THE UNIVERSITY OF OKLAHOMA

GRADUATE COLLEGE

A SPECTROSCOPIC INVESTIGATION OF THE SELF ASSOCIATION AND HYDRATION OF LACTAMS IN CARBON TETRACHLORIDE

A DISSERTATION

SUBMITTED TO THE GRADUATE FACULTY

in partial fulfillment of the requirements for the

degree of

DOCTOR OF PHILOSOPHY

BY

JOHN DAVID WORLEY

Norman, Oklahoma

1964

A SPECTROSCOPIC INVESTIGATION OF THE SELF ASSOCIATION AND HYDRATION OF LACTAMS IN CARBON TETRACHLORIDE

APPROVED BY

Harold E. Affsprung

J. C. Colbert

J. D. Christian

R. J. Gleason

J. Rud Nielson

DISSERTATION COMMITTEE

ACKNOWLEDGEMENTS

The author wishes to acknowledge and express his sincere appreciation to the following people for their assistance in making this investigation.

Drs. Harold E. Affsprung and Sherril D. Christian for suggesting this problem and for their encouragement, patience and very helpful suggestions throughout the course of this investigation.

James Johnson for the Karl Fischer titrations and for the use of his constant activity solutions.

Ronald Lynch and Gary Wood for the use of their equilibration data in the evaluation of the absorptivity of the hydrated 2-Pyrrolidone species.

Delbert Mueller for the use of his least squares computer program in fitting the hydration data.

Dr. A. J. Weinheimer and Bob Middlebrook for providing the 2-Pyrrolidone, N-Methyl-2-pyrrolidone and apparatus used for the purification of these substances.

The National Institutes of Health for providing the necessary funds for carrying out this investigation.

Finally the author would like to express his gratitude to his mother and his wife whose love and encouragement made all this possible. He would also like to acknowledge the many people in the graduate school with whom he was fortunate enough to be friends. There is no substitute for them.

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A SPECTROSCOPIC INVESTIGATION OF THE SELF ASSOCIATION *AND HYDRATION OF LACTAMS IN CARBON TETRACHLORIDE*

CHAPTER I

INTRODUCTION

"It is hardly an exaggeration to say that in the chemistry of living systems the hydrogen bond is as important as the carbon-carbon bond."[1] This statement from the introductory chapter of Pimentel and McClellan's book "The Hydrogen Bond" illustrates the importance which has been attached to hydrogen bonding in systems of biological interest. Pauling, Corey and Branson in 1951 proposed two helical structures for proteins which were stabilized by means of intramolecular hydrogen bonds between peptide groups.[2] Following this in 1953 Watson and Crick proposed the helical duplex model for the genetically important deoxyribonucleic (DNA) acid.[3] The two strands of the duplex were held together by intermolecular hydrogen bonds between purine and pyrimidine bases. Both of these models have been confirmed from X-ray diffraction studies on the solid crystalline substance.

Since the nucleic acids and proteins normally function within the cell, which is surrounded by an aqueous medium, it is important to know more about the behavior of these macromolecules in aqueous solutions.

Inasmuch as water itself is a very good hydrogen bonding substance there is every reason to believe it may play an important part in the molecular configuration assumed by the biopolymer. This has been well illustrated by X-ray studies of solid deoxyribonucleic acid films under varying conditions of humidity.[4,5] Four different diffraction patterns were observed in differing relative humidity regions indicating that the presence of variable amounts of water can cause DNA to assume one of four structural forms.

Falk, Hartman and Lord have made a comprehensive study of this phenomenon using spectroscopic and gravimetric techniques.[6,7,8] From their data they assert a sequence of hydration sites within the DNA polymer where the structure of the adsorbed water is similar to that of liquid water. They maintain that at low relative humidities the structure of DNA assumes a disordered state similar to that found when DNA is denatured by heating. On this basis they conclude that an ordered structure, where the purine and pyrimidine bases are stacked on top of each other, is the stable configuration in solution since, this kind of structure offers the maximum number of water-water and water-base interactions.

Peter Geiduschek and co-workers have studied the denaturation of DNA in acidic media. The results of these studies indicate that the DNA configuration remains intact even though a large number of basic sites in the purine and pyrimidine bases are protonated. This constitutes further evidence that the helix is stabilized by means other than interpeptide hydrogen bonds.[9,10,11]

Klotz has reviewed the hydration of proteins and used the "iceberg"

concept of Frank and Evans to explain the behavior of protein solutions.[12] This theory states that water assumes an icelike structure in the vicinity of the hydration sites in the protein. This, of course, is in direct contradiction to the previous assertion that the water of hydration has a liquid like structure.

Considerable light can be thrown on the effects of hydrogen bonding within macromolecules by the study of simple molecules of low molecular weight containing within their structure, functional groups representative of particular segments of biopolymers. Synthetic polypeptides and amides have been very useful for carrying out this kind of investigation since they contain the very important peptide group.

Pimentel and McClellan have defined a hydrogen bond as follows: "A hydrogen bond exists between a functional group A-H and an atom or group of atoms B in the same or different molecule when (a) there is evidence of bond formation and (b) when there is evidence that this new bond linking A-H and B specifically involved the hydrogen atom already bonded to A."[13] There are many experimental techniques which can provide information about (a) but only a few can give information concerning (b). Some experimental methods commonly used to detect association, but which cannot satisfy part (b) of the above definition, are studies of dielectric, cryoscopic, distribution, conductometric, solubility and acoustic phenomena.

Spectroscopic methods readily satisfy both parts of Pimentel and McClellan's definition. The near infrared region of the spectrum from 1 μ to 4 μ has been found to be one of the most useful regions for

making spectral studies of hydrogen bonding. It is here that absorption due to the A-H stretching vibration occurs. The effect of hydrogen bond formation on this particular mode of vibration is both pronounced and easily studied since, the hydrogen bonded, A-H...B, stretching absorption shift occurs in the same near infrared region but significantly removed from the absorption of the unbonded A-H mode. The effect of hydrogen bonding on the A-H stretching mode may be summarized by the following points:[14]

1. The stretching mode and its harmonics are shifted to lower frequencies.

2. The stretching mode is broadened and so are its harmonics.

3. The absorptivity increases when a hydrogen bond is formed.

The thermodynamic parameters have been evaluated by measuring changes in the absorbance bands of the free and associated species with varying concentration and temperature. This technique has been reviewed by Mecke[15] for alcohols and phenols.

Very few quantitative studies have actually been carried out on amide systems. Tsuboi studied the spectra of N-methylacetamide (NMA) and δ-valerolactam in carbon tetrachloride as functions of concentration and temperature.[16] He found a sharp band at 2.88 μ for the NMA in carbon tetrachloride which he assigned to the free N-H stretching motion. Two broad bands appearing at 2.97~3.0 μ and 3.22 μ were assigned to vibrational modes of the associated species. The δ-valerolactam, a cyclic amide, also exhibited three absorption bands. The band at 2.92 μ was assigned to the free N-H stretching mode and the bands at 3.11 μ and 3.24 μ were assigned to the associated molecules. Tsuboi

concludes that the 3.0 μ to 2.97 μ shift of the associated band of NMA with lowering concentration and increasing temperature is an indication of association by linear polymerization. No frequency shift with change in concentration was observed for the lactam.

Mizushima has made a detailed study of NMA in carbon tetrachloride solution using Raman, infrared and ultraviolet spectra as well as dipole moment techniques.[17] His conclusions are in accord with those of Tsuboi. He asserts that NMA is present in the "trans" configuration and that association occurs through formation of a linear polymer while the lactams form cyclic dimers.

Klemperer and co-workers studied the spectra of N-ethylacetamide (NEA), N-n-butylacetamide (NBA) and γ-butyrolactam as functions of concentration in carbon tetrachloride.[18] They also studied the spectra of these substances at fixed concentrations in other solvents. The concentration range studied using the lactam was 10^{-2} to $4(10)^{-1}$ molar. For the nitrogen substituted amides the concentration range was $7.8(10)^{-4}$ to $1.6(10)^{-1}$ molar. The free N-H bands were found at 2.880 μ (NEA), 2.898 μ (NBA) and 2.897 μ (lactam). Vibrational bands of the associated species were found at 2.970~3.022 μ and 3.242~3.218 μ for NEA and 2.985~3.012 μ and 3.242 μ for NBA. The γ-butyrolactam showed bands at 3.106 μ and 3.211 μ which were identified as N-H vibrational modes and were assigned to the associated molecules. Klemperer's data support Mizushima's theory that secondary amides form linear polymers while lactams form cyclic dimers.

R. M. Badger and Hector Rubalcava have made a quantitative spectral study of the self association of propionamide in carbon tetrachloride.[19]

They found seven absorption bands in the spectrum and assigned one each to the symmetric and asymmetric stretching vibrations of the amino hydrogens in the monomer and the remaining five vibrational modes to a dimer. They used the dimer band at 3.13 μ to determine the dimerization constant (K_2) and heat of dimerization (ΔH_{Dim}). K_2 was found to be 45 molar^{-1} units at 25°C and ΔH_{Dim} had the value -7.85 kcal/mole.

Klotz and Franzen have evaluated equilibrium constants, free energies, enthalpies and entropies of association from spectroscopic studies for N-methylacetamide in a variety of solvents.[20,21] For carbon tetrachloride they found an association constant (K_{Ass}) at 25° of 4.7 molar^{-1} units and a heat of association (ΔH_{Ass}) of -4.2 kcal/mole. Using dioxane as a solvent values of K_{Ass} = .52 and ΔH_{Ass} = -0.8 kcal/mole were found. In water a value of K_{Ass}=.005 and ΔH_{Ass}=0.0 were established. These calculations have led Klotz and Franzen to the conclusion that the intrinsic stability of interpeptide hydrogen bonds in aqueous solution is very small and that the stabilizing factor for proteins in solution must be something other than interpeptide linkages.

Tsuboi in 1951 investigated the system valerolactam-carbon tetrachloride and determined the dimerization constant and heat of dimerization[22] of the lactam. The geometry of this molecule makes it very favorable for dimer formation since, the C=O and N-H functional groups are held in the "cis" position by the ring structure. Tsuboi found the absorbance of both bands assigned to the associated species at 3.11 μ and 3.24 μ to vary with temperature and concentration. However, the ratio of the absorbance of the two bands remained constant over the entire concentration and temperature ranges studied. This is taken as

strong evidence for the existence of only one major associated species in carbon tetrachloride solution.

Assuming a cyclic dimer Tsuboi evaluated the association constant from absorbance measurements of the 3.11 μ dimer band.[22] He found a K_2 of 432 molal^{-1}units. By determining K_2 at several temperatures and using the van't Hoff relation he evaluated the heat of dimerization and found ΔH_{Dim} to be -10.3 kcal/mole or -5.0 kcal/H-bond.

R. C. Lord and T. J. Porro made a near infrared study of the dimerization of ϵ-caprolactam in CCl_4.[23] They found the free N-H at 2.915 μ and three bands for the associated molecule at 3.027 μ , 3.108 μ and 3.238 μ . They also found that the absorbance ratios of the 3.238 μ band and the 3.027 μ band to the 3.108 μ band were constants with respect to changes in temperature and concentration. Using both monomer and dimer absorption bands they have evaluated association constants and heats of dimerization. At 25° they found the association constant to have a value of 168 molal^{-1} units. The heat of dimerization was -5.46 kcal/mole or -2.73 kcal/hydrogen bond.

Other investigations of amide association in solution include vapor pressure studies of secondary amides in benzene by Davies and Thomas.[24] They evaluated association constants and enthalpies and found that the secondary amides apparently associate to form linear polymers in benzene as well as CCl_4. The halogenated primary and secondary amides, trichloroacetamide and N-methyltrichloroacetamide appear to form trimers in benezene.

Christian, Affsprung and Taylor have studied the effects of small amounts of water dissolved in organic solvents on the monomer-dimer

equilibrium of carboxylic acids.[25] Assuming the presence of free water, acid monomer, dimer and monomer monohydrate in the solvent benzene they derived equations from which an anhydrous association constant could be evaluated. Comparison of this K_2 with apparent association constants calculated from distribution studies revealed that the anhydrous K_2 was two or three times larger. This illustrates the remarkable effect small amounts of water can have on the self association of hydrogen bonding solutes in an organic solvent. These facts suggested an excellent method for studying the hydration of molecules of biological interest in non-hydrogen bonding solvents. Such studies, however, require a knowledge of the solute properties of water in these solvents.

Greinacher, Lüttke, and Mecke have studied the infrared spectrum of water in carbon tetrachloride and a number of other solvents.[26] Two absorption bands are found in the near infrared region for gaseous H_2O. They occur at 2.728 μ and 2.665 μ and are assigned to the symmetric and asymmetric stretching motions respectively. These bands are found at 2.770 μ and 2.700 μ in carbon tetrachloride and although there is a small shift toward longer wavelengths it is considered too small to be anything more than a solvation effect due to the CCl_4. Liquid water shows a single broad band at 2.941 μ.

