

Available online at www.sciencedirect.com



Spectrochimica Acta Part A 60 (2004) 2843-2852

SPECTROCHIMICA ACTA PART A

www.elsevier.com/locate/saa

Structure investigation, spectral, thermal, X-ray and mass characterization of piroxicam and its metal complexes

M.A. Zayed, F.A. Nour El-Dien, Gehad G. Mohamed*, Nadia E.A. El-Gamel

Chemistry Department, Faculty of Science, Cairo University, Giza, Egypt

Received 19 August 2003; accepted 2 December 2003

Abstract

 $[M(H_2L)_2](A)_2, yH_2O$ (where H_2L : neutral piroxicam (Pir), A: Cl⁻ in case of Ni(II) or acetate anion in case of Cu(II) and Zn(II) ions and y = 0-2.5) and $[M(H_2L)_3](A)_z, yH_2O$ (A: SO_4^{2-} in case of Fe(II) ion (z = 1) or Cl⁻ in case of Fe(III) (z = 3) and Co(II) ions (z = 2) and y = 1-4) chelates are prepared and characterized using elemental analyses, IR, magnetic and electronic reflectance measurements, mass spectra and thermal analyses. IR spectra reveal that Pir behaves a neutral bidentate ligand coordinated to the metal ions through the pyridyl-N and carbonyl-O of the amide moiety. The reflectance and magnetic moment measurements reveal that these chelates have tetrahedral, square planar and octahedral geometrical structures. Mass spectra and thermal analyses are also used to confirm the proposed formulae and the possible fragments resulted from fragmentation of Pir and its chelates. The thermal behaviour of the chelates (TGA and DTA) are discussed in detail and the thermal stability of the anhydrous chelates follow the order Ni(II) \cong Cu(II) < Fe(II) < Zn(II) < Fe(III) < Co(II) chelates. The water molecules are removed in the first step while the Pir molecule is removed in the second and subsequent steps. X-ray powder diffraction was also used as a confirmatory tool to elucidate the crystallinity of the chelates.

© 2004 Elsevier B.V. All rights reserved.

Keywords: Piroxicam; Chelates; IR; TGA and DTA; Magnetic; Solid reflectance spectra; XRD; Mass spectra

1. Introduction

In addition to being an important article of commerce as a nonsteriodal anti-inflammatory agent [1], piroxicam (Pir) (4-hydroxy-2-methyl-*N*-2-pyridyl-2H-1,2-benzothiazine-3carboxamide) was an intrinsically interesting chemical compound by virtue of possessing four different heteroatom sites for positioning the two substituents necessary to complete its partial structure in Fig. 1.

Recently, it had been reported that the metabolic product of Pir in the biological system showed highly photosensitizing properties [2]. In order to investigate the possible role of the metal ions in the photosensitizing effect of Pir in biological system, the interaction of Cu(II) and other first row transition-metal ions using different spectroscopic methods had been studied. A variety of recent observations indicated that copper complexes when administered in conjunction with anti-inflammatory drugs exhibit synergistic activity [3]. It had also been found that the Cu complexes of some anti-arthritic drugs are themselves more active as anti-inflammatory agents than their parent compounds [4,5]. However, inspite of that, the preparation of the transition metal complexes of Pir had been reported earlier [6–8] and by screening the literature survey, no studies concerning their thermal and mass are reported. This led to give the main idea of the studies.

Hence, one of the goals of this paper is to calculate the stability constant values of Pir chelates with Fe(II), Fe(III), Co(II), Ni(II), Cu(II) and Zn(II) ions. From the structural data it is possible to discern how they are coordinated to the metal ions, then to hypothesize how they can be coordinated in biological systems. The study aims also to determine the coordination capacity of Pir that incorporates served binding sites via the IR, magnetic moment, conductance and solid reflectance spectra measurements. Thermal analyses techniques (TGA and DTA) were also used. The crystallinity of the chelates are confirmed by using X-ray powder diffraction technique.

^{*} Corresponding author. Fax: +202-572556.

E-mail address: ggenidy@hotmail.com (G.G. Mohamed).



Fig. 1. Piroxicam system diagram and positional notation the planer configuration shown is EZE, from left to right.

