Base Strength and Solute–Solvent Interaction of Metal Extractants: Hydroxamic Acids

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The protonation behaviour of versatile metal extractants, N-aryl-substituted hydroxamic acids, has been studied in aqueous perchloric acid. pK_{BH^+} values have been calculated according to linear free energy relationship and excess acidity methods. The partition data of these reagents between various organic solvent–aqueous systems obtained have been correlated with regular solution theory. Primary medium activity coefficients, free energies of transfer of a solute in solvent and $1000\rho/M$ values are determined. The solvent polarity effect is also discussed.

Hydroxamic acids have been used extensively as metal extractants.[1-5] These reagents can exist in two tautomeric forms (Scheme 1). Spectroscopic evidence indicates that the keto form predominates.[6] These metal extractants are isolated as naturally occurring compounds[7] and are the products of some photochemical reactions.[8] Besides their wide-ranging applications in various fields, viz. analytical,[9,10] agricultural,[11] nuclear,[12] in polymer chemistry,[13] in flotation techniques,[14] as corrosion inhibitors,[15] in therapeutic applications having antimalarial and antibacterial activity,[16,17] as mental tonic and anti-inflammatory agents,[18] as inhibitors of stone formation in urinary tract and in DNA biosynthesis,[19] their physicochemical properties have received little attention.[20-25]

Many papers have appeared in the literature describing the colour reaction performed by these reagents when complexed with various metal ions, but the effect of solvent on distribution constants has not been treated theoretically, although it constitutes one of the fundamental features of solvent extraction. Thus an attempt is made to correlate the distribution data with theoretical calculations based on regular solution theory. The effect of solvent properties such as dielectric constant (ε) and solvent polarity (π^*) on experimental distribution constant is also discussed. Values of solubility parameter (δ_C), free energy of transfer (ΔG_t°), activity coefficient (γ_C), density (ρ), molar volume (V_1) and $1000\rho/M$ are determined.

These metal extractants serve as weak organic bases in the presence of strong acidic solutions; their dissociation constants, pK_{BH^+}, have been calculated in perchloric acid following the linear free energy approach and excess acidity method. The percentage of protonated hydroxamic acid and their distribution ratios as a function of perchloric acid are also discussed.

Experimental

Reagents and instrumentation. Hydroxamic acids were prepared according to a reported method[26] and purified by crystallisation from benzene. A saturated solution of ammonium metavanadate was prepared in distilled water. The chloroform used was shaken six times with equal volumes of water to remove the ethanol (present as a preservative) and distilled. It was stored in an amber bottle in a cool place. Concentrated perchloric acid was standardised against standard sodium hydroxide. Acid solutions of appropriate concentrations were made up by diluting concentrated perchloric acid with distilled water to give the required acidity. All other chemicals used were of analytical grade.

A Spekol model EK-1 instrument (Carl Zeiss, Jena)
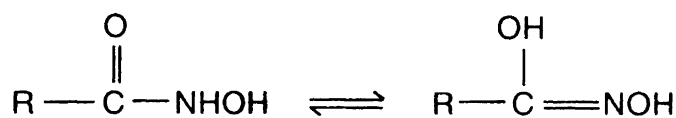
Scheme 1.

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was used for the measurement of absorbance, and a DMS 100S UV–VIS Varian spectrophotometer was used for recording ultraviolet spectra using 10-mm matched silica cells. All the calculations were carried out on a Pentium 166 MMX computer.

Measurement of K_D *and D.* For the measurement of the distribution constant, K_D, an aliquot or organic solvent solution containing hydroxamic acid and an equal volume of aqueous phase were shaken for 30 min at $30 \pm 1\,^\circ\text{C}$ in an extraction tube. After centrifuging, the concentrations of both phases were determined spectrophotometrically.

To measure the distribution ratios, D, 10 ml of organic solvent containing the hydroxamic acid were equilibrated with 10–25 ml of the 0–10 M perchloric acid for 5 min following a rapid technique.[20]

The reagent concentrations were determined colorimetrically using the vanadium (V) method.[27] This method has the advantage that side-reactions do not necessarily invalidate the measurements.

