

Critical Phenomena in Ionic Fluids

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The properties in the critical region, including those of the coexisting phases, have been widely investigated for neutral-molecule fluids. For ionic fluids, the investigation of these properties has been very difficult because, in most cases, the critical points occur at very high temperatures. This report summarizes several recent studies of ionic fluids, both experimental and theoretical, including the discovery of a model ionic fluid with a critical point at 44 °C and the measurement of its properties.

My focus is on fluids where the electrostatic forces are the dominant attractive interactions related to the phase separation. Such fluids include simple salts, such as NaCl, above their melting points, together with certain two-component systems where the ionic interactions are the primary factor in the phase separation. There are also fluids where ions are present but are incidental to the phase separation; examples of this sort will be mentioned but not discussed in detail.

In addition to knowledge of various physical and structural properties of the phases in equilibrium, there is particular interest in the exact shape of the coexistence curve as the critical point is approached. This is usually expressed by the critical exponent β defined by

$$(\rho_1 - \rho_v) = B(T_c - T)^\beta \quad (1)$$

where ρ_1 and ρ_v are the liquid and vapor densities, respectively, and B is a constant. This expression is exact in the limit as T approaches the critical temperature T_c ; it may also be a good approximation over a range of temperature. For liquid-liquid phase separation in two-component systems, the difference in composition takes the place of the difference in density in eq 1. The value of the exponent β is of primary theoretical interest, as will be discussed below.

In preparation for consideration of near-critical properties, I first present in the next section the pertinent information about the phase diagram and concerning the ionic vapor and the ionic liquid derived by theory for a simple model and by experiment for the simple salt NaCl. This section includes extrapolations to the critical point for the simple model and for NaCl, but there is considerable uncertainty in these extrapolations. Next, I consider the few examples of ionic fluids where experimental measurements extend close to the critical temperature and define the shape of the coexistence curve. In only one of these cases was it feasible to attain high precision very close to T_c ; the results for

that system are described in detail. The available theory for near-critical properties is compared with these experimental results.

Phase Diagram Vapor and Liquid Properties

The phase diagram on a corresponding states basis is qualitatively the same for an ionic as for a neutral fluid. Figure 1 shows this pattern with the reduced variables $T_r = T/T_c$ and $V_r = V/V_c$. The hard-sphere model with diameter a and charge Ze is a useful approximation. Then a consideration of distance and energy scaling factors yields the following relationships for the critical properties of different fluids:

$$V_c'' = V_c'(a''/a)^\delta \quad (2)$$

$$T_c'' = T_c'(Z''/Z)^2(\epsilon'a'/\epsilon''a'') \quad (3)$$

In eq 3 the relative permittivity or dielectric constant ϵ is introduced for the solvent in ionic solutions; for pure ionic fluids $\epsilon = 1$. The approximate locations for familiar aqueous solutions are indicated on Figure 1. The large value of ϵ places simple aqueous salt solutions at a very high reduced temperature where ion dissociation is large and Debye-Huckel theory is a good approximation. At lower reduced temperatures the ions are largely paired in dilute systems, while at high concentration a typical fused-salt pattern develops with several ions of the opposite sign around a given ion. The phase diagram of the ionic fluid has been discussed by Friedman and Larsen¹ and in my more general paper on ionic fluids.² Note that, for aqueous systems, ϵ decreases so rapidly with increase in T that a lower critical point is expected for 2-2 electrolytes at high T .

For precise discussions of the primitive (hard-sphere) model, it is useful to define dimensionless concentration and temperature as follows:

$$c^* = ca^3 \quad (4)$$

$$T^* = T(4\pi\epsilon_0\epsilon ak/Z^2e^2) \quad (5)$$

In eq 5, k is the Boltzmann constant and ϵ_0 is the permittivity of free space. Much of the literature on ionic fluids is written in terms of electrostatic units, whereupon the $4\pi\epsilon_0$ factor disappears. Also, in many papers the reciprocal of T^* is used with the symbol β^* ; I have chosen to use T^* since it is easier to understand and possible confusion of β^* with the unrelated critical exponent β of eq 1 is avoided. Theoretical values for c^* and T^* at the critical point will be discussed below, following separate considerations of the vapor and the liquid.

