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Dean
THE PARTIAL PRESSURE OF HYDROGEN CHLORIDE FROM ITS SOLUTIONS IN NON-AQUEOUS SOLVENTS AT 25° C.

III. ETHYLENE GLYCOL SOLUTIONS

BY

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A THESIS

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INTRODUCTION

The study of the partial pressure of hydrogen chloride from solutions in various non-aqueous solvents has already been started by Schmelzle\(^1\) and Westfall\(^2\). These investigators carried out their work using benzene and nitrobenzene as solvents. Their results show that solutions of hydrogen chloride in benzene and nitrobenzene agree with Henry's law. Also a method of calculating the equilibrium constants of hydrogen chloride in such solvents was devised.

Since then, Zellhoefer and coworkers\(^3\) have made an extensive study of the solubility of halogenated hydrocarbons (e.g. CHCl\(_2\)F and types such as CH\(_2\)X, CH\(_2\)X\(_2\), CHX\(_3\), CH\(_4\), C\(_2\)X\(_3\)) in a large variety of solvents at different pressures.

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Let us consider monofluordichlormethane whose partial pressure was studied from numerous non-aqueous solvents. When the experimental pressure of the solute equals that calculated from Raoult's law we have a solution which agrees with the law. If the experimental pressure is greater than this theoretical value we have a positive deviation from Raoult's law. Likewise, if the experimental pressure is less than this value, we have a negative deviation. A positive deviation indicates lower solubility and a negative deviation indicates higher solubility than the theoretical. In Table I are given a few of the results of Zeilhoefer and co-workers. 4

<table>
<thead>
<tr>
<th>Solvent</th>
<th>mol fraction</th>
<th>Raoult's Law mol fraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Glycol</td>
<td>0.05</td>
<td>0.381</td>
</tr>
<tr>
<td>Acetic Acid</td>
<td>0.286</td>
<td>&quot;</td>
</tr>
<tr>
<td>Aniline</td>
<td>0.385</td>
<td>&quot;</td>
</tr>
<tr>
<td>Nitro benzene</td>
<td>0.471</td>
<td>&quot;</td>
</tr>
<tr>
<td>Dioxane</td>
<td>0.479</td>
<td>&quot;</td>
</tr>
<tr>
<td>Dimethylaniline</td>
<td>0.695</td>
<td>&quot;</td>
</tr>
</tbody>
</table>

On examining Table I one notices that monofluorodichloromethane has a positive, nitrobenzene a negative, and aniline practically no deviation from the theoretical value of Raoult's law. The difference in solubility was explained in terms of solvation and hydrogen bond formation, indicated in the following manner,

\[ (O - H \leftrightarrow X) \]  

(1)

where \( X \) denotes the donor (of electrons) atom in the solvent molecule. Positive deviations from Raoult's law are explained in terms of association in the solvent, which offsets the hydrogen bond formation between solute and solvent. Negative deviations are explained by hydrogen bond formation offsetting any association effect. While the solubility of monofluorodichloromethane was determined at only three pressures the results seem to follow Henry's law quite closely, in all solvents.

That the nature of the solute is also an important factor is indicated by a number of facts: 1. Zellhoefer and coworkers found that the successive replacement of three of the hydrogen atoms of methane resulted in each case in an increase in solubility, and for the type \( \text{CHX}_3 \) the theoretical solubility in many solvents
was found to be greatly exceeded. This fact seems to indicate that the activity of the hydrogen ion has an appreciable effect on the solubility. In the scale of hydrogen ion activity the halogenated hydrocarbons and hydrogen chloride are at opposite ends. 2. That monofluorodichloromethane was found to be more soluble in glacial acetic acid than in glycol, while preliminary studies of hydrogen chloride in glycol indicates that it is more soluble in this solvent than in acetic acid. It seems worth while then to make a similar study using hydrogen chloride as the solute, instead of monofluorodichloromethane.

Another reason why this system should prove interesting is in connection with types of solvents. The available data for the partial pressure of hydrogen chloride in various solvents, shows that the solvents fall into two classes.

(1) Those in which hydrogen chloride deviates from both the laws of Raoult and of Henry.

e. Westfall, P. F., loc. cit.
(2) Those in which hydrogen chloride follows the law of Henry but deviates from the law of Raoult. The only solvent that has been studied in the first group is water. The second group includes benzene, nitrobenzene, and glacial acetic acid. Chesterman investigated the action of hydrogen chloride solutions on metals and has found that these solutions can be divided into three classes:

(1) Those in such solvents as water, alcohols, and ketones.