The development of techniques for the control of water activity in non-hydrogen bonding solvents such as carbon tetrachloride have permitted spectroscopic studies to be carried out on the hydration of various molecules.[27] The principal advantage of this approach is that specific hydration effects can be observed in the near infrared region. Studies carried out in aqueous solutions are obviously hindered by the fact that

water itself is such a good hydrogen bonding substance that no specific interactions can be studied. By controlling the activity of water in various solutions of amide in carbon tetrachloride it is possible to determine the effect of water on the self association of the amide. This has permitted the calculation of hydration constants of the monomeric and associated species in solution.

CHAPTER II

OBJECTIVES

The objectives of this research effort were threefold:

1. A spectral evaluation of the thermodynamics parameters for the association of 2-pyrrolidone in anhydrous carbon tetrachloride.

2. To determine the effect of water on the above equilibria and the calculation of hydration constants at 25°C.

3. To determine and compare the effect of water on the non-selfassociating compound N-methyl-2-pyrrolidone and to calculate hydration constants for this system at 25°C.

CHAPTER III

EXPERIMENTAL

The 2-pyrrolidone (γ -butyrolactam) was a product of the General Aniline and Film Corporation. It was purified by vacuum distillation through a 73 cm Vigreaux column. After a forerun of approximately 20 ml. had been collected the receivers were changed, and a constant boiling middle fraction of 50 ml was taken as the purified product. This was followed by redistillation through a smaller Vigreaux column (10 cm). The product was a colorless, viscous liquid at temperatures slightly above room temperature. The melting point range of the sample was 25-26°C. The purified product was stored in a stoppered flask in an evacuated desiccator.

The N-methyl-2-pyrrolidone was also a product of the General Aniline and Film Corporation. It was purified by a single vacuum distillation through the small Vigreaux column. The product was colorless and somewhat more volatile than the 2-pyrrolidone. The sample was stored in the same manner as the 2-pyrrolidone.

Special precautions were always taken in handling the sample to assure its dryness. In the purification process the vacuum was released by bleeding air back into the system through a 40 cm glass column filled with $MgClO_4$. The samples were immediately stoppered and placed in a

desiccator so that exposure to the atmosphere was minimized.

The solvent, carbon tetrachloride, was purified by the method of Krchma and Williams.[28] Sulfide impurities were removed by refluxing with mercury for a period of eight hours. This was followed by successive washings with concentrated sulfuric acid, 50% potassium hydroxide and water. The carbon tetrachloride was then placed over anhydrous $CaCl_2$ to dry for a period of twenty four hours. The final step in the purification process was a distillation through a 30 plate Oldershaw column. A special receiver containing $MgClO_4$ or $CaSO_4$ was used to collect the solvent after an initial forerun had been removed. The reflux ratio of the forerun was 30:1. The bulk of the solvent was collected at a reflux ratio of 10:1. The boiling point range, uncorrected for pressure, was 75-76°C.

All spectra were recorded on a Beckman Model DK-1 double beam ratio recording spectrophotometer. The operating range of the instrument was from 180 mμ to 3500 mμ. The radiation source for the near infrared region of the spectrum was a 6 volt incandescent tungsten lamp. The prism was cut from a single silica crystal and had a high dispersion in the near infrared region. A lead sulfide cell served as the detector system. The instrument was equipped with a modified Brown potentiometric strip recorder which recorded linearly in percent transmittance and absorbance.

The cells were products of the Beckman company and were of one, two and five centimeter path length. They were optically matched having silica faces with a high transparency in the near infrared region. The body or barrel of the cells were made from Vycor. Each cell came equipped

with a ground glass stopper.

Temperature control in the two and five centimeter cells was maintained by mounting them in a specially constructed cell holder. Water from a constant temperature water bath was circulated through the cell holder thereby holding the cells and their contents at any desired temperature. Melnick, who constructed the cell holder, has completely described it in his doctoral dissertation.[29]

A Beckman 92527 Temperature Regulated Cell Holder was used to maintain a constant temperature in the one centimeter cells. This device had two 50 watt heaters wired in parallel and installed in a special block accomodating cells from one to five centimeters in length. The temperature was selected and maintained through the operation of a thermoswitch. However, the temperature could be controlled equally well by simply circulating water from a constant temperature bath through the block. Temperature could be maintained in this way to $\pm$ .2°C within the cell as confirmed with a thermometer.

Two water baths were used for different phases of this study. The first bath consisted of a six gallon container filled with distilled water, a combination heating and stirring tower and a circulatory pump. A 250 watt light bulb painted black to eliminate nonuniform radiant heating was used as the fine control for heating the bath. The light bulb was connected to a "Thermistemp" temperature regulator with a thermistor as a sensing element. The temperature could be kept accurate to $\pm$ .5°C with this bath.

The second bath was a Haake Series "F" circulator and constituted a great improvement. The filling volume was from 1.2 to 1.7 liters

of distilled water. The temperature was controlled by two heaters of 500 watts for rapid, coarse heating and 100 watts for fine heating. The temperature could be set at any desired level by use of a thermoregulator. The principle advantage of this bath was that the small volume of water permitted rapid heating or cooling so that a large temperature range could be covered in a brief period of time. Temperature accuracy using this bath was $\pm$ .2°C.

During the first phase of this study all manipulations and operations were carried out in a dry box. Since solution preparation and sample handling was a tedious and slow task in the dry box, a new method of achieving a dry system was actively sought. The invention of equilibrators adapted especially for spectral studies proved to be a major step forward in relieving this problem.

A typical equilibrator has been dipicted in Figure 1. It was constructed from a 10 x 30 male ground glass joint which would fit directly into the spectral cells. The reservoir (A) was filled through (B) with a constant activity solution of sulfuric acid or with a drying agent such as P_2O_5. Small rubber septums were used to plug (B) so that a closed system was formed.

The equilibrators could be used in this fashion to put known amounts of water into amide-carbon tetrachloride solutions or they could be used to take water out of solutions. Figures 2 and 3 show two common experimental operations using the equilibrators.

Figure 2 shows the equilibrator being used to keep carbon tetrachloride in the reference cell dry. It was, of course, essential in making hydration studies of this type to have an absolutely dry reference

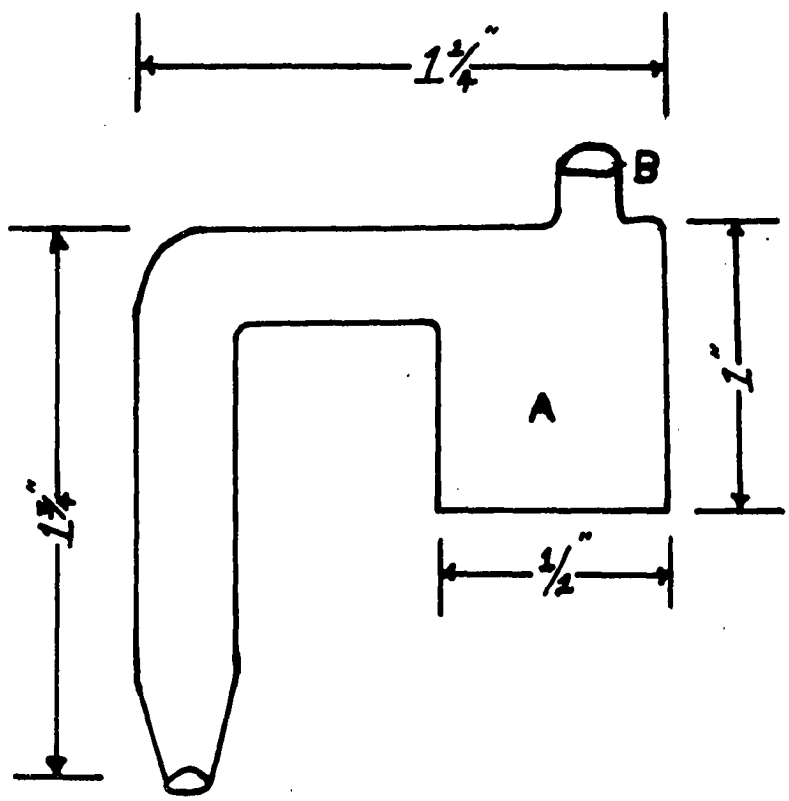


Figure 1. Equilibrator

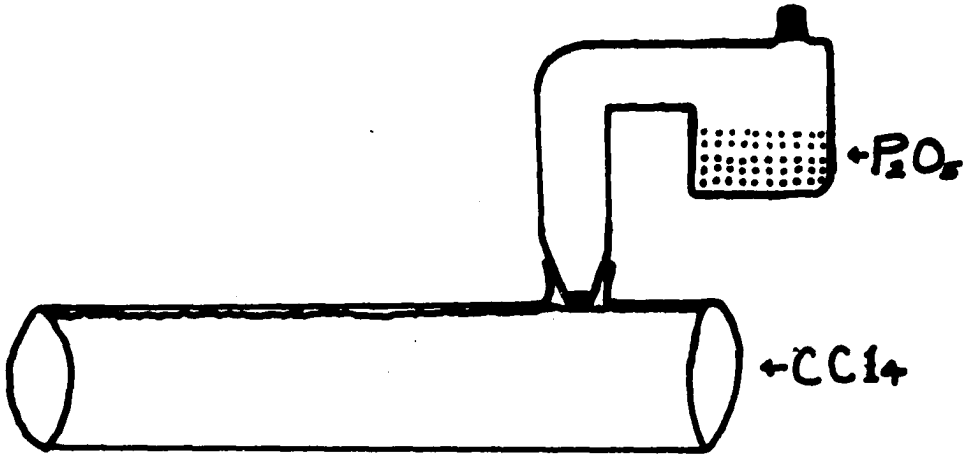


Figure 2. Drying Technique

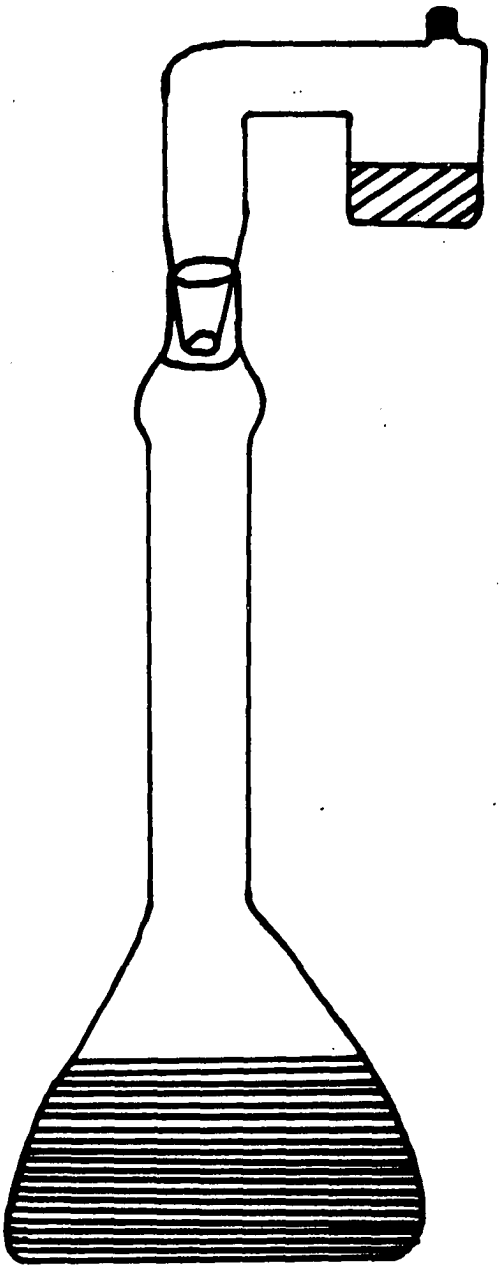


Figure 3. Equilibrating Technique

and this could always be assured by leaving the equilibrator in place during all spectral investigations.

Figure 3 shows the equilibrator attached to a 10 ml volumetric flask. The equilibrators fit 25 and 50 ml volumetric flasks equally well, and were often used with them. The equilibrator in Figure 3 contains a solution of sulfuric acid of known water activity. After equilibrium has been achieved the water activity will be the same in the carbon tetrachloride solution as in the sulfuric acid solution.

The time required for equilibrium to be established was of the order of twelve hours. Actual equilibration periods were always longer than this to assure the attainment of complete equilibrium.

The equilibrium process was generally carried out in volumetric flasks rather than in the cells themselves. Although the former requires the somewhat hazardous transfer of the equilibrated sample from the flask to the cell, it was chosen because it was faster. In the flasks as many samples as desired could be equilibrated with a large number of water activities. Only one solution at a time could be equilibrated in the cell. Gains and losses of water by the sample in the transfer process were minimized by use of a dry hypodermic syringe. The equilibrator cap was removed, and enough sample to fill the cell taken. The cell was filled by placing the hypodermic needle on the bottom of the cell and running in the solution.