Ionic Vapor. It is well-known that the saturated vapor of NaCl and other alkali halides comprises primarily ion-pair molecules, together with a substantial portion of dimeric M_2X_2 molecules with a ring structure. The experimental measurements extend to about 1700

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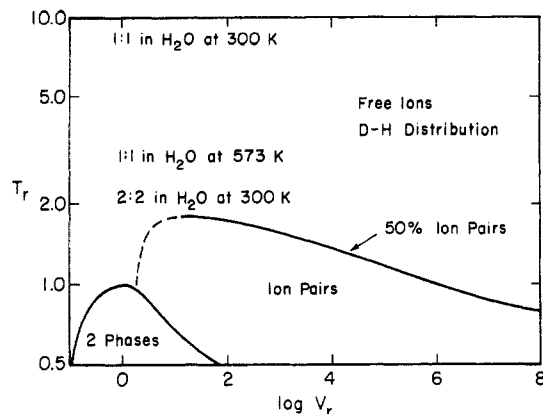


Figure 1. The phase boundaries and critical point for an ionic system on a corresponding states basis. Also shown are the approximate locations for familiar aqueous solutions and the boundary for 50% ion pairing.

K, but the molecular properties of the monomer (ion pair) are known in such detail that accurate calculations can be made for higher temperatures.³ For the ring M_2X_2 , the experimental knowledge of molecular properties is limited and approximate; also, one expects open-chain M_2X_2 and larger clusters to become significant at higher temperatures. Hence, extrapolation of total vapor properties for NaCl at higher temperatures is uncertain without further theoretical guidance.

In 1983, Gillan⁴ and Tani and Henderson⁵ independently presented theories for the ionic vapor based on equilibria among clusters. Gillan considered clusters up to M_3X_3 but under ideal-gas conditions, while Tani and Henderson limited themselves to ion pairs and triplets but included the effects of electrostatic interactions and excluded volumes among clusters. Subsequently, Schreiber and I⁶ combined the advantageous features of each of these treatments.

A cluster (s,t) comprises s positive ions and t negative ions that are linked either directly or indirectly by separations of distance less than $2a$, i.e., twice the hard-core diameter. The properties of the $(1,1)$ ion pairs were derived by Bjerrum,⁷ who obtained a closed expression in terms of the exponential integral. Gillan⁴ used Monte Carlo (MC) methods to calculate the properties of the larger clusters, both neutral and singly charged, up to $(3,3)$, $(3,2)$, and $(2,3)$. Following Tani and Henderson,⁵ the mean spherical approximation was used to estimate interactions between clusters in the vapor, while standard statistical mechanics yielded the equilibrium distribution of clusters.

Figure 2 shows the calculated distributions of particles among clusters for two near-critical temperatures.⁶ The dominance of the $(1,1)$ cluster for concentrations from 10^{-5} to 10^{-2} is clear. The population of single ions decreases rapidly with decrease in temperature. At concentrations above 0.01, the populations of large clusters, both neutral and charged, increase rapidly as the general pattern shifts to that of an expanded ionic liquid. The approximations of this cluster calculation limit its validity to concentrations below about 0.02. We