2. Experimental

2.1. Materials and measurements

All reagents and solvents were of the analytical grade. Elemental analyses were carried out at the Microanalytical center at Cairo University. IR spectra were recorded on a Perkin-Elmer FT-IR type 1650 spectrophotometer in wavenumber region $4000-200 \,\mathrm{cm}^{-1}$. The solid reflectance spectra were measured on a Shimadzu 3101PC spectrophotometer. The molar magnetic susceptibilities were measured on powdered samples using the Faraday method. The diamagnetic corrections were made by Pascal's constant and Hg[Co(ScN)₄] was used as calibrant. The molar conductance of solid complexes in DMF was measured using Sybron–Barnstead conductometer (Meter-PM.6. E =3406). The thermogravimetric analysis (TGA and DTGA) and differential thermal analysis (DTA) were carried out in dynamic nitrogen atmosphere (20 ml min^{-1}) with a heating rate of 10 °C min⁻¹ using Shimadzu TGA-50H and DTA-50H thermal analyzer, respectively. The mass spectra were recorded by the EI technique at 70 eV using MS-5988 GS-MS Hewlett-Packard instrument in the Microanalytical Center, Cairo University. The X-ray powder diffraction analyses were carried out by using Rigku Model ROTAFLEX Ru-200. Radiation was provided by copper target (Cu anode 2000 W) high intensity X-ray tube operated at 40 KV and 35 MA. Divergence and the receiving slits were 1 and 0.1, respectively.

2.2. Synthesis of the complexes

A hot ethanol solution $(60 \,^{\circ}\text{C})$ of the respective metal chlorides ((Fe(II), Fe(III), Co(II) and Ni(II)) or acetates (Cu(II) and Zn(II)) (25 ml, 0.5 mM) was added to hot ethanol solution $(60 \,^{\circ}\text{C})$ of Pir (1 mM, 25 ml). The resulting mixture was stirred under reflux for 30 min. The complexes were precipitated as microcrystalline powders. They were removed by filtration, washed with hot ethanol followed by diethylether and dried in a vacuum desiccator over anhydrous calcium chloride. The m.p and analytical data are collected in Table 1.

3. Results and discussion

Some physical properties of the synthesized complexes are listed in Table 1. The complexes are less soluble than the Pir ligand in solvents, such as alcohols, tetrahydrofuran, dichloromethane, and acetone. The elemental analyses are satisfactory in the acceptable range of $\pm 0.5\%$ error and show that the complexes have a ligand-to-metal ratio of 2:1 for Ni(II), Cu(II) and Zn(II) ions and 3:1 for Fe(II), Fe(III) and Co(II) ions, in accordance with the results obtained by molar ratio method. Mass spectra of Pir show that the expected molecular ion. The molecular ion in the mass spectra of Cu(II) and Zn(II) chelates agree with 2:1 ligand-to-metal stocichiometric ratio.

3.1. Calculation of metal complexes stability constants

The stability constants of the metal chelate are evaluated spectrophotometrically using recommended methods [9]. The calculated log *K* values for the Fe(II)-, Fe(III)-, Co(II)-, Ni(II)-, Cu(II)- and Zn(II)-chelates of Pir are given in Table 2. The stability of Pir–metal chelates follows the sequence: Cu(II) > Co(II) > Zn(II) \cong Ni(II) > Fe(III) > Fe(II).

3.2. IR spectra and mode of bonding

Detailed interpretation of IR spectra and the effect of binding of Fe(II), Fe(III), Co(II), Cu(II) and Zn(II) ions on the vibration frequencies of the free Pir is discussed. The IR spectrum of Pir shows well-defined peaks at 3385 and 3336 cm⁻¹ are assigned to the ν (N–H) and ν (O–H) vibrations, respectively. In the wave number range $1700-1000 \,\mathrm{cm}^{-1}$, some of the most characteristic bands of these groups are found. The so-called amide I band, which represents mainly the ν (C=O) stretching mode, is seen as a very strong band at $1630 \,\mathrm{cm}^{-1}$, with a weak shoulder at 1612 cm⁻¹ in free Pir. In Pir complexes, this band is shifted to higher or lower frequencies or disappear, supporting the participation of the amide carbonyl group of Pir in the coordination to metal ion. This band appears broadened in most of the complexes by the simultaneous appearance of the $\delta(H_2O)$ vibration in the same range, a fact which obstructs the clear visualization of the effective displacement of the ν (C=O) band [7].

The band located at 1572 cm^{-1} is assigned to the ν (C=N) stretching vibration of pyridyl nitrogen. Coordination of the pyridyl nitrogen is indicated by a $3-26 \text{ cm}^{-1}$ shift to a higher wavenumber or $55-50 \text{ cm}^{-1}$ shift to a lower wavenumber in all the chelates [10–12]. The participation of pyridine in complex formation can be established from the shift of the ring deformation found at 604 and 405 cm⁻¹ in the free pyridine to higher frequencies [13]. In the present case, the in-plane ring deformation (604 cm⁻¹ in free pyridine) is probably overlapped by the 623 cm^{-1} . After complexation, the band at 623 cm^{-1} disappears and a new band at