All solutions in this study were prepared by weighing, and the concentrations are on the molal basis. At first solutions were made up in the dry box, but with the advent of equilibrators this technique was abandoned. Stock solutions were prepared by weighing a volumetric flask,

adding a small amount of the non-volatile solute and reweighing. The flask was then filled with dry CCl_4 and reweighed. Aliquots of the stock solution were subsequently diluted with dry CCl_4 and weighed. The concentration range for all studies carried out in the one, two and five cm cells was from 10^{-4} molal to $(3.0)(10^{-2})$ molal. Once the solutions were made up they could be easily dried by use of the equilibrators.

Solutions were scanned on the DK-1 in the following manner. The sample and reference cells were filled with dry carbon tetrachloride and capped with equilibrators filled with phosphorus pentoxide. The temperature was adjusted to the desired level and the zero and 100% adjustments made at the wavelength of the free N-H absorption at 2.879 μ. The sample cell was then filled with dry amide solution and scanned from 3500 mμ to 2500 mμ at a rate of 30 mμ /inch. A period of at least ten minutes was allowed for temperature equilibration before scanning. The slit was opened to a value of 0.175 mm at the absorption maxima. The sensitivity was always maximized to give the smallest possible slit width. A germanium filter with cutoff at about 1.9 μ was placed in the light path to remove stray light.

The wet systems were treated in an analogous fashion to the dry ones except the instrument was standardized at a wavelength of 2680 mμ, the absorption mode arising from the water-amide complex and giving a quantitative measure of this species. The mechanical slit width at this wavelength was 0.075 mm.

For the dry system all recordings were made in percent transmittance and later converted to absorbance values using log tables. The wet systems were recorded directly in absorbance.

The actual technique of equilibrating the various solutions of amide with water was a simple one. A 50 ml solution of known amide concentration was subdivided into six aliquot parts consisting of about eight ml each. These were placed in 10 ml flasks and were capped with equilibrators containing solutions of different water activities. The flasks with equilibrators in place were placed in quart fruit jars and capped with canning lids. The jars were submerged in a water bath at 25°C and allowed to stand for a period exceeding 24 hours to reach equilibrium.

After the equilibration process was completed the flasks were taken from the bath and the equilibrators removed. The flasks were immediately capped with ground glass stoppers and their absorbances measured. No solution stood longer than two hours before the absorbance measurement was made.

CHAPTER IV

THEORY

Dimerization of 2-Pyrrolidone in Anhydrous Carbon Tetrachloride

If it is assumed that 2-pyrrolidone associates in anhydrous carbon tetrachloride according to the reaction

$$2M \rightleftharpoons D \qquad (a)$$

where M and D represent the monomer and dimer, the equilibrium constant may be represented by

$$K_2 = \frac{m_D}{m_M^2} . \qquad (1)$$

The symbols m_M and m_D represent the molal concentrations of the monomer and dimer respectively. The concentration range is assumed to be low enough so actual amide concentrations can be used in place of activities.

The equilibrium constant for this reaction may be evaluated from spectral data by several methods. The total amide concentration is measurable and may be related to the monomer and dimer concentrations by the expression,

$$m_f^A = m_M + 2m_D . \qquad (2)$$

Assuming that the band at 2.879 μ is a function of monomer

concentration only, and that Beer's law is obeyed, we may write

$$A_M = \epsilon_M b m_M \qquad (3)$$

where A_M is the absorbance measured at 2.879 μ, ϵ_M is the absorptivity of the monomer and b is the light path length in the cell. Combining and rearranging equations (1), (2) and (3) the following expression may be obtained

$$\frac{m_f^A}{A_M} = \frac{2K_2}{(\epsilon_M b)^2} A_M + \frac{1}{\epsilon_M b} . \qquad (4)$$

Equation (4) is the equation of a straight line with slope equal to $2K_2/(\epsilon_M b)^2$ and intercept equal to $1/\epsilon_M b$. It is clear that a plot of the experimentally determinable quantities m_f^A/A_M vs A_M will be linear and the absorptivity of the monomer and the association constant for (1) may be calculated from the intercept and slope of the line.

A second method of evaluating equilibrium constants from spectral data utilizes Beer's law at three different wavelengths. If Beer's law is obeyed for the monomer, dimer and the C-H absorption peaks at 2.879 μ, 3.095 μ and 3.345 μ the following expressions are valid

$$A_M = \epsilon_M b m_M \qquad (5)$$

$$A_D = \epsilon_D b m_D \qquad (6)$$

and

$$A_{C-H} = \epsilon_{C-H} b m_f^A . \qquad (7)$$

Solving (5), (6) and (7) for monomer, dimer and total amide concentration

and substituting in equation (2) we have

$$\frac{A_{C-H}}{\epsilon_{C-H}b} = \frac{A_M}{\epsilon_M b} + \frac{2A_D}{\epsilon_D b} . \qquad (8)$$

Rearranging (8) gives the following

$$\frac{A_{C-H}}{A_M} = \frac{2\epsilon_{C-H}}{\epsilon_D} \frac{A_D}{A_M} + \frac{\epsilon_{C-H}}{\epsilon_M} . \qquad (9)$$

If A_{C-H}/A_M vs A_D/A_M is plotted, a straight line with slope $2\epsilon_{C-H}/\epsilon_D$ and intercept $\epsilon_{C-H}/\epsilon_M$ will result. The absorptivity of the C-H band, ϵ_{C-H}, may be determined from a Beer's law plot of the absorbance vs the total molal concentration of the amide. The monomer and dimer absorptivities may then be computed from the slope and intercept of the plot of A_{C-H}/A_M vs A_D/A_M. Having determined these quantities it is then possible to evaluate an association constant for each experimental point from the equation.

$$K_2 = \frac{\frac{A_D}{\epsilon_D b}}{\frac{A_M^2}{(\epsilon_M b)^2}} = \frac{A_D}{A_M^2} \frac{\epsilon_M^{2b}}{\epsilon_D} . \qquad (10)$$

If this type of experiment is repeated at several temperatures and if the absorptivity of the monomer can be assumed to be temperature independent, the heat of dimerization can be evaluated by making use of van't Hoff relation

$$\left[\frac{\partial \ln K_2}{\partial T}\right]_p = \frac{\Delta H_{Dim}}{RT^2} . \qquad (11)$$

By taking the logarithm of both sides of expression (10) and substituting into the integrated form of the van't Hoff expression

$$\log K_2 = \frac{-\Delta H_{Dim}}{2.3\ R}\ \frac{1}{T} + B, \qquad (12)$$

it is seen that a new equation

$$\log \frac{A_D}{A_M^2}\ \frac{\epsilon_M^{2b}}{\epsilon_D} = \frac{-\Delta H_{Dim}}{2.3\ R}\ \frac{1}{T} + B \qquad (13)$$

results. B is a constant of integration. Rearranging equation (13) gives

$$\log \frac{A_D}{A_M^2} = -\frac{\Delta H_{Dim}}{(2.3)R}\ \frac{1}{T} + B', \qquad (14)$$

from which it is seen that a plot of $\log A_D/A_M^2$ vs $1/T$ will yield the heat of dimerization from the slope of the resulting line.

The free energy and entropy are calculated from the fundamental thermodynamic expressions

$$\Delta F^\circ = - RT \ln K_2 \qquad (15)$$

and

$$\Delta S^\circ = \frac{\Delta H_{Dim} - \Delta F^\circ}{T}. \qquad (16)$$

Calculation of Hydration Constants for the System 2-Pyrrolidone-Water-Carbon Tetrachloride

To determine hydration constants for 2-pyrrolidone in carbon

tetrachloride, it is necessary to develop equations for treating spectral and water solubility data. It has been shown that water exists as the monomer in carbon tetrachloride.[30] By knowing the activity of water in the carbon tetrachloride solution it is always possible to know the concentration of free water from an application of Henry's law.

$$\frac{P_{Water}}{P^{o}_{Water}} = a_W = K_{Henry} \; m_{H_2O} \qquad (17)$$

where a_W is the water activity and K_{Henry} is Henry's constant. From a consideration of the spectrum it is evident that the asymmetric and symmetric stretching vibrations of water in carbon tetrachloride both obey Beer's law

$$A^{2660}_{asym} = \epsilon^{2660}_{H_2O} b m_{H_2O} = k\, a_W \qquad (18)$$

and

$$A^{2735}_{sym} = \epsilon^{2735}_{H_2O} b m_{H_2O} = k'\, a_W \qquad (19)$$

where k and k' are constants of proportionality containing the Henry's law constant, the absorptivity of the peak and the path length of the cell. Either of these two absorption bands may be used for analysis of free water in the system.

When water is added to a 2-pyrrolidone-carbon tetrachloride solution, striking changes are observed in the spectrum. Two new

absorption bands appear at 2680 mμ and 2860 mμ. The 2680 mμ band appears first as a shoulder on the long wavelength side of the asymmetric stretching mode at 2660 mμ. The band increases in intensity with increasing concentration of water or of 2-pyrrolidone. The 2860 band is broad and partially masked by the 2879 mμ and 3097 mμ bands at high concentration of lactam.

If the absorbance at the 2680 mμ absorption band is a function of total hydrate concentration only, the following expression may be written

$$A_{2680} = \epsilon_{Hyd} b m_{Hyd} \qquad (20)$$

where ϵ_{Hyd} is the absorptivity of each of the hydrated species and m_{Hyd} is the total molal concentration of all hydrated species. It is assumed that all hydrated species have the same absorptivity at 2680 mμ since no shift in the absorption frequency with change in amide concentration or with variation of water activity is observed.

Assuming that both the monomer and the dimer form a monohydrate the following expression is valid

$$m_f^{H_2O} = m_{M\cdot H_2O} + m_{D\cdot H_2O} + m_{H_2O} \qquad (21)$$

where $m_f^{H_2O}$ is the formal molal concentration of water, $m_{M\cdot H_2O}$ is the molal concentration of monomer monohydrate, $m_{D\cdot H_2O}$ is the molal concentration of the dimer hydrate and m_{H_2O} is the molal concentration of free monomeric water. From (20) we have the expression

$$m_{Hyd} = \frac{A_{2680}}{\epsilon_{Hyd} b} = m_{M \cdot H_2O} + m_{D \cdot H_2O} \qquad (22)$$

The formal concentration of amide is given by

$$m_f^A = m_{M \cdot H_2O} + 2m_{D \cdot H_2O} + m_M + 2m_D \qquad (23)$$

If equation (22) is subtracted from (23) it is seen that the monomer monohydrate term is eliminated, but that a dimer monohydrate term remains. The resulting expression is

$$m_f^A - \frac{A_{Hyd}}{\epsilon_{Hyd} b} = m_M + 2K_2 m_M^2 + m_{D \cdot H_2O} \qquad (24)$$

The factor $A_{Hyd}/\epsilon_{Hyd} b$ corrects the formal amide concentration for that portion of the amide which is tied up as monomer monohydrate and dimer monohydrate. But from equation (23) an additional dimer monohydrate term remains in expression (24). On the other hand, if equation (22) is multiplied by a factor of two and subtracted from equation (23), the term $-m_{M \cdot H_2O}$ remains in the resulting expression. Hence, for the purpose of determining a good approximation to $m_M + 2K_2 m_M^2$, $1\frac{1}{2}$ times equation (22) was subtracted from equation (23). This resulted in a "dry" m_f^A concentration from which the monomer concentration could be calculated using the known dimerization constant. The absorptivity, ϵ_{Hyd}, was found to be 200 molal^{-1} cm^{-1}, from a plot of A_{Hyd} vs. m_{Hyd}. The concentration of hydrate was obtained using the Karl Fischer method.[31]

Assuming hydration to occur in the following manner

$$H_2O + M \rightleftharpoons M\cdot H_2O \qquad (b)$$

and

$$H_2O + 2M \rightleftharpoons D\cdot H_2O \qquad (c)$$

hydration constants may be written as

$$K_{M\cdot H_2O} = \frac{m_{M\cdot H_2O}}{m_M \cdot m_{H_2O}} \qquad (25)$$

and

$$K_{2M\cdot H_2O} = \frac{m_{D\cdot H_2O}}{m_M^2 \, m_{H_2O}}. \qquad (26)$$

Rewriting (25) and (26) by substituting from equation (17) and solving for $m_{D\cdot H_2O}$ and $m_{M\cdot H_2O}$ and substituting these values in (22), it is seen that

$$\frac{A_{Hyd}}{\epsilon_{Hyd}{}^b} = K'_{M\cdot H_2O} \; a_W m_M + K'_{2M\cdot H_2O} \; a_W m_M^2 \qquad (27)$$

is obtained. A plot of $A_{Hyd}/a_W \epsilon_{Hyd}{}^b m_M$ vs. m_M yields an initial estimate of $K'_{M\cdot H_2O}$ from the intercept and $K'_{D\cdot H_2O}$ from the slope. It should be pointed out that $K'_{M\cdot H_2O}$ and $K'_{D\cdot H_2O}$ differ from the hydration constants in equations (25) and (26) by the fact that the former have been divided by the Henry's law constant. Hydration constants corresponding to those of equations (25) and (26) may be obtained by multiplying the slope and intercept of equation (27) by K_{Henry}/K_{Dim} and K_{Henry} respectively.

