Effect of Pressure on the Electronic Structure of Ferric Hydroxamates and Ferrichrome A

D. C. GRENOBLE AND H. G. DRICKAMER

School of Chemical Sciences and Materials Research Laboratory, University of Illinois, Urbana, Ill. 61801

ABSTRACT The effect of pressure up to 175 kilobars on the electronic structure of three ferric hydroxamates and on ferrichrome A has been studied by optical absorption and Mössbauer resonance. The ferric ion was reduced to ferrous ion with pressure, as has been previously observed for various compounds. For the hydroxamates, the amount of reduction correlated very well with the location and shift of the metal-to-ligand charge transfer peak. This is entirely consistent with a previously presented theory. The results for ferrichrome A did not fit quantitatively into the series. Since the shape of the potential well is almost certainly different for this compound, this result is not surprising.

Hydroxamic acids have received wide attention among analytical chemists because of their ability to form highly colored ferric coordination compounds. The chemical formula of a primary hydroxamic acid is R—CO—NH(OH), where R can represent various chemical groups. Coordination to the ferric ion is through the —CO—N(OH)— group after removal of the proton. There is renewed biochemical interest in hydroxamic acids as a result of finding this characteristic group in natural products from such organisms as aerobic microbial cells. The biological function of hydroxamates is to transfer iron through metabolic channels, and present it for incorporation into the porphyrins and other iron-containing enzymes and proteins. The biological transfer of iron is accomplished by a reduction of the ferric iron, which is bound tightly by the hydroxamic acid groups, to ferrous iron, which is only weakly bound and can easily be removed from the compound. Neilands (1–4) has written a review article and several papers concerning the chemical and biological aspects of hydroxamic acids.

In this study, three model compounds for the biological compounds and one biological hydroxamate, ferrichrome A, were studied under different pressures. The model compounds are tris(acetohydroxamato)iron(III), tris(benzohydroxamato)iron(III), and tris(salicylhydroxamato)iron(III). For brevity, these compounds are referred to as Fe(AHA)₃, Fe(BHA)₃, and Fe(SHA)₃, respectively. The iron atom in these compounds is octahedrally coordinated to six oxygen atoms (5), and thus requires three hydroxamic acid moieties for coordination to the iron. In contrast, ferrichrome A is a hexadentate iron compound. The crystal structure of this compound has also been determined (6).

Abbreviations: AHA, tris(acetohydroxamato); BHA, tris(benzohydroxamato); SHA, tris(salicylhydroxamato).

METHODS

The hydroxamates were prepared by methods available in the literature (7, 8), using iron enriched to 90% in ⁵⁷Fe, and gave appropriate analyses. The unenriched iron in the ferrichrome A was replaced by enriched iron by the method of Neilands (3). The techniques used in the high-pressure studies, Mössbauer resonance and optical absorption, have been described (9–11).

