Reaction of Ethoxide with *p*-Nitrophenyl Benzenesulphonate. Catalysis and Inhibition by Alkali Metal Ions; Contrast with *p*-Nitrophenyl Diphenylphosphinate

Erwin Buncel* and Marko J. Pregel

Department of Chemistry, Queen's University, Kingston, Ontario, Canada K7L 3N6

In the reaction of *p*-nitrophenyl benzenesulphonate with ethoxide in ethanol, some alkali metals (Na^+, K^+, Cs^+) cause a rate acceleration, while Li^+ causes a rate retardation; the results are analysed in terms of interactions of metal ions with the ground state and the transition state of the reaction.

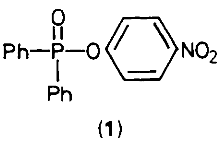
The effects of alkali metal ions on organic reactions have not received a great deal of attention, despite the fact that these ions are ubiquitous in biological systems. We recently reported on the observation of alkali metal ion catalysis in the reaction of *p*-nitrophenyl diphenylphosphinate (**1**) with ethoxide ion in ethanol, and on the consequent rate retarding effect of crown ethers and cryptands on this process.[1]

We now report on the contrasting findings that we have observed in the reaction of alkali metal ethoxides with *p*-nitrophenyl benzenesulphonate (**2**). Comparison of the two systems is expected to highlight the differences in reactivity between phosphorus-[2] and sulphur-based[3] esters with similar structures.

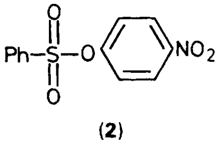
In the phosphinate ester system, the reactivity order was found to be LiOEt > NaOEt > KOEt > EtO^- and crown ethers and cryptands exhibited a rate retarding effect on the reactions of LiOEt, NaOEt, and KOEt.[1]

For the *p*-nitrophenyl benzenesulphonate–ethoxide system, the reactivity order is different (KOEt > CsOEt > NaOEt > EtO^- > LiOEt), as shown in Figure 1. The data for alkali metal ethoxides in the presence of excess complexing agents represent the reaction of free ethoxide, since complexing agents break up alkali metal ethoxide ion pairs by sequestering the metal ion as a coronate or cryptate.† Thus, the observed rate retarding effect of complexing agents in the case of KOEt indicates that the KOEt ion pair is more reactive than free ethoxide. Conversely, added cryptand has a rate accelerating effect in the case of LiOEt, indicating that the LiOEt ion pair is less reactive than free ethoxide. This behaviour contrasts with that of the phosphinate ester (**1**)–ethoxide system, where LiOEt, NaOEt, and KOEt are more reactive than free ethoxide ion.[1]

The overall kinetic effect of a given metal ion is the result of a balance of its association with the ground state and with the transition state.[4] Assuming that association of M^+ with the esters can be neglected, then the ground state association (*i.e.* $M^+ + EtO^- \rightleftarrows MOEt$) will be the same in systems (**1**) and (**2**). The observed reactivity differences must therefore be



(1)



(2)

---

† On addition of 2.2.2 cryptand to the reaction of KOEt with (**2**), the rate *decreases* until one equivalent of complexing agent is added, at which point a minimum value is reached and further addition has no effect. When 2.1.1 is added to the reaction of LiOEt with (**2**), the rate *increases* until one equivalent of cryptand is added, then levels off to a maximum value which is identical to the minimum value seen in the reaction of KOEt with excess 2.2.2. This is assumed to correspond to the reaction of free ethoxide with (**2**) since all metal ions should be complexed at these concentrations of added cryptands.

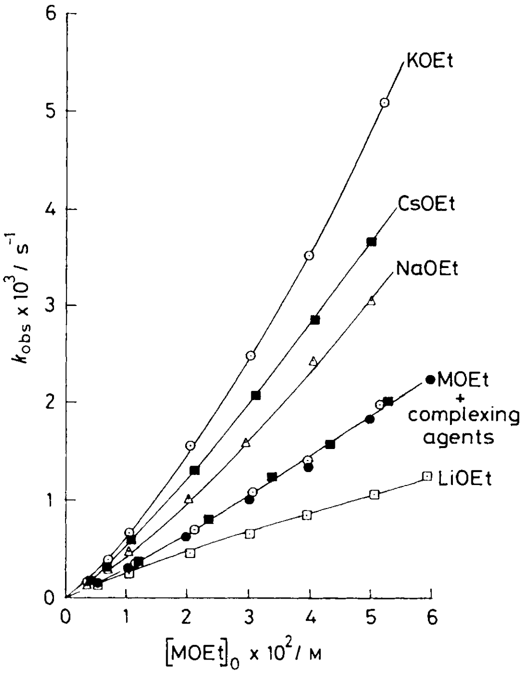


Figure 1. Variation of the observed rate constant with the total base concentration for the reaction of (**2**) with various ethoxide species in anhydrous ethanol at 25 °C. The line labelled 'MOEt + complexing agents' is made up of data for KOEt in the presence of excess 18-crown-6 (●), KOEt with excess 2.2.2 cryptand (⊙), and LiOEt with excess 2.1.1 cryptand (■).

explained by differences in the interactions of alkali metal ions with the two transition states.

Theoretical association constants of the metal ions for the transition states can be calculated from the kinetic data using a thermodynamic cycle.[5] The association constants of various metal ions which have been obtained in this way are compared in Table 1.