Using a generalized least squares computer program an error function is calculated using equation (27) and the initial estimates of $K_{M\cdot H_2O}$

and $K_{D\cdot H_2O}$. By varying $K_{M\cdot H_2O}$ and $K_{D\cdot H_2O}$ the error function may be minimized and the best fitted values of the hydration constants obtained.

Calculation of Hydration Constants for the System N-Methyl-2-pyrrolidone-Water-Carbon Tetrachloride

The spectrum of solutions of N-methyl-2-pyrrolidone and water in carbon tetrachloride is nearly identical to that of 2-pyrrolidone and water in carbon tetrachloride except the former system shows no absorption bands due to self-association. Consequently, the broad 2860 mμ band due to the water-amide complex is relatively free from interfering absorption.

Assuming hydration occurs in the following manner

$$H_2O + M' \rightleftharpoons M'H_2O \qquad (d)$$

the hydration constant may be expressed as

$$K_{Hyd} = \frac{m_{M'\cdot H_2O}}{(m_f^A - m_{M'\cdot H_2O})m_{H_2O}} \qquad (28)$$

where $m_{M'\cdot H_2O}$ represents the molal concentration of the amide-water complex, m_f^A is the total amide concentration and m_{H_2O} is the free water. If both of the bands at 2680 mμ and 2860 mμ are assumed to be solely produced by the complex the following expressions are valid

$$A_{2680} = \epsilon'_{Hyd} b m_{M'\cdot H_2O} \qquad (29)$$

and

$$A_{2860} = \epsilon''_{Hyd} b m_{M\cdot H_2O} . \qquad (30)$$

Substituting (17) and (29) into (28) and rearranging, an equation of the form

$$\frac{m_f^A}{A_{2680}} = \frac{1}{K'_{Hyd}\,\epsilon_{Hyd}{}^b}\,\frac{1}{a_W} + \frac{1}{\epsilon_{Hyd}{}^b} \qquad (31)$$

is obtained. K'_{Hyd} now contains the Henry's law constant, and all other terms have been previously defined. Plotting m_f^A/A_{2680} vs. $1/a_W$ yields the absorptivity of the complex and the hydration constant from the intercept and the slope. A linear least squares program gave values of K'_{Hyd} and ϵ'_{Hyd} which gave the best fit to the data. Similar operations carried out with the 2860 mμ absorption band should give a new absorptivity, ϵ''_{Hyd}, and the same hydration constant, K'_{Hyd}.

CHAPTER V

RESULTS

The spectra of 2-pyrrolidone in anhydrous carbon tetrachloride at several temperatures are shown in Figure 4. The sharp band at 2879 mμ has been assigned to the free N-H stretching vibration of the monomer. The two broad peaks at 3095 mμ and 3210 mμ are assigned to the dimer. The ratio of these two peaks is constant with varying temperature and concentration indicating that both bands arise from the same associated species. The peak at 3345 mμ has its origin in the C-H stretching vibrations. It should be emphasized that there is no observed shift of frequency or alteration in appearance for any of these peaks over the temperature and concentration ranges studied. Liddel and Becker have asserted that the absorptivity for monomeric methanol in carbon tetrachloride is a function of temperature, and changes by as much as 20-50% over a 60° temperature range.[32,33] No such changes were observed in the absorptivities of the 2-pyrrolidone and the N-methyl-2-pyrrolidone used in this study. Data were treated by a method similar to that used by Affsprung, Christian and Melnick[34] for carboxylic acids, and it was observed that the absorptivity for the free N-H stretching absorption maxima varied in a random fashion with no systematic dependence on temperature. Although this method of data analysis was

developed in some detail in Chapter IV, the results of this method will not be presented here since the only information of value gained from it was that there was no temperature dependence of the absorptivities. The actual absorptivities and equilibrium constants were calculated using a much more reliable method.

Tables I and II contain the data used in determining the absorptivities of the C-H (3345 mμ), monomer (2879 mμ) and dimer (3095 mμ) absorption peaks. The C-H absorptivity was determined from the slope of the line of a Beer's law plot. Figure 5 shows a plot of these data in the form A_{C-H} vs. A_D/A_M. The monomer and dimer absorptivities were obtained from the intercept and the slope of the line, which was fitted by a linear least squares method using weighting factors proportional to the value of A_{C-H}. Table III presents the values of the monomer, dimer and C-H absorptivities evaluated by this method.

Tables IV and V present monomer absorbances, dimer absorbances and the ratio of A_D/A_M^2 for the two and five centimeter cell data. Table VI contains equilibrium constants calculated using the absorptivities in Table III and the ratios of A_D/A_M^2 in Tables V and VI. Figure 6 is a plot of these equilibrium constants on a logarithm scale versus the reciprocal of the absolute temperature. The dashed line represents a least squares fit of the five centimeter cell data. The dot-dash line is a least squares fit for all data taken in the two centimeter cells. The solid line is a least squares fit representing all experimental points. The heat of dimerization may be calculated from the slopes of the lines.

Figure 7 depicts the near infrared spectrum of water in carbon tetrachloride at several concentrations. The band at 2660 mμ arises from

the asymmetric stretching vibrations of the water molecule, while the band at 2735 mμ has its origin in the symmetric vibrational mode. No frequency shift or change in the half-band width for either of these peaks was observed over a 50° temperature range. This is a strong indication that water does not self associate in carbon tetrachloride. Tables VII and VIII contain data which were used to verify the obedience to Beer's law at these wavelengths. The formal water concentrations in carbon tetrachloride were determined by Karl Fischer titration. The absorptivity of the peak at 2660 mμ was found to have a value of 29 $molar^{-1}_{cm^{-1}}$.

The spectra of 2-pyrrolidone in dry carbon tetrachloride and that of the same solution of 2-pyrrolidone in carbon tetrachloride equilibrated with solutions of constant water activities of 0.476 and 0.807 are shown in Figure 8. Note that there is very little change in the intensities of the monomer, dimer and C-H peaks. A new peak appearing at 2680 mμ has been assigned to the hydrated species. A second band broader in character appears at 2860 mμ. This band is also assigned to the hydrated species. A more detailed discussion of the origins of these two bands will be given in Chapter VI. The absorbance due to the band at 2860 mμ is difficult to measure in the case of 2-pyrrolidone because of interfering absorption from the free (N-H) and hydrogen bonded (N-H...O) stretching modes of the monomer and dimer. A differential spectrum shows this band clearly, since nearly all of the monomer and dimer absorptions cancel out.

Table IX contains the raw and corrected absorbances of the 2680 mμ peak for a series of solutions of 2-pyrrolidone in carbon tetrachloride

over a range of concentrations and water activities. Corrections were made for the infringement of the free water peak at 2660 mμ on the hydrate by subtracting the contribution of free water to the absorbance at 2680 mμ using a series of spectra of water in carbon tetrachloride. These blanks were run at the same water activities as those used in equilibrating the 2-pyrrolidone solution. The spectrum of the appropriate blank was placed beneath the spectrum of 2-pyrrolidone-water-carbon tetrachloride and adjusted until the symmetric peaks of the free water at 2735 mμ matched exactly. Then a perpendicular line was drawn along the ordinate through the center of the 2680 mμ peak until it intersected the wing of the 2660 mμ peak. The absorbance value at this intersection of the 2-pyrrolidone-water-carbon tetrachloride spectrum was subtracted from the peak absorbance value at 2680 mμ giving a value corrected for infringement and background. Figure 9 shows a plot of the corrected absorbance at 2680 mμ versus the formal concentration of 2-pyrrolidone over a range of water activities. The lines were calculated from a least squares treatment of the data.

Fragmentary evidence from equilibrium studies on the 2-pyrrolidone-water-carbon tetrachloride system using the Karl Fischer method indicate that the absorptivity of the hydrated species is about 200 $molal^{-1}cm^{-1}$. This value was used to obtain initial estimates of the hydration constants for reactions (b) and (c) in Chapter IV. It should be pointed out that this value cannot be taken as completely conclusive, since there are not enough data available to obtain a value with any large degree of certainty.

Figure 10 depicts the spectrum of N-methyl-2-pyrrolidone at three water activities. The broad band at 2860 mμ is very prominent in these spectra. The peak at 2680 mμ and the band at 2860 mμ are attributed to

the hydrated species. The broad band at 2910 mμ is apparently due to the first overtone of the carbonyl stretching mode.

Table X contains the raw and corrected absorbances for the 2680 mμ peak along with the concentrations of N-methyl-2-pyrrolidone and the range of water activities. Table XI contains the absorbance values of the peak at 2860 mμ. These values are taken directly from the recorded spectra.

Tables XII and XIII contain the data in a form used for the determination of the hydration constants from the 2680 mμ and 2860 mμ peaks respectively. In the first instance the intercept was forced to a value which would give an absorptivity equal to that of the hydrated species of the 2-pyrrolidone-water-carbon tetrachloride system. The absorptivity for the 2860 mμ band was determined from the average of the ratios of the 2680 mμ and 2860 mμ peak absorbances. This ratio was found to have a value of 1.219 and corresponds to an absorptivity at 2860 mμ of 164 $molal^{-1}cm^{-1}$. These data are illustrated graphically in Figure 11. It is interesting to note that the hydration constant for the N-methyl-2-pyrrolidone system is almost identical to that of the hydration constant for the 2-pyrrolidone monomer.

Table XIV presents a summary of the calculated thermodynamic constants at 25°C for all systems studied in this investigation. Since the wet systems were studied at only one temperature the enthalpies and entropies could not be calculated.

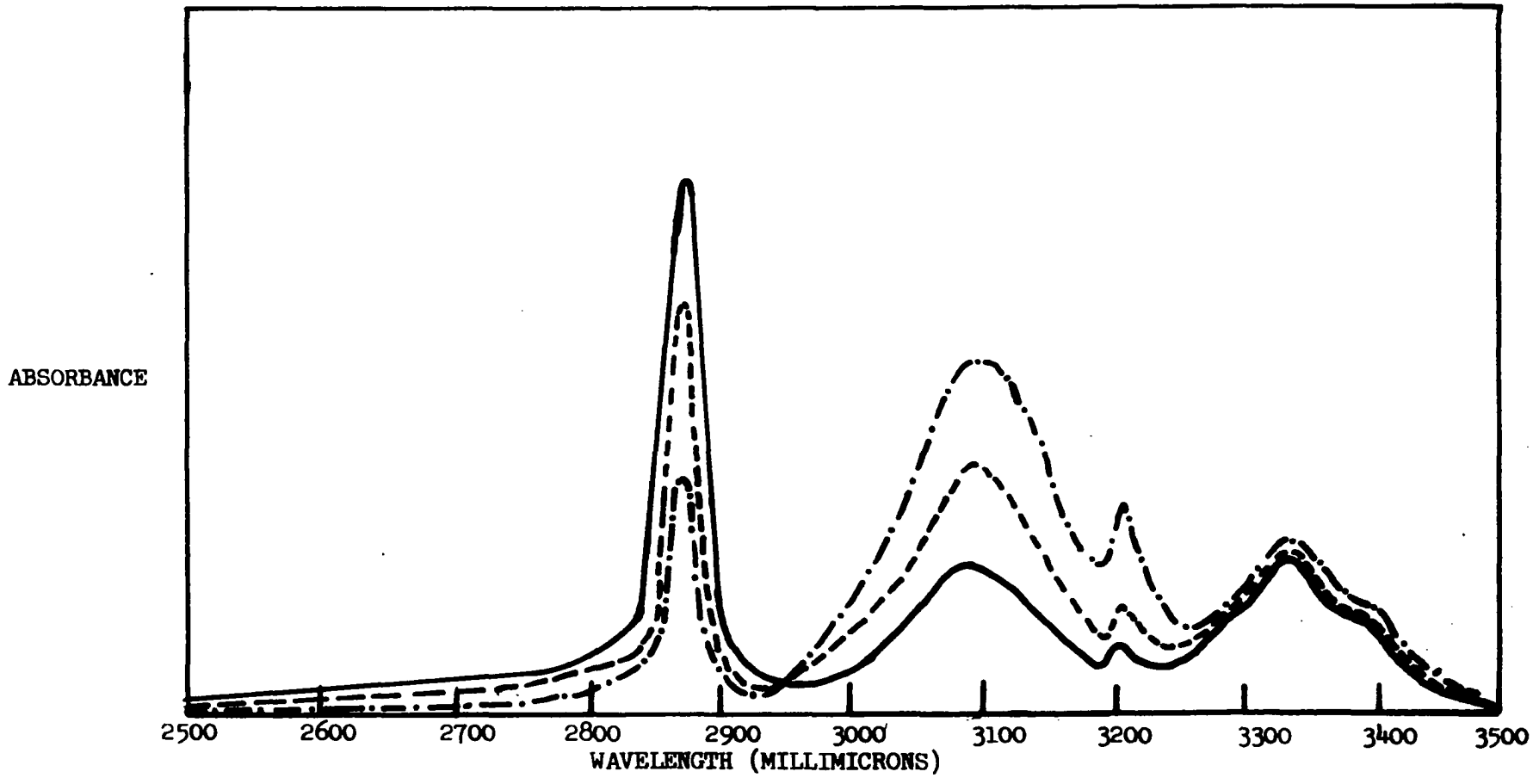


Figure 4. Spectra of Dilute Solution of 2-pyrrolidone in Carbon Tetrachloride at 20°(—·—·—), 30°(— — —) and 40°(——).