STUDIES WITH MODEL COMPOUNDS

The optical absorption of Fe(AHA)₃ in the visible region is shown in Fig. 1 as a function of pressure. The optical transition is a ligand-to-metal charge transfer. The initial charge transfer energy in Fe(AHA)₃ is at 23.25 kK(kiloKayser). The optical spectra were normalized to 1.0 at the peak maximum for each pressure. The shift to lower energy at 50% of the peak maximum is 1.19 kK (0.175 eV) in 140 kbar. Still further out on the tail, at about 25% of the peak maximum, the shift is 1.9 kK in 140 kbar. In contrast the peak maximum shows only a modest shift (less than 0.2 kK at 90% of the maximum absorption). The optical spectra of Fe(BHA)₃ and Fe(SHA)₃ are similar to Fe(AHA)₃ although the initial energies of the charge transfer bands are different. At 1 atm, the peak maxima for Fe(AHA)₃, Fe(BHA)₃, and Fe(SHA)₃ are, respectively, 23.25, 22.22, and 21.95 kK; and the energies at 50% of maximum absorption are 19.8, 18.9, and 18.55 kK, respectively. The shifts to lower energy at 50% of peak maximum in 140 kbar for Fe(AHA)₃, Fe(BHA)₃, and Fe(SHA)₃ are 1.19, 1.25, and 1.25 kK, respectively, so that the changes in optical spectra with pressure for these three compounds are very similar.
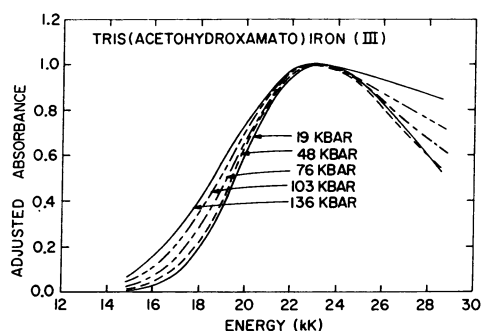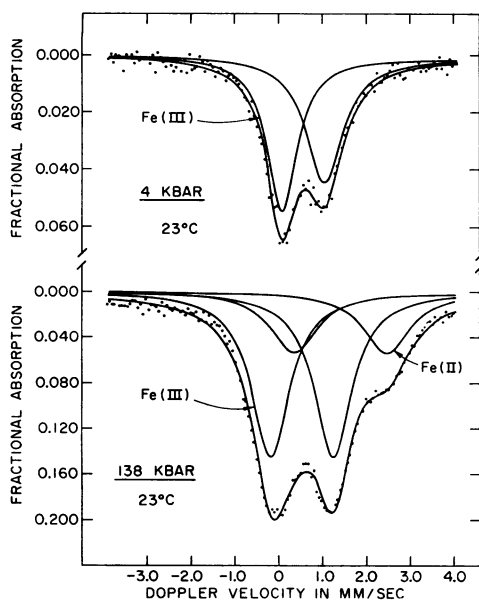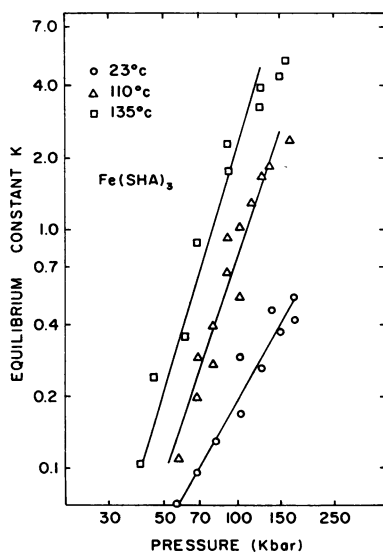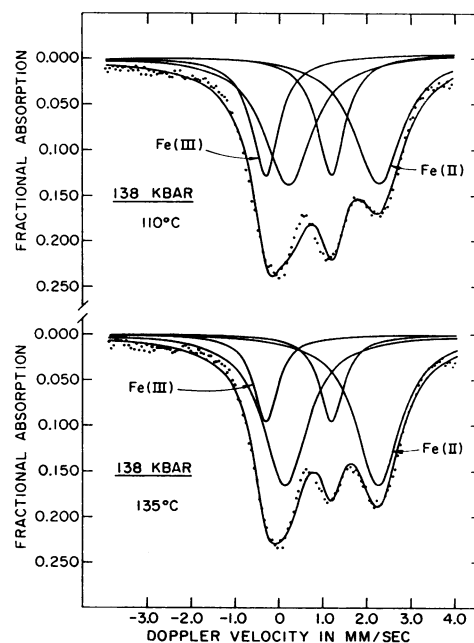

Fig. 1. Optical absorption spectra vs. pressure for Fe(AHA)₃.

Fig. 2. (a) Mössbauer spectra of Fe(SHA)₃.

Fig. 2. (b) Mössbauer spectra of Fe(SHA)₃ (contd.).

