Inorg. Chem. **1981**, *20*, 3279–3283

Contribution from the Materials and Molecular Research Division,
Lawrence Berkeley Laboratory, University of California, Berkeley, California 94720

Electronic Solution Spectra for Uranium and Neptunium in Oxidation States III to VI in Anhydrous Hydrogen Fluoride

M. BALUKA, N. EDELSTEIN,* and T. A. O'DONNELL*

Received October 14, 1980

Spectra have been recorded for solutions in anhydrous hydrogen fluoride (AHF) of uranium and neptunium in oxidation states III to VI. The spectra for U(III), Np(III), and Np(IV) in AHF are very similar to those in acidified aqueous solution, but that for U(IV) suggests that the cationic species is UF_2^{2+}. The AHF spectra for the elements in oxidation states V and VI are not comparable with those of the formally analogous aqueous solutions, where the elements exist as well-defined dioxo cations. However, the AHF spectra can be related to spectra in the gas phase, in the solid state, or in nonaqueous solvents for each element in its appropriate oxidation state.

Introduction

Gillespie[1] and others have shown that superacid media are particularly effective in producing stable solutions of relatively simple cationic species, frequently with charges much higher than those which can exist in water. Thus cations such as S_8^{2+}, Se_4^{2+}, and I_2^+ exist in superacid media, whereas related species in aqueous solution are oxo or hydroxo acids or anions. Anhydrous hydrogen fluoride (AHF) is an extremely acidic solvent, and its acidity is increased markedly by addition of simple Lewis acids such as BF_3, AsF_5, or SbF_5. In this solvent, solvation of cations and fluoro complexation can occur, but there is no corresponding process to oxo complexation, which accounts for much of the complex speciation in aqueous solutions. Furthermore, most anions are solvolyzed or protonated in AHF. The only anions other than fluoride which do exist in AHF are those derived from the Lewis acids of the systems, e.g., BF_4^-, AsF_6^-, and SbF_6^-, and it is shown spectroscopically in this work that they appear to have no observable complexing effect on cations in AHF solution. There is some experimental evidence that fluoro complexation of cations in AHF is not as strong as hydroxo complexation in water, and the summation of all of these considerations has led to the postulation that intrinsically AHF should provide simpler cationic speciation than does water.[2]

Until fairly recently, it has not been possible to investigate spectral or other properties in AHF solution of d or f transition elements over a wide range of oxidation states. As indicated above, most anions are solvolyzed with fluoride, and the fluorides of transition-metals in low oxidation states are sparingly soluble in AHF. It is now possible to circumvent the problem of the insolubility of low-oxidation-state fluorides of heavy metals and to study their cations in solution. The difluorides, NiF_2 and CoF_2, when suspended in AHF and exposed to about 2 atm of BF_3 (a relatively weak Lewis acid in HF), dissolved within minutes to give concentrated solutions (up to 1 M), and electronic spectra were recorded.[3] These spectra indicate that the solutions contain cations, solvated by HF molecules with the same coordination geometry as in water. It appears that fluoride ions in the insoluble fluorine-bridged polymeric binary fluorides have been converted to BF_4^- anions, which are then in solution as the counterions to the solvated cations.

Similar treatment of PrF_3 and NdF_3 with HF and BF_3 after many hours gave much less concentrated solutions (about 5 $\times 10^{-2}$ M), whose spectra[3] indicated formation of simple HF-solvated cations, with BF_4^- counterions. Presumably, as the charge on the cation increases, the polymeric fluoride lattice, now more favored energetically, is harder to disrupt to produce solvated ions.

In the preliminary stages of the present work, application of BF_3 at 2 atm to a suspension of UF_4 in HF produced no discernible amounts of uranium(IV) in solution. The interactions of BF_3 in HF with di-, tri-, and tetrafluorides are consistent with the action of a relatively weak Lewis acid. Obviously, stronger Lewis acids such as SbF_5 and AsF_5 would be more likely to produce HF solutions of U(IV). These acids had not been used in the initial study[3] because they, unlike BF_3, are potential oxidants in HF and it was not known at that time whether they might oxidize Ni(II) or Co(II).

For dissolution of UF_4, AsF_5 was used initially as the Lewis acid and was maintained at about 2 atm over a stirred suspension of UF_4 in the HF to provide over several days a green solution, the spectrum of which was recorded and is shown in Figure 1b. SbF_5 was also used subsequently as a Lewis acid to produce U(IV) solutions in AHF, and identical spectra were recorded.