TABLE I

2-Pyrrolidone-Carbon Tetrachloride

Data used to evaluate the absorptivity, ϵ_{C-H}, at 3345 mμ in cells of five centimeter path length

t°C.	$(m_f^A)(10^3)$ molal units	A_{C-H}	t°C.	$(m_f^A)(10^3)$ molal units	A_{C-H}
10.0	0.908	0.224	30.0	1.010	0.239
	.712	.178		0.785	.194
	.496	.126		0.649	.156
	.307	.077		0.552	.130
	.144	.052		0.389	.093
				0.233	.055
				0.142	.036
15.0	1.010	0.246	40.0	0.952	0.214
	0.649	.163		.829	.188
	0.552	.137		.683	.154
	0.389	.102		.586	.129
	0.233	.061		.437	.100
	0.142	.033		.281	.060
				.142	.031
25.0	0.819	0.188			
	.531	.119			
	.302	.066			
	.194	.043			

TABLE II

2-Pyrrolidone-Carbon Tetrachloride

Data used to evaluate monomer and dimer absorptivities at 2879 mμ and 3097 mμ in cells of five centimeter path length

t°C.	$(m_f^A)(10^3)$ molal units	A_M	A_D	A_{C-H}/A_M	A_D/A_M
10.0	0.908	0.458	0.354	0.489	0.773
	.712	.405	.238	.440	.580
	.496	.317	.155	.397	.489
	.307	.215	.069	.358	.321
	.144	.137	.037	.380	.270
15.0	1.010	0.550	0.346	0.447	0.629
	0.785	.463	.238	.419	.514
	0.649	.406	.185	.401	.456
	0.552	.357	.137	.384	.384
	0.389	.270	.095	.378	.352
	0.233	.192	.038	.318	.198
	0.142	.119	.018	.277	.151
25.0	0.819	0.482	0.202	0.390	0.419
	.531	.393	.102	.303	.260
	.302	.226	.037	.292	.164
	.194	.155	.018	.277	.116

TABLE II Continued -

t° C.	$(m_f^A)(10^3)$ molal units	A_M	A_D	A_{C-H}/A_M	A_D/A_M
30.0	1.010	0.595	0.256	0.402	0.430
	0.785	.508	.194	.382	.382
	0.649	.432	.135	.361	.312
	0.552	.375	.099	.347	.264
	0.389	.298	.060	.312	.201
	0.233	.187	.022	.294	.118
	0.142	.118	.009	.305	.076
40.0	0.952	0.593	0.191	0.361	0.322
	.829	.532	.151	.353	.284
	.683	.457	.109	.337	.239
	.586	.403	.089	.320	.221
	.437	.321	.053	.312	.165
	.281	.223	.024	.269	.108
	.142	.123	---	.252	---

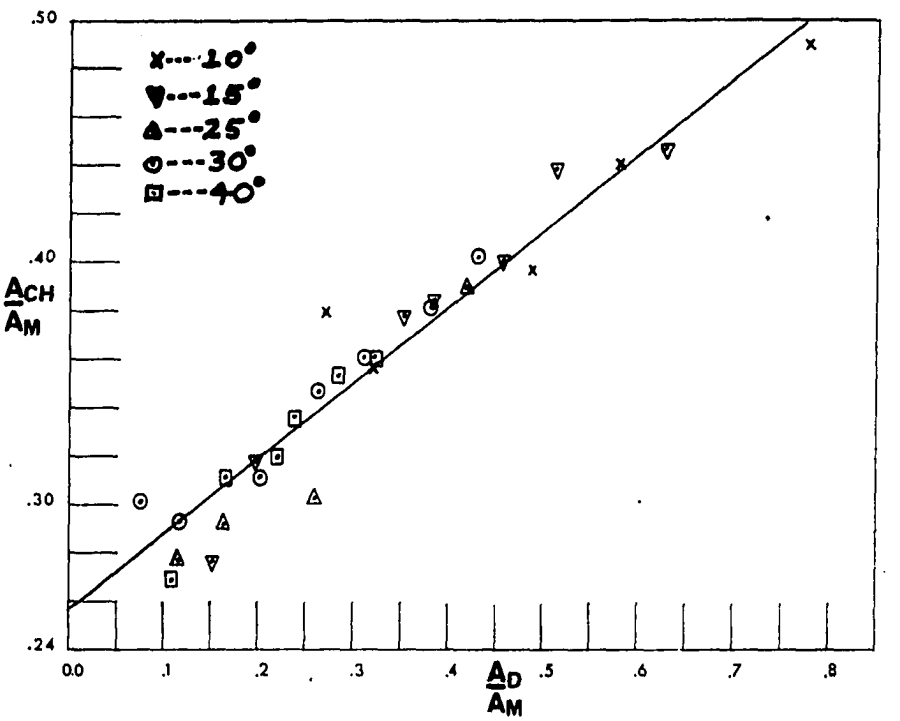


Figure 5. Relation of the ratio of the Absorbance at 3345 mμ to the Absorbance at 2879 mμ and the Absorbance at 3095 mμ to that at 2879 mμ at the Temperatures Indicated.

TABLE III

2-Pyrrolidone-Carbon Tetrachloride

Best values of absorptivities using computer least squares intercept and slope

ϵ_{C-H} molal^{-1} cm^{-1} (3345 mu)	ϵ_{M} molal^{-1} cm^{-1} (2879 mu)	ϵ_{D} molal^{-1} cm^{-1} (3095 mu)
47±1	183+8	303+24

TABLE IV

2-Pyrrolidone-Carbon Tetrachloride

Data used to evaluate K_2 and ΔH_{Dim} in cells of two centimeter path length

$(m_f^A)(10^3)$ molal units	t°C.	A_D	A_M	A_D/A_M^2
1.641	19.7	0.243	0.289	2.909
	25.6	.215	.302	2.357
	32.5	.181	.327	1.693
	38.3	.156	.339	1.357
	44.7	.132	.349	1.084
1.798	27.0	0.201	0.342	1.718
	32.5	.175	.354	1.396
	38.5	.146	.367	1.084
	43.4	.129	.371	0.937
1.990	23.5	0.239	0.339	2.080
	28.7	.210	.348	1.734
	36.7	.169	.375	1.202
	42.6	.137	.384	0.929
	47.2	.126	.385	0.850

TABLE IV Continued -

$(m_f^A)(10^3)$ molal units	t°C.	A_D	A_M	A_D/A_M^2
2.263	20.1	0.315	0.357	2.472
	25.6	.285	.374	2.038
	30.2	.254	.399	1.595
	34.5	.228	.403	1.404
	40.7	.196	.419	1.116
	47.8	.162	.439	0.840
2.863	24.6	0.480	0.449	2.381
	30.3	.428	.461	2.014
	35.1	.389	.487	1.640
	39.5	.367	.516	1.378
3.430	11.1	0.654	0.409	3.910
	15.3	.618	.440	3.192
	20.1	.592	.463	2.762
	22.6	.567	.478	2.482
	31.8	.502	.521	1.849
	41.7	.420	.554	1.368

TABLE V

2-Pyrrolidone-Carbon Tetrachloride

Data used to determine ΔH_{Dim} and K_2 in cells of five centimeter path length

t°C.	$(m_f^A)(10^3)$ molal units	A_D	A_M	$(A_D/A_M^2)2.5$
10.0	0.908	0.354	0.458	4.220
	.712	.238	.405	3.628
	.496	.155	.317	3.855
	.307	.069	.215	3.732
	.144	.037	.137	4.928
15.0	1.010	0.346	0.550	2.828
	0.785	.238	.463	2.775
	0.649	.185	.406	2.805
	0.552	.137	.357	2.688
	0.389	.095	.270	3.258
	0.233	.038	.192	2.578
	0.142	.018	.119	3.178
25.0	0.819	0.202	0.482	2.172
	.531	.102	.393	1.620
	.302	.037	.226	1.810
	.194	.018	.155	1.872

TABLE V Continued -

$t^{o}C.$	$(m_f^A)(10^3)$ molal units	A_D	A_M	$(A_D/A_M^2)2.5$
30.0	1.010	0.256	0.595	1.808
	0.785	.194	.508	1.880
	0.649	.135	.432	1.808
	0.552	.099	.375	1.760
	0.389	.060	.298	1.690
	0.233	.022	.187	1.572
	0.142	.009	.118	1.615
40.0	0.952	0.191	0.593	1.358
	.829	.151	.532	1.335
	.683	.109	.457	1.305
	.586	.089	.403	1.370
	.437	.053	.321	1.285
	.281	.024	.223	1.208
	.142	---	.123	---

TABLE VI

2-Pyrrolidone-Carbon Tetrachloride

Data used to evaluate ΔH_{Dim} and K_2 in cells of two centimeter path length

$(m_f^A)(10^3)$ molal units	t°C.	K molal^{-1} units
1.641	19.7	643
	25.6	521
	32.5	374
	38.3	299
	44.7	239
1.798	27.0	379
	32.5	308
	38.5	239
	43.4	207
1.990	23.5	459
	28.7	383
	36.7	265
	42.6	205
	47.2	187

TABLE VI Continued -

$(m_f^A)(10^3)$ molal units	t°C.	K. molal^{-1} units
2.263	20.1	546
	25.6	450
	30.2	352
	34.5	310
	40.7	246
	47.8	185
2.863	24.6	526
	30.3	445
	35.1	362
	39.5	304
3.430	11.1	864
	15.3	705
	20.1	610
	22.6	548
	31.8	408
	41.7	302

TABLE VI Continued -

$(m_f^A)(10^3)$ molal units	t°C.	K molal^{-1} units
0.908	10.0	932
.712		801
.496		851
.307		824
.144		1089
1.010	15.0	624
0.785		613
0.649		619
0.552		594
0.389		720
0.233		569
0.142		702
0.819	25.0	480
.531		358
.302		400
.194		413

TABLE VI Continued -

$(m_f^A)(10^3)$ molal units	$t^oC.$	K $molal^{-1}$ units
1.010	30.0	399
0.785		415
0.649		399
0.552		388
0.389		373
0.233		347
0.142		356
0.952	40.0	300
.829		295
.683		288
.586		302
.437		283
.281		266
.142		---

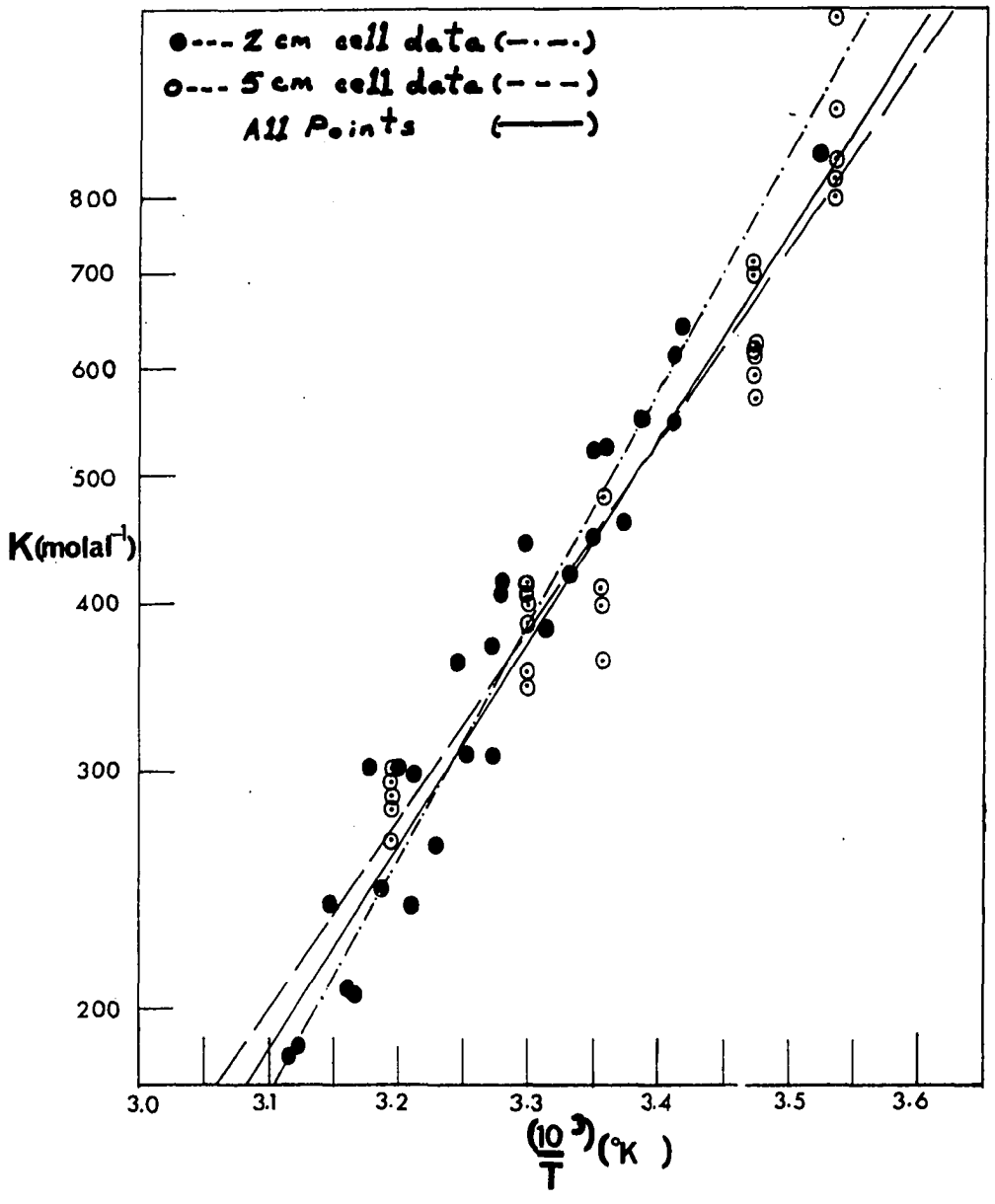


Figure 6. Variation of Log K_2 with 1/T for the System 2-pyrrolidone - Carbon Tetrachloride.

