

The formula $\{H_3N-H-OH\}^0$ has been written to suggest that an un-ionized molecular entity similar to an ion pair exists in solution; this entity arises from the presumed hydrogen bonding between the nitrogen and oxygen atoms from the ammonia and water molecules. While this series of equilibria is reminiscent of a similar set for carbon dioxide and water, reference 4, there is no evidence and little reason to believe that with ammonia any of these steps is slow. Indeed, reaction 1 is undoubtedly a fast one to which only a Langevin time lag could be attributed; the same is true of reaction 2. Nevertheless, and presumably this may be attributed to the enhanced stability conferred by the hydrogen bond, there is a finite concentration of dissolved but unhydrated ammonia in the solution together with an amount of undissociated ammonium hydroxide, and, finally, an equilibrium concentration of ions, as is represented by equations 1 and 2. These facts were clearly established by Moore and Winmill,⁷ from whose work on ammonia and alkyl substituted amine solutions the hydrogen bond concept arose.

The ratio of the concentrations of ammonia and undissociated ammonium hydroxide, equation 1, is independent of concentration if the concentrations be small, and may be termed B .

$$B = [NH_3]/[NH_4OH] \quad (3)$$

The value of B may be determined by measuring at at least three temperatures the ionization and the partition of ammonia between water and an immiscible solvent such as toluene. Moore and Winmill⁷ thus determined how much of the undissociated ammonia in solution was present as dissolved ammonia and how much as ammonium hydroxide. Sidgwick⁸ states that at 25° the value of B is 0.885. Using the relations

$$K(0)(\text{true}) = [NH_4^+][OH^-]/[NH_4OH] \quad (4)$$

$$K(0)(\text{apparent}) = \frac{[NH_4^+][OH^-]}{[NH_4OH] + [NH_3]} = \frac{[NH_4^+][OH^-]}{[NH_4OH](1 + B)} \quad (5)$$

$$= K(0)(\text{true})/(1 + B) \quad (6)$$

and the value of $K(0)$ (apparent) at 25° from reference 6, we find that the method of Moore and Winmill gives $K(0)$ (true) = 3.34×10^{-5} , while Moore and Winmill themselves report $K(0)$ (true) at 20° from their measurements to be $5.2 \pm 1.3 \times 10^{-5}$. Bates and Pinching⁶ give the values of $K(0)$ (apparent) at 20° and at 25° as 1.710 and 1.774×10^{-5} . The difference between the computed quantity 3.34×10^{-5} and Moore and Winmill's value of $5.2 \pm 1.3 \times 10^{-5}$ is thus too large to be accounted for by the limits of error specified, and is in the wrong direction as a function of temperature, unless other unaccounted potent factors affecting the true ionization constant in a different manner than the apparent constant were present.

We may employ the method described in reference 4 to obtain $K(0)$ (true) from the high field conductance data. From a series of plots of the coefficients A_2 and A_3 against b (see eq. 21, ref. 4),

(7) T. Moore and T. Winmill, *J. Chem. Soc.*, **101**, 1635 (1912); also, **91**, 1373 (1907).

(8) N. V. Sidgwick, "The Chemical Elements and their Compounds," Oxford University Press, New York, N. Y., 1950, p. 660.

the reduced slope Q and the function ρ_0 (eq. 17, ref. 4) are obtained. From the data presented in Fig. 1 ρ_0 has the average value 0.1639, and, with the value for $K(0)$ (apparent) = 1.77×10^{-5} , $K(0)$ (true) is found to be $6.3 \pm 0.5 \times 10^{-5}$. The precision of measurement is at least ten times poorer than for carbon dioxide, as may be seen from the discussion of errors in ref. 4. This value for $K(0)$ (true) must be compared with those in the paragraph above; it is larger than any of the values of $K(0)$ (true) either quoted or computed from other data. A larger $K(0)$ bespeaks a smaller Wien effect. It is indeed possible that electrolytic impurities might cause the measurements here reported to be low, although every precaution has been taken to exclude contaminants. It seems more probable that the rapidity of reactions 1 and 2 would give a high field conductance increase larger than would truly represent the equilibrium concentrations of the several ionic species at low fields and thus a $K(0)$ (true) which is too small. This possibility hinges upon the relative speeds of reactions 1 and 2, direct information on which is not available. Experiments at a variety of pulse lengths might help elucidate this point. It is planned to extend these measurements to a range of temperatures for which values of $K(0)$ (apparent) are available in order to test the internal consistency of data derived from the high field conductance measurements.

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Thiohydantoins of Amino Acids¹

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Recent reports³ have shown that some thiosemicarbazones have an inhibitory effect on the growth of vaccinia viruses. Thompson and Wilkin have shown that the phenylalanine antagonist, β -2-thienylalanine, prevented the multiplication of vaccinia virus in chick embryonic tissues.⁴ 5-Substituted thiohydantoins possess certain structural similarities to both of these types of virus inhibitors, having an α -aminocarbonyl and an N-substituted thiourea portion in the same molecule. In pursuing a program of virus chemotherapy⁵ we have synthesized a number of thiohydantoins in order to determine whether these compounds had antiviral activity. Tests conducted by Dr. R. L. Thompson at the Indiana University Medical

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(2) Abstracted from the thesis of Wesley L. Archer, to be submitted to Indiana University in partial fulfillment for the Degree of Doctor of Philosophy.

(3) Cf. R. L. Thompson, S. A. Minton, Jr., J. E. Officer and G. A. Hitchings, *J. Immunology*, **70**, 229 (1953); D. Hamre, J. Bernstein and R. Donovan, *Proc. Soc. Exp. Biol. Med.*, **73**, 275 (1950).

(4) R. L. Thompson and M. L. Wilkin, *ibid.*, **68**, 434 (1948).

(5) Cf. E. Campaigne, *et al.*, *THIS JOURNAL*, **75**, 988 (1953).

