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Keyphrases

Molluscicides—structure-activity relationship
 Snails—species of *Biomphalaria*, *Bulinus*
 Schistosoma—infesting organism
P-Phenylenediamine—color reagent
 Colorimetric analysis—spectrophotometry

Salt Effects in the Spectrophotometric Examination of a Complex of *N'*-Methylnicotinamide Cation

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The extent of complex formation, as determined spectrally, between the *N'*-methylnicotinamide cation and the electron donor, 8-chlorotheophyllinate anion, was found to vary greatly when different salts were used to adjust ionic strength. The degree of complex formation when sodium perchlorate was used was found to be much smaller than when sodium acetate was used. An examination of these salt effects indicated that the observed behavior was explicable on the basis that ion pair formation occurred between the pyridinium cation and perchlorate anion.

THE NECESSITY of maintaining a constant ionic strength in studying the behavior of ionic species in solution is widely recognized. It is not, perhaps, so widely appreciated that a so-called inert salt introduced for this purpose can exert an effect quite different from that anticipated solely on the basis of ionic strength variation. Recent studies in this laboratory involving the interaction of *N'*-methylnicotinamide cation (NMN) with 8-chlorotheophyllinate anion (CT) have demonstrated quite dramatically that neutral salts can influence observed behavior independently of their contribution to the maintenance of ionic strength. As will be seen, the extent of interaction of NMN with CT, as detected spectrally, was found to vary considerably depending on whether sodium acetate or sodium perchlorate was used to maintain a constant ionic strength. The observed behavior can be reasonably explained on the basis that extensive ion pair formation occurred between perchlorate anion and NMN.

Ion pairing of quaternary nitrogen compounds is well documented. For example, Azzari and Kraus (1) reported the formation of ion pairs between the tetrabutyl ammonium cation and such anions as iodide, bromide, and nitrate in water at 25°. Fuoss (2) showed that ionic associations could be treated by mass law considerations, and he and his co-workers (3) reported the ion pair formation constant for the tetraethyl ammonium-picrate ion pair to be 0.8 L. mole⁻¹ in water. Kosower (4), in a description of charge transfer interactions, asserted that the major contributing form to the ground state complex between 1-methyl-pyridinium cation and iodide anion was the ion pair.

EXPERIMENTAL

Materials—The chloride salt of NMN was prepared by the method of Huff and Perlzweig (7). Nicotinamide was reacted with an excess of methyl iodide in alcohol to yield NMN as the iodide salt, m.p. 204° uncorrected [reported, 205.5–206.5° (7)]. The iodide salt was converted to the chloride salt by reacting the NMN iodide with an aqueous slurry of silver chloride. The resulting aqueous solution of NMN chloride was filtered and concentrated to yield white crystals of NMN chloride. The chloride salt was recrystallized from hot alcohol, and gave a m.p. of 241° uncorrected [reported, 237–238° (7)].

8-Chlorotheophylline was purchased from the Aldrich Chemical Co. and used without further puri-

Received December 18, 1967, from the Department of Pharmaceutics, School of Pharmacy, State University of New York at Buffalo, Buffalo, NY 14214.

Accepted for publication March 18, 1968.

The Computing Center at the State of New York at Buffalo is partially supported by NIH grant FR-00126 and NSF grant GP-7318.

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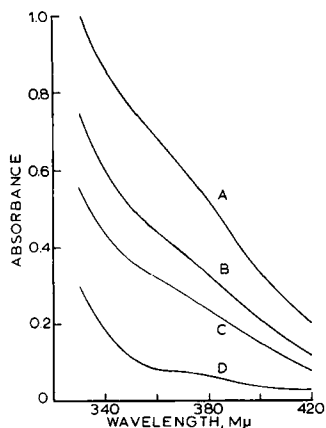


Fig. 1—Plot illustrating the influence of added salts on the spectral characteristics of solutions at a pH of 8.0 containing 0.1 M CT and 0.005 M NMN. Curve A represents the spectrum of the mixture in the absence of added salt. Curve B is the spectrum of the mixture in the presence of 0.9 M sodium acetate. Curve C is the spectrum of the mixture in the presence of 0.9 M sodium perchlorate. Curve D is the spectrum of 0.1 M CT alone.

fication. Other chemicals were reagent grade and were obtained from Fisher Scientific Co.