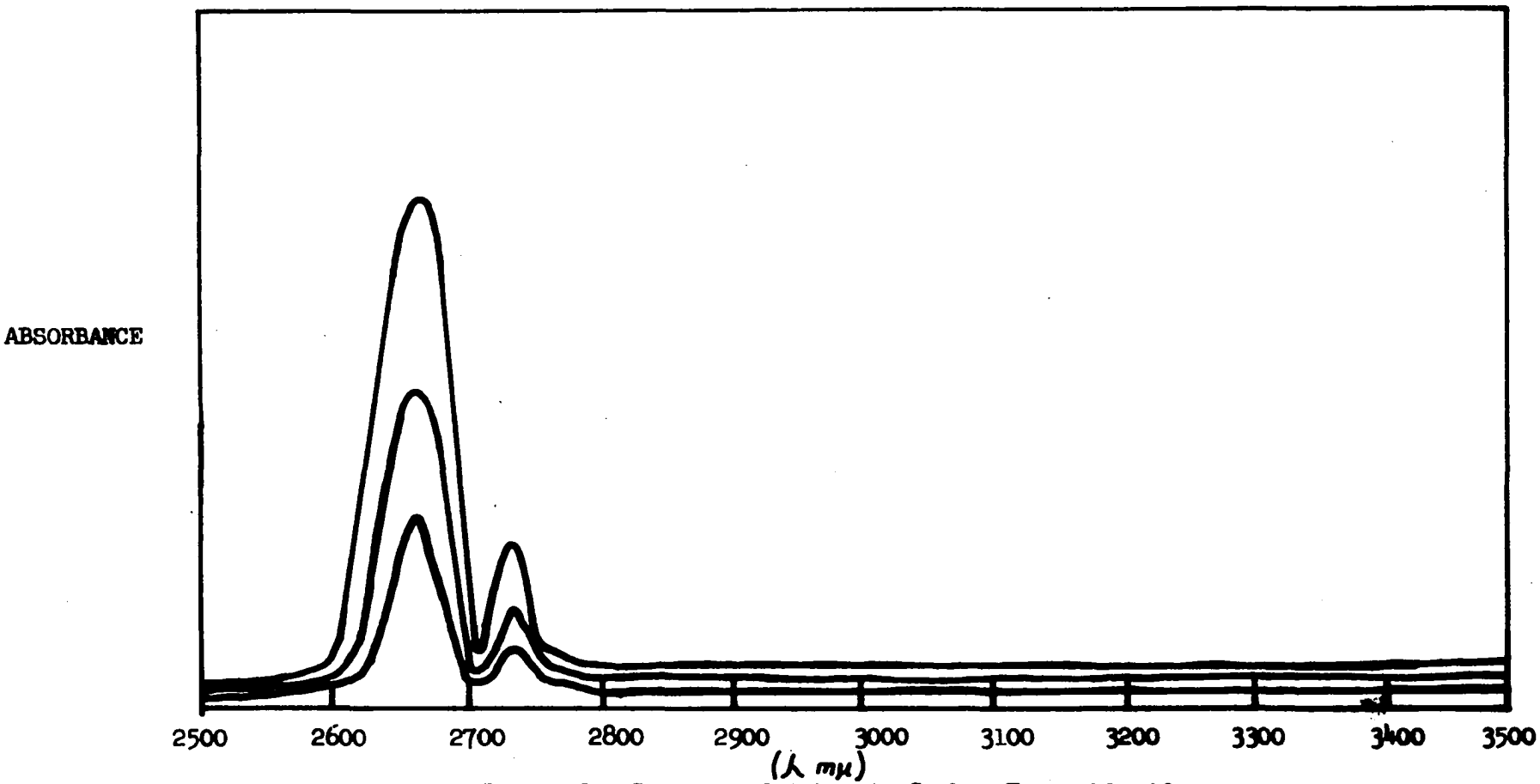


Figure 7. Spectra of Water in Carbon Tetrachloride at Three Concentrations and at 25°C.

TABLE VII

Water-Carbon Tetrachloride

Data used to verify Beer's law in cells of 1 centimeter path length. These data were used to correct for infringement of the free water peak on the hydrate peak.

a_W	$A_{2660\ m\mu}$	$A_{2735\ m\mu}$
0.256	0.095	0.029
.476	.154	.051
.615	.203	.076
.693	.224	.082
.807	.268	.103
.890	.285	.111
.945	.312	.121

TABLE VIII

Water-Carbon Tetrachloride

Data used to verify Beer's law in cells of 2 centimeter path length. Formal concentrations of water were determined by Karl Fischer titration.

$A_{2660\ m\mu}$	$(M_f^{H_2O})(10^3)$ molar units
0.111	2.23
.190	3.24
.253	4.42
.362	5.98
.488	8.44
.551	9.22

$$\epsilon_{H_2O} = 29\ \text{molar}^{-1}\text{cm}^{-1}$$

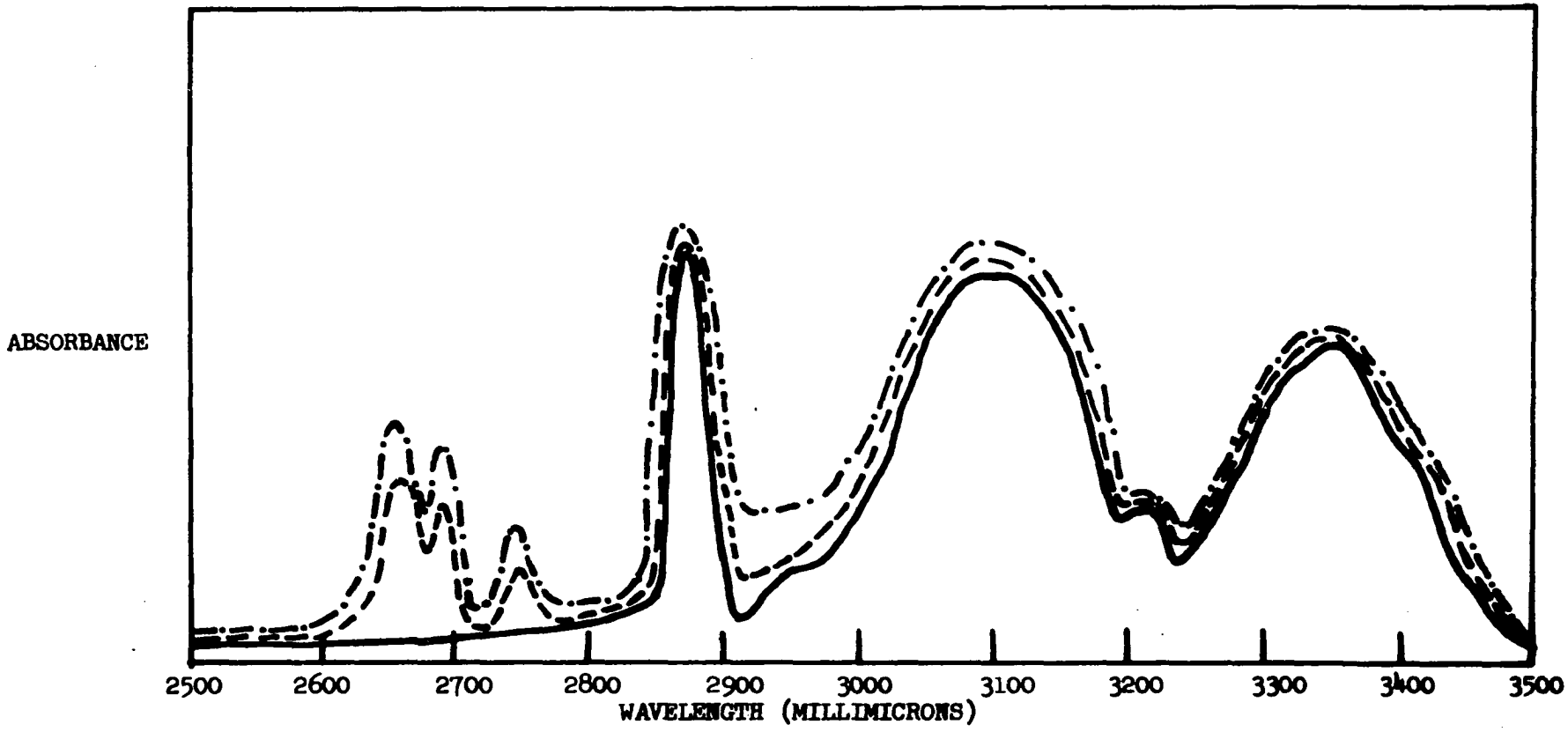


Figure 8. Spectra of a Single Dilute Solution of 2-pyrrolidone in Carbon Tetrachloride Under Anhydrous Conditions(——) and at Water Activities of 0.476(– – –) and 0.807(—·—·—).

TABLE IX

2-Pyrrolidone-Water-Carbon Tetrachloride

Data used to evaluate hydration constants for 2-pyrrolidone monomer and dimer

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
0.747	0.256	0.037	0.099
0.839		.042	.098
0.921		.047	.099
1.121		.052	.112
1.330		.053	.120
1.427		.055	.130
1.579		.068	.137
1.984		.075	.147
2.157		.096	.159
2.653		.102	.181
0.747	0.476	0.063	0.150
0.839		.057	.144
0.921		.062	.151
1.121		.078	.171
1.330		.089	.177
1.427		.106	.195
1.579		.113	.213
1.984		.132	.232
2.157		.147	.245
2.653		.174	.279

TABLE IX Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
0.747	0.615	0.071	0.189
0.839		.081	.192
0.921		.097	.205
1.121		.107	.230
1.330		.123	.243
1.427		.131	.254
1.579		.160	.287
1.984		.187	.319
2.157		.203	.349
2.653		.247	.390
0.747	0.693	0.100	0.215
0.839		.092	.202
0.921		.103	.229
1.121		.127	.248
1.330		.135	.263
1.427		.156	.280
1.579		.180	.320
1.984		.203	.346
2.157		.232	.362
2.653		.279	.420

TABLE IV Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
0.747	0.807	0.104	0.241
0.839		.111	.242
0.921		.113	.261
1.121		.147	.286
1.330		.160	.301
1.427		.172	.321
1.579		.200	.351
1.984		.234	.392
2.157		.270	.416
2.653		.312	.472
0.747	0.945	0.121	0.275
0.839		.133	.277
0.921		.140	.305
1.121		.162	.327
1.330		.187	.365
1.427		.206	.363
1.579		.232	.405
1.984		.270	.451
2.157		.291	.453
2.653		.333	.514

