Calorimetric Investigations of Hydrogen Bond and Charge Transfer Complexes

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1 Introduction

On the basis of formidable experimental evidence it is believed that complex formation between like and/or unlike molecules takes place in many systems. This intermolecular complex formation is significant in many areas of chemistry and biochemistry and consequently has attracted considerable attention.

Complexes are usually classified according to the intermolecular interaction primarily responsible for their formation. Hydrogen bond complex formation, first recognized explicitly by Latimer and Rodebush[1] in 1920, has been discussed comprehensively in a number of monographs.[2-4] The theory of the hydrogen bond has been reviewed recently.[5] Current interest in charge transfer complex formation began with the spectroscopic studies of Benesi and Hildebrand[6] in 1948. The vast amount of research since this pioneering work has been discussed elsewhere.[7-12]

The thermodynamic properties associated with complex formation are of fundamental importance. Reliable values are necessary for many chemical and biochemical applications, and also for testing such extrathermodynamic relationships as the Badger–Bauer correlation[13] of hydrogen bond enthalpy

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[1] W. M. Latimer and W. H. Rodebush, *J. Amer. Chem. Soc.*, 1920, **42**, 1419.

[2] 'Hydrogen Bonding', ed. D. Hadzi and H. W. Thompson, Pergamon Press, Oxford, 1959.

[3] G. C. Pimentel and A. L. McClellan, 'The Hydrogen Bond', W. H. Freeman, San Francisco, 1960.

[4] S. N. Vinogradov and R. H. Linnell, 'Hydrogen Bonding', Van Nostrand Reinhold Co., New York, 1971.

[5] P. A. Kollman and L. C. Allen, *Chem. Rev.*, 1972, **72**, 283.

[6] H. A. Benesi and J. H. Hildebrand, *J. Amer. Chem. Soc.*, 1948, **70**, 3978; *ibid.*, 1949, **71**, 2703.

[7] R. S. Mulliken and W. B. Person, 'Molecular Complexes', Wiley-Interscience, New York, 1969.

[8] R. Foster, 'Organic Charge-Transfer Complexes', Academic Press, London, 1969.

[9] M. A. Slifkin, 'Charge Transfer Interactions of Biomolecules', Academic Press, London, 1971.

[10] G. Briegleb, 'Elektronen-Donator-Acceptor-Komplexe', Springer-Verlag, Berlin, 1961.

[11] L. J. Andrews and R. M. Keefer, 'Molecular Complexes in Organic Chemistry', Holden-Day, San Francisco, 1964.

[12] J. Rose, 'Molecular Complexes,' Pergamon Press, Oxford, 1967.

[13] R. M. Badger and S. H. Bauer, *J. Chem. Phys.*, 1937, **5**, 839.

with i.r. shift and the correlation[14] of charge transfer bond energy with the enhancement of dipole moment.

A wide range of techniques, discussed in the books referred to above, has been applied to investigations of the thermodynamics of complex formation. Following a review of the data available in 1959, Pimentel and McClellan[3] commented that '. . . the most outstanding conclusion is that additional and systematic studies are needed.' With the substantial number of investigations since then and the ever greater appreciation of some of the complicating factors, the situation has improved. But even in 1969 Foster[8] wrote that 'none of the methods currently used yields results in which, for all systems, high confidence can be placed.' Given this situation and the importance of the thermodynamics of complex formation, all new methods for their investigation deserve scrutiny. In this paper we discuss a number of calorimetric methods developed in recent years. As these possess certain advantages over traditional methods, they represent a significant contribution to the investigation of complex formation in solution. Because they have been introduced only in the last few years, calorimetric methods have received brief mention only in the more recent[4,8,9] of the general references given above.

2 Thermodynamic Properties of Complex Formation

At first we consider the formation of only one complex of 1:1 stoicheiometry between unlike molecules as represented by

$$A + B \rightleftharpoons AB \qquad (1)$$

It is customary (and reasonable[8,11,15]) to take all activity coefficients to be unity so that the equilibrium constant for this reaction can be expressed in terms of the equilibrium composition of the solution. For example, we commonly express the equilibrium composition in terms of concentrations (molarities) and thence obtain an equilibrium constant that we represent by K_c. Equilibrium compositions may also be expressed in terms of mole fractions and molalities, thus leading to equilibrium constants represented by K_z and K_m, respectively.

It should be emphasized that the assumption of all activity coefficients equal to unity (or the nearly equivalent assumption of a ratio of activity coefficients equal to unity) is common to most methods used in investigation of complex formation in non-electrolyte systems. In particular, we emphasize that similar assumptions about activity coefficients are used in the various spectroscopic methods and also the calorimetric methods discussed here.