Density measurements. A hydrostatic weighing method was used for density measurements. All the weighings were corrected for air buoyancy, and the densities were calculated by the formula

$$D_{HA} = \frac{W_{air}}{W_{air} - W_{water}} (D_{water} - D_{air}) + D_{air}$$

where D_{HA} is the density of hydroxamic acid, D_{water} is the density of water, D_{air} is the density of air, W_{air} is the weight of the sample in air, and W_{water} is the weight of the sample in water. The density values are useful for the determination of molar volume and $1000\rho/M$ where M is the molecular weight.

All the data reported here are averages of at least six separate determinations.

Results and discussion

N-Arylhydroxamic acids act as weak organic bases in the presence of strong acidic solutions. The dissociation constants, pK_{BH^+}, of very weak bases are usually determined spectrophotometrically by measuring, in concentrated aqueous acid solution, the ionisation ratios $C_{BH^+}/C_B = I$ for the equation

$$BH^+ \underset{}{\overset{pK_{BH^+}}{\rightleftharpoons}} B + H^+ \qquad (1)$$

where C_B and C_{BH^+} are the molar concentration of the base and conjugate acid, respectively. In the present investigation, values of the ionisation ratio measured at various perchloric acid concentrations are used to determine pK_{BH^+} values of these metal extractants. Perchloric acid is a monobasic acid and is much more completely dissociated than sulfuric acid in concentrated solutions.

pK_{BH^+} are obtained from an analysis of the variation of some physical property of the substrate with changing acid concentration. The change observed with increasing acidity is from a spectrum characteristic of the free base to that of the protonated base. Protonation equilibria cannot be studied by the UV method because of insufficient differences between the spectra of the substrate and the conjugate acid. The same observation was reported by Buglass et al.[28] Therefore I is estimated by the equation that describes the variation of distribution ratio, D, with changing acidity $I = (K_D - D)/D$. For distribution of the hydroxamic acid between a non-polar, inert, organic solvent and an aqueous acid solvent where the hydroxamic acid is essentially unprotonated, the thermodynamic distribution constant is $K_D = [B]_0/[B]_{aq}$. The observed distribution ratio for the hydroxamic acid between the same organic solvent and an aqueous acid layer in which protonation occurs is $D = [B]_0/[B]_{aq} + [BH^+]$. Combining these equations for protolysis equilibrium gives

$$D = (K_D - Dh_0)/K_a \qquad (2)$$

where $H_0 = -\log h_0$ and $pK_{BH^+} = -\log K_a$. Because of small activity coefficient effects, the thermodynamic distribution constant, K_D, is best estimated by eqn. (2) by a linear plot of D vs. Dh_0 or by calculating from least-squares methods rather than from the distribution constant for water and organic solvent. D is the distribution ratio, measured between the organic solvent and perchloric acid solutions. As an organic solvent carbon tetrachloride was chosen because of its favourable physical properties. The data for K_D and D are presented in Tables 1 and 2, respectively.

Protonation data have been analysed by two different methods.

Bunnett and Olsen[29] have formulated a free energy relationship [eqn. (3)]

$$\log I + H_0 = \phi(H_0 + \log C_{H^+}) + pK_{BH^+} \qquad (3)$$

for the determination of pK_{BH^+}, and this relates the Hammett acidity function H_0 to the ionisation ratio, I.

Table 1. pK_{BH^+} of N-aryl-substituted hydroxamic acids by BOM and EAM methods at $30 \pm 1\,^\circ\text{C}$.

Hydroxamic acid	K_D	BOM				EAM			
		pK_{BH^+}	ϕ	r	σ	pK_{BH^+}	m^*	r	σ
N-o-Tolylbenzo-	18.60	-0.84 ± 0.025	0.66 ± 0.015	0.999	0.05	-1.17 ± 0.006	0.45 ± 0.016	0.997	0.10
N-o-Tolyl-4-methylbenzo-	81.84	-1.12 ± 0.090	0.61 ± 0.036	0.985	0.09	-1.47 ± 0.005	0.57 ± 0.041	0.987	0.10
N-p-Tolyl-4-bromobenzo-	73.20	-0.96 ± 0.005	0.59 ± 0.031	0.993	0.10	-1.32 ± 0.035	0.53 ± 0.030	0.993	0.09
N-p-Tolyl-3,5-dinitrobenzo-	21.94	-0.74 ± 0.025	0.55 ± 0.029	0.993	0.10	-1.10 ± 0.011	0.56 ± 0.028	0.994	0.09

Table 2. Distribution ratio and percentage protonation of N-aryl-substituted hydroxamic acids as a function of perchloric acid concentration.