Since UF_4 dissolved very slowly in HF–AsF_5 to give a solution only about 5×10^{-2} M in U(IV), it was decided to treat metallic uranium in the form of turnings with HF–AsF_5 in an attempt to achieve more rapid solution. This solution did color quite rapidly, being green in reflected light and red in transmitted light, the latter color indicating the possible presence of some U(III) in solution. This observation is in contrast to the situation in H_2O since aqueous solutions of U(III) require complexation by ligands such as chloride for reasonable stability, and even those solutions gradually reduce water and form U(IV).[4] The next step was to treat metallic uranium in HF and BF_3 at 2 atm. Within minutes the solution became pink and, after overnight stirring, was a strong magenta color and was about 0.1 M in uranium. The UV–visible spectrum of this solution is shown in Figure 1a.

UF_5 suspended in AHF was treated with the two Lewis acids AsF_5 and SbF_5 to produce U(V) solutions. UF_6 is relatively soluble in AHF and spectra were recorded of the simple solution and of solutions containing AsF_5 and SbF_5 (see Figure 1c).

Solutions containing neptunium in oxidation states III–VI in the presence of the different Lewis acids were prepared essentially as described above for the uranium systems. As the acidity of the HF solutions was increased by addition of Lewis acids of increasing strength, the stability of solutions of the two hexafluorides decreased. The spectrum for a so-

---

* To whom correspondence should be addressed: N.E., University of California; T.A.O., Department of Chemistry, University of Melbourne, Melbourne, Australia.

(1) Gillespie, R. J. *Acc. Chem. Res.* **1968**, *1*, 202.
(2) O'Donnell, T. A. *J. Fluorine Chem.* **1978**, *11*, 467.
(3) Barraclough, C. G.; Cockman, R. W.; O'Donnell, T. A. *Inorg. Chem.* **1977**, *16*, 673.
(4) Katz, J. J.; Seaborg, G. T. "The Chemistry of the Actinide Elements"; Wiley: New York, 1957; p 174.

0020-1669/81/1320-3279$01.25/0

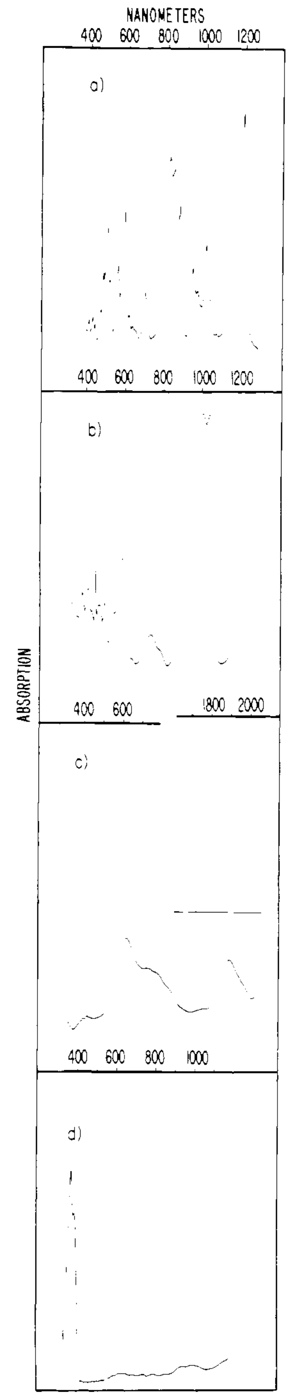

Figure 1. Optical spectra of (a) U(III)/AHF–AsF_5 (2 atm), (b) U(IV)/AHF–AsF_5 (2 atm), (c) UF_5/AHF–AsF_5 (1.5 atm), and (d) UF_6/AHF–AsF_5 (1.5 atm).

lution of Np(VI) in HF–SbF_5 on standing for many hours changed progressively to that for Np(V), and similar but less extensive reduction occurred in the U(VI) system. There is now evidence[5] to suggest that strongly protonating media leach unsaturated low-molecular-weight fractions from Kel-F to provide reducing species in AHF solutions.

