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Predicting Stability Constants for Uranyl Complexes Using Density Functional Theory

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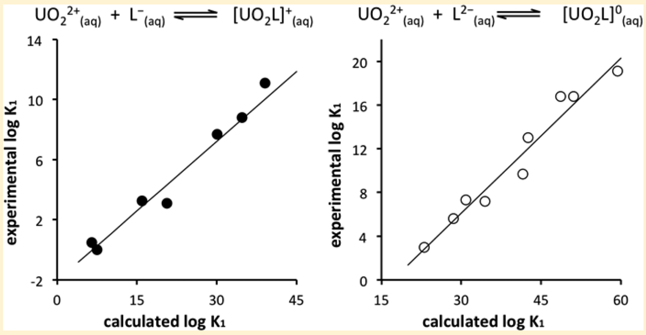
Ⓢ *Supporting Information*

ABSTRACT: The ability to predict the equilibrium constants for the formation of 1:1 uranyl/ligand complexes (log K_1 values) provides the essential foundation for the rational design of ligands with enhanced uranyl affinity and selectivity. We use density functional theory (B3LYP) and the integral equation formalism polarizable continuum model (IEF-PCM) to compute aqueous stability constants for UO_2^{2+} complexes with 18 donor ligands. Theoretical calculations permit reasonably good estimates of relative binding strengths, while the absolute log K_1 values are significantly overestimated. Accurate predictions of the absolute log K_1 values (root-mean-square deviation from experiment <1.0 for log K_1 values ranging from 0 to 16.8) can be obtained by fitting the experimental data for two groups of mono- and divalent negative oxygen donor ligands. The utility of correlations is demonstrated for amidoxime and imide dioxime ligands, providing a useful means of screening for new ligands with strong chelating capability to uranyl.



■ INTRODUCTION

Understanding the speciation of metal ions in aqueous solution is the basis for predicting and controlling the behavior of metal ions in biological systems, the environment, analytical chemistry, extraction metallurgy, and other industrial processes.[1] This understanding also provides the essential foundation for the rational design of ligands with enhanced metal ion affinity and selectivity. Central to this endeavor, an important and fundamental characterization of ligand reactivity involves the determination of equilibrium constants for the formation of simple 1:1 metal/ligand complexes with a series of metal ions. Such values, which are referred to as stability constants or formation constants, are determined most frequently in aqueous solution. In practice, stability constants are normally defined as the ratio of concentrations rather than activities. In the case of 1:1 complexes, which are the focus of this article, the aqueous stability constant is defined as the concentration of the metal complex, ML, divided by the concentrations of the metal ion, M, and the ligand, L, as shown by eq 1. Because they may span many orders of magnitude, stability constants are usually reported in logarithmic form, log K_1. Individual log K_1 values quantify the strength of the metal–ligand interaction, and the difference between log K_1 values for two metal ions measures the degree of selectivity.

$$M + L \rightleftharpoons ML \qquad K_1 = \frac{[ML]}{[M][L]} \tag{1}$$

A large number of stability constants for metal complexes have been determined and tabulated.[2−4] Yet, despite an extensive research effort in this field, available log K_1 data represent a very small fraction of conceivable complexes. Given the number of possible metal ions and an infinite variety of potential ligands, it is not feasible to determine experimentally all log K_1 values. Moreover, experimental limitations, such as redox stability, hydrolysis, and solubility, make it difficult or impossible to measure the stability constant for certain ML complexes. For these reasons, there is a need for reliable methods that allow one to estimate unknown log K_1 values.

Over the past 50 years, many techniques for estimating unknown stability constants from existing log K_1 data have been reported.[1,5] Successful approaches include (i) linear correlations between log K_1 for one group of reactions and log K_1 for a second group of reactions, (ii) correlations between log K_1 and metal ion properties such as ionic radius, formal charge, electronegativity, or ionization potential, and (iii) parametric equations in which coefficients are adjusted to reproduce a training set of log K_1 data. Although useful, none of these approaches are universally applicable and most require the availability of experimental log K_1 data that are closely related to the unknown system before an accurate prediction is possible. This requirement is problematic with the design of novel ligands for which no data is available.