Table 1 Analytical and physical data of piroxicam and their metal chelates

Compound	Chemical formula	Colour	M.P. (°C)	Found (Calcd.)%				Molecular wt. found (Calcd.)	$\mu_{\rm eff}$ (BM)	$\frac{\Lambda_{\rm m} \Omega^{-1}{\rm cm}^2}{{\rm mol}^{-1}}$	
				С	Н	Ν	S	М			
Piroxicam [Pir]	C ₁₅ H ₁₃ N ₃ O ₄ S	White	198	54.50 (54.38)	3.70 (3.90)	12.34 (12.68)	9.55 (9.66)	_	331 (331)	_	_
Fe(II)-Pir	[Fe(Pir)3]SO4·H2O	Reddish brown	115	46.41 (46.43)	3.60 (3.53)	11.20 (10.80)	11.42 (11.00)	4.81 (4.82)	_	4.98	83
Fe(III)-Pir	[Fe(Pir) ₃]Cl ₃ ·4H ₂ O	Reddish brown	174	43.50 (43.99)	3.62 (3.82)	10.50 (10.26)	7.71 (7.82)	4.60 (4.57)	_	5.30	320
Co(II)-Pir	[Co(Pir)3]Cl2·H2O	Orange	206	47.45 (47.33)	3.36 (3.59)	11.20 (11.04)	8.62 (8.41)	5.38 (5.17)	-	4.95	153
Ni(II)–Pir	[Ni(Pir)2]Cl2·H2O	Light green	219	44.35 (44.44)	3.25 (3.33)	10.25 (10.37)	7.81 (7.90)	7.22 (7.28)	(810)	Diam	165
Cu(II)-Pir	$[Cu(Pir)_2](Ac)_2$	Olive green	193	48.00 (48.37)	3.62 (3.79)	10.10 (9.95)	7.75 (7.58)	7.38 (7.52)	842 (843.5)	1.81	140
Zn(II)-Pir	$[Zn(Pir)_2](Ac)_2 \cdot H_2O$	Lemon	236	47.59 (47.27)	4.05 (4.17)	9.50 (9.73)	7.65 (7.42)	8.09 (7.53)	863 (863)	Diam	117

 Table 2

 Stability constants of di- and tri-valent metal–Pir chelates

Metal	log K	Metal	$\log K$
Fe ^{II} Fe ^{III} Co ^{II}	$5.26 \pm 0.48 \\ 5.42 \pm 0.17 \\ 5.91 \pm 0.31$	Ni ^{II} Cu ^{II} Zn ^{II}	$5.49 \pm 0.78 \\ 6.96 \pm 0.47 \\ 5.50 \pm 0.27$

 $[M^{n+}-Pir] = 2.5 \times 10^{-5} M; t = 25 \circ C.$

672–656 cm⁻¹ is found in the chelates, as previously reported for pyridine complexes of divalent ions of the first row transition metal series [14]. Detecting the shift of the out-of-plane $\rho(py)$ at 405 cm⁻¹ is difficult because the spectra are rich for the complexes.

Two bands, located at 1351 and 1040 cm^{-1} , are assigned to the antisymmetric and symmetric stretching vibrations of the SO₂ group, respectively. These two SO₂ bands are shifted to lower or higher frequencies in complexes, in accordance to the assignment of Cini et al. [15]. As the SO₂ group is not involved in metal binding, this shift to higher frequencies must be related to important hydrogen bonding effects. Similar effects are also known to occur in divalent metal saccharinates (a molecule with a structure very similar to that of the benzothiazine part of Pir) [16,17]. Another possible explanation for this shift may be due to the electronic density changes on the sulfur atom and in the ring after complex formation.

The ν (M–N) bands for the pyridyl nitrogen appeared in the wave number range 564–572 cm⁻¹ for the di- and trivalent metal ion chelates [14]. New peaks of weak or medium intensity are observed in the wave number range 358–386 cm⁻¹ which are attributed to ν (M–O) vibrations of binary chelates [18]. The vibrations of the sulphate ion in Fe(II)-chelate appear at 1160 and 629 cm⁻¹. This indicates the uncoordinated nature of sulphate group [19].

