BASIC PROPERTIES OF ANIONS IN MIXED SOLVENTS

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by

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DEDICATION

To my parents, with thanks.

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ABSTRACT

The thesis is primarily concerned with the effect of solvent on the kinetic and thermodynamic properties of anions. The solvent systems used are mixtures of water with dimethylsulphoxide, trifluoroethanol and acetonitrile.

There are two main sections of the thesis. The first deals with thermodynamic aspects of the work and includes values for the free energies of transfer of silver hydroxide and silver fluoride from water to the mixed solvents together with the pK_a values of many acids in the solvent mixtures. (It is shown that the fluoride ion-selective electrode is well suited for use in all the solvent systems studied). Variations in the thermodynamic parameters are discussed in terms of changes in anionic solvation.

The second section is concerned with the effect of solvent on the rates of several proton-transfer reactions involving carbon acids and anion bases. Rate constants for such reactions were found to be very sensitive to solvent composition and, in particular, increased rapidly with the DMSO content of DMSO-H_2O mixtures. The corresponding equilibrium constants were often largely independent of solvent composition. The rates and equilibria are discussed in terms of the Brönsted relationship and it is shown that apparent anomalies in Brönsted coefficients for several of the reactions studied disappear upon transfer to DMSO-H_2O mixtures of sufficiently high DMSO concentration. It is argued that anomalous values of Brönsted coefficients in water result from specific solvation effects.

The primary kinetic hydrogen isotope effects (i.e. k^H/k^D) for several of the above reactions were also studied. Comparisons of k^H/k^D with ΔG^o of reaction lead to the conclusion that variations in the latter are not solely responsible for the observed changes in the isotope effects; this is discussed in terms of solvent involvement in the proton/deuteron-transfer reaction.

A section dealing with the mechanism of the observed oxidation of dimethylsulphoxide by bromine in aqueous solution is also included.

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GENERAL INTRODUCTION

It has long been known that the chemical properties of species in solution are dependent upon the nature of the solvent[1]. This is particularly true of ionic species and has stimulated much research into the properties of electrolyte solutions in non-aqueous solvents[2,3,4]. More recently, interest has turned to the use of mixed solvent systems and it is with such media that this thesis is concerned.

Two important classes of solvents are the so-called 'protic' and 'dipolar aprotic' solvents. The former possess relatively acidic protons and, as such, are able to take part in hydrogen bonding with suitable species (examples of such solvents include water, methanol and formamide). The latter are highly polar in nature and are incapable of donating hydrogen atoms which are labile enough to form hydrogen bonds (examples are acetone, acetonitrile and dimethyl sulphoxide). It should be noted that extensive ion aggregation occurs in any medium whose dielectric constant is insufficiently high and, because of this, only solvents with dielectric constants greater than about fifteen are differentiated in this way[2].

Ionic solvation can vary greatly between these two classes of solvents. Thus, the estimated values of the solvation energies of anionic species in protic solvents are greater than those for cations of comparable size[5]. This has been attributed to the extra stabilisation of the anion by hydrogen bonding with the solvent[2]. However, in aprotic dipolar solvents the order is reversed and the corresponding figures are less for anions than for cations of similar size[6]. These effects give rise to large, and often spectacular, differences in the reactivities of ions, and especially of anions, in different solvent systems. The work in this thesis is primarily concerned with the effect of solvent composition on the activity of anions - as reflected in the values of the free energy of transfer of silver salts from water to various mixed solvent systems, the variation with solvent composition of the acid strength (pK_a) of various species and the variation with solvent of the rates of several proton transfer reactions. Because such variations are particularly marked when we compare data obtained in water with those available using dimethyl sulphoxide[2] (hereafter referred to as DMSO), most of the work undertaken has involved the use of water-DMSO mixtures. However, to show that the changes observed are not specifically confined to this solvent system, data have also been obtained using mixtures of water with the aprotic, dipolar acetonitrile and, for comparison purposes, with the protic trifluoroethanol.

The thesis itself is divided into two main sections. The first deals with the effect of solvent on the thermodynamic properties of selected species, whilst the second is concerned with variations in kinetic parameters of various proton-transfer reactions due to the changes in solvent composition. The latter section includes a discussion on relationships between rates and equilibria in the solvent mixtures and also the variations observed in the kinetic hydrogen isotope effects (k_H/k_D) for several of the reactions studied. A section dealing with the mechanism of the observed oxidation of DMSO by bromine is also included.

SECTION I

THERMODYNAMIC PARAMETERS IN MIXED SOLVENT SYSTEMS

PART A

THERMODYNAMICS OF ION SOLVATION

I. INTRODUCTION

Anions behave as much stronger bases in dipolar aprotic solvents than in protic solvents. Because many organic reactions involve the base catalysed breakage of a carbon-hydrogen bond as the rate determining step, the rates of many such reactions can be greatly increased by the use of dipolar aprotic, rather than protic, solvents[7]. Such increases in rate are particularly striking when the anionic species involved are small, with high charge density (e.g. Cl^-, Br^-, N_3^- and, to a lesser extent I^-).

Changes in the free energy of any species on transfer from one medium to another are often expressed in terms of the free energy of transfer of the species (ΔG_{Tr}) between solutions of equal concentrations of species in the two solvents. Values for the transfer of anions from water to dipolar aprotic solvents clearly show that the small anions with a high charge density are most affected by the change[4]. This is in agreement with kinetic data.

Using size and charge density as suitable criteria, the two anions most likely to show large variations in behaviour with solvent, are fluoride and hydroxide. However, data for these two species are very limited. The present part of the thesis presents additional free energy data for these two anions. The few results which are already available would seem to agree with predictions. Thus, the estimated free energies of transfer of fluoride ion from water to dimethylformamide and propylene carbonate[9] are 3-4 kcal mol^{-1} larger than the corresponding figures for chloride ion. Also, the work by Das and Kundu on the autoprotolysis constant of water in DMSO and DMSO-water mixtures[10] would suggest that the free energy of hydroxide ion rises very rapidly with increasing DMSO content of the solvent.

In practice the free energies of transfer of ionic species can only be obtained from data relating to the transfer energies of whole electrolytes. This gives the sum of the free energies of transfer of each of the individual ions involved - to obtain the value for each ion then requires the use of one extrathermodynamic assumption. There are several assumptions which are commonly made; nearly all result in comparable values of ΔG_{Tr} for individual ions[11].

In this study the free energies of transfer of silver fluoride and silver hydroxide between water and several water-DMSO mixtures have been obtained. (For comparison purposes, the effect of using trifluoroethanol rather than DMSO in the mixtures has also been studied.) The values of ΔG_{Tr} so obtained can be compared with those already known for several other silver salts and because the cation is constant throughout the series, such comparisons reflect changes in anion solvation only and thus eliminate the need for any extrathermodynamic assumptions.

2. BASIC THEORY

The chemical potential, μ, of any species in a solvent X at a fixed temperature and pressure can be defined by the equation:

$$\mu_X = \mu_X^o + RT \ln a_X \qquad 1.1$$

where μ_X^o is the standard chemical potential of the species (at the same temperature and pressure) in solvent X, R the universal gas constant, T the absolute temperature and a_X the activity of the species in solvent X. The activity is usually expressed in terms of the concentration of the species multiplied by an activity coefficient, γ. Thus:

$$\mu_X = \mu_X^o + RT \ln [\text{species}]_X \gamma_X \qquad 1.2$$

As the solution becomes sufficiently dilute solute-solute interactions decrease and $\gamma_X \rightarrow 1$. The standard state to which μ_X^o refers is a hypothetically ideal solution of unit concentration of the species concerned, at the necessary temperature and pressure.

Any change in the solvent will cause a change in the chemical potential of the species. Such changes are most easily represented by the free energy of transfer, ΔG_{Tr}, of the species concerned. This quantity refers to the free energy change for the following process at a temperature T and one atmosphere pressure:-

$$\begin{array}{lcl} \text{1 mole of species from} & & \text{1 mole of species in a} \\ \text{a hypothetically ideal} & \xrightarrow{\Delta G_{Tr}} & \text{hypothetically ideal} \\ \text{solution of unit species} & & \text{solution of unit species} \\ \text{concentration in} & & \text{concentration in} \\ \text{solvent X} & & \text{solvent Y} \end{array}$$

where Y is the new solvent system.

If the standard state of the species in any solvent is taken as the hypothetically ideal solution of unit species concentration in that solvent, (as implied by equation 1.2) then:

$$\Delta G_{Tr} = \mu^{o}_{Y} - \mu^{o}_{X} \qquad 1.3$$

where μ^{o}_{Y} and μ^{o}_{X} refer to the appropriate standard chemical potentials in solvents Y and X respectively. Using this approach, ΔG_{Tr} actually represents the standard free energy change for the process.

However, if the standard state is maintained as the hypothetically ideal solution of unit species concentration in solvent X, irrespective of the actual solvent used, then:

$$\Delta G_{Tr} = (\mu^{o}_{X} + RT \ln {}^{X}\gamma^{Y}) - \mu^{o}_{X}$$

$$\text{i.e.} \quad \Delta G_{Tr} = RT \ln {}^{X}\gamma^{Y} \qquad 1.4$$

where ${}^{X}\gamma^{Y}$ is described either as a solvent activity coefficient[2], medium effect[12], distribution coefficient[13] or degenerate activity coefficient[14].

It is easily seen that:

$$\Delta G_{Tr} = \mu^{o}_{Y} - \mu^{o}_{X} = RT \ln {}^{X}\gamma^{Y} \qquad 1.5$$

In most of the present work the first approach is used and this being the case, the free energies of transfer quoted represent standard free energy changes. The molar concentration scale is used throughout, the values of ΔG_{Tr} being dependent upon the concentration scale used.

3. FREE ENERGY DATA FOR SILVER FLUORIDE

(a) Theory of method

The free energy of transfer for whole electrolytes, as well as neutral molecules, can be determined by standard thermodynamic procedures using solubilities or E.M.F. values of suitable cells. The latter method was adopted for use with silver fluoride. The two half cells used comprised a fluoride ion selective electrode in contact with a solution of potassium fluoride and a silver electrode in contact with a solution of silver nitrate. The two were connected by a salt bridge.

The former half cell responds according to the equation:

$$E_{F^-} = E^{o}_{F^-} - \frac{RT}{F} \ln a_{F^-}$$

where $E^{o}_{F^-}$ is a constant for the half cell, R the gas constant, T the absolute temperature, F the Faraday constant and "a" the activity.

For our purposes the fluoride half cell is represented by LaF_3/F^- [15], although it should be noted that the actual cell process is not fully understood.

The silver half cell is represented by Ag^+/Ag with the half cell reaction being:

$$Ag^+ + e^- \rightleftharpoons Ag$$

Thus, for the overall cell $LaF_3/\bar{F}||Ag^+/Ag$ the corresponding Nernst equation is:

$$E = E^\circ + \frac{RT}{F} \ln a_{Ag^+} a_{\bar{F}} \qquad 1.6$$

where E is the observed cell EMF and E° the standard EMF.

Using our definition of the standard state of any species in solution as the hypothetically ideal solution of unit species concentration in the solvent used, then for two solvents X and Y:-

$$E_X = E_X^\circ + \frac{RT}{F} \ln[Ag^+]_X[F^-]_X(\gamma_{Ag^+})_X(\gamma_{F^-})_X$$

$$E_Y = E_Y^\circ + \frac{RT}{F} \ln[Ag^+]_Y[F^-]_Y (\gamma_{Ag^+})_Y(\gamma_{F^-})_Y$$

$$\therefore E_Y - E_X = (E_Y^\circ - E_X^\circ) + \frac{RT}{F}\left\{\ln\frac{[F^-]_Y[Ag^+]_Y}{[F^-]_X[Ag^+]_X} + \ln\frac{(\gamma_{Ag^+})_Y(\gamma_{F^-})_Y}{(\gamma Ag^+)_X(\gamma_{F^-})_X}\right\}$$

Making the concentrations of Ag^+ and F^- the same in the two half cells and remembering that the activity coefficient of an ion for a 1:1 electrolyte in solution can be represented by the mean ionic activity coefficient, $\gamma_\pm$, which for a given solvent is dependent (to a first approximation) only upon the ionic strength, then:

$$E_Y - E_X = (E^o_Y - E^o_X) + \frac{RT}{F}\left\{\ln\frac{[Ag^+]^2_Y}{[Ag^+]^2_X} + \ln\frac{(\gamma_\pm)^2_Y}{(\gamma_\pm)^2_X}\right\} \qquad 1.7$$

Now:

$$(\Delta G^o_X - \Delta G^o_Y) = (\Delta G_{Tr})_{Ag^+} + (\Delta G_{Tr})_{F^-}$$

$$= (\Delta G_{Tr})_{AgF}$$

and, because the change in free energy for the cell process is simply related to the observed EMF by $\Delta G = -nEF$ (n representing the number of electrons involved in the change which will be unity in this case), then:

$$(E^o_Y - E^o_X)F = (\Delta G_{Tr})_{AgF} \qquad 1.8$$

If we substitute equation 1.8 into 1.7:

$$(\Delta G_{Tr})_{AgF} = (E_Y - E_X)F - RT\left\{\ln\frac{[Ag^+]^2_Y}{[Ag^+]^2_X} + \ln\frac{(\gamma_\pm)^2_Y}{(\gamma_\pm)^2_X}\right\} \qquad 1.9$$

Equation 1.9 can now be used to obtain ΔG_{Tr} for silver fluoride because all the necessary parameters can be either measured or calculated.

It should be noted that two implicit assumptions concerning the ion selective electrode have been made in the derivation of equation 1.9, neither of which are obviously true. Thus, we have assumed that:

(a) the EMF produced by the fluoride ion selective electrode, when coupled with a reference electrode, changes by $\frac{RT}{F}$ volts for any ten-fold change in fluoride ion activity, irrespectively of solvent.

(b) that the constant $E^{o}_{F^-}$ is independent of solvent.

In actual fact, experimental data obtained in various solvent systems, collected together in appendix I, clearly show that, to within our experimental accuracy, both these assumptions are valid. In the work following, equation 1.9 is therefore used without further justification.

To solve equation 1.9 necessitates the evaluation of the mean ionic activity coefficients for the solutions used. The use of the semi-empirical equation of Davies[16] relating the ionic activity coefficients to ionic strength allows us to do this - thus, for singly charged ions:

$$\lg \gamma_{ion} (= \lg \gamma_{\pm}) = - \frac{AI^{\frac{1}{2}}}{1 + I^{\frac{1}{2}}} + 0.3\ AI \qquad 1.10$$

where A is a constant dependent only upon temperature and the solvent, and I is the ionic strength. The latter can be calculated from the formula $I = \frac{1}{2} \Sigma\, c_i z_i^{\,2}$ where z_i represents the charge on any ion i and c_i its concentration in solution. The constant A is defined by:

$$A = \frac{1.823 \times 10^6}{(\varepsilon T)^{3/2}}\ , \qquad \varepsilon \text{ is the}$$

dielectric constant of the medium and T the absolute temperature.

Thus, to use the Davies equation in the present work requires that the dielectric constants of the various solvent mixtures be known. In the case of DMSO-water mixtures these values are available[17] but this is not true of trifluoroethanol-water mixtures. In the latter case the dielectric constant of the solvent mixture has been assumed to be a linear function of mole fraction trifluoroethanol. The dielectric constant for the latter is taken as 26.7 [4].

As is usual in this kind of study the temperature used throughout the work is 25°C.

(b) Experimental

The fluoride ion selective electrode is basically similar in construction to the normal glass pH electrode except that the membrane consists of a single-crystal rare earth fluoride such as LaF_3, NdF_3 or PrF_3 [18]. The internal reference solution usually contains both fluoride and chloride ions. In the present study the model used was the Beckman Fluoride Selection Electrode (Model number 39600) which utilises a LaF_3 membrane.

The fluoride ion solutions were made up from BDH laboratory reagent grade potassium fluoride which had been dried by heating under vacuum for several days. The silver ion solutions were prepared from BDH "AnalaR" grade silver nitrate, whilst the salt bridge was made up from "AnalaR" potassium nitrate.

The solvents were purified as follows:
Koch-Light puriss grade DMSO was dried over molecular sieves and then fractionally distilled from calcium hydride under reduced pressure (typically, the boiling point is 86°C at 11 mm.Hg pressure). The middle fraction was retained for use and stored in a refrigerator.

Aldrich Chemical Company trifluoroethanol was dried over anhydrous sodium sulphate and fractionally distilled, the middle fraction again being collected for use (B.pt. 74-75°C, in agreement with the literature value).

Water was deionised and distilled from potassium permanganate. It was then boiled for thirty minutes and cooled under nitrogen before use.

All solvent mixtures were made up according to volume, any small change in total volume arising from the mixing of the two solvents being ignored.

The experimental procedure was the same for all solvent compositions. Thus, the fluoride electrode was placed into a beaker containing a solution of potassium fluoride of known concentration whilst the silver electrode (consisting of a suitable length of silver wire which was always cleaned with a mild abrasive before use) was placed into a beaker containing silver nitrate solution of identical concentration, and solvent composition, to the fluoride ion solution. Both beakers were suspended in a water bath which was maintained at 25(± 0.2)°C. Contact between beakers was made by means of a salt bridge of 0.1M potassium nitrate solution. Determination of the cell EMF with water as solvent used aqueous potassium nitrate in the salt bridge. However, for the DMSO-water and trifluoroethanol-water mixtures the bridge solvent contained appropriate mixtures of the solvents.

The EMF produced by the cell was measured using a Radiometer pH/mV meter, model number 26. The solutions were stirred occasionally and about half an hour was allowed for equilibrium to be established before final steady readings were taken (in actual fact steady values of the EMF were often reached within ten minutes). The determination of the cell EMF was performed in duplicate and the average value was taken; these duplicate values were almost always reproducible to within 1 mV.

(c) Results

The free energy of transfer of silver fluoride from water to any solvent mixture, s, can be obtained from equation 1.9 which, in this context, can be written as:

$$(\Delta G_{Tr})_{AgF} = (E_s - E_{aq})F - RT\left\{ \ln\frac{[Ag^+]_s^2}{[Ag^+]_{aq}^2} + \ln\frac{(\gamma_\pm)_s^2}{(\gamma_\pm)_{aq}^2} \right\}$$

The results are tabulated in tables 1.1 and 1.2. The value of RT is taken as 2478 joules mol^{-1} at 298°K and F as 96,496 coulombs mol^{-1}. Values of $\gamma_\pm$ are derived from the Davies equation using the values of A shown in the tables.

TABLE 1.1 - THE EVALUATION OF (ΔG_{Tr}) of AgF FROM WATER TO VARIOUS DMSO-WATER MIXTURES at 25°C

Vol.% DMSO in solvent	A Parameter for Davies Equation	$[Ag^+]$, $[F^-]$ (mol dm^{-3})	EMF of the cell $LaF_3\|F^-\|\|Ag^+\|Ag$ (mV)	Lg $\gamma_\pm$ [a]	ΔG_{Tr} [b] (kcal mol^{-1})
0	0.509	0.020	503	-0.061	0.00
50	0.584	0.010	559	-0.053	2.09
70	0.650	0.006	596	-0.049	3.54
80	0.695	0.004	627	-0.042	4.71
90	0.841	0.002	671	-0.036	6.53

a The ionic strength of any solution is equal to the concentration of Ag^+, F^- quoted

b Calculated from equation 1.9

TABLE 1.2 - THE EVALUATION of ΔG_{Tr} of AgF FROM WATER TO A SOLUTION OF 70% BY VOLUME TRIFLUOROETHANOL AND 30% WATER AT 25°C

Volume % Trifluoroethanol in solvent	$[Ag^+],[F^-]$ mol dm^{-3}	lg $\gamma_{\pm}$ [a]	EMF of the cell $LaF_3\|F^-\|\|Ag^+\|Ag$ (mV)	ΔG_{Tr} [b] (kcal mol^{-1})
0	0.02	-0.061	503	0.00
70	0.006	-0.055	591	3.43

a The ionic strength of any solution is the concentration of Ag^+,F^- quoted. The value of A in the Davies equation is taken as 0.509 for water and 0.770 for the mixture

b Calculated from equation 1.9

The free energies of transfer so obtained have units of joules mol^{-1} - however, to facilitate comparison with data already available, values have been converted into kcal mol^{-1} (1 kcal = 4,184 joules).

The solvent composition is quoted in terms of the percentage, by volume, of each component - Appendix II shows how this correlates with the molar percentage of each component for DMSO-water mixtures. The EMF quoted refers to the cell $LaF_3|\bar{F}||Ag^+|Ag$.

Discussion of the results is deferred until the free energy data for silver hydroxide have been presented.

4. FREE ENERGY DATA FOR SILVER HYDROXIDE

(a) Theory of method

Only values for transfer from water to DMSO-water mixtures have been obtained for silver hydroxide. The method employed utilises data already available in the literature and, in principle at least, is very simple. Thus, Madec[19] has already obtained accurate values of the solvent activity coefficient ratios $({}^{aq}\gamma^{s}_{Ag^+}/{}^{aq}\gamma^{s}_{H^+})$ for the transfer of Ag^+ and H^+ from water to DMSO-water mixtures (s) at 25°C. Data are available for the complete range of solvent mixtures and therefore, using equation 1.4, we can calculate the corresponding differences in the free energies of transfer of the ions from water to any required DMSO-water mixture. Thus

$$(\Delta G_{Tr})_{Ag^+} - (\Delta G_{Tr})_{H^+} = RT\{\ln {}^{aq}\gamma^{s}_{Ag^+} - \ln {}^{aq}\gamma^{s}_{H^+}\} \qquad 1.11$$

Using the results of Fiordiponti et al[20] we can obtain figures for the corresponding combined free energies of transfer of H^+ and OH^-. This, together with the fact that the free energy of transfer of an electrolyte is the sum of the free energies of transfer of the constituent ions, enables us to calculate the free energy of transfer of silver hydroxide from water to the mixture at 25°C by using equation 1.12.

$$(\Delta G_{Tr})_{AgOH} = \left\{ (\Delta G_{Tr})_{Ag^+} - (\Delta G_{Tr})_{H^+} \right\} + \left\{ (\Delta G_{Tr})_{OH^-} + (\Delta G_{Tr})_{H^+} \right\} \qquad 1.12$$

Before using the results of Fiordiponti et al to supply the free energy data required by 1.12, it is necessary to define the standard states of the solvent components themselves (in this case, water and DMSO). This, at temperature T and pressure P, is taken as the pure liquid concerned and consequently the activity of water in any mixture of water and DMSO at temperature T and pressure P can be obtained from equation 1.13:-

$$(a_{H_2O})_s \approx \frac{P_m}{P_o} = (X_{H_2O})_s (f_{H_2O})_s \qquad 1.13$$

where $(X_{H_2O})_s$ represents the mole fraction in solvent mixture s, $(f_{H_2O})_s$ the corresponding activity coefficient, P_m the partial vapour pressure of water above the mixture and P_o the vapour pressure of pure water (both liquids being under total pressure P and at temperature T). Clearly as $X_{H_2O} \to 1$, $f_{H_2O} \to 1$.

Using the standard state of water as defined above and that of an electrolyte species in the mixture, at the same temperature and pressure, as the hypothetically ideal solution of unit species molality in that solvent, Fiordiponti has calculated values for the equilibrium constant (defined by equation 1.14) for the dissociation of water ($H_2O \rightleftharpoons H^+ + OH^-$) in various DMSO-water mixtures at 25°C and one atmosphere pressure.

$$K_{a/a} = \frac{[OH^-]_s[H^+]_s(\gamma_\pm)_s^2}{(a_{H_2O})_s} \qquad 1.14$$

The activity of water is obtained by substitution of the appropriate vapour pressure data into equation 1.13. Using this value in equation 1.14 assumes that the activity of water is independent of ionic strength - this is true at the low ionic strengths ($\simeq 10^{-3}$M) used in Fiordiponti's experiments.

The values of $K_{a/a}$ so obtained allow us to calculate the standard free energy change for the dissociation of water in the corresponding DMSO-water mixture, by utilisation of equation 1.15

$$(\Delta G^o)_s = -RT \ln(K_{a/a})_s \qquad 1.15$$

where T = 298°K.
Because the standard state for water is independent of the solvent system, s, it follows that:

$$(\Delta G^o)_s - (\Delta G^o)_{aq} = (\Delta G_{tr})_{H+} + (\Delta G_{Tr})_{OH-}$$

and substitution of equation 1.15 into the above results

TABLE 1.3 - THE EVALUATION of ΔG_{Tr} OF SILVER HYDROXIDE FROM WATER TO VARIOUS DMSO-WATER MIXTURES AT 25°C

Volume % DMSO in solvent	$-\lg K_a/a$[a]	$(\Delta G_{Tr})_{OH^-} + (\Delta G_{Tr})_{H^+}$[b] kcal mol^{-1}	$\lg {}^{aq}\gamma^{s}_{Ag^+} - \lg {}^{aq}\gamma^{s}_{H^+}$[c]	$(\Delta G_{Tr})_{Ag^+} - (\Delta G_{Tr})_{H^+}$[d] kcal mol^{-1}	$(\Delta G_{Tr})_{AgOH}$[e] kcal mol^{-1}
0	14.00	0.00	0.00	0.00	0.00
50	15.64	2.23	+ 0.10	+ 0.14	2.37
70	17.22	4.38	- 0.45	- 0.61	3.77
80	18.20	5.72	- 0.80	- 1.09	4.63
90	19.80	7.90	- 1.38	- 1.80	6.10

a Reference 20
b From equation 1.16
c Reference 19
d From equation 1.1[illegible]
e From equation 1.12

in the required expression:

$$(\Delta G_{Tr})_{H^+} + (\Delta G_{Tr})_{OH^-} = -RT \ln\frac{(K_{a/a})_s}{(K_{a/a})_{aq}} \qquad 1.16$$

Equations 1.16 and 1.11 can now supply the parameters required by equation 1.12 for a successful evaluation of the free energy of transfer of silver hydroxide at 25°C from water to solvent mixture s.

(b) Results

The value of RT at 298°K is again taken as 2478 joules mol^{-1} and the final values of ΔG_{Tr} for silver hydroxide have been converted into units of kcal mol^{-1} (1 kcal = 4,184 joules).

The data are presented in table 1.3.

5. DISCUSSION

The values of the free energies of transfer of several silver salts from water to DMSO-water mixtures are collected together in table 1.4, together with Madec's estimated values for the silver ion itself[21].

Clearly, for any solvent mixture, the ΔG_{Tr} values for the various silver salts represent relative free energies of transfer of the corresponding anions. It is evident, therefore, that the free energies of transfer of small anions, with high charge density, are

TABLE 1.4 - VALUES FOR FREE ENERGY OF TRANSFER OF VARIOUS SILVER SALTS FROM WATER TO DMSO-WATER MIXTURES AT 25°C

Vol.% of DMSO in solvent	$(\Delta G_{Tr})_{AgX}$ (kcal mol^{-1})							Estimated [c] $(\Delta G_{Tr})_{Ag^+}$ (kcal mol^{-1})
	X = F^- [a]	OH^- [a]	Cl^- [b]	Br^- [b]	I^- [b]	AcO^- [b]	N_3^- [b]	
50	2.09	2.37	0.0		-1.7	1.8	-0.7	-2.4
70	3.54	3.77	0.2		-3.4	2.4	-1.3	-5.6
80	4.71	4.63	0.3		-4.2	2.9	-1.6	-7.0
90	6.53	6.10	0.5		-5.2	3.7	-1.9	-8.8
100			1.2	-1.6	-5.2	7.0	-1.0	-10.6

a Values obtained from present study

b Values (in kcal mol^{-1}) obtained from reference 21 where

$$\Delta G_{Tr} = \frac{RT}{4184} \ln {}^{H_2O}\gamma^{s}_{AgX}$$

c Values (in kcal mol^{-1}) obtained from reference 21 where

$$(\Delta G_{Tr})_{Ag^+} = \frac{RT}{4184} \ln {}^{H_2O}\gamma^{s}_{Ag^+}$$

indeed much greater than those of larger anions with lower charge density - this effect, as anticipated, is particularly striking in the case of hydroxide and fluoride anions. Thus, the free energies of transfer of the anions from water to any given DMSO-water mixture are, in order of magnitude, OH^-, $F^- >> Cl^- > N_3^- > Br^- > I^-$. Also it would appear that the smaller the anion, the more steeply does the ΔG_{Tr} value rise with increasing DMSO content of the solvent. The estimated free energies of transfer of the silver ion indicate that an increase in anion free energy invariably occurs on addition of DMSO to an aqueous solution thereof.

All of the above trends are consistent with the suggestion that, in water, considerable stabilisation of the anions can result from hydrogen bonding with the solvent - in general the greater the charge density of the anion the greater the stabilisation. Addition of an aprotic dipolar solvent, such as DMSO, will cause a decrease in hydrogen bonding and an increase in anion free energy. The greater the charge density on the anion the greater the increase. (The very large increases in the free energy of acetate compared to chloride presumably reflects the considerable stabilisation of the carboxylate anion by hydrogen bonding in water.)

Trifluoroethanol possesses a more acidic proton than water[22] and, as such, it could be anticipated that anion solvation, with respect to water, should be enhanced.