(2) Those in esters, ethers, and haloforms.

(3) Those in benzene, carbon disulfide, and carbon tetrachloride.

Since the first group of Chesterman corresponds to the first group above, it follows that ethylene glycol solutions of hydrogen chloride should give a pressure solubility relationship similar to that for water.

It will be the purpose of this paper then to present a study of the partial pressure of hydrogen chloride from solutions of hydrogen chloride in ethylene glycol which was undertaken in order to add to the data of this kind; to determine whether glycol

solutions of hydrogen chloride behave like aqueous solutions; and to attempt to evaluate the equilibrium constant for the hydrogen chloride-solvent-base reaction.

EXPERIMENTAL

Materials

Ethylene glycol—Ethylene glycol used was an Eastman Kodak Company product. Its index of refraction at 25°C using sodium light was 1.4301, which indicates that it contained a trace of water.

Glycol hydrogen chloride solutions—The glycol solutions of hydrogen chloride were prepared by saturating the glycol with hydrogen chloride gas generated by dropping C. P. quality, 95%, sulfuric acid on analytical reagent quality potassium chloride.

Standard Solutions—Constant boiling hydrochloric acid was used as the primary standard. Acid solutions of normality suitable for titrating the hydrogen chloride glycol solutions of various concentra-

tions were prepared by diluting the constant boiling acid. The carbonate free sodium hydroxide solutions of corresponding normalities were compared with the acid solutions. Methyl red was the indicator used.

**Apparatus.**—The apparatus used was the same as that used by Schmelzle\(^8\) and Westfall.\(^9\) A diagram of it is shown in Figure I. The two bulbs \(C\) and \(D\) are connected by means of a mercury sealed stopcock \(A\). All plugs and stopcocks were lubricated with "Cello Grease", which was found not to be appreciably affected by the glycol solutions (see below). Two apparatus of this kind were used. The volumes of the bulbs "D" were found by determining the weight of water necessary to fill them at \(25^\circ\). The volume of the bulb for apparatus No. 1 was 150.5±.1 milliliters. The volume of the bulb for Apparatus No. 2 was found to be 150.3±.1 milliliters.

**Procedure.**—The central stopcock (\(A\) in Figure I) was closed and the glycol solutions poured into bulb \(C\). The ground glass stoppers were inserted, sealed with mercury, while the bulb \(D\) was evacuated. The stopcock

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FIGURE I: APPARATUS FOR MEASURING VAPOR PRESSURE
at E was closed, and a rubber plug placed over the open end of the tube at E in order to prevent contamination of the tip by the water in the constant temperature bath. The stopcock at A was opened and the apparatus placed in the bath controlled at 25.00 ± 0.02°. After equilibrium had been established, the apparatus was removed from the constant temperature bath and stopcock A was closed immediately. The plug at E was removed, the end of the tube placed in a beaker of freshly distilled water, and the stopcock B opened to allow about 25 milliliters of water to be drawn into the bulb to absorb the vapor. The water was permitted to stand in the bulb about fifteen minutes. The solution was then washed into a beaker and the hydrogen chloride titrated with standardized sodium hydroxide. The partial pressure of the hydrogen chloride was calculated using the gas law equation

\[ pv = nRT \]  

where n is the number of moles of hydrogen chloride found in the vapor; R the molal gas constant in liter-millimeters (62.37); T the absolute temperature (298.1° K); and v the volume of the bulb D. The simple gas law equation may be used without introducing appreciable
error, since it has been pointed out that hydrogen chloride deviated from the simple gas laws by less than one percent.\textsuperscript{10}

A sample of the glycol solution was obtained by removing the two ground glass stoppers from the bulb \(C\) and pouring out the solution through a tube fitted with a ground glass joint. One to three gram samples were collected in tightly stoppered, weighing bottles, quickly reweighed, and opened under the surface of water. The amount of hydrogen chloride in the sample was found by titrating with a sodium hydroxide solution, and the molalities of the solutions were calculated from these results.