The work reported here allows a comprehesive comparison of the spectra of simple species in AHF of uranium and neptunium in oxidation states III–VI. It is not possible to do so in aqueous solution due to the formation of neptunyl and uranyl ions for oxidation states V and VI. Spectroscopic evidence has been obtained to show that complexation of cations in AHF by the fluoro anions BF_4^-, AsF_6^-, and SbF_6^- is insignificant.

Experimental Section[6]

Preparation of Uranium and Neptunium Solutions. Oxidation State III. Treatment with HF–BF_3 of metallic uranium in the form of turnings, which had extensive surface oxidation, led to the formation of large amounts of U(IV) as insoluble UF_4. Consequently a small cube (about 0.5 g) of the metal and triply distilled HF (10–15 ml) were added to a Kel-F trap fitted with a Kel-F valve and provided with a Teflon-coated magnetic stirrer. BF_3, which has only small solubility in HF at room temperature, was applied until the total pressure in the trap was 3 atm, i.e., the vapor pressures of HF and BF_3 above the liquid were approximately 1 and 2 atm, respectively. Within minutes, a pink color developed in the solution and this became a strong magenta color overnight. There was some apparent formation of UF_4. This is consistent with surface oxide on the metal providing a source of U(IV), which cannot be held in solution by the weak Lewis acidity of BF_3 and yields a precipitate of UF_4. Therefore, metallic uranium was treated for a short time with HF and BF_3 or AsF_5. This procedure removed the surface contaminants and some metal, and the resulting solution was poured off. There was evidence in these preliminary steps that U(III) solutions in HF reacted with the surface of relatively new Kel-F traps. The remaining clean metal was treated with HF and BF_3 and the solution filtered into a Kel-F trap, which had been treated previously, first with ClF_3 and, after careful removal of ClF_3, with HF and H_2. Under these conditions magenta solutions of U(III) in HF were obtained which were stable indefinitely. Weighing of the trap and residue after removal of the HF and BF_3 indicated that the solution was approximately 0.1 M in U(III). AsF_5 at a partial pressure of 2 atm also gave a solution containing U(III).

Similar procedures produced stable solutions of neptunium(III). So that the possible complexing effects of BF_4^-, AsF_6^-, and SbF_6^- in HF solution could be investigated spectroscopically, the Lewis acids AsF_5 and SbF_5 were used to produce solutions of Np(III). AsF_5 was added progressively until the partial pressure of AsF_5 in the system was 2 atm, i.e., a solution approximately 5 M in AsF_6 was formed (see below). A concentrated solution (10 M) of SbF_5 in HF was also used to give a Np(III) solution.

Oxidation State IV. AsF_5 was applied to a suspension of UF_4 in a Kel-F tube, as above, until the total pressure became steady at 3 atm. The Kel-F trap was evacuated and weighed initially and then was weighed after each addition of HF and AsF_5. Weight increments indicated that the solution was approximately 5 M in AsF_5 (mostly as AsF_6^-) on equilibration.

After several days of stirring much green solid remained, but the pale green supernatant solution was filtered through Teflon filter paper (Millipore Corp.) held in a Kel-F filter-cartridge into another Kel-F trap. Weighing of the residue in this second trap after removal of HF and AsF_5 indicated that the solution was about 5×10^{-2} M in U(IV).

A green solution containing U(IV) was also produced with SbF_5 at a concentration of 3 M.

Np(IV) solutions were prepared from NpF_4 by using AsF_5 at a partial pressure of 2 atm and SbF_5 (10 M in solution).

Oxidation State V. UF_5 and NpF_5 were prepared by reduction of UF_6 and NpF_6 with PF_3 at −78 °C[7,8]. The solids, suspended in stirred HF, were treated with AsF_5 at a partial pressure of 1.5 atm and with SbF_5 at concentrations of 0.4, 1, and 10 M.

Oxidation State VI. The hexafluorides are readily soluble in AHF at room temperature.[9] For comparisons with those uranium fluorides which needed Lewis acids to cause dissolution, the UF_6 solution was treated with AsF_5 (1.5 atm) and SbF_5 (10 M).

NpF_6 was prepared by the simplified method reported for the preparation of RuF_6 and PuF_6.[10] A solution about 0.15 M in the hexafluoride was prepared and treated with BF_3 (2 atm), AsF_5 (2 atm), and SbF_5 (10 M).

---

(5) Cockman, R. W.; O'Donnell, T. A., unpublished observations.