Given developments in computational methods and the vast increase in computational power over the past decade, the question arises as to whether it is now possible to calculate aqueous stability constants. One approach for making this prediction is summarized in Figure 1. Experimental concen-

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tration-based log K_1 values are related to free energy change for the reaction, ΔG^*_{aq}, by eq 2

$$\log K_1 = \log \frac{[ML]}{[M][L]} \approx \log \frac{a_{ML}}{a_M a_L} = \frac{-\Delta G_{aq}}{2.303RT} \tag{2}$$

where a_{ML}, a_M, and a_L are the activities of the species involved in the equilibrium. In principle, ΔG^*_{aq} can be computed using the thermodynamic cycle shown in Figure 1. In this case, ΔG^*_{aq} is given by eq 3

$$\begin{aligned} \Delta G_{aq} = {} & G^*_{aq}([ML(OH_2)_{m-n}]^{x+y}) + nG^\circ_{aq}(H_2O) \\ & - G^*_{aq}([M(OH_2)_m]^x) - G^*_{aq}(L^y) \end{aligned} \tag{3}$$

where for each species $G^*_{aq} = G^\circ_g + \Delta G^{\circ\to*} + \Delta G^*_{solv}$. In this approach, the calculated gas-phase free energies (G°_g) are defined with the standard state being that of an ideal gas at 1 atm, whereas the free energies of solvation (ΔG^*_{solv}) represent the change in free energy for transfer of 1 mol of solute from the gas phase to the aqueous phase at a standard state of 1 M. The conversion of an ideal gas from 1 atm (24.46 L mol^{-1}) to 1 M (1 mol L^{-1}) is given by

$$\begin{aligned} \Delta G^{\circ\to*} &= -T\Delta S^{\circ\to*} = RT\ \ln(V_0/V^*) = RT\ \ln(24.46) \\ &= 1.89\ \text{kcal/mol} \quad (T = 298.15\ \text{K}) \end{aligned} \tag{4}$$

Applying eq 4 to each gas-phase reactant and product yields the correction given in the upper leg of the thermodynamic cycle shown in Figure 1. With water, an additional correction, $G^l_{aq} = G^*_{aq} + RT\ln([H_2O])$, is needed if the pure solvent $H_2O_{(l)}$ is adopted as the reference state for the solvent in the lower leg of the thermodynamic cycle. Here, $RT\ln([H_2O]) = 2.38$ kcal/mol is a free energy change associated with moving a solvent from a standard-state solution phase concentration of 1 M to a standard state of the pure liquid, 55.34 M. Although methods for calculating both G°_g and ΔG^*_{solv} have been available for some time, it is only in the past few years that the predictive capability of this approach has been systematically and critically evaluated for metal complexes.

$$\begin{tikzcd}
{[M(OH_2)_m]^x_{(g)}} \arrow[d, "\Delta G^*_{solv}"] & + & L^y_{(g)} \arrow[d, "\Delta G^*_{solv}"] \arrow[r, "\Delta G^\circ_g + (n-1)\Delta G^{\circ\to*}"] & {[ML(OH_2)_{m-n}]^{x+y}_{(g)}} \arrow[d, "\Delta G^*_{solv}"] & + & n\,H_2O_{(g)} \arrow[d, "\Delta G^*_{solv}"] \\
{[M(OH_2)_m]^x_{(aq)}} & + & L^y_{(aq)} \arrow[r, "\Delta G_{aq} - nRT\ln([H_2O])"] & {[ML(OH_2)_{m-n}]^{x+y}_{(aq)}} & + & n\,H_2O_{(l)}
\end{tikzcd}$$

Figure 1. Thermodynamic cycle showing how ΔG_{aq} is related to ΔG°_g and ΔG^*_{solv}.

