soc.*, 1979, **101**, 4672.
- J. L. Jensen and P. A. Lenz, *J. Am. Chem. Soc.*, 1978, **100**, 1291.
- E. H. Cordes, *Prog. Phys. Org. Chem.*, 1967, **4**, 24.
- L. G. Sillen and A. E. Martell, 'Stability Constants,' Chem. Soc. Spec. Publ. Nos. 17 (1964) and 25 (1971).
- D. P. N. Satchell and R. S. Satchell, *Ann. Rep. Prog. Chem.*, 1978, **A**, 25.