For a *dilute solution* of A and B in some solvent S, the most common situation experimentally, the K values and ΔG° values are related as follows:

$$K_z = K_m(M_S)^{-1} = K_c(V_{m,S})^{-1} \qquad (2)$$

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[14] H. Ratajczak and W. J. Orville-Thomas, *J. Chem. Phys.*, 1973, **58**, 911.

[15] K. Denbigh, 'The Principles of Chemical Equilibrium', 2nd edn., Cambridge University Press, London, 1966.

$$\Delta G_z^\circ = \Delta G_m^\circ + RT\ln(M_S) = \Delta G_c^\circ + RT\ln(V_{m,S}) \qquad (3)$$

where M_S is the molar mass of solvent S and $V_{m,S}$ is the molar volume of S. The standard enthalpy of complex formation at infinite dilution (ΔH°) is given by the van't Hoff equation,

$$\Delta H^\circ = -R\left(\frac{\partial \ln K_m}{\partial 1/T}\right)_p \qquad (4)$$

It follows from equation (2) that ΔH° is related to the temperature coefficients of K_z, K_m, and K_c as follows:

$$-R\left(\frac{\partial \ln K_z}{\partial 1/T}\right)_p = -R\left(\frac{\partial \ln K_m}{\partial 1/T}\right)_p = \Delta H^\circ \qquad (5)$$

$$-R\left(\frac{\partial \ln K_c}{\partial 1/T}\right)_p = \Delta H^\circ + RT^2\alpha_S \qquad (6)$$

where α_S is the thermal expansivity of the solvent. The term $RT^2\alpha_S$[16,17] in equation (6) is mistakenly omitted from many calculations; with $\alpha_S \approx 10^{-3}\ K^{-1}$ for 'typical' solvents at 300 K, $RT^2\alpha_S \approx 1$ kJ mol^{-1} (*i.e.*, often about 10% of ΔH°). The standard entropies of complex formation are related by

$$\Delta S_z^\circ = \Delta S_m^\circ - R\ln(M_S) = \Delta S_c^\circ - R\ln(V_{m,S}) \qquad (7)$$

3 Determination of Thermodynamic Properties

The thermodynamic properties of complex formation have usually been obtained from measurements of some equilibrium constant and its variation with temperature. The methods used have been reviewed.[2-4,7-12]

The equilibrium constant has most frequently been obtained from spectroscopic measurements (i.r., u.v., visible, n.m.r.) carried out over a range of compositions of one or both of the reactants. Thermodynamic procedures such as distribution between immiscible liquids and measurement of various colligative properties, dielectric studies, polarography, *etc.*, have been used to a much lesser extent. In all of these methods it is usual to incorporate the assumption that activity coefficients may be approximated to unity. As most measurements have involved dilute solutions of A and B in some 'inert' solvent, the molarity equilibrium constant K_c has been most widely used. (This is often experimentally more convenient than K_m or K_z and is entirely 'proper' when ΔH° is calculated correctly as noted above.)

Perusal of published equilibrium constants suggests that in many cases little confidence can be placed in the results. Reported equilibrium constants for a

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[16] E. J. King, 'Acid–Base Equilibria', Pergamon Press, Oxford, 1965.

[17] E. M. Woolley, J. G. Travers, B. P. Erno, and L. G. Hepler, *J. Phys. Chem.*, 1971, **75**, 3591.

particular complex frequently differ by substantially more than the assigned uncertainties; for example, in Table 1 we list equilibrium constants for the 1,3,5-trinitrobenzene–hexamethylbenzene complex at 293 K obtained from u.v. studies. The situation has been improved by critical analyses, such as that of Person,[22] of the various spectroscopic procedures. With careful experimental work and data analysis, reliable equilibrium constants can be obtained.

Table 1 *Equilibrium constants for the formation of the* 1,3,5-*trinitrobenzene–hexamethylbenzene complex at* 293 K

Solvent	K_c/dm^3 mol^{-1}	*Reference*
cyclohexane	17.50 ± 0.20	18
	13.5 ± 0.4	19
carbon tetrachloride	4.87 ± 0.08	18
	5.7 ± 0.3	19
	7.10 ± 0.05	20
chloroform	0.92 ± 0.10	18
	0.76 ± 0.05	19
	0.81	21

Having measured equilibrium constants over a temperature range, the value of ΔH° can be obtained by using equations (4)—(6). (The values of ΔH° commonly lie between -10 and -30 kJ mol^{-1}.) Since differentiation procedures always result in a loss of precision, ΔH° (and $\Delta S_c{}^\circ$) will be less precisely known than $\ln K_c$. King[16] has described a thorough and rigorous analysis of the relationship between random errors in equilibrium constants and the thermodynamic properties derived by differentiation. Even with equilibrium constants obtained from careful spectroscopic studies, the propagation of errors can lead to standard deviations in ΔH° of 1—2 kJ mol^{-1}. As King's statistical analysis excludes systematic errors, the standard deviations obtained represent lower limits for the total uncertainties. In view of the above, it is not surprising that the discrepancies between published results for ΔH° are often large, sometimes exceeding 100%. [Extensive comparisons of K and ΔH° values obtained in different ways are given in refs. 27, 28, 32, 35, 36, 39, 44, and 64, discussed later in this review.]