Accuracy of method.—Preliminary experiments were also performed for the purpose of determining the accuracy of the method and to establish the time required for the attainment of equilibrium. Saylor\textsuperscript{11} used this method in determining the partial pressure of hydrogen chloride from benzene solutions at \(30^\circ\) and arbitrarily chose five to seven days for the establishment of equilibrium. Schmelzle\textsuperscript{12} was able to duplicate the

\begin{itemize}
  \item \textsuperscript{10} Gray and Burt, \textit{J. Chem. Soc.}, \textbf{95} 1633 (1909).
  \item \textsuperscript{11} Saylor, John H., \textit{loc. cit.}
  \item \textsuperscript{12} Schmelzle, A. F., \textit{loc. cit.}
\end{itemize}
measurements of Saylor allowing two days for the systems to reach equilibrium. Velehradsky\textsuperscript{13} used the above apparatus for measuring the vapor pressure of glacial acetic acid at $25^\circ$. He was able to obtain values concordant with those in the literature allowing only one day. When longer periods of time were taken, the method gave inconsistent values. The inconsistency of these results were attributed to an attack on the stopcock grease (Cello Grease) by the acetic acid, which made it impossible to separate the liquid phase from the vapor phase by means of stopcock $A$ while the vapor was being rinsed from bulb $D$. In order then to obtain a better check on the method, the partial pressure of aqueous constant boiling hydrochloric acid (7.0 molal) was measured allowing varying periods of time for establishment of equilibrium. The results obtained are given in Table II.

\textsuperscript{13} Velehradsky, E., \textit{Bachelor's Thesis}, Creighton University (1939).
TABLE II

Partial Pressure of Hydrogen Chloride From Aqueous Constant Boiling Hydrochloric Acid

<table>
<thead>
<tr>
<th>Time (hrs.)</th>
<th>Pressure HCl (mm.)</th>
</tr>
</thead>
<tbody>
<tr>
<td>23</td>
<td>0.40</td>
</tr>
<tr>
<td>25</td>
<td>0.40</td>
</tr>
<tr>
<td>46</td>
<td>0.27</td>
</tr>
<tr>
<td>46</td>
<td>0.37</td>
</tr>
<tr>
<td>48</td>
<td>0.49</td>
</tr>
<tr>
<td>70</td>
<td>0.40</td>
</tr>
<tr>
<td>196</td>
<td>0.42</td>
</tr>
</tbody>
</table>

mean 0.39 ± 0.07

The values of the partial pressure of 7 molal hydrochloric acid interpolated from the data of Randall and Young\textsuperscript{14} is 0.35 mm. From these results it may be concluded that by means of this method, vapor pressures may be determined with an accuracy of 0.1 mm. and that even one day is sufficient for establishment of equilibrium.

In measuring the partial pressure of hydrogen chloride from glycol solutions, however, two days were

\textsuperscript{14} Randall, Merle, and Young, Leona E., loc. cit.
allowed. Blank experiments were also performed to determine the efficacy of the stopcock lubricant (Cello Grease). The procedure for these experiments was essentially that described above except that stopcock A remained closed while the apparatus was in the constant temperature bath for periods up to two days. By means of these experiments it was found that "Cello Grease" was suitable for this system and also for hydrogen chloride benzene solutions.

RESULTS

The results obtained are given in Table III, and are shown graphically in Figure II and Figure III. In the table the values in parentheses are obtained from a large scale plot. In Figure II the molality is plotted against the pressure. The curve shows great similarity to the curve which is obtained for hydrogen chloride solutions in water. The curve for water solutions has the same shape as the curve for glycol, but is displaced to the right. Had the solubility of hydrogen chloride in glycol corresponded to Henry's law, the curve would have been a straight line, cutting the origin.

In Figure III have been plotted the logarithm
### TABLE III

The Partial Pressure of Hydrogen Chloride From Ethylene Glycol and The Equilibrium Constant of The Acid-Solvent-Base Reaction