Following the approach illustrated in Figure 1, electronic structure calculations at various levels of theory with two common solvation models, COSMO-RS and SMD, were used to calculate aqueous log K_1 values for complexes of a series of divalent transition metals with simple ligands such as imidazole and acetate.[6] Reference values for the gas-phase free energies were obtained at the CCSD(T)/aug-cc-pVTZ level of theory and used to benchmark less expensive methods such as RI-MP2 and various DFT functionals. Even at the highest levels of theory, comparison of calculated and experimental stability constants revealed large discrepancies ranging from −18 to +11 orders of magnitude prompting the authors to conclude "although the agreement between the experimental (as derived from the stability constants) and calculated values is often within 5 kcal/mol, in more complicated cases, it may exceed 15 kcal/mol"[6a] and "the accuracy of "absolute" values is still out of reach".[6b] These studies demonstrate that the prediction of stability constants for metal−ligand complexes is far from straightforward, with errors arising from the offset of two large terms, gas-phase complexation energies and solvation energies. It was noted that an accurate description of ion solvation remains a main stumbling block to the quantitative prediction of log K_1 values. This observation is reinforced by recent efforts to compute stability constants for a series of uranyl complexes with phosphate ligands[7] and changes in the free energy for water exchange reactions in the uranyl ion[8] where it was concluded that the prediction of the solution energetics still requires the continued development of solvent modeling approaches because small changes in the models can lead to large changes in reaction energies in solution. One strategy to improve the accuracy of solvation calculations for ions is to include chemically important solute−solvent interactions explicitly into the quantum chemical model. For example, explicit modeling of the second hydration shell around divalent metal ions was critical for the accurate estimation of ion solvation free energies and reaction free energies in aqueous solution.[7−9] The limitation of this approach is that as the number of solvent molecules increases, an adequate sampling of solute−solvent clusters becomes increasingly difficult.

Although it would appear that it is not yet possible to compute absolute log K_1 values for multivalent metal ions with a predictable degree of accuracy using the approach embodied in Figure 1 and eq 3 when treating only the first hydration shell explicitly, it has been observed that calculated energy differences can correctly predict the relative stability for series of related metal complexes. An early example[10] is provided by the hydrolysis of trivalent metal ions, Al^{3+}, Sc^{3+}, Ti^{3+}, V^{3+}, Cr^{3+}, Mn^{3+}, Fe^{3+}, Ga^{3+}, and Y^{3+}, as shown in eq 5

$$[M(OH_2)_6]^{3+} \rightleftharpoons [M(OH)(OH_2)_5]^{2+} + H^+ \tag{5}$$

Values of ΔE were calculated at the B3LYP/DZVP2 level of theory. A plot of ΔE versus experimental log K was linear suggesting that for this series of complexes the dominant effects giving rise to the relative basicities are captured by this level of theory.

In another study, DFT calculations were used to compute the ΔG_g values for the equilibrium involving the exchange of a water molecule for ammonia with wide range of metal ions[11] as shown in eq 6

$$[M(OH_2)_6]^{n+} + NH_3 \rightleftharpoons [M(NH_3)(OH_2)_5]^{n+} + H_2O \tag{6}$$

Separate linear correlations between calculated ΔG_g and experimental ΔG_{aq} for metal ions of different charges (M^{2+}, M^{3+}, and M^{4+}) were obtained. The benefit of such correlations was demonstrated by using them to estimate log K_1 values for metal complexes that are impossible to determine experimentally. In a recent example, a B3LYP/6-31+G(d)/CPCM model was used to correlate aqueous log K_1 values for Fe^{3+} complexes with a series of substituted 3-hydoxypyridin-4-ones.[12] To simplify calculations and decrease computing time, a value termed ΔG^* was obtained as the difference between G^*_{aq} for FeL^{2+} in the absence of any coordinated water molecules and G^*_{aq} for L^-. A plot of ΔG^* versus experimental log K_1 gave a linear correlation that predicted experimental log K_1 values spanning a range of 6 orders of magnitude to within an average

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error of ±0.32 log units. It is important to note that in the aforementioned modeling studies[10−12] the formal charge of a ligand was constant. This allowed proper calibration of applied solvation models to provide a reasonable estimate of the differential effect of solvation on selected model reactions.

Motivated by research toward the design of ligands for the extraction of the uranyl ion from seawater,[13] we evaluated the use of the thermodynamic cycle shown in Figure 1 to compute aqueous stability constants for UO_2^{2+} complexes with negative oxygen donor ligands. In what follows, we document a straightforward approach that is able to rank-order the stability constants for 13 complexes spanning nearly 17 orders of magnitude. Although calculated aqueous log K_1 values are overestimated by 12−39 log units, a plot of experimental log K_1 versus calculated log K_1 for two groups of ligands with a formal charge of −2 and −1 give linear correlations, $R^2 \geq 0.950$, allowing the reliable prediction of log K_1 values for three reported amidoxime and imide dioxime ligands. The approach is utilized to predict the log K_1 values for two novel oxime-based uranophiles.