A number of methods have been advanced recently for evaluation of ΔH° (and usually K) from results of calorimetric measurements. With most of these procedures, ΔH° is obtained from calorimetric results at a single temperature and without recourse to the van't Hoff equation. It thus appears that these calorimetric methods offer significant advantages over other approaches which

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[18] C. C. Thompson and P. A. D. de Maine, *J. Phys. Chem.*, 1965, **69**, 2766.
[19] R. Foster, *J. Chem. Soc.*, 1960, 1075.
[20] G. Briegleb and J. Czekalla, *Z. Elektrochem.*, 1955, **59**, 184.
[21] C. E. Castro, L. J. Andrews, and R. M. Keefer, *J. Amer. Chem. Soc.*, 1958, **80**, 2322.
[22] W. B. Person, *J. Amer. Chem. Soc.*, 1965, **87**, 187.

rely on differentiation procedures. Lamberts[23] has provided a general review of calorimetric investigations of molecular complexes in solution while Christensen *et al.*[24–26] have reviewed applications of titration calorimetry to a variety of reactions in solution. Our discussion is intended to supplement these earlier reviews.

It is convenient to divide the calorimetric methods into those in which complex formation takes place in dilute solution in some 'inert' solvent and those involving no solvent.

4 Complex Formation in Inert Solvents

In this section we consider the complex formation represented by the equation

$$\mathrm{A(S)} + \mathrm{B(S)} \rightleftharpoons \mathrm{AB(S)} \qquad (8)$$

in which (S) indicates that the preceding species is in dilute solution in the solvent S. Following usual practice, we will use the molarity equilibrium constant K_c. With the notation given in Table 2, K_c is given by

$$K_c = \frac{c}{(c_\mathrm{A} - c)(c_\mathrm{B} - c)} \qquad (9)$$

Table 2 *Summary of symbols used*

	A	B	AB	S
Stoicheiometric number of moles	N_A	N_B		N_S
Stoicheiometric mole fractions	x_A	x_B		x_S
Stoicheiometric molarities	c_A	c_B		c_S
Equilibrium numbers of moles	$N_\mathrm{A}-n$	$N_\mathrm{B}-n$	n	N_S
Equilibrium mole fractions	z_A	z_B	z_AB	z_S
Equilibrium molarities	$c_\mathrm{A}-c$	$c_\mathrm{B}-c$	c	c_S

This equation can be rearranged to

$$c^2 - (c_\mathrm{A} + c_\mathrm{B} + K_c^{-1})c + c_\mathrm{A}c_\mathrm{B} = 0 \qquad (10)$$

Calorimetric methods involve measurement of the enthalpy change Q associated with formation of n mol of complex AB according to the reaction represented by equation (8). This measured Q is related to the standard enthalpy of complex formation at infinite dilution, ΔH°, by

$$Q = n\Delta H^\circ = cV\Delta H^\circ \qquad (11)$$

in which V represents the total volume of the solution. The enthalpy change of

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[23] L. Lamberts, *Ind. chim. belge*, 1971, **36**, 347.

[24] J. J. Christensen, J. Ruckman, D. J. Eatough, and R. M. Izatt, *Thermochim. Acta*, 1972, **3**, 203.

[25] D. J. Eatough, J. J. Christensen, and R. M. Izatt, *Thermochim. Acta*, 1972, **3**, 219.

[26] D. J. Eatough, R. M. Izatt, and J. J. Christensen, *Thermochim. Acta*, 1972, **3**, 233.

interest is obtained experimentally as a difference between a principal experimental enthalpy change and effects extraneous to the complex formation, such as enthalpies of solution or dilution. A widely used procedure is illustrated by the following cycle:

$$\begin{tikzcd}
\begin{array}{l}\text{A (pure)} + \text{B (S, dilute)} \\ + \\ \text{S}\end{array} \arrow[d, "Q_{\text{II}}"] \arrow[r, "Q_{\text{I}}"] & \text{AB (S, dilute)} \\
\text{A (S, dilute)} + \text{B (S, dilute)} \arrow[ur] &
\end{tikzcd}$$

A (pure) represents either pure solid A or pure liquid A. The desired enthalpy change Q for reaction (8) is obtained as the difference between the two experimental enthalpy changes Q_{I} and Q_{II}:

$$Q = Q_{\text{I}} - Q_{\text{II}} \qquad (12)$$

In some procedures, A(pure) is replaced by A(S, concentrated).

All of the techniques to be discussed in this section are based on a combination of equations (9) or (10) with equation (11). In each case we derive the fundamental equation and indicate the method used in its application. No attempt has been made to list all of the systems to which these methods have been applied, but some key references are given.