The work of Parker et al[11] supports this viewpoint and presumably, therefore, the value for $(\Delta G_{Tr})_{AgF}$ from water to 70/30 trifluoroethanol-water of 3.43 kcal mol^{-1} reflects considerable cation destabilisation in this solvent. Work currently in progress in this laboratory should provide further data for silver salts in this solvent system.[23]

PART B

ACID STRENGTHS IN MIXED SOLVENTS

1. INTRODUCTION

Most of the remainder of this thesis is concerned with proton-transfer reactions in mixed solvent systems - and in particular with the effect of solvent on rates and equilibria of such reactions. Clearly then we require the relative acidities of the species involved.

Using the definitions proposed by Brönsted[24], an acid is any species having the tendency to lose a proton whereas a base is any species which has the tendency to gain a proton. The ionisation of any acid, HA, can be represented by:

$$HA \rightleftharpoons H^+ + A^-$$

where A^- is the "conjugate base" of HA and H^+ represents the proton. In solution the free energies of the various species will depend upon strengths of their interaction with solvent. The equation, however, does not require us to know the exact nature of the solvated species and is therefore applicable to any solvent we wish to use.

The equilibrium constant for the above process, K_a, at any temperature is defined by:

$$K_a = \frac{a_{H^+}\, a_{A^-}}{a_{HA}}$$

and, converting to the more familiar pK_a scale:

$$pK_a = -\lg \; K_a = -\lg \frac{a_{H^+}\, a_{A^-}}{a_{HA}} \qquad 1.17$$

where 'a' represents activity. These values of pK_a represent the relative acidities of the species concerned - the lower the pK_a the further to the right does the equilibrium lie and the more acidic is the compound.

If we define the standard state of any species in any solvent as the hypothetically ideal solution of unit concentration using that solvent, we can justifiably discuss variations with solvent composition of the pK_a of a compound. Thus:

$$(pK_a)_{aq} = -\lg \frac{[H^+]_{aq}[A^-]_{aq}(\gamma_\pm)^2_{aq}}{[HA]_{aq}(\gamma_{HA})_{aq}} \qquad 1.18$$

$$(pK_a)_s = -\lg \frac{[H^+]_s[A^-]_s(\gamma_\pm)^2_s}{[HA]_s(\gamma_{HA})_s} \qquad 1.19$$

where s, as usual, represents any new solvent or solvent mixture.

Remembering that $\Delta G^o = -RT\ln K_a$, then it can easily be shown that the difference between the pK_a of the acid in solvent s, and that in water is related to the free energies of transfer (from water to solvent s) of the components, by the equation:

$$(pK_a)_s - (pK_a)_{aq} = \frac{(\Delta G_{Tr})_{H^+} + (\Delta G_{Tr})_{A^-} - (\Delta G_{Tr})_{HA}}{2.303\ RT} \qquad 1.20$$

Thus, any variations with solvent of the pK_a of any acid reflect the changes in the solvation of the species involved and, utilising the results of the previous section of the thesis, the effect of anion solvation on overall pK_a values can be estimated.

2. THEORETICAL ASPECTS OF pK_a DETERMINATIONS

Many methods are available for the evaluation of pK_a's in aqueous solution[25,26]. However, once out of this medium we are more restricted in those which we can use (this is primarily due to lack of sufficient standard data in these solvent systems). In the present work most of the pK_a's have been obtained by electrochemical means, although restricted use has been made of the spectrophotometric technique. Each is now considered in turn.

(a) Electrochemical

Equations 1.18 and 1.19 show that the pK_a may be obtained from the equation:

$$pK_a = -\lg[H^+]\gamma_{\pm} - \lg\frac{[A^-]\gamma_{\pm}}{[HA]\gamma_{HA}}$$

We know that, using the standard states previously defined, $(-\lg[H^+]\gamma_{\pm})$ represents the pH of solution and if we use the approximation that, in dilute solution, $\gamma_{HA} = 1$

$$pK_a = pH - \lg\frac{[A^-]}{[HA]} - \lg\gamma_{\pm} \qquad 1.20$$

Thus, determination of the pH of a solution of known $[A^-]:[HA]$ ratio enables us to evaluate pK_a (the value of $\gamma_{\pm}$ can be obtained from the Davies equation, 1.10).

The pH of the solution is almost invariably obtained from a cell involving a glass (or hydrogen) electrode coupled with a reference electrode. Combinations used to date, for non-aqueous media, have included calomel-hydrogen[27], calomel-glass[28,29,30] and Ag,AgCl-glass[31]; where work has been duplicated comparable values of pK_a have been obtained. In the present study use was made of the calomel and glass electrodes. In aqueous solution the EMF of the cell, calomel||glass, can be obtained from the equation:

$$E_{aq} = (E^{o}_{aq} - E_{cal}) - 2.303\,\frac{RT}{F}\,pH_{aq}$$

where E^{o}_{aq} is the cell constant for the glass electrode half cell and E_{cal} the EMF of the standard calomel half cell. If we change to solvent s and refer our standard state to that solvent, as indicated by equations 1.18 and 1.19, then, utilising eqn. 1.8, we can write:

$$E_{s} = (E^{o}_{s} - E_{cal}) - 2.303\,\frac{RT}{F}\,pH_{s}$$

where $$E^{o}_{s} = E^{o}_{aq} + \frac{(\Delta G_{Tr})_{H^+}}{F} + \Delta E_{j}$$

ΔE_j represents any changes in junction potentials due to change in solvent. The derivation of this equation assumes that the Nernst equation is still obeyed in all the solvent mixtures.*

---

* This has been shown to hold for very many solvent mixtures including most of those used in the present

work. Thus, it is true for all DMSO water mixtures[30,32] (providing that sufficient time is allowed for equilibrium to be established) and also for acetonitrile[33] and acetonitrile-water mixtures[23]. The present work confirms that the assumption is also valid for the trifluoroethanol-water mixtures used.

Clearly then, once out of water, the function we measure using a conventional pH meter is not an absolute value of pH_s but $pH_s + \frac{(\Delta G_{Tr})_{H+}}{2.303RT} + \frac{F\Delta E_j}{2.303RT}$, a function we shall call $(pH_s)_{obs}$. Thus, to obtain absolute values of pH in any solvent other than water we must evaluate this constant "correction factor" which will now be referred to as "c". This is easily achieved by measurement of $(pH_s)_{obs}$ for known concentrations of H^+. Thus:

$$(pH_s)_{obs} = pH_s + c = -lg[H^+]_s - lg(\gamma_\pm)_s + c$$

and so $$c = (pH_s)_{obs} + lg[H^+]_s + lg(\gamma_\pm)_s \qquad 1.21$$

Once the correction term is known, all the parameters required for the solution of equation 1.20 can be either measured experimentally or calculated and the pK_a obtained.

In practice the pK_a values are obtained at several different buffer ratios (i.e. $[A^-]:[HA]$ values) and the mean value evaluated. The most convenient way to obtain such data is to titrate a known concentration of acid, HA, with a standard solution of hydroxide MOH (all solutions having the same solvent composition, of course) and to record the pH values at several points during the titration. The equations defining such a system at any point in the titration are:-

$$[HA] + [A^-] = a \qquad (i)$$

$$[M^+] = b \qquad (ii)$$

$$[A^-] + [OH^-] = [M^+] + [H^+] \text{ (charge neutrality)} \qquad (iii)$$

$$K_a = \frac{[H^+][A^-]\gamma_\pm^2}{[HA]} \qquad (iv)$$

$$K_w = [H^+][OH^-]\gamma_\pm^2 \qquad (v)$$

Combining (ii) and (iii) gives:

$$[A^-] = b + [H^+] - [OH^-] \qquad (vi)$$

Thus, using solutions of sufficient concentration, gives us the condition that $|[H^+] - [OH^-]| << b$ and so:

$$\frac{[A^-]}{[HA]} = \frac{b}{a - b}$$

This, combined with the measured pH, allows us to calculate pK_a from equation 1.20. The inequality above can be tested using appropriate values of pH and K_w, and holds for nearly all the work carried out in this thesis.

(b) Spectrophotometric

The determination of ionization constants by u.v. or visible spectrophotometry depends upon the direct determination of the ratio of molecular species (neutral molecule) to ionized species in a non-absorbing buffer solution. Thus, if we add a small quantity of HA to a solution of buffer (whose constituent acid is represented by HB) then the following equilibrium constants must hold:

$$(K_a)_{HB} = \frac{[H^+][B^-]}{[HB]} \gamma_{\pm}^{2}$$

$$(K_a)_{HA} = \frac{[H^+][A^-]}{[HA]} \gamma_{\pm}^{2}$$

Combining these two equations and solving for H^+ gives:

$$(K_a)_{HB} = (K_a)_{HA} \frac{[HB]}{[B^-]} \frac{[A^-]}{[HA]}$$

and so:

$$(pK_a)_{HA} = (pK_a)_{HB} + \lg \frac{[B^-]}{[HB]} - \lg \frac{[A^-]}{[HA]} \qquad 1.22$$

Clearly, the pK_a of HB in the particular solvent system used must be known and thus, the method is of only limited applicability to mixed solvent systems. Usually HB is chosen so that it has a pK_a of similar magnitude to that expected for HA. Thus, if we use buffer ratios (i.e. $[B^-]/[HB]$) close to unity, HA will be approximately half neutralised at equilibrium and so errors in the determined pK_a value will be minimised.

If equal volumes of HA are syringed into matched cells containing hydrochloric acid (such that $-\lg[H^+] < pK_{HA} - 2$), sodium hydroxide (such that $-\lg[H^+] > pK_{HA} + 2$) and a buffered solution of known buffer ratio then the resulting solutions contain HA, A^- and a HA/A^- mixture respectively. The absorbance of each solution at a wavelength λ (chosen such that only HA and A^- absorb) is obtained from the usual Beer-Lambert law and is given by:

$$A_{HA} = \varepsilon_{HA}\, m$$

$$A_{A^-} = \varepsilon_{A^-}\, m$$

$$A = \varepsilon_{HA}[HA] + \varepsilon_{A^-}[A^-] \qquad (\text{where } [HA] + [A^-] = m)$$

'ε' represents the extinction coefficient of the species concerned at wavelength λ and m is the same concentration of HA, A^- or mixture.

Thus:

$$\frac{A_{A^-} - A}{A - A_{HA}} = \frac{\varepsilon_{A^-}(m - [A^-]) - \varepsilon_{HA}[HA]}{\varepsilon_{HA}([HA] - m) + \varepsilon_{A^-}[A^-]} = \frac{[HA]}{[A^-]}$$

and substitution into equation 1.22 gives:

$$(pK_a)_{HA} = (pK_a)_{HB} + \lg\frac{[B^-]}{[HB]} + \lg\frac{A_{A^-} - A}{A - A_{HA}} \qquad 1.23$$

Thus, if we measure the required absorbances, equation 1.23 can be used to give the pK_a required. It is usual to use several different buffer ratios and calculate the mean value for the pK_a.

The wavelength λ (the so-called "analytical wavelength") is chosen such that the difference between the absorbances of the two species is at maximum (i.e. the absorbance is due primarily to either HA or A^-). The rough value of the pK_a required for the accurate procedure outlined in this section is obtained from separate trial experiments using different buffer systems of varying pH.

The Beer-Lambert law was found to hold in all solvent systems used.

3. EXPERIMENTAL

Only basic experimental points will be discussed here - details specific to individual compounds are presented in the results section dealing with the species concerned. All mixed solvent solutions were prepared by dilution of suitable aqueous stock solutions with the organic solvent in question.

(a) Electrochemical

An Activion Ltd. (pH 0-14) glass, calomel combination electrode and Radiometer (Type 26) pH meter were used throughout this work.

At the start of a day's work, and before any potentiometric work was undertaken, the pH scale was calibrated using standard aqueous buffer solutions[34]. Normally, it was sufficient to use solutions of 0.05M potassium hydrogen phthalate (pH 4.008 at 25°C) and 0.01M borax (pH 9.180 at 25°C) - however, in those cases where the absolute pH of solution to be recorded was high enough to lie outside this range, the scale was calibrated using 0.01M borax and a solution 0.132M in NaOH and 0.08M in KCl (pH 13.00 at 25°C).

After the pH scale had been standardised the electrode was soaked for about two hours in the particular solvent system to be used. The first task was then to determine the solvent correction factor, c. Thus, the observed, steady pH values, in the required mixed solvent, for at least three perchloric acid solutions of differing concentrations (commonly $7 \times 10^{-2} - 7 \times 10^{-4}$M) were recorded and the values for the solvent correction factor determined from equation 1.21. The mean value was then calculated and used for all work in that solvent system. As might be anticipated, this correction factor was found to be independent of the particular combination electrode used. The determination of c was, of course, only performed at the start of work in any particular solvent system. Thereafter, once the electrode had been soaked in the solvent (when not in use it was suspended in a beaker of distilled water) the required titration was immediately begun.

The assembly used for the above, and all subsequent pH determinations, consisted of a small beaker surrounded by a water jacket through which was circulated water at 25 (± 0.1)°C. The vessel was fitted with a tight rubber bung through which passed the combination electrode and an inlet tube for nitrogen gas. All solutions were continuously stirred by means of a small magnetic stirrer and nitrogen bubbled through to minimise carbon dioxide absorption.

For the potentiometric titration itself, the required volume of HA solution was pipetted into the cell followed by the required alliquot(s) of standard base solution (the solvent composition of both solutions being the same). Sodium hydroxide was used throughout this work, any interference of the sodium ion with the glass membrane being negligible for this model of glass electrode. The final steady pH reading after addition of each aliquot of base was taken - the time required for equilibration varied tremendously with the nature of the compound under study, and was particularly long in those cases where proton transfer from HA to $\bar{O}H$ was slow (e.g. 2-nitropropane, where at least one and a half hours is required in aqueous solution). Where transfer was fast, at least thirty minutes was allowed for temperature equilibration of the system.

When repeated, equivalent pH values were invariably reproducible to within 0.04 pH units. At the end of a day's work using any particular solvent mixture, the electrode was soaked in water for several hours and the pH scale rechecked using the buffer solutions employed for the original calibration. Very rarely did the drift in either value exceed 0.02 pH units. This is in agreement with previous observations of the behaviour of glass electrodes in non-aqueous media[35], and demonstrates the applicability of the glass electrode for use in such media.

(b) Spectrophotometric

The procedure for all solvent mixtures was the same. Solutions of 0.01M hydrochloric acid, 0.01M sodium hydroxide and the required buffer solutions (all having the same, requisite solvent composition) were prepared. Two cm^3 of each were pipetted into silica spectroscopic cells which were then placed into the spectrophotometer cell block - in this case the instrument was a Gilford 2400S u.v./visible spectrophotometer. The cell block can accommodate four cells and is thermostatted by water circulating from a water bath. The temperature used throughout the present work was 25 ($\pm$ 0.1)$^{\circ}$C.

The cells were thermostatted for forty minutes and then equal volumes (commonly, about 20 μl) of HA syringed into each. The cells were shaken, returned to the cell block and the steady absorbances recorded (as in the electrochemical work, the time for equilibration varied with the compound under study). When non-matching cells were used, the relevant correction factors had to be obtained - this was achieved very easily, just before the addition of HA. Thus, the absorbance of one of the cells in the cell block (it did not matter which) was corrected to zero at the wavelength, λ, to be used. The absorbances of the other cells relative to this zero were then measured - these figures represented correction factors which had to be added to, or subtracted from the measured absorbance once HA had been added and equilibrium reached. This, and all subsequent operations, relies upon the fact that none of the components in solution, at this stage in the procedure, absorb at this λ.

Usually the absorbances of HA/A^- mixtures in solutions of three different buffer ratios were determined and the values of pK_a in this solvent obtained from equation 1.23. The mean value could then be calculated.

4. MATERIALS

Solvents:	The solvents (water, DMSO, trifluoroethanol and acetonitrile) were purified by procedures already described in this thesis.
Perchloric Acid:	BDH AnalaR grade material was diluted with water and the accurate concentration determined by titration with standard sodium hydroxide solution.
Sodium Hydroxide, Acetic Acid: $NaOH, CH_3CO_2H$	Both were made up by appropriate dilution of standard BDH concentrated volumetric solutions (these being manufactured using AnalaR grade materials).
Monochloroacetic Acid: CH_2ClCO_2H	BDH AnalaR reagent was dissolved in water and, because of the hygroscopic nature of the solid, the exact concentration determined by titration with standard sodium hydroxide solution.
2-nitropropane: $(CH_3)_2CHNO_2$	Because of the method used in the kinetic study of this compound, material of very high purity was required. BDH reagent grade material was dried over calcium sulphate

and then twice fractionally distilled. The middle fraction was retained in each case (B.pt. 120°C, in agreement with the literature value). However, g.l.c. and n.m.r. analysis revealed small amounts of nitroethane and 1-nitropropane as impurities. Final purification of the sample was achieved by spinning band distillation (50% recovery rate) - g.l.c. analysis (Perkin-Elmer F11 instrument using a 15% Apiezon L on chromosorb 80/60 mesh column) showed no impurities, as did n.m.r. analysis (Perkin-Elmer R20 instrument).

<u>Nitroethane</u>: $CH_3CH_2NO_2$ BDH Reagent grade material was dried over anhydrous magnesium sulphate and fractionally distilled, the middle fraction being retained for use (B.pt. 114°C, Literature value 114°C). N.m.r. analysis revealed negligible proportions of proton containing impurities and g.l.c. (silicone on chromosorb 80-120 mesh and 15% apiezon L on chromosorb 80-60 mesh columns) showed $\geq$ 97% purity.

<u>1:1 Dinitroethane</u>: $CH_3CH(NO_2)_2$ This was prepared by the method of Kaplan and Shechter[36] which involves the oxidative nitration of the silver salt of the nitroethane anion. Nitroethane, sodium nitrite and silver nitrate react together, in basic solution, according to the scheme·

$$CH_3CH_2NO_2 + OH^- \rightleftharpoons [CH_3\bar{C}HNO_2] + H_2O$$

$$[CH_3\bar{C}HNO_2] + 2Ag^+ + NO^-_2 \rightarrow CH_3CH(NO_2)_2 + 2Ag$$

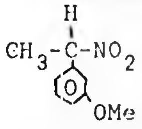
The crude dinitroethane so obtained was extracted with ether-benzene and recovered by distillation under reduced pressure. N.m.r. analysis revealed only minor amounts of impurities; potentiometric titration with standard hydroxide solution indicated a purity of 93% and did not reveal the presence of any other weak acid.

Ethyl-2-methylacetoacetate:

$CH_3.CO.CHMe.CO_2Et$

Aldrich Chemicals Ltd. technical grade material was fractionally distilled under reduced pressure (b.pt. 72°C at 4.5 mm Hg). The middle fraction was collected - however, subsequent g.l.c. analysis revealed ≥ 5% impurity. Thus, the sample was distilled under vacuum using the spinning band technique. G.l.c. and n.m.r. analysis showed the final product to be ≥ 99% pure; the only remaining impurity was identified as ethylacetoacetate. The relevance of this impurity to the kinetic behaviour of this compound is discussed in the corresponding section of the thesis.

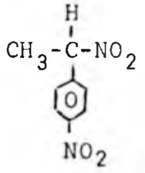
1-(3'-methoxyphenyl), 1-nitroethane:



This was prepared by the general method reported by Bordwell et al[37] for the preparation of 1-aryl-1-nitroethanes. Basically, it consists of the oxidation of the appropriate oxime with peroxytrifluoroacetic acid - this gives the required nitroethane derivative. The oxime itself is prepared by reaction of the correspondingly substituted acetophenone with hydroxylamine.

In this instance the crude nitroethane derivative so prepared (liquid) was passed down a silica column using 10% ether-hexane as solvent. The appropriate fraction was collected, the solvent removed and the final product obtained by distillation under reduced pressure. N.m.r. analysis of the product showed only minor traces of impurity and potentiometric titration with standard base solution (using 80% by volume DMSO in the solvents throughout) showed the compound to be 100% pure, to within experimental error.

1-(4'-nitrophenyl),1-nitroethane:



This was prepared by the same general method as the 3' methoxyphenyl compound. In this case, though, the crude product (solid) was recrystallised several times from carbon tetrachloride. N.m.r. analysis of the resulting crystals revealed no other proton containing impurities.

All other materials used were of AnalaR grade.

When not in use all organic materials were stored in a refrigerator.

5. RESULTS

The results are presented for each compound in turn and are all collected together, in the case of water-DMSO mixtures, in Table 1.18.

In those cases where the electrochemical method was used and titrations repeated, the value of pH reported in the tables for any particular point is the mean of the observed values. In the tables, the acid under study is always represented by HA, for convenience.

Mean ionic activity coefficients are calculated from the Davies equation (1.10). The necessary A parameters are known accurately for DMSO-H_2O mixtures only; in all other cases they are obtained assuming a linear relationship between dielectric constant and mole fraction of organic component of the solvent (see earlier in text).

The necessary solvent correction factors for the electrochemical method are presented in Table 1.5.

Acetic acid, monochloroacetic acid

The electrochemical method was used throughout. Results are given in Tables 1.6 and 1.7 for acetic acid and Table 1.8 for monochloroacetic acid.

2-nitropropane

The electrochemical method was used throughout. Only DMSO-H_2O mixtures were used; the time required for equilibration decreased markedly with increasing DMSO content of the solvent.

In 80% by volume DMSO steady pH values

were reached within 15 minutes; however, a steady upward drift started about 15 minutes later.* Thus, for this particular solvent composition it was not practicable to perform a standard potentiometric titration and, instead, the appropriate volumes of 2-nitropropane and sodium hydroxide were mixed and the steady pH value at this buffer ratio determined. This was repeated for several different buffer ratios and the pK_a calculated as usual. Results are recorded in Table 1.9.

Nitroethane

Results were obtained by the electrochemical method and are tabulated in Tables 1.10 and 1.11. In water and DMSO-water mixtures steady pH values were reached within 15 minutes and were steady for a further ten minutes or so; thereafter, upward drift ensued.* Thus, the same experimental procedure was adopted for use in these mixtures as in the case of 2-nitropropane in DMSO rich mixtures (see above).

1-1 Dinitroethane

Only DMSO-H_2O mixtures were used. The pK_a in water was roughly checked using electrochemical means and was found to agree with the literature value. The spectrophotometric technique was used for all the solvent mixtures, non-reproducible behaviour being observed in corresponding potentiometric work. Results are presented in Table 1.12.

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* The exact cause of these drifts are unknown, although they presumably reflect instability of the nitroalkane anion in these particular media.

Ethyl-2-methylacetoacetate

Again, only DMSO-H_2O mixtures were studied and the electrochemical method used throughout. Although no difficulties were encountered in the mixtures, problems did arise using water as solvent - thus, as soon as hydroxide was added to the substrate solution, the pH decreased steadily and no constant value was reached. (It is of interest to note that cleavage of ethylacetoacetate occurs in basic solution.[38])

In order to obtain a rough value of pK_a in water the acid solution was added to the reaction vessel and thermostatted; hydroxide was then added and a graph of pH against time was constructed. The pH at zero time was obtained by suitable extrapolation and used in the calculation of pK_a. This was repeated using several differing volumes of added hydroxide and the mean value of pK_a calculated. Agreement with the value, obtained by spectroscopic means[44], is excellent. Results are given in Tables 1.13 and 1.14.

1-(3' methoxyphenyl),1-nitroethane

Only DMSO-H_2O mixtures were studied. Spectrophotometry was used for aqueous solutions; the buffer system was made up by acidifying Tris(hydroxymethyl)methylamine (Tris for short) to give a Tris $\overset{+}{H}$ - Tris buffer*. All other work employed the electrochemical technique. Results are in Tables 1.15 and 1.16.

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* Because, in this case, the buffer acid is charged (i.e. Tris $\overset{+}{H}$) equation 1.22 must be amended and becomes:

$$(pK_a)_{HA} = (pK_a)_{Tris\ \overset{+}{H}} + \lg\frac{[Tris]}{[Tris\ \overset{+}{H}]} + \lg\frac{A_{A^-}-A}{A-A_{HA}} - 2\lg\gamma_{\pm} \qquad 1.24$$

As long as the amount of HA added is very small, lg $\gamma_{\pm}$ can be calculated from the Davies equation with I - [Tris $\overset{+}{H}$].

1-(4' nitrophenyl),1-nitroethane

Only DMSO-H_2O mixtures were used and the results are given in Table 1.17. The pK_a was determined by a potentiometric titration of an approximately 5 x 10^{-3}M solution of substrate with standard base. A graph of pH V volume NaOH added, was drawn and the end point determined. The pH at $[\bar{A}] = [HA]$ (i.e. 50% ionisation of HA) was then read off the graph and pK_a calculated.

TABLE 1.5 - SOLVENT CORRECTION FACTORS FOR THE GLASS ELECTRODE WHEN USED IN MIXED SOLVENT SYSTEMS AT 25°C

Solvent Composition [a]				
Vol.% H_2O	Vol.% DMSO	Vol.% TFE	Vol.% MeCN	Correction Factor, c.
50	50	-	-	0.65 ± 0.03
30	70	-	-	0.95 ± 0.04
20	80	-	-	1.14 ± 0.08
50	-	50	-	-0.40 ± 0.02
20	-	80	-	-1.66 ± 0.01
50	-	-	50	-0.24 ± 0.04

a Abbreviations:

Dimethylsulphoxide	(DMSO)
Trifluoroethanol	(TFE)
Acetonitrile	(MeCN)

b Calculated from equation 1.21, using standard perchloric acid solutions

TABLE 1.6 - THE pK_a OF ACETIC ACID IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol. 0.08M NaOH added to 20 cm^3 0.04M HA (cm^3)	$(pH_s)_{obs}$	$lg\frac{[A^-]}{[HA]}$	$lg\gamma_{\pm}$ [a]	pK_a [b]	Mean Value of pK_a
70	1.00	7.06	-0.95	-0.038	7.10	7.11 ± 0.06
	2.00	7.42	-0.60	-0.048	7.12	
	3.00	7.66	-0.37	-0.056	7.14	
	4.00	7.85	-0.17	-0.063	7.08	
	5.00	8.01	0.00	-0.068	7.13	
80	1.00	8.14	-0.95	-0.040	7.98	7.95 ± 0.12
	2.00	8.46	-0.60	-0.050	7.97	
	3.00	8.67	-0.37	-0.060	7.96	
	5.00	9.00	0.00	-0.074	7.93	
	7.00	9.34	0.37	-0.081	7.91	

a Calculated from Davies equation (1.10) using A parameters listed in Table 1.1 and taking I as $[Na^+]$ in solution

b Calculated from equation 1.20 where $(pH_s) = (pH_s)_{obs} - c$ The values of c are obtained from Table 1.5

TABLE 1.7 - THE pK_a OF ACETIC ACID IN TRIFLUOROETHANOL - AND ACETONITRILE-WATER MIXTURES AT 25°C

Vol.% TFE in solvent [a]	Vol.% MeCN in solvent [b]	Vol. 0.20M NaOH added to x cm^3 0.1M HA (cm^3)	$(pH_s)_{obs}$	$\lg\frac{[A^-]}{[HA]}$	$\lg\gamma_{\pm}$ [c]	pK_a [d]	Mean Value pK_a
50		1.00	4.69	-0.60	-0.07	5.76	5.78 ± 0.03
(x = 10.00)		3.00	5.47	0.18	-0.11	5.79	
		0.40	4.75	-0.60	-0.09	7.10	7.11 ± 0.03
80		0.80	5.16	-0.17	-0.11	7.10	
(x = 4.00)		1.10	5.53	0.18	-0.13	7.14	
		0.05	5.19	-0.60	-0.08	6.09	6.19 ± 0.10
	50	0.10	5.67	-0.17	-0.09	6.17	
	(x = 5.00)	0.15	6.16	0.18	-0.10	6.32	

a TFE = Trifluoroethanol

b MeCN = Acetonitrile

c Calculated from Davies equation taking A as 0.624, 0.930 and 0.640 for solutions containing, by volume, 50% TFE, 80% TFE and 50% MeCN respectively. $I = [Na^+]$.

d Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$. The values of c are obtained from Table 1.5.

TABLE 1.8 - THE pK_a of MONOCHLOROACETIC ACID IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol.0.08M NaOH added to 20 cm^3 0.04M HA (cm^3)	$(pH_s)_{obs}$	$\lg\frac{[A^-]}{[HA]}$	$\lg\gamma_{\pm}$ [a]	pK_a [b]	Mean Value pK_a
70	1.00	5.02	-0.95	-0.038	5.06	5.06 ± 0.05
	2.00	5.38	-0.60	-0.048	5.08	
	3.00	5.59	-0.37	-0.056	5.07	
	5.00	5.94	0.00	-0.068	5.06	
	7.00	6.30	0.37	-0.073	5.05	
80	1.00	5.66	-0.95	-0.04	5.51	5.56 ± 0.10
	2.00	6.04	-0.60	-0.05	5.55	
	3.00	6.29	-0.37	-0.06	5.58	
	5.00	6.64	0.00	-0.07	5.57	
	7.00	7.02	0.37	-0.08	5.59	

a Calculated from Davies equation (1.10) using A parameters listed in Table 1.1 and taking I as $[Na^+]$ in solution.

b Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$.