The Mössbauer spectra of Fe(SHA)$_3$ at 4 kbar and 23°C, and at 138 kbar and 23, 110, and 135°C are shown in Figs. 2a and b. At low pressure the ferric peaks are asymmetric, indicating the presence of a spin–spin relaxation effect. This is similar to the spin–spin relaxation observed in hemin, which has been treated theoretically by Blume (12). The ratio of peak widths at 4 kbar is about 1.25; by 60 kbar, the ratio has decreased to 1.05. At higher pressures the peaks appeared to be symmetric. At high pressure the ferric iron reduces to ferrous iron; with increasing temperature the conversion increases significantly. The equilibrium constant K, defined as $K = C_{II}/C_{III}$, where C_{II} and C_{III} are the fraction of ferrous and ferric sites, is shown in Fig. 3 as a function of pressure for Fe(SHA)$_3$. Fig. 4 shows a comparison of the conversions for the three hydroxamates at 23 and 135°C. Note that for all three compounds there is a marked increase in conversion

with increased temperature. The isomer shifts and quadrupole splittings of the ferric and ferrous iron in the three compounds are given in Table 1. The Fe(III) in these compounds is in the high-spin configuration, with typical values of about 0.4 mm/sec for isomer shift. The quadrupole splittings, ranging from 0.70 to 0.94 mm/sec at 4 kbar, are somewhat high for high-spin Fe(III) in octahedral symmetry, but reflect the inequivalence of the oxygen atoms coordinating to the iron, since each hydroxamic acid group contributes one negative oxygen ion and one covalent oxygen atom to the octahedral coordination sphere. The Fe(III) isomer shift decreases by about 0.035 mm/sec in 175 kbar in Fe(AHA)$_3$ and Fe(SHA)$_3$, but in Fe(BHA)$_3$ the isomer shift first increases by about 0.05 mm/sec in the first 25 kbar, and then decreases by about 0.03 mm/sec in the next 150 kbar. It should be noted, as
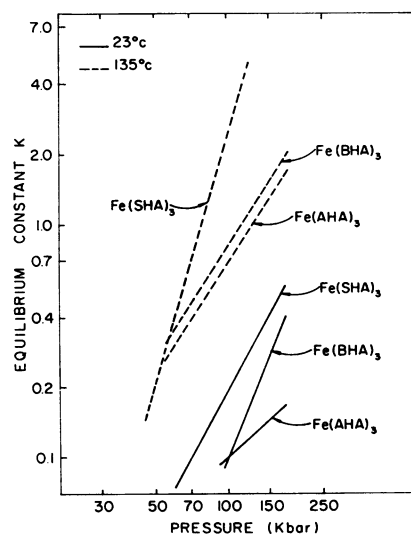

Fig. 3. LnK vs. LnP for Fe(SHA)₃.

Fig. 4. Lnk vs. LnP for ferric hydroxamates.