<table>
<thead>
<tr>
<th>m</th>
<th>p (atm.)</th>
<th>kV</th>
<th>k²K</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.333</td>
<td>0.0008</td>
<td>0.02122</td>
<td>2.898</td>
</tr>
<tr>
<td>(2.000)</td>
<td>(0.0019)</td>
<td>(0.02185)</td>
<td>(2.898)</td>
</tr>
<tr>
<td>2.269</td>
<td>0.0025</td>
<td>0.02204</td>
<td>2.899</td>
</tr>
<tr>
<td>2.453</td>
<td>0.0032</td>
<td>0.02306</td>
<td>2.898</td>
</tr>
<tr>
<td>2.721</td>
<td>0.0046</td>
<td>0.02498</td>
<td>2.910</td>
</tr>
<tr>
<td>3.023</td>
<td>0.0074</td>
<td>0.02846</td>
<td>2.898</td>
</tr>
<tr>
<td>3.070</td>
<td>0.0075</td>
<td>0.02820</td>
<td>2.896</td>
</tr>
<tr>
<td>3.112</td>
<td>0.0086</td>
<td>0.02980</td>
<td>2.898</td>
</tr>
<tr>
<td>3.128</td>
<td>0.0079</td>
<td>0.02842</td>
<td>2.899</td>
</tr>
<tr>
<td>(4.000)</td>
<td>(0.0210)</td>
<td>(0.03623)</td>
<td>(2.898)</td>
</tr>
<tr>
<td>4.658</td>
<td>0.0357</td>
<td>0.04056</td>
<td>2.897</td>
</tr>
<tr>
<td>6.572</td>
<td>0.1347</td>
<td>0.05584</td>
<td>2.891</td>
</tr>
<tr>
<td>6.630</td>
<td>0.1386</td>
<td>0.05622</td>
<td>2.905</td>
</tr>
<tr>
<td>6.924</td>
<td>0.1716</td>
<td>0.05983</td>
<td>2.898</td>
</tr>
<tr>
<td>8.780</td>
<td>0.4241</td>
<td>0.07417</td>
<td>2.898</td>
</tr>
</tbody>
</table>
FIGURE II: THE PARTIAL PRESSURE OF HYDROGEN CHLORIDE FROM GLYCOL SOLUTIONS
Figure III: The partial pressure of hydrogen chloride from solutions in: (1) benzene (2) nitrobenzene (3) acetic acid (4) ethylene glycol.
of the pressure against the logarithm of the molality for hydrogen chloride solutions in benzene (1), nitrobenzene (2), glacial acetic acid (3), and glycol (4). The curves for benzene, nitrobenzene, and acetic acid, all have a slope of one and consequently correspond to Henry's law. The curve for glycol definitely does not have a slope of one and is not a straight line, at least at lower concentrations.

DISCUSSION

Calculation of activity coefficients.— For the calculation of activity coefficients the method of Lewis and Randall\(^\text{15}\) was employed. The calculations involved are as follows:

\[ a \propto p \]  

Equation (3) simply states that the activity is proportional to the pressure, and it may be changed to an equality by the insertion of a proportionality constant, which for convenience is selected as \( k^2 \). We obtain then

\[ k^2 a_2 = p \]  

\(^{15}\) Lewis and Randall, *Thermodynamics and The Free Energy of Chemical Substances*, (1923).
By definition the activity of the reacted hydrogen chloride is given by

$$a_2 = m y_+ m y_-$$  \hspace{1cm} (5)

or

$$a_2 = (y_x m)^2$$  \hspace{1cm} (6)

where \(a\) represents the activity of the reacted hydrogen chloride. By substituting Equation (6) into Equation (4) we obtain

$$k^2 y^2 m^2 = p$$  \hspace{1cm} (7)

or

$$k y' = \frac{p^{\frac{1}{2}}}{m}$$  \hspace{1cm} (8)

It will be noticed that instead of arriving at a value for the activity coefficient, the activity coefficient times a proportionality constant is obtained by this procedure. The value of this constant could not be derived from the data in this paper and there was no data available in the literature which could be used. Consequently, for the present, we must be satisfied with the quantity \(k y'\), rather than the activity coefficient.
Calculation of the dissociation constant for hydrogen chloride in glycol. It seems there are two possible reactions between hydrogen chloride and a solvent. The first reaction results in the formation of a solvated proton and a chloride ion, which is represented by the equation

\[ \text{HCl} + \text{S} = \text{HS}^+ + \text{Cl}^- \]  
(9)

where S is the solvent and HS is the solvated proton. The second reaction involves the formation of a hydrogen bonded compound, or a stable ion pair, which is represented by the equation

\[ \text{HCl} - \text{S} = \text{(H-S-Cl)} \]  
(10)

where S is again the solvent, and S-H-Cl represents the hydrogen bonded compound or the stable ion pair.