■ METHODS

Electronic structure calculations were performed with the Gaussian 09, revision B.01,[14] and NWChem, version 6.3,[15] program packages. We used the B3LYP flavor of density functional theory throughout the study.[16] The standard Stuttgart small-core (SSC) 1997 relativistic effective core potential (RECP) was used for uranium, replacing 60 core electrons to account for scalar relativistic effects.[17] The valence electrons in this basis set are represented by a contracted [8s/7p/6d/4f] basis. The standard 6-311++G** basis set with diffuse functions was employed for the light atoms. Frequency calculations were performed at the B3LYP/SSC/6-31+G* basis set to verify that geometries were minima and to compute zero point energies and thermal corrections using the rigid rotor−harmonic oscillator approximation. These values were used to calculate ΔG_g and gas phase log K_1 values in accord with eq 7 below. B3LYP with the Stuttgart small-core RECP has been widely used for actinyl complexes, yielding geometries and energetics that are in good agreement with experimental data.[18]

Using the gas phase geometries, we calculated aqueous solvation free energies, ΔG^*_{solv}, with the integral equation formalism polarizable continuum model (IEF-PCM) method[19] at the B3LYP/SSC/6-31+G* level using the default atomic radii of the united force field (UFF). As shown in Figure 1, a standard state correction of $\Delta G^{\circ\to *}$ (eq 4) was applied for each reactant and product and a concentration correction of $RT \ln([H_2O])$ was applied for each water molecule liberated on formation of the complex. Optimized atomic coordinates and energies for all reported structures, as well as a sample Gaussian input file, are provided as Supporting Information.

■ RESULTS AND DISCUSSION

A series of 13 negative oxygen donor ligands, Figure 2, were selected for study based on several criteria. The availability of accurate aqueous stability constant data for the formation of uranyl complexes was determined by examination of Smith and Martell's compilation of Critical Stability Constants series[2,4] and Grenthe's treatise on the chemical thermodynamics of uranium.[20] Data at 25 °C were selected to cover as wide a range of log K_1 values as possible for rigid ligands with structures capable of exhibiting only one conformation either in the free state or when bound to the uranyl cation. This was done to minimize potential contributions arising from the presence of multiple conformational states, either in the bound or free form of the ligand, and to eliminate uncertainty in the selection of initial coordinates for geometry optimizations. In cases where log K_1 values were not available at zero ionic strength, they were corrected to zero ionic strength using the Davies equation.[21] The experimental log K_1 values, listed in Table 1, span a range of nearly 17 orders of magnitude, ranging from a low of 0.0 for nitrate to a high of 16.8 for catecholate.

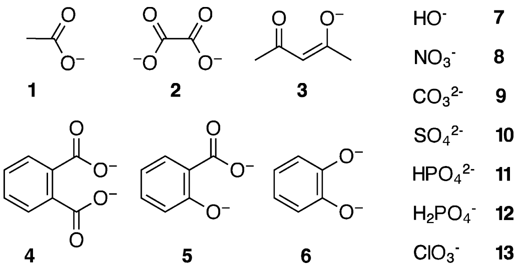

Figure 2. Negative oxygen donor ligands.

Table 1. Comparison of Experimental, Calculated, and Predicted log K_1 Values

ligand	exptl, aq[a] log K_1	calcd, g[e] log K_1	calcd, aq[f] log K_1	abs. error	pred, aq[g] log K_1	abs. error
1 acetate	3.1[b]	154.5	20.6	17.5	4.3	1.2
2 oxalate	7.3[b]	290.1	30.8	23.5	6.5	0.8
3 acetylacetonate	7.7[b]	163.6	30.1	22.4	7.3	0.4
4 phthalate	5.6[b]	276.8	28.5	22.9	5.4	0.2
5 salicylate	13.0[b]	293.1	42.6	29.6	12.0	1.0
6 catecholate	16.8[b]	309.5	51.1	34.3	16.1	0.7
7 OH^-	8.8[c]	180.8	34.7	25.9	8.7	0.1
8 NO_3^-	0.0[c]	133.0	7.5	7.5	0.3	0.3
9 CO_3^{2-}	9.7[c]	321.2	41.5	31.8	11.6	1.9
10 SO_4^{2-}	3.0[c]	287.8	23.1	20.1	2.8	0.2
11 HPO_4^{2-}	7.2[c]	301.1	34.5	27.3	8.2	1.0
12 $H_2PO_4^-$	3.3[b]	145.5	15.9	12.7	2.9	0.4
13 ClO_3^-	0.5[d]	133.8	6.5	6.0	0.0	0.5

[a] Corrected to zero ionic strength with the Davies equation.[21] [b] Reference 2. [c] Reference 19. [d] Reference 4. [e] Calculated from ΔG_g. [f] Calculated from ΔG_{aq}. [g] Predicted from correlations shown in Figure 3.