(6) For a review of the properties of fluorine and fluorides and the techniques for handling these materials, see: O'Donnell, T. A., "Comprehensive Inorganic Chemistry"; Pergamon Press: Oxford, 1973; Vol. 2, pp 1009–1106 and the references therein.

(7) O'Donnell, T. A.; Rietz, R.; Yeh, S., unpublished observations.

(8) Baluka, M.; Yeh, S.; Banks, R.; Edelstein, N. *Inorg. Nucl. Chem. Lett.* **1980**, *16*, 75.

(9) Frlec, B.; Hyman, H. H.; *Inorg. Chem.* **1967**, *6*, 1596.

(10) Burns, R. C.; O'Donnell, T. A. *J. Inorg. Nucl. Chem.* **1980**, *42*, 1613.

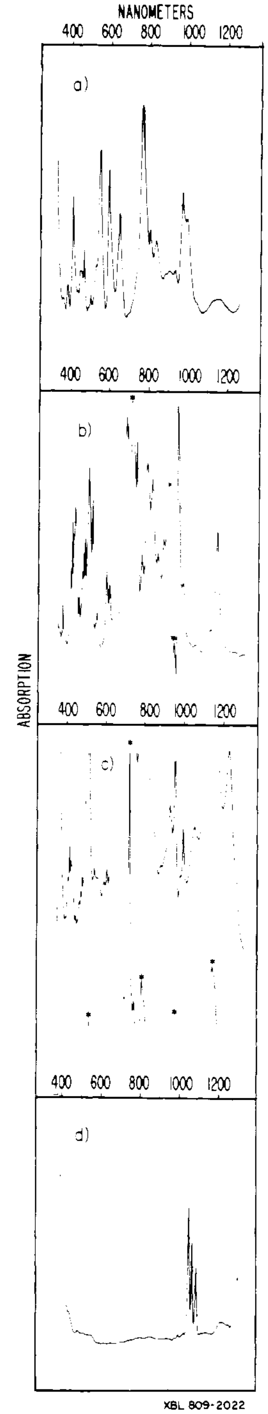

Figure 2. Optical spectra of (a) Np(III)/AHF–AsF_5 (2 atm), (b) Np(IV)/AHF–AsF_5 (2 atm), (c) NpF_5/AHF–AsF_5 (2 atm), (d) NpF_6/AHF–BF_3 (2 atm). The asterisk above certain peaks represents absorption occurring on the second absorption scale (OD between 1 and 2) of the Cary 17.

Measurements. The cells and procedures were essentially as reported earlier,[3] and a Cary 17 spectrophotometer was used for recording spectra.

Results and Discussion

Spectra of aqueous solutions are available for comparison with the corresponding ones in AHF for the simple aquo cations of U(III) and U(IV)[11] and for Np(III) and Np(IV)[12] in strongly acid noncomplexing solutions, and these have been compared with the AHF spectra (Figures 1 and 2) of the same elements in oxidation states III and IV.

Table I. Positions of Major Peaks in Spectra of U(III) in H_2O and AHF (nm)

U^{3+} in 1 M DCl_4-D_2O[11]	U^{3+} in AsF_5-HF
1219	1198
1064	1010
	967
967	942
898	876
876	844
	820
	700
636	591
615	558
	520
522	503
507	482
	477
	442
451	429
	411

Table II. Positions of Major Peaks in Spectra of Np(III) in H_2O and AHF (nm)

Np^{3+} in 2 M $HClO_4$-H_2O[12]	Np^{3+} in AsF_5-HF
1363	1338
991	966
849	829
786	760
660	643
600	586
552	544

Cohen and Carnall[11] recorded the aqueous spectrum for U(III) after reduction of U(VI) to U(III) in 1.0 M $DClO_4$ in a spectral cell. As had been found with the spectra of tripositive lanthanides in AHF,[3] similarities between the aqueous and AHF spectra for U(III) are very marked. The major difference in this case is that the relative intensities of the two principal bands in each solvent are reversed. The positions of the bands for each solvent are given in Table I. Peak positions are shifted by about 20–40 nm to higher energies in passing from H_2O to AHF. The shifts to higher energies in moving from aqueous to AHF solutions are smaller for lanthanides (3–5 nm), but those for d-transition elements are very much larger (50–100 nm) and are to lower energies.[3] The aqueous and AHF spectra are very similar to those recorded for U(III) in formic acid.[13,14]

The spectra for Np(III) and Np(IV) in AHF (Figure 2) are very similar to those for Np(III) and Np(IV) in H_2O.[12] Again, there are relatively small shifts in peak positions to higher energies in AHF (see Table II), although these are smaller for Np(IV) than for Np(III).