3.3. Magnetic susceptibility and electronic spectra measurements

In the solid reflectance spectrum of Fe(III)-complex, one band is observed at $21.93 \times 10^3 \text{ cm}^{-1}$, which may be assigned to the ${}^{6}A_{1g} \rightarrow T_{2g}$ (G) transition in octahedral geometry of the complex [20]. The magnetic moment of the Fe(III)-complex ($\mu_{eff} = 5.30 \text{ BM}$) is within the range of values corresponding to octahedral geometry. The diffused reflectance spectra of Fe(II) complex shows two absorption bands at 12.20×10^3 and 22.22×10^3 cm⁻¹ which are assigned to ${}^{5}T_{2g} \rightarrow {}^{5}Eg$ transition and charge transfer, respectively [21,22]. The magnetic moment of the Fe(II) complex $(\mu_{\rm eff} = 4.98 \, \rm BM)$ is within the range of octahedral geometry. The d-d spectrum of the Ni(II) complex shows two bands at 16.89×10^3 and 20.92×10^3 cm⁻¹ which are assigned to ${}^{1}A_{1g} \rightarrow {}^{1}A_{2g}$ and ${}^{1}A_{1g} \rightarrow {}^{1}B_{1g}$ transitions, respectively, in a square planar geometry [23,24]. The diamagnetic nature of the complex lends support to the above geometry. The electronic spectrum of the Co(II) complex gives three bands at 12.27×10^3 , 15.04×10^3 and 17.59×10^3 cm⁻¹ wavenumber regions, respectively, which are assigned to the transitions ${}^4T_{1g}(F) \rightarrow {}^4T_{2g}(F)(\nu_1), {}^4T_{1g}(F) \rightarrow {}^4A_{2g}(F)(\nu_2)$ and ${}^4T_{1g}(F) \rightarrow {}^4T_{2g}(P)(\nu_3)$, respectively, suggesting that there is an octahedral geometry around Co(II) ion [25–27]. The magnetic susceptibility measurement value is 4.95 BM is an indication of octahedral geometry [28].

Cu(II) complex exhibits two broad bands at $16.67 \times 10^3 \text{ cm}^{-1}$ and 14.71×10^3 assigned to ${}^2\text{Eg} \rightarrow {}^2\text{T}_{2g}$ and ${}^2\text{B}_{1g} \rightarrow {}^2\text{A}_{1g}$ transitions as well as a shoulder band in the range $17.54 \times 10^3 \text{ cm}^{-1}$ characteristic of a square planner geometry for Cu(II) complexes with $d_{x^2-y^2}$ ground state [29,30]. The magnetic moment of the Cu(II)-complex is 1.81 BM, which support this geometry [31]. Zn(II) complex is diamagnetic and is likely to be tetrahedral.

3.4. Mass spectra

The mass spectrum of Pir shows a well-defined parent peak at m/z = 331 (M⁺) with a relative intensity = 22%. The C₃H₃NO⁺ ion, m/z = 69, RI = 100% is the base peak in the spectrum of Pir. The parent ion and other fragments obtained by cleavage in different positions in the Pir molecule and the possible fragment ions are shown in Scheme 1. It is found that, the molecular ion peaks agree to a great extent with those previously published [32]. This means that Pir drug is a pure substance and ready for the use as a reagent in this study.

The mass histograms of Cu(II)- and Zn(II)-complexes with Pir show parent peaks in their mass spectra at m/z =845 (MH⁺_{1.5}, RI = 46%) and 863 (M⁺, RI = 24%) for the Cu(II)- and Zn(II)-complexes, respectively (as given in Table 3). The mass spectra of both chelates show a molecular ion peak with m/z = 173 (RI = 59 and 26% for Cu(II)and Zn(II)-chelates, respectively) which can be assigned to C₁₀H₇NO₂⁺ ion. The different molecular ion peaks given in Table 3 can be attributed to the loss of metal oxide, hydrated water, anions and different ions resulted from Pir molecule decomposition.

3.5. Thermal analyses (TGA and DTA)

The TGA curve of Pir refers to two stages of mass losses at temperature ranges ~170–360 and 360–650 °C. These stages involved mass losses of 92.15% (f. 92.26%) and 7.85% (f. 7.65%) for the first and second steps of decomposition, respectively. These mass losses may be due to the successive losses of $C_{13}H_{11}N_3O_4S$ and C_2H_2 at the given temperature ranges, respectively. DTA curve referes to five endothermic and three exothermic peaks are strong and centered at 200 (endo-), 212 (exo-), 243 (endo-), and 266 °C (exo-). The remaining peaks are weak and centered at 370 (endo-), 492 (exo-), 562 (endo-) and 620 °C (endo-).



Scheme 1. Mass fragmentation pattern of Pir.