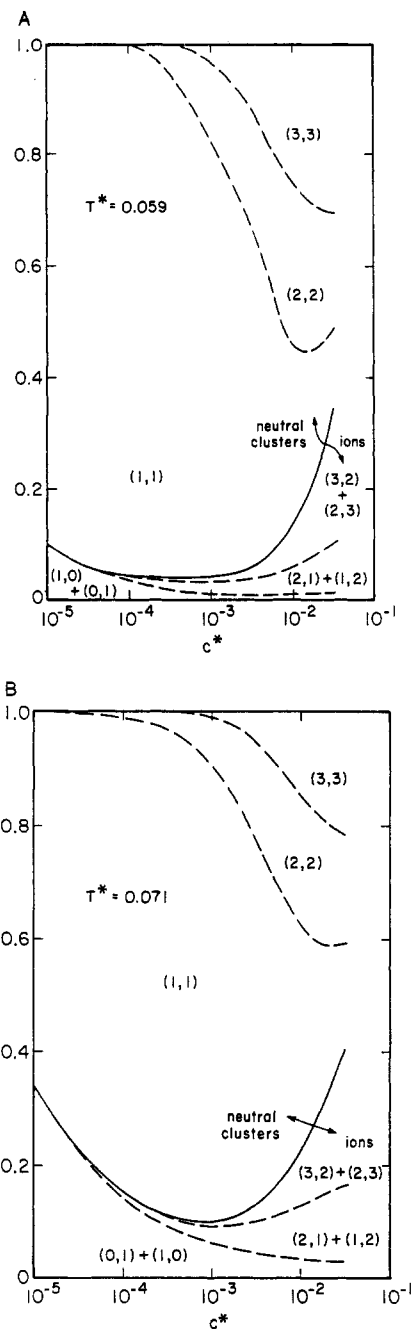


Figure 2. The structure of ionic vapor at two near-critical temperatures, as indicated by the distribution of particles among clusters at $T^* = 0.059$ (A) and at $T^* = 0.071$ (B), as a function of reduced concentration. The vertical distance between lines gives the fraction of particles in the indicated cluster. The charged clusters are at the bottom, and the solid line gives the division between the totals for charged and for uncharged clusters.

will see that this limit is below the critical concentration; hence, estimates of critical properties can only be obtained by interpolation to the properties of the ionic liquid.

Ionic Liquid. the density of liquid NaCl was measured over the range 1100–1700 K by Kirshenbaum et al.,⁸ who give the equation

$$d = 2.061 (1 - T/4330.7) \text{ g}\cdot\text{cm}^{-3} \quad (6)$$

This equation implies a surprisingly large and constant rate of decrease of density with temperature. This

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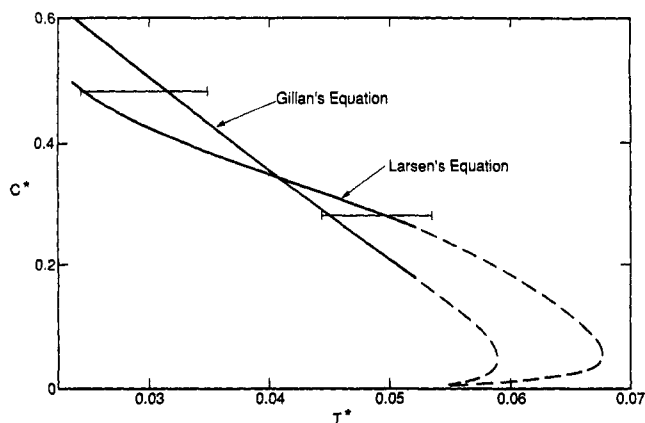


Figure 3. Approximate curves for the concentration of the saturated ionic liquid from equations of Gillan⁴ and Larsen,¹¹ both based on Monte Carlo calculations by Larsen.¹¹ Also shown are the estimate for the vapor density and possible extrapolations through a critical region. See text for further details.

pattern differs from that of neutral fluids which show a much smaller rate of decrease at low temperature, but a rate increasing with temperature. Thermodynamic properties have been calculated for liquid NaCl on the basis of various measurements to 1300 K and the assumption of constant heat capacity at higher temperatures.³

For the liquid primitive model, the primary information is an array of Monte Carlo calculations by Larsen.⁹ These were made on a fixed-volume basis and do not yield directly the properties at saturation pressure; also, there are some problems of nonergodicity.¹⁰ Thus, the total uncertainty in the results is not a simple statistical matter but must be estimated with several aspects in mind. Figure 3 shows for the two most pertinent values of c^* at which calculations were made, 0.286 and 0.479, the ranges of temperature within which the saturated liquid curve should lie. Also shown are the curves from the two equations fitted to the MC data. The first is a complex general equation of state from Larsen,¹¹ which, however, is known to be in error at vapor densities because it fails to recognize the strong ion pairing. The second equation is one fitted by Gillan⁴ specifically to the MC points nearest to saturation conditions; it is simple in form as follows:

$$c^* = 0.951(1 - 15.5_6 T^*) \quad (7)$$

One notes that this linear dependence on T^* is the same as the linear dependence on T of eq 6 for the experimental density of liquid NaCl.

Critical Point for the Ionic Fluid. Figure 3 also shows the probable curves connecting the vapor density through a critical point to the liquid density for each of the two equations for the liquid. There is, of course, some uncertainty in the exact shape near the critical point; this is discussed in detail in ref 6 for the curve connecting to the Larsen equation where this uncertainty is large. But it is clear that the major uncertainty is in the curve for the saturated liquid below 0.05 in T^* . It seems probable that the true critical point for the primitive ionic fluid lies between the two locations shown in Figure 3; hence, we conclude that c_c^* is near 0.05 and T_c^* probably lies between 0.059 and 0.068.