Although the AHF spectra for U(III), Np(III), and Np(IV) are very similar to the aqueous counterparts, the U(IV) spectrum in HF–AsF_5 is different from the spectrum recorded for U(IV) in 1 M $DClO_4$–D_2O by Cohen and Carnall,[11] the shifts in positions of major peaks being 100 nm or greater. Their spectrum was reproduced in $HClO_4$ medium by Stein and co-workers,[15] who then added aqueous HF progressively to their 0.03 M U(IV) solution until the HF concentration was 0.080 M, observing progressive changes in the spectra. They

---

[11] Cohen, D.; Carnall, W. T. *J. Phys. Chem.* **1960**, *64*, 1933.

[12] Hagan, P. C.; Cleveland, J. M. *J. Inorg. Nucl. Chem.* **1966**, *28*, 2905.

[13] Drozkzynski, J. *J. Mol. Struct.* **1973**, *19*, 465.

[14] After this work was completed the following report was found which states that treatment of U metal with HF and SbF_5 yielded a solid residual which was $U(SbF_6)_3$: Tarvero, M. Rapport CEA-R3205, 1967, p 51.

[15] Stein, L.; Williams, C. W.; Fox, I. M.; Gebert, E. *Inorg. Chem.* **1966**, *5*, 662.

Table III. Positions of Vibrational Fine Structure in U(IV) Solutions (nm)

U^{4+} in AsF_5-HF	0.03 M U^{4+} in 1 M $HClO_4$ and 0.08 M HF^a
390	405
420	426
439	452
458	469
478	486
	495
506	525
560	590
600	610
650	626
	642

[a] From ref 15.

postulated the very marked shifts in peak positions and the development of vibrational fine structure in the range 400–650 nm indicated the formation of UF_2^{2+} as fluoride was added. They also showed that, as fluoride was added progressively to a 1 M H_2SO_4 solution containing U(IV), similar spectral characteristics developed before the precipitation of a compound, which was amorphous but which had a composition close to $UF_2SO_4{\cdot}2H_2O$. They ultimately crystallized $UF_2SO_4{\cdot}2H_2O$ from a solution initially 6 M in H_2SO_4 and 0.9 M in HF.

Figure 1b shows the spectrum for U(IV) in HF–AsF_5. It is remarkably similar to the spectrum marked E in Figure 2 of the paper by Stein et al., i.e., the one postulated as being due to UF_2^{2+}, suggesting that the species UF_2^{2+}, whose spectra and stabilities Stein et al. compare with those of NpO_2^+ and PuO_2^{2+}, is the dominant uranium-containing species in these HF solutions. A comparison of the vibrational fine structure for the two solutions is given in Table III. Supporting evidence for this proposal is that the involatile green residue, obtained after evaporation of free HF and AsF_5 from the U(IV) solution and pumping at room temperature, showed a weight increase over the original weight of UF_4, which corresponded with the formation of a compound having the formula $UF_2(AsF_6)_2$.

It is interesting to speculate that the stability of the isoelectronic species UF_2^{2+}, NpO_2^+, and PuO_2^{2+} could provide a rationalization of an apparent paradox in the aqueous chemistry of U, Np, and Pu in high oxidation states. It has long been recognized that Np(V), present as NpO_2^+, is very much more stable toward disproportionation than the corresponding dioxo species for U and Pu. With the assumption that two favored species for U are UO_2^{2+} and UF_2^{2+} (in aqueous solutions containing fluoride), the inclusion of NpO_2^+ in the isoelectronic series could be used as a basis for consideration of the apparently anomalous stability of Np(V) in solution. On the other hand, there is nothing in the known solution chemistry of U(IV) and Np(IV) to suggest that the speciation of the two elements in the same oxidation state should be so different in AHF.