The dehydration step in Fe(III)–Pir complex is found in the temperature range 65-170 °C (wt. loss: calcd. 5.87%, f. 6.27%). While, in case of Ni(II)–Pir and Fe(II)–Pir complexes, this step occurs in the temperature range 25–80 °C (wt. loss: calcd. 2.22%, f. 2.35%) and 25–110 °C (wt. loss: calcd. 1.55%, f. 1.61%) for Ni(II) and Fe(II)-complexes, respectively. In addition, the dehydration step in Zn(II)-complex, the dehydration step is immediately accompanied by a loss of C₂H₆ molecule within the temperature range 30–150 °C (wt. loss: calcd. 5.79%, f. 5.83%). The Co(II)-Complex losses water of hydration in the temperature range 50–200 °C (wt. loss: calcd. 7.80%, f. 7.63%) is accompanied by the loss of Cl₂ gas. It is obvious from the DTA curves (Table 4) that, the loss of water molecules of hydration appears endothermic and/or endothermic accompanied by exothermic peaks in the temperature ranges mentioned above. For Co(II)-complex, the loss of water of hydration and Cl₂ gas molecules appear as

Table 3 Mass spectral data of Cu(II)- and Zn(II)-Pir metal chelates

Compound	<i>m</i> / <i>z</i> Found (calcd.)	RI (%)	Fragments
[Cu(Pir) ₂](Ac) ₂	845 (843.5) 662 (661.5) 569.5 (570.5) 516.5 (515.5) 400 (401) 283 (282) 211 (210) 173 (173) 152.5 (153) 139 (138) 117 (117) 98 (97) 71 (69)	46 59 50 45 68 59 59 59 64 100 59 68 55	$\begin{array}{l} [C_{34}H_{32}CuN_6O_{12}S_2]^+\\ [C_{30}H_{26}CuN_6O_6S]^+\\ [C_{28}H_{23}CuN_6O_4]^+\\ [C_{25}H_{20}CuN_6O_3]^+\\ [C_{19}H_{19}N_3O_5S]^+\\ [C_{10}H_8N_3O_5S]^+\\ [C_{9}H_8NO_3S]^+\\ [C_{10}H_7NO_2]^+\\ [C_{7}H_5O_2S]^+\\ [C_{6}H_2O_2S]^+\\ [C_{8}H_5O]^+\\ [C_{4}H_3NS]^+\\ [C_{4}H_3NS]^+\\ [C_{4}H_3NS]^+\\ \end{array}$
[Zn(Pir) ₂](Ac) ₂ ·H ₂ O	863 (863) 741 (742) 532 (532) 351 (349) 316 (317) 263 (264) 175 (173) 144 (145) 106 (106) 81 (80)	24 30 20 29 29 39 26 25 31 100	$\begin{array}{l} [C_{34}H_{34}N_6O_{13}S_2Zn]^+\\ [C_{31}H_{29}N_6O_8S_2Zn]^+\\ [C_{19}H_{21}N_3O_9SZn]^+\\ [C_{14}H_{10}N_3O_2SZn]^+\\ [C_{10}H_8N_2O_4SZn]^+\\ [C_{15}H_{10}N_3O_2]^+\\ [C_{10}H_7NO_2]^+\\ [C_{9}H_5O_2]^+\\ [C_{7}H_6O]^+\\ [C_{5}H_6N]^+ \end{array}$

endo- and exothermic peaks in the temperature range 50–200 °C.

On the other hand, the coordinated water molecules are directly bonded to the metal ion. Hence, they are eliminated at higher temperature range than those of water of hydration. It is recommended in literature [33] that the water of coordination is usually eliminated in the temperature ranges 100-350 °C.

The complexes of Pir with di- and tri-valent metal ions undergo thermal decomposition in a series of overlapping reactions with multiple products. The interpretation of the TGA and DTA curves in terms of discrete stages with associated weight losses requires a great deal of imagination and certain amount of wishful thinking. The organic part (Pir) together with the anions (Cl⁻, Ac⁻ or SO₄²⁻) in the moiety of complexes may decompose in more than two steps with the possibility of the formation of more than one intermediate. These intermediates may finally decompose to stable metal oxides or sulfides (see Table 4). It is shown from these results that, the calculated and estimated mass losses are comparable.

The DTA curves show a medium to strong exothermic and endothermic peaks; the position of each peak is given in Table 4, accompanying the decomposition of the anions and Pir molecule to different gases like H_2O , O_2 , CO_2 , CO, NO, NO_2 , SO_2 , SO_3 , etc. The Fe(II)-, Fe(III)- and Zn(II)-chelates show a medium to strong exo- or endothermic peaks in the DTA curves which may be attributed to phase transitions. The position and nature of the peaks (exo- or endo-) are given in Table 4.

On the basis of the above observations, the thermal decomposition of the metal chelates under investigation may be generally summarized in the proposed Scheme 2.

-First decomposition step is represented as the following:



In the second and subsequent steps of decomposition, the loss of the anions, coordinated water and ligand molecules takes place, as follows:

$$[M(H_2L)_z](A)_x \xrightarrow{80-1200 \text{ C}} MX + zH_2L + xA$$

Final decomp.

where MX = metal oxide (MO) in case of Fe(III), Co(II), Ni(II) and Cu(II).

