

if k_1 (and k_{-1} since the equilibrium isotope effect is small) has $k_D/k_T = 2.8$, the observed values of $k_D/k_T = 1.2-1.5$ require that $k_{-1}/k_2 = 10-3$; that is, that $k_{-1}/k_{\text{exptl}} = 11-4$.

Internal return cannot be important for exchanges with CsCHA or for benzylic positions with LiCHA because the experimental primary isotope effects are too high. If the internal return explanation were correct for exchange of aryl deuterium with LiCHA, the derived values for $k_1(\alpha\text{-cumene})/k_1(\text{benzene})$ would be smaller than the experimental ratio by the 11-4 factors

given above; that is, the LiCHA relative rate, 0.13-0.35, would now be far lower than the CsCHA relative rate, 1.3. Although this is a logically permissible result, it seems more probable that the relative rates for the forward steps with the two catalysts are more nearly comparable and, consequently, that $k_2 \gg k_{-1}$ and $k_{\text{exptl}} = k_1$ for all of our systems. The latter alternative requires that the observed low isotope effect for exchange of aryl hydrogens with LiCHA be a real effect for k_1 and leads to the interpretation advanced earlier.^{4a}

Acidity in Nonaqueous Solvents. I. Picolinium Ions in Methanol¹

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Free energies, enthalpies, and entropies of ionization of a series of α -substituted 4-picolinium ions in pure anhydrous methanol have been determined by the differential potentiometric method. Three points of general interest emerge from the data. First, the effects of alkyl group substituents are remarkably small; second, pK values are very nearly equal to those found for aqueous solutions; and third, there is no correlation between free energies and enthalpies of ionization for these acids.

Introduction

Very little data exists on the free energies, enthalpies, and entropies of ionization of Brønsted acids in nonaqueous solutions. Such data must be acquired if any interpretation of solvent effects is to be complete. In a continuing study to shed light on the nature of solvent effects on reactivity, we have measured the dissociation constants of a series of α -substituted 4-picolinium ions in pure methanol solution at 0.1 and 25.0°. From these data, free energies, enthalpies, and entropies of ionization are evaluated.

Results

The data obtained are summarized in Table I. Reported pK values are not corrected for ionic strength, and, therefore, standard states are implicitly referred to ca. 10^{-3} M solution in methanol. Under the assumptions of Debye-Hückel theory, no corrections are necessary for reference to a standard state of infinitely dilute solution.

The limitations of the differential potentiometric method for determining pK values have been discussed by Grunwald.² The pK values for the picolines in methanol fall within the optimum range for this method, and the values reported in Table I should have an ac-

Table I. Thermodynamic Data for the Reaction $4\text{-XCH}_2\text{C}_6\text{H}_4\text{NH}^+ \rightleftharpoons 4\text{-XCH}_2\text{C}_6\text{H}_4\text{N} + \text{H}^+$

	pK		ΔG°	ΔH	ΔS°
	0.1°	25.0°	(25°), kcal./ mole	(25°), kcal./ mole	(25°), e.u.
H	6.541	6.090	8.300	6.7	-5.2
CH ₃	6.550	6.090	8.300	6.9	-4.8
C ₂ H ₅	6.534	6.073	8.280	6.9	-4.7
C ₆ H ₅	6.228	5.794	7.900	6.5	-4.7
OH	6.333	5.846	7.970	7.3	-2.3
CN	5.002	4.565	6.220	6.5	+1.0
NH ₃ ⁺	4.13	3.77	5.13	5.5	+1.2

curacy approaching the precision of 0.01 unit reported in the Experimental Section of this paper.

The accuracy of the enthalpies and entropies reported is more difficult to evaluate. According to the analysis of Petersen, Margraf, and Ross,³ however, the accuracy is certainly not worse than $\pm 10\%$, and is quite probably better than $\pm 5\%$. The enthalpies are thus constant within experimental error for all of the compounds studied except those with the hydroxyl and ammonium substituents.

Discussion

Three points of general interest emerge from the data of Table I.

First, the effects of alkyl group substituents are remarkably small. Several years ago, Brown⁴ reported pK values for 4-methyl-, 4-ethyl-, and 4-isopropylpyridine in aqueous solution. The pK value was 6.02 for all three compounds. These two pieces of information may now be added to the growing body of data⁵ in support of our earlier suggestion that alkyl groups have negligible inductive effects.⁶

(3) R. C. Petersen, J. H. Margraf, and S. D. Ross, *ibid.*, **83**, 3819 (1961).

(4) H. C. Brown and X. R. Mihm, *ibid.*, **77**, 1723 (1955).

(5) H. D. Holtz and L. M. Stock, *ibid.*, **86**, 5188 (1964), and other references cited therein.

(1) Presented at the 149th National Meeting of the American Chemical Society, Detroit, Mich., April 1965.