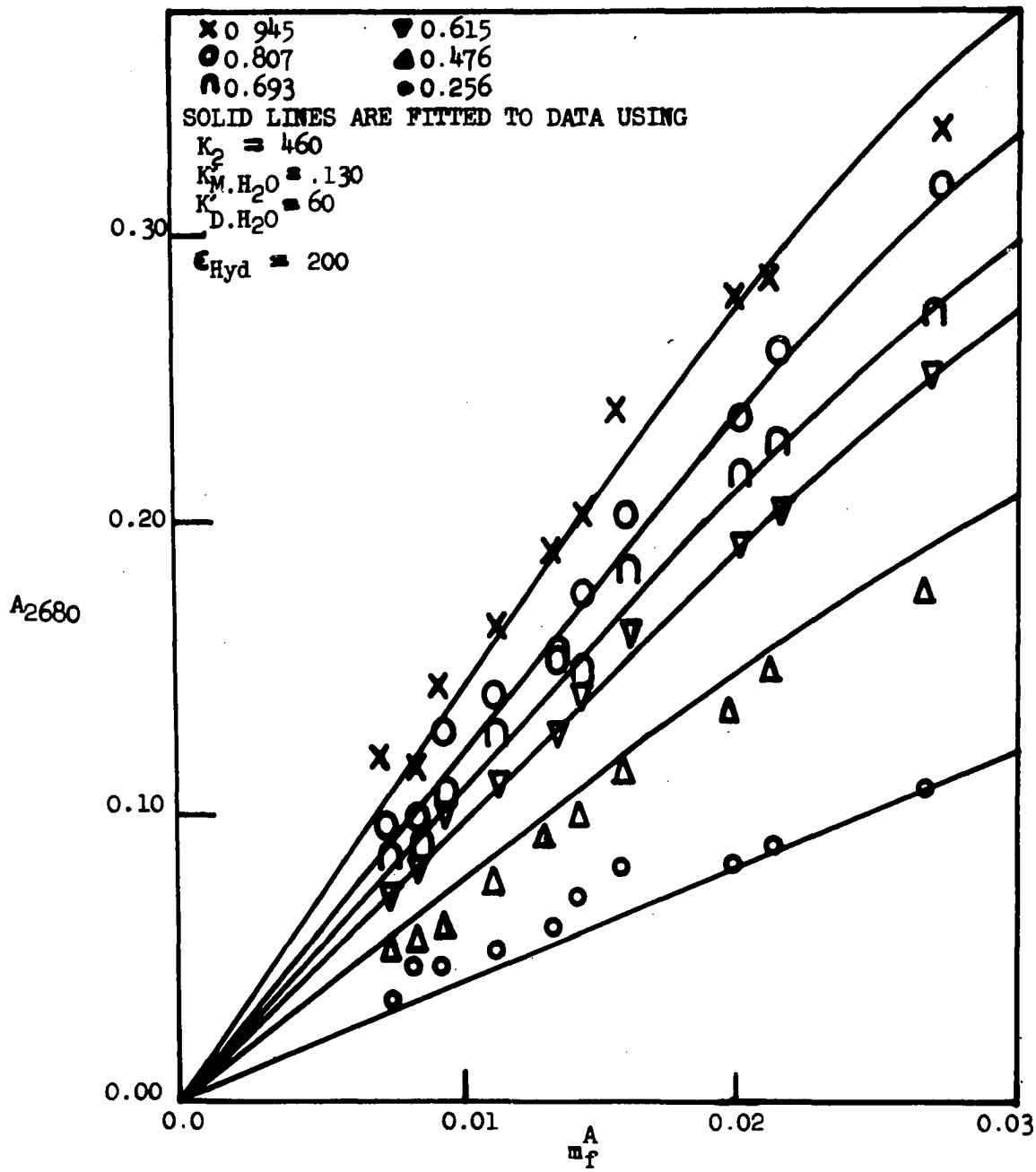


Figure 9. Relation of the Corrected Peak Absorbance at 2680 mμ and the Formal 2-pyrrolidone Concentration at the Indicated Water Activities.

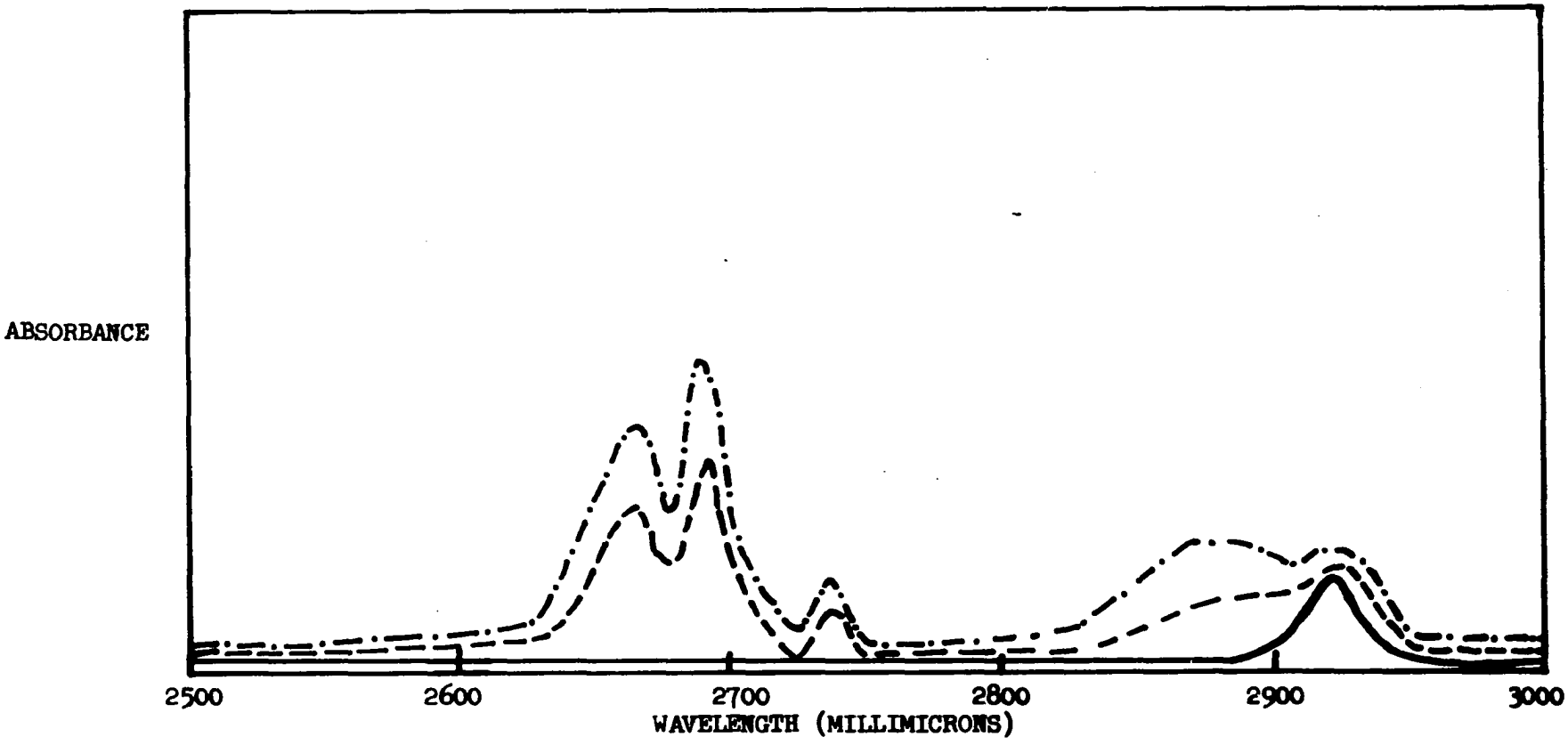


Figure 10. Spectra of a Single Dilute Solution of N-Methyl 2-pyrrolidone in Carbon Tetrachloride at Water Activities of 0.0(——), 0.476(– – –) and 0.807(–·–·–).

TABLE X

N-Methyl-2-pyrrolidone-Water-Carbon Tetrachloride

Corrected and raw absorbance values at 2680 mμ in cells of one centimeter path length

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
1.002	0.256	0.060	0.137
1.269		.072	.131
1.505		.085	.150
1.685		.097	.158
1.993		.122	.181
2.230		.134	.195
2.473		.180	.259
2.921		.222	.302
1.002	0.476	0.094	0.192
1.269		.117	.204
1.505		.131	.231
1.685		.146	.241
1.993		.173	.271
2.230		.178	.276
2.473		.250	.359
2.921		.275	.387

TABLE X Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
1.002	0.615	0.139	0.253
1.269		.149	.273
1.505		.177	.299
1.685		.201	.326
1.993		.221	.340
2.223		.246	.359
2.473		.322	.453
2.921		.360	.505
1.002	0.693	0.154	0.279
1.269		.163	.291
1.505		.193	.332
1.685		.217	.345
1.993		.240	.380
2.223		.256	.381
2.473		.351	.498
2.921		.404	.550

TABLE X Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2680\ m\mu}$	$A^{Raw}_{2680\ m\mu}$
1.002	0.807	0.161	0.320
1.269		.188	.340
1.505		.204	.372
1.685		.252	.402
1.993		.302	.457
2.230		.287	.428
2.473		.396	.573
2.921		.429	.621
1.002	0.945	0.174	0.353
1.269		.203	.378
1.505		.256	.431
1.685		.283	.482
1.993		.335	.524
2.230		.377	.534
2.473		.444	.661
2.921		.559	.738

TABLE XI

N-Methyl-2-pyrrolidone-Water-Carbon Tetrachloride

Peak absorbance at 2860 mμ at 25°C. in cells of one centimeter path length

$(m_f^A)(10^2)$ molal units	a_W	$A^{Corr}_{2860\ m\mu}$
1.002	0.256	0.050
1.269		.052
1.505		.071
1.685		.078
1.993		.090
2.230		.103
2.473		.153
2.921		.188
1.002	0.476	0.076
1.269		.090
1.505		.110
1.685		.122
1.995		.142
2.230		.150
2.473		.212
2.921		.234

TABLE XI Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A_{2860\ m\mu}$
1.002	0.615	0.148
1.269		.178
1.505		.213
1.685		.247
1.993		.290
2.230		.301
2.473		.380
2.921		.437
1.002	0.693	0.111
1.269		.131
1.505		.160
1.685		.176
1.993		.209
2.230		.215
2.473		.292
2.921		.330

TABLE XI Continued -

$(m_f^A)(10^2)$ molal units	a_W	$A_{2860\ m\mu}$
1.002	0.807	0.130
1.269		.153
1.505		.183
1.685		.206
1.993		.246
2.230		.236
2.473		.333
2.921		.372
1.002	0.945	0.148
1.269		.178
1.505		.213
1.685		.247
1.993		.290
2.230		.301
2.473		.380
2.921		.437

TABLE XII

N-Methyl-2-pyrrolidone-Water-Carbon Tetrachloride

Data used for the determination of the hydration constant in cells of one centimeter path length

m_f^A/A_{2680}^{Corr}	a_W	$1/a_W$
0.167	0.256	3.906
.176		
.177		
.173		
.163		
.166		
.137		
.131		
0.106	0.476	2.100
.108		
.114		
.115		
.125		
.098		
.106		
0.0720	0.615	1.626
.0851		
.0850		
.0838		
.0901		
.0906		
.0768		
.0811		

TABLE XII Continued -

m_f^A/A_{2680}^{Corr}	a_W	$1/a_W$
0.0650	0.693	1.443
.0778		
.0779		
.0776		
.0830		
.0871		
.0704		
.0723		
0.0622	0.807	1.239
.0675		
.0737		
.0668		
.0659		
.0777		
.0624		
.0680		
0.0575	0.945	1.058
.0625		
.0587		
.0595		
.0594		
.0591		
.0556		
.0522		

TABLE XIII

N-Methyl-2-pyrrolidone-Water-Carbon Tetrachloride

Data used for the evaluation of the hydration constant from the absorbance at 2860 mμ at 25°C.