The equilibrium constant for the first reaction is calculated as follows. By definition

\[ K = \frac{a_{\text{HS}^+} \cdot a_{\text{Cl}^-}}{a_{\text{HCl}}} \]  
(11)

where \( a_{\text{HS}^+} \) is the activity of the solvated proton, \( a_{\text{Cl}^-} \) the activity of the chloride ion, and \( a_{\text{HCl}} \) the activity of the unreacted hydrogen chloride. We assume
that is the activity of the unreacted hydrogen chloride equals its molality, since hydrogen chloride is a neutral molecule and also since its concentration is very low. By making use of the law of Raoult we arrive at the equation \[ \text{m}_{\text{HCl}} = \frac{\text{m}_0 \cdot \rho}{\rho_0} \] (13)

where \( m_0 \) is the number of mols of glycol in 1000 grams of solvent; \( \rho \) the partial pressure of hydrogen chloride from solution; and \( \rho_0 \) the vapor pressure of liquid hydrogen chloride at 25°, (46.8 atm.). \[ K = \frac{(\delta m)^2}{m_0 \cdot \rho / \rho_0} \] (14)

Here the concentration of the ions has been taken as equal to the concentration of the hydrogen chloride solution, since the concentration of unreacted hydro-

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16. Since the mol fraction of unreacted hydrogen chloride is very small in comparison with unity, it is neglected (see Equation (18) below).

gen chloride is very low, and the reacted hydrogen chloride assumed to be completely dissociated. This equation is the one which fits the data for water and should be applicable to our data for glycol. However, as mentioned, we are forced to be satisfied with the quantity $k / \gamma$ rather than the activity coefficient. If Equation (14) is multiplied by $k^2$ it becomes

$$k^2K = \frac{(k \gamma m)^2 \rho}{m_o \rho}$$

that is, a constant is obtained which is proportional to the equilibrium constant and should make it possible to tell whether as expected the glycol is similar to water solutions of hydrogen chloride.

The equilibrium constant for the reaction represented in Equation (10) is calculated as follows. By definition

$$K = \frac{a_{SHCl}}{a_{HCl}}$$

where $a_{SHCl}$ is the activity of the hydrogen bonded compound or the stable ion pair, and $a_{HCl}$ is the activity of the unreacted hydrogen chloride. We assume

$$a_{SHCl} = m_{SHCl}$$
since we have no knowledge of its activity coefficients in the various solvents.

The molality of unreacted hydrogen chloride is obtained using Raoult's law,

\[
m = m_s \frac{p_0}{p_0 - p - m_{HCl}}
\]

(18)

and the molality of reacted hydrogen chloride is then given by

\[
m_{HCl} = m - m_{HCl}
\]

(19)

where \( m \) is the total molality of hydrogen chloride in the solution. Substitution of equations (12), (17), (18), (19) into equation (16) and algebraic rearrangement gives

\[
K = \frac{m (p_0 - p) - m_{HCl} p}{m_{HCl} p}
\]

(20)

This is the equation that fits the data for benzene, nitrobenzene and also glacial acetic acid. However, it does not fit the data for water and glycol solutions of hydrogen chloride.

As previously stated, the activity coefficient of hydrogen chloride in glycol could not be found. However, an attempt was made to determine the activity
coefficient from theoretical considerations. The limiting Debye-Hückel theory was employed. It may be written

$$\log \gamma = \frac{A \sqrt{m}}{D^{3/2}}$$  \hspace{1cm} (21)

where $\gamma$ is the activity coefficient; $A$ is a constant for a 1:1 electrolyte at a given temperature; $m$ is the molal concentration. We may write then

$$\log \gamma_w = \frac{A \sqrt{m}}{D_w^{3/2}}$$  \hspace{1cm} (22)

where $\gamma_w$ is the activity coefficient of hydrogen chloride in water; $D_w$ is the dielectric constant. We may also write

$$\log \gamma_g = \frac{A \sqrt{m}}{D_g^{3/2}}$$  \hspace{1cm} (23)

where $\gamma_g$ is the activity coefficient of hydrogen chloride in glycol; $D_g$ is the dielectric constant. We may then obtain

$$\log \gamma_g = \frac{D_w^{3/2} \log \gamma_w}{D_g^{3/2}}$$  \hspace{1cm} (24)

by dividing Equation (22) by Equation (23) and solving for $\log \gamma_g$. 
By means of Equation (24) the activity coefficient of hydrogen chloride in glycol in a 2 molal solution was calculated using the value of Randall and Young for the activity coefficient of hydrogen chloride in a 2 molal water solution. Then from this value and the value for the quantity $k$ obtained for the 2 molal hydrogen chloride glycol solution, the constant small $k$, was evaluated. Then using this value the equilibrium constant for hydrogen chloride in glycol was obtained from the average of the values for $k^2K$ given in Table III. These calculations are summarized in Table IV.