[Perchloric acid]/M	N-o-Tolylbenzo-		N-o-Tolyl-4-methylbenzo-		N-p-Tolyl-4-bromobenzo-		N-p-Tolyl-3,5-dinitrobenzo-	
	D	% Protonation	D	% Protonation	D	% Protonation	D	% Protonation
1	42.03	13.17	118.00	7.38	90.83	10.32	30.20	16.03
2	22.00	30.87	98.25	18.99	69.23	25.31	24.92	35.97
3	14.02	51.71	76.12	35.99	62.81	44.83	17.51	57.40
4	9.99	71.05	59.29	56.30	45.51	65.06	11.00	75.53
5	7.21	85.76	41.01	75.98	31.83	82.05	6.00	88.34
6	5.10	94.44	22.01	89.91	21.75	92.80	3.96	95.53
7	3.81	98.29	17.38	96.80	12.05	97.76	2.00	98.64
8	2.40	99.65	7.78	99.33	7.12	99.53	1.01	99.72
9	1.22	99.95	4.78	99.91	2.75	99.94	0.52	99.96
10	0.40	99.99	1.20	99.99	0.87	99.99	0.15	99.99

Such plots are usually linear. The slope ϕ is a measure of the susceptibility of the equilibrium to changing acid concentrations.

Another extrapolative method is the excess acidity method, which is an earlier approach proposed by Marziano and Passerin,[30] further developed by Yates and Cox.[31] This method uses overlapping indicators but does not involve an acidity function. Thus proton concentration C_{H^+}, and concept of excess medium acidity X, are summarised as

$$\log I - \log C_{H^+} = m^*X + pK_{BH^+} \qquad (4)$$

where X represents the excess acidity and is the difference between the observed acidity and that which the system would have, if it were ideal. Values of X as a function of weight percent composition are available for the aqueous perchloric acid systems.[32] The slope parameter m^* expresses the sensitivity of the substrate to the changing acidity.

Table 1 lists the values of K_D and pK_{BH^+}, obtained from BOM and EAM methods, values of ϕ and m^* parameters, the standard deviation, σ, and the correlation coefficient, r, for the protonation of N-arylhydroxamic acids in perchloric acid. Protonation data have been analysed in 4–10 M perchloric acid solutions. Insufficient protonation at lower acidity hinders the experimental determination of protonation constants at this stage. An examination of the data in Table 1 reveals that the EAM pK_{BH^+} is higher than pK_{BH^+} values from a linear free energy relationship. The ϕ values (0.55–0.66) and the values of the slope m^* (0.45–0.57) for hydroxamic acid are all $\ll 1$, indicating that these bases are not behaving as Hammett bases. Values of similar magnitude have been observed for amides.[28] Typical protonation m^* values for the carbonyl group are around 0.6,[33] and m^* corresponds to $(1 - \phi)$. This correspondence seems less operative in perchloric acid solutions. It is observed that substitution by an electron-releasing group, as in o-tolyl-4-methylbenzohydroxamic acid, causes an increase in m^* because the substituent stabilizes the protonated base, thus increasing the interaction with

the solvent. At the same time it decreases the basicity of the conjugate acid of hydroxamic acid. Nitro-groups are strongly electron-withdrawing oxygen-containing groups, which would be expected to interact strongly and specifically with the solvent through formation of strong O–H–O-type hydrogen bonds. With increasing acidity the water content decreases and so solvation decreases, and the influence of substituent on the reaction centre should go up. It is also concluded that the conjugate acid of nitro-substituted hydroxamic acid is stronger as compared to halogen derivatives. Although the bromo derivative is also electron withdrawing it would be expected to produce a weaker and less specific solute–solvent interaction. Thus, for the weaker bases like N-aryl-substituted hydroxamic acids, BOM and EAM methods give slightly different pK_{BH^+} values. A comparison between the values obtained from BOM and EAM for these bases shows a constant difference of 0.3 pK_{BH^+} units. Furthermore, it is interesting to notice that the pK_{BH^+} values of similar compounds, the weak amides, also furnished different values by a different approach, when the ionisation ratios are determined in perchloric acid solutions.[34] Differences in pK_{BH^+} values have also been found for other weak bases.