= metal sulphide (MS) in case of Fe(II) and Zn(II).

Scheme 2. Thermal decomposition of Pir complexes.

<i>M.A</i> .
Zayed
et
al. ,
⁷ Spectrochimica Acta Part A 60
(200
4) 2843–2852

Table 4 Thermal analyses (TGA and DTA) results of Pir-binary chelates

Temperature range (°C)	Peak temperature in DTA (°C)	No. of decomposed steps	Mass loss % found (calcd.)	Total mass loss % found (calcd.)	Assignment	Metallic residue
25-110	-100(+)	1.0	1.61 (1.55)		Loss of H ₂ O	FeS
110-700	-248(-), 279(+), 370(-), 390(+), 432(-),	2.0	90.84 (90.88)	92.45 (92.43)	Loss of SO ₄ ²⁻ and Pir molecules	
	500(+), 483(-), 548(+), 659(-)					
	-934(-), 1141(+)				Phase transition	
65-170	-120(+), 147(-)	1.0	6.27 (5.87)		Loss of 4H ₂ O	FeO
170-650	-190(+), 206(-), 252(-), 270(+), 380(+),	2.0	87.82 (88.27)	94.09 (94.14)	Loss of 1.5 Cl ₂ and Pir molecules	
	468(-), 630(+), 659(+)					
	-805(+), 825(-)				Phase transition	
50-200	-71(+), 85(-), 135(+), 150(-), 190(-)	1.0	7.63 (7.80)		Loss of H ₂ O and Cl ₂	CoO
200-1200	-361(+), 540(-), 656(+), 766(+), 790(+),	2.0	85.54 (85.62)	93.17 (93.42)	Loss of Pir molecules	
	935(+), 946(-), 1020(+), 1030 (-)					
25-80	-40(+), 75(-)	1.0	2.35 (2.22)		Loss of H ₂ O	NiO
80-1200	-91(+), 113(-), 180(+), 195(-), 275(-),	1.0	88.47 (88.52)	90.82 (90.74)	Loss of Cl ₂ and Pir molecules	
	330(-), 380(+), 405(-), 445(+), 516(+)					
80-1200	-237(-), 260(+), 470(-), 500(-), 891(+),	3.0	90.58 (90.76)	90.58 (90.76)	Loss of acetate and Pir molecules	CuO
	1050(-), 1090(-), 1150(-)					
30-150	-45(+), 75(+), 86(-), 130(+)	1.0	5.83 (5.79)		Loss of H ₂ O and C ₂ H ₆	ZnS
150-750	-170(+), 185(-), 276(+), 390(-), 405(-),	3.0	82.10 (82.97)	87.93 (88.76)	Loss of 2CO ₂ and Pir molecules	
	586(-), 680(+), 710(-), 730(+), 748(-),					
	755(-)					
	-790(-), 849(-), 952(+), 1183(+)			Phase transition		
	Temperature range (°C) 25–110 110–700 65–170 170–650 50–200 200–1200 25–80 80–1200 80–1200 30–150 150–750	Temperature range (°C)Peak temperature in DTA (°C)25-110 $-100(+)$ 110-700 $-248(-), 279(+), 370(-), 390(+), 432(-), 500(+), 483(-), 548(+), 659(-))$ $-934(-), 1141(+)$ 65-170 $-120(+), 147(-)$ 170-650 $-190(+), 206(-), 252(-), 270(+), 380(+), 468(-), 630(+), 659(+))$ $-805(+), 825(-)$ 50-200 $-71(+), 85(-), 135(+), 150(-), 190(-)$ $200-1200$ 20-1200 $-361(+), 540(-), 656(+), 766(+), 790(+), 935(+), 946(-), 1020(+), 1030(-)$ 25-80 $-40(+), 75(-)$ 80-1200 $-91(+), 113(-), 180(+), 195(-), 275(-), 330(-), 380(+), 405(-), 445(+), 516(+)$ 80-1200 $-237(-), 260(+), 470(-), 500(-), 891(+), 1050(-), 1090(-), 1150(-)$ 30-150 $-45(+), 75(+), 86(-), 130(+)$ 150-750 $-170(+), 185(-), 276(+), 390(-), 405(-), 586(-), 680(+), 710(-), 730(+), 748(-), 755(-)$ $-790(-), 849(-), 952(+), 1183(+)$	$\begin{array}{c ccccccccccccccccccccccccccccccccccc$	$\begin{array}{c c c c c c c c c c c c c c c c c c c $	$ \begin{array}{c c c c c c c c c c c c c c c c c c c $	$ \begin{array}{c c c c c c c c c c c c c c c c c c c $

(+): endothermic, (-): exothermic.