Figures 1 and 2 also give spectra for U(V), U(VI), Np(V), and Np(VI) in AHF in the presence of appropriate Lewis acids. Not surprisingly, these spectra are quite different from the formally analogous ones reported for aqueous solutions of these elements in oxidation states V and VI in which dioxo cations are formed. The UO_2^{2+} entity is stable in AHF. When UO_2F_2 is dissolved in HF–AsF_5, the spectrum in the 340–350-nm region, although much better resolved, is very similar to that for UO_2^{2+} in aqueous $HClO_4$ solution, and a strong Raman signal, characteristic of UO_2^{2+}, is obtained for the HF–AsF_5 solution.[16] Qualitatively, the present spectra resemble appropriate spectra of related solids or spectra in the gas phase or in those nonaqueous solvents in which the dioxo cations are not formed. The spectrum of U(V) in acidic AHF is similar to those reported by Ryan[17] for U(V) in halo complexes of U(V). As expected, there is no absorbance in the visible or the near-UV region by solutions of UF_6, the only spectral feature in the UV resulting from charge transfer. Earlier conductance and Raman spectroscopic studies have shown that UF_6 dissolves in AHF as the molecular entity.[9] The AHF solution spectrum for NpF_5 resembles that for a fluorocarbon mull of solid $CsNpF_6$,[18] and that for NpF_6 in neutral HF or in HF containing the Lewis acid BF_3 is closely related to the gas-phase spectrum.[19] AsF_5, a stronger Lewis acid than BF_3, could not be used reliably in AHF solutions of NpF_6 for the reasons given below.

UF_6 and NpF_6 both have reasonable solubility in neutral HF, whereas the pentafluorides required the presence of a Lewis acid for dissolution. So that the spectra for both elements would be strictly comparable in all oxidation states, the three Lewis acids used in this study were added to both hexafluoride solutions. BF_3 had no effect on either spectrum. The UF_6 spectrum was unchanged by the addition of AsF_5, but the NpF_6 spectrum was made more complex. The addition of SbF_5, particularly in high concentration (10 M) and on standing, caused the observed spectra to be very different from those for both hexafluorides. After about 1 day the spectrum for NpF_6–SbF_5–HF had changed to being almost identical with that for NpF_5–SbF_5–HF. As mentioned earlier, strongly acidic HF can protonate unsaturated material in Kel-F (presumably low-molecular-weight fractions of the polymer) and leach these from the Kel-F container into solution.[5] This material could then reduce Np(VI) to Np(V). This proposed behavior is consistent with the spectroscopic observation in the present work that NpF_6 solutions were reduced more readily than UF_6 in SbF_5–HF solutions and that AsF_5–HF reduced Np(VI) but not U(VI). The known chemistry of actinide hexafluorides shows that NpF_6 is a stronger oxidant than UF_6.[6]

In the Experimental Section of this paper it was reported that solutions were prepared for both metals in a wide range of oxidation states by using the three Lewis acids BF_3, AsF_5, and SbF_5 with each system where possible; i.e., each of the solutions was prepared, whenever possible, containing the anions BF_4^-, AsF_6^-, or SbF_6^- or the appropriate polymeric anions. Except for the specific chemical reduction referred to above for U(VI) and Np(VI) in highly acidic solutions, no significant spectroscopic differences were observed that could be attributed to the counteranions in any series of spectra; i.e., complexation by fluoro anions is not significant, nor is there any detectable fluoro complexation of the actinide ions in moving from the fairly acidic HF–AsF_5 solutions to the weakly acidic HF–BF_3 solutions. Such effects, although small, have been observed for spectra of cations of d transition metals in AHF solutiions of differing acidity.[5]

So that the actinide spectra would be as strictly comparable as possible and effects due to introduction of impurities into HF–SbF_5 solutions could be eliminated, all spectra given in Figures 1 and 2 were taken in HF–AsF_5 except that for Np(VI), which is susceptible to reaction with the impurities leached from the Kel-F by HF–AsF_5. In this case HF–BF_3 was used.

Acknowledgment. We are grateful for financial support from the American Educational Foundation Fulbright Program (M.B.) and to the Council of the University of Mel-

---

(16) Barraclough, C. G.; Cockman, R. W.; O'Donnell, T. A. *Inorg. Nucl. Chem. Lett.* **1981**, *17*, 83.

(17) Ryan, J. L. *J. Inorg. Nucl. Chem.* **1971**, *33*, 153.

(18) Varga, L. P.; Brown, J. D.; Reisfeld, M. J.; Cowan, R. D.; *J. Chem. Phys.* **1970**, *52*, 4233.