TABLE 1. *Mössbauer parameters for ferric hydroxamates**

	Pressure (kbar)								
	4	25	50	75	100	125	150	175	
Fe(SHA)₃									
Fe(III)	0.440	0.450	0.445	0.430	0.427	0.420	0.413	0.406	(23°C)
Fe(II)	—†	—	—	1.36	1.35	1.34	1.33	1.32	
Fe(AHA)₃									
Fe(III)	0.415	0.414	0.414	0.411	0.410	0.403	0.393	0.381	(23°C)
Fe(II)	—	—	—	1.39	1.38	1.37	1.36	1.35	
Fe(BHA)₃									
Fe(III)	0.370	0.426	0.431	0.426	0.422	0.415	0.408	0.401	(23°C)
Fe(II)	—	—	—	—	1.35	1.31	1.27	1.23	
Ferrichrome A									
Fe(III)	0.195	0.265	0.390	0.410	0.394	0.370	0.350	0.335	(23°C)
Fe(II)	—	—	1.202	1.200	1.200	1.198	1.195	1.194	
Fe(SHA)₃									
Fe(III)	0.94	1.11	1.23	1.31	1.36	1.40	1.43	1.45	(23°C)
Fe(II)	—	—	—	2.29	2.20	2.17	2.15	2.14	
Fe(III)	—	—	1.26	1.30	1.42	1.47	1.51	1.53	(110°C)
Fe(II)	—	—	—	2.05	2.02	2.03	2.05	2.06	
Fe(III)	—	—	1.32	1.41	1.46	1.50	1.51	1.53	(135°C)
Fe(II)	—	—	2.08	2.04	2.06	2.09	2.12	2.14	
Fe(AHA)₃									
Fe(III)	0.70	0.99	1.13	1.21	1.25	1.28	1.30	1.33	(23°C)
Fe(II)	—	—	—	2.32	2.31	2.30	2.28	2.27	
Fe(III)	—	—	1.19	1.32	1.41	1.48	1.51	1.55	(135°C)
Fe(II)	—	—	—	1.97	1.98	2.00	2.01	2.03	
Fe(BHA)₃									
Fe(III)	0.70	0.91	1.06	1.17	1.25	1.32	1.36	1.39	(23°C)
Fe(II)	—	—	—	—	2.31	2.25	2.19	2.14	
Fe(III)	—	1.06	1.18	1.27	1.36	1.42	1.47	1.48	(105°C)
Fe(II)	—	—	—	—	1.99	1.98	1.97	1.97	
Fe(III)	—	—	1.06	1.17	1.25	1.32	1.36	1.39	(135°C)
Fe(II)	—	—	1.86	1.86	1.85	1.85	1.84	1.84	
Ferrichrome A									
Fe(III)	0.37	0.59	1.01	1.20	1.32	1.39	1.46	1.50	(23°C)
Fe(II)	—	—	2.49	2.44	2.42	2.40	2.40	2.39	

* Data given as (mm/sec) isomer shifts (quadrupole splitting) relative to iron metal.

† —, not present.

discussed below, that Fe(BHA)₃ showed markedly more asymmetric peaks than the other two model compounds, indicating some difference in its structure, which might account for the unusual behavior of its isomer shift. However, the fitting of such asymmetric peaks is difficult, and we cannot eliminate the possibility that the peculiar behavior in the low-pressure region is in part an artifact of the fitting. The Fe(III) quadrupole splittings all increased by about 0.6 mm/sec in 175 kbar. The Fe(II) isomer shift of about 1.35 mm/sec, and the large quadrupole splitting of about 2.20 mm/sec, clearly indicates that the Fe(II) produced at high pressure is high spin.

As discussed above, Fe(SHA)₃ showed asymmetric peaks at low pressures, and with increasing pressure the asymmetry decreased. Fe(BHA)₃ showed even greater relaxation effects, and its ratio of peak widths was about 1.56 at 4 kbar, as compared to 1.25 for Fe(SHA)₃. With increasing pressure, the relaxation effect in Fe(BHA)₃ also decreased, and by 70 kbar the ratio was down to 1.09. By 90 kbar, the ratio appeared to be one. Fe(AHA)₃, in contrast to the other two compounds, showed only a small relaxation effect, so the peaks were fit as symmetric at all pressures. This relaxation effect was also

observed in some ferric hydroxamates studied by Epstein *et al.* (13). They found that Fe(AHA)₃ and Fe(SHA)₃ showed modest relaxation effects, while Fe(BHA)₃ exhibited a much larger effect. The asymmetry is associated with the relative population of ground and excited nuclear states. As the distance between iron sites decreases with increasing pressure, we would expect both an increase in splitting of the levels and a decrease in relaxation time. Both these factors would serve to decrease the asymmetry.

STUDIES WITH FERRICHROME A

Ferrichrome A was also studied at different pressures, both optically and with Mössbauer resonance. The optical spectrum is very much like that of the model compounds described above. The charge transfer peak is at 22.22 kK, and at 50% of maximum absorption the energy is 18.9 kK. The shift of the charge transfer peak at 50% of maximum absorption is 1.5 kK to lower energy, some 0.25 kK larger than the shifts in the model compounds.