These values are intriguing because two of the values of association constants for the phosphinate transition state are much larger than all other values. More significantly, the trends in metal ion association constants are different. The order is $\mathrm{K^+ > Cs^+ > Na^+ > Li^+}$ for the sulphonate and $\mathrm{Li^+ > Na^+ > K^+ > Cs^+}$ for the phosphinate. The two esters are believed to react by similar mechanisms. However, the way in which the metal ion interacts with the transition state is apparently not the same in both cases. It is expected that the explanation underlying these differing results arises from differences in the geometric and electronic structures of the two transition states.

An obvious difference between the two transition states in the reactions of (**1**) and (**2**) is that they do not contain the same number of ligand oxygens. In the phosphinate transition state, the metal ion can interact with a single negatively charged oxygen. Simple cation–anion interactions are expected to result in stronger association with smaller cations, as observed. On the other hand, in the sulphonate transition state, interactions with two oxygens are possible. If the oxygens are far enough apart, interaction of a large metal ion with two oxygens (chelation) may be stronger than interaction of a small metal ion with only one oxygen, leading to a reverse trend in association constants with the transition state.

Another difference between the two transition states is the extent of charge delocalization. The phosphinate transition state has its negative charge localized on one oxygen. This allows strong interactions with metal ions. If the interaction of the transition state with the metal ion is stronger than the interaction of the metal ion with solvent, then a solvent molecule is displaced from the co-ordination sphere of the metal ion and a contact ion pair is formed. Association here involves the interaction of the transition state with bare metal ions, which increase in size in the order $\mathrm{Li^+ < Na^+ < K^+ < Cs^+}$.[7] Since the smaller cations should interact more strongly with the anionic transition state, this would lead to the association constants decreasing in the order $\mathrm{Li^+ > Na^+ > K^+ > Cs^+}$, which is identical to the observed ordering (Table 1).

Table 1. Association constants of transition states for metal ions ($K_{a'}$, mol^{-1} dm^3).

	$\mathrm{Li^+}$	$\mathrm{Na^+}$	$\mathrm{K^+}$	$\mathrm{Cs^+}$
Sulphonate transition state	133	284	430	409
Phosphinate transition state	5188	1207	347	266

In contrast, the sulphonate transition state, which has its charge delocalized over two oxygens, has weaker interactions with metal ions and is not able to displace a solvent molecule from the co-ordination sphere of the metal ion. The transition state interacts with solvated metal ions and these increase in size in the order $\mathrm{Cs^+ < K^+ < Na^+ < Li^+}$.[8] This would result in association constants decreasing in the order $\mathrm{Cs^+ > K^+ > Na^+ > Li^+}$. This is similar to the observed order, $\mathrm{K^+ \geqslant Cs^+ > Na^+ > Li^+}$ (Table 1).

Further work is in progress on metal ion effects in related systems (*e.g.* carboxylic esters) in order to investigate metal ion effects further and bring additional evidence to bear on the problem. The results of such studies may shed light on the role of metal ions in biological systems.[6]

This research was supported by the N.S.E.R.C. The awards to M. J. P. of a Postgraduate Scholarship by the N.S.E.R.C. and a Graduate Award by Queen's University are gratefully acknowledged.

Received, 31st May 1989; Com. 9/02265A

References

1 E. Buncel, E. J. Dunn, R. A. B. Bannard, and J. G. Purdon, *J. Chem. Soc., Chem. Commun.*, 1984, 162; E. J. Dunn and E. Buncel, *Can. J. Chem.*, 1989, **67**, 1440.

2 R. D. Cook, C. E. Diebert, W. Schwarz, P. C. Turley, and P. Haake, *J. Am. Chem. Soc.*, 1973, **95**, 8088; M. Skoog and W. P. Jencks, *ibid.*, 1983, **105**, 3356; N. Bourne, E. Chrystiuk, A. M. Davis, and A. Williams, *ibid.*, 1988, **110**, 1890.

3 J. L. Kice, C. Walters, and S. Bunton, *J. Org. Chem.*, 1974, **39**, 346; N. A. Suttle and A. Williams, *J. Chem. Soc., Perkin Trans. 2*, 1983, 1563.

4 E. Buncel and E. A. Symons, *J. Am. Chem. Soc.*, 1976, **98**, 656; E. Buncel and H. Wilson, *Acc. Chem. Res.*, 1979, **12**, 42; M. H. Abraham, P. L. Grellier, J.-L. M. Abboud, R. M. Doherty, and R. W. Taft, *Can. J. Chem.*, 1988, **66**, 2673; M. H. Abraham, P. L. Grellier, A. Nasehzadeh, and R. A. Walker, *J. Chem. Soc., Perkin Trans. 2*, 1988, 1717; R. L. Schowen, 'Catalytic Power and Transition-State Stabilization,' ch. 2 of 'Transition States of Biochemical Processes,' eds. R. D. Gandour and R. L. Schowen, Plenum Press, New York, 1978.

5 J. L. Kurz, *J. Am. Chem. Soc.*, 1963, **85**, 987.

6 C. H. Suelter, 'Monovalent Cations in Enzyme-Catalyzed Reactions,' ch. 7 of 'Metal Ions in Biological Systems,' vol. 3, ed. H. Sigel, Marcel Dekker, New York, 1974.

7 F. A. Cotton and G. Wilkinson, 'Advanced Inorganic Chemistry,' 5th edn., Wiley, New York, 1988.

8 Solvated radii were calculated using the Stokes equation from the data of R. Kay and D. Evans, *J. Phys. Chem.*, 1966, **70**, 2325.