Equation (10) is a quadratic equation in c and can be solved to give

$$2c = (c_{\text{A}} + c_{\text{B}} + K_c^{-1}) - [(c_{\text{A}} + c_{\text{B}} + K_c^{-1})^2 - 4c_{\text{A}}c_{\text{B}}]^{\frac{1}{2}} \qquad (13)$$

[The other root of equation (10) is unacceptable.]
Using a K_c value determined independently (for example, from spectroscopic measurements), equation (13) permits calculation of c values for a series of stoicheiometric molarities c_{A} and c_{B}. These values of c, together with the corresponding Q values, can be used in equation (11), according to which a plot of Q *vs.* cV will be linear with slope ΔH° (and zero intercept). This procedure was introduced by Arnett and co-workers.[27] The results obtained have been compared[27,28] with those obtained using another calorimetric method (to be discussed in the next section) and with spectroscopic results. While dependent upon other studies for values of K_c, this method does not involve trial-and-error calculations as do most of the techniques capable of yielding both K_c and ΔH° from calorimetric measurements alone. The ΔH° value obtained is very sensitive to variations in K_c when the latter is small.

Methods advanced to obtain simultaneously both K_c and ΔH° from equations

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[27] E. M. Arnett, T. S. S. R. Murty, P. von R. Schleyer, and L. Joris, *J. Amer. Chem. Soc.*, 1967, **89**, 5955.

[28] E. M. Arnett, L. Joris, E. Mitchell, T. S. S. R. Murty, T. M. Gorie, and P. von R. Schleyer, *J. Amer. Chem. Soc.*, 1970, **92**, 2365.

(10) and (11) have involved trial-and-error methods of calculation. The limits of such methods arising from the propagation of errors have been discussed by Cabani and Gianni,[29] who have shown that a good fit between calculated and experimental data can be compatible with fictitious values of K_c and ΔH° if the experiments are conducted with a constant concentration of one of the reactants. Procedures to reveal systematic errors are listed and a weighting procedure, which takes into account the different propagation of errors for different experimental conditions, is proposed.[29]

Equations (11) and (13) can be combined to give

$$\Delta H^\circ = \frac{2\,Q/V}{(c_A + c_B + K_c^{-1}) - [(c_A + c_B + K_c^{-1})^2 - 4c_Ac_B]^{\frac{1}{2}}} \qquad (14)$$

Using this equation, Lamberts and Zeegers-Huyskens[30] and others[31] have obtained both K_c and ΔH° by way of trial-and-error calculations. For each set of experimental results (c_A, c_B, V, Q), ΔH° values are calculated from equation (14) for a range of values of K_c. From plots of ΔH° *vs.* c_A or c_B for each of these K_c values, one selects as the 'best' value of K_c that which gives a constant value of ΔH° (*i.e.*, a line of zero slope).

A method introduced by Bolles and Drago[32] and extensively used by Drago and co-workers[33,34] involves the following combination of equations (10) and (11):

$$K_c^{-1} = \frac{1}{V}\frac{Q}{\Delta H^\circ} + c_Ac_BV\frac{\Delta H^\circ}{Q} - (c_A + c_B) \qquad (15)$$

Again both K_c and ΔH° are obtained by trial-and-error calculations. For each set of experimental results (c_A, c_B, V, Q), ΔH° values are calculated for a range of values of K_c, and K_c^{-1} is plotted against ΔH°. The curves thus obtained, one for each set of experimental results, should ideally intersect at a single point, this being the unique solution of equation (15). A statistical procedure for this method of data analysis[35] has been applied. Comparisons between calorimetric and spectroscopic results have been given.[32,35]

Another approach, making use of calorimetric measurements at two temperatures, has been proposed by Neerinck *et al.*[36] Ignoring $RT^2\alpha_S$ in equation (6) and assuming ΔH° to be temperature independent over the temperature range of interest, one obtains

$$\ln[K_c(T_1)/K_c(T_2)] = \frac{\Delta H^\circ}{R}\left(\frac{1}{T_2} - \frac{1}{T_1}\right) \qquad (16)$$

[29] S. Cabani and P. Gianni, *J. Chem. Soc.* (*A*), 1968, 547.
[30] L. Lamberts and F. Zeegers-Huyskens, *J. Chim. phys.*, 1963, 435.
[31] L. Abello and G. Pannetier, *Bull. Soc. chim. France*, 1967, 3752.
[32] T. F. Bolles and R. S. Drago, *J. Amer. Chem. Soc.*, 1965, **87**, 5015.
[33] R. M. Guidry and R. S. Drago, *J. Amer. Chem. Soc.*, 1973, **95**, 759.
[34] F. L. Slejko and R. S. Drago, *Inorg. Chem.*, 1973, **12**, 176.
[35] T. D. Epley and R. S. Drago, *J. Amer. Chem. Soc.*, 1967, **89**, 5770.
[36] D. Neerinck, A. Van Audenhaege, L. Lamberts, and P. Huyskens, *Nature*, 1968, **218**, 461.