The Mössbauer spectra of ferrichrome A at 4 kbar and 138 kbar is shown in Fig. 5. Note that the lower-pressure spectrum is broadened considerably more than the model compound
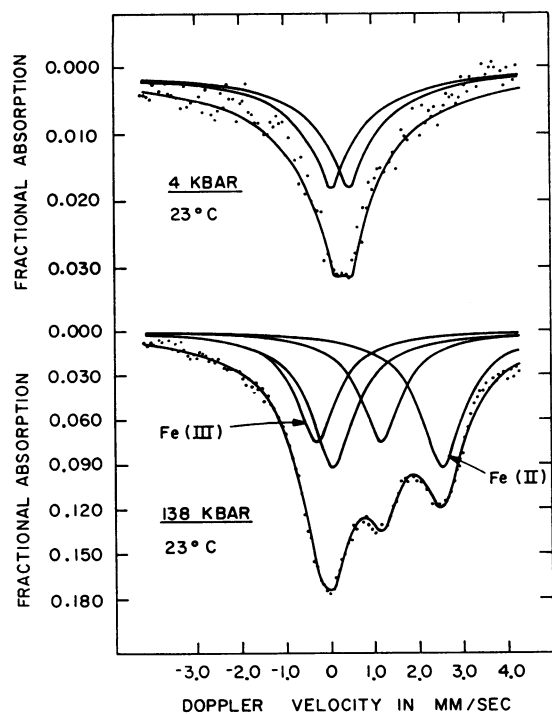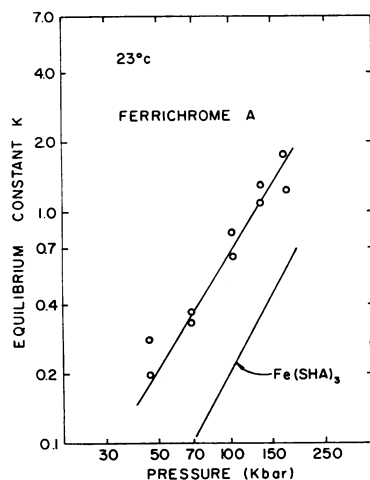
Fig. 5. Mössbauer spectra of ferrichrome A.

TABLE 2. *Parameters A and B for* $K = AP^B$

Compound	T(°C)	A	B	Pressure range (kbar)
Fe(SHA)₃	23	3.66×10^{-5}	1.86	60–170
	110	6.64×10^{-7}	3.03	50–150
	135	1.45×10^{-7}	3.62	40–110
Fe(BHA)₃	23	8.08×10^{-7}	2.54	100–175
	105	2.35×10^{-5}	2.03	90–175
	135	3.96×10^{-4}	1.66	60–175
Fe(AHA)₃	23	1.30×10^{-3}	0.94	100–175
	135	4.19×10^{-4}	1.60	55–175
Ferrichrome A	23	3.12×10^{-4}	1.67	40–175

mm/sec. The isomer shift and quadrupole splitting of both Fe(III) and Fe(II) in ferrichrome A are given in Table 1, and the parameters A and B are given in Table 2 for all the ferric hydroxamates. Values of the parameters for ferrichrome should be considered as somewhat qualitative since the spectra were broadened by the hyperfine interaction.

The reduction of ferric iron to ferrous iron requires a transfer of an electron from the ligands to the metal. As previously mentioned, the charge transfer energies for Fe(AHA)₃, Fe(BHA)₃, and Fe(SHA)₃ are 23.25, 22.22, and 21.95 kK, respectively. Referring back to Fig. 4, we can see that this is also the order of increasing conversion, so there is a definite correlation between the energy of the charge transfer band and the amount of conversion observed in a series of closely similar compounds. This is entirely consistent with our theory presented previously (15, 16), in which we suggested that the reduction occurred because the excited state (the Fe(II) plus ligands with a hole) decreases in energy with respect to the ground state (Fe(III) ion plus normal ligands) with increasing pressure, until a thermal transfer of an electron becomes possible. Because of the Franck-Condon principle, the optical transfer of an electron occurs vertically on a configuration coordinate diagram in contrast to the thermal process, as shown in Fig. 7. Nevertheless, the correlation between optical and thermal processes might be expected as long as the shapes of the potential wells are very similar from