In those cases in which disagreement between the two methods is observed, the question arises as to which method gives the more accurate estimate of pK_{BH^+}. Although the need for individual acidity functions can be overcome by the Bunnett–Olsen and Cox–Yates methods, the BOM method has been found to be less reliable when used with an acidity function other than H_0.[36] However, the practical and theoretical advantages of the EAM method should render this method most suitable in less concentrated acid solutions, but in the case of very weak bases, for which more concentrated acid solutions are necessary, the validity of the EAM method is questionable. It has been observed that, particularly in the case of weak organic bases, BOM works better than EAM. At the same time, EAM is capable of providing mechanistic information, which the other method cannot. The use of an excess acidity function

allows one to give a physical meaning to the slope parameter m^*, which measures the hydrogen bonding solvation of the protonated base.

Table 2 presents data for percentage protonation and distribution ratios, D, between an organic solvent (carbon tetrachloride) and an increasing perchloric acid concentration (1–10 M). As the EAM does not involve an acidity function, the percentage of protonated hydroxamic acid is calculated on the basis of BOM pK_{BH^+} values, following the equation

$$\% \text{ Protonated} = \left(\frac{h_0}{h_0 + K_{BH^+}} \right) \times 100 \qquad (5)$$

Going from water to increasing perchloric acid concentrations, the degree of protonation increases at different rates depending on the structure of hydroxamic acid. At the same time, as the value of perchloric acid concentration increases, D decreases. This suggests that protonated species is hydrophilic in nature.

The acid–base properties of a substance are important parameters, especially in the investigation of acid–base catalysed reactions. Evaluation of the relative basicities of weak organic bases has attracted much interest both as a source of information about electronic structures and as a means of interpreting the reactivities of acid-catalysed reactions. An understanding of the acid–base interaction enables one to suggest optimum conditions for separating similar compounds by extraction with acid.

N-Arylhydroxamic acids are versatile metal extractants. In solvent extraction systems, if we know the fundamental parameters of the solvent, the aqueous solution and the solute, we shall be able to pre-estimate the behaviour and solute and, moreover, the role of solvent and diluent. For this, the K_D values for the partition of these metal extractants were determined for a series of organic solvent–aqueous systems. The solvents employed are of small solubility in water and with various dipole moments as listed in Table 3. An attempt is made to correlate the values of K_D, for a number of organic solvents, with some physical parameter characteristics of the solvents, to study solute–solvent interaction. Such parameters include dielectric constants (ε) and solvent polarity (π^*). The solubility parameter values (δ_C), primary medium activity coefficient (γ_C) and free energy of transfer (ΔG_t°) of a solute in a solvent of these metal extractants are discussed.

Hydroxamic acids are sparingly soluble in water but highly soluble in various organic solvents. Owing to this solubility factor, for determining K_D, hydroxamic acids were taken in such a quantity that their distribution in both the phases were up to that limit, so as to produce no difficulty in analysing the phases. K_D values can be regarded as the distribution constants for the monomeric form of these reagents, as there is no significant variation in K_D involved, with total reagent concentration. At the same time, the ionisation constants of these reagents in water are small, pK_a being between 8–14, and so the

concentration of their anionic species in the aqueous phase is negligible. K_D values were determined to obtain information about the suitability of the solvents for extraction work. Further K_D values can be correlated with regular solution theory.

A regular solution is defined as one 'which has no chemical interaction between solute and solvent molecules and no change in states of association and orientation by mixing'. To such solutions, Hildebrand's theory of regular solution[37] can be applied, according to which the solubility parameter, δ, of a substance is defined as

$$\delta = (E^V/v)^{1/2} \qquad (6)$$

where $\delta E^V/v$ is the energy of vaporisation per unit volume.

The heat of mixing of solute and solvent depends on the differences of their δ values. The solubility parameters of these metal extractants are calculated following the equation

$$\ln \frac{\phi_{1,\text{org}}}{\phi_{2,\text{aq}}} = \frac{V_1}{RT} [(\delta_C - \delta_{\text{aq}})^2 - (\delta_C - \delta_{\text{org}})^2]$$
$$+ V_1 \left(\frac{1}{V_{\text{org}}} - \frac{1}{V_{\text{aq}}} \right) \qquad (7)$$

where δ_C is the solubility parameter of solute, ϕ_1 is the volume fraction of the solute, ϕ_2 is the volume fraction of the solvent in the solution, V_1 is the molar volume of the solute, δ_{aq} is the solubility parameter of aqueous phase, V_{org} is the volume of organic solvent and V_{aq} is the volume of aqueous phase for all organic solvents.