where $K_c(T)$ is the molarity equilibrium constant at temperature T. Substituting equation (9) into equation (16) and rearranging gives

$$\Delta H^\circ = \frac{RT_1T_2}{T_1 - T_2} \ln \frac{c_1(c_A - c_2)(c_B - c_2)}{c_2(c_A - c_1)(c_B - c_1)} \tag{17}$$

on the basis of experiments arranged so that the stoicheiometric molarities c_A and c_B are the same at temperatures T_1 and T_2. In this equation c_1 and c_2 represent the molarities of the complex at temperatures T_1 and T_2, respectively. Substituting $c = Q/V\Delta H^\circ$ from equation (11) into equation (17) gives an equation from which ΔH° can be obtained by trial and error. Having obtained ΔH°, equation (9) leads to K_c. Results from this approach were compared[36] with those obtained by the calorimetric method of Bolles and Drago[32] discussed above; the ΔH° values agree well but there are considerable discrepancies between the K_c values.

All of the above methods capable of yielding both K_c and ΔH° involve trial-and-error methods of calculation. A new and potentially better procedure, which we propose here, avoids the necessity for trial-and-error calculations. When the stoicheiometric molarities c_A and c_B (at constant temperature) are chosen so that their sum is always the same we define a constant χ by

$$\chi = c_A + c_B + K_c^{-1} \tag{18}$$

Combining equations (15) and (18) and rearranging gives

$$\frac{Vc_Ac_B}{Q} = -\frac{1}{(\Delta H^\circ)^2} \cdot \frac{Q}{V} + \frac{\chi}{\Delta H^\circ} \tag{19}$$

according to which a plot of Vc_Ac_B/Q *vs.* Q/V should be linear with slope $-1/(\Delta H^\circ)^2$ and intercept $\chi/\Delta H^\circ$. From such a plot we can therefore obtain ΔH° and, if c_A and c_B are chosen so that $(c_A + c_B)K_c$ is not considerably greater than unity, K_c can also be evaluated. We believe that this approach could prove to be preferable to those dependent upon trial-and-error methods of calculation. A thorough test of this method, which appears to be especially suited to flow calorimetry, should involve evaluation of K_c and ΔH° from measurements at different values of $(c_A + c_B)$.

5 Complex Formation Studies without Inert Solvents

In this section we will be concerned with the mixing of pure liquid A and pure liquid B to give a liquid mixture, all at constant temperature and pressure:

$$N_A \text{ mol A } (T, p) + N_B \text{ mol B } (T, p) \rightarrow \text{mixture } (T, p) \tag{20}$$

The influence of complex formation reactions, such as that represented by equation (1), on the thermodynamic properties of liquid mixtures has long been recognized but remains difficult to assess. At the beginning of this century Dolezalek[37] and his followers tried to account for all deviations from ideal solution behaviour in terms of chemical equilibria between the like and/or unlike molecules. This approach was soon shown to be chemically unrealistic and numerically inadequate for many systems. But it has also been shown that complex formation as represented by chemical reaction equations such as (1) is the principal interaction in many solutions. Properties of these solutions may therefore be realistically and usefully interpreted in terms of the thermodynamics of complex formation as described below.

In this section it is convenient to use the mole fraction equilibrium constant K_z (again assuming the activity coefficient ratio to be unity):

$$K_z = z_{AB}/z_A z_B \tag{21}$$

Combining this assumption with the Gibbs–Duhem equation[15] leads to

$$\gamma_{z,AB} = \gamma_{z,A} = \gamma_{z,B} = 1 \tag{22}$$

The mixture of the three chemical species (A, B, AB) therefore behaves as an ideal solution and all deviations from ideality observed experimentally for a mixture of N_A (stoicheiometric) mol A with N_B (stoicheiometric) mol B can be attributed to the formation of the complex AB. In the present situation, the common assumption of activity coefficient ratio equal to unity is therefore equivalent to assuming that the system is an 'ideal associated solution'.[38] The enthalpy change associated with the mixing process represented by equation (20) is supposed to arise entirely from the formation of the complex.

An alternative formulation of the discussion above is useful in that it leads directly to a relationship between the enthalpy of mixing and the thermodynamics of complex formation.[39] In this formulation we discuss the thermodynamics of mixing in terms of 'physical' (non-complexing) and 'chemical' contributions, with the latter due to complex formation.[40-42] According to this model, the molar enthalpy of mixing, ΔH_m, can be expressed as

$$\Delta H_m = \Delta H_m^{\text{phys}} + \Delta H_m^{\text{chem}} \tag{23}$$

in which ΔH_m^{phys} and ΔH_m^{chem} represent the molar 'physical' and 'chemical' contributions. For systems with large exothermic enthalpies of mixing ($-\Delta H_m$ greater than *ca.* 1.5 kJ mol^{-1} at $x = 0.5$), it is usually reasonable to neglect

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[37] F. Dolezalek, *Z. phys. Chem.*, 1908, **64**, 727.