(19) Eistenstein, J. C.; Pryce, M. H. L. *Proc. R. Soc. London, Ser. A.* **1960**, *255*, 181.

bourne for granting sabbatical leave (T.A.O.). We wish to thank R. N. R. Mulford of Los Alamos National Laboratory for providing the samples of Np metal used in this work.

This work was supported by the Division of Chemical Sciences, Office of Basic Energy Sciences, U.S. Department of Energy, under Contract No. W-7405-Eng-48.

Registry No. AHF, 7664-39-3; U(III), 22578-81-0; U(IV), 16089-60-4; U(V), 22537-60-6; U(VI), 22541-40-8; Np(III), 21377-65-1; Np(IV), 22578-82-1; Np(V), 22537-61-7; Np(VI), 22541-66-8; UF_2^{2+}, 18785-29-0.

Contribution from the Department of Chemistry,
Michigan State University, East Lansing, Michigan 48824

ESR Studies of Some Oxotetrahalo Complexes of Vanadium(IV) and Molybdenum(V)

K. K. SUNIL and M. T. ROGERS*

Received October 24, 1980

ESR spectra of $[VOF_4]^{2-}$ and $[MoOF_4]^-$ have been studied in single crystals of $(NH_4)_2SbF_5$ and spectra of $[MoOCl_4]^-$ in single crystals of $(NH_4)_2SbCl_5$. The spin-Hamiltonian parameters of these pentacoordinated complexes have been obtained and compared with those for the corresponding hexacoordinated species. Molecular orbital parameters for the penta- and hexacoordinated species obtained from experimental **g**- and **A**-tensor components have been compared with values calculated by the MS–SCF–Xα method.

Introduction

The transition-metal oxohalo complexes of the type $[MOX_5]^{m-}$, where M = V, Nb, Cr, Mo, and W and X = F, Cl, Br, and I, have been the subject of detailed ESR studies.[1-20] The **g**, metal hyperfine, and ligand hyperfine tensors have been used to investigate the nature of bonding in these compounds. On the other hand, very little work has been done on the corresponding pentacoordinated oxohalo complexes.[21,22] Since ESR spectroscopy provides a very sensitive probe for the detection and measurement of the effects of small changes in bonding, we have undertaken a fairly detailed study of the single-crystal ESR spectra of $[VOF_4]^{2-}$, $[MoOF_4]^-$, and $[MoOCl_4]^-$ in the hope of getting a better understanding of the differences in the bonding between the penta- and hexacoordinated transition-metal oxohalo complexes.

The pentacoordinated transition-metal complexes form a class of compounds of considerable interest with the structures possessing a diversity of forms between the two limiting symmetries of trigonal bipyramidal (D_{3h} symmetry) and square pyramidal (C_{4v} symmetry).[22-24] The energy barrier between these two structures is predicted[25] to be small for species with five equivalent ligands, and examples are observed to occur in both symmetry classes.[26,27] On the other hand, complexes having an axial ligand different from the other four tend to form square-pyramidal complexes.[24] The pentacoordinated oxohalo complexes under investigation here fall in the latter category.

ESR studies of VO^{2+} and Cr^{3+} in $(NH_4)_2SbCl_5$[21] and Fe^{3+} in $(NH_4)_2SbF_5$[28] have been found to provide pentacoordinated species $[VOCl_4]^{2-}$, $[CrCl_5]^{2-}$, and $[FeF_5]^{2-}$, respectively, while for Cr^{3+} in K_2SbF_5 the species which predominates[29] is trigonally-distorted $[CrF_6]^{3-}$. ESR spectra of $[MoOCl_4]^-$ have been studied in solution as well as in a diluted single crystal of $[AsPh_4][NbOCl_4]$.[32] We have carried out the ESR studies of the fluoro complexes of VO^{2+} and MoO^{3+} in ammonium

---

(1) N. S. Garif'yanov, *Dokl. Akad. Nauk SSSR*, **155**, 385 (1954).

(2) K. D. Bowers and J. Owen, *Rep. Prog. Phys.*, **18**, 304 (1955).

(3) N. S. Garif'yanov and V. N. Fedotov, *Zh. Strukt. Khim.*, 3, 711 (1962).