**Table IV**

| Data Used For The Calculation of Dissociation Constant of Hydrogen Chloride In Glycol |
|---------------------------------|--------------|
| $m$                             | 2.0          |
| $\gamma_w$                     | 1.032        |
| $\gamma_g$                     | 1.100        |
| $k'$                           | 0.02185      |
| $k$                            | 0.020        |
| $k^2K$                         | 2.898        |
| $K$                            | $7.245 \times 10^3$ |
In Table V are listed the values for the dissociation constants of hydrogen chloride in various solvents. The constants for hydrogen chloride solutions in benzene, nitrobenzene, and glacial acetic acid were calculated using Equation (20). The constants for hydrogen chloride in water and glycol were calculated using Equation (15), and Equation (14).

### TABLE V

<table>
<thead>
<tr>
<th>Solvent</th>
<th>K</th>
</tr>
</thead>
<tbody>
<tr>
<td>Benzene</td>
<td>0.8</td>
</tr>
<tr>
<td>Nitrobenzene</td>
<td>2.0</td>
</tr>
<tr>
<td>Glacial Acetic Acid</td>
<td>7.1</td>
</tr>
<tr>
<td>Ethylene Glycol</td>
<td>$7.24 \times 10^3$</td>
</tr>
<tr>
<td>Water</td>
<td>$1.86 \times 10^7$</td>
</tr>
</tbody>
</table>

### SUMMARY AND CONCLUSIONS

1. The partial pressure of hydrogen chloride from its solutions in ethylene glycol at 25° has been determined over the concentration range 1.333 to 8.780 molal.
2. The results obtained are, as expected, similar to those obtained from water solutions in that they deviate from both the laws of Henry and Raoult.

3. Approximate values for the activity coefficients of hydrogen chloride in ethylene glycol at 25° are obtainable from the data in this paper.

4. Making use of the approximate values for the activity coefficient of hydrogen chloride in ethylene glycol its equilibrium constant for the acid-solvent-base reaction was calculated. A constant value is obtained using Equation (14) which corresponds to the equilibrium constant for a reaction represented by the Equation

\[ \text{HCl} + \text{S} = \text{HS}^+ + \text{Cl}^- \]

Here again glycol solutions correspond to water solutions, in which solvent the equilibrium constant is also obtainable from Equation (14). This seems to indicate that the hydrogen chloride which is reacted with the solvent is completely ionized in both water and ethylene glycol. Both glycol and water solutions of hydrogen chloride differ from its solutions in benzene, nitrobenzene and glacial acetic acid for which the equilibrium constant is obtained using Equation (20)
which corresponds to the reaction

\[ \text{HCl} - S = S\text{-H-Cl} \]

where \( S^H\text{Cl} \) may represent either a hydrogen bonded compound or a stable ion pair. This seems to be a confirmation of the hypothesis that there are two kinds of acid-solvent reactions, and it might be concluded that the first step in the reaction is the formation of a hydrogen bonded compound. Then if the force between the proton and the solvent molecule is stronger than that between the proton and the acid anion, the resulting compound will be completely dissociated into its ions providing the dielectric constant is high enough. If the dielectric constant is not high enough, the substance may be dissociated but the ions form stable pairs. This may be the situation that exists in benzene, nitrobenzene and glacial acetic acid.  

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If on the other hand the bond between the solvent and the proton is weaker than that between the solvent and the acid anion, then only a slight reaction takes place between the acid and the solvent, and no dissociation takes place regardless of the dielectric constant of the medium, as illustrated in the case of hydrogen chloride in nitrobenzene which has a high dielectric constant but in which there is only a slight reaction and one in which ionization does not take place.


Kraus and Hooper, "Dielectric Constant of Some Solutions of Electrolytes in Benzene." *Journal of the American Chemical Society*, XLVI, 2265-68 (1934).