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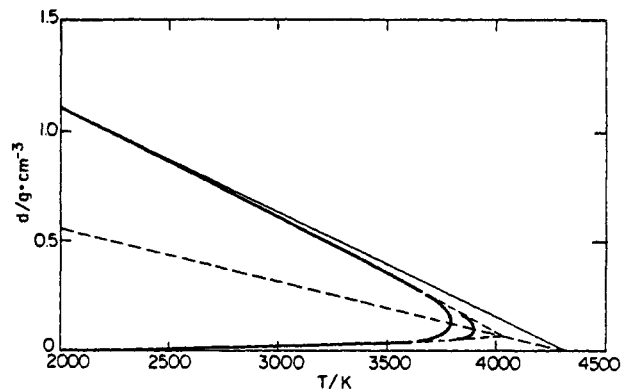


Figure 4. An extrapolation of the liquid and vapor densities of NaCl to higher temperature with alternate estimates for the critical point.

Figure 4 shows a similar extrapolation for NaCl, which I reported earlier.² It is based on eq 6 for the liquid, together with calculated values for the vapor using experimental molecular properties for the dimer and Gillan's values for higher clusters.⁴ The more recent cluster calculations⁶ increase slightly the estimates for vapor density and favor the lower among the alternate critical temperatures indicated in Figure 4. Thus, I now recommend as probable values $T_c = 3800 \pm 200$ K and $d_c = 0.14 \pm 0.02$ g·cm⁻³, although the absolute uncertainties are certainly larger.

Experimental Near-Critical Data for Ionic Fluids

Since direct experiments have not been feasible at the very high critical temperatures of simple ionic fluids such as NaCl, we turn to somewhat more complex fluids with lower critical temperatures which should model most properties of simple ionic fluids.

Ammonium Chloride. NH₄Cl was shown by Bubach and Franck¹² to have a critical point at 1155 K. They showed that the liquid was ionic, with properties very similar to those of KCl. The vapor, however, is primarily NH₃ + HCl rather than NH₄Cl ion-pair molecules. This difference in the vapor lowers the critical temperature and raises the critical pressure greatly as compared to the values expected for a simple fluid of NH₄⁺ and Cl⁻ ions. It is interesting to note that the liquid and vapor densities of NH₄Cl fit eq 1 with $\beta = 1/2$. But one must be cautious in attributing this or any other property of NH₄Cl to ionic fluids more generally.

Two-Component Ionic Fluids. From eq 3 one notes the effects of change of Z , ϵ , or a on the critical temperature. This relationship considers only the electrostatic effects; other types of interparticle forces may be present and may even dominate a phase separation. Thus, an aqueous system with a little NaCl shows a vapor-liquid phase equilibrium and a critical point determined primarily by the properties of H₂O. As the fraction of NaCl increases, the role of the electrostatic effects increases and the critical line for concentrated NaCl-H₂O is consistent with the simple ionic-fluid model.² At intermediate compositions, an equation based on pure water properties with perturbation terms for the effect of NaCl of KCl is effective.^{13,14}

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Liquid-liquid phase separation and a critical point have been reported recently for two rather similar aqueous systems involving large organic ions. Specifically, the salts are tetra-*n*-butylammonium thiocyanate¹⁵ and tetra-*n*-pentylammonium bromide,¹⁶ with critical points of 423 and 409 K, respectively, for their solutions in H₂O. Consideration of eq 3 or of other equations for ionic solutions^{17,18} shows that electrostatic forces have only a minor role in the phase separation in these systems; the principal factor is the hydrophobic interactions of the large alkyl groups.

For salts with multiply charged ions, ionic effects can cause phase separation in aqueous systems at high temperature. From eq 3 or Figure 1, one can estimate that a 2-2 salt in water might have a lower critical point near 600 K, and indeed, aqueous UO₂SO₄ does show that behavior,^{18,19} with $T_c = 559$ K. Similar lower critical points have been observed for aqueous solutions of alkali-metal phosphates²⁰ and possibly other salts. In UO₂SO₄ and the phosphates, however, there is substantial hydrolysis with acid excess in the dilute phase and acid deficiency in the concentrated phase. Thus, they are really three-component systems and are not good models of simple ionic fluids.