Following the approach illustrated in Figure 1, ΔG_{aq} values were calculated for the equilibrium shown in eq 7,

$$[UO_2(OH_2)_5]^{2+} + L^{x-} \rightleftharpoons [UO_2L(OH_2)_{5-n}]^{2-x} + nH_2O \tag{7}$$

These calculations require initial atomic coordinates for each species. Given that the geometry of the predominant solvated uranyl ion in aqueous solution, $[UO_2(H_2O)_5]^{2+}$, is known[8,22] and the anionic ligands, L^{x-}, have only one populated conformation, a possible ambiguity regarding the geometry of species present in eq 7 occurs only with the $[UO_2L(H_2O)_{5-n}]^{2-x}$ complex. There are two issues: (i) whether L^{x-} interacts in a unidentate or bidentate fashion and (ii) the number of water molecules that are displaced when the ligand coordinates the metal. These issues were addressed computationally.

With the exception of hydroxide, **7**, it was possible to locate minima for all cases for a bidentate form that displaced two water molecules. In some cases, it was also possible to locate a minimum corresponding to a unidentate form that displaced only one water molecule. At the B3LYP/SSC/6-311++G**

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level of theory with solvent effects included using the IEF-PCM model, the bidentate complexes with the uranyl ion were the most stable, with the exception of complexes with $H_2PO_4^-$ and ClO_3^- ligands that showed a slight energetic preference for monodentate over bidentate coordination. The more stable form for each ligand was selected for the ΔG_{aq} calculations.

Similarly, with respect to the number of water molecules bound to the uranyl ion, calculations confirmed that for the set of ligands shown in Figure 2 the most stable arrangement in the gas phase was obtained when there were a total of five oxygen atoms occupying the equatorial coordination sites. In other words, four water molecules with hydroxide and three water molecules with all other ligands. This finding is consistent with crystal structure data that establishes five to be the most observed equatorial coordination number for the uranyl cation.[23] Although there are many examples of uranyl complexes in which six oxygen atoms are coordinated in the equatorial positions, almost without exception these complexes involve two or more four-membered chelate rings, for example, $[UO_2(NO_3)_2(OH_2)_2]^0$ or $[UO_2(AcO)_3]^-$. In these cases, the higher coordination can be attributed to the decreased steric crowding resulting from the small bite distances associated with four-membered chelate rings.

The reaction free energies for eq 7 were evaluated for **1−13** in both the gas phase and aqueous solution by application of the IEF-PCM model to compute ΔG^*_{solv} terms. The results were converted to log K_1 values (eq 2), which are summarized in Table 1. As expected for complexes that are formed by the combination of oppositely charged species, the calculated gas phase log K_1 values in the absence of solvent correction overestimate the experimental values by >100 orders of magnitude. A plot of experimental log K_1 versus calculated gas phase log K_1 (Figure 3, top) shows two groupings of data, one for monovalent ligands and one for divalent ligands. On fitting lines through each group of data, it can be seen that the overestimation of calculated gas phase values for divalent ligands is approximately double that of monovalent ligands, reflecting the dominant influence of electrostatics on ΔG°_g.

Including solvation effects via the IEF-PCM solvation model greatly eliminates the difference in the calculated log K_1 values for two groups of monovalent and divalent ligands (Figure 3, bottom). While the order of ligand stability is correctly determined (even across each group), the calculated log K_1 values are significantly overestimated, with an absolute error ranging from 7.5 to 34.3 kcal/mol (Table 1). These results are consistent with previous studies[6−12] indicating that accurate computation of absolute complexation free energies for reactions involving multicharged ions is a difficult task. In addition to deficiencies in the applied density functional theory and the associated basis set, the results of calculations are highly sensitive to the choice of a cluster and solvation model. A part of the problem is that continuum dielectric models are often not adequate when dealing with solutes that have concentrated charge densities with strong local solute−solvent interactions. As a result, electronic structure calculations for ions within the context of pure dielectric continuum models are often prone to make large errors in the hydration free energies.[24]

Using the default IEF-PCM solvation model in Gaussian optimized for neutral solutes leads to a underestimation of $\Delta G^*_{solv}(OH^-)$ and $\Delta G^*_{solv}(CH_3COO^-)$ by 21.3 and 12.2 kcal/mol, respectively, compared with the recommended "experimental" values obtained using the cluster pair approximation.[25] The deviation from experiment for dianionic ligands is expected to be even more significant. Thus, if the stability of $[UO_2(OH_2)_5]^{2+}$ and L^{x-} (x = 1,2) species on the left-hand side of the equilibrium 7 is underestimated, then the log K_1 values calculated from this equilibrium will be overestimated.