TABLE 1.9 - THE pK_a of 2-NITROPROPANE IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent		Vol. xM NaOH added to y cm^3 zM HA (cm^3)	$(pH_s)_{obs}$	$lg\frac{[A^-]}{[HA]}$	$lg\gamma_{\pm}$ [a]	pK_a [b]	Mean Value pK_a
70	x= 0.20 y= 20.00 z= 0.101	5.00	11.81	-0.009	-0.095	10.97	
		6.08	11.94	-0.165	-0.103	10.93	10.96 ± 0.08
		7.00	12.12	0.354	-0.104	10.94	
		9.00	12.96	0.913	-0.109	11.01	
80[c]	x= 0.10 y= 22.00 z= 0.10	9.96	13.05	-0.08	-0.108	12.10	
	x= 0.10 y= 4.40 z= 0.10	2.99	13.48	0.33	-0.123	12.14	12.07 ± 0.07
	x= 0.10 y= 10.90 z= 0.10	9.96	13.90	1.02	-0.132	11.87	

a Calculated from Davies equation (1.10) using A parameters listed in Table 1.1 and taking I as $[Na^+]$ in solution

b Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$

c Because of pH drift, the various HA/OH^- mixtures were made up separately and then pH recorded before the drift set in (see text)

TABLE 1.10 - THE pK_a OF NITROETHANE IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol. 0.20M[a] NaOH added to 20 cm³ xM HA		$(pH_s)_{obs}$	$\lg\frac{[A^-]}{[HA]}$	$\lg\gamma_{\pm}$ [b]	pK_a [c]	Mean Value pK_a
	x	(cm³)					
0	0.095	1.00	7.97	-0.93	-0.04	8.94	9.03 ± 0.07
	0.094	3.00	8.70	-0.33	-0.06	9.09	
	0.094	4.00	8.77	-0.14	-0.07	8.98	
	0.095	5.00	8.96	0.05	-0.08	8.99	
	0.095	7.00	9.51	0.45	-0.08	9.14	
50	0.095	1.00	9.84	-0.93	-0.05	10.17	10.19 ± 0.05
	0.099	3.00	10.41	-0.35	-0.07	10.18	
	0.098	6.00	10.97	0.20	-0.09	10.23	
70	0.097	2.00	11.94	-0.58	-0.07	11.64	11.67 ± 0.07
	0.096	4.00	12.41	-0.15	-0.08	11.69	
	0.093	2.00	12.01	-0.56	-0.07	11.70	
	0.108	1.00	11.60	-0.94	-0.06	11.65	

a Because of pH drift in solutions after standing for 30 minutes, each HA/OH^- mixture was made up separately and pH determined immediately (see text)

b From Davies equation (1.10) using A parameters from Table 1.1

c From equation 1.20 where $pH_s = (pH_s)_{obs} - c$

TABLE 1.11 - THE pK_a OF NITROETHANE IN TRIFLUOROETHANOL- AND ACETONITRILE-H_2O MIXTURES AT 25°C

Vol.% TFE in solvent[a]	Vol.% MeCN in solvent[b]	Vol. 0.20M NaOH added to x cm^3 yM HA			$(pH_s)_{obs}$	$lg\frac{[A^-]}{[HA]}$	$lg\gamma_{\pm}$[c]	pK_a[d]	Mean Value pK_a
		x	y	(cm^3)					
50		10.00	0.10	1.00	9.15	-0.60	-0.07	10.23	10.26 ± 0.04
		"	"	2.00	9.62	-0.17	-0.10	10.29	
		"	"	3.00	9.92	0.18	-0.11	10.25	
80		5.00	0.104	0.50	9.21	-0.60	-0.09	11.56	11.50 ± 0.06
		"	"	1.00	9.58	-0.17	-0.11	11.52	
		"	"	1.50	9.80	0.17	-0.14	11.43	
	50	4.00	0.093	0.04	9.81	-0.56	-0.08	10.69	10.63 ± 0.10
		"	"	0.08	10.21	-0.12	-0.09	10.66	
		"	"	0.12	10.42	-0.23	-0.10	10.53	

a TFE - Trifluoroethanol

b MeCN = Acetonitrile

c Calculated from Davies equation (1.10) using A parameters listed in Table 1.7. $I = [Na^+]$

d Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$ The values of c are shown in Table 1.5

TABLE 1.12 - THE pK_a of 1:1 DINITROETHANE IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Concentration[a] of species in solution before addition of HA ($mol.dm^{-3}$)	Absorbance of[b] cell contents after HA addition (at 385 nm)	lg (Buffer[c] ratio)	pK_a [d]	Mean Value pK_a
	0.01M KOH	1.22	-	-	
	0.01M HCl	0.00	-	-	
70	0.10M B, 0.10M HB	0.44	0.00	5.32	5.32 ± 0.05
	0.10M B, 0.05M HB	0.64	0.30	5.32	
	0.10M B, 0.025M HB	0.84	0.60	5.31	

a Volume of solution in each cell was 2 mls.
Abbreviations: B = monochloroacetate anion
HB = monochloroacetic acid

b 20 μl of a 0.01M stock solution of dinitroethane in DMSO were syringed into the cells. The absorbances quoted have been corrected for non-matching of the cells (see script)

c This refers to the ratio B/HB

d Obtained from equation 1.23, where the pK_a of monochloroacetic acid in an aqueous mixture with 70% by volume DMSO, is taken as 5.06 (c.f. Table 1.8)

TABLE 1.13 - THE pK_a OF ETHYL-2-METHYLACETOACETATE IN WATER AT 25°C

Volume 0.20M NaOH added to 20 cm^3 0.104M HA [a] (cm^3)	pH [b]	$\lg\frac{[A^-]}{[HA]}$ [c]	$\lg\gamma_\pm$ [d]	pK_a	Mean Value pK_a
1.00	11.21	-1.16	-0.04	12.41	12.41 ± 0.05
5.00	12.10	-0.27	-0.08	12.44	
7.00	12.27	-0.07	-0.08	12.38	

a Quoted concentration is corrected for 1% impurity in the sample of HA used

b The pH values quoted represent the values at zero time (see text)

c In this particular instance, because of the high pK_a of the compound with respect to pK_w, the approximation that $|[H^+] - [OH^-]| << b$, where b is the concentration of added NaOH, breaks down. Thus, if the original concentration of HA is represented by a, the expression we must use for K_a is:

$$K_a = \frac{[H^+](b-[OH^-])}{a-b+[OH^-]}\gamma_\pm^2$$

The $[OH^-]$ required can be obtained from $[OH^-] = \frac{10^{pH-14}}{\gamma_\pm}$

The values used for I must now take into account $[OH^-]$ as well as $[Na^+]$

TABLE 1.14 - THE pK_a OF ETHYL-2-METHYLACETOACETATE IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol .020M NaOH[a] added to 20cm^3 xM HA (cm^3)		$(pH_s)_{obs}$	$\lg\frac{[A^-]}{[HA]}$	$\lg\gamma_\pm$[b]	pK_a[c]	Mean Value pK_a
	x						
	0.100	1.00	11.89	-0.95	-0.049	12.24	
	"	3.00	12.55	-0.37	-0.072	12.34	
50	"	5.00	12.89	0.00	-0.086	12.33	12.30±0.10
	"	6.00	13.04	0.18	-0.088	12.30	
	"	7.00	13.22	0.37	-0.091	12.29	
	0.108	1.00	12.59	-0.98	-0.054	12.67	
	"	3.00	13.21	-0.42	-0.081	12.75	
70	"	5.00	13.53	-0.07	-0.095	12.74	12.72±0.06
	"	6.00	13.67	0.10	-0.103	12.73	
	"	7.00	13.82	0.27	-0.104	12.71	

a. Quoted concentrations are corrected for 1% impurity in sample of HA used.

b. Calculated from Davies equation (1.10) using A parameters listed in Table 1.1 and I as $[Na^+]$ in solution.

c. Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$.

TABLE 1.15 - THE pK_a of 1-(3'METHOXYPHENYL),1-NITROETHANE IN WATER AT 25°C

Concn.of species in solution before addition of HA (mol dm^{-3})[a]	Absorbance of cell contents after addition of HA (using 285nm)[b]	lg (Buffer Ratio)[c]	2 lg $\gamma_{\pm}$[d]	pK_a[e]	Mean Value pK_a
0.10M KOH	0.963	-	-	-	
0.01M HCl	0.158	-	-	-	
0.05M HB, 0.15M B	0.919	0.476	-0.152	7.46	
0.05M HB, 0.05M B	0.872	0.000	-0.152	7.33	7.40 ± 0.07
0.05M HB, 0.025M B	0.805	-0.302	-0.152	7.31	
0.05M HB, 0.01M B	0.670	-0.699	-0.152	7.48	

a. Volume of solution in each cell was 2 mls.
Abbreviations: HB = $Tris^{+}H$
B = Tris

b. 5µl of the nitroalkane were syringed into the cell. The absorbances quoted have been corrected for non-matching of the cells.

c. This refers to [B]/[HB]

d. Obtained from Davies equation (1.10) putting I = [Tris H^{+}]

e. Obtained from equation 1.24. The pK_a of $Tris^{+}H$ is taken as 8.075 (reference 39).

TABLE 1.16 - THE pK_a of 1-(3'METHOXYPHENYL),1-NITROETHANE IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol.0.02M NaOH added to 10 cm^3 0.019M HA (cm^3)	$(pH_s)_{obs}$	$lg\frac{[A^-]}{[HA]}$	$lg\gamma_{\pm}$ [a]	pK_a [b]	Mean Value pK_a
50	2.00	9.01	-0.58	-0.03	8.97	8.93 ± 0.05
	4.00	9.41	-0.15	-0.04	8.95	
	6.00	9.75	0.22	-0.04	8.92	
	8.00	10.22	0.70	-0.05	8.92	
80	2.00	11.51	-0.58	-0.04	10.98	10.94 ± 0.11
	4.00	11.88	-0.15	-0.05	10.94	
	6.00	12.25	0.22	-0.05	10.95	
	8.00	12.68	0.70	-0.06	10.90	

a. Calculated from Davies equation (1.10) using A parameters shown in Table 1.1 and I as $[Na^+]$ in solution.

b. Calculated from equation 1.20 where $pH_s = (pH_s)_{obs} - c$.

TABLE 1.17 - THE pK_a of 1-(4'NITROPHENYL),1-NITROETHANE IN DMSO-MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol.0.02M NaOH added to 10 cm^3 xM HA (cm^3)		$(pH_s)_{obs}$	Vol.NaOH [a] needed for 100% ionisation of HA (cm^3)	$(pH_s)_{obs}$ at [b] $[A^-] = [HA]$	pK_a [c]
	x					
		1.00	7.25			
		2.00	7.78			
50	$\approx 1.3 \times 10^{-2}$	3.00	8.19	5.70 ± 0.20	8.12 ± 0.05	7.40 ± 0.08
		4.00	8.65			
		5.00	9.65			
		1.00	9.60			
		2.00	10.44			
80	$\approx 5.3 \times 10^{-3}$	2.10	10.58			
		2.20	10.74	2.38 ± 0.10	9.75 ± 0.05	8.64 ± 0.13
		2.30	10.88			
		2.40	11.28			
		2.45	11.65			
		2.50	12.52			
		2.55	13.65			

a. Obtained from the graph of $(pH_s)_{obs}$ V Vol.NaOH added.

b. This value of $(pH_s)_{obs}$ is then read off the graph
Calculated from the equation:

$$pK_a = (pH_s) - \lg \gamma_{\pm}$$

where (pH_s) is the absolute pH at 50% ionisation
i.e. $(pH_s)_{obs} - c$.

c. $\lg \gamma_{\pm}$ is obtained from Davies equation (1.10).

TABLE 1.18 - COMPARISON OF PRESENT RESULTS WITH PREVIOUSLY REPORTED DATA FOR DMSO-H_2O MIXTURES[a]

Vol.% DMSO in solvent	pK_a of Acid[b]							
	AcOH	ClAcOH	2NP	NE	DNE	EMAA	m-OMe	p-NO_2
0	$(4.75)^c$	$(2.89)^c$	7.74	9.03 $(8.60)^d$	$(5.24)^e$	12.41 $(12.42)^f$	7.40	$(6.63)^g$
50	$(5.82)^h$	$(4.20)^h$	-	10.19	-	12.30	8.93	7.40
70	7.12 $(7.08^h,$ $6.99^i)$	5.06 $(4.81)^h$	10.96	11.67	5.32	12.72	-	-
80	7.95 $(8.10^h,$ $8.02^i,$ $8.00^j)$	5.56 $(5.66^h,$ $5.69^j)$	12.07	-	-	-	10.94	8.64
100	$(12.6)^k$	-	$(16.2)^{\ell}$	$(16.4)^{\ell}$	$(6.6)^{\ell}$	-	-	-

For footnotes see over:

Footnotes to Table 1.18.

a. No corresponding data are available for MeCN- or TFE-H_2O mixtures although the pK_a of AcOH in MeCN is reported to be 22.4 (reference 40).

b. Brackets indicate values from previous studies. Abbreviations are as follows: AcOH(acetic acid), ClAcOH(monochloroacetic acid), 2NP(2-nitropropane), NE(nitroethane), DNE(dinitroethane), EMAA(ethyl-2-methylacetoacetate), m-OMe(1-3'methoxyphenyl-1-nitroethane), p-NO_2(1-4'nitrophenyl-1-nitroethane).

c. Reference 25. d. References 41,42. e. Reference 43.

f. Reference 44 (Spectrophotometric method, temperature 25.7°C).

g. Reference 37 (Reported values have been corrected to infinite dilution).

h. Reference 27 (Temperature 20°C).

i. Reference 31. j. Reference 45. k. Reference 46.

ℓ. Reference 47.

6. pK_a DETERMINATIONS FOR HYDROGEN FLUORIDE

(a) Theory

Hydrogen fluoride in solution not only ionises as an acid but also associates with fluoride ion to form bifluoride ion[48,49]. Thus, in solution, the following equilibria are established:

$$HF \rightleftharpoons H^+ + F^- \quad \text{with } K_a = \frac{[H^+][F^-]}{[HF]}\gamma_{\pm}^2$$

and $$HF + F^- \rightleftharpoons HF_2^- \quad \text{with } K_{HF_2^-} = \frac{[HF_2^-]}{[HF][F^-]}$$

This means that the determination of the pK_a for HF is a more complicated matter than for the acids we have considered so far. We now need to know not only hydrogen ion activity but also fluoride ion activity. Fortunately, both are readily obtained - the former by appropriate use of the glass electrode and the latter by use of the fluoride ion selective electrode. Such a procedure has already proved to be successful for the pK_a evaluation of HF in aqueous solution[50].

If we titrate a known concentration of potassium fluoride with standard perchloric acid then at any point in the titration the concentrations of the species in solution are related through the equations:

$$m = [HF] + [F^-] + 2[HF_2^-] \qquad 1.25$$

(where "m" is the original concentration of potassium fluoride, corrected for the dilution involved) and:

$$[H^+] + [K^+] = [HF_2^-] + [F^-] + [OH^-] + [ClO_4^-]$$

which is the condition for charge neutrality in solution. Because $[OH^-] << [K^+], [F^-]$ and remembering that $[K^+] = m$, the above equation becomes:

$$[ClO_4^-] - [H^+] = m - [HF_2^-] - [F^-]$$

$$= [HF] + [HF_2^-] \qquad 1.26$$

Since $[H^+]$ and $[F^-]$ are determined experimentally from the measured ion activities ($a = \text{concn.}\gamma_\pm$) and because $[ClO_4^-]$ and 'm' are easily calculated then equations 1.25 and 1.26 enable us to evaluate $[HF_2^-]$ and $[HF]$. All the parameters required for evaluation of pK_a of HF are then available.

A few points must now be raised concerning the determination of fluoride ion concentration. This is achieved by using the fluoride ion electrode in conjunction with a calomel reference. The resulting cell, $LaF_3|F^-||$calomel, responds according to the equation:

$$E = E^o + \text{const.lg } a_{F^-} \qquad \text{(see Appendix 1)}$$

Thus, if we place the electrodes in a solution of potassium fluoride the measured potential, E_o, is given by:

$$E_o = E^o + \text{const.lg}(a_{F^-})_o$$

If we then add x cm^3 of perchloric acid the observed potential, E_x, is given by:

$$E_x = E^o + \text{const.lg}(a_{F^-})_x$$

Therefore, $\frac{E_o - E_x}{\text{const.}} = \lg(a_{F^-})_o - \lg(a_{F^-})_x$. 1.27

Knowing $\lg(a_{F^-})_o$, E_o, E_x and the constant we can calculate $[F^-]_x$, remembering that activities and concentrations are readily interconvertible ($a_{F^-} = [F^-]\gamma_{\pm}$). The constant is determined by separate calibration of the fluoride electrode for each solvent system used (see experimental section following).*

A differential method, such as the above, for the evaluation of fluoride ion activity at any point in the titration has the advantage of being independent of the experimentally derived value of E^o - thus, increasing the accuracy in the determination of the value of a_{F^-}.

The hydrogen ion concentrations in solution are determined by means of a glass combination electrode, as previously described.

(b) Experimental

Materials were purified by procedures already described in this thesis. Only DMSO-H_2O mixtures were studied.

The procedure for measurement of the absolute pH ($-\lg[H^+]\gamma_{\pm}$) of any solution, pH_s, was as described earlier with $pH_s = (pH_s)_{obs} - c$.

The fluoride ion selective electrode was calibrated prior to use in each solvent mixture (cf. Appendix 1)

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* At 25°C this constant nearly always equals the theoretically expected value of 59 mV - see Appendix 1.

Solutions of known fluoride ion concentration (commonly between 10^{-1} and 10^{-4}M) and required solvent composition were prepared and the EMF of the cell, $LaF_3|F^-||$calomel, measured for each $[F^-]$. A plot of EMF against a_{F^-} was then constructed for each solvent mixture and the gradient (i.e. the required constant) found. Values are reported in Table 1 of Appendix 1.

The basic experimental procedure for the potentiometric titration was as reported earlier in the present section. This time, however, both glass combination and fluoride ion electrodes were placed into solution; simultaneous measurements for a_{H^+} and a_{F^-} could then be made. Throughout this work fluoride ion concentrations $\leq$ 0.01M were used - hydrogen fluoride produced during the course of the titration was then insufficiently concentrated to attack the membrane of the glass electrode.

The temperature used throughout was 25.0 (± 0.1)°C. All pH and EMF measurements were recorded with a Radiometer Type 26 pH/mV meter. pH values were reproducible to within 0.04 pH units and EMF values to within 2 mV.

A rough check on the pK_a of HF in water used 0.01M potassium fluoride and 0.012M perchloric acid solutions. The amount of bifluoride ion present was negligible and the pK_a evaluated was 3.16 ± 0.02 which is in excellent agreement with Kresge and Chiang's value of 3.165 ± 0.007 (reference 50).

(c) Results

The results are given in Tables 1.19 - 1.21. Values of $K_{HF_2^-}$ are also included. Calculations of the latter, however, are heavily dependent upon the value of the bifluoride ion concentration - small changes in this figure can cause large variations in the value of $K_{HF_2^-}$. Because the bifluoride ion concentration is usually very small ($\leq 10^{-3}$M) in the experiments reported here, the errors in this value, and therefore in $K_{HF_2^-}$, are expected to be large.

Table 1.22 compares pK_a variations observed in the case of HF with those of the carboxylic acids already studied.

TABLE 1.19 - THE pK_a of HF in 50/50 (v/v) DMSO-H_2O MIXTURES AT 25°C

Vol. 0.012M $HClO_4$ added to 20 cm^3 0.01M KF (cm^3)	"m"[a] (mol dm^{-3})	I (mol dm^{-3})	Lg $\gamma_\pm$[b] (mol dm^{-3})	pH_s^-[c]	$[\overset{+}{H}]$[d] (mol dm^{-3})	EMF of cell[e] $LaF_3 \vert \bar{F} \vert \vert LaF_3$ (mV)	$[\bar{F}]$[f] (mol dm^{-3})	[HF][g] (mol dm^{-3})	$[HF_2^-]$[g] (mol dm^{-3})	pK_a[h]	$K_{HF_2^-}$[i]
0.00	10.00×10^{-3}	10.00×10^{-3}	-5.3×10^{-2}			160					
4.00	8.33×10^{-3}	8.3×10^{-3}	-4.9×10^{-2}	5.38	4.7×10^{-6}	147	6.04×10^{-3}	1.71×10^{-3}	2.9×10^{-4}	4.88	28
8.00	7.14×10^{-3}	7.1×10^{-3}	-4.5×10^{-2}	4.86	1.53×10^{-5}	132	3.35×10^{-3}	3.05×10^{-3}	3.7×10^{-4}	4.89	36
12.00	6.25×10^{-3}	6.5×10^{-3}	-4.3×10^{-2}	4.34	5.06×10^{-5}	110	1.38×10^{-3}	4.03×10^{-3}	4.2×10^{-4}	4.85	74
16.00	5.56×10^{-4}	5.5×10^{-3}	-4.0×10^{-2}	3.40	4.35×10^{-4}	60	2.0×10^{-4}	4.44×10^{-3}	4.6×10^{-4}	4.81	53
									Mean Value	4.86 (±0.03)	≈47

a "m" = $[HF] + [\bar{F}] + [HF_2^-]$

b Calculated from Davies equation (1.10) where A is 0.584

c This has been corrected for solvent effect and represents $-\lg [H] \gamma_\pm$ (see eqn. 1.21)

d Obtained from $pH_s^- = -\lg [H] \gamma_\pm$

e In 50/50 DMSO/H_2O the cell responds according to the equation $E = E^o + 58.5 \lg a_{F^-}$ (see appendix I)

f Obtained from equation 1.27 where $\lg (a_{F^-})_o = -2.053$

g From equations 1.25 and 1.26

h Calculated from $pK_a = \frac{[\overset{+}{H}][\bar{F}]}{[HF]} \gamma_\pm^2$

i $K_{HF_2^-} = \frac{[HF_2^-]}{[HF][\bar{F}]}$

TABLE 1.20 - THE pK_a of HF in 70/30 (v/v) DMSO-H_2O MIXTURES AT 25°C

Vol. 0.012M $HClO_4$ added to 20 cm^3 0.01M KF (cm^3)	$10^3 m$ [a] ($mol\ dm^{-3}$)	$10^3 I$ ($mol\ dm^{-3}$)	$10^2 \lg\gamma_\pm$ [b]	pH_s [a]	$10^7[H^+]$ [a] ($mol\ dm^{-3}$)	EMF of cell $LaF_3\|F^-\|\|cal$ (mV)	$10^4[F^-]$ [d] ($mol\ dm^{-3}$)	$10^3[HF]$ [a] ($mol\ dm^{-3}$)	$10^3[HF_2^-]$ [a] ($mol\ dm^{-3}$)	pK_a [a]	$K_{HF_2^-}$ [a]
0.00	10.00	10.00	-5.6			265					
8.00	7.14	7.1	-5.0	6.68	2.09	229	23.01	2.02	1.41	6.72	303
12.00	6.25	6.2	-4.8	6.05	8.87	204	8.36	3.54	0.96	6.78	326
16.00	5.56	5.5	-4.5	4.35	45.0	119	0.27	5.15	0.19	6.71	136[e]
									Mean Value	6.74 (±0.03)	≈310

a As in Table 1.19

b From Davies equation (1.10) with A = 0.650

c In 70/30 DMSO/H_2O the cell responds according to the equation $E = E^o + 57 \lg a_{F^-}$

d Obtained from equation 1.27 with $\lg (a_{F^-})_o = -2.056$

e This reflects the uncertainty in the low value of $[HF_2^-]$ and is ignored in the evaluation of the mean value

TABLE 1.21 - THE pK_a of HF in 80/20 (v/v) DMSO-H_2O MIXTURES AT 25°C

Vol. 0.012M $HClO_4$ added to 20 cm^3 0.01M KF (cm^3)	$10^3 m$ [a] (mol dm^{-3})	$10^3 I$ (mol dm^{-3})	$10^2 \lg \gamma_{\pm}$ [b]	pH_s [a]	EMF of cell [c] $LaF_3\|F^-\|\|cal$ (mV)	$10^3 [F^-]$ [d] (mol dm^{-3})	$10^3 [HF]$ [a] (mol dm^{-3})	$10^3 [HF_2^-]$ [a] (mol dm^{-3})	pK_a [a]	$K_{HF_2^-}$ [a]
0.00	10.00	10.00	-6.0		340					
8.00	7.14	7.1	-5.2	8.26	302	2.27	1.99	1.44	8.25	318
12.00	6.25	6.2	-5.1	7.46	271	0.69	3.39	1.11	8.20	475
0.00	10.00	10.00	-6.0		335					
8.00	7.14	7.1	-5.2	8.22	298	2.42	2.14	1.29	8.22	250
12.00	6.25	6.2	-5.1	7.45	269	0.79	3.49	1.01	8.27	366
								Mean Value	8.23 (±0.03)	≈352

a As in Table 1.19

b Calculated from Davies equation (1.10) using A as 0.695

c In 80/20 DMSO/H_2O the cell responds according to: $E = E^o + 60 \lg a_{F^-}$

d Calculated from equation 1.27 with $\lg (a_{F^-})_o = -2.060$

TABLE 1.22 - COMPARISON OF pK_a VALUES FOR HF AND CARBOXYLIC ACIDS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	pK_a of acid [a]			$K_{HF_2^-}$ [b]
	HF	AcOH	ClAcOH	
0	3.16 [c]	4.75 [d]	2.89 [d]	8 [e]
50	4.86	5.82 [f]	4.20 [f]	≈ 47
70	6.74	7.12	5.06	≈ 310
80	8.23	7.95	5.56	≈ 352

a All values quoted are from present study unless stated otherwise. Abbreviations: AcOH(acetic acid), ClAcOH(monochloroacetic acid)

b Defined by $K_{HF_2^-} = \frac{[HF_2^-]}{[HF][F^-]}$

c Reference 50

d Reference 25

e Reference 51

f Reference 27

7. DISCUSSION

It can be seen from Table 1.18 that, where work has been duplicated, correlation with pK_a values from previous studies is generally within 0.1 pK units. This indicates the reliability of the methods adopted for use in the present work. The one exception is nitroethane in aqueous solution - here a discrepancy of 0.43 is observed. However, the method used in the previously reported determination[41] would seem to be rather suspect; when repeated under conditions identical to those reported by the authors, a continual drift in pH was observed with no steady value being reached.

From the data presented in the tables for the neutral acids studied, it is clear that changing the nature of the solvent can induce major changes in pK_a values. Such behaviour can be rationalised in light of the results obtained for the free energy of transfer of ions reported earlier in the thesis. Thus, we expect variations in the pK_a of an acid HA on transfer from water to a solvent mixture to reflect mainly the changes in the free energies of the ions H^+ and A^-. The rapid increase in the pK_a's of many species with increasing DMSO content of the solvent would imply that anion destabilisation in this solvent considerably outweighs the corresponding decrease in proton free energy. Conversely, rapid increase in the pK_a values upon the addition of the highly protic trifluoroethanol to the aqueous solution would imply that proton destabilisation ~~is~~ outweigh~~ed~~s ~~by~~ the considerable anion stabilisation (the latter presumably arising through increased incidence of hydrogen bonding with solvent). Acetonitrile is known to solvate both anions and cations[4] poorly and so the observed increases in the pK_a of HA with increasing acetonitrile concentration of the medium presumably reflects both anion and cation destabilisation in this solvent.

Comparison of the pK_a values of the acids in water and, say, 70/30 (v/v) DMSO-H_2O mixtures (Tables 1.18 and 1.22) shows that the nature of the anion greatly influences the pK_a changes observed. Thus, the differences in pK_a of HA (i.e. $(pK_a)_s - (pK_a)_{aq}$) decrease in the order HF > NE > AcOH > 2NP, ClAcOH >> EMAA, DNE (the abbreviations are as used in Table 1.18). Such comparisons should reflect primarily changes in anion free energy only and, because anions of high charge density are most affected by the addition of DMSO (see first part of thesis), we should expect pK_a variations to be greater for, say, nitroethane than dinitroethane or for hydrogen fluoride compared with acetic acid. This is clearly in accordance with experimental results. Further support for this argument comes when we compare values of $(pK_a)_{HF}$ and $K_{HF_2^-}$ (Table 1.22). As the DMSO content of the solvent increases so the values of both pK_a and $K_{HF_2^-}$ rise. This presumably reflects the much greater changes in fluoride free energy compared with the bulkier bifluoride ion.

It is believed that increases in acidity for a series of related acids (e.g. substituted benzoic acids) for a given solvent result in many instances from the stabilising effect of increased charge delocalisation in the anions. In light of our present discussion it could be expected, therefore, that such increases should be more pronounced in dipolar aprotic solvents, where solvation effects are relatively small, than water where they are large and work in opposition to the delocalisation effects. It is thus encouraging to note that data available for a series of substituted benzoic acids in various non-aqueous solvents[4] and DMSO-H_2O mixtures[52] show that the pK_a differences in dipolar aprotic media are considerably greater than they are in water.

