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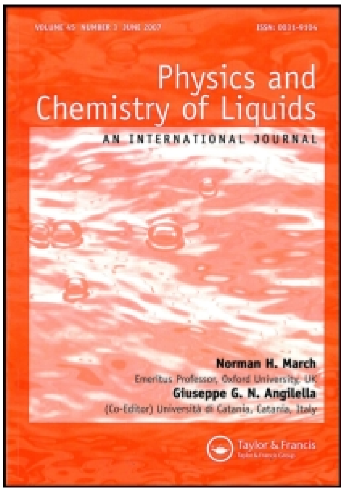

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Solubility of the Pesticide Monuron in Organic Nonelectrolyte Solvents. Comparison of Observed Versus Predicted Values Based upon Mobile Order Theory

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SOLUBILITY OF THE PESTICIDE MONURON IN ORGANIC NONELECTROLYTE SOLVENTS. COMPARISON OF OBSERVED VERSUS PREDICTED VALUES BASED UPON MOBILE ORDER THEORY

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Experimental solubilities are reported at 25.0°C for monuron (also called 3-(4-chlorophenyl)-1,1-dimethyl urea) dissolved in 18 different organic nonelectrolyte solvents containing ether-, chloro-, hydroxy-, ester, methyl- and *t*-butyl-functional groups. Results of these measurements, combined with published literature data, are used to test the applications and limitations of expressions derived from Mobile Order theory. For the 21 nonalcoholic solvents for which predictions could be made computations show that Mobile Order theory does provide fairly reasonable estimates of the saturation mole fraction solubilities. Average absolute deviation between predicted and observed values is 48.4%. Monuron solubilities in the alcohol solvents are used to calculate stability constants for presumed solute–solvent hydrogen bonds that are believed to occur in solution.

Keywords: Pesticide; Monuron solubilities; Organic nonelectrolyte solvents; Solubility predictions

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INTRODUCTION

Free energy of solvation is an important thermodynamic variable that quantifies the free energy difference between a molecule in the gas phase and the molecule dissolved in a solvent. Free energies of solvation provide valuable information regarding molecular interactions between dissolved solute and surrounding solvent molecules, and can be used to calculate numerical values of partition coefficients that describe the equilibrium distribution of a solute between two immiscible liquid phases. For example, the octanol/water partition coefficient is the free energy of solvation of the solute molecule in wet 1-octanol minus its free energy of solvation in water. Solvation free energies and partition coefficients are of critical importance in many pharmaceutical, environmental and chemical engineering applications. Solute partitioning between two immiscible phases is the basis for all chromatographic separations. Correlations have been derived for predicting brainblood partitioning of organic solutes from solutes' measured and/or calculated free energies of solvation [1,2], and for estimating aqueous solubilities, soil adsorptions, bioaccumulations and toxicities of organic compounds from experimental octanol–water partition coefficient data [3–11].

Historically, many of the very early studies focussed exclusively on developing correlational equations based upon upon octanol/water partition coefficients. Recent studies have shown that the octanol–water partition coefficient may not necessarily be the best indicator of how likely it is for a particular solute molecule to penetrate a lipid bilayer, skin, brain or central nervous system, or to accumulate in different tissues and body organs. Experimental studies have been expanded to include additional organic solvents, as well as aqueous micellar solvent media, and to use solute descriptors calculable from structural considerations and/or easily measured thermodynamic quantities.

In this regard, Abraham and coworkers [12–19] developed expressions for describing the partition of solutes between water and a given solvent

$$\log P = c + r \cdot R_2 + s \cdot \pi_2^{\mathrm{H}} + a \cdot \sum \alpha_2^{\mathrm{H}} + b \cdot \sum \beta_2^{\mathrm{H}} + v \cdot V_x \qquad (1)$$

and between the gas phase and a given solvent

$$\log L = c + r \cdot R_2 + s \cdot \pi_2^{\mathrm{H}} + a \cdot \sum \alpha_2^{\mathrm{H}} + b \cdot \sum \beta_2^{\mathrm{H}} + l \cdot \log L^{16}. \quad (2)$$