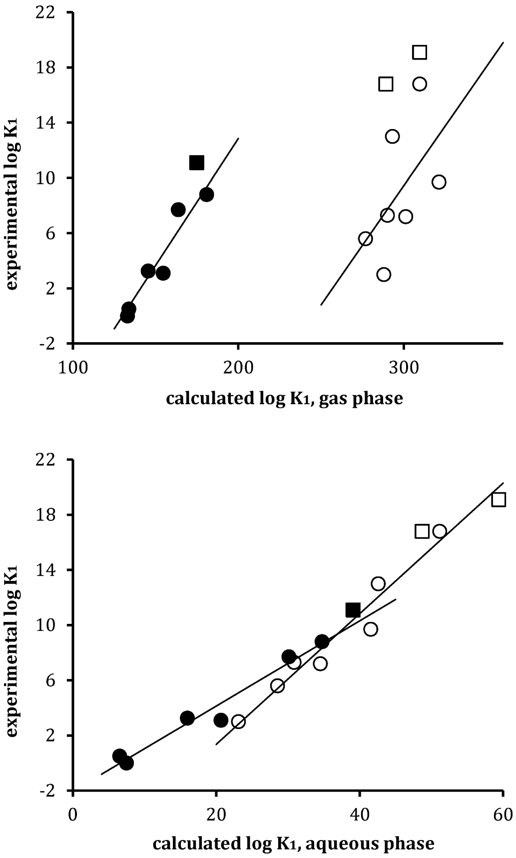


Figure 3. Plots of experimental log K_1 versus calculated log K_1 in gas phase (top) and aqueous solution (bottom) for negative oxygen donor ligands **1−13** (● for monoanions and ○ for dianions, see Table 1) and oxime ligands **14−16** (■ for monoanions and □ for dianions, see Table 2). Linear correlations in the bottom plot are for negative oxygen donors **1−13**: exptl log K_1 = 0.309 × calcd log K_1 − 2.047, with R^2 = 0.967 (monovalent ligands) and exptl log K_1 = 0.473 × calcd log K_1 − 8.108, with R^2 = 0.950 (divalent ligands). The data for oxime ligands **14−16** were not used to derive the correlations.

As follows from Figure 3 (bottom), the "raw" log K_1 values calculated directly from eqs 2 and 3 show a significant correlation (R^2 = 0.926) with the experimental data. However, the agreement with experiment can be significantly improved by a separate fitting of experimental data for L^- and L^{2-}

$$pK_1'(L^-) = -2.05 + 0.309 \times pK_1(L^-), \quad R^2 = 0.967 \tag{8}$$

$$pK_1'(L^{2-}) = -8.11 + 0.473 \times pK_1(L^{2-}), \quad R^2 = 0.950 \tag{9}$$

This is not surprising given the vast difference in the magnitude of the average solvation correction for reaction 7 with monovalent (174.8 kcal/mol) and divalent (349.6 kcal/mol) ligands. Regressions 8 and 9 account for the deficiency of theoretical models in calculating log K_1 values for complexation reactions with a change of the sum of absolute formal charges on reactants and products of $\Delta|q|$ = 2 and $\Delta|q|$ = 4, respectively. For log K_1 values spanning a range of nearly 17 log units, the root-mean-square deviation (RMSD) from experiment is 0.82

log units, which provides a computationally viable and accurate method to predict the absolute values of 1:1 stability constants for uranyl complexes.

However, this approach is not without limitations when it is applied for divalent ligands with well-separated charges. Indeed, in the limit of a large separation between the two anionic groups, the solvation free energy of a divalent ligand approaches the sum of the solvation free energies of the two subsystems. Consequently, the differential effect of solvation for complexation of a divalent ligand with largely separated functional groups can be treated as if it were the sequential binding of the two separate subsystems. Since each step involves a change of $\Delta|q| = 2$ in the sum of the absolute charges between reactants and products, log K_1 values for divalent ions with well-separated charges should be better described by eq 8. For divalent anions with linkers of intermediate length, the stability constants are expected to be intermediate between the lower- and upper-bound estimates provided by eqs 8 and 9.