A phenomenon frequently observed in dipolar aprotic solvents is that of homoconjugation[40,53,54] where the anion A^- associates with the neutral molecule HA by means of hydrogen bonding i.e. $A^- + HA \rightleftharpoons (A\bar{H}A)$. This reflects, presumably, the very high activity of anions in dipolar aprotic solvents when compared with water. The only molecule showing such behaviour in the present study is hydrogen fluoride and in this case homoconjugation occurs even in water - reflecting the ability of fluoride ion to form very strong hydrogen bonds with suitable species.

Previous authors have observed an empirically linear relationship between pK_a and the mole fraction of DMSO in the DMSO-H_2O mixtures[32,45]. This is true for most of the acids studied although slight deviations were reported to occur at the very highest DMSO concentrations for acetic and benzoic acids. The species used in the present study all give such linear plots - this means, of course, that fairly accurate pK_a values can be obtained for almost any desired solvent composition. There are no obvious theoretical grounds for predicting such a relationship and it is not found in any other solvent system.

SECTION II

KINETIC STUDIES IN MIXED SOLVENT SYSTEMS

PART A

KINETICS AND MECHANISM OF THE OXIDATION OF DMSO WITH BROMINE IN AQUEOUS SOLUTION

1. INTRODUCTION

In subsequent sections of this thesis results are presented for proton-transfer reactions of the type $SH + B^- \rightleftharpoons S^- + BH$, where B^- represents the base used, and SH the substrate. Rate constants for the formation of S^- are often determined by using iodine or bromine to remove S^- as soon as it is formed (see, for example, reference 55). However, when this method was applied to solutions containing DMSO-H_2O as solvent a rapid reaction was observed between bromine and DMSO (such a reaction was found to be negligibly slow with iodine). It was decided to investigate the kinetics and mechanism of the bromine-DMSO reaction in aqueous solution.*

Tagaki et al[56] have reported the isolation of small amounts of diphenylsulphone when diphenylsulphoxide is allowed to react for several hours with bromine in water-acetic acid mixtures. Apart from this, however, the possibility of oxidation of sulphoxides, R_2SO, to sulphones, R_2SO_2, by halogens (X_2) does not seem to have received much attention - although much interest has been shown in the oxidation of organic sulphides, R_2S, to the corresponding sulphoxides using halogens[57]. It is generally assumed that the latter reaction proceeds via the halogenosulphonium ion as in equation 2.1:

$$X_2 + R_2S \rightleftharpoons R_2\overset{+}{S}X + X^- \xrightarrow{H_2O} R_2SO + 2H^+ + 2X^- \qquad 2.1$$

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* This investigation was carried out in conjunction with Dr. B. G. Cox and the results have already been published (B. G. Cox and A. Gibson J.C.S. Perkin II 1973 1355).

Tagaki et al propose that the mechanism for the oxidation of diphenylsulphoxide to diphenylsulphone involves either direct reaction between HOBr and the sulphoxide or a scheme analogous to equation 2.1.

In the present study we investigate the stoichiometry and kinetics of the bromine-DMSO reaction in aqueous solution, in an attempt to establish the reaction mechanism.

2. MATERIALS

Dimethylsulphoxide was purified by distillation under reduced pressure from calcium hydride, as previously described.

All inorganic materials were of AnalaR grade.

3. EXPERIMENTAL AND RESULTS

(a) Stoichiometry of reaction

The stoichiometry of the reaction was shown to be:

$$Me_2SO + Br_2 + H_2O \rightarrow Me_2SO_2 + 2Br^- + 2H^+,$$

the consumption of Br_2 and production of H^+ being determined by standard titration procedures. The production of dimethylsulphone, using acetate buffer to speed up the reaction, was followed by comparing nmr spectra of standard solutions of sulphoxide and sulphone with those of reaction mixtures. A Perkin-Elmer R20 instrument was used.

(b) Order of reaction with respect to Br_2 and DMSO

Reactions were carried out in the presence of excess sodium bromide (0.05M), perchloric acid (0.044M) and DMSO ($5 \times 10^{-2} - 3 \times 10^{-1}$M). Added bromine concentrations were about 1-3 x 10^{-3}M. The rate of disappearance of bromine in these solutions was followed using a Gilford 2400 spectrophotometer set at 389 nm. At this wavelength the absorption is due to both Br_2 and Br_3^- formed in the equilibrium $Br_2 + Br^- \rightleftharpoons Br_3^-$. Because the reaction is first order in bromine, however, (see below) the extinction coefficients are not used in calculating the rate constants.

The experimental procedure was the same in all cases: about 2 cm^3 of the aqueous solution of NaBr, $HClO_4$ and DMSO were placed in the spectrophotometer cell. The cell was then placed in the spectrophotometer cell block and thermostatted (at 25 ± 0.2°C) for forty minutes. About 10 μl of saturated aqueous bromine solution were then syringed into the cell, the cell contents stirred and the reaction followed.

Under the above conditions the reaction was found to be first order in bromine concentration over at least four half lives*. The rate law, for a given [DMSO], was found to be:

$$-\frac{d[Br_2^*]}{dt} = k'[Br_2^*]$$

where k' is the observed first-order rate constant and $[Br_2^*]$ represents the total bromine concentration in solution (i.e. $[Br_2^*] = [Br_2] + [Br_3^-]$). The constant k' was found to vary with DMSO according to the equation: $k' = k_e[DMSO]$, showing the reaction to be first order in DMSO also. Values of k' and k_e are given in Table 2.1.

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* i.e. A plot of $-lg\ A_t$ against t (A_t is the absorbance at time t) is linear over at least four half lives.

The reaction rate was found to be independent of acid concentration in solutions containing up to 0.2M perchloric acid and sufficient sodium perchlorate to maintain the ionic strength at 0.21M. However, for a constant acid concentration, a significant increase in rate was observed as the ionic strength was increased by adding sodium perchlorate. Thus, values of 10^3 k_e of 8.2, 9.3 and 10.1 dm^3 mol^{-1} s^{-1} were found when the ionic strength was 0.1, 0.15 and 0.2M respectively.

(c) Reaction rate in acetic acid-acetate buffers

The reactions were carried out in acetate buffers with acetate concentrations in the range 0.02 to 0.10M. The experimental procedure was the same as outlined in the previous section. DMSO concentrations were 0.0225M and initial bromine concentrations ca. 1 x 10^{-3}, [NaBr] = 0.055M. Because the DMSO concentration decreased by up to 10% during reaction, the mean value of the initial and final concentrations was used when calculating k_e (see paragraph b). The ionic strength was maintained at 0.2M by addition of sodium perchlorate.

Second order rate constants, k_e, in the various buffer solutions are recorded in Table 2.2. These show that the reaction rate at any particular acetate concentration is independent of the buffer ratio and that k_e can be represented by the equation:

$$k_e = k_o + k_{OAc^-} [OAc^-]$$

where k_o (obtained from a plot of k_e V[OAc^-]) has a value of

0.01 dm^3 mol^{-1} s^{-1} and k_{OAc^-} a value of 15.4 dm^6 mol^{-2} s^{-1}. The value for the former should equal that of k_e for the uncatalysed reaction at I = 0.2M. Agreement is excellent.

(d) Effect of variation of bromide ion concentration

DMSO concentrations were either 0.0225 (for lower bromide ion concentrations) or 0.0563M. Initial bromine concentrations were about 1 x 10^{-3}M. Reactions were carried out in an acetate-acetic acid buffer with [NaOAc] = [HOAc] = 0.04M whilst the ionic strength was maintained at 0.2M by the addition of sodium perchlorate. The observed second-order rate constants, k_e, are recorded in Table 2.3.

(e) Reactions in deuterium oxide

The effect of changing the solvent from H_2O to D_2O was studied in solutions with and without added acetate catalyst. For the uncatalysed reactions, solutions contained 0.096M $HClO_4$ (or $DClO_4$), 0.05M NaBr and 0.113M DMSO ($DClO_4$ was prepared by diluting 12M $HClO_4$ with D_2O). The final reaction solvent contained 98% D_2O. The observed rate constants were $k_e(H_2O)$ = 9.29 x 10^{-3} and $k_e(D_2O)$ = 3.47 x 10^{-3} dm^3 mol^{-1} s^{-1} i.e. $k_e(H_2O)/k_e(D_2O)$ = 2.7.

In the case of acetate catalysed reactions [NaOAc] = [HOAc] = 0.04M, [NaBr] = 0.055M and [DMSO] = 0.0563M (using H_2O) or 0.0274M (using D_2O). The final solvent contained 95% D_2O and the observed rate constants were $k_e(H_2O)$ = 5.87 x 10^{-1} and $k_e(D_2O)$ = 3.30 x 10^{-1} dm^3 mol^{-1} s^{-1} with $k_e(H_2O)/k_e(D_2O)$ = 1.8.

(f) Effect of high concentrations of DMSO

The rate of loss of bromine was measured in solutions containing high concentrations of DMSO (up to 85 volume %)

so that the DMSO is involved not only as reactant but also as a solvent component.

All solutions contained 0.05M NaBr and 0.048M $HClO_4$. Initial bromine concentrations were again ca 1×10^{-3}M. The observed first-order rate constants, k', are listed in Table 2.4 - the formal second-order rate constants have been obtained by dividing the first-order rate constants by the molarity of DMSO.

In the solvent mixture containing 80 volume % DMSO the effect of varying bromide ion concentration was also studied. The ionic strength was maintained at 0.10M by the addition of $HClO_4$. Observed first-order rate constants, k', are listed in Table 2.5.

The rate of oxidation of DMSO in 80 volume % DMSO was also measured in the presence of acetate ion. The solutions contained 0.05M NaBr and the initial bromine concentration was ca 5×10^{-4}M. [NaOAc] = [HOAc] and was usually 9×10^{-3}M and the average concentration during the reaction was taken as 8.5×10^{-3}M. The observed first-order rate constant was k' = 0.268 ($\pm$ 0.002) s^{-1} which corresponds to a value of k_{OAc^-} of 2.76 $dm^6\ mol^{-2}\ s^{-1}$ where

$$k_{OAc^-} = \frac{k'}{[OAc^-][DMSO]}$$

. The latter value may be compared with the value of 15.4 $dm^6\ mol^{-2}\ s^{-1}$ in aqueous media.

TABLE 2.1 - RATE CONSTANTS FOR THE OXIDATION OF DMSO WITH BROMINE AT 25°C [a]

[DMSO] mol dm^{-3}	10^4 x k' [b] s^{-1}	10^3 x k_e [c] dm^3 mol^{-1} s^{-1}
0.0563	4.76	8.45
0.0113	9.33	8.26
0.169	13.7	8.11
0.225	18.0	8.00
0.282	22.7	8.05

a Measured by B. G. Cox with [NaBr] = 0.055M and [$HClO_4$] = 0.048M

b Obtained from a plot of -lg A_t against t, the gradient being k'. A_t represents the absorbance at time t (using 389nm throughout)

c Calculated from $k_e = \frac{k'}{[DMSO]}$

TABLE 2.2 - RATES OF OXIDATION OF DMSO WITH BROMINE IN ACETATE-ACETIC ACID BUFFER[a] AT 25°C

[OAc$^-$] (mol dm^{-3})	[OAc$^-$]/[HOAc]	10 x $(k_e)_{obs}$ [b] (dm^3 mol^{-1} s^{-1})	10 x $(k_e)_{calc}$ [c] (dm^3 mol^{-1} s^{-1})
0.02	0.2	3.30	3.18
0.02	1.0	3.13	3.18
0.04	0.2	6.23	6.26
0.04	1.0	6.23	6.26
0.06	0.2	8.97	9.34
0.06	1.0	9.49	9.34
0.06	5.0	9.12	9.34
0.08	0.2	12.1	12.4
0.08	1.0	12.2	12.4
0.08	5.0	12.3	12.4
0.10	0.2	15.3	15.5
0.10	1.0	16.2	15.5
0.10	5.0	15.3	15.5

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[a] Measured by B. G. Cox using [NaBr] = 0.055M and $I = 0.2M$

[b] Calculated from method given in Table 2.1

[c] Calculated from $k_e = 0.01 + 15.4[OAc^-]$

TABLE 2.3 - THE EFFECT OF BROMIDE ION CONCENTRATION ON THE RATE OF OXIDATION OF DMSO WITH BROMINE [a] AT 25°C

[NaBr] (mol dm^{-3})	0.025	0.045	0.064	0.084	0.104	0.124	0.143	0.163	0.183	0.203
$10^2 \times k_e$ (dm^3 mol^{-1} s^{-1})	176	88.7	48.5	33.5	22.6	15.6	12.6	10.50	8.53	6.90

a Measured by B. G. Cox using [HOAc] = [NaOAc] = 0.04M and ionic strength = 0.2

b Calculated by procedure given in Table 2.1

TABLE 2.4 - EFFECT OF HIGH DMSO CONCENTRATIONS ON THE RATE OF OXIDATION OF DMSO WITH BROMINE [a] AT 25°C

Volume % DMSO	[DMSO] (mol dm^{-3})	10^4 x k' [b] (s^{-1})	10^4 x k_e [c] (dm^3 mol^{-1} s^{-1})
1.8	0.282	23.8	84.5
10.0	1.44	117	81.3
20.0	2.87	160	55.7
40.0	5.75	103	17.9
60.0	8.63	18.9	2.19
80.0	11.4	0.73	0.064
85.0	12.1	0.23	0.019

a Using [NaBr] = 0.05M and [$HClO_4$] = 0.048M

b Obtained by method shown in Table 2.1

c Calculated from k_e = k'/[DMSO]

TABLE 2.5 - EFFECT OF VARIATION OF [Br^-] ON THE RATE OF OXIDATION OF DMSO WITH BROMINE [a] IN 80/20 (v/v) DMSO/H_2O AT 25°C

[NaBr] (mol dm^{-3})	10^5 x k' [b] (s^{-1})	10^7 x k'[NaBr]^2 (mol^2 dm^{-6} s^{-1})
0.027	26.5	1.93
0.052	7.3	1.97
0.077	2.92	1.73
0.095	2.18	1.97

a Ionic strength made up to 0.1M using $HClO_4$

b Calculated as in Table 2.1

4. DISCUSSION

The results obtained strongly support a mechanism for the oxidation as shown below**.

$$Br_2 + Me_2S{=}O \underset{fast}{\overset{K}{\rightleftharpoons}} Me_2BrS^+{=}O + Br^- \qquad (a)$$

$$Me_2BrS^+{=}O + H_2O(+\ B) \xrightarrow[slow]{k_1} Me_2S(=\overset{+}{O}H){=}O + H^+ \ (\text{or } BH^+) + Br^- \qquad (b)$$

$$Me_2S(=\overset{+}{O}H){=}O \xrightarrow[fast]{k_2} Me_2SO_2 + H^+ \qquad (c)$$

where B represents the added (excess) base and k_1 includes the concentrations of water and base.

This mechanism (in the case of low $[Me_2BrS^+{=}O]$ and excess $[Br^-]$) predicts the rate law:

$$-\frac{d[Br_2^*]}{dt} = k_e[Br_2^*][DMSO] \qquad 2.2$$

with $k_e = k_1K/[Br^-](1 + K_e[Br^-])$

where K_e is the equilibrium constant for the formation of tribromide ion.

Equation 2.2 can easily be derived; applying the steady state principle to the intermediate gives:

$$[Me_2\ BrS^+O] = K[Me_2SO][Br_2]/[Br^-]$$

** The structure of the intermediate in (a) and (b) may in fact be Me_2S^+-OBr rather than $Me_2BrS^+{=}O$. As hydrolysis of both could lead to the product Me_2SO_2 it is difficult to choose between them.

The rate law is then given by:

$$-\frac{d[Br_2^*]}{dt} = k_1K[Me_2SO][Br_2]/[Br^-]$$

However, $[Br_2^*] = [Br_3^-] + [Br_2] = Br_2(1 + K_e[Br^-])$.
Thus:

$$-\frac{d[Br_2^*]}{dt} = \frac{k_1K[Me_2SO][Br_2^*]}{[Br^-](1 + K_e[Br^-])}$$

which is analogous to equation 2.2.

Equation 2.2 predicts that a plot of $\{k_e(1 + K_e[Br^-])\}^{-1}$ against $[Br^-]$ should be a straight line passing through the origin. Such a plot is shown in Figure 2.1 using data from Table 2.3 and a value of K_e of 16.8[58]. It can be seen that the plot is linear over a 10-fold variation of bromide ion concentration, supporting the proposed mechanism.

The proposed mechanism predicts that the uncatalysed reaction, because of the pre-equilibrium involved, should show a positive salt effect (i.e. an increase in rate with increasing ionic strength). This is observed. Also consistent is the observation of catalysis by acetate ion (as in the case of bromine oxidation of organic sulphides[56]) and the observed deuterium isotope effects** - both indicate a proton transfer occurring during

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** The interpretation of the deuterium solvent isotope effects is complicated by the unknown effect of solvent on the equilibrium (a) and free energies of the species involved in (b) but would seem to be too high to be attributed solely to a medium effect.

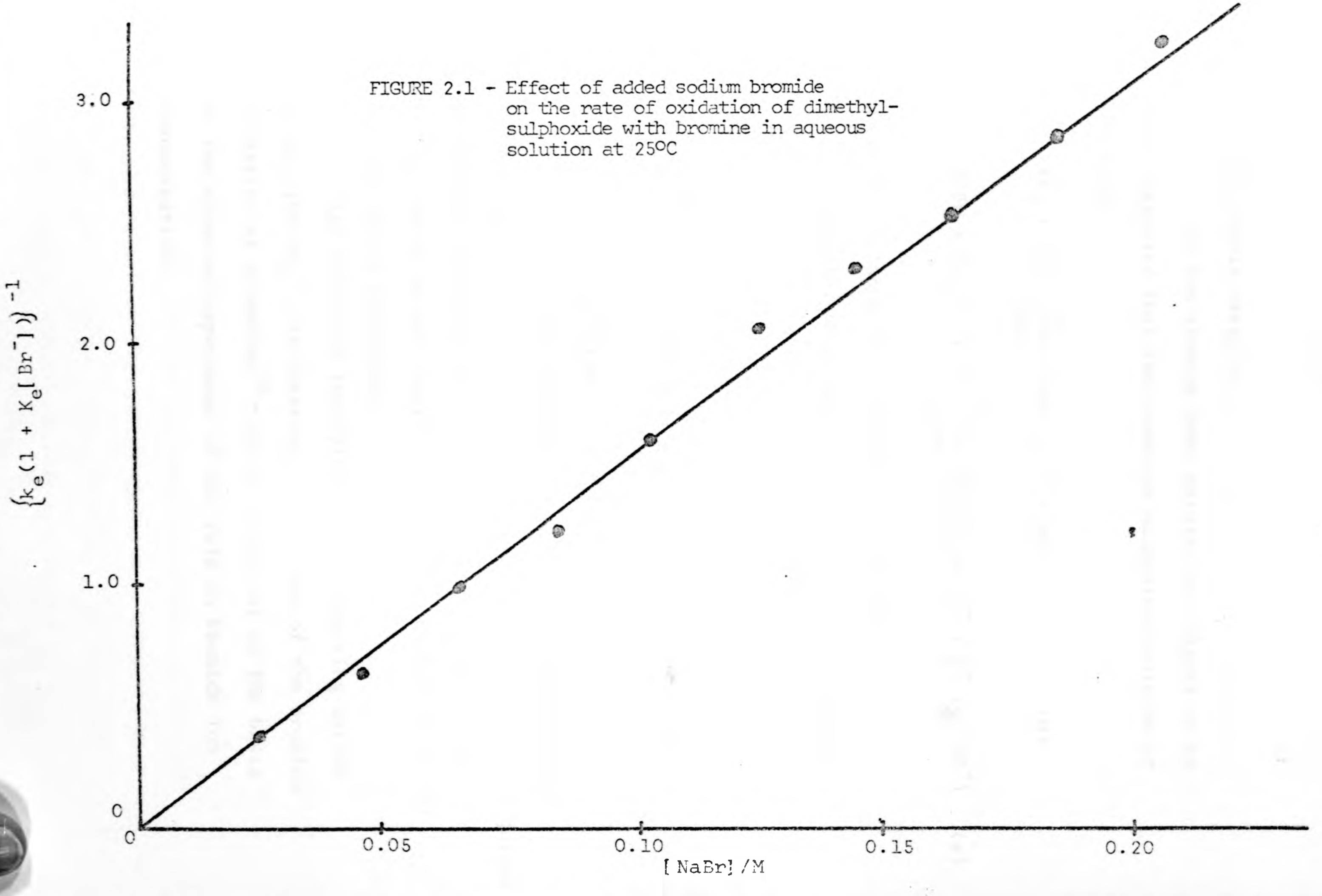

FIGURE 2.1 - Effect of added sodium bromide on the rate of oxidation of dimethylsulphoxide with bromine in aqueous solution at 25°C

the hydrolysis step, b.

As has already been pointed out Tagaki et al have suggested that the reaction mechanism could be of the form:

$$Br_2 + H_2O \underset{fast}{\overset{K'}{\rightleftharpoons}} HOBr + H^+ + Br^- \qquad (d)$$

$$HOBr + Me_2SO\ (+\ B) \xrightarrow[slow]{k_3} Me_2SO_2 + Br^- + H^+\ (or\ BH^+) \qquad (e)$$

where k_3 includes the concentration of base.

However, under present conditions where [HOBr] will be low, such a mechanism predicts a rate equation of the form:

$$-\frac{d[Br_2^*]}{dt} = k_e[Me_2SO][Br_2^*] \qquad 2.3$$

with $k_e = k_3K'/\ [H^+][Br^-](1 + K_e[Br^-])$.

Although the dependence on bromide concentration is the same as for the previous scheme such a mechanism requires an inverse dependence of the rate upon hydrogen ion concentration which is not observed. Thus, the mechanism of (d) and (e) can be eliminated.

Any mechanism involving rate determining attack by Br_2 (or Br_3^-), as observed in the case of the bromine oxidation of aldehydes[59], can be ruled out on the basis of the observed dependence of the rate on bromide ion concentration.

Considered in light of the results of Section I of this thesis, the effect of increasing DMSO concentration on the rate of reaction (Tables 2.3 and 2.4) is also consistent with the mechanism in equations a-c. This will now be discussed in detail.

We expect the equilibrium constant for (a) to decrease markedly with increasing DMSO content of the solvent mixtures because of the poorer bromide ion solvation. This will cause a decrease in the equilibrium value of $[Me_2BrS^+O]$ and thus an overall decrease in the rate of reaction (given by $k_1[Me_2BrS^+O]$) should be observed. The free energy of water decreases significantly upon addition of DMSO[60] and this should tend to cause a reduction in the value of k_1, which will complement the effect on the rate of reaction caused by the reduction in $[Me_2BrS^+O]$. It is encouraging to note that, for the uncatalysed reaction, a reduction in k_e by a factor of 4×10^3 is observed in going from water to an aqueous solution 85% by volume of which is DMSO (Tables 2.1 and 2.4).

As with bromide ion we expect the acetate ion free energy to rise sharply as the DMSO content of the solvent increases. Such an effect should tend to counterbalance the decrease in the rate of the catalysed reaction predicted on the basis of factors like those outlined above for the uncatalysed reaction. The results show this to be the case; for the catalysed reaction a decrease in rate by a factor of 5.6 is observed in going from water to a solution 85% by volume DMSO. This is to be compared with a factor of 4×10^3 in the uncatalysed reaction.

Finally it should be noted that the equilibrium constant, K_e, for the process $Br_2 + Br^- \rightleftharpoons Br_3^-$ should increase with increasing DMSO content of the solvent because of the much larger increase in bromide free energy with respect to that of Br_3^-.* Thus, at high enough concentration of DMSO we might well expect to reach the situation where $K_e >> 1$. This being the case, equation 2.2 predicts the relationship:

$$k_e = \frac{k_1 K}{K_e} [Br]^{-2}$$

It is satisfactory to note that this was found to be the case for the one solvent mixture used (80% by volume DMSO with [NaBr] $\geqslant$ 0.025M). Results are shown in Table 2.5.

Thus, the mechanistic scheme a-c readily accommodates all the experimental observations listed in Tables 2.1 - 2.5.

It may be noted that, because the oxidation is very fast in the presence of acetate ions, the reaction may prove to be a reasonable method for the preparation of sulphones as an alternative to the more usual oxidation with hydrogen peroxide[62]. It is also interesting to note that Corey et al[63] have recently used DMSO-chlorine mixtures as reagents for the oxidation of primary and secondary alcohols to carbonyl compounds. Reacting the alcohol, RR'CHOH, with DMSO-Cl_2 at -45°C in dichloromethane as solvent they produced

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* Such increases are observed for the corresponding $I_2 + I^- \rightleftharpoons I_3^-$ equilibrium[61] and, from results presented earlier, for the equilibrium $HF + F^- \rightleftharpoons HF_2^-$.

a complex, which upon addition of triethylamine yielded the ketone RR'CO. The proposed mechanism involves reactions and intermediates which are entirely analogous to those reported here for the bromine oxidation of DMSO in aqueous solution, although in the absence of H.OH the attack on the intermediate is by RR'CH.OH. The reactions do, of course, differ in the nature of the final step. The reported scheme can be summarised as shown below:

$$(CH_3)_2SO + Cl_2 \longrightarrow (CH_3)_2ClS^+{=}O + Cl^-$$

$$\Big\downarrow \text{RR'CHOH at } -45^{\circ}C$$

$$(CH_3)_2S(={\overset{+}{O}}CHRR'){=}O + Cl^- + \overset{+}{H}$$

$$\Big\downarrow Et_3N \text{ at } -45^{\circ}C$$

$$RR'C{=}O + Et_3NHCl + (CH_3)_2SO$$

Finally, it ought to be pointed out that, even though the rate of oxidation in the absence of added catalyst decreases significantly with increasing DMSO content of the solvent, bromine still cannot be used as a scavenging reagent in such media. This is because in the presence of the bases used in the corresponding proton-transfer reactions the oxidation would still be too rapid for the bromination of the anions to be followed.

PART B

PROTON-TRANSFER REACTIONS IN
MIXED SOLVENT SYSTEMS

1. INTRODUCTION

During the last fifty years examination of the rapidly accumulating data on reaction rates and equilibria has led to the formulation of several empirical correlations. Their general form is a linear relationship between the logarithms of the rate or equilibrium constants for a given reaction (denoted by the general symbol k_i^A) using one group of reactants and those for the same reaction* involving a group of structurally similar reactants (k_i^B). The relationship may be expressed[64] as:

$$\lg k_i^A = m \lg k_i^B + c \qquad 2.4$$

where m is the slope and c the intercept of the straight line observed. Clearly, such relationships can correlate equilibrium constants, rate constants or both.

Equilibrium constants (K) are related to the overall standard free energy change in the reaction by $\Delta G^o = -RT\ln K$, and specific rate constants (k') can, according to transition-state theory[65], be expressed in terms of a free energy of activation ($\Delta G^\dagger$) by the equation:

$$\ln k' = \ln\frac{kT}{h} - \frac{\Delta G^\dagger}{RT} \qquad 2.5$$

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* The cases considered would usually involve the same reaction for a series of similar substrates, the same reaction for a given substrate with differing reagents or a single reaction under a variety of environmental conditions (e.g. solvent changes).

T is the absolute temperature, k the Boltzmann constant, h the Planck constant and R the gas constant. It therefore follows that the empirical correlation of reactivity changes by means of equation 2.4 is equivalent to a linear free energy relationship (LFER) since it can be expressed:

$$\Delta G^{A} = n\Delta G^{B} + d \qquad 2.6$$

where the symbol G refers to a standard free energy, be it ΔG^{o} or $\Delta G^{\dagger}$. The relationship between n and m and between d and c obviously depends upon whether the reactivity comparison is expressed in terms of equilibrium constants, rate constants or both.

Many examples of LFER's are now known[64]. The first one to be introduced was that of Brönsted and Pedersen in 1924; this relationship - the Brönsted relation - differs from most others in that it relates rates and equilibria for the same reaction.

The Brönsted relation correlates the acid or base strength of catalysts with their effectiveness in reactions subjected to general acid or base catalysis respectively; it was developed largely as a result of the studies by Brönsted and Pedersen[66] on the general base-catalysed decomposition of nitramide. The relationship may be expressed by the following equations:

$$\lg k_A = \lg G_A + \alpha \lg K_A \qquad 2.7$$

$$\lg k_B = \lg G_B - \beta \lg K_A \qquad 2.8$$

where k_A and k_B are the catalytic constants in some process for acid (A) and base (B) catalysis, respectively; K_A is the dissociation constant for acid A (equation 2.7) or the conjugate acid of the base B (equation 2.8); G_A, G_B, α and β are constants that depend upon the reaction process and the conditions of temperature and solvent.

For a simple proton-transfer reaction of the type $SH + B^- \underset{k_A}{\overset{k_B}{\rightleftharpoons}} S^- + BH$ equation 2.8 can be written:

$$\begin{aligned} \lg k_B &= \lg G - \beta \lg K_{BH} \\ &= \lg G - \beta \lg \frac{K_{SH}}{K} \end{aligned}$$

where K_{SH} is the dissociation constant of the substrate and K the overall equilibrium constant for the reaction (i.e. k_B/k_A).