With a nonaqueous second component, however, it is possible to select a system that models rather accurately the simple ionic fluid, yet has a relatively low critical temperature. One chooses large ions which depart not too far from spherical shape and have their charge either buried in the center or widely distributed on the periphery of the ions. Similarly, the solvent should provide electrical dipole moments without having any localized interaction with the ions. Also, if all electrical effects were removed, the two components would have nearly ideal mixing properties. My associates and I found two systems meeting these criteria. In each case the ion size increased from about 2.7 Å in NaCl to about 6 or 7 Å and the permittivity from 1 to a value near 3 or 4. From eq 3, one calculates a reduction of T_c from 3800 K to the range 300-500 K. The properties of these two systems are presented next.

Tetra-*n*-butylammonium Picrate-1-Chloroheptane. This fluid²¹ has a critical point at 414.4 K and a coexistence curve represented by the equation

$$(x_1'' - x_1') = 0.028(414.4 - T)^{1/2} \quad (8)$$

This clearly indicates $\beta = 1/2$; however, the measurements by visual observation of meniscus appearance and disappearance were subject to an uncertainty of 0.1 K. Efforts to make more precise measurements were frustrated by a slow decomposition.

Other properties, including electrical conductance, were measured²² for this picrate system at slightly su-

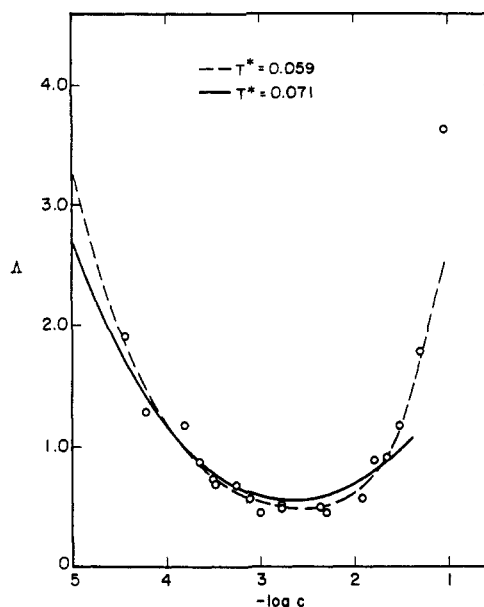


Figure 5. Comparison of experimental conductivity data with theoretical predictions for the primitive model. See text for details.

percritical temperatures. The most interesting of these properties is the conductivity, which can be compared with predictions based on the cluster calculations for the primitive model (see Figure 2). This is shown in Figure 5. There are two adjustments. First is the limiting conductance, which is assumed to be the same for all charged species. Second is the relationship of the concentration scale of the primitive model to that of the experimental system. Clearly the better fit is that for $T^* = 0.059$. In that case $\Lambda_0 = 46$ and the comparison of concentration scales indicates an ion diameter of 7 Å; both values are reasonable for these large ions.

The excellent agreement of experimental with calculated conductance over the full range of validity of the calculations offers strong confirmation of this experimental system as representative of a simple ionic fluid. Thus, it is interesting to note that the conductance increases with concentration, reaching a value near 8 at the critical point. At still higher concentration, one must consider viscosity changes, and the conductance-viscosity product is the appropriate indicator of ionic character. This $\eta\Lambda$ product continues to increase over the full range from the critical composition to the pure fused salt.

R₃R'N⁺R₃R'B⁻-R₂O. Since we were unable to make precision measurements very close to the critical point with our first model system, we sought another with even lower critical temperature and, we hoped, complete chemical stability. A promising candidate for the salt was the alkylammonium alkylboride with the triethyl-*n*-butyl structure on each ion, which was reported to be liquid at room temperature by Ford et al.²³ Ford and Hart²⁴ report high electrical conductance and other salt-like properties for this compound in appropriate solvents. In comparison with the picrate salt investigated earlier, the preparation of this boride was a more demanding task, but was successfully carried out by Dr. Singh, who then studied its solubility relationships in

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Table I
Alternate Sets of Parameters for Eq 10

β	T_c , K	B_0	B_1	B_2
0.5	317.693 ₂	0.351 ± 0.009	-3.4 ± 1.2	27 ± 14
0.326	317.691 ₀	0.058 ± 0.002	24.5 ± 2.7	-123.3 ± 28.4

various solvents.²⁵ With diphenyl ether it shows a critical point at 317.7 K, a very convenient temperature for precise experimentation.