3.6. Calculation of activation thermodynamic parameters of pir chelates

The thermodynamic activation parameters of decomposition processes of dehydrated complexes namely activation energy (E^*), enthalpy (ΔH^*), entropy (ΔS^*) and Gibbs free energy change of the decomposition (ΔG^*) were evaluated graphically by employing the Coast–Redfern relation [34]:

$$\log\left[\frac{\log\{W_{\rm f}/(W_{\rm f}-W)\}}{T^2}\right] = \log\left[\frac{AR}{\theta E^*}(1-\frac{2RT}{E^*})\right]$$
$$-\frac{E^*}{2.303RT}$$

where $W_{\rm f}$ is the mass loss at the completion of the reaction, W the mass loss up to temperature T; R the gas constant, E^* the activation energy in kJ mol⁻¹, θ the heating rate and $(1-(2RT/E^*)) \cong 1$. A plot of the left-hand side of equation 1 against 1/T gives a slope from which E^* was calculated and A (Arrhenius factor) was determined from the intercept. The activation energy for the elimination of water molecules of hydration from the Fe(II)-chelate is equal 138.3 kJ mol⁻¹. Regarding the Fe(III)-chelate, the activation of the first step is equal $47.20 \text{ kJ mol}^{-1}$. Meanwhile, the data listed in Table 5 show that the Co(II)-, Ni(II)- and Zn(II) chelates are found to have activation energy for the elimination of water molecules of hydration are equal to 3, 17.4 and $34.69 \text{ kJ mol}^{-1}$, respectively. From the activation energy values, one can concluded that, the water and anion molecules are easy to be eliminated from the chelates according to the following order: Co(II) > Ni(II) > Zn(II) >Fe(III) > Fe(II) chelates. According to the kinetic data obtained fro DTG curves, all the complexes have -ve entropy, which indicates that activated complexes have more ordered systems than reactants.

3.7. X-ray powder diffraction

The corresponding spectrum of the X-ray powder diffraction pattern is presented in Fig. 2. By comparing the X-ray chart of Pir with that previously published [27,28] and the interplanar distances (d(Å)), it is concluded that Pir exists in the cubic form and not in the needle form. XRD analysis for chelates shows that the coordination of Pir alone or in the presence metal ions (Fe(II), Fe(III), Co(II), Ni(II), Cu(II) and Zn(II)) changes the XRD pattern of the ligand. This means that the metal ions are not fitted in the same phase of Pir. Therefore, the non-similarity of XRD pattern between the metal ions and complexes suggests that these complexes have a different phase structure than the Pir.

3.8. Structural interpretation

The structure of the complexes was confirmed by the IR, magnetic, solid reflectance, conductance and XRD analysis.



Fig. 2. X-ray diffraction pattern of (a) Pir, (b) Fe(II), (c) Fe(III), (d) Co(II), (e) Ni(II), (f) Cu(II) and (g) Zn(II) chelates.

Table 5 Thermodynamic data of the thermal decomposition of metal chelates with Pir

Complex	Decomposed range (°C)	E^* (kJ mol ⁻¹)	$A (S^{-1})$	$\Delta S^* \; (\mathrm{J}\mathrm{K}^{-1} \; \mathrm{mol}^{-1})$	$\Delta H^* \text{ (kJ mol}^{-1}\text{)}$	$\Delta G^* (\text{kJ mol}^{-1})$
[Fe(II)(Pir)3]SO4·H2O	25–110	138.3	1.72×10^{5}	-15.69	137.8	38.64
	110-330	193.6	5.09×10^{17}	-11.40	191.3	188.2
	330-700	212.5	1.55×10^{13}	-0.38	208.3	208.1
[Fe(III)(Pir)3]SO4·4H2O	65–170	47.20	2.08×10^{5}	-16.15	46.35	47.99
	170–365	52.71	1.05×10^{4}	-20.11	50.47	55.88
	365-650	28.85	5.59×10^{2}	-30.92	22.91	44.99
[Co(II)(Pir)3]Cl2·H2O	50-200	3.0	1.64×10^{2}	-32.58	0.656	2.93
	200-540	9.05	5.16×10^{2}	-30.22	6.337	16.21
	540-1200	33.30	1.64	-30.01	26.293	51.53
[Ni(II)(Pir) ₂]Cl ₂ ·H ₂ O	25-80	17.40	1.97×10^{2}	-25.46	16.52	19.22
	80-1200	11.34	1.62×10^{3}	-31.66	7.74	21.44
[Cu(II)(Pir) ₂](Ac) ₂	80–300	121.1	1.4×10^{7}	-12.86	118.9	112.4
	300-600	131.3	6.86×10^{7}	-11.97	127.0	133.2
	600–1200	62.00	1.88×10^{2}	-27.72	53.84	81.0
[Zn(II)(Pir)2](Ac)2·H2O	30–150	34.69	5.12×10^{3}	-20.04	33.67	36.12
	150–333	115.4	1.12×10^{10}	-6.30	113	114.8
	333–531	79.3	6.42×10^{4}	-18.82	75.51	84.11
	531–750	6.29	2.5×10^2	-34.16	0.40	27.10