Fe(SHA)₃ shown in Fig. 2. This broadening in ferrichrome A has been attributed by Wickman *et al.* (14) to the presence of a magnetic hyperfine structure. At room temperature, the hyperfine interaction is produced by a field having a shorter relaxation time than at low temperatures, and the hyperfine interaction dissolves into the broad line shown. For comparison to the model compounds, the spectra were fit as a quadrupole split pair at low pressure. At high pressure, the iron of ferrichrome A was also reduced, in greater yields than in the model compounds, as shown in Fig. 6, where the conversion of the ferrichrome at 23°C is compared to that of Fe(SHA)₃. Ferrichrome A was not run at higher temperatures because of the possibility of decomposition. The Fe(II) produced is also high spin, with an isomer shift of about 1.2

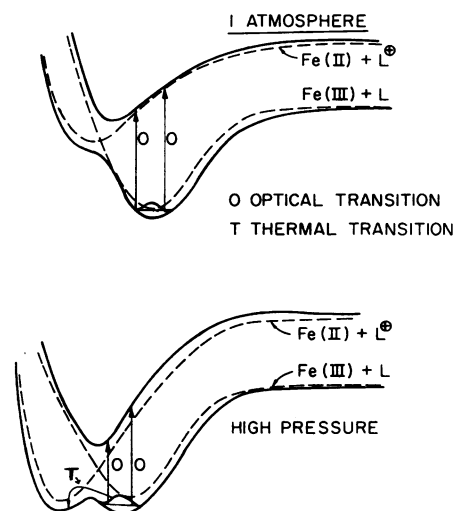Fig. 6. LnK vs. LnP for ferrichrome A.

Fig. 7. Schematic configuration coordinate diagram.

compound to compound. It is important for this correlation that the compounds be similar, since ferrichrome A, with its charge transfer energy at 22.22 kK, is similar in the sense of being oxygen coordinated; however, the structure is sufficiently different so that the potential well is modified and the conversion is not quantitatively consistent with that of the hydroxamates.

1. Neilands, J. B., *Struct. Bonding (Berlin)*, **1**, 59 (1966).
2. Neilands, J. B., *Bacteriol. Rev.*, **21**, 101 (1957).
3. Emery, T. E., and J. B. Nielands, *J. Amer. Chem. Soc.*, **82**, 3658 (1960).
4. Garibaldi, J. A., and J. B. Neilands, *J. Amer. Chem. Soc.*, **77**, 2429 (1959).
5. Linder, H. J., and S. Göttlicher, *Acta Crystallog.*, **B25**, 832 (1969).
6. Zalkin, A., J. D. Forrester, and D. H. Templeton, *Science*, **146**, 261 (1964).
7. Weinland, R. F., and G. Baier, *Berichte*, **53B**, 685 (1920).
8. Hantzch, A., and C. H. Desch, *J. Liebig's Ann. Chemie*, **323**, 1 (1902).
9. Fitch, R. A., T. E. Slykhouse, and H. G. Drickamer, *J. Opt. Soc. Amer.*, **47**, 1015 (1957).
10. Drickamer, H. G., and A. S. Balchou, *Modern Very High Pressure Techniques*, ed. R. Wentorf (Butterworths, London, 1962).
11. Debrunner, P., R. W. Vaughan, A. R. Champion, J. Cohen, J. Moyzis, and H. G. Drickamer, *Rev. Sci. Instrum.*, **37**, 1310 (1966).
12. Blume, M., *Phys. Rev. Lett.*, **18**, 305 (1967).
13. Epstein, L. M., and D. K. Straub, *Inorg. Chem.*, **8**, 453 (1969).
14. Wickman, H. H., M. P. Klein, and D. A. Shirley, *Phys. Rev.*, **152**, 345 (1966).
15. Drickamer, H. G., G. K. Lewis, and S. C. Furry, *Science*, **163**, 885 (1969).
16. Drickamer, H. G., V. C. Bastron, D. C. Fisher, and D. C. Grenoble, *J. Solid State Chem.*, **2**, 94 (1970).