If we change B to a structurally similar base B', then the above equation becomes:

$$\lg k_{B'} = \lg G - \beta \lg \frac{K_{SH}}{K'}$$

Subtracting the two above equations then gives us:

$$\lg \frac{k_B}{k_{B'}} = \beta \lg \frac{K}{K'}$$

and this (using equation 2.5) can be written:

$$\Delta G^{\dagger\prime} - \Delta G^{\dagger} = \beta(\Delta G^{o\prime} - \Delta G^{o})$$

i.e. $$\beta = \frac{\delta \lg k}{\delta \lg K} = \frac{\delta \Delta G^{\dagger}}{\delta \Delta G^{o}} \qquad 2.9$$

where $\delta\Delta G^{\dagger}$ represents the difference in free energies of activation for the process and $\delta\Delta G^o$ the corresponding difference in standard free energies of reaction.

A similar treatment can be applied to the Brönsted α coefficient for the acid catalysed back reaction (k_A) and it can be shown (see page 491 of reference 67, for example) that, for such a case, $\alpha + \beta = 1$. It should be noted that, because acid-base catalysis involves proton transfer between the catalyst and substrate, Brönsted-type relations also apply to the variation of reaction velocity with acid-base strength of the substrate. The values of α and β so obtained, however, are not necessarily equal to those obtained by varying the nature of the catalyst.

The two limiting values of β, as defined by equation 2.9, would normally be expected to arise when (a) the bases become so strong that the forward reaction is diffusion controlled (in which case $\beta = 0$) and (b) when the bases become so weak that the back reaction becomes diffusion controlled (in which case $\beta = 1$). Such considerations led Brönsted and Pedersen to believe that the values of α and β should always be positive and lie between zero and unity. Certainly all of their own experimental values lay between these limits.

Despite the above, however, a number of Brönsted coefficients lying outside these limits are now known[68,69,70]. Such deviations are not consistent with any simple interpretation of the Brönsted relation in terms of potential energy curves or with the usual form of the Marcus equation which relates free energies of activation and standard free energy of reaction (page 216 of reference 48). Attempted explanations of these deviations usually depend upon the occurrence of effects specific to either reactants, products or transition state[48,67,68].

In Section I of this thesis it was shown that acid strengths are considerably influenced by the nature of the solvent, this governing the relative solvation (and therefore free energies) of the species involved in the equilibrium. Particularly strong solvation was shown to occur in water relative to many other solvents. Similarly, it was pointed out that the rates of very many reactions are highly sensitive to changes in solvent. It is somewhat surprising, therefore, that no one has fully considered the extent to which Brönsted coefficients (which are, of course, dependent upon both rates and equilibrium constants), and particularly the "anomalous" Brönsted coefficients, are due to the effects of solvation in water. The present part of the thesis deals, therefore, with the effect of changing solvent on the rates, equilibria and Brönsted β coefficients for reactions of the type:

$$SH + B^- \rightleftharpoons S^- + BH \ .$$

The substrates used are sufficiently acidic for accurate determinations of their pK_a to be made and the results have already been presented in Section I. They include several substituted nitroalkanes (and also 1:1 dinitroethane), substrates for which anomalous Brönsted values have been found in aqueous solution[43,67].

It should be noted that the results presented in the following pages refer to reaction between a neutral substrate and negatively charged base. Both the kinetic and thermodynamic behaviour of systems of different charge type (e.g. $SH + B \rightleftharpoons S^- + BH^+$) is expected to differ from that observed here. It should also be remembered that in those cases where the substrate has more than one dissociable proton (or when the base has more than one equivalent point at which a proton can be attached), it is usual to apply statistical corrections to the observed rates and equilibria (see pages 197-199 of reference 48). However, in the work presented here this problem rarely arises and even when it does (e.g. in comparisons of data for nitroethane and dinitroethane) the resulting corrections have a negligible effect on α and β values. Such statistical factors will not be mentioned further.

2. DETERMINATION OF REACTION RATES - BASIC THEORY

If the concentrations of B^- and BH are held constant in the proton transfer:

$$SH + B^- \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} S^- + BH$$

then the reaction can be characterized by the two first-order rate constants k_1 and k_{-1}, the values of which will depend upon the actual concentrations of acid and base used. k_1 can be expressed in the form $k_1 = k_B[\bar{B}]$ where k_B is the characteristic velocity constant for proton transfers from

SH to the base B. Similarly $k_{-1} = k_{BH}[BH]$. These forward and reverse rate constants are related to each other through K_{SH}, the dissociation constant of the substrate, and K_{BH}, the dissociation constant of the conjugate acid to base B:

$$\frac{k_B}{k_{BH}} = \frac{[S^-][BH]}{[SH][B^-]} = \frac{K_{SH}}{K_{BH}} \qquad 2.10$$

(where, at sufficiently low ionic strengths, the activity coefficients of the species cancel out).

If we have more than one base present in the reaction solution then the observed first-order rate constant, k_{obs}, for the production of S^- is of the form $k_{obs} = \sum_i k_{Bi}[Bi]$. The above proton transfer carried out in aqueous solution, for instance, will have not only B^- but also H_2O and OH^- present as bases. The first-order rate constant for the production of S^- is therefore $k_{obs} = k_o + k_{OH}[OH^-] + k_{B^-}[B^-]$ where k_o includes the concentration of water.

Considerations such as the above must always be borne in mind when interpreting kinetic data for proton-transfer reactions.

The method used to obtain a value for any catalytic constant, k_{Bi}, is primarily governed by the relative concentrations of S^- and SH present at equilibrium.*

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* The ratio of S^-:SH present at equilibrium can be calculated from equation 2.10 which can be written:

$$\lg\frac{[S^-]}{[SH]} = pH - pK_{SH}$$

At pH = pK_{SH} + 2, 99% ionisation of SH is achieved whereas at pH = pK_{SH} - 2, the figure is only 1%.

Clearly, if the value of S^-/SH at equilibrium is sufficiently high (usually ≥ 0.1) then the forward reaction occurs to an extent which, in principle, enables the rate of approach to equilibrium to be measured. This can then be used to calculate k_{Bi}. However, in many instances (and particularly in many of the cases considered in this thesis) negligible reaction would occur before equilibrium was reached, the final value of S^-/SH being very low. Under conditions such as the latter a "scavenger" is used to react with the anion and stop the equilibrium being established; the usual reagents used for this purpose are bromine and iodine[55]. Each of these cases (i.e. direct observation of approach to equilibrium and scavenging) will now be considered in detail.

(a) Direct observation of approach to equilibrium.

For the general proton-transfer reaction:

$$SH + \sum_i Bi \overset{K}{\rightleftharpoons} S^- + \sum_i Ai$$

where the charges on Bi and Ai have been omitted for clarity, we may write for the rate of production of S^-:

$$-\frac{d[SH]}{dt} = \frac{d[S^-]}{dt} = \sum_i k_{Bi}[Bi][SH] - \sum_i k_{Ai}[Ai][S^-]$$

where the symbolism is the same as that used previously. Because $\sum_i [Bi]$ and $\sum_i [Ai]$ are kept constant throughout the reaction then the overall integrated rate equation is:

$$\ln([S^-]_e - [S^-]) = -(\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])t + \text{const.} \qquad 2.11$$

where $[S^-]_e$ is the concentration of anion at equilibrium and $[S^-]$ is the concentration at any time, t.

A derivation of this equation is given in Appendix III.*

As an example of the above consider the case of a proton-transfer reaction involving SH and acetate ion in aqueous solution. Equation 2.11 can then be written:

$$\ln([S^-]_e)-[S^-]) = -(k_o' + k_{OAc}[OAc^-] + k_{HOAc}[HOAc])t + \text{const.}$$

where k_o' includes all the first-order rate constants for the proton transfers involving SH, S^-, H_2O, OH^-. Defining $r = \frac{[OAc^-]}{[HOAc]}$ and remembering equation 2.10 then:

$$\ln([S^-]_e-[S^-]) = -\left\{k_o' + [OAc^-]k_{OAc}(1 + \frac{K_{HOAc}}{rK_{SH}})\right\}t + c$$

Plotting $\ln([S^-]_e - [S^-])$ against t will give a straight line. If the gradient is measured for several different values of $[OAc^-]$ at constant r then a plot of gradient against $[OAc^-]$ will give another straight line whose slope will be $k_{OAc}\ (1 + \frac{K_{HOAc}}{rK_{SH}})$. Knowing K_{HOAc}, K_{SH} and r enables k_{OAc} to be calculated.

If the equilibrium ratio of $[S^-]:[SH]$ is sufficiently high, of course, the observed rate law will

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* If we follow the rate of production of S^- spectrophotometrically then, because changes in the anion concentration are proportional to the observed changes in absorbance, equation 2.11 becomes:

$$\ln(A_e-A) = -(\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])t + \text{const.} \qquad 2.12$$

be of the form $-\frac{d[SH]}{dt} = \sum_i k_{Bi}[Bi][SH]$.

This, for instance, is the case in the present work using hydroxide as the base.

(b) Scavenging Technique.

If we denote the halogen by X_2 then the reaction scheme for any scavenging reaction may be written:

$$SH + \sum_i Bi \rightleftharpoons S^- + \sum_i Ai$$

$$S^- + X_2 \xrightarrow{k} \text{Products}$$

where one mole of substrate reacts with one mole of halogen.

Applying the normal steady state approximation to S^- gives:

$$[S^-] = \frac{(\sum_i k_{Bi}[Bi])[SH]}{\sum_i k_{Ai}[Ai] + k[X_2]}$$

The rate of reaction is then given by:

$$-\frac{d[SH]}{dt} = \frac{k[X_2](\sum_i k_{Bi}[Bi][SH])}{\sum_i k_{Ai}[Ai] + k[X_2]} = -\frac{d[X_2]}{dt}$$

If $k[X_2] >> \sum_i k_{Ai}[Ai]$ then

$$-\frac{d[SH]}{dt} = -\frac{d[X_2]}{dt} = (\sum_i k_{Bi}[Bi])[SH] \qquad 2.13$$

and the reaction rate is independent of halogen. If this inequality holds then the scavenging technique can usually be used to obtain values for k_{Bi}.

It should be noted that halide ions, X^-, are produced during the reaction which means that the equilibrium

$X^- + X_2 \overset{K}{\rightleftharpoons} X_3^-$ is established. As long as $[X^-]$ is in excess, however, the kinetics are still very simple and because all the halogen species are in rapid equilibrium we may write:

$$-\frac{d[SH]}{dt} = -\frac{d[X_2^*]}{dt} = \sum_i k_{Bi}[Bi][SH]$$

where $[X_2^*] = [X_2] + [X_3^-] = [X_2](1 + K[X^-])$

with $[X^-] >> [X_2],[X_3^-]$.

The overall rates themselves are, of course, independent of the particular halogenating species taking part in reaction, be it X_2, X_3^- or both.

As an example of a scavenging reaction let us consider the case of proton transfer between SH and a base B^- in aqueous solution. The kinetic behaviour is governed by the relative concentrations of SH, B and X_2. There are two cases of particular importance and each will now be considered in turn.

(i) $\underline{[SH], [B^-] >> [X_2^*]}$

In such a case:

$$-\frac{d[X_2^*]}{dt} = (k_o' + k_B[B^-])[SH] = \text{constant}$$

where the symbols have the same meaning as before. For this reaction pseudo zero-order kinetics are observed.

The rate of disappearance of X_2^* is usually followed spectrophotometrically by recording the changes in absorbance at a wavelength λ, where only the halogen species absorb (it does not matter whether absorption is due to X_2, X_3^- or both). To convert the absorbances into concentrations of X_2^* required by the previous equation necessitates the use of the Beer-Lambert Law. The required extinction coefficient (ε^*) at this λ and $[X^-]$ is easily obtained by measuring the absorbance of a solution of known $[X_2^*]$ in a 1 cm cell, whence $\varepsilon^* = \text{Absorbance}/[X_2^*]$. Clearly then for any reaction followed at this λ and $[X^-]$ the concentration of X_2^* at any point is given by:

$$[X_2^*] = A/\varepsilon^*$$

Thus, our kinetic expression becomes:

$$-\frac{dA}{dt} = \varepsilon^*(k_o' + k_B[B^-])[SH] \qquad 2.14$$

A plot of A against t will give a straight line of slope $(k_o' + k_B[B^-])\varepsilon^*[SH]$. If we perform the reaction at several different values of $[B^-]$ then a plot of $(k_o' + k_B[B^-])$ against $[B^-]$ will give a straight line of slope k_B. $[SH]$ and ε^* must be known before k_B can be calculated.

Such zero-order scavenging clearly requires the substrate to be of very high purity since small amounts of reactive impurity can severely complicate interpretations of the kinetic data. Also the value of ε^* is solvent dependent and so changes in the reaction medium necessitate the redetermination of ε^*.

(ii) $[B^-] >> [SH] < [X_2*]$

The reaction is followed as in the case just discussed. However, this time, [SH] is no longer constant throughout the reaction and pseudo first-order kinetics are observed:

$$-\frac{d[SH]}{dt} = (k_o' + k_B[B^-])[SH]$$

Because the concentration of added X_2 is greater than that of SH, then [SH] at any time must equal $[X_2*]-[X_2*]_e$ (where $[X_2*]_e$ is the concentration of total halogen at the end of reaction). Therefore:

$$-\frac{d[SH]}{dt} = -\frac{d[X_2*]}{dt} = (k_o' + k_B[B^-])([X_2*]-[X_2*]_e).$$

Integration of the above expression gives:

$$\ln([X_2*]-[X_2*]_e) = (k_o' + k_B[B^-])t + \text{const.}$$

As before $[X_2*] = A/\varepsilon*$ and so:

$$\ln(A-A_e) = -(k_o' + k_B[B^-])t + \text{const.} \qquad 2.15$$

A plot of $\ln(A-A_e)$ against t will give a straight line of slope $-(k_o' + k_B[B^-])$ and once figures have been obtained for various values of $[B^-]$ then a plot of $-(k_o' + k_B[B^-])$ against $[B^-]$ will give us k_B.

The kinetic expression using first-order scavenging is independent of $\varepsilon*$ and [SH] which means that no new parameters are required if the reaction medium is changed.

It ought to be pointed out that (a) any scavenging reaction is intrinsically reversible (see for example reference 71) and that (b) at some stage in reaction the scavenging condition ($k[X_2] >> \sum_i k_{Ai}[Ai]$) will obviously break down. However, it is usually possible to achieve reaction conditions such that these effects cause deviations from expected kinetic behaviour only at the very end of reaction and hence can be ignored.

3. EXPERIMENTAL

(a) Materials.

The materials used were the same as those already mentioned in Section I of this thesis and details of preparations etc. will not be repeated here.

All inorganic materials (except KF) were again of AnalaR grade.

(b) Determination of extinction coefficients.

As has already been pointed out, values for the extinction coefficients relating to X_2* must be known before catalytic constants can be obtained from zero-order scavenging reactions.

The present study used iodine as the scavenging reagent and the reactions were always carried out at 353 nm with $[I^-] = 0.01M$. Because only I_3^- absorbs at this wavelength the value of ε* at any given temperature in any medium (using 353 nm) is given by:

$$\varepsilon^* = \left\{1 + \frac{1}{K[I^-]}\right\}^{-1}\varepsilon$$

where K is the equilibrium constant for tri-iodide formation ($I_2 + I^- \rightleftharpoons I_3^-$) and ε is the extinction coefficient of I_3^- at this wavelength and temperature. Because the values of K and ε in aqueous solution are 710 and 2.64×10^4 respectively at 25°C [72] the expression for ε_{aq}^* under these conditions becomes:

$$\varepsilon_{aq}^* = 2.64 \times 10^4 \left(1 + \frac{1.48 \times 10^{-3}}{[I^-]}\right)^{-1} .$$

At $[I^-] = 0.01M$, $\varepsilon_{aq}^* = 2.314 \times 10^4\ mol^{-1}\ dm^3\ cm^{-1}$.

The value of ε^* (at 353 nm, $[I^-] = 0.01M$ and 25°C) in any solvent other than water can be obtained by comparing the observed absorbance for a solution of known $[I_2^*]$ using water as solvent (A_{aq}) with that of a solution of equal $[I_2^*]$ using the solvent mixture (A_s). We can then use the equation:

$$\frac{A_{aq}}{A_s} = \frac{2.314 \times 10^4}{\varepsilon_s^*} \qquad 2.16$$

to obtain the value of ε_s^*.

In practice separate solutions of 0.02M KI and 10^{-4}M I_2 were made up in both water and the required solvent mixture. To each of four silica spectroscopic cells was then added 1 cm³ of aqueous KI and 1 cm³ of aqueous I_2. These four cells were then placed into the cell block of a Gilford 2400S spectrophotometer and thermostatted at

25 (± 0.2)°C. Steady values of the absorbances were then recorded.* The mean value was then calculated. Similarly the value of the absorbance in the solvent mixture was obtained and the required extinction coefficient calculated from equation 2.16.

To within experimental error, the values of ε^* in trifluoroethanol- and acetonitrile-water mixtures were found to be the same as that in water. Variations were found in DMSO-H_2O mixtures, however, and the results are given in Table 2.6.

(c) Kinetic Measurements.

The particular method used for a given substrate-base reaction is included in the results section. The reactions were followed with a Gilford 2400 or 2400S spectrophotometer, or a Durrum-Gibson stopped flow spectrophotometer depending upon the reaction rate. All measurements were made at 25 (± 0.2)°C.

When the Gilford spectrophotometers were used the usual experimental procedure was as follows: into each of four 1 cm. cells were pipetted 2 cm^3 of reaction solution containing the necessary concentration of buffer (i.e. B^- and BH mixture). First-order rate constants were usually obtained for at least four differing base concentrations.

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* The cells used were not matched and so the necessary correction factors had to be applied to the absorbance readings. The procedure for doing this was as outlined in the part of this thesis dealing with the determination of pK_a values by spectrophotometry.

If the reaction was to be followed by the scavenging technique then these solutions were also 0.01M in KI and about 10^{-4}M in total iodine. The cells were then thermostatted for forty minutes in the cell block, after which the required amount of pure substrate (or solution thereof in the appropriate solvent) was syringed into the first cell; the volume added was usually 10-30µl. The cell contents were then shaken (or stirred) and the reaction followed. The procedure was then repeated for the three remaining solutions; however, if the velocities of the reactions were suitably slow the spectrophotometer could be used to monitor all four cell reactions at the same time because it could be set to record the absorbance of each cell in turn at preset time intervals. Whenever a reaction was repeated, rate constants were almost invariably reproducible to within 5%.*

In those cases where the approach to equilibrium was followed by stopped flow the reaction procedure was as follows: one syringe was filled with the standard buffer solution and the other with the solution of substrate (the solvent in both being the same, of course). Care was taken to ensure that no air bubbles were in solution.

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* It ought to be noted that where zero-order scavenging with iodine was performed allowance always had to be made for the slow spontaneous loss of iodine which always occurred in the buffer solutions. This was achieved by performing separate experiments and subtracting the slope of the absorbance against time plot for the "buffer reaction" from the gradient obtained in the corresponding scavenging reaction. Only rarely did the correction cause more than a 5% decrease in the observed value of k_B.

The machine was then activated and the resulting oscilloscope trace photographed using the Polaroid camera. This procedure was repeated for each required base concentration and solvent composition.

The stopped flow instrument does not give a direct measurement of absorbance, the oscilloscope trace representing a plot of the photomultiplier output voltage against time (where the voltage is proportional to the intensity of the light beam emerging from the reaction compartment). It is therefore related to the absorbance of solution through the equation:

$$A = \ln \frac{I_o}{I} = \text{constant} - \ln V$$

where A is the absorbance of solution at the wavelength used, I_o the intensity of incident light, I the intensity of the light having passed through solution and V the corresponding voltage. Because of this relationship equation 2.12 can now be written:

$$\ln([S^-]_e - [S^-]) = \ln(-A+A_e) = \ln(\ln\frac{V}{V_e})$$

$$= -(\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])t + \text{const.} \qquad 2.17$$

where V_e is the steady voltage at equilibrium.

Values of V at differing time intervals were read off the oscilloscope trace and equation 2.17 then used to obtain the necessary rate constants (cf. treatment of equation 2.11).

Rate constants obtained from stopped flow measurements were almost always reproducible to within 5%.

4. RESULTS

2-Nitropropane:

The rates of deprotonation of 2-nitropropane (2NP) were obtained by zero-order scavenging of the anion with iodine; the reactions were carried out in solutions with [2NP] in the range $5 \times 10^{-3} - 5 \times 10^{-2}$M, $[I^-]$ was 0.01M and initial iodine concentrations were $\leq 10^{-4}$M. The rates were measured in DMSO-H_2O mixtures using acetate, monochloroacetate and fluoride buffers with anion concentrations normally in the range $2 \times 10^{-2} - 10^{-1}$M. Fluoride buffers with $[F^-]/[HF]$ ca.40 (pH = 4.7 in water) were used to avoid complications from HF_2^- formation; monochloroacetate buffers had $[ClOAc^-]/[HClOAc] = 50$, the iodination being reversible under conditions of low pH. The iodination of 2NP is known to be reversible, in fact, when either $[I^-]$ or $[H^+]$ is high[72] but under the conditions used here equation 2.14 was found to hold over at least 90% of reaction.

In the solvent mixtures containing DMSO the first-order water and hydroxide catalytic constants (i.e. k_o' of equation 2.14) were found to be negligible compared with $k_B[B^-]$. The results are presented in Tables 2.7-2.10.

Our value for k_{OAc} in aqueous solution of 2.71×10^{-6} dm^3 mol^{-1} s^{-1} is in close agreement with Davies' value[73] of 2.64×10^{-6} dm^3 mol^{-1} s^{-1} but differs from that of 3.02×10^{-6} dm^3 mol^{-1} s^{-1} reported by Bell and Goodall[74]. Both groups used zero-order iodinations as in the present study; however, the sample of 2NP used by the latter contained reactive nitroethane impurity which complicated the determination of an accurate k_{OAc} value.

Nitroethane:

Rates of proton transfer from nitroethane (NE) to acetate ion were measured in DMSO-, trifluoroethanol- and acetonitrile-water mixtures. The experimental methods and conditions were the same as for 2NP; substrate concentrations were 0.146M. Results are given in Tables 2.11-2.14 together with equilibrium constants for the process, calculated from equation 2.10.

1,1-Dinitroethane:

The reaction of 1,1-dinitroethane (DNE) with acetate ion in DMSO-H_2O mixtures was followed by observing the appearance of the dinitroethane anion at 385nm using stopped flow spectrophotometry. [DNE] was ca. 5×10^{-5}M and $[OAc^-]$ was in the range 0.02 to 0.10M with $[OAc^-] = 2[HOAc]$.

The integrated rate equation used to calculate k_{OAc} was:

$$\ln(\ln\frac{V}{V_e}) = -\{k_{OAc}[OAc^-](1 + \frac{K_{HOAc}}{2K_{DNE}})\}t + \text{const.}$$

(cf. equation 2.17). This relationship was found to hold over at least four half lives. Kinetic and thermodynamic data for the reaction between dinitroethane and acetate ions are given in Tables 2.15-2.16 together with the corresponding results for nitroethane.

Ethyl-2-methylacetoacetate:

Catalytic constants, k_{OAc} and k_F, were obtained for proton transfers from ethyl-2-methylacetoacetate (EMAA) to acetate and fluoride ions respectively. Zero-order iodinations were used throughout with [EMAA] = 9.77×10^{-4}M, [I^-] = 0.01M and initial [I_2] ca. 10^{-4}M. Base concentrations ranged from 0.01 to 0.10M with [OAc^-]/[HOAc] = 2 and [F^-]/[HF] ≈ 40. Allowance was always made when calculating k_B for the EMAA consumed during reaction (commonly about 10%).

Under conditions such as the above equation 2.14 was found to hold over at least 70% of reaction. A small rapid loss of iodine at the beginning of reaction was interpreted in terms of a few parts per cent of ethylacetoacetate impurity[75].

The value of k_{OAc} in aqueous solution was also determined using first-order scavenging with bromine. [Br^-] was 0.01M, [EMAA] was ca. 10^{-3}M and initial [Br_2] ca. 3×10^{-3}M. The value for k_{OAc} of 4.31×10^{-3} dm mol^{-1} s^{-1} so obtained was in excellent agreement with the value of 4.26×10^{-3} dm^3 mol^{-1} s^{-1} obtained by the above zero-order scavenging technique. Both values, however, are somewhat higher than the previously reported value of 3.23×10^{-3} dm^3 mol^{-1} s^{-1}.[76]

Results for EMAA are presented in Tables 2.17-2.19.

1-(3' methoxyphenyl),1-nitroethane:

The rates of reaction of this substrate (m-OMe) with hydroxide ion in DMSO-H_2O mixtures were obtained spectrophotometrically by observing the rate of production of anion. λ_{max} for the anion absorption was very sensitive to the DMSO content of solution and consequently reactions were followed at 285nm using water and 50/50 (v/v) DMSO-H_2O mixtures but at 325nm for 80/20 (v/v) DMSO-H_2O mixtures.

The [m-OMe] used in reactions was ca. 6×10^{-4}M and $[OH^-]$ ranged from 10^{-2} - 10^{-3}M. Stopped flow spectrophotometry was used for all DMSO-H_2O mixtures. The rate law in all cases (cf. equation 2.11 and related discussion) was of the form:

$$-\frac{d[\text{m-OMe}]}{dt} = \frac{d[\text{m-OMe}^-]}{dt} = k_{OH}[OH^-][\text{m-OMe}]$$

(where m-OMe^- is the nitroalkane anion) and the corresponding first-order plots (cf. equations 2.12 and 2.17) were linear over at least five half lives.

Rates were also measured for reaction with acetate and monochloroacetate ions in 80/20 (v/v) DMSO-H_2O mixtures. Zero-order scavenging with iodine was used throughout with $[I^-] = 0.01$M, initial $[I_2]$ ca. 10^{-4}M and [m-OMe] = 7.1×10^{-3}M. Base concentrations varied from 10^{-2} to 10^{-3}M with $[OAc^-]/[HOAc] = 4$ and $[ClOAc^-]/[HClOAc] = 50$. High buffer ratios were needed because the iodinations were reversible under conditions of high $[H^+]$. Also, in the case of monochloroacetate proton-transfer reactions, it was necessary to reduce $[I^-]$ to 5×10^{-3}M before satisfactory scavenging could be obtained. In the latter case ε^* had to be determined and, using the procedure already outlined, the value in such 80/20 (v/v) DMSO-H_2O mixtures with $[I^-] = 5\times10^{-3}$M was found to be 2.185×10^4 mol^{-1} dm^3 cm^{-1}.

For all scavenging reactions equation 2.14 was found to hold over at least 90% of reaction. It was not possible to determine rates of reaction in aqueous solution, however, because the iodinations were then too reversible.

Kinetic and thermodynamic results are presented in Tables 2.20 and 2.22. The value for k_{OH} of 8.48 $dm^3\ mol^{-1}\ s^{-1}$ in water is in excellent agreement with that of 8.46 $dm^3\ mol^{-1}s^{-1}$ reported by Bordwell et al[37].

1-(4' nitrophenyl),1-nitroethane:

The rates of proton transfer from this species (p-NO_2) to hydroxide ion using DMSO-H_2O mixtures were measured by methods identical with the m-OMe compound. [p-NO_2] used was ca. 2×10^{-5}M and the wavelength adopted was 383nm in water and 50/50 (v/v) DMSO-H_2O mixtures but 450nm in 80/20 (v/v) DMSO-H_2O mixtures.

The rate constant for the deprotonation of p-NO_2 by acetate ion in 80/20 (v/v) DMSO-H_2O mixtures was also measured. The proportion of nitroalkane ionised at equilibrium was sufficiently high to allow the reaction to be followed by observing the production of nitroalkane anion. This was followed spectrophotometrically using a Gilford 2400S instrument and a wavelength of 460nm. Acetate concentrations were $\geqslant 10^{-2}$M with $[OAc^-]/[HOAc] = 4$ and [p-NO_2] ca. 5×10^{-4}M. The integrated rate equation used for determination of k_{OAc} was:

$$\ln(A_e - A) = -\{k_o' + [OAc^-]k_{OAc}(1 + \frac{K_{HOAc}}{4K_{p\text{-}NO_2}})\}\,t + c.$$

The symbols have the same meaning as in equation 2.12. The above relationship was found to hold over at least three half-lives of the reaction.

No rates of proton transfer to acetate or monochloroacetate could be determined in aqueous solution because of the reversibility of the iodination reaction.