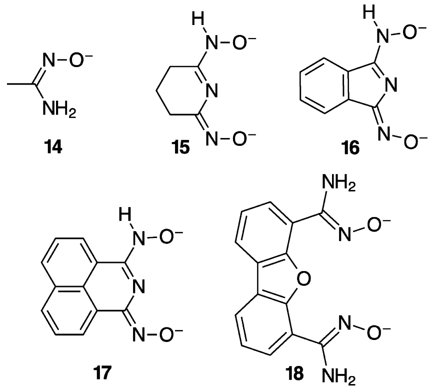
Additionally, we tested the ability of M06 and M06-L density functionals[26] with the SSC/6-311++G** basis set and the IEF-PCM and SMD solvation models to predict the log K_1 values for the initial set of 13 negative oxygen donor ligands. The results, Table S3 of Supporting Information, indicate that the B3LYP/SSC/6-311++G** method combined with the IEF-PCM solvent model provides the best overall performance, with an RMSD between the calculated and experimental log K_1 values of 0.82 log units. The M06/SSC/6-311++G** method combined with the SMD solvent model gives the second best performance (an RMSD of 0.88 log units), while the M06-L density functional yields less satisfactory results (an RMSD of 1.31−1.36 log units).

Having developed the log K_1 calculation method that achieves an accuracy of ±1.0 log units, we can now begin to examine the binding affinity to UO_2^{2+} for a set of ligands that could be potentially used to selectively extract uranium from seawater. Oxime-based ligands are of interest because they are among very few ligands able to selectively bind UO_2^{2+} in the presence of a large excess of competing metal ions under seawater conditions.[13] DFT calculations have been applied to investigate the interaction between amidoximate ligands and the uranyl ion;[13a,c,e−g] however, a comparison of the predicted and experimental stability constants has not been made. The results of calculations for amidoxime (**14**, **18**) and imide dioxime (**15**−**17**) ligands are summarized in Table 2. Experimental log K_1 data for **14**−**16** have been reported.[13e,27,28]

We note that experimental log K_1 values depend on the accurate measurement of the ligand pK_a's. However, a traditional potentiometric titration cannot provide a reliable measurement of the pK_a values over 12. Because of this limitation, there is considerable discrepancy in the reported pK_a values for **14**, ranging from 10.6 to 13.3,[27b,29] which has been resolved only recently.[13h] Given the strongly basic nature of the oxime functional group, it is to be expected that the uncertainty of log K_1 values for **14**−**16** will be larger than that for **1**−**13**.

In our initial study, we identified η^2 binding with the N−O oximate bond to uranyl as the energetically preferred coordination mode for amidoximate ligands.[13,30] Present calculations also support the η^2 coordination of two amidoximate groups linked by a dibenzofuran moiety (**18**). Consistent with previous calculations[13,31] and X-ray diffraction data,[31] imide dioxime ligands **15**−**17** show tridentate binding with uranyl ions through two oxime oxygen atoms and a ring nitrogen atom.

Table 2. log K_1 Values for UO_2^{2+} Complexes with Oxime Ligands



ligand	exptl, aq[a] log K_1	pred, aq[b] log K_1	abs. error
14	11.1[c]	10.0	1.1
15	19.1[d]	20.0	0.9
16	16.8[e]	14.9	1.9
17		17.1	
18		19.9 ± 2.2[f]	

[a] Corrected to zero ionic strength with the Davies equation.[21] [b] Predicted from regression lines given by eqs 8 and 9 and shown in Figure 3. [c] Reference 27. [d] Reference 28. [e] Reference 13e. [f] Due to significant charge separation, predicted as an average of regression lines given by eqs 8 and 9.

The number of displaced water molecules is controlled by the denticity of **14**−**18**. Calculation indicated that when one additional water molecule was placed in the inner-sphere equatorial plane of $[UO_2(Ligand)(H_2O)_2]$ (Ligand = **15**−**18**), either the complex was unstable, or the binding to uranyl was significantly less favorable compared with the hydrogen-bonding interaction with the negatively charged oxime O donor atom in the second coordination shell. For example, the uranyl complex with **18** exhibits a compact η^2 coordination of the two N−O bonds and can attain the coordination number six by having two water molecules in the equatorial plane, but the calculated log K_1 value is lower than that obtained by having one water molecule in the equatorial plane.