Methods—The influence of sodium acetate and sodium perchlorate on the extent of complex formation between NMN and CT was evaluated by using either neutral salt, or a combination of the two, to adjust the ionic strength of solutions containing the interactants to a value of 1.0. These solutions were prepared by dissolving appropriate amounts of 8-chlorotheophylline, pK_a 5.5 (6), and neutral salts in about 35 ml. of water, using enough sodium hydroxide to adjust the pH to a value of 8.0. These were placed in 50-ml. volumetric flasks and aliquots of a stock solution of NMN as the chloride salt were added to the flasks which were then made to volume to gain the desired concentrations. The ionic strength of the resultant solutions was taken as the sum of the concentrations of sodium acetate, sodium perchlorate, and sodium 8-chlorotheophyllinate. The concentration of NMN as the chloride was negligible compared to the concentrations of other ionic species. The resultant solutions were brought to 40° in a water bath. The extent of complex formation was measured spectrally at 40° using a

Beckman, model DB, spectrophotometer equipped with a thermostated cell compartment. The large increase in absorbance at 360 $m\mu$ of solutions of CT on the addition of NMN, a species with no visible spectrum, was taken as a measure of the extent of interaction between NMN and CT. Absorbance values were corrected for the small contribution due to CT. Representative spectra are shown in Fig. 1.

RESULTS AND DISCUSSION

Spectral evidence for a postulated complex of NMN and CT is given in Fig. 1. It can be seen that the extent of complex formation was decreased by increasing the ionic strength from a value of 0.1 to a value of 1.0 with the addition of sodium acetate or sodium perchlorate. Though a decrease with increasing ionic strength was expected on a theoretical basis, the marked difference in the extent of interaction observed when the different salts were employed was unexpected. The results of a number of experiments are given in Table I, all of which illustrate the same behavior.

That the extent of interaction between NMN and CT is influenced differently by the two salts can be explained on the basis of a rather simple model which assumes that NMN forms ion pairs with perchlorate anion. As a first approximation, based on the much larger influence of sodium perchlorate, it was assumed that acetate anion was not involved in ion pair formation. The model neglects possible differences in activity coefficients, hydration, and ion size parameters of unassociated NMN and CT in the presence of the different salts, and attributes the larger influence of perchlorate to a competitive effect. The model is summarized by



where $\text{NMN:CT} = 1:1$ complex formed between NMN and CT, K = association constant for the formation of NMN:CT , $\text{NMN:ClO}_4 = 1:1$ ion pair formed between NMN and perchlorate, G = ion pair formation constant for NMN:ClO_4 .

The absorbance, A , due to the complex, NMN:CT can be written as

$$A = \epsilon(\text{NMN:CT}) = \epsilon K(\text{NMN})(\text{CT}) \quad (\text{Eq. 3})$$

where ϵ = molar absorptivity of the complex.

The absorbance can be stated in terms of total NMN, $(\text{NMN})_t$, as

TABLE I—ABSORBANCE VALUES AT 360 $m\mu$ DUE TO COMPLEX FORMED BETWEEN NMN AND CT IN VARIOUS SOLUTIONS ADJUSTED TO AN IONIC STRENGTH OF 1.0

Solution Composition				Observed	Absorbance ^a Calculated	Difference
(CT)	(NMN) _t	(ClO ₄)	(Ac)			
0.10	0.0050	0.00	0.90	0.359	0.374	-0.016
0.10	0.0050	0.50	0.40	0.289	0.280	+0.009
0.10	0.0050	0.90	0.00	0.238	0.233	+0.005
0.15	0.0050	0.00	0.85	0.515	0.505	+0.010
0.15	0.0050	0.30	0.55	0.415	0.426	-0.011
0.15	0.0050	0.85	0.00	0.341	0.332	+0.009
0.14	0.0025	0.00	0.86	0.241	0.242	-0.001
0.14	0.0025	0.43	0.43	0.193	0.190	+0.003
0.14	0.0025	0.86	0.00	0.142	0.157	-0.015
0.21	0.0025	0.00	0.79	0.303	0.319	-0.016
0.21	0.0025	0.40	0.40	0.253	0.262	-0.009
0.21	0.0025	0.79	0.00	0.214	0.222	-0.008

^a Corrected for the absorbance contribution due to CT.

$$A = \frac{\epsilon K(\text{NMN})_i(\text{CT})}{1.0 + K(\text{CT}) + G(\text{ClO}_4)} \quad (\text{Eq. 4})$$

Equation 4 can be rewritten as

$$\frac{1}{A} = \frac{1}{\epsilon K(\text{NMN})_i(\text{CT})} + \frac{1}{\epsilon(\text{NMN})_i} + \frac{G(\text{ClO}_4)}{\epsilon K(\text{NMN})_i(\text{CT})} \quad (\text{Eq. 5})$$