[38] I. Prigogine and R. Defay, 'Chemical Thermodynamics', translated by D. H. Everett, Longmans Green and Co., London, 1950.

[39] L. G. Hepler and D. V. Fenby, *J. Chem. Thermodynamics*, 1973, **5**, 471.

[40] R. Anderson, R. Cambio, and J. M. Prausnitz, *J. Amer. Inst. Chem. Eng.*, 1962, **8**, 66.

[41] I. D. Watson and A. G. Williamson, *J. Sci. Ind. Res., India*, 1965, **24**, 615.

[42] R. L. Scott, 'Weak Complexes: Molecular Models and Thermodynamic Evidence', paper read before the Chemical Society, Exeter, 1967.

ΔH_m^{phys} in the above expression. On this basis, we proceed to relate molar enthalpies of mixing to the equilibrium constant K_z and the enthalpy of complex formation ΔH°. To do so we use $n = Q/\Delta H^\circ$ from equation (11) and

$$\Delta H_m = Q/(N_{\mathrm{A}} + N_{\mathrm{B}}) \tag{24}$$

which relates the measured enthalpy Q and stoicheiometric amounts of A and B to the molar enthalpy of mixing that is usually reported.

Using the notation summarized in Table 2 with $N_{\mathrm{S}} = 0$, equation (21) can be written as

$$K_z = \frac{n(N_{\mathrm{A}} + N_{\mathrm{B}} - n)}{(N_{\mathrm{A}} - n)(N_{\mathrm{B}} - n)} \tag{25}$$

Substituting $n = Q/\Delta H^\circ$ from equation (11) in (25) and rearranging yields

$$\frac{N_{\mathrm{A}}N_{\mathrm{B}}}{Q} = -\frac{K_z + 1}{K_z(\Delta H^\circ)^2}\,Q + (N_{\mathrm{A}} + N_{\mathrm{B}})\frac{K_z + 1}{K_z\Delta H^\circ} \tag{26}$$

Further substitution of $Q = \Delta H_m(N_{\mathrm{A}} + N_{\mathrm{B}})$ from (24) in (26) and use of the definition of stoicheiometric mole fraction leads to

$$\frac{x_{\mathrm{A}}x_{\mathrm{B}}}{\Delta H_m} = -\frac{K_z + 1}{K_z(\Delta H^\circ)^2}\,\Delta H_m + \frac{K_z + 1}{K_z\Delta H^\circ} \tag{27}$$

According to this equation, a plot of $x_{\mathrm{A}}x_{\mathrm{B}}/\Delta H_m$ *vs.* ΔH_m should be linear; K_z and ΔH° can be evaluated from the slope and intercept. We have applied equation (27) to investigation of complex formation between chloroform and triethylamine[39] and between chloroform and various ethers.[43]

It is also possible to relate differential or partial molar enthalpies of solution to K_z and ΔH° in ways that permit useful evaluation of these latter quantities from results of calorimetric measurements.[44] We begin by identifying $\bar{L}_{\mathrm{A}}$ and $\bar{L}_{\mathrm{B}}$ as partial molar enthalpies of solution of A and B, respectively. A formal definition is given by

$$\bar{L}_{\mathrm{A}} = \left(\frac{\partial Q}{\partial N_{\mathrm{A}}}\right)_{T,p,N_{\mathrm{B}}} \tag{28}$$

First, we will be concerned with partial molar enthalpies of solution at infinite dilution ($\bar{L}_{\mathrm{A}}^\circ$ and $\bar{L}_{\mathrm{B}}^\circ$). These limiting partial molar enthalpies of solution may be obtained[45] almost directly from 'small increment' solution calorimetry or

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[43] N. F. Pasco, D. V. Fenby, and L. G. Hepler, *Canad. J. Chem.*, in the press.

[44] T. Matsui, L. G. Hepler, D. V. Fenby, *J. Phys. Chem.*, 1973, **77**, 2397.

[45] H. C. Van Ness, 'Classical Thermodynamics of Non-Electrolyte Solutions', Pergamon Press, New York, 1964.