$A_f/A_{2860\ m\mu}$	a_W	$1/a_W$
0.200	0.256	3.906
.244		
.211		
.216		
.221		
.216		
.161		
.155		
0.131	0.476	2.100
.141		
.136		
.138		
.140		
.148		
.116		
.124		
0.099	0.615	1.626
.105		
.103		
.110		
.110		
.112		
.092		
.096		

TABLE XIII Continued -

m_f^A/A_{2860}	a_W	$1/a_W$
0.0900	0.693	1.443
.0968		
.0940		
.0957		
.0953		
.1030		
.0846		
.0885		
0.0770	0.807	1.239
.0829		
.0822		
.0817		
.0810		
.0944		
.0742		
.0785		
0.0677	0.945	2.100
.0712		
.0706		
.0682		
.0740		
.0650		
.0668		

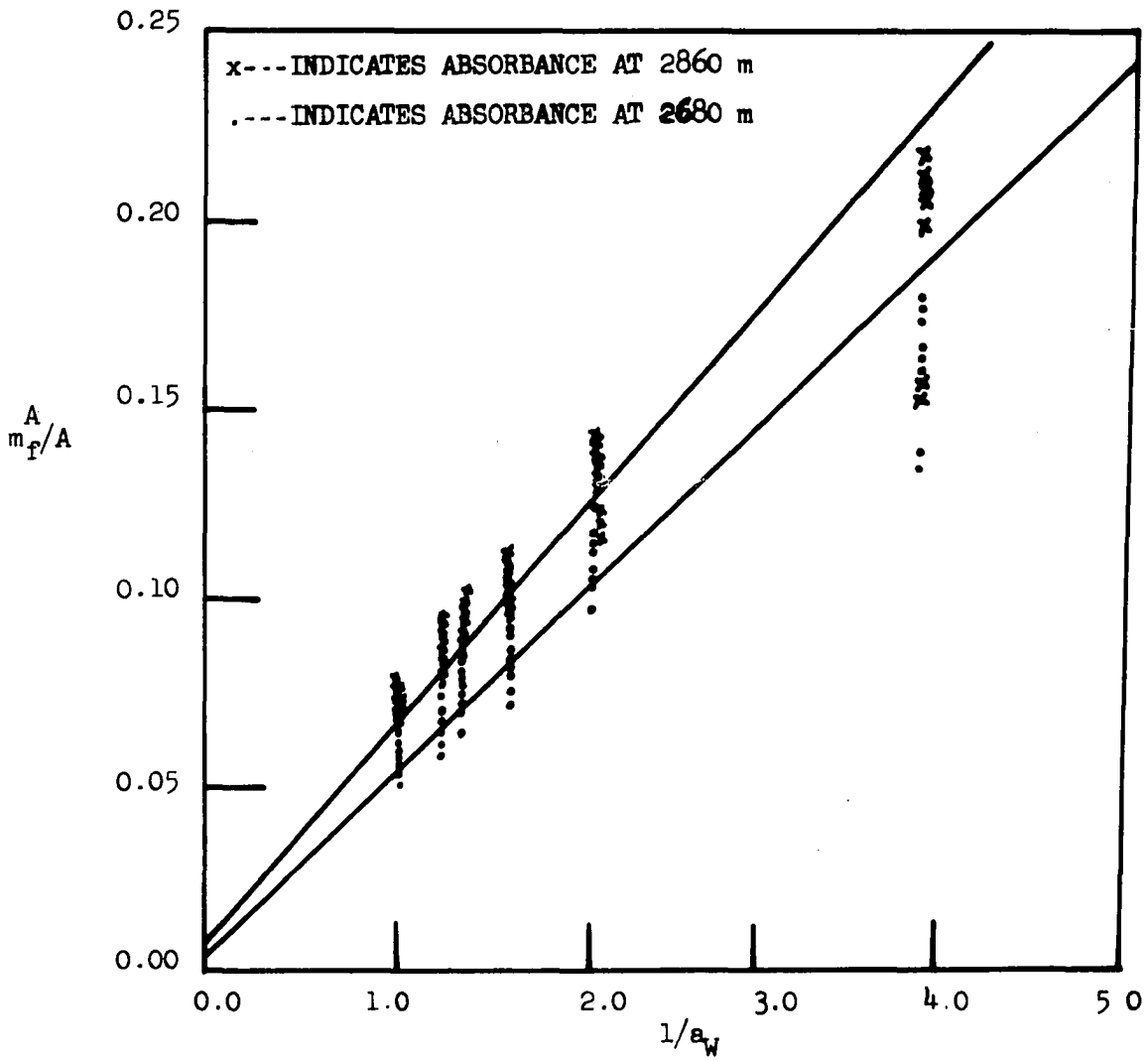


Figure 11. Relation of the Ratio of the Formal N-Methyl-2-pyrrolidone Concentration to the Absorbance at 2860 mμ and the Reciprocal of the Water Activity.

TABLE XIV

A summary of the thermodynamic constants at 25°C. for the lactams (a) 2-Pyrrolidone and N-Methyl-2-pyrrolidone, (b) δ-Valerolactam and (c) ϵ-Caprolactam

Reaction	K molal^{-1}	$-\Delta H^{\circ}$ kcal/mole	$-\Delta F^{\circ}$ kcal/mole	$-T\Delta S^{\circ}$ eu/mole
(a)				
$2M \rightleftharpoons D$	460 ± 11%	7.00	3.63	11.3
$M + W \rightleftharpoons M \cdot W$	15 ± 23%	----	1.60	----
$D + W \rightleftharpoons D \cdot W$	15 ± 17%	----	1.60	----
$M' + W \rightleftharpoons M' \cdot W$	10.9 ± 11%	----	1.47	----
(b)				
$2M \rightleftharpoons D$	432	10.3	----	----
(c)				
$2M \rightleftharpoons D$	168	5.46	----	----

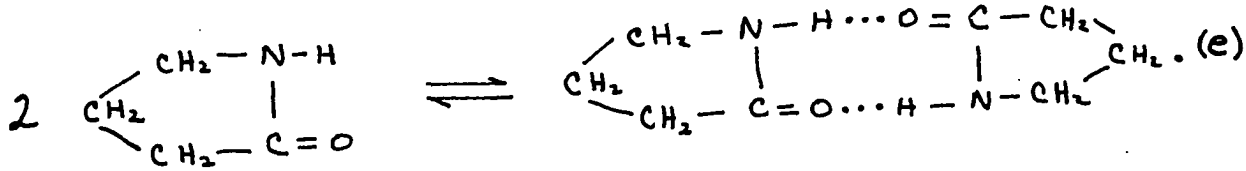
CHAPTER VI

DISCUSSION AND CONCLUSIONS

The cyclic lactam, 2-pyrrolidone, has proved to be an excellent hydrogen bonding solute for studying monomer-dimer equilibria in solution. The geometry of the 2-pyrrolidone structure is such that only cyclic polymers are likely to exist in dilute solution, and this makes the hydrogen bond equilibria much easier to study than in the systems involving non-cyclic amides, which tend to associate through linear chains of indefinite length. Both Tsuboi[16] and Mizushima[17] have established from spectral and dipole moment studies that a simple amide such as N-methylacetamide associates through linear polymers while the lactam (cyclic amide) δ-valerolactam forms only a cyclic polymer. Tsuboi has concluded from spectral investigations on δ-valerolactam solutions in carbon tetrachloride that the only associated species present in measurable concentrations is the cyclic dimer.[22] Lord and Porro have given a similar interpretation of the association of ϵ-caprolactam in carbon tetrachloride.[23]

One of the main objectives of this work was to study quantitatively the hydrogen bonding equilibria of 2-pyrrolidone in a solvent with minimal solvent-solute interactions under anhydrous conditions. This objective has been accomplished. The results of this investigation

are in accord with the assumption that 2-pyrrolidone associates in carbon tetrachloride according to the reaction



(e)

It is clear from the development of the theory in Chapter IV that the presence of measurable concentrations of cyclic species other than the dimer would have invalidated equations (2), (9) and (10). The linearity of the curves shown in Figures 5 and 6 lend strong support to the hypothesis that only a dimer is present. The fact that A_D/A_M^2 does not vary systematically with concentration can be taken as a strong indication that no non-cyclic forms are present. This, of course, leaves only a cyclic dimer.

Table XIV includes the thermodynamic constants calculated from spectral data for the dry system, as well as the results reported by Tsuboi[22] and Lord and Porro.[23] The results in Table XIV indicate that the thermodynamic constants are comparable in magnitude. ΔH_{Dim} values imply that the energy of a hydrogen bond between the amido hydrogen and the carbonyl oxygen is in the range of 3 to 5 kcal/hydrogen bond, which is consistent with enthalpy values reported for the association of other amides in carbon tetrachloride. Values of K_2 are uniformly greater than dimerization constants for most non-cyclic aliphatic amides in carbon tetrachloride.[35]

The association constant listed in Table XIV for 2-pyrrolidone

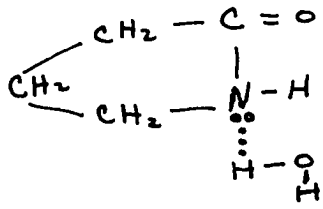
at 25°C. was taken from the solid line of Figure 6. The standard deviation of the points from the calculated line corresponds to an uncertainty in K_2 of about 11%. The strength of the hydrogen bond in the 2-pyrrolidone dimer was evaluated from the slope of the solid line in Figure 6 with an uncertainty of $\pm$ 1.5 - 2.0 kcal/mole.

A second major objective of this research was to make quantitative studies on the hydration of a peptide group. The compound 2-pyrrolidone was chosen for these studies since it contains a peptide linkage and since its behavior under anhydrous conditions had been studied and was well understood. Quantitative hydration studies were also made on the compound N-methyl-2-pyrrolidone because it shows no self association in solution and because of its structural resemblance to 2-pyrrolidone.

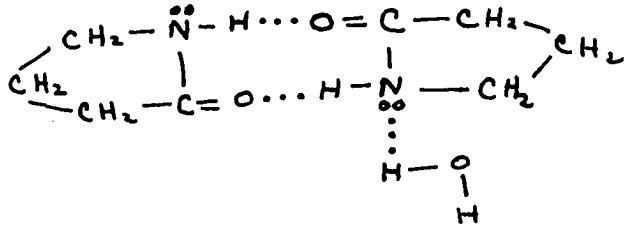
These investigations have confirmed the presence of hydrates for both the 2-pyrrolidone and the N-methyl-2-pyrrolidone. They further indicate that both monomer and dimer form monohydrates and that this hydration occurs primarily at the basic ring nitrogen. It is clear from a comparison of the spectra of N-methyl-2-pyrrolidone-water-carbon tetrachloride and 2-pyrrolidone-water-carbon tetrachloride that the amino hydrogen is relatively unimportant in the hydration process. On the basis of qualitative spectra it would appear that compounds containing the carbonyl group alone form relatively small concentrations of hydrate compared to compounds which contain a basic nitrogen atom. In the case of the 2-pyrrolidone dimer, hydration at the basic nitrogen would seem highly probable. Partition and water solubility studies of carboxylic acids in benzene have indicated that the acid dimer does

not hydrate to a significant degree.[25] Water solubility studies of 2-pyrrolidone in CCl_4 have shown that the dimer does hydrate, and from structural comparisons of the two dimeric species one can conclude that the presence of the basic nitrogens in the pyrrolidone dimer is the reason it can hydrate.

A plausible explanation of the appearance of two absorption bands due to the hydrated species may be made on the basis of the following structures:

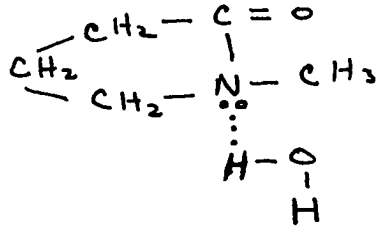


(f)



(g)

and



(h)

The band at 2680 mμ is sharp and only slightly shifted to longer wavelengths. It is proposed here that this band arises from the weak perturbation forces acting on the free or "dangling" (O-H) stretching vibrations in the hydrate. On the other hand the broad 2860 mμ peak is readily explained as arising from the vibrations of the hydrogen

bonded (N:...H-O) group. Diamond has given a similar interpretation to the spectra of various mineral acid hydrates in carbon tetrachloride.[36]

It is difficult to make any kind of argument to explain the relative values of the hydration constants obtained in this investigation. This is especially true in view of the uncertainty involved in each of them. The calculated values, $K_{M \cdot H_2O} = 15$ molal^{-1} and $K_{D \cdot H_2O} = 15$ molal^{-1}, for the 2-pyrrolidone monomer and dimer were obtained from a grid of standard deviations calculated from a computer least squares treatment of the data. These values of K gave the lowest standard deviation of the points from the line. The standard deviation was (4.496)(10-5). This method was used by Chii Lin in treating data of a similar nature.[37] From the appearance of the contour lines on the grid obtained in this work the $K_{M \cdot H_2O}$ value was found to have an uncertainty of ±23% while $K_{D \cdot H_2O}$ could be known to within ±17%.

The value of $K_{M' H_2O}$ for the N-methyl-2-pyrrolidone is in good agreement with the hydration constant for the 2-pyrrolidone monomer. This would be expected to be the case since each molecule has an equal number of hydration sites available and since their structures are so similar. Using $K_{M' H_2O} = 10.9$ molal^{-1} the standard deviation was approximately .01 which corresponds to an error of 6% in the measured absorbance value.

Since the hydration studies were carried out at only one temperature there was no way to evaluate the enthalpy. Hence the hydrogen bond strength for the bond (N:...H-O) could not be determined. Qualitatively one may argue that the magnitude of the frequency shift of the free water peak to that of the hydrogen bonded absorption peak indicates that

a hydrogen bond of comparable strength to that in the 2-pyrrolidone dimer has been formed. A 264 cm^{-1} shift is observed for hydrate formation, which a 245 cm^{-1} frequency shift is observed in the formation of the 2-pyrrolidone dimer. Although several attempts have been made to establish a quantitative relationship between ΔH and the observed frequency shift no successful relationship has been developed which will work for all types of hydrogen bonds. However, there is a considerable amount of evidence to show that such a relationship does exist for a series of like compounds.[38]

The use of equilibrators and equilibration techniques developed in this investigation presents a method of great promise for studying hydration phenomena and for obtaining quantitative data on the hydration process. Further refinement of the techniques to allow for temperature variations will allow the hydrogen bond strengths of the hydrates to be determined. The use of some solute other than water to equilibrate through the vapor phase is also promising. The compounds pyrrolidine and acetone are examples in point.

In conclusion it should be pointed out that unless more data are obtained on systems such as those studied here it is going to be very difficult indeed to learn anything about the very important role of water in biological systems. The availability of such information can be of considerable help in understanding hydration reactions and in establishing the structure of proteins.

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