Equation 5 is essentially the same as that derived by Merrifield and Phillips (8) for a case in which both a solvent and an electron donor interacted with an electron acceptor but the solvent-acceptor complex did not contribute to the absorbance of the system. When $(\text{NMN})_i$ and (CT) are held constant, plots of $1/A$ against (ClO_4) should be straight lines. Such plots are given in Fig. 2. The value of $G/\epsilon K$ can be calculated by multiplying the slopes of the lines by the appropriate $(\text{NMN})_i(\text{CT})$ product. The intercepts of the lines in Fig. 2, multiplied by the appropriate $(\text{NMN})_i$, when plotted against $1/(\text{CT})$ should give a straight line with intercept $1/\epsilon$ and slope $1/\epsilon K$. This plot, which corresponds to a Benesi-Hildebrand (5) plot, is given in Fig. 3. The values for ϵ , K , and G calculated in this manner are 359 L. mole⁻¹ cm.⁻¹, 2.69 L. mole⁻¹, 0.92 L. mole⁻¹.

Treatments such as the one used here depend heavily upon a knowledge of the value $1/\epsilon$. Since this value is small, a small absolute error in the determination of $1/\epsilon$ can lead to large errors in ϵ , and in the association constants. An attempt was made to evaluate the association constants independently of a determination of ϵ . This can be done by considering two experimental points at one time. It can be shown that

$$\frac{A_1}{A_2} = \frac{(\text{NMN}1)_i(\text{CT}1)[1.0 + K(\text{CT}2) + G(\text{ClO}_42)]}{(\text{NMN}2)_i(\text{CT}2)[1.0 + K(\text{CT}1) + G(\text{ClO}_41)]} \quad (\text{Eq. 6})$$

where 1 and 2 refer to two different experimental systems. The values for K and G were generated, according to Eq. 6, on an IBM 7044 computer. The values for K and G were found to be 2.67 L. mole⁻¹, and 0.86 L. mole⁻¹. The corresponding value of ϵ was calculated to be 355 L. mole⁻¹ cm.⁻¹.

The calculated absorbance values given in Table I are based upon the computer generated values of K , G , and ϵ . The random distribution of the difference cited in Table I indicates that the observed differences in the effects of sodium acetate and sodium perchlorate can be reasonably explained on the basis of ion pair formation between NMN and perchlorate. The treatment presented here does not preclude the possibility of ion pairing between acetate anion and NMN but does suggest that such ion pairing is small compared with the association between NMN and perchlorate.

It is interesting to note that the definition of a particular ionic behavior may vary quantitatively depending upon the "inert" salt used for the purpose of adjusting ionic strength. In the present investigation, the apparent association constant for the NMN:CT complex in systems in which sodium perchlorate was the salt used was found to be 1.52 L. mole⁻¹ by the Benesi-Hildebrand (5) treatment. When sodium acetate was the salt used the apparent association constant was found to be 2.59 L. mole⁻¹. This observation makes it clear that the effects of neutral salts cannot always be judged solely on the

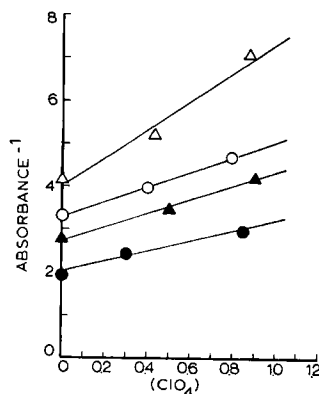


Fig. 2—Plot, based on Eq. 5, illustrating the dependence of complex formation on perchlorate concentration. Solutions contained: 0.14 M CT and 0.0025 M NMN (Δ), 0.21 M CT and 0.0025 M NMN (\circ), 0.10 M CT and 0.005 M NMN (\blacktriangle), 0.15 M CT and 0.005 M NMN (\bullet).

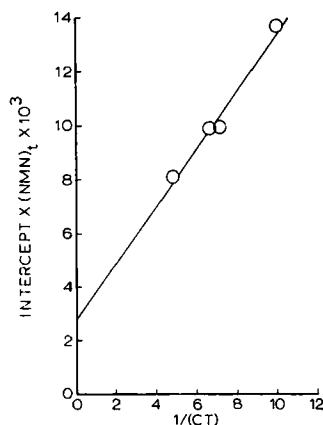


Fig. 3—Plot used to determine K and ϵ . See text.

basis of their contribution to the maintenance of ionic strength.

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Keyphrases

N-Methylnicotinamide—8-chlorotheophyllinate complex
Salt effects—complex formation
Ionic strength—salt maintenance
UV spectrophotometry—analysis