(2) E. Grunwald, *J. Am. Chem. Soc.*, **73**, 4934 (1951).

Second, it is interesting that the pK values in methanol are nearly identical with those reported by Brown for aqueous solution. Bell⁷ has tabulated pK values in water and methanol for a number of primary ammonium ions and notes that these acids show an average increase of 1.2 pK units in going from water to methanol. Grunwald⁸ has reported, however, that the tertiary amine, *N,N*-dimethylaniline, has essentially the same pK in water and methanol even though the value in 80% methanol is 1.2 units lower. Perhaps the equality of pK values in water and pure methanol is characteristic of tertiary amines. It is also interesting to note that the entropies of ionization in methanol for the picolinium ions are very nearly equal to the entropies of ionization in water for various ammonium ions reported by Bell.⁷

Third, there is no correlation between free energies and enthalpies of ionization in this series. The situation is quite different from that found for the ionization of anilinium ions in aqueous solution where the enthalpies and free energies show parallel changes with substituent.⁹ Our present data are similar to those found by Hepler¹⁰ for the ionization of phenols in aqueous solution, and are consistent with his interpretation of the origin of enthalpy and entropy effects in such reactions.

A plot of pK vs. σ_1 for the compounds in Table I gives a ρ value of -2.8 . The hydroxyl- and ammonium-substituted compounds fall rather badly off the plot. We have commented earlier on the behavior of charged substituents,¹¹ and Jaffé¹² has remarked on the variable substituent constant for the hydroxyl group.

We would also like to point out that our present studies have verified the reversibility of the glass electrode in pure methanol solution, and that we have found that junction potentials are essentially constant throughout our measurements. It is thus possible to calibrate a pH meter using a buffer composed of any of the compounds reported in Table I, and thereby to carry out direct measurements of ionization constants for other acids in pure methanol solution.

Experimental Section

Materials. Reagent grade methanol (Baker Analyzed reagent) was further purified by drying with Linde Molecular Sieves 3A, followed by passage through a column of strongly basic, then strongly acidic ion-exchange resins. The purified methanol was determined to have less than $3 \times 10^{-6} M$ acidic or basic impurities by potentiometric titration, and less than $10^{-3} M$ water by Karl-Fischer titration.

A solution of $0.025 M$ anhydrous perchloric acid in methanol was prepared by dissolving sodium perchlorate in purified methanol and slowly passing the resulting solution through a column of strongly acid

ion-exchange resin. The conversion to perchloric acid was shown to be quantitative, and the solutions contained less than 20 p.p.m. of water.

The ion-exchange resins were commercially obtained. The resins were first washed with distilled water and sucked dry on a Büchner funnel, then dried under vacuum at approximately 100° . The dried resins were washed thoroughly with dry methanol before use. All operations during purification and preparation of acid were carried out under a dry nitrogen atmosphere.

Sodium methoxide solutions were prepared by adding freshly cut sodium metal to purified methanol and were standardized by titration of potassium hydrogen phthalate.

Perchloric acid solutions were standardized by titration with the methanolic sodium methoxide.

With the exception of the cyano-substituted compound, the α -substituted picolines were obtained commercially and purified immediately before use by distillation or crystallization. The melting points and boiling points checked closely with literature values.

The synthesis of 4-pyridineacetonitrile was accomplished by the method of Schulze.¹³ The identity of the product was confirmed by infrared spectrum and melting point.

Determination of pK Values. Grunwald's differential method² was employed to determine the absolute ionization constants of all except the ammonium-substituted compound. The pK for this latter compound was determined from half-neutralization potentials together with the standardization potentials determined from the other members of the series.

A Beckman research pH meter Model 1019 was used in conjunction with either a Corning or Beckman high pH glass electrode and a Leeds and Northrup calomel electrode filled with methanolic KCl. The reversibility of the electrode system was demonstrated by the fact that in titration of bases, the quantity $\log [P/(100 - P)]$, where P is the percentage of base titrated vs. potential of the cell, gave linear plots with precisely theoretical slopes at both 0.1 and 25.0° .

Titration to determine the pK values utilized $10^{-3} M$ solutions of the picolines titrated with the $0.025 M$ perchloric acid. Titrations were carried out in a jacketed, stoppered vessel under an atmosphere of purified nitrogen. Water from a constant temperature bath was circulated through the jacket to maintain temperatures within 0.05° of the reported temperature.

Water content of the solutions after titration seldom exceeds $10^{-3} M$, and was never greater than $3 \times 10^{-3} M$.

At least two runs, often more, were made for each compound at each temperature. A precision of better than $0.01 pK$ unit was achieved for all compounds except that with the ammonium substituent. This latter compound showed maximum deviations of as much as $0.03 pK$ unit.

Determinations of the pK of 4-pyridylcarbinol were carried out at 0.1 , 9.9 , and 25.0° . Enthalpies calculated from any two temperatures were identical, within experimental error.

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