(less accurately) from 'integral' enthalpies of mixing such as those already represented by ΔH_m. We have recently compared $\bar{L}°$ values obtained by way of these two approaches.[44]

In the limit $N_A \rightarrow 0$ it follows from equation (25) that $K_z \rightarrow n/(N_A - n)$. Consequently, in this limit, the fraction of A that is complexed is given by

$$n/N_A = K_z/(1 + K_z) \tag{29}$$

As all enthalpy changes have been attributed to the formation of the complex,

$$\bar{L}_A° = (n/N_A)\Delta H° = K_z \Delta H°/(1 + K_z) \tag{30}$$

or

$$\Delta H° = (1 + K_z)\bar{L}_A°/K_z \tag{31}$$

Because the system A + B is symmetrical when AB is the only complex, we also have

$$\Delta H° = (1 + K_z)\bar{L}_B°/K_z \tag{32}$$

Equation (26), a quadratic in Q, can be solved and the result differentiated with respect to N_A to give a general expression for $\bar{L}_A$ as previously described.[44] In the limit $N_A \rightarrow 0$ this general expression reduces to equation (30), which has already been derived in a simpler way. In the case $N_A = N_B$, the general expression reduces to[44]

$$\Delta H° = 2\,\bar{L}_{0.5}/[1 - (1 + K_z)^{\frac{1}{2}}] \tag{33}$$

in which $\bar{L}_{0.5}$ represents the partial molar enthalpy of solution of either A or B in solution with $x_A = x_B = 0.5$.

We have applied equations (31)—(33) to investigation of complex formation in the chloroform–triethylamine system in two ways.[44] First, calorimetrically determined $\bar{L}$ values have been combined with previously reported K_z values to yield the desired $\Delta H°$. Second, these simultaneous equations have been solved (either graphically or analytically) to yield values of both $\Delta H°$ and K_z.

For those cases in which $K_z \gg 1$, equations (29), (31), and (32) reduce to

$$n/N_A \approx 1 \tag{34}$$

and

$$\bar{L}_A° \approx \Delta H° \tag{35}$$

In such cases it follows from equation (34) that (almost) all of the A molecules (in infinitely dilute solution in B) are involved in complex formation. The application of equation (35) to evaluation of $\Delta H°$, while appealingly simple, requires caution as it is only applicable if the equilibrium constant is sufficiently large.

Arnett *et al.*[27,28] have applied equation (35) to a large number of systems for which it appears to be a valid approximation. In this work pure A was a solid, making it necessary to obtain $\bar{L}_A{}^\circ$ indirectly. This was done by introducing an 'inert' solvent S and a 'model compound' M, a compound similar to A but unable to undergo complex formation with B (*e.g.* A = phenol, M = anisole or fluorobenzene). The value of $\bar{L}_A{}^\circ$ was obtained by using the expression

$$\bar{L}_A{}^\circ = [\bar{L}_A{}^\circ{}_{(s)}\,(\mathrm{B}) - \bar{L}_M{}^\circ{}_{(s)}\,(\mathrm{B})] - [\bar{L}_A{}^\circ{}_{(s)}\,(\mathrm{S}) - \bar{L}_M{}^\circ{}_{(s)}\,(\mathrm{S})] \qquad (36)$$

in which $\bar{L}_X{}^\circ{}_{(s)}$ (Y) represents the partial molar enthalpy of solution of solid X in infinitely dilute solution in the liquid Y. Although the results reported by Arnett *et al.*[27,28] are in good agreement with the results obtained using their method discussed earlier in this review, [equations (11) and (13)], a critical analysis[46] has indicated that the ΔH° values obtained are strongly dependent on the choice of both the 'inert' solvent and the 'model compound'.

Duer and Bertrand[46] have modified equation (36) to determine the difference in the standard enthalpies of formation of two complexes, AB and AC (both in the same solvent):

$$\Delta H^\circ(\mathrm{AB}) - \Delta H^\circ(\mathrm{AC}) = [\bar{L}_A{}^\circ{}_{(s)}\,(\mathrm{B}) - \bar{L}_M{}^\circ{}_{(s)}\,(\mathrm{B})] - [\bar{L}_A{}^\circ{}_{(s)}(\mathrm{C}) - \bar{L}_M{}^\circ{}_{(s)}(\mathrm{C})] \qquad (37)$$

This procedure eliminates the effect of the 'inert' solvent. The ΔH° differences obtained were found to be nearly independent of the choice of 'model compound' M; hence this appears to be a useful procedure for obtaining relative ΔH° values.

The approximation represented by equations (34) and (35) has also been implied in other investigations[47-49] in which its validity is questionable.

6 Further Complex Formation

In many systems complex formation in addition to that represented by equation (1) takes place. Some of the methods discussed in the last two sections can be extended to these more complicated situations.

Systems in which AB and AB_2 complexes are present appear to be quite common. McGlashan and Rastogi,[50] assuming that such systems behave as ideal associated solutions, have proposed a method whereby the two equilibrium constants can be obtained from vapour pressure measurements. These equilibrium constants, or those obtained in any other way, can then be combined with results of calorimetric measurements to yield the standard enthalpies of complex

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formation.[50-53] The method based on partial molar enthalpies of solution that was discussed in the last section has been extended[44] to allow the standard enthalpies of formation of both the AB and AB_2 complex to be evaluated from previously reported equilibrium constants and calorimetric $\bar{L}_A^\circ$ and $\bar{L}_B^\circ$ values.