At low concentration of the solute, $\phi_{\text{aq}} = \phi_{\text{org}} = 1$, and for the low concentration of solute used in the present investigation, it can be assumed that

$$\frac{\phi_{1,\text{org}}}{\phi_{1,\text{aq}}} = K_D \qquad (8)$$

where K_D is the experimental distribution constant between organic and aqueous phases. K_D can thus be related directly to the right-hand side terms of eqn. (7).

The values of K_D of N-aryl-substituted hydroxamic acids seem to depend on solvent polarity. A log–log plot of K_D values against the dielectric constants of the organic solvents gives a parabolic relationship as shown in Fig. 1. The Kamlet–Taft[38] treatment is based on the method of solvatochromic comparison. They developed a β scale of solvent hydrogen-bond acceptor (HBA) basicities and an α scale of solvent hydrogen bond donor (HBD) acidities. They also developed a π^* scale of solvent polarity for solvent–solute interaction. The π^* scale of solvent polarity is so called because it is based on the effect of solvent on the electronic spectral transitions of various nitroaromatic compounds. A plot of $\log K_D$ against π^* values of the solvent yields a straight line as observed in Fig. 2.

Distribution of a solute between two immiscible solvents has also been used to determine the free energy of
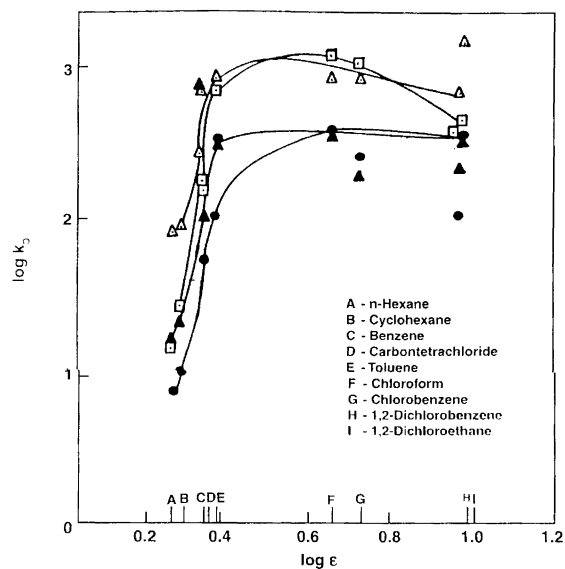
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Fig. 1. log–log plot of K_D vs. ε for: (●) N-o-tolybenzo-, (□) N-o-tolyl-4-methylbenzo-, (▲) N-p-tolyl-4-bromobenzo- and (△) N-p-tolyl-3,5-dinitrobenzo-hydroxamic acids.

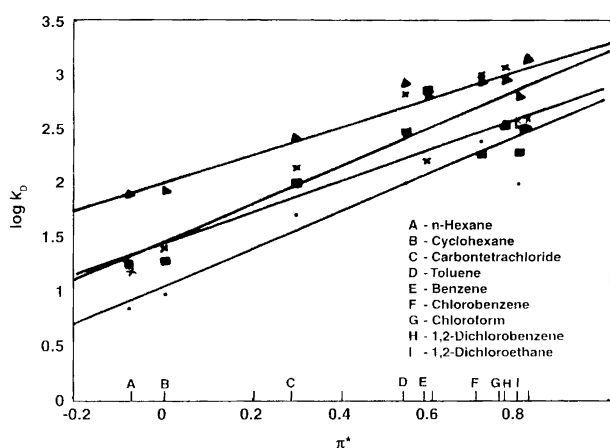Fig. 2. Plot of log K_D vs. π^* for: (●) N-o-tolybenzo-, (×) N-o-tolyl-4-methylbenzo-, (▲) N-p-tolyl-4-bromobenzo- and (■) N-p-tolyl-3,5-dinitrobenzo-hydroxamic acids.

transfer ΔG_t°, of a nonelectrolyte. Thus

$$\Delta G_t^\circ = -RT \ln(C_2/C_1) \tag{9}$$

where $C_2/C_1 = K_D$.