Results for p-NO_2 are presented in Tables 2.21 and 2.22. The value determined for k_{OH} in aqueous solution of 61.0 $dm^3\ mol^{-1}\ s^{-1}$ is in close agreement with the value of 57.9 $dm^3\ mol^{-1}\ s^{-1}$ reported by Bordwell et al.[37]

TABLE 2.6 - VALUES OF EXTINCTION COEFFICIENTS FOR IODINE SOLUTIONS IN DMSO-H_2O MIXTURES AT 25°C, USING [I^-] = 0.01M AND A WAVELENGTH OF 353nm

Vol.% DMSO in solvent	0	50	60	70	80	85	90	95
10^{-4} ε*[a] (mol^{-1} dm^{+3} cm^{-1})	2.314	2.376	2.388	2.254	2.240	2.231	2.221	1.873

a Calculated from equation 2.16

TABLE 2.7 - REACTION OF 2-NITROPROPANE (2NP) WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$10^2[OAc^-]$ (mol dm^{-3})	$10^2[2NP]$ (mol dm^{-3})	k_{obs} [a] (s^{-1})	10^5k_{OAc} [b] (dm^3 mol^{-1} s^{-1})
0	2.00	6.53	6.02×10^{-8}	0.271
	4.00	6.95	11.4×10^{-8}	
	6.00	5.91	16.5×10^{-8}	
	10.00	2.62	22.6×10^{-8}	
50	1.90	0.72	4.89×10^{-7}	2.56
	3.90	0.72	9.8×10^{-7}	
	5.80	0.72	14.8×10^{-7}	
	7.80	0.72	20.5×10^{-7}	
	9.80	0.72	25.5×10^{-7}	
60	1.90	0.60	9.87×10^{-7}	5.56
	3.90	0.60	21.1×10^{-7}	
	5.80	0.60	31.8×10^{-7}	
	7.80	0.60	43.4×10^{-7}	
	9.80	0.60	48.7×10^{-7}	
70	1.90	0.68	2.95×10^{-6}	17.95
	3.90	0.68	5.90×10^{-6}	
	5.80	0.68	10.4×10^{-6}	
	7.80	0.68	14.3×10^{-6}	
	9.80	0.68	17.5×10^{-6}	
80	1.90	0.70	12.4×10^{-6}	54.7
	3.90	0.70	22.1×10^{-6}	
	5.80	0.70	35.4×10^{-6}	
	7.80	0.70	41.9×10^{-6}	
	9.80	0.70	52.3×10^{-6}	
90	1.90	0.64	2.06×10^{-4}	633
	3.90	0.64	3.98×10^{-4}	
	5.80	0.64	5.50×10^{-4}	
	7.80	0.64	6.73×10^{-4}	
95	1.00	0.70	1.85×10^{-4}	1470
	1.90	0.70	3.58×10^{-4}	
	3.90	0.70	6.60×10^{-4}	
	5.80	0.70	8.90×10^{-4}	
	7.80	0.70	11.4×10^{-4}	

Notes continued:

TABLE 2.7 (contd.)

a Calculated from equation 2.14 where $k_{obs} = (k_o' + k_{OAc}[OAc^-])$. The required values of ε^* at 0.01M KI are obtained from Table 2.6.

b From plot of k_{obs} against $[OAc^-]$

TABLE 2.8 - REACTION OF 2-NITROPROPANE (2NP) WITH MONOCHLOROACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	[ClOAc$^-$] (mol dm^{-3})	[2NP] (mol dm^{-3})	10^7 k_{ClOAc} [a] (dm^3 mol^{-1} s^{-1})
0	0.376	0.063	2.7
30	0.375	0.169	6.1
50	0.374	0.169	15.0
70	0.190	0.169	78.0
90	0.083	0.169	910

[a] Calculated from equation 2.14 using values of ε^* given in Table 2.6 and ignoring k_o'

TABLE 2.9 - REACTION OF 2-NITROPROPANE (2NP) WITH FLUORIDE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$10^2[F^-]$ (mol dm^{-3})	10^2[2NP] (mol dm^{-3})	k_{obs} [a] (s^{-1})	$10^6 k_F$ [a] (dm^3 mol^{-1} s^{-1})
0	3.99	8.54	1.49×10^{-8}	0.255
	7.96	7.53	2.49×10^{-8}	
	11.94	10.04	3.35×10^{-8}	
	15.92	5.94	4.63×10^{-8}	
	19.90	6.94	5.55×10^{-8}	
50	2.00	0.72	1.58×10^{-7}	9.00
	4.02	0.72	3.01×10^{-7}	
	6.02	0.72	5.18×10^{-7}	
	8.03	0.72	7.25×10^{-7}	
	10.03	0.72	9.47×10^{-7}	
70	2.00	0.68	2.53×10^{-6}	146
	4.02	0.68	5.14×10^{-6}	
	6.02	0.68	8.74×10^{-6}	
	8.03	0.68	11.44×10^{-6}	
	10.03	0.68	15.56×10^{-6}	
80 [b]	2.00	0.70	1.61×10^{-5}	781
	4.02	0.70	3.56×10^{-5}	
	8.03	0.70	7.04×10^{-5}	
	10.03	0.70	8.78×10^{-5}	

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[a] See footnotes to Table 2.7

[b] It was found impossible to study solutions of > 80 vol % DMSO because of the low solubility of KF in such media. The use of $(CH_3)_4NF$ rather than KF was equally as unsuccessful

TABLE 2.10 - COLLECTED DATA FOR PROTON TRANSFERS FROM 2-NITROPROPANE TO ACETATE, MONOCHLOROACETATE AND FLUORIDE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$6+\lg k_{OAc}$	$7+\lg k_{ClOAc}$	$7+\lg k_F$	$pK_{2NP}-pK_{HOAc}$ [b]	$pK_{2NP}-pK_{HClOAc}$ [b]	$pK_{2NP}-pK_{HF}$ [b]
0	0.43	0.43	0.42	3.0	4.9	4.6
50	1.41	1.18	1.95	3.7	5.3	4.6
70	2.25	1.89	3.16	3.9	5.9	4.3
80	2.74	-	3.89	4.1	6.5	3.9
90	3.80	2.96	-	-	-	-
100	-	-	-	4.2	-	-

a For the equilibrium $SH + B^- \overset{K}{\rightleftharpoons} S^- + BH$, $K = K_{SH}/K_{BH}$ where K_{SH} is the dissociation constant for substrate and K_{BH} that for conjugate acid BH (cf. equation 2.10). Therefore: $-\lg K = pK = pK_{SH} - pK_{BH}$. Values of pK are quoted in the table

b Obtained from Tables 1.18 and 1.22

TABLE 2.11 - REACTION OF NITROETHANE (NE) WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	10[NE] (mol dm^{-3})	10^2[OAc$^-$] (mol dm^{-3})	k_{obs} [a] (s^{-1})	$10^5 k_{OAc}$ [a] (dm^3 mol^{-1} s^{-1})
	3.00	2.00	7.30×10^{-7}	
	3.00	4.00	13.8×10^{-7}	
0	3.00	6.00	20.0×10^{-7}	3.26
	3.00	8.00	27.1×10^{-7}	
	3.00	10.00	33.6×10^{-7}	
	1.46	2.00	2.58×10^{-6}	
	1.46	4.00	5.32×10^{-6}	
30	1.46	6.00	6.21×10^{-6}	9.94
	1.46	8.00	9.11×10^{-6}	
	1.46	10.00	10.3×10^{-6}	
	1.46	2.00	8.65×10^{-6}	
	1.46	4.00	17.4×10^{-6}	
50	1.46	6.00	26.7×10^{-6}	42.1
	1.46	8.00	34.4×10^{-6}	
	1.46	9.00	37.8×10^{-6}	
	1.46	2.00	7.29×10^{-5}	
	1.46	3.00	10.00×10^{-5}	
70	1.46	4.00	12.7×10^{-5}	295
	1.46	5.00	15.8×10^{-5}	
	1.46	6.00	18.6×10^{-5}	

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[a] See footnotes to Table 2.7

TABLE 2.12 - REACTION OF NITROETHANE (NE) WITH ACETATE ION IN TRIFLUOROETHANOL-WATER MIXTURES AT 25°C

[NE] = 0.146 mol dm^{-3}

Vol.% TFE in solvent	10^2[OAc$^-$] (mol dm^{-3})	10^6[k_{obs}]^a (s^{-1})	10^5[k_{OAc}] ^{a} (dm^3 mol^{-1} s^{-1})
25	2.00	1.44	2.89
	4.00	2.06	
	6.00	2.59	
	8.00	3.23	
50	2.00	1.26	1.68
	4.00	1.62	
	6.00	1.94	
	7.00	2.10	
	8.00	2.27	
65	1.00	1.01	1.49
	2.00	1.17	
	4.00	1.44	
	6.00	1.78	
80	1.00	0.80	0.86
	2.00	0.88	
	3.00	0.96	
	4.00	1.06	

a Calculated as shown in Table 2.7.
ε^* is 2.314 x 10^4 mol^{-1} dm^3 cm^{-1} throughout

TABLE 2.13 - REACTION OF NITROETHANE (NE) WITH ACETATE ION IN ACETONITRILE-WATER MIXTURES AT 25°C

[NE] = 0.146 mol dm^{-3}

Vol.% MeCN in solvent	10^2[OAc$^-$] (mol dm^{-3})	10^6 k_{obs} [a] (s^{-1})	10^5 k_{OAc} [a] (dm^3 mol^{-1} s^{-1})
30	10.00	5.61	5.61
50	2.00	1.14	5.91
	4.00	2.37	
	6.00	3.50	
	8.00	5.87	
65	6.00	4.72	7.89
80	1.00	1.00	9.92
	2.00	2.07	
	3.00	9.72	
	4.00	13.88	

a Calculated as shown in Table 2.7 with $\varepsilon^* = 2.314 \times 10^4$ mol^{-1} dm^3 cm^{-1}

TABLE 2.14 - COLLECTED DATA FOR PROTON TRANSFERS BETWEEN NE AND ACETATE ION IN DMSO- AND TFE-H_2O MIXTURES[a] AT 25°C

Vol.% DMSO in solvent	Vol.% TFE in solvent	5 + lg k_{OAc}	$pK_{NE} - pK_{HOAc}$ [b]
0	0	0.51	4.1
50		1.63	4.2
70		2.47	4.5
100		-	4.4
	50	0.22	4.4
	65	0.17	-
	80	-0.07	4.4

a NE = nitroethane, TFE = trifluoroethanol

b See footnotes to Table 2.10. Values obtained from Tables 1.7, 1.11 & 1.18.

TABLE 2.15 - REACTION OF 1,1-DINITROETHANE (DNE) WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$10^2[OAc^-]$ ($mol\ dm^{-3}$)	$\frac{[OAc^-]}{[HOAc]}$	k_{obs} [a] (s^{-1})	k_{OAc} [b] ($dm^3\ mol^{-1}\ s^{-1}$)
0				2.16 [c]
50	1.60	2.00	0.97	1.43×10^2
	3.20	2.00	2.63	
	4.80	2.00	3.44	
70	1.00	2.00	15.9	3.55×10^3
	2.00	2.00	29.9	
	3.00	2.00	46.7	

a Defined by $k_{obs} = k_o' + k_{OAc}[OA_c^-](1 + \frac{K_{HOAc}}{2K_{DNE}})$

- cf. equation 2.11 and subsequent discussion

b Calculated from plot of k_{obs} against $[OAc^-]$. Values of K_{HOAc} and K_{DNE} are obtained from Table 1.18

c Reference 43

TABLE 2.16 - COLLECTED DATA FOR PROTON TRANSFERS FROM NITROETHANE AND 1,1-DINITROETHANE TO ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

	1,1-Dinitroethane		Nitroethane		
Vol.% DMSO in solvent	pK_{DNE} [a]	$5 + \lg k_{OAc}$	pK_{NE} [a]	$5 + \lg k_{OAc}$	β [b]
0	5.24	5.33	8.81 [c]	0.51	1.35
50	5.30	7.15	10.19	1.63	1.13
70	5.32	8.55	11.67	2.47	0.96
100	6.6	-	16.4	-	-

a See Table 1.18

b $\beta = \dfrac{\lg k_{OAc}(DNE) - \lg k_{OAc}(NE)}{pK_{NE} - pK_{DNE}}$; (cf. equation 2.9)

c Average of pK_{NE} values available

TABLE 2.17 - REACTION OF ETHYL-2-METHYLACETOACETATE (EMAA) WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

[EMAA] = 9.28×10^{-4} mol dm^{-3} [a]

Vol.% DMSO in solvent	10^2[OAc^-] (mol dm^{-3})	k_{obs} (s^{-1})	10^2 k_{OAc} [b] (dm^3 mol^{-1} s^{-1})
	1.97	9.75×10^{-5}	
	3.95	17.9×10^{-5}	
0	5.93	26.7×10^{-5}	0.426
	7.90	36.2×10^{-5}	
	9.80	43.3×10^{-5}	
	2.00	2.76×10^{-4}	
	3.90	5.28×10^{-4}	
30	5.90	7.77×10^{-4}	1.31
	7.90	10.38×10^{-4}	
	9.80	13.14×10^{-4}	
	1.97	0.83×10^{-3}	
	3.95	1.73×10^{-3}	
50	5.93	2.36×10^{-3}	4.18
	7.90	3.30×10^{-3}	
	8.90	3.73×10^{-3}	
	1.97	1.97×10^{-3}	
60	3.95	3.41×10^{-3}	7.98
	5.93	4.95×10^{-3}	
	6.90	5.52×10^{-3}	
	1.00	2.43×10^{-3}	
	2.00	4.71×10^{-3}	
70	3.00	6.88×10^{-3}	22.30
	3.90	9.05×10^{-3}	
	4.90	10.90×10^{-3}	

[a] This represents [EMAA] after correcting for amount consumed during reaction

[b] See footnotes to Table 2.7

TABLE 2.18 - REACTION OF ETHYL-2-METHYLACETOACETATE (EMAA) WITH FLUORIDE ION IN DMSO-H_2O MIXTURES AT 25°C

[EMAA] = 9.41×10^{-4} mol dm^{-3} [a]

Vol.% DMSO in solvent	$10^2[F^-]$ (mol dm^{-3})	k_{obs} [b] (s^{-1})	k_F [b] (dm^3 mol^{-1} s^{-1})
0	2.00	2.28×10^{-5}	3.52×10^{-4}
	3.90	3.90×10^{-5}	
	5.90	3.61×10^{-5}	
	7.90	4.45×10^{-5}	
	9.80	4.91×10^{-5}	
50	2.00	2.15×10^{-4}	9.26×10^{-3}
	3.90	3.55×10^{-4}	
	5.90	5.50×10^{-4}	
	7.90	7.34×10^{-4}	
	9.80	8.68×10^{-4}	
70	2.00	2.50×10^{-3}	1.39×10^{-1}
	3.90	4.81×10^{-3}	
	5.90	8.20×10^{-3}	
	9.80	12.99×10^{-3}	

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[a] This concentration has been corrected for [EMAA] consumed during reaction

[b] See footnotes to Table 2.7

TABLE 2.19 - COLLECTED DATA FOR PROTON TRANSFERS FROM EMAA[a] TO ACETATE AND FLUORIDE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$2 + \lg k_{OAc}$	$2 + \lg k_{F}$	$pK_{EMAA} - pK_{HOAc}$[b]	$pK_{EMAA} - pK_{HF}$[b]
0	-0.37	-1.35	7.7	9.4
30	0.12	-	-	-
50	0.62	-0.03	6.5	7.5
60	0.90	-	-	-
70	1.35	1.14	5.6	6.0

a EMAA = Ethyl-2-methylacetoacetate

b See footnotes to Table 2.10

TABLE 2.20 - REACTION OF 1-(3' METHOXYPHENYL), 1-NITROETHANE (m-OM_e) WITH HYDROXIDE, ACETATE AND MONOCHLOROACETATE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	10^2[OH^-] (mol dm^{-3})	10^2[OAc^-][a] (mol dm^{-3})	10^2[$ClOAc^-$][a] (mol dm^{-3})	k_{obs}[b] (s^{-1})	k_{OH}[c] ($dm^3mol^{-1}s^{-1}$)	k_{OAc}[c] ($dm^3mol^{-1}s^{-1}$)	k_{ClOAc}[c] ($dm^3mol^{-1}s^{-1}$)
	0.50			4.44×10^{-2}			
0	1.00			8.35×10^{-2}	8.48		
	2.00			16.95×10^{-2}			
	0.50			0.461			
50	1.50			1.24	84.6		
	2.50			2.09			
	0.10			4.54			
80	0.15			6.34	4.50×10^{3}		
	0.20			9.62			
		0.50		2.02×10^{-4}			
80		1.00		4.12×10^{-4}		4.17×10^{-2}	
		1.50		6.41×10^{-4}			
			5.00	1.55×10^{-5}			
80			10.00	3.04×10^{-5}			3.04×10^{-4}
			15.00	4.58×10^{-5}			

Footnotes continued:

TABLE 2.20 (contd.)

a $[m\text{-OMe}] = 7.10 \times 10^{-3}$ mol dm^{-3}

b For OH^- : $k_{obs} = k_{OH}[OH^-]$; see equation 2.11 and discussion.

For carboxylate anions: $k_{obs} = (k_o' + k_B[B^-]$; see footnotes to Table 2.7 (For monochloroacetate reaction $[I^-] = 0.005$M and $\varepsilon^* = 2.185 \times 10^4$ mol^{-1} dm^3 cm^{-1})

c Calculated from plots of k_{obs} against $[B^-]$

TABLE 2.21 - REACTION OF 1-(4' NITROPHENYL). 1-NITROETHANE (p-NO_2) WITH HYDROXIDE AND ACETATE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$10^3[OH^-]$ (mol dm^{-3})	$10^2[OAc^-]$ (mol dm^{-3})	k_{obs} [a] (s^{-1})	k_{OH} [b] (dm^{3}mol^{-1}s^{-1})	k_{OAc} [c] (dm^{3}mol^{-1}s^{-1}
0	1.00		4.74×10^{-2}	61.0	
	2.00		11.90×10^{-2}		
	4.00		25.90×10^{-2}		
50	15.0		19.9	1.42×10^3	
	25.0		30.8		
80	1.00		1.13×10^2	1.35×10^5	
	1.50		2.05×10^2		
	2.00		2.85×10^2		
80		1.00	3.23×10^{-2}		1.60
		1.50	5.12×10^{-2}		
		2.00	7.68×10^{-2}		

a For OH^- : $k_{obs} = k_{OH}[OH^-]$

For OAc^- : $k_{obs} = k_o' + k_{OAc}[OAc^-]\left(1 + \frac{K_{HOAc}}{4K_{p\text{-}NO_2}}\right)$

since $[OAc^-]/[HOAc] = 4$

b Calculated from plot of k_{obs} against $[OH^-]$

c Calculated from plot of k_{obs} against $[OAc^-]$; values for K_{HOAc} and $K_{p\text{-}NO_2}$ are obtained from Table 1.18

TABLE 2.22 - COLLECTED DATA FOR PROTON TRANSFERS FROM 1-ARYLNITROETHANES TO HYDROXIDE AND ACETATE IONS IN DMSO-H_2O MIXTURES AT 25°C

	p-NO_2		m-OMe			p-NO_2		m-OMe	
Vol.% DMSO in solvent	pK_a[a]	lg k_{OH}	pK_a[a]	lg k_{OH}	β[b]	pK_a[a]	lg k_{OAc}	pK_a[a]	lg k_{OAc}
0	6.63	1.79	7.40	0.93	1.11[c]				
50	7.40	3.15	8.93	1.93	0.80				
80	8.64	5.13	10.94	3.65	0.64				
80					0.69	8.64	0.20	10.94	-1.38

a See Table 1.18

b Calculated from $\beta = \dfrac{\lg k_B(p\text{-}NO_2) - \lg k_B(m\text{-}OMe)}{pK_a(m\text{-}OMe) - pK_a(p\text{-}NO_2)}$

c This value of β obtained from rates and equilibria for reaction between hydroxide and only two substrates agrees well with the value of 1.14 obtained by Bordwell et al[37], who measured corresponding rates and equilibria for 14 1-arylnitroethanes with a total pK_{SH} variation of 1.08

5. DISCUSSION

Variations of reaction rate with solvent will depend upon the relative effects of solvent upon the free energies of reactants and transition state. The changes for any simple proton-transfer reaction of the type $SH + B^- \rightleftharpoons S^- + BH$ would therefore be expected to depend primarily upon the relative effects of solvent upon anion base and the larger transition state anion. The conclusions reached in Section I of this thesis would, however, predict that the effect of solvent upon anion base free energy should be the predominant factor. This being the case we should expect anion bases of highest charge density to show the greatest increases in reaction rate upon addition of DMSO to aqueous reaction solvent. The results clearly show this to be true (cf. the comparisons in Table 2.10 which show that the changes in reaction rate for proton transfers from 2-nitropropane to various anions increase in the order $ClOAc^- < OAc^- << F^-$). It should also be noted that for trifluoroethanol we expect a decrease, and for acetonitrile an increase, in reaction rate (for a given proton transfer) upon addition to any aqueous reaction solvent. Both are observed (see Tables 2.13 and 2.14).

We shall now turn our attention to the relationship between equilibria and rates for the proton-transfer reactions in mixed solvent systems and, in particular, to the Brönsted relationship in such media. It has already been shown that the Brönsted β coefficients relating the effect of variations in the nature of B^- or SH on the rate and equilibria of reaction is given by

$$\beta = \frac{\delta \lg k}{\delta \lg K} = \frac{\delta \Delta G^{\ddagger}}{\delta \Delta G^{o}} \qquad \text{(eqn. 2.9)}$$

where k is the rate constant for proton transfer from substrate to base and K ($K = K_{SH}/K_{BH}$) is the equilibrium constant for the reaction in question.

It has been suggested that Brönsted coefficients can also be obtained by varying the medium in which the reaction is carried out[76] and obtaining rates and equilibria in the new solvent. This approach has been used in several cases, usually with DMSO-H_2O mixtures and using values of pK_{SH}-pK_{BH} estimated from measured H_- acidity functions (see, for example, reference 77). Interpretations of these Brönsted coefficients assume that the effects of solvent on the free energy of the system varies monotonically with the extent of reaction. Results presented here, however, would indicate that this cannot be the case. Tables 2.10 and 2.14, for example, show that, although the pK_a's of both substrates and conjugate acids vary considerably with increasing amounts of DMSO (or trifluoroethanol) in the solvent, there is very little change in the overall equilibrium constant for the reaction (expressed as pK_{SH}-pK_{BH}). Despite this, the rates increase rapidly with increasing content of the organic component in DMSO-H_2O mixtures and decrease in trifluoroethanol-H_2O mixtures.

The above behaviour can be readily explained in terms of the large differences in the interaction of water (relative to dipolar aprotic solvents such as DMSO) with anions of high charge density (e.g. F^-, RCO_2^-, RNO_2^-) and large polarisable anions such as the transition state anions. Thus, the addition of DMSO to water raises the free energy of acetate and introalkane anions by a similar amount relative to the transition state anion, leading to large rate increases but relatively small changes in equilibrium constant. Trifluoroethanol being more acidic than water would be expected to stabilise the smaller anions even further relative to the transition state anions and so cause reductions in rate of reaction, as observed. Effects such as these prevent the effect of solvent on the free energy of the system from varying monotonically with the extent of proton transfer in the reaction. Such behaviour is perhaps not unexpected since it has been known for some time that the rates of the reactions $^*X^- + RX \rightleftharpoons RX^* + X^-$ (where X and X* are isotopes) are very sensitive to the nature of the solvent[3] although the equilibrium constant must be independent of the reaction medium. It is clear, therefore, that interpretation of Brönsted coefficients obtained by solvent variation is a complicated matter.

We shall now return to Brönsted β coefficients obtained by the traditional methods of keeping substrate constant and altering the base (the same solvent being used throughout) or vice-versa. For the system to exhibit a "normal" β value (i.e. $0 \leq \beta \leq 1$) it is necessary for the effects of the substituent on the free energy of the system to vary monotonically with the extent of reaction[78]. A value of $\beta > 1$ indicates that the effect of the substituent on $\Delta G^{\dagger}$ is greater than ΔG°. The relevance of solvation effects, such as those discussed above, to the explanation of "anomalous" Brönsted coefficients found in water can be seen from the results presented in Tables 2.16 and 2.22. Considering, for example, the results shown in Table 2.16: as the fraction of DMSO in the solvent is increased so the difference in pK_a of dinitroethane and nitroethane increases going from 3.6 pK_a units in water to 9.8 pK_a units in DMSO. This presumably reflects the greater loss in solvation energy of the mononitroalkane anion relative to the dinitroethane anion with its more highly dispersed charge*, as pointed out in Section I of the thesis. Similar free energy variations are not expected for the transition states since, for both reactions, these have highly dispersed charges. Thus, the β values decrease and become "normal" as the DMSO content of the solvent increases. The results in Table 2.22 can be explained in an analogous manner.

The above explanation of anomalous Brönsted coefficients is shown schematically in Figure 2.2, where curve I represents nitroethane (or the m-OMe substituted 1-arylnitroethane) and curve II represents the dinitroethane (or p-NO_2 substituted 1-arylnitroethane). The full lines indicate the reaction in, for example, the gas phase and the dotted lines in water. For simplicity the curves in the two phases have been moved relative to each other so that the energy of the reactants can be aligned; the solvent is

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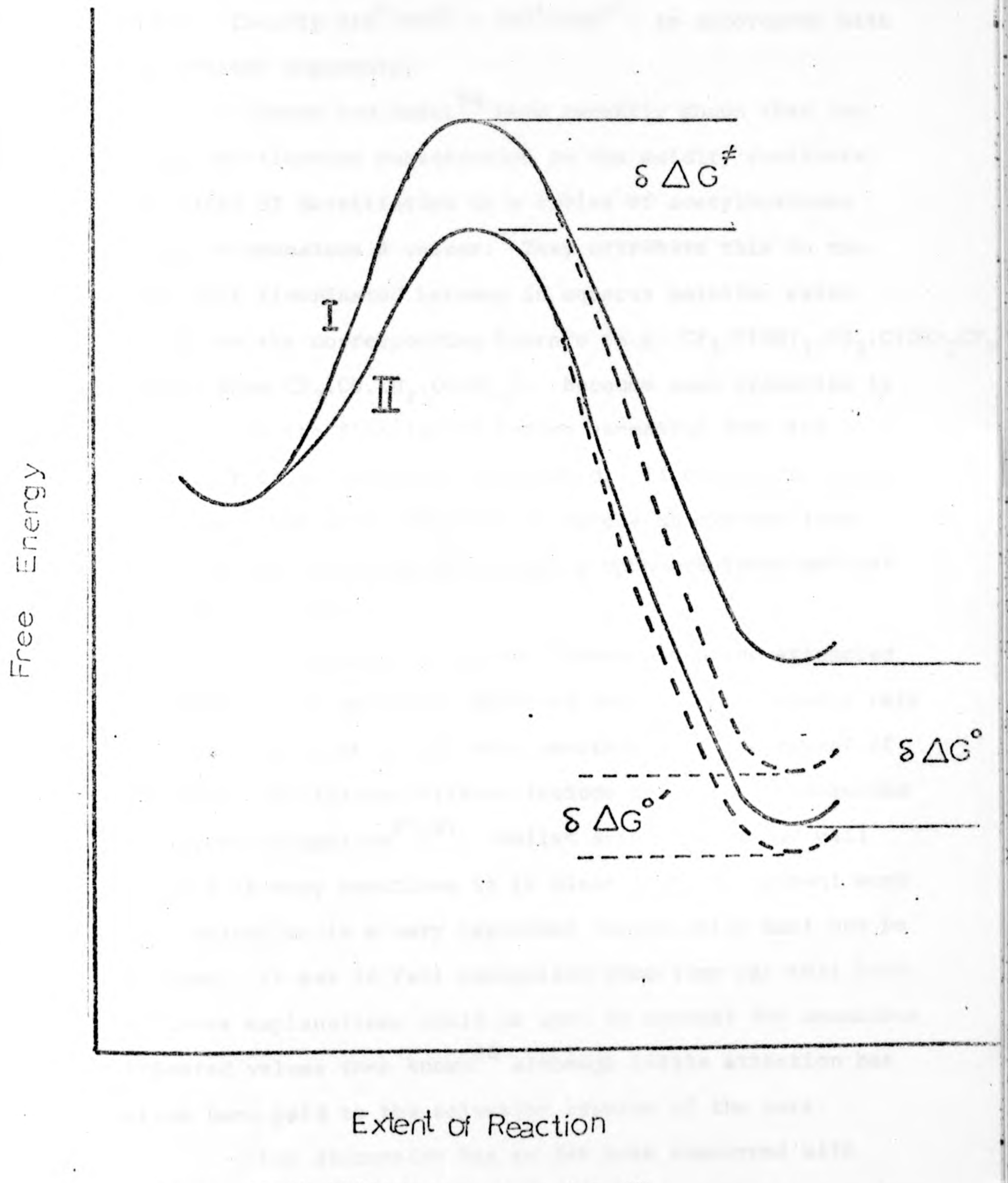
* Similar factors are presumably responsible for the ΔpK_a in the gas phase between, for example, dichloroacetic and acetic acids being ca.15 compared with 3.5 in water.[79]

FIGURE 22 - Brönsted coefficients and solvent effects



shown as having no effect on the free energy of the transition state. Clearly $\delta\Delta G^{\ddagger}/\delta\Delta G^{o} < \delta\Delta G^{\ddagger}/\delta\Delta G^{o\prime}$, in accordance with the present arguments.