Thus, we selected this system $R_3R'N+R_3R'N-R_2''O$ with R = ethyl, R' = *n*-butyl, and R'' = phenyl as our second model ionic fluid. It proved to be completely stable at its critical temperature, and its other characteristics meet the criteria for a model of a simple ionic fluid. Diphenyl ether has a permittivity in the expected range. The initial series of measurements²⁵ of the coexisting compositions yielded the equation

$$(x_1'' - x_1') = 0.0310(317.7 - T)^{1/2} \quad (9)$$

Thus, we find $\beta = 1/2$ again for this case and obtain an amplitude of 0.031, which is very similar to the value 0.028 for the picrate system.

Precision Measurements Very Near the Critical Point. It was now possible to undertake precision measurements to millikelvin accuracy in temperature.²⁶ We chose to measure the refractive index by the method of Jacobs et al.²⁷ The sample was enclosed in a cell of triangular prismatic shape with two optically flat window surfaces. The difference in refractive index is known to be a good approximation to the difference in volume fraction and is, therefore, a good order parameter. The sample composition was 0.0531 in mole fraction, which is the critical composition found earlier for the boride system.

In all, 31 two-phase measurements²⁶ were made in the range from less than 5 mK to 1.4 K below T_c . An upper limit (317.696 K) was obtained for T_c from the lowest temperature definitely showing no phase separation, but the exact T_c within the range 317.693 ± 0.003 K is best determined from an expression representing the two-phase data. It is useful to represent the data with a Wegner series in $t = 1 - T_c$; this has the form²⁸

$$|n' - n''| = B_0 t^\beta (1 + B_1 t^\Delta + B_2 t^{2\Delta} + \dots) \quad (10)$$

Classical theory leads to the same form,²⁸ in each case Δ is 0.5. Alternate fits were made with $\beta = 0.5$, the classical value, and 0.326, the Ising value which pertains for nonionic fluids.^{28,29} Table I gives the values and standard deviations of the amplitudes B_0 , B_1 , and B_2 . The standard deviation of fit is 0.0001 on either basis, and the deduced T_c in each case is acceptable in that it lies within the range 317.693 ± 0.003 K given above.

Figure 6 shows the measured values²⁶ and calculated curves for eq 10 on the two bases with $\beta = 1/2$ or 0.326. The marked effect of a change of only 2 mK in T_c is apparent for the measurements close to T_c . The deviations follow a random pattern; indeed, for many measurements the deviation has the same sign for either basis.

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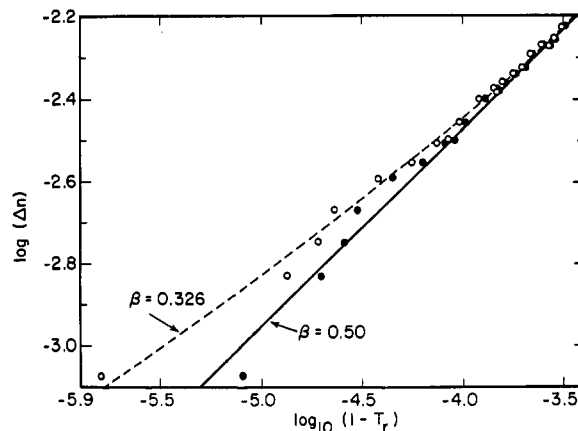


Figure 6. Alternate nonlinear least-squares fits of eq 10 to the experimental measurements. The curves are calculated whereas the solid circles are obtained from the T_c for $\beta = 0.5$ and the open circles from the T_c for $\beta = 0.326$. Reprinted with permission from ref 26. Copyright 1990 American Institute of Physics.

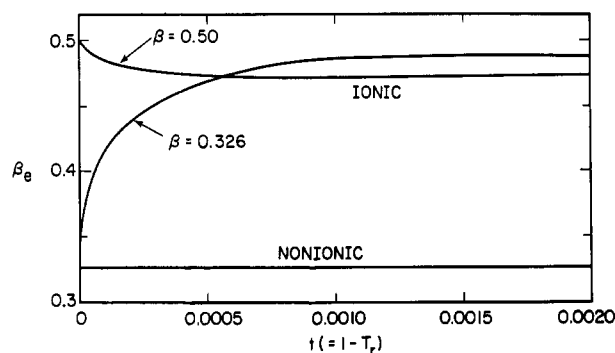


Figure 7. Curves of effective β given by the alternate equations of Table I and Figure 6. The curve for a typical nonionic liquid-liquid system is also shown. Reprinted with permission from ref 26. Copyright 1990 American Institute of Physics.