Hildebrand has shown that the primary medium activity coefficient of a nonelectrolyte solute γ_C is given by the equation[39]

$$RT \ln \gamma_C = V_1(\delta_C - \delta)^2 \tag{10}$$

where V_1 and δ_C are the molar volume and solubility parameter of the solute, respectively.

To calculate the molar volume, V_1 ($V_1 =$ mol. wt/density), the densities of these hydroxamic acids were determined at $30 \pm 1\,^\circ\text{C}$. The method was found convenient as well as accurate provided some precautions were taken. The solubility parameters, δ, for water and

Table 3. Various physical parameters of N-aryl-substituted hydroxamic acids

Solvents	Type of solvent	N-o-Tolybenzo-				N-o-Tolyl-4-methylbenzo-				N-p-Tolyl-4-bromobenzo-				N-p-Tolyl-3,5-dinitrobenzo-			
		K_D	δ_C	$\gamma_C \times 10^3$	$\Delta G_t^\circ \times 10^3$	K_D	δ_C	$\gamma_C \times 10^3$	$\Delta G_t^\circ \times 10^3$	K_D	δ_C	$\gamma_C \times 10^3$	$\Delta G_t^\circ \times 10^3$	K_D	δ_C	$\gamma_C \times 10^3$	$\Delta G_t^\circ \times 10^3$
n-Hexane	NHB	6.70	14.23	6503.83	−1.16	14.66	14.21	7208.24	−1.62	75.00	14.07	4292.48	−2.63	15.70	14.22	70127.28	−1.68
Cyclohexane	NHB	8.95	14.62	673.27	−1.34	23.79	14.58	701.74	−1.91	81.82	14.44	431.37	−2.69	18.21	14.55	4041.94	−1.77
Carbon tetrachloride	NHB	49.22	14.62	133.22	−2.38	131.20	14.58	136.80	−2.94	244.44	14.52	117.96	−3.36	95.46	14.72	1502.68	−2.76
Benzene	Ar–NHB–HBA	138.50	15.09	105.37	−2.99	151.10	14.82	44.47	−3.00	625.00	14.78	38.64	−3.93	647.51	14.73	121.60	−3.96
Chlorobenzene	Ar–NHB–HBA	237.50	14.85	11.92	−3.34	972.00	14.71	7.92	−4.14	857.54	14.73	9.09	−4.12	176.00	14.96	73.89	−3.17
1,2-Dichlorobenzene	I	96.25	15.10	4.90	−2.79	340.30	15.02	6.14	−3.48	626.25	14.92	3.27	−3.92	185.63	15.12	19.10	−3.20
Toluene	Ar–NHB–HBA	95.00	14.70	59.22	−2.78	641.20	14.54	41.15	−3.85	795.45	14.50	35.74	−4.06	289.00	14.68	323.31	−3.43
Chloroform	NHB–HBD	341.70	14.79	18.88	−3.55	1148.00	14.68	14.36	−4.24	875.00	14.70	17.13	−4.12	338.50	14.88	136.43	−3.53
1,2-Dichloroethane	NHB	328.00	14.93	5.26	−3.53	401.85	15.11	11.21	−3.61	1343.20	14.86	5.07	−4.38	306.70	15.02	31.17	−3.47
Density/g l⁻¹		1.1428				1.2104				1.5048				1.3773			
Molar volume (V_1)/l		198.63				199.12				203.28				230.15			
1000ρ/M		5.03				5.08				4.92				4.34			

for organic solvents were obtained from the literature.[39] The density of water was taken as 0.995 670 at 30 °C.

In Table 3 are given the values of K_D, δ_C, γ_C, ΔG_t°, ρ, V_1 and $1000\rho/M$ for these metal extractants in a number of common organic solvents.

The data for solubility parameters of these reagents showed that the calculated values of δ_C for the solvents investigated here are reasonably constant. This indicates that the properties of hydroxamic acid solutions in these solvents are at least qualitatively similar to those postulated for regular solutions. Although the quantitative application of these equations is difficult because of the dependence on factors involving a small difference between relatively closely matched solubility parameters, to apply its approach qualitatively, consideration of the solubility parameter, as an important factor in the evaluation of organic solvents for extraction, can be quite useful. A knowledge of solubility parameter values will prove fruitful, not only in choosing appropriate solvents for extraction but also in extending the range of extraction methods.

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