Jones and Patel[80] have recently shown that the effect of fluorine substitution on the acidity constants and rates of detritiation in a series of acetylacetones leads to anomalous β values. They attribute this to the fact that fluorinated ketones in aqueous solution exist mainly as the corresponding hydrate (e.g. $CF_3.C(OH)_2.CH_2.C(OH)_2CF_3$ rather than $CF_3.CO.CH_2.CO.CF_3$); because such hydration is specific to the fluorinated ketone anomalous Brönsted behaviour is not entirely unexpected. It could, of course, be argued that such hydration is merely an extreme form of solvation operating specifically upon reactants and not transition state.

As pointed out in the introduction the attempted explanations of anomalous Brönsted coefficients usually rely upon effects which do not vary smoothly with the extent of reaction. Postulated effects include induction, mesomerism and hyperconjugation[67,81]. Whilst such effects may well operate in many reactions it is clear from the present work that solvation is a very important factor which must not be ignored. It was in fact recognised some time ago that both of these explanations could be used to account for anomalous Brönsted values then known[68] although little attention has since been paid to the solvation aspects of the work.

Our discussion has so far been concerned with the contribution of solvation to anomalous Brönsted behaviour (those cases in which $0 > \beta > 1$). However, it is equally likely that these solvation effects occur, albeit to a lesser extent, even when "normal" behaviour is observed. This would apply equally when the substrate is held constant and the base varied as well as when the base is altered with a constant substrate. Clearly then such factors must be borne in mind before attempting any detailed interpretations of β values in terms of transition state structure etc. (see, for example, page 185 of reference 76).

NOTE: Much of the work reported in this part of the thesis has already been published (B. G. Cox and A. Gibson J.C.S. Chem.Comm. 1974 638, B. G. Cox and A. Gibson, in press).

PART C

PRIMARY KINETIC HYDROGEN ISOTOPE EFFECTS
IN PROTON-TRANSFER REACTIONS

1. INTRODUCTION

Following the discovery of deuterium in 1932[82] and tritium in 1934[83] it was shown that replacement of hydrogen by either of these isotopes could frequently cause major changes in the equilibrium, and particularly in the velocity, of reactions in which bonds involving these species are made or broken. Since then many hundreds of papers dealing with such matters have been published (for general reviews see references 84-87). In this thesis we are primarily concerned with the kinetics of proton/deuteron transfer reactions of the type $SL + B^- \rightleftharpoons S^- + BL$ (where $L = H$ or D). The ratios of the rates of proton and deuteron transfer from SL to B^- (i.e. k^H/k^D) are called primary kinetic hydrogen isotope effects.

Kinetic hydrogen isotope effects are most conveniently discussed in the framework of transition state theory[65]. It is assumed that the reactants come together to form a transition state located at the top of an energy barrier to reaction which is at the summit of a pass connecting reactants and products*. The transition state is regarded as a normal molecule in equilibrium with reactants and products except that one of its vibrations is replaced by an internal translation.

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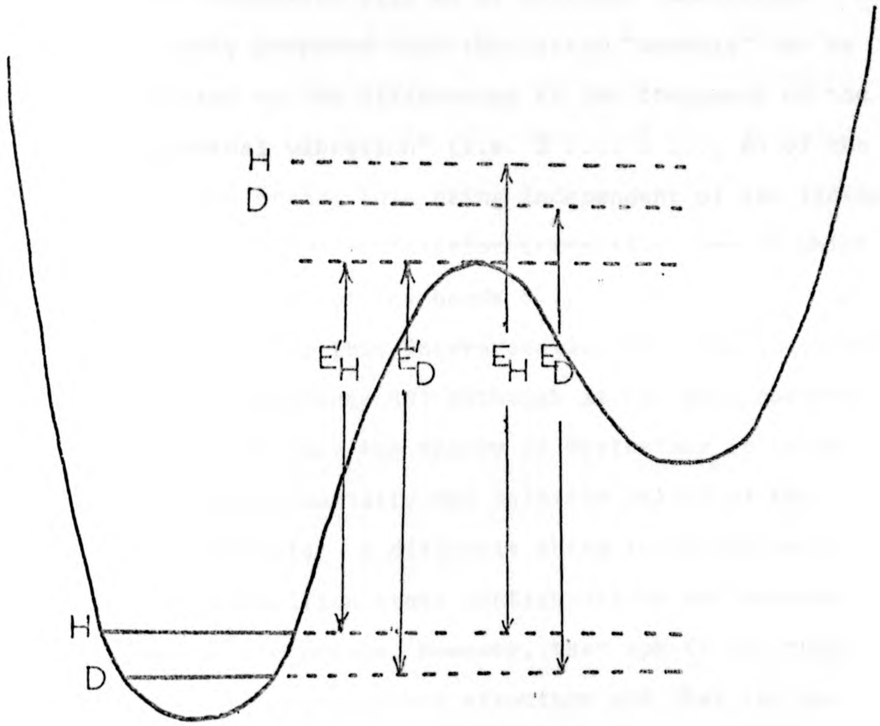
* Such a model has been accepted without comment in the previous part of the thesis.

The variation of the free energy of the system with the extent of reaction for a proton transfer of the type $SL + B^- \rightleftharpoons BL + S^-$ (where L = H or D) is represented diagrammatically in Figure 2.3. Because intermolecular and interatomic forces are unaffected by changes in nuclear mass the same curve has been used for both hydrogen isotopes. The lowest energy levels of the initial and final states will depend upon the isotopic mass as illustrated by the zero-point energies for H and D indicated in the figure. Ordinarily, it might be supposed that the activation energies for transfer of proton or deuteron from SL to B^- would be E_H' and E_D' respectively. However, this ignores the zero-point energies of the corresponding transition states arising as they do from vibrations dependent upon the mass of the isotopic species being transferred. Therefore, the activation energies are correspondingly increased and are represented by E_H and E_D. (For a full discussion of all of these points see Chapter 12 of reference 48).

The above semi-classical treatment would predict that the maximum permitted isotope effect would arise when the reactant species are such that the zero-point energies of the two transition states are approximately equal. On this basis the maximum isotope effect which should be observed at ordinary temperatures is ca. 10[87]. However, many values of $k^H/k^D \geq 10$ (at 298°K) are known and there are also many cases where unexpectedly large variations in k^H/k^D are observed for only small structural changes in either base or acid. Both of these apparent anomalies can be explained

FIGURE 2.3 - Zero-point energies and hydrogen isotope effect for a proton-transfer reaction



on the basis of quantum mechanical tunnelling of the proton (an effect which is more important for the proton than the deuteron)[48,87,88]. This explanation leads to the prediction that, for a given system, the maximum isotope effect should occur at ΔG = 0 since at this point the tunnel correction will be at maximum. Westheimer[89] had previously proposed that the latter "anomaly" can be attributed to the differences in the frequency of the "symmetrical vibration" (i.e. $\overleftarrow{S}$ .... $\overset{?}{L}$ .... $\vec{B}$) of the transition state, this being independent of the isotope (L) for a symmetrical transition state (i.e. one in which the force constants of the bonds S.... L and L .... B are equal). The validity of this interpretation has been questioned (page 270 reference 48) although it has been assumed by many authors. To test the theory of Westheimer it is necessary to vary experimentally the relative values of the two pertinent force constants: a difficult thing to do because we have no idea of transition state configurations or frequencies. It is usual[74] to assume, however, that ΔpK (= pK_{SH}-pK_{BH}) will reflect transition state structure and that the maximum isotope effect should occur at ΔpK (i.e. ΔG^o) = 0.

Both of the above two theories have led people to seek correlations between k^H/k^D and ΔG^o (or pK) for the reaction and, in some cases, there is evidence for a maximum isotope effect in the neighbourhood of $\Delta pK = 0$. The variations in ΔpK are commonly caused by introducing substituents into one, or both, of the reactants; studies so far reported include proton-abstraction by bases from carbonyl compounds[75,90] and nitroalkanes[43,74,91] and, also base-catalysed elimination reactions[92]. In other investigations ΔpK for a given reaction has been changed by varying the composition of mixtures of DMSO with water or alcohols. This procedure has, for example, been applied to the inversion of (-)-menthone[77] and ionization of nitroethane[93], both in the presence of hydroxide ions.

The work reported here investigates further the relationships between k^H/k^D and ΔpK and, in particular, the variation of both parameters with solvent changes. Reactions have been carried out using several substrates and bases in DMSO-H_2O, trifluoroethanol-H_2O and acetonitrile-H_2O mixtures at 25°C.

2. MATERIALS

Unless otherwise stated all materials were prepared and/or purified by methods given elsewhere in this thesis. All deuterated materials were stored in a refrigerator when not in use.

Acetone: $(CH_3)_2CO$	Koch-Light Ltd. AnalaR grade material was used without further purification. No proton-containing impurities were revealed by nmr analysis; glc analysis showed >99% purity.
2-nitropropane-2-d_1 $(CH_3)_2CDNO_2$	Base (OAc^-) catalysed exchange of 2NP with D_2O in DMSO was used here. 25 cm^3 of D_2O (≥99.8%D; 2.5 moles), 10 cm^3 2NP (0.11 moles),

50 cm^3 DMSO and 2.25g AnalaR sodium acetate (0.03 moles) were heated at 70°C for three hours. The 2NP-2-d_1 was extracted with 2 x 25 cm^3 volumes of ether; the combined extracts were dried with Na_2SO_4 and the ether removed by distillation. Nmr analysis of the resultant material showed that 94% deuteration had been achieved (70% recovery). Small amounts of ether and DMSO were present as impurities. This sample of deuterated 2NP was then exchanged a second time using 2g sodium acetate, 20 cm^3 D_2O and 40 cm^3 DMSO. Separation of the final deuterated material was achieved by the same methods as before and in the fractional distillation the fraction boiling at 119°C collected; a 50% recovery was achieved in this final exchange. Nmr analysis of the sample showed it to be a mixture of 2NP-2d_1 (97.6%) and 2NP (2.4%).

Nitroethane-1-d_2:
$CH_3CD_2NO_2$

The method used was basically the same as that reported by Bell and Goodall[74]. 35 cm^3 of D_2O (≥99.8%D; 1.75 moles), 5 cm^3 nitroethane (0.067 moles) and 0.03g potassium hydroxide (5.3 x 10^{-4} moles) were refluxed for 24 hours. The organic layer was extracted with ether, dried over $MgSO_4$ and the ether removed by distillation. The resulting sample (88%-CD_2-) was exchanged a second time and again extracted with ether.

In the final distillation of the combined organic extracts the fraction boiling at 118°C was collected for use. Nmr analysis showed only $CH_3CD_2NO_2$ to be present.

Ethyl-2-methyl-acetoacetate-2-d_1:
$CH_3.CO.CDMe.CO_2Et$

The method used was identical to the uncatalysed exchange process reported by Bell and Crooks[75]. An approximately 0.10M solution of the ester in D_2O ($\geq$99.8%D) was prepared and left for four days at 35°C. Nmr analysis of the resulting solution showed that 99% deuteration had been effected.

Acetone-d_6:
$(CD_3)_2CO$

Fluorochem Ltd. acetone-d_6 was used without further purification. Nmr analysis showed $\geq$99.7% purity.

3. EXPERIMENTAL

The rates of ionization of the deuterated compounds were followed by zero-order scavenging with iodine as previously described for the undeuterated species. To minimise errors in the isotope effects rates of deprotonation and dedeuteration were determined simultaneously using identical buffer/solvent mixtures. It was usual to use only one base concentration in determining k_B^H/k_B^D, chosen so that the observed first-order rate constant, k_{obs} (cf. equation 2.14), was given by $k_{obs} = k_B[B^-]$, the contribution of k'_0 being negligible.

The experimental procedure was as follows: equal volumes of buffer in the required solvent were pipetted into the spectrophotometric cells, iodine added and the cells thermostatted as before. Reaction in each case was then initiated by addition of equal volumes of either the H or D compound (the same syringe being used in both cases). This process was repeated at least once, and the mean values of observed first-order rate constants used in calculating k^H/k^D; the values of k_{obs} were invariably reproducible to within 5%. In trifluoroethanol-H_2O mixtures it was found necessary to draw full catalytic plots in order to obtain $k_B{}^D$ accurately because of the magnitude of $k'_0{}^D$ with respect to $k_B{}^D$.

Because 2-nitropropane-2-d_1 and ethyl-2-methyl-acetoacetate-2-d_1 could not be prepared in very high isotopic purity and since the observed values of k^H/k^D were in the range 5-10, it was necessary to consider the contribution made to the observed rate of reaction by the ionization of undeuterated material. This is fairly easily obtained because, if we follow a sufficiently small amount of reaction (as is the case here) then the observed zero-order rate constant, k', is given by[74]:

$$k' = k^H[SH]_o + k^D[SD]_o$$

where $k^H = \sum_B k_B{}^H[B]$, $k^D = \sum_B k_B{}^D[B]$

and $[SH]_o$, $[SD]_o$ are the concentrations of undeuterated and deuterated materials respectively at zero time (cf. equation 2.14). If we carry out the proton transfer with base B^- then equation 2.14 becomes:

$$-\frac{dA}{dt} = \varepsilon^* \left\{ (k_o'^{H} + k_B^{H}[B^-]) \frac{[SH]_o}{[SD]_o} + k_o'^{D} + k_B^{D}[B^-] \right\} [SD]_o \qquad 2.18$$

where A again represents the absorbance (at 353 nm) of the solution at time t and ε^* is the extinction coefficient of I_2 solution at 353 nm and $[I^-]$ of 10^{-2}M. Therefore, knowing k^H and the isotopic purity ($[SH]_o/[SD]_o$), the value of k^D (and thence k_B^D) can be calculated.

Experimental details specific to individual substrates are given in the results section. The values of k^H/k^D for any reaction were, in almost all cases, reproducible to within 5%.

4. RESULTS

2-nitropropane-2-d_1

Values of k^H/k^D were determined in several DMSO-H_2O mixtures using acetate, monochloroacetate and fluoride ions as bases. The base concentrations used to determine k_B^D were usually in the range 7-10 x 10^{-2}M. Equation 2.18 was found to hold over at least 90% of reaction. Results are given in Tables 2.23 and 2.24.

Nitroethane-1-d_2

Primary kinetic hydrogen isotope effects were determined in DMSO-H_2O, trifluoroethanol-H_2O and acetonitrile-H_2O

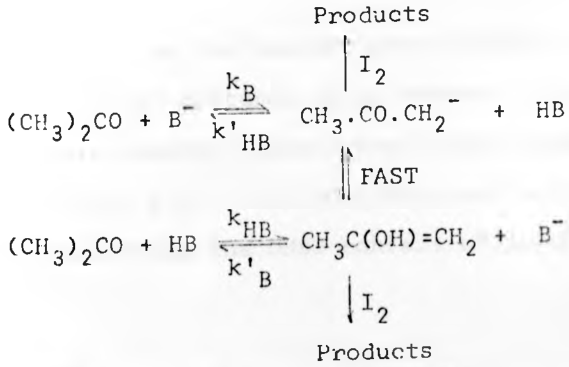
mixtures using acetate ion as the base. Acetate ion concentrations were in the range 2-10 x 10^{-2}M. Results are presented in Tables 2.25-2.27; equation 2.18 held over 90% of reaction in all cases.

Ethyl-2-methylacetoacetate-2-d_1

Results were obtained in DMSO-H_2O mixtures with acetate and fluoride ions. Base concentrations were in the range 2-10 x 10^{-2}M and full catalytic plots were used to obtain $k_B{}^D$ for most solvent mixtures. Equation 2.18 was found to hold over at least 70% of reaction and results are given in Tables 2.28-2.30.

Acetone

The catalytic constants were determined from zero-order scavenging experiments using iodine, as reported for several other carbon acids. The study of acetone, however, is complicated by the fact that acetone can react with both base and acid to produce an equilibrium mixture of the corresponding enol and enolate anion (both of which react with iodine)*. The kinetic scheme can be represented by:



$$\begin{tikzcd}
 & \text{Products} \\
(CH_3)_2CO + B^- \arrow[r, shift left, "k_B"] & CH_3.CO.CH_2^- + HB \arrow[l, shift left, "k'_{HB}"] \arrow[u, "I_2"'] \arrow[d, shift left] \\
(CH_3)_2CO + HB \arrow[r, shift left, "k_{HB}"] & CH_3C(OH)=CH_2 + B^- \arrow[l, shift left, "k'_B"] \arrow[u, shift left, "FAST"'] \arrow[d, "I_2"] \\
 & \text{Products}
\end{tikzcd}$$

---

* For full discussion see chapter 8 of reference 48.

with the enol-enolate species in rapid equilibrium compared with the rate of iodination of the species. Under conditions of zero-order scavenging the rate law becomes:

$$-\frac{d[I_2^*]}{dt} = -\frac{d[SH]}{dt} = (k_o' + k_B[B^-] + k_{HB}[HB])[SH]$$

$$= \left\{k_o' + [B^-](k_B + \frac{k_{HB}}{r})\right\}[SH]$$

where k_o' includes the H_2O, OH^- and H^+ catalysed rate constants at this buffer ratio, r ($r = [B^-]/[HB]$). k_o' was found to be negligible in the DMSO-H_2O mixtures studied.

The observed first-order rate constant for reaction (k_{obs}) can be written:

$$k_{obs} = k_o' + [B^-](k_B + \frac{k_{HB}}{r}) \ .$$

To determine values of k_B the reaction must be studied at, at least, two differing buffer ratios. If k_{obs} is plotted against $[B^-]$ at each buffer ratio, the gradient of each plot will give us the parameters necessary to obtain k_B.

In the present study rate constants were determined in DMSO-H_2O mixtures using acetate ion as the base. The acetate concentrations were in the range 2-10 x 10^{-2}M with r = 4 and 0.67. The contributions of k_{HOAc} to k_{obs} became negligible as the DMSO content of solution increased and for solutions of highest DMSO concentration it was possible to use only one buffer solution (r = 4). [SH] was 2.94×10^{-1}M throughout, $[I^-]$ was 10^{-2}M and initial $[I_2]$ ca. 5×10^{-5}M. Excellent zero-order scavenging was observed in all cases.

It ought to be noted that, in water, a small correction must be applied to the observed first order rate constant, k'_{obs}, to allow for the contribution to it of an observed term taking the form $k[HOAc][OAc^-]$; this has been attributed to the existence of a concerted reaction mechanism in which the transition state contains two catalyst molecules, one acid and one base (see, for example, page 149 of reference 48). If we use the value of $k = 3.33 \times 10^{-7}$ dm^6 mol^{-2} s^{-1} obtained by Bell and Jones[94] then the rate of reaction in water is given by:

$$-\frac{d[I_2^*]}{dt} = -\frac{d[SH]}{dt} = \Big\{k'_o + [B^-](k_{OAc} + \frac{k_{HOAc}}{r}) + 3.37 \times 10^{-7}[HOAc][OAc^-]\Big\}[SH]$$

with $k'_{obs} = k_{obs} + 3.33 \times 10^{-7} [HOAc][OAc^-]$. In all cases $0 \leq (k'_{obs} - k_{obs})/(k'_{obs}) \leq 0.1$. Such corrections were not applied in the DMSO-H_2O mixtures, plots of k_{obs} (no correction applied) against $[OAc^-]$ being linear.

Results are presented in Table 2.31; the value of $k_{OAc} = 25.0 \times 10^{-8}$ dm^3 mol^{-1} s^{-1} found in water is in excellent agreement with Bell and Jones' value[94] of 25.3×10^{-8} dm^3 mol^{-1} s^{-1} and that of 24.3×10^{-8} dm^3 mol^{-1} s^{-1} found by Bell and Lidwell[95].

Acetone-d_6

The experimental procedure used was the same as that for the undeuterated material; full plots of k_{obs} against OAc^- were drawn in all cases. In water the values of k'_{obs} (see previous discussion concerning the undeuterated species) were corrected for $k[HOAc][OAc]$ by the same method used for the H compound. In the absence of a value of k for the deuterated species (and because the resulting corrections are small, anyway) it was assumed that, for solutions of the same $[OAc^-]$ and $[HOAc]$:

$$\left\{\frac{k'_{obs}-k_{obs}}{k'_{obs}}\right\}_H = \left\{\frac{k'_{obs}-k_{obs}}{k'_{obs}}\right\}_D .$$

No corrections were applied to observed first-order rate constants in DMSO-H_2O mixtures, plots of k_{obs} against $[OAc^-]$ again being linear.

Excellent scagenging was observed throughout and results are presented together with those for the H compound in Table 2.31.

TABLE 2.23 - REACTION OF 2-NITROPROPANE (2NP) and 2-NITROPROPANE-2-d_1 WITH ACETATE, MONOCHLOROACETATE AND FLUORIDE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	10^2[2NP]^a (mol dm^{-3})	10^2[2NP2d_1]^{a,b} (mol dm^{-3})	10^2[OAc^-] (mol dm^{-3})	10^2[$ClOAc^-$] (mol dm^{-3})	10^2[F^-] (mol dm^{-3})	k^H_{obs} ^{c} (s^{-1})	k^D_{obs} ^{d} (s^{-1})	$\frac{k^H_{OAc}}{k^D_{OAc}}$	$\frac{k^H_{ClOAc}}{k^D_{ClOAc}}$	$\frac{k^H_F}{k^D_F}$
0					10.00	$2.56\text{x}10^{-8}$	$5.46\text{x}10^{-9}$			4.8
			10.00			$2.78\text{x}10^{-7}$	$2.89\text{x}10^{-8}$	9.6		
				37.6		$1.01\text{x}10^{-8}$	$1.20\text{x}10^{-9}$		8.4	
30	16.90	16.50	40.00			$2.46\text{x}10^{-6}$	$2.50\text{x}10^{-4}$	9.6		
	16.90	16.50		37.5		$2.28\text{x}10^{-7}$	$2.81\text{x}10^{-8}$		8.2	
50	0.681	0.640			7.80	$5.15\text{x}10^{-7}$	$1.12\text{x}10^{-7}$			4.6
	14.40	14.00	9.88			$2.61\text{x}10^{-6}$	$2.78\text{x}10^{-7}$	9.4		
	16.90	16.50		37.4		$5.61\text{x}10^{-7}$	$7.75\text{x}10^{-8}$		7.3	
70	0.681	0.640			7.80	$10.00\text{x}10^{-6}$	$2.12\text{x}10^{-6}$			4.7
	2.454	3.747	7.11			$1.46\text{x}10^{-5}$	$1.89\text{x}10^{-6}$	7.7		
	16.90	16.50		19.0		$1.48\text{x}10^{-6}$	$1.85\text{x}10^{-7}$		8.0	
80	0.681	0.640			7.80	$6.26\text{x}10^{-5}$	$1.29\text{x}10^{-5}$			4.8
	3.514	3.115	6.76			$5.66\text{x}10^{-5}$	$7.86\text{x}10^{-6}$	7.2		
90	0.681	0.640	9.80			$4.28\text{x}10^{-4}$	$6.30\text{x}10^{-5}$	6.7		
	16.90	16.50		8.30		$7.56\text{x}10^{-6}$	$1.21\text{x}10^{-6}$		6.3	
95	0.681.	0.640	7.90			$1.32\text{x}10^{-3}$	$2.20\text{x}10^{-4}$	6.0		

Footnotes continued:

TABLE 2.23 (contd.)

a Concentrations of substrates are only quoted in those cases where the same [SH] was used for all runs at the quoted $[B^-]$

b This is uncorrected for the presence of 2NP

c Defined by $k_{obs} = k_B[B^-]$ and obtained from equation 2.14; the required values of ε^* are given in Table 2.6. The quoted values of k_{obs} are the mean of at least two experimental determinations and have been corrected for any "buffer reaction".

d This represents the first-order constants after correction for 2.4% H compound (cf. equation 2.18).

TABLE 2.24 - EQUILIBRIA AND KINETIC HYDROGEN ISOTOPE EFFECTS FOR PROTON TRANSFERS FROM 2-NITROPROPANE TO ACETATE, MONOCHLOROACETATE AND FLUORIDE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$pK_{2NP}-pK_{HOAc}$ [a]	k^H_{OAc}/k^D_{OAc}	$pK_{2NP}-pK_{HClOAc}$ [a]	k^H_{ClOAc}/k^D_{ClOAc}	$pK_{2NP}-pK_{HF}$ [a]	k^H_F/k^D_F
0	3.0	9.6	4.9	8.4	4.6	4.8
30	-	9.6	-	8.2	-	-
50	3.7	9.4	5.3	7.3	4.6	4.6
70	3.9	7.7	5.9	8.0	4.3	4.7
80	4.1	7.2	6.5	-	3.9	4.8
90	-	6.7	-	6.3	-	-
95	-	6.0	-	-	-	-

[a] Obtained from Tables 1.18 and 1.22

TABLE 2.25 - REACTION OF NITROETHANE (NE) AND NITROETHANE-1-d_2 (NEd_2) WITH ACETATE ION IN DMSO- AND ACETONITRILE-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol.% MeCN in solvent	[NE], [NE-d_2] (mol dm^{-3})	[OAc$^-$] (mol dm^{-3})	k^H_{obs} [a] (s^{-1})	k^D_{obs} [a] (s^{-1})	k^H_{OAc}/k^D_{OAc}
0		0.146	0.100	3.36×10^{-6}	5.70×10^{-6}	5.9
30		0.146	0.100	1.03×10^{-6}	1.64×10^{-6}	6.3
50		0.146	0.090	3.78×10^{-5}	5.07×10^{-6}	7.5
70		0.146	0.060	1.86×10^{-4}	2.87×10^{-6}	6.50
	0	0.146	0.100	3.36×10^{-6}	5.70×10^{-6}	5.9
	30	0.146	0.100	5.61×10^{-6}	7.51×10^{-7}	7.5
	50	0.146	0.080	5.87×10^{-6}	6.66×10^{-7}	8.8
	65	0.146	0.060	4.73×10^{-6}	5.70×10^{-7}	8.3
	70	0.146	0.030	2.88×10^{-6}	3.74×10^{-7}	7.7

[a] Defined by $k_{obs} = k_{OAc}[OAc^-]$ and obtained from equation 2.14 using values of ε* given in Table 2.14. The quoted values of k_{obs} represent the mean value of at least two experimental determinations and have been corrected for any "buffer reaction".