It is also informative to consider the effective exponent β_e which is defined by

$$\beta_e = \partial \ln |n' - n''| / (\partial \ln t) \quad (11)$$

This was evaluated from eq 10 for each set of parameters in Table I and gave the curves shown in Figure 7.

It is apparent from these results and the earlier measurements²⁵ further from the critical point that β_e is between 0.47 and 0.50 for the full range of t down to 0.001 and β_e is consistent with either curve on Figure 7 for smaller t . Thus it seems virtually certain that these data are concordant with a limiting exponent of 0.50. But the data are also concordant with the curve showing an extremely sudden shift in β_e in the range $t < 0.0001$ and a limiting β at the value for neutral fluid systems of 0.326. On the latter basis, however, the Wegner expansion coefficients are larger by at least an order of magnitude than any reported heretofore. For single-component neutral fluids, B_1 is usually near 1.0, while for two-component liquid systems, B_1 is often 0,³⁰ thus, the present value $B_1 = 24.5$ is anomalous.

Theory for Critical Exponents of an Ionic Fluid

Kholodenko and Beyerlein^{31,32} presented very recently a treatment in which the partition function for the ionic

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fluid was mapped onto that for the well-known spherical model.³³ This predicts $\beta = 1/2$, in full agreement with our results. But it indicates nonclassical values for ν , δ , and other critical exponents that are yet to be measured.

Hafskjold and Stell³⁴ give an excellent review of theory for ionic fluids to 1982. But their treatment for the critical exponents was inconclusive and was based on considerations for simple fluids with a single type of particle, as was our recent discussion.²⁶ It now appears^{31,35} that an electrolyte with its long-range-force pattern involving repulsive forces between like particles

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and attractive forces between unlike particles is fundamentally different and that arguments by analogy to simple fluids are not valid.

There are many opportunities for further research on ionic fluids. For example, measurements of properties related to the other critical exponents would be most welcome, as well as theoretical advances giving predictions over a range of temperature, etc., in the critical region.

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Determining the Functional Conformations of Biologically Active Peptides

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In recent years, the development of microscale protein sequencing and cloning technologies has provided a wealth of information on the amino acid sequences of polypeptides. However, knowledge of the chemical composition of a peptide is of little value in understanding its biological function, unless information concerning its active conformation is also available. In cases where the peptide chains are long enough, they fold into stable globular structures composed of helices, sheets, turns, and loops that can often be characterized in detail by X-ray crystallography and, increasingly, by NMR spectroscopy. This type of structural characterization usually provides an adequate basis for selecting analogues (mutants) that may be prepared by genetic engineering and expression, in order to understand and manipulate function in greater detail.

Unfortunately, many biologically active peptides do not maintain their functional conformations when isolated in a crystalline form or in aqueous solution. They do not fold into independently stable globular structures with a close-packed hydrophobic core, but depend

instead upon extensive intermolecular interactions in order to define their active states. Most peptides consisting of fewer than 50 amino acid residues and having no more than one disulfide bridge fall into this category, including a wide variety of peptide hormones that are of great pharmacological and therapeutic interest. These peptides exist in aqueous solution in multiple, rapidly interchanging conformational states and can be expected to adopt different conformations in different environments. Determining the active conformation of such peptides has, therefore, been particularly difficult, since they are often large enough to have complex conformations that contain elements of secondary and even tertiary structure, and they cannot generally be studied directly in their functional environment. This Account describes three complementary approaches to this problem.

Study of Conformation Induction in Peptides at Model Interfaces

Since the biological activities of many flexible, multiconformational peptides are generally expressed upon binding to a cell surface receptor or other biological interface, their functional conformations are most likely to include structural elements that bind to these interfaces. The potential importance of this type of conformation induction at interfaces was introduced in a general form by Kaiser and Kezdy,¹ who suggested that amphiphilic secondary structures, having one surface comprising predominantly hydrophobic ele-

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