The dependent variables in Eqs. (1) and (2) are the $\log P$ (the partition coefficient of solute(s) between water and a given solvent) and $\log L$ (Ostwald solubility coefficient). The independent variables are the solute descriptors as follows: R_2 and V_x refer to the excess molar refraction and McGowan volume of the solute, respectively, $\sum \alpha_2^{\mathrm{H}}$ and $\sum \beta_2^{\mathrm{H}}$ are measures of the solute's hydrogen-bond acidity and hydrogen-bond basicity, π_2^{H} denotes the solute's dipolarity/polarizability descriptor, and $\log L^{16}$ is the solute's gas phase dimensionless Ostwald partition coefficient into hexadecane at 298 K. The Ostwald partition coefficient, L, is the inverse of the Henry's law constant (Pa m^3 mole^{-1}). It should be noted that the various c, r, s, a, b, v and l coefficients depend on the solvent phase under consideration. The r-coefficient gives the tendency of the phase to interact with solutes through polarizability-type interactions, mostly via electron pairs. The s-coefficient is a measure of the solvent phase dipolarity/polarity, while the a- and b-coefficients represent the solvent phase hydrogen-bond basicity and hydrogen-bond acidity, respectively. The l- and v-coefficients are a combination of the work needed to create a solvent cavity wherein the solute will reside, and the general dispersion interaction energy between the solute and solvent phase. In the case of partition coefficients, where two solvent phases are involved, the c, r, s, a, b, v and l coefficients represent differences in the solvent phase properties.

Several earlier studies [20–23] developed the computational methodology for determining the various solute descriptors from measured solubility data for crystalline nonelectrolyte solutes dissolved in organic solvents for which the solvent regressional coefficients were known. Solutes studied included *trans*-stilbene, diuron, monuron, buckminsterfullerene and ferrocene. Diuron and monuron had the larger numerical values of the $\sum \alpha_2^{\mathrm{H}}$ and $\sum \beta_2^{\mathrm{H}}$ solute descriptors. Continued development of additional correlation equations requires the establishment of large solubility and activity coefficient databases for each solvent system to be studied. The databases should contain solutes that span as wide of a range of solute descriptors as possible. For this reason, we are in the process of measuring solubility data for several crystalline organic compounds. In the present communication we report monuron (also called 3-(4-chlorophenyl)-1,1-dimethyl urea) solubilities at 25°C in 18 different organic solvents of varying polarity and hydrogen bonding capability. Results of these measurements, combined with our previously reported monuron solubility data [23], are used to further test the applications and limitations of predictive expressions derived from Mobile Order theory.

MATERIALS AND METHODS

Monuron (Aldrich, 99%) was used as received. Methylcyclohexane (Aldrich, 99 + %, anhydrous), *tert*-butylcyclohexane (Aldrich, 99 + %), 2-butanol (Aldrich, 99 + %, anhydrous), 2-pentanol (Acros, 99 + %), 2-methyl-2-butanol (Acros, 99 + %), 2-methyl-1-propanol (Aldrich, 99 + %, anhydrous), 3-methyl-1-butanol (Aldrich, 99 + %, anhydrous), 4-methyl-2-pentanol (Acros, 99 + %), 2-ethyl-1-hexanol (Aldrich, 99 + %), 2-methyl-1-pentanol (Aldrich, 99%), 2-methyl-2-propanol (Arco Chemical Company, 99 + %), cyclopentanol (Aldrich, 99%), ethylbenzene (Aldrich, 99.8%, anhydrous), tetrahydrofuran (Aldrich, 99.9%, anhydrous), dibutyl ether (Aldrich, 99%), dichloromethane (Aldrich, 99.8%, anhydrous), ethyl acetate (Aldrich, HPLC, 99.9%) and butyl acetate (Aldrich HPLC, 99.7%) were stored over molecular sieves and distilled shortly before use. Gas chromatographic analysis showed monuron to have a purity of 99.5 mole percent, and solvent purities to be 99.7 mole percent or better.