We note that dianionic forms of **15**−**17** can be obtained by deprotonating either oxime O or central imide N atom. The relative stability of possible tautomeric forms calculated in the field of continuum solvents is given in Figure S1 of Supporting Information. In agreement with the previous study,[13e] the results indicate that the central imide N remains protonated in the free base, while one of the oxime nitrogen atoms is protonated in the metal-ion complex. Log K_1 calculations are reported using the lowest-energy states for the free and bound forms in solution.

The calculated log K_1 values for **14**−**18** are listed in Table 2. As evident from Figure 3, the plotted points (marked as empty and filled squares for mono- and divalent ligands, respectively) lie on straight lines corresponding to the least-squares fittings of the two groups of mono- and divalent negative oxygen donor ligands. The results establish that correlations obtained for oxoacid ligands **1**−**13** show their utility for predicting the absolute log K_1 values for amidoxime **14**, coordinating in an η^2 fashion and imide dioximes **15** and **16**, involving bonding to the imide N. Thus, our approach can be used as a screening tool to provide good estimates of log K_1 values for new ligands before they are synthesized and tested in the laboratory.

In the quest for new ligands with high binding affinity and selectivity toward uranium in natural seawater, two new candidate ligands for UO_2^{2+} are examined: 1,8-naphthalimide dioxime, **17**, and dibenzofuran bis-amidoxime, **18**. During the preparation of this manuscript, ligand **17** and its corresponding uranyl complex have been synthesized and characterized spectroscopically by Bernstein et al.[31] A tridentate binding motif to uranyl has been confirmed by single-crystal X-ray diffraction, but no binding constants have been measured. Our calculations predict **17** to have UO_2^{2+} binding affinity ($\log K_1$ = 17.1 (see Table 2)) intermediate between **15** and **16**, with the former imide dioxime being one of the most strongly chelating ligands.[28] An analog of **18**, in which the dibenzofuran linkage is replaced by a fluorene linkage, was identified in our previous work[13c] as one of the top bis-amidoxime architectures for uranyl ion complexation. Calculations verify that **18** shows one of the highest binding affinities for the uranyl ion ($\log K_1 = 19.9 \pm 2.2$) reported for amidoxime-based ligands. Efforts on synthesis, measurements of the complexation properties, and incorporation of **18** into a polymer fiber are underway in our laboratory.

SUMMARY

The design of new host architectures within a polymer fiber that recognize and selectively extract uranium under seawater conditions plays a critical role in increasing sorption capacity of existing adsorbents. A key step in predicting ligand efficiency at sequestering uranium is the ability to predict accurately the log K_1 values for the uranyl and competing metal ions. In this work, we employed density functional theory (B3LYP) using the Stuttgart small-core relativistic effective core potential for uranium and the IEF-PCM solvation model to calculate the log K_1 values for 18 negative oxygen and oxime-based donor ligands. Theoretical calculations correctly determine the order of ligand stability but significantly overestimate the absolute log K_1 values. Accurate account of the ion solvation free energy poses a formidable challenge to the quantitative prediction of log K_1 values. We presented a computational protocol that can achieve good accuracy (below 1 log unit) by employing linear least-squares fitting of the calculated log K_1 values to the experimental data for 13 negative oxygen donor atoms. Separate fitting of mono- and divalent ligands is required to ensure accurate results for a wide range of log K values (from 0 to 16.8). The results verify that correlations obtained for oxoacids can provide accurate results for amidoxime and imide dioxime ligands. This extends the utility of our approach to predict the log K_1 values and screen for new amidoxime-based ligands with strong uranyl ion binding. Calculations suggest that **18** is a promising candidate exhibiting one of the highest binding affinities for the uranyl ion.

ASSOCIATED CONTENT

Ⓢ Supporting Information

Complete refs 14 and 15a, sample Gaussian09 input file, performance of the B3LYP, M06, and M06-L density functionals combined with the IEF-PCM and SMD solvation models in reproducing the log K_1 values for negative oxygen donor ligands, the relative conformational stability of **15**−**18** in the free form calculated in the field of continuum solvent, and Cartesian coordinates of ligands and metal−ligand complexes accompanied by electronic and Gibbs free energies of all species obtained with the B3LYP functional. This material is available free of charge via the Internet at http://pubs.acs.org.

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Notes

The authors declare no competing financial interest.

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