(4) N. F. Garif'yanov, V. N. Fedotov, and N. S. Kucheryaenko, *Izv. Akad. Nauk SSSR, Ser. Khim.*, **4**, 743 (1964); *Bull. Acad. Sci. USSR, Div. Chem. Sci.* (*Engl. Transl.*), 589 (1964).

(5) N. S. Garif'yanov, *Dokl. Akad. Nauk. SSSR*, **135**, 385 (1964).

(6) K. DeArmond, B. B. Garrett, and H. S. Gutowsky, *J. Chem. Phys.*, **42**, 1019 (1965).

(7) D. I. Gyabchikov, I. N. Marov, Y. N. Dubrov, V. K. Belyaeva, and D. N. Ermakov, *Dokl. Akad. Nauk SSSR*, **169**, 1107 (1966); *Dokl. Phys. Chem.* (*Engl. Transl.*), **169**, 795 (1966).

(8) J. T. C. Van Kemenade, J. L. Verbeek, and P. F. Cornaz, *Recl. Trav. Chim., Pays-Bas*, **85**, 629 (1966).

(9) H. Kon and N. H. Sharpless, *J. Phys. Chem.*, **70**, 105 (1966).

(10) M. M. Abraham, J. P. Abriata, M. E. Foglio, and E. Pasquini, *J. Chem. Phys.*, **45**, 2069 (1966).

(11) J. L. Verbeek and P. F. Cornaz, *Recl. Trav. Chim. Pays-Bas*, **86**, 209 (1967).

(12) J. L. Verbeek and A. T. Vink, *Recl. Trav. Chim. Pays-Bas*, **86**, 913 (1967).

(13) R. D. Downing and J. F. Gibson, *J. Chem. Soc. A*, **1**, 655 (1967).

(14) J. L. Verbeek, "Ligand Hyperfine Structure in the ESR Spectra of the Ions $[CrOF_5]^{2-}$ and $[MoOF_5]^{2-}$", Thesis, Eindhoven, 1968.

(15) P. T. Manoharan and M. T. Rogers, *J. Chem. Phys.*, **49**, 5510 (1968).

(16) P. T. Manoharan and M. T. Rogers, *J. Chem. Phys.*, **49**, 3912 (1968).

(17) J. T. C. Van Kemenade, "Ligand Hyperfine Interactions in Oxyhalides of Pentavalent Chromium, Molybdenum and Tungsten", Thesis, Eindhoven, 1970.

(18) J. R. Shock and M. T. Rogers, *J. Chem. Phys.*, **58**, 3356 (1973).

(19) L. A. Dalton, R. D. Bereman, and C. H. Brubaker, Jr., *Inorg. Chem.*, **8**, 2477 (1969).

(20) C. D. Garner, L. H. Hill, F. E. Mabbs, D. L. McFadden, and A. T. McPhail, *J. Chem. Soc., Dalton Trans.*, 853 (1977).

(21) J. M. Flowers, J. C. Hempel, W. E. Hatfield, and H. H. Dearman, *J. Chem. Phys.*, **58**, 1479 (1973).

(22) E. L. Muetterties and R. A. Schunn, *Q. Rev.*, **20**, 245 (1966).

(23) C. Furlani, *Coord. Chem. Rev.*, 3, 141 (1968).

(24) J. A. Ibers, *Annu. Rev. Phys. Chem.*, **16**, 380 (1965).

(25) J. Zemann, *Z. Anorg. Allgem. Chem.*, **324**, 241 (1963).

(26) (a) M. Mari, Y. Saito, and T. Watanabe, *Bull. Chem. Soc. Jpn.*, **34**, 295 (1961); (b) N. K. Raymond, D. V. Meek, and J. A. Ibers, *Inorg. Chem.*, **7**, 1111 (1968).

(27) (a) J. J. Alexander and H. B. Gray, *J. Am. Chem. Soc.*, **89**, 3356 (1967); (b) K. G. Caulton, *Inorg. Chem.*, **7**, 392 (1968).

(28) C. J. Radnell, J. R. Pilbrow, S. Subramanian, and M. T. Rogers, *J. Chem. Phys.*, **62**, 4948 (1975).

(29) J. C. Hempel, D. Klassen, W. E. Hatfield, and H. H. Dearman, *J. Chem. Phys.*, **58**, 1487 (1973).

0020-1669/81/1320-3283$01.25/0