The chelation is brought about by the pyridine nitrogen and C=O of amide group of Pir. From the reflectance spectra and magnetic moment measurements, the geometrical structure of the chelates were proposed and found to be octahedral, tetrahedral or square planar.

As a general conclusion, Pir behaves as neutral bidentate ligands in all the chelates. The structure of the chelates can be given as follows:



(Water of hydration and anions are excluded for simplicity)

References

- J.R Prous, Annual Drug Data Report, Science, Barcelona, 1982, p. 197.
- [2] A. Western, J.R. Van Camp, R. Benasson, E.J. L, I.E. Kochevar, PhotoChem. Photobiol. 46 (1987) 469.
- [3] R.K. Crouch, T.W. Kensler, L.W. Oberley, J.R.J. Sorenson, in: K.D. Karlin, J. Zubieta (Eds.), Biological and Inorganic Copper Cheimstry, vol.1, Adenine Press, Guilderland, NY, 1985, p. 1397.
- [4] (a) J.R.J. Sorenson, J. Med. Chem. 19 (1976) 135;
 (b) J.R.J. Sorenson (Ed.), Inflammatory Diseases and Copper, Humana Press, Clifton, NJ, 1982, p. 289.
- [5] D.H. Brown, W.E. Smith, J.W. Teape, A.J. Lewis, J. Med. Chem. 23 (1980) 729.

- [6] R. Cini, G. Giorgi, A. Cinquantini, C. Rossi, M. Sabat, Inorg. Chem. 29 (1990) 5197.
- [7] E. Santi, M.H. Torre, E. Kremer, S.B. Etcheverry, E.J. Baran, Vibrat. Spectrosc. 5 (1993) 285.
- [8] S. El-Khateeb, S. Abdel Fattah, S. Abdel Razeg, M. Tawakkol, Anal. Lett. 22 (1989) 101.
- [9] A.D. Taneja, K.P. Srivastova, J. Inorg. Nucl. Chem. 33 (1971) 2678.
- [10] G.G. Mohamed, F.A. Nour El-Dien, N.E.A. El-Gamel, J. Thermal Anal. 67 (2000) 135.
- [11] E.Q. Castro, S. Berenes, N.B. Behrens, A.T. Behrens, A.T. Benavides, R. Contreras, H. Noth, Polyhedron 19 (2000) 1479.
- [12] W.R. Paryzek, E. Luks, Polyhedron 13 (1994) 899.
- [13] K. Nakamoto, Infrared and Raman Spectra of Inorganic and Coordination Compounds, fourth ed., Wiley, New York, 1980.
- [14] D.X. West, J.K. Swearingen, J.V. Martinez, S.H. Ortega, A.K. El-Sawaf, F.V. Meurs, A. Castineiras, I. Garcia, E. Bermejo, Polyhedron 18 (1999) 2919.
- [15] R. Cini, G. Giorgi, A. Cinquantini, C. Rossi, M. Sabat, Inorg. Chem. 29 (1990) 5197.
- [16] S.Z. Haider, K.M.A. Malik, J. Bangladesh Acad. Sci. 6 (1982) 119.
- [17] E.G. Ferrer, S.B. Etcheverry, E.J. Baran, Monatch. Chem. 124 (1993) 355.
- [18] G.G. Mohamed, M.A. Zayed, N.E.A. El-Gamel, Spectrochim. Acta Part (A) 58 (2002) 3167.
- [19] A.P. Mishra, S.K. Srivastava, V. Srivastava, J. Indian, Chem. Soc. 74 (1997) 487.
- [20] M.M. Moustafa, J. Thermal. Anal. 50 (1997) 463.
- [21] F.A. Cotton, G. Wilkinson, C.A. Murillo, M. Bochmann; Advanced Inorganic Chemistry, sixth ed., Wiley, New York, 1999.
- [22] A. Hazel, C.J. Mckenzie, L.P. Nielsen, Polyhedron 19 (2000) 1333.
- [23] A.Z. El-Sonbati, A.A. El-Bindary, E.M. Mabrouk, Trans. Met. Chem. 17 (1992) 66.