Calorimetric results have also been used in investigations of various self-association reactions that can be represented by

$$nA(S) \rightleftharpoons A_n(S) \qquad (38)$$

in which (S) indicates that monomeric A and associated species A_n are in solution in some solvent S.

We first call attention to the pioneering investigation of Schellman.[54] He used osmotic coefficients for evaluation of the equilibrium constant for dimerization of urea in aqueous solution, and then combined the equilibrium constant with enthalpies of dilution for evaluation of ΔH° of dimerization. Subsequent investigations by Stokes have provided further experimental results and considerably more extensive theoretical analysis of self-association of urea in aqueous solution[55] and in various other solvents.[56] We also call attention to the work of Gill *et al.*,[57-60] who have derived useful relationships between enthalpies of dilution and thermodynamic functions for self-association reactions such as represented by equation (38). They have investigated several aqueous systems of biochemical interest.

Many investigations[2-4] have provided convincing evidence that carboxylic acids form hydrogen-bonded dimers in more or less 'inert' solvents. On the assumption that only monomer–dimer equilibria need be considered, i.r. and other non-calorimetric measurements have led to many equilibrium constants for dimerization reactions in dilute solutions. Although some ΔH° values have been calculated by way of the van't Hoff equations (4)—(6), uncertainties appear to be large and there are substantial disagreements between results of different investigators.

Woolley *et al.*[61,62] have measured enthalpies of dilution of acetic acid in carbon tetrachloride and benzene and of some chloro-substituted acetic and propionic acids in carbon tetrachloride. They have calculated K_c and ΔH° values for dimerization reactions, with K_c values in reasonable agreement with those derived from earlier non-calorimetric investigations. The ΔH° values from these calorimetric investigations[61,62] are probably more accurate than those derived from equilibrium constants at several temperatures.

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Although it appears well established that monomers and dimers are the principal carboxylic acid species in dilute solutions, recent i.r. measurements and 'factor analysis' of the results by Bulmer and Shurvell[63] provide evidence that larger associated species are also present in all but the most dilute solutions. At present we have no quantitative information about these larger species, nor do we know how much error in thermodynamic properties of dimerization reactions results from failure to consider trimers, *etc.*

Factor analysis of results from spectroscopic investigations of self-association of ROH compounds (discussed below) is certainly desirable, and it may be that this important data treatment technique can be extended to apply to enthalpies of dilution.

Numerous investigations have provided convincing evidence for self-association of ROH compounds in various solvents, but there is widespread disagreement as to identification of predominant associated species, and few reliable data for well-defined association reactions are available. We now discuss phenol as typical of this class of compounds.

Many, but not all, of the i.r. and other spectroscopic investigations (see refs. 17 and 64 for a review) of phenol in various 'inert' solvents indicate almost conclusively that self-associated species larger than dimers are important even in dilute solutions. This conclusion is supported by data for the distribution of phenol between water and carbon tetrachloride.[17] Finally, enthalpies of dilution[17,64] also show that there must be self-associated species larger than dimers. Calculations[17,64] based on these calorimetric results indicate that trimers may be the predominant self-associated species in most solutions [$n = 3$ in the reaction represented by (38)], a result that is in agreement with some spectroscopic investigations. But a recent thorough i.r. investigation by Whetsel and Lady[65] suggests that dimerization and stepwise further self-association to various n-mers is the most realistic representation. Although there seems to be no reason to question the reliability of either the calorimetric or spectroscopic experimental results, it is obvious that one or both of the conflicting interpretations[17,64,65] must be in error. Possible means of resolving present uncertainties have been suggested.[64]

7 Conclusions

Investigations based on methods discussed in this paper have shown that it is possible to obtain ΔH° or in some cases *both* K and ΔH° for reactions of type represented by equations (1) and (8) from results of calorimetric measurements at a single temperature. The most important general advantage of the calorimetric methods isthat they permit evaluation of ΔH° without recourse to the van't Hoff equation and the uncertainties derived from differentiation of experimental results. It is probable that the 'best' thermodynamic functions for

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reactions of type (1) and (8) can be obtained as a result of combination of equilibrium constants determined by spectroscopic (or other non-calorimetric) means with calorimetric enthalpies. For more complicated reaction schemes such as those which involve complexes represented by AB_2 or A_n, it is generally impracticable to obtain all of the desired thermodynamic functions from only one kind of measurement. In these cases it is especially advantageous to combine calorimetric results with those derived from spectroscopic measurements, vapour pressures, or other non-calorimetric measurements.

We are pleased to acknowledge the assistance of the Donors of the Petroleum Research Fund, administered by the American Chemical Society, the National Research Council of Canada, and the Research Committee of the New Zealand Grants Committee.