Excess solute and solvent were placed in amber glass bottles and allowed to equilibrate in a constant temperature water bath at 25.0± 0.1°C for at least three days (often longer). Attainment for equilibrium was verified both by repetitive measurements after several additional days and by approaching equilibrium from supersaturation by pre-equilibrating the solutions at a higher temperature. Aliquots of saturated monuron solutions were transferred through a coarse filter into a tared volumetric flask to determine the amount of sample and diluted quantitatively with 2-propanol for spectrophotometric analysis at either 254 nm (alkane solvents and dibutyl ether) or 286 nm (all other solvents) on a Bausch and Lomb Spectronic 2000. Concentrations of the dilute solutions were determined from a Beer–Lambert law absorbance versus concentration working curve for nine standard solutions. The calculated molar absorptivity of monuron at 250 nm varied slightly with concentration, ranging from a value of $\varepsilon \approx$ 20,600 L mol^{-1} cm^{-1} (1.90×10^{-5} molar) to $\varepsilon \approx$ 19,880 L mol^{-1} cm^{-1} (7.75×10^{-5} molar). At 278 nm the calculated molar absorptivity of $\varepsilon \approx$ 1.185 L mol^{-1} cm^{-1}was constant over the concentration range of 2.58×10^{-4} molar to 1.29×10^{-3} molar.

Experimental molar concentrations were converted to (mass/mass) solubility fractions by multiplying by the molar mass of monuron, volume(s) of volumetric flask(s) used and any dilutions required to place the measured absorbances on the Beer–Lambert law absorbance versus concentration working curve, and then dividing by the mass of the saturated solution analyzed. Mole fraction solubilities were computed from (mass/mass) solubility fractions using the molar masses of the solute and solvent. Experimental mole fraction solubilities, X_A^{sat}, are listed in Table I for monuron dissolved in 41 organic solvents. Numerical values were obtained using monuron as received from the supplier, and they represent the average of between four and eight independent determinations. The measurements were reproducible to within ±2%. We did recrystallize several grams of the solid from methanol for solubility determinations in a few select solvents (2-butanol, methylcyclohexane, tetrahydrofuran and butyl acetate). To within the stated experimental uncertainty of ±2% there was no difference in the experimental solubilities using the recrystallized and unrecrystallized samples.

RESULTS AND DISCUSSION

Solvents listed in Table I include both noncomplexing alkanes and self-associating alcohols. Of the many solution models proposed in recent years, mobile order theory is perhaps the only one that is capable of describing solute behavior in such a wide range of solvent mixtures. The basic model [24–31] assumes that all molecular groups perpetually move, and that neighbors of a given kind of external atom in a molecule constantly change identity. All molecules of a given kind dispose of the same volume, equal to the total volume V of the liquid divided by the number N_A of molecules of the same

TABLE I Comparison between experimental monuron mole fraction solubilities and predicted values based upon mobile order theory

Organic solvent	$(X_A^{sat})^{exp,a}$	*Data reference*	$(X_A^{sat})^{calc}$	*% Dev*[b]
n-Hexane	0.00005489	[23]	0.00005058	−7.9
n-Heptane	0.00005565	[23]	0.00005673	1.9
n-Octane	0.00006571	[23]	0.00007112	8.2
n-Nonane	0.00007811	[23]	0.00009228	18.1
n-Decane	0.00010076	[23]	0.0001008	0.0
n-Hexadecane	0.00008653	[23]	0.0001825	110.9
Cyclohexane	0.00005088	[23]	0.00007086	39.3
Methylcyclohexane	0.00007012	This work	0.00008531	21.7
2,2,4-Trimethylpentane	0.00004697	[23]	0.00003662	−22.0
tert-Butylcyclohexane	0.00009920	This work	0.0001491	50.3
Benzene	0.001365	[23]	0.003444	152.3
Toluene	0.001155	[23]	0.001761	52.5
Ethylbenzene	0.0007931	This work	0.001612	103.2
Dibutyl ether	0.001383	This work	0.0009972	−27.9
Tetrahydrofuran	0.02643	This work	0.004502	−83.0
1,2-Dichloroethane	0.006743	[23]	0.01102	63.4
Dichloromethane	0.009436	This work	0.01015	7.6
Chloroform	0.01240	[23]	0.003180	−74.4
Carbon tetrachloride	0.0003361	[23]	0.0007384	119.7
Butyl acetate	0.008675	This work	0.004923	−43.3
Ethyl acetate	0.01007	This work	0.009278	−7.9
Methanol	0.01264	[23]	0.01236[c]	
Ethanol	0.01142	[23]	0.01133[c]	
1-Propanol	0.01287	[23]	0.01277[c]	
2-Propanol	0.008095	[23]	0.01085[c]	
1-Butanol	0.01358	[23]	0.01427[c]	
2-Butanol	0.009261	This work	0.009343[c]	
2-Methyl-1-propanol	0.01060	This work	0.01069[c]	
2-Methyl-2-propanol	0.006547	This work	0.007075[c]	
1-Pentanol	0.01483	[23]	0.01512[c]	
2-Pentanol	0.01056	This work		
3-Methyl-1-butanol	0.01249	This work		
2-Methyl-2-butanol	0.004726	This work		
1-Hexanol	0.01496	[23]	0.01462[c]	
2-Methyl-1-pentanol	0.01206	This work		
4-Methyl-2-pentanol	0.008105	This work		
1-Heptanol	0.01478	[23]	0.01433[c]	
1-Octanol	0.01457	[23]	0.01400[c]	
2-Ethyl-1-hexanol	0.01096	This work		
1-Decanol	0.01320	[23]	0.01314[c]	
Cyclopentanol	0.01534	This work		

[a] Numerical values represent the average of between four and eight independent determinations, with the measurements being reproducible to ± 2.0%.

[b] Deviations (%) = $100\,[(X_A^{sat})^{calc} - (X_A^{sat})^{exp}]/(X_A^{sat})^{exp}$.

[c] Numerical values of K_{Oi} used were: $K_{Oi} = 2{,}200\,cm^3\,mol^{-1}$ for methanol; $K_{Oi} = 4{,}000\,cm^3\,mol^{-1}$ for ethanol and 2-propanol; $K_{Oi} = 6{,}000\,cm^3\,mol^{-1}$ for 1-propanol; $K_{Oi} = 7{,}000\,cm^3\,mol^{-1}$ for 1-butanol; $K_{Oi} = 8{,}000\,cm^3\,mol^{-1}$ for 2-butanol; $K_{Oi} = 9{,}500\,cm^3\,mol^{-1}$ for 1-pentanol; and $K_{Oi} = 14{,}000\,cm^3\,mol^{-1}$ for 2-methyl-1-propanol, 2-methyl-2-propanol and for 1-hexanol through 1-decanol.

kind, i.e., Dom $A = V/N_A$. The center of this domain perpetually moves. The highest mobile disorder is achieved whenever groups visit all parts of their domain without preference. Preferential contacts lead to deviations with respect to this "random" visiting. This is especially true in the case of hydrogen-bonding as specific interactions results in a specific orientation of the "donor" molecule with respect to an adjacent "acceptor" molecule.

In the case of an inert crystalline solute dissolved in a self-associating solvent, mobile order theory expresses the volume fraction saturation solubility, Φ_A^{sat}, as

$$\begin{aligned}\ln \Phi_A^{sat} = {} & \ln a_A^{solid} - 0.5(1 - V_A/V_{solvent})\Phi_{solvent} \\ & + 0.5 \ln[\Phi_A^{sat} + \Phi_{solvent}(V_A/V_{solvent})] \\ & - \Phi_{solvent}^2 V_A(\delta'_A - \delta'_{solvent})^2 (RT)^{-1} \\ & - r_{solvent}(V_A/V_{solvent})\Phi_{solvent}, \end{aligned} \qquad (3)$$

where the $r_{solvent}(V_A/V_{solvent})\Phi_{solvent}$ term represents the contributions resulting from hydrogen-bond formation between the solvent molecules. For most of the published applications, $r_{solvent}$ was assumed to be unity for strongly associated solvents with single hydrogen-bonded chains such as monofunctional alcohols, to be two for water or diols, and to equal zero for nonassociated solvents such as saturated hydrocarbons. A more exact value for alcoholic solvents can be calculated based upon

$$r_{solvent} = (K_{solvent}\Phi_{solvent}/V_{solvent})/(1 + K_{solvent}\Phi_{solvent}/V_{solvent}) \qquad (4)$$

with a numerical value of $K_{solvent} = 5{,}000\,\text{cm}^3\,\text{mol}^{-1}$ assumed for all monofunctional alcohols.

If complexation does occur between the crystalline solute and solvent

$$\begin{aligned}\ln \Phi_A^{sat} = {} & \ln a_A^{solid} - 0.5(1 - V_A/V_{solvent})\Phi_{solvent} \\ & + 0.5 \ln[\Phi_A^{sat} + \Phi_{solvent}(V_A/V_{solvent})] \\ & - \Phi_{solvent}^2 V_A(\delta'_A - \delta'_{solvent})^2 (RT)^{-1} \\ & + \ln[1 + \Phi_{solvent}(K_{Asolvent}/V_{solvent})] \end{aligned} \qquad (5)$$

then an additional term involving the solute–solvent equilibrium constant, K_{ASolvent}, must be introduced to describe the solubility enhancement that arises as a result of specific interactions. A slightly more complex expression applies in the case of solute complexation with a self-associating solvent. The expression's mathematical form depends to a large extent upon the number and type of hydrogen-bond acid and hydrogen-bond base functional groups in the solute molecule being studied.

In the present case, monuron contains an -NC(O)NH- moiety which could conceivably act as both a hydrogen-bond base (through the lone electron pairs on the oxygen and nitrogen atoms) and hydrogen-bond acid. Examination of the solubility data in Table I reveals that monuron is twenty times more soluble in dibutyl ether $(X_{\mathrm{A}}^{\mathrm{sat}} = 0.001383)$ than in the two saturated hydrocarbons, *n*-octane ($X_{\mathrm{A}}^{\mathrm{sat}} = 0.00006571$) and *n*-nonane ($X_{\mathrm{A}}^{\mathrm{sat}} = 0.00007811$) of comparable size. The large solubility enhancement observed by the introduction of the ether functional group into an alkane solvent suggests formation of a monuron-dibutyl ether association complex. If such complexation does occur then monuron's -NC(O)NH- functional group must act as a hydrogen-bond acid. The ether oxygen on dibutyl ether has no choice other than to act as a hydrogen-bond base. We do not believe, however, that the observed solubility enhancement results from complexation. As shown in the third and fourth columns of Table I, Eq. (3) of Mobile Order theory predicts the mole fraction solubility of monuron in dibutyl ether to within 28% without introducing any solute–solvent stability constants. Deviations of similar magnitude are noted in the case of the saturated hydrocarbon solvents, which are incapable of forming solute–solvent complexes. Based upon these observations we feel that if monuron does interact specifically with alcohol solvent molecules in solution, the -NC(O)NH- functional group will act as a hydrogen-bond base, rather than as a hydrogen-bond base.

An "O contribution" (see Ruelle *et al.* [30,31])

$$\mathrm{O} = \sum \nu_{\mathrm{Oi}} \ln\{1 + K_{\mathrm{Oi}}[(\Phi_{\mathrm{solvent}}/\mathrm{V}_{\mathrm{solvent}}) - \nu_{\mathrm{Oi}}(\Phi_{\mathrm{A}}^{\mathrm{sat}}/\mathrm{V}_{\mathrm{A}})]\} \qquad (6)$$

must be added to Eq. (3) to describe the effect on the solubility of H-bonds formed between the hydrogen-bond base sites on the solute and hydrogen-bond acid sites on the solvent. Each particular H-bond interaction like oxygen–alcohol (or nitrogen–alcohol) contributes to increase solubility, and is characterized by the group interaction stability constant K_{Oi}. For crystalline proton-acceptor solute molecules having limited solubility ($\Phi_{\mathrm{A}}^{\mathrm{sat}} \approx 0$) the predictive expression becomes

$$\begin{aligned} \ln \Phi_{\mathrm{A}}^{\mathrm{sat}} = {} & \ln a_{\mathrm{A}}^{\mathrm{solid}} - 0.5(1 - V_{\mathrm{A}}/V_{\mathrm{solvent}})\Phi_{\mathrm{solvent}} \\ & + 0.5\ \ln[\Phi_{\mathrm{A}}^{\mathrm{sat}} + \Phi_{\mathrm{solvent}}(V_{\mathrm{A}}/V_{\mathrm{solvent}})] \\ & - \Phi_{\mathrm{solvent}}^{2} V_{\mathrm{A}}(\delta'_{\mathrm{A}} - \delta'_{\mathrm{solvent}})^{2}\,(RT)^{-1} \\ & - r_{\mathrm{solvent}}(V_{\mathrm{A}}/V_{\mathrm{solvent}})\Phi_{\mathrm{solvent}} \\ & + \sum v_{\mathrm{Oi}} \ln\{1 + K_{\mathrm{Oi}}(\Phi_{\mathrm{solvent}}/V_{\mathrm{solvent}})\}, \end{aligned} \tag{7}$$

where v_{Oi} is the number of active and independent type *i* proton-acceptor sites on the solute molecule. Equation (7) assumes specific interactions in solution, but not in the solid state. The crystalline phase in equilibrium with the saturated solution is taken to be pure monuron. The existence of molecular association complexes in solution does not necessarily require that the complex must also exist in the solid phase. Packing constraints in the crystalline phase may prevent the formation of solid molecular complexes. The symbols δ'_{A} and $\delta'_{\mathrm{solvent}}$ denote the modified solubility parameters of the solute and solvent, respectively, V_{i} is the molar volume, and $a_{\mathrm{A}}^{\mathrm{solid}}$ is the activity of the solid solute. This latter quantity is defined as the ratio of the fugacity of the solid to the fugacity of the pure hypothetical supercooled liquid. The numerical value of $a_{\mathrm{A}}^{\mathrm{solid}}$ can be computed from

$$\begin{aligned} a_{\mathrm{A}}^{\mathrm{solid}} = {} & -\Delta H_{\mathrm{A}}^{\mathrm{fus}}(T_{\mathrm{mp}} - T)/(RTT_{\mathrm{mp}}) + (\Delta C_{\mathrm{p,A}}/RT)(T_{\mathrm{mp}} - T) \\ & - (\Delta C_{\mathrm{p,A}}/T)\ln(T_{\mathrm{mp}}/T) \end{aligned} \tag{8}$$

the solute's molar enthalpy of fusion, $H_{\mathrm{A}}^{\mathrm{fus}}$, and heat capacity difference, $\Delta C_{\mathrm{p,A}}$, at the normal melting point temperature, T_{mp}. Contributions from nonspecific interaction are incorporated into mobile order theory through the $\Phi_{\mathrm{solvent}}^{2} V_{\mathrm{A}}(\delta'_{\mathrm{A}} - \delta'_{\mathrm{solvent}})^{2}\,(RT)^{-1}$ term.

It is noted that other research groups have treated nonspecific interactions differently, and have assumed solution models other than the Scatchard–Hildebrand solubility parameter theory. For example, Ruelle [32] in predicting solubilities in systems involving hydrogen-bond formation between a dissolved solute and surrounding solvent molecules, modified the Scatchard–Hildebrand expression by multiplying it by the fraction of time during which the solute is not bound to the solvent [i.e., during which the distribution between the solvent and unbound solute molecules can still be considered to occur at random]. Ruelle's treatment further assumed that nonspecific interactions involving the bound solute were negligible. The theoretical justification for Ruelle's modification was not given, and it is not clear to us what mathematical form the integral $(\Delta G_{\mathrm{ASolvent}})_{\mathrm{phys}}$ for the binary solution would have to take in order to give

$$\begin{aligned}(\Delta G_{\mathrm{A}})_{\mathrm{phys}} = {} & \{1/[1.0 + \max(K_{\mathrm{Oi}}, K_{\mathrm{OHi}})(\Phi_{\mathrm{solvent}}/V_{\mathrm{solvent}})]\} \\ & \cdot \Phi^2_{\mathrm{solvent}} V_{\mathrm{A}}(\delta'_A - \delta'_{\mathrm{solvent}})^2 \end{aligned} \tag{9}$$

whenever $(\Delta G_{\mathrm{ASolvent}})_{\mathrm{phys}}$ is differentiated with respect to the number of moles of solute present. The differentiation is required in deriving the solubility equation. Readers are reminded that any modification to the $(\Delta G_{\mathrm{A}})_{\mathrm{phys}}$ expression must also show up in the corresponding $(\Delta G_{\mathrm{solvent}})_{\mathrm{phys}}$ expression, the latter expression being obtained by differentiating $(\Delta G_{\mathrm{ASolvent}})_{\mathrm{phys}}$ with respect to the number of molecules of solvent. In Eq. (9) $\max(K_{\mathrm{Oi}}, K_{\mathrm{OHi}})$ stands for the association constant governing the strongest intermolecular H-bond displayed by the molecular groups in solution. We have elected to use the Scatchard–Hildebrand solubility parameter theory, rather than the Ruelle modification, because we have serious reservations about whether nonspecific interactions for bound molecules are truly negligible as assumed by Ruelle in proposing Eq. (9).

Predictive application of Eqs. (3), (5) and (7) is relatively straightforward. First, an average numerical value of $\delta'_{\mathrm{monuron}} = 24.37\ \mathrm{MPa}^{1/2}$ is computed by requiring that each equation (with $r_{\mathrm{solvent}} = 0$ and/or $\mathrm{K}_{\mathrm{ASolvent}} = 0$) perfectly describes monuron mole fraction solubility data in *n*-hexane ($\delta'_{\mathrm{monuron}} = 24.30\ \mathrm{MPa}^{1/2}$), *n*-heptane ($\delta'_{\mathrm{monuron}} = 20.38\ \mathrm{MPa}^{1/2}$), and *n*-octane ($\delta'_{\mathrm{monuron}} = 24.44\ \mathrm{MPa}^{1/2}$). The numerical value of $a_{\mathrm{A}}^{\mathrm{solid}} = 0.01891$ is calculated using Eq. (6) with $\Delta H_{\mathrm{A}}^{\mathrm{fus}} = 29{,}460\ \mathrm{J\,mol}^{-1}$ and $T_{\mathrm{mp}} = 447.6\ \mathrm{K}$ [33]. The two heat capacity terms in Eq. (8) were set equal to zero as we were unable to find heat capacity data for the subcooled liquid solute. A numerical value of

$V_{\text{monuron}} = 152.8\,\text{cm}^3\,\text{mol}^{-1}$ was used for the molar volume of the hypothetical subcooled liquid solute. The solute's molar volume was estimated using a group contribution method and molar volumes of chlorobenzene and organic molecules having amide and methyl functional groups.

Table I summarizes the predictive ability of Mobile Order theory for the various organic solvents for which both monuron solubility data and modified solubility parameters could be found. Solvent molar volumes and modified solubility parameters are listed in Table II. We were unable to find modified solubility parameters for all of the secondary and branched alcohols. The modified solubility parameters account for only nonspecific interactions, and in the case of the alcoholic solvents the hydrogen-bonding contributions have been removed. Numerical values of δ'_{solvent} were obtained from published compilations [27, 28, 30, 31], and were either deduced by regressing actual solubility data of solid *n*-alkanes in organic solvents in accordance with the configurational entropic model of Huyskens and Haulait-Pirson [34] or estimated using known values for similar organic solvents. Examination of the entries in Table I reveals that Mobile Order theory does provide fairly reasonable (though by no means perfect) estimates of the solubility behavior of monuron in a wide range of nonalcoholic organic solvents. Average absolute deviation between predicted and observed values is 48.4%. Expressed on a natural logarithmic mole fraction scale, Mobile Order theory predicts the monuron solubilities to within an average absolute deviation of ± 0.47 ln units. The arithmetic average of the $\ln X_A^{\text{sat}}$ difference is 0.04, which indicates a slight systematic error in the predicted values.

Alcoholic solvents were excluded from all statistical analyses because the observed mole fraction solubility was used to determine the K_{Oi} stability constant for each monuron–alcohol solvent pair in solution using Eq. (7) with v_{Oi} set equal to one. Hence, the numerical entries in the third column of Table I represent back-calculated mole fraction solubilities, rather than predicted values. The numerical value of K_{Oi} was found to increase with increasing alcohol size, ranging from a lower limit of about $K_{\text{Oi}} \approx 2{,}200\,\text{cm}^3\,\text{mol}^{-1}$ for methanol to an upper value of about $K_{\text{Oi}} \approx 14{,}000\,\text{cm}^3\,\text{mol}^{-1}$ for 1-hexanol through 1-decanol. This observation was not unexpected in that monuron has slightly larger mole fraction solubility in the larger alcohols

TABLE II Solvent and solute properties used in mobile order predictions

Component (i)	$V_i/(\mathrm{cm^3\ mol^{-1}})$	$\delta_i'/(\mathrm{MPa^{1/2}})$[a]
n-Hexane	131.51	14.56
n-Heptane	147.48	14.66
n-Octane	163.46	14.85
n-Nonane	179.87	15.07
n-Decane	195.88	15.14
n-Hexadecane	294.12	15.61
Cyclohexane	108.76	14.82
Methylcyclohexane	128.32	15.00
2,2,4-Trimethylpentane	166.09	14.30
tert-Butylcyclohexane	173.9	15.50
Benzene	89.4	18.95
Toluene	106.84	18.10
Ethylbenzene	123.1	18.02
Dibutyl ether	170.3	17.45
Tetrahydrofuran	81.4	19.30
Methanol	40.7	19.25
Ethanol	58.7	17.81
1-Propanol	75.10	17.29
2-Propanol	76.90	17.60
1-Butanol	92.00	17.16
2-Butanol	92.4	16.60
2-Methyl-1-propanol	92.8	16.14
2-Methyl-2-propanol	94.3	15.78
1-Pentanol	108.6	16.85
1-Hexanol	125.2	16.40
1-Heptanol	141.9	16.39
1-Octanol	158.3	16.38
1-Decanol	191.6	16.35
Ethyl acetate	98.5	20.79
Butyl acetate	132.5	19.66
1,2-Dichloroethane	78.8	20.99
Tetrachloromethane	97.08	17.04
Dichloromethane	64.5	20.53
Chloroform	80.7	18.77
Monuron[b]	152.8[c]	24.37[d]

[a] Tabulated values are taken from a compilation given in Ruelle *et al.* [27,28,30,31].

[b] The numerical value of $a_\mathrm{A}^\mathrm{solid} = 0.01891$ was calculated from Eq. (8) using $\Delta H_\mathrm{A}^\mathrm{fus} = 29{,}460\,\mathrm{J\,mol^{-1}}$ and $T_\mathrm{mp} = 447.6\,\mathrm{K}$ [33].

[c] Numerical value of $V_\mathrm{monuron} = 152.8\,\mathrm{cm^3\,mol^{-1}}$ was estimated using a group contribution method and molar volumes of chlorobenzene and molecules having amide and methyl functional groups.

[d] Numerical value was calculated using the measured monuron mole fraction solubilities in *n*-hexane, *n*-heptane and *n*-octane, in accordance with Eqs. (3) and (5); with $r_\mathrm{solvent} = 0$ and/or $K_\mathrm{ASolvent} = 0$.

than in the much smaller methanol and ethanol solvents. From a strictly mathematical point-of-view, this requires that the "O contribution" in Eq. (7) must become larger with increasing alcohol. Moreover, the increased "O contribution" must offset the fact that the negative $-\Phi^2_{\text{solvent}} V_A(\delta'_A - \delta'_{\text{solvent}})^2 (RT)^{-1}$ term becomes more negative as δ'_{solvent} moves further away from δ'_{monuron}. This is what happens as one progresses in alcohol size from methanol to 1-decanol. The modified solubility parameter of methanol ($\delta'_{\text{methanol}} = 19.25\ \text{MPa}^{1/2}$) is much closer to the modified solubility parameter of monuron ($\delta'_{\text{monuron}} = 24.37\,\text{MPa}^{1/2}$) than is the modified solubility parameter of 1-decanol ($\delta'_{\text{decanol}} = 16.35\,\text{MPa}^{1/2}$). Without the "O contribution" and without a larger K_{Oi} value, Mobile Order theory would have predicted that monuron would be significantly less soluble in 1-decanol than in methanol. Such predictions would be inconsistent with the observed mole fraction solubility data.

Readers are reminded that in evaluating the applicability of Mobile Order theory one must realize that many of these particular systems are highly nonideal, and that the experimental solubility data covers over a 1,700-fold range in mole fraction. Had an ideal solution been assumed, then the predicted mole fraction solubility would be $X_A^{\text{sat}} = a_A^{\text{solid}} = 0.01891$ for each solvent. The ideal solution approximation corresponds to a considerably larger average absolute deviation of 14,140% between predicted and observed values for the nonalcoholic solvents studied.

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