TABLE 2.26 - REACTION OF NITROETHANE (NE) AND NITROETHANE-1-d_2 (NE d_2) WITH ACETATE ION IN TRIFLUOROETHANOL-H_2O MIXTURES AT 25°C

Vol.% TFE in solvent	[NE],[NE-d_2] (mol dm^{-3})	$10^2[OAc^-]$ [a] (mol dm^{-3})	$10^6 k^H_{obs}$ [b] (s^{-1})	$10^7 k^D_{obs}$ [b] (s^{-1})	$10^6 k^H_{OAc}$ [c] ($dm^3 mol^{-1} s^{-1}$)	$10^6 k^D_{OAc}$ [c] ($dm^3 mol^{-1} s^{-1}$)	k^H_{OAc}/k^D_{OAc}
0							5.9
25	0.146	2.00	1.44	1.13	28.90	3.08	9.3
	0.146	4.00	2.06	1.42			
	0.146	6.00	2.59	1.65			
	0.146	8.00	3.23	1.83			
50	0.146	2.00	1.26	2.06	16.80	1.77	9.5
	0.146	4.00	1.62	2.43			
	0.146	6.00	1.94	2.74			
	0.146	8.00	2.27	3.17			
65	0.146	1.00	1.01	1.27	14.90	1.84	7.9
	0.146	2.00	1.17	1.52			
	0.146	4.00	1.44	1.93			
	0.146	6.00	1.78	2.27			
80	0.146	1.00	0.796	0.90	8.57	1.33	6.4
	0.146	2.00	0.879	1.08			
	0.146	3.00	0.962	1.22			
	0.146	4.00	1.063	1.33			

Footnotes continued:

TABLE 2.26 (contd.)

a [OAc^-]/[HOAc] = r

b Calculated from equation 2.14; values quoted have been corrected for the buffer reaction

c From a plot of k_{obs} against [B^-]

TABLE 2.27 - EQUILIBRIA AND KINETIC HYDROGEN ISOTOPE EFFECTS FOR PROTON TRANSFERS FROM NITROETHANE TO ACETATE ION IN DMSO-, ACETONITRILE- AND TFE-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	Vol.% MeCN[a] in solvent	Vol.% TFE[b] in solvent	$pK_{NE}-pK_{HOAc}$[c]	k^H_{OAc}/k^D_{OAc}
0			4.1	5.9
30			-	6.3
50			4.2	7.5
70			4.5	6.5
	0		4.1	5.9
	30		-	7.5
	50		4.44	8.8
	65		-	8.3
	80		-	7.7
		0	4.1	5.9
		25	-	9.3
		50	4.4	9.5
		65	-	7.9
		80	4.4	6.4

a MeCN = acetonitrile

b TFE = trifluoroethanol

c From Tables 1.7, 1.11 and 1.18

TABLE 2.28 - REACTION OF ETHYL-2-METHYLACETOACETATE (EMAA) AND ETHYL-2-METHYLACETOACETATE-2-d_1 (EMAA $2d_1$) WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	10^4[EMAA]^a (mol dm^{-3})	10^4[EMAA $2d_1$]^{a,b} (mol dm^{-3})	10^2[OAc$^-$] (mol dm^{-3})	k^H_{obs} ^{c} (s^{-1})	k^D_{obs} ^{c} (s^{-1})	k^H_{OAc} ($dm^3mol^{-1}s^{-1}$)	k^D_{OAc} ($dm^3mol^{-1}s^{-1}$)	k^H_{OAc}/k^D_{OAc}
0	9.28	9.20	9.80	$3.84\text{x}10^{-4}$	$6.35\text{x}10^{-5}$			6.0
30	9.28	9.20	2.00	$2.76\text{x}10^{-4}$	$4.60\text{x}10^{-5}$	$1.31\text{x}10^{-2}$	$2.25\text{x}10^{-3}$	5.8
			3.90	$5.28\text{x}10^{-4}$	$8.68\text{x}10^{-5}$			
			5.90	$7.77\text{x}10^{-4}$	$13.9\text{x}10^{-5}$			
			7.90	$10.38\text{x}10^{-4}$	$18.2\text{x}10^{-5}$			
			9.80	$13.14\text{x}10^{-4}$	$22.0\text{x}10^{-5}$			
50	9.28	9.20	1.97	$0.826\text{x}10^{-3}$	$1.27\text{x}10^{-4}$	$4.18\text{x}10^{-2}$	$6.65\text{x}10^{-3}$	6.3
			3.95	$1.73\text{x}10^{-3}$	$2.65\text{x}10^{-4}$			
			5.93	$2.36\text{x}10^{-3}$	$3.54\text{x}10^{-4}$			
			7.90	$3.30\text{x}10^{-3}$	$5.16\text{x}10^{-4}$			
			8.90	$3.73\text{x}10^{-3}$	$5.49\text{x}10^{-4}$			
60	9.28	9.20	1.97	$1.97\text{x}10^{-3}$	$3.15\text{x}10^{-4}$	$7.38\text{x}10^{-2}$	$1.40\text{x}10^{-2}$	5.7
			3.95	$3.41\text{x}10^{-3}$	$5.89\text{x}10^{-4}$			
			5.93	$4.95\text{x}10^{-3}$	$8.49\text{x}10^{-4}$			
			6.90	$5.52\text{x}10^{-3}$	$10.25\text{x}10^{-4}$			
70	9.28	9.47	1.00	$2.43\text{x}10^{-3}$	$3.81\text{x}10^{-4}$	$2.23\text{x}10^{-1}$	$4.06\text{x}10^{-2}$	5.5
			2.00	$4.71\text{x}10^{-3}$	$8.25\text{x}10^{-4}$			
			3.00	$6.88\text{x}10^{-3}$	$12.9\text{x}10^{-4}$			
			4.90	$10.9\text{x}10^{-3}$	$20.1\text{x}10^{-4}$			

Footnotes continued:

TABLE 2.28 (contd.)

a These have been corrected for the substrate consumed during reaction
b This is uncorrected for the presence of 1% H material
c See footnotes to Table 2.23

TABLE 2.29 - REACTION OF ETHYL-2-METHYLACETOACETATE (EMAA) AND ETHYL-2-METHYLACETOACETATE (EMAA $2d_1$) WITH FLUORIDE ION IN DMSO-H_2O MIXTURES AT 25°C [a]

Vol.% DMSO in solvent	10^4[EMAA] (mol dm^{-3})	10^4[EMAA-$2d_1$] (mol dm^{-3})	10^2[F^-] (mol dm^{-3})	k^H_{obs} (s^{-1})	k^D_{obs} (s^{-1})	k^H_F (dm^{3}mol^{-1}s^{-1})	k^D_F (dm^{3}mol^{-1}s^{-1})	k^H_F/k^D_F
0	9.41	9.69	2.00	2.28×10^{-5}	6.18×10^{-5}	3.52×10^{-4}	1.17×10^{-4}	3.0
			3.90	3.90×10^{-5}	8.68×10^{-5}			
			5.90	3.61×10^{-5}	11.1×10^{-5}			
			7.90	4.45×10^{-5}	13.4×10^{-5}			
			9.80	4.91×10^{-5}	15.4×10^{-5}			
50	9.41	9.69	2.00	2.15×10^{-4}	6.71×10^{-5}	9.26×10^{-3}	3.42×10^{-3}	2.7
			3.90	3.55×10^{-4}	12.8×10^{-5}			
			5.90	5.50×10^{-4}	20.8×10^{-5}			
			7.90	7.34×10^{-4}	26.7×10^{-5}			
			9.80	8.68×10^{-4}	33.6×10^{-5}			
70			2.00	2.50×10^{-3}	1.06×10^{-3}	1.39×10^{-1}	5.59×10^{-2}	2.5
			3.90	4.81×10^{-3}	2.17×10^{-3}			
			5.90	8.20×10^{-3}	3.65×10^{-3}			
			9.80	13.0×10^{-3}	5.36×10^{-3}			

[a] See footnotes to Table 2.28

TABLE 2.30 - EQUILIBRIA AND KINETIC HYDROGEN ISOTOPE EFFECTS FOR PROTON TRANSFERS FROM EMAA[a] TO ACETATE AND FLUORIDE IONS IN DMSO-H_2O MIXTURES AT 25°C

Vol.% DMSO in solvent	$pK_{EMAA}-pK_{HOAc}$ [b]	k^H_{OAc}/k^D_{OAc}	$pK_{EMAA}-pK_{HF}$ [b]	k^H_F/k^D_F
0	7.7	6.0	9.4	3.0
30	-	5.8	-	-
50	6.5	6.3	7.5	2.7
60	-	5.7	-	-
70	5.6	-	6.0	2.5

a EMAA = ethyl-2-methylacetoacetate

b From Table 1.18

TABLE 2.31 - REACTION OF ACETONE AND ACETONE-d_6 WITH ACETATE ION IN DMSO-H_2O MIXTURES AT 25°C

[Acetone] = 0.294M,
[Acetone d_6] = 0.293M

Vol.% DMSO in solvent	[OAc^-] =4[HOAc] (mol dm^{-3})	[OAc^-] =.67[HOAc] (mol dm^{-3})	k^H_{obs} [a] (s^{-1})	k^D_{obs} [a] (s^{-1})	k^H_{OAc} [b] ($dm^3 mol^{-1} s^{-1}$)	k^D_{OAc} [b] ($dm^3 mol^{-1} s^{-1}$)	k^H_{OAc}/k^D_{OAc}
0	0.10		2.67×10^{-7} [c]	3.07×10^{-8} [c]	2.50×10^{-7}	7.97×10^{-8}	3.1 [d]
		0.099	3.52×10^{-7} [c]	8.55×10^{-8} [c]			
50	0.020		1.47×10^{-8}	0.47×10^{-8}	6.92×10^{-7}	3.02×10^{-7}	2.3
	0.040		2.91×10^{-8}	1.05×10^{-8}			
	0.059		4.30×10^{-8}	1.68×10^{-8}			
	0.079		5.39×10^{-8}	2.40×10^{-8}			
	0.098		7.48×10^{-8}	2.94×10^{-8}			
		0.030	2.74×10^{-8}	0.80×10^{-8}			
		0.040	3.67×10^{-8}	1.23×10^{-8}			
		0.050	4.31×10^{-8}	1.38×10^{-8}			
70	0.01		1.75×10^{-8}	0.72×10^{-8}	1.93×10^{-6}	8.46×10^{-7}	2.3
	0.02		3.86×10^{-8}	1.66×10^{-8}			
	0.03		5.16×10^{-8}	2.51×10^{-8}			
	0.04		7.89×10^{-8}	3.35×10^{-8}			
	0.05		10.41×10^{-8}	4.42×10^{-8}			
		0.02	4.50×10^{-8}	1.66×10^{-8}			
		0.01	1.99×10^{-8}	0.72×10^{-8}			

Table continued:

TABLE 2.31 (contd.)

[Acetone] = 0.294M,
[Acetone d_6] = 0.293M

Vol.% DMSO in solvent	$[OAc^-] = 4[HOAc]$ (mol dm^{-3})	$[OAc^-] = .67[HOAc]$ (mol dm^{-3})	k^H_{obs} [a] (s^{-1})	k^D_{obs} [a] (s^{-1})	k^H_{OAc} [b] ($dm^3 mol^{-1} s^{-1}$)	k^D_{OAc} [b] ($dm^3 mol^{-1} s^{-1}$)	k^H_{OAc}/k^D_{OAc}
90	0.02		1.88×10^{-7}	0.86×10^{-7}	9.70×10^{-6}	3.70×10^{-6}	2.6
	0.04		3.38×10^{-7}	1.25×10^{-7}			
	0.06		5.84×10^{-7}	2.34×10^{-7}			
	0.08		7.99×10^{-7}	2.73×10^{-7}			

a $k_{obs} = \left\{ k_o' + [OAc^-](k_{OAc} + \frac{k_{HOAc}}{r}) \right\}$ where $r = [OAc^-]/[HOAc]$

b Obtained from a plot of k_{obs} against $[OAc^-]$

c These have been corrected for $k[OAc^-][HOAc]$ term (see text); contributions due to k_o' have been ignored

d This is somewhat lower than the value of 4.8 reported by Reitz and Kopp[96]

5. DISCUSSION

The results presented (see Tables 2.24, 2.26, 2.29 and 2.30) show that no correlations are observed in the mixed solvent systems between k^H/k^D and ΔpK; in some cases large variations in isotope effects with little change in ΔpK are observed, whilst in others the isotope effects remain constant with varying ΔpK. This is most strikingly illustrated using the results presented in Table 2.26 which show the variation in k^H/k^D for proton transfer from nitroethane to acetate ion in various mixed solvent systems. In all the differing solvent mixtures the isotope effects go through a distinct maximum as the concentration of the organic component of the solvent is increased; however, the ΔpK in all cases is almost constant at 4.4 pK units.

Such results clearly show that the variation of isotope effects with solvent cannot be related solely to the effect of solvent on ΔG as was first supposed. The most obvious reason for this is that the solvent molecules themselves interact with the polar transition state, the extent of this interaction depending upon the nature of the acid and the base. Such effects would give rise to a solvent dependency of transition state zero-point energies and, hence, of the isotope effects also. In addition, however, it is also possible that the solvent could affect the magnitude of any tunnel effect present in the reaction. This is because, if the solvent interaction becomes so strong that the proton transfer through the energy barrier is accompanied by the simultaneous translation or rotation of one or more solvent molecules, then the "effective mass" of the proton during transfer will be increased and the tunnel correction correspondingly reduced. The higher this mass, the smaller the resulting tunnel correction and the lower, therefore, the isotope effect.* Thus, the isotope effect for many proton-transfer reactions in differing solvents will depend not only upon ΔG^o for reaction but also upon the effects of solvent on transition state zero-point energies and upon any tunnel correction which may have to be applied.

Such effects as the above are expected to be of importance even in those cases where variations in ΔpK are brought about by altering substrates and/or bases but keeping the solvent constant. It is interesting to note, for example, that the large tunnel corrections frequently found for sterically hindered systems in aqueous solution have been attributed to a higher and steeper energy barrier in such cases: according to the present view, however, the role of the hindering groups is to exclude water molecules and thus preserve a low effective mass for the proton.[87]

The results of the present study are very difficult to interpret quantitatively in terms of the above arguments because in mixed solvents we do not know the nature of the solvent species involved with the transition state or

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* Any increase in effective mass would also affect the zero-point energy of the transition state and would, in fact, lead to an increase in the isotope effect. However, such effects are expected to be small compared with the solvent dependency of the tunnel corrections.[87]

proton during transfer. Clearly what are required are data for proton transfers in differing pure solvents. The few results already available are consistent with the factors outlined so far. Thus, Caldin and Mateo[97] have studied the reaction of 4-nitrophenylmethane with tetramethylguanidine in several aprotic solvents. They found that k^H/k^D at 25°C varied between 11 and 50 although the actual rates bore no relation to the rather small variations in ΔG (ca. 3 kcal mol^{-1}). The highest isotope effects were found in the more non-polar solvents and these were explained in terms of increased tunnel corrections in such solvents. This would be expected on the basis of dipolar solvents interacting more strongly than non-polar solvents with the proton during transfer and hence causing a greater value of the effective mass.

In conclusion then, it can be seen from the results of this second section of the thesis that the rates, equilibria and kinetic hydrogen isotope effects of proton-transfer reactions are all highly dependent upon the nature of the solvent. Care must, therefore, be exercised in correlating Brönsted coefficients and isotope effects with transition state structure.[77,98]

APPENDIX I

THE FLUORIDE ION SELECTIVE ELECTRODE IN MIXED SOLVENT SYSTEMS

INTRODUCTION AND THEORY

The fluoride electrode, first discovered by Frant and Ross (1966), has proved to be a highly successful ion selective electrode in aqueous solution and more than a hundred papers dealing with its use and properties have so far been published. (Very little work has used non-aqueous or mixed solvent systems.) The electrode senses activity, rather than concentration, of the fluoride ion in solution and, when coupled with a reference electrode, responds according to the equation:

$$E_{obs} = E^{o} - \text{const.}\ln a_{F^-}$$

where E_{obs} is the observed cell EMF, E^{o} is a constant determined by the internal filling solution ($E^{o}_{F^-}$) and the reference electrode, and a_{F^-} is the activity of the fluoride ion in solution.

The value of the constant at any temperature nearly always equals the constant $\frac{RT}{F}$ appearing in the more common Nernst equation, and consequently the electrode is almost invariably described in terms of the Nernst-type equation:

$$E_{obs} = E^{o} - \frac{RT}{F} \ln a_{F^-}$$

This equation holds for fluoride concentrations ranging from saturated solution to about 10^{-5} molar, although interference can occur if hydroxide ion concentration becomes too high. However, at the hydroxide ion concentrations used in the present work this problem does not arise.

Use of the fluoride ion electrode to determine the free energies of transfer of silver fluoride from water to mixed solvent systems requires:

(a) evidence that the electrode still gives a Nernst-like response to a_{F^-} in the new solvent (i.e. a change in E_{obs} of $\frac{RT}{F}$ volts for a ten fold change in a_{F^-})

(b) that the constant $E^{o}_{F^-}$ is, to within experimental accuracy, independent of the solvent used.

The former can be established by observing the variation with fluoride ion activity of the EMF for the cell $LaF_3, F^-||$calomel. If point (a) is valid then a plot of EMF against a_{F^-} ($a_{F^-} = [F^-]\gamma_{\pm}$) should give a straight line of gradient close to 59 mV at 25^{o}C. This has been found to be true for ethanol-water mixtures (Lingane, 1968) and also isopropanol-water mixtures (Bazzelle, 1971).

However, the second point (which is now considered in detail) is more difficult to establish since it necessarily requires the use of an electrode which responds specifically to fluoride ion activity and whose chemistry is clearly understood. In this study use was made of the lead amalgam-lead fluoride electrode, first introduced by Jahn-Held and Jellinek in 1936 and used extensively by De Vries during the 1940's. The half cell can be represented as Pb(Hg), PbF_2(solid)$|F^-$

and the cell process by:

$$Hg + Pb^{2+} + 2e \rightleftharpoons Pb(Hg)$$

The corresponding Nernst equation is:

$$E_{Pb(Hg)} = E'_{Pb(Hg)} - \frac{RT}{2F} \ln \frac{1}{a_{Pb^{2+}}}$$

where E' includes the standard reduction potential of Pb^{2+} and a term involving the activity of the amalgam (this, of course, being constant for a given amalgam).

Remembering that the equilibrium constant, K_{sp}, for $PbF_2 \rightleftharpoons Pb^{2+} + 2F^-$ must be maintained then:

$$a_{Pb^{2+}} = K_{sp}\ (a_{F^-})^{-2}$$

and so:

$$E_{Pb(Hg)} = E^{o}_{Pb(Hg)} - \frac{RT}{F} \ln a_{F^-}$$

(where $E^{o} = E' + \frac{RT}{2F} \ln K_{sp}$).

Thus, the electrode responds directly to a_{F^-}. Using water as solvent, we may write:

$$E_{Pb(Hg)aq} = E^{o}_{Pb(Hg)} - \frac{RT}{F} \ln(a_{F^-})_{aq} \qquad (i)$$

If we define the standard state of any species as the hypothetically ideal aqueous solution of unit concentration, irrespective of the solvent used, then the corresponding Nernst equation for any solvent system s becomes

$$E_{Pb(Hg)s} = E^{o}_{Pb(Hg)} - \frac{RT}{F} \ln (a_{F^-})_{s} \qquad (ii)$$

Because no junction potentials are introduced by changing solvent, the value of E^o is constant throughout and all changes in E arise from changes in fluoride ion activity.

Turning now to the fluoride ion selective electrode half cell - this can be represented as $LaF_3|F^-$ with the corresponding Nernst equation:

$$E_{F^-} = E^o_{F^-} - \frac{RT}{F} \ln a_{F^-}$$

Using water as solvent we may write:

$$(E_{F^-})_{aq} = E^o_{F^-} - \frac{RT}{F} \ln(a_{F^-})_{aq} \qquad \text{(iii)}$$

and using the standard states just defined for solvent s:

$$(E_{F^-})_s = E'_{F^-} - \frac{RT}{F} \ln(a_{F^-})_s \qquad \text{(iv)}$$

where the constant E'_{F^-} may or may not be equal to $E^o_{F^-}$ (see below).

Coupling together the half cells involving fluoride ion electrode and lead fluoride electrode produces the overall cell $Pb(Hg), PbF_2|F^-||F^-|LaF_3$. If we use the same fluoride ion concentration in both half cells*, for a given solvent, then we may write:

$$E_{aq} = \text{(iii)} - \text{(i)} = E^o_{F^-} - E^o_{Pb(Hg)} \qquad \text{(v)}$$

$$\text{and} \quad E_s = \text{(iv)} - \text{(ii)} = E'_{F^-} - E^o_{Pb(Hg)} \qquad \text{(vi)}$$

where E represents the overall cell EMF. We now have a means of establishing whether or not $E^o_{F^-} = E'_{F^-}$, because if we find that $E_{aq} = E_s$, then $E^o_{F^-}$ and E'_{F^-} must also be equal and point (b) is valid.

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* The concentrations of fluoride used should be high enough to avoid complications arising from the dissociation of PbF_2 in the amalgam electrode.

RESULTS AND DISCUSSION

The results are shown in Tables 1 and 2.

For all the solvents used, save the higher acetonitrile-water mixtures, the EMF of the $LaF_3|F^-||$calomel cell is proportional to the fluoride ion activity (Table 1) with the proportionality constant always very similar in magnitude to that expected from the Nernst equation. Thus, no significant errors are introduced by using a Nernst-type equation to describe the fluoride ion electrode half cell in our solvent mixtures.

The results in Table 2 clearly show that, to within experimental accuracy (± 2 mV), the EMF of the cell Pb(Hg), $PbF_2|F^-||F^-|LaF_3$ using equal concentrations of fluoride in both half cells, is constant and thus independent of solvent. This can only be the case if the cell constant, $E^o_{F^-}$ is independent of solvent, as pointed out in the previous section.

Thus, both assumptions concerning the fluoride ion selective electrode made in the free energy of transfer work are valid.

It should perhaps be noted at this point that Stelting and Manaha have recently investigated the behaviour of the fluoride electrode in aqueous mixtures containing up to 20 volume per cent of acetonitrile. They explain the solvent induced changes in the EMF of the cell, $LaF_3|F^-||$calomel, entirely in terms of the changes in solubility product of the LaF_3 in the solvent mixtures*. Clearly, the present study

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* This is theoretically impossible because, at the fluoride ion concentrations used, such differences in EMF are totally independent of the solubility product of LaF_3 (cf. eqns. v and vi.)

TABLE 1 - VARIATION OF THE EMF OF THE CELL $LaF_3|F^-||$CALOMEL WITH FLUORIDE ION CONCENTRATION AND SOLVENT COMPOSITION

Vol.% of each solvent compo- nent [a]	EMF (mV) of the cell $LaF_3|F^-||$calomel using $[F^-]$, in mol dm^{-3}, shown below								Gradient of the graph $(E_{obs}$ v. $a_{F^-})$ [d] (mV)
	0.10	0.08	0.02	0.01	0.005	0.002	0.001	0.0001	
H_2O									56
50/50 DMSO/H_2O [b]				156	137		100		58.5
80/20 DMSO/H_2O [b]				334			278	218	60
50/50 MeCN/H_2O [c]		195	162						58
70/30 MeCN/H_2O [c]	210		190			142			e
70/30 TFE/H_2O [c]		170	137						58

a Solvent abbreviations: DMSO(dimethyl sulphoxide), MeCN(acetonitrile), TFE(trifluoroethanol)

b Temperature used, 25°C

c Carried out at room temperature (≈ 23°C)

d a_{F^-} calculated from $a_{F^-} = [F^-]\gamma_{\pm}$, where $\gamma_{\pm}$ is obtained from Davies equation using appropriate A values and taking I as the concentration of F^- shown

e The plot is non-linear and is especially deviant at lower $[F^-]$. This can be attributed to ion pairing in the mixtures of high [MeCN].

TABLE 2 - EFFECT OF FLUORIDE ION CONCENTRATION AND SOLVENT COMPOSITION ON THE EMF OF THE CELL $Pb(Hg), PbF_2|F^-||F^-|LaF_3$ AT ROOM TEMPERATURE

Vol.% of each solvent component [a]	EMF (mV) of the cell $Pb(Hg)\,|\,PbF_2|F^-||F^-|LaF_3$ at Room Temperature, using $[F^-]$ ($mol\ dm^{-3}$) shown below						
	0.15	0.10	0.08	0.05	0.02	0.01	0.001
H_2O		421				429[b]	459[b]
50/50 $DMSO/H_2O$	425			425		423	441[c]
80/20 $DMSO/H_2O$		421		426			
50/50 $MeCN/H_2O$			421		420		
70/30 $MeCN/H_2O$[d]		421			420		
70/30 TFE/H_2O			418		419		

a Solvent abbreviations as in Table 1.

b These values are probably anomalous, due to dissociation of PbF_2, in the amalgam half cell, at low $[F^-]$. (K_{sp} in water is 7.8×10^{-8}).

c This is probably anomalous for the same reason as b. K_{sp} is likely to decrease with increasing DMSO content of the solvent and so deviations are now expected only at the lowest $[F^-]$, as observed.

d Any ion pairing has no overall effect on the cell EMF, because fluoride ion activity is still the same in both half cells.

would suggest that the entire variations in EMF arise from variations in fluoride ion free energy; a point apparently not considered by the authors. They also fail to consider changes in junction potential between calomel electrode and the fluoride solution - a situation which is deliberately eliminated in our present study.

EXPERIMENTAL

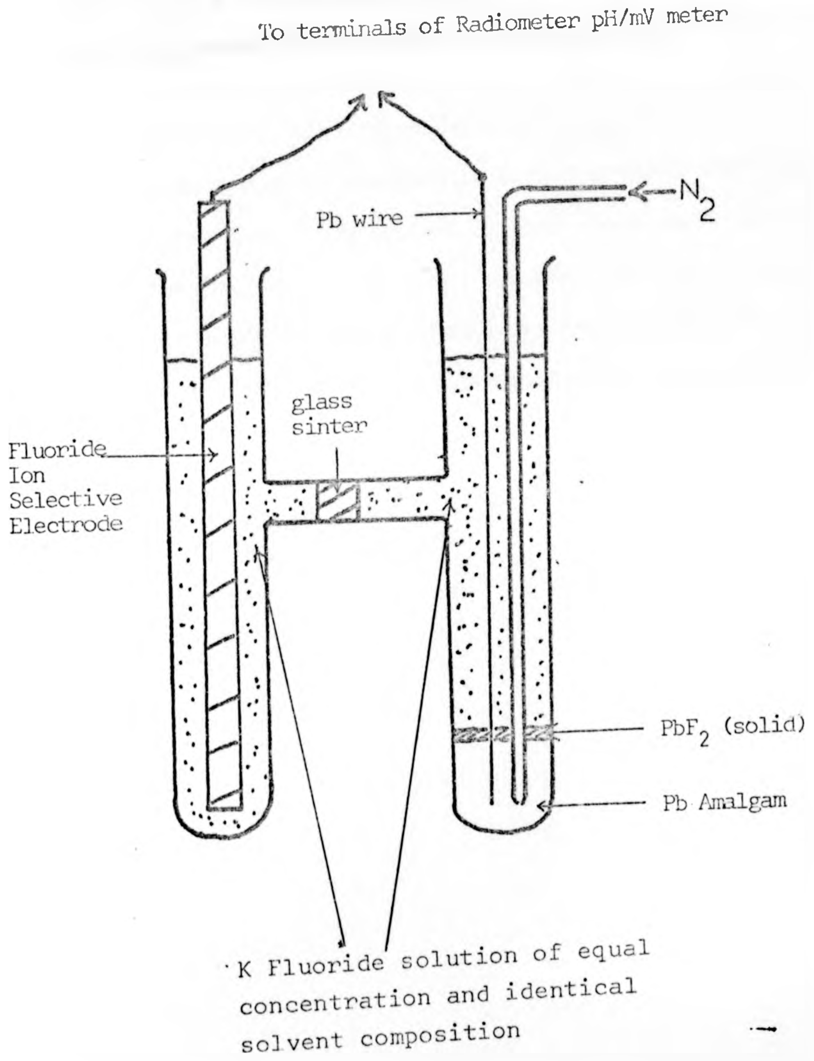
The lead amalgam was 5% by weight of lead, and was prepared by the method described by Vogel. Lead fluoride was prepared by treating a solution of lead acetate (BDH Laboratory Reagent) with a slight excess of analytical grade sodium fluoride, and filtering off the fine precipitate of lead fluoride. The latter was washed with water and dried overnight at 200°C.

Fluoride solutions were prepared from BDH potassium fluoride (Laboratory Reagent) purified by procedures already described in this thesis (as were the DMSO, acetonitrile and water used).

Fisons Ltd. acetonitrile (spectroscopic grade) was used without further purification.

The amalgam cell was made up as shown in the accompanying figure. Nitrogen was bubbled through the amalgam for thirty minutes before any readings were taken - this prevented formation of lead oxide during cell equilibration. The steady values of cell EMF, at room temperature, were recorded - values were reproducible to within two millivolts.

ILLUSTRATION OF THE Pb(Hg), $PbF_2|F^-\|F^-|LaF_3$ CELL



The $LaF_3|F^-||$calomel cells were prepared as described in the section on the solvent dependence of the pK_a of hydrofluoric acid. Cell EMF's were almost always reproducible to within one millivolt.

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APPENDIX II

DMSO-H_2O MIXTURES IN TERMS OF VOLUME % AND MOLE % OF EACH COMPONENT AT 25°C

Volume % DMSO	Mole % DMSO*
10	2.7
20	5.9
40	14.4
50	20.2
60	27.5
70	38.5
80	50.2
85	58.9
90	69.4
95	83.0

$$* \text{ Mole } \% = 10^2 \frac{(\text{Vol.\% x Density})_{DMSO}(\text{Mol.Wt.DMSO})^{-1}}{(\text{Vol.\% x Density})_{DMSO}(\text{Mol.Wt.DMSO})^{-1} + (\text{Vol.})_{H_2O}(\text{Mol.Wt.}H_2O)^{-1}}$$

$$= 10^2 \frac{(1.096 \text{ Vol.\%})_{DMSO}(78.13)^{-1}}{(1.096 \text{ Vol.\%})_{DMSO}(78.13)^{-1} + (\text{Vol.\%})_{H_2O}(18)^{-1}}$$

APPENDIX III - KINETIC EXPRESSIONS FOR THE EQUILIBRIUM $SH + \sum_i Bi \rightleftharpoons S^- + \sum_i Ai$

Using the same symbolism as given in Section II of this thesis then:

$$-\frac{d[SH]}{dt} = \frac{d[S^-]}{dt} = \sum_i k_{Bi}[Bi][SH] - \sum_i k_{Ai}[Ai][S^-]$$

and defining the following variables:

$a = [SH]$ = concn. of SH at time t

$b = [S^-]$ = concn. of S^- at time t

$a_e = [SH]_e$ = concn. of SH at equilibrium

$b_e = [S^-]_e$ = concn. of S^- at equilibrium

enables us to write, from the stoichiometry of the reaction, $x = a - a_e = b_e - b$ since $a + b = a_e + b_e$.

The above kinetic expression can now be written:

$$\frac{d[S^-]}{dt} = \frac{db}{dt} = -\frac{dx}{dt} = \sum_i k_{Bi}[Bi][SH] - \sum_i k_{Ai}[Ai][S^-]$$

$$= (\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])x + (\sum_i k_{Bi}[Bi])a_e$$

$$- (\sum_i k_{Ai}[Ai])b_e$$

But at equilibrium the rate of forward and reverse reactions are equal and so

$$(\sum_i k_{Ai}[Ai])b_e = (\sum_i k_{Bi}[Bi])a_e \ .$$

Therefore:

$$-\frac{dx}{dt} = (\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])x$$

Integration of this equation between the limits $x = x$, $t = t$ and $x = a_o$, $t = 0$ gives us:

$$\ln x = -(\sum_i k_{Bi}[Bi] + \sum_i k_{Ai}[Ai])t + \text{const}$$

and because $\ln([S^-]_e - [S^-]) = \ln x$ then the above expression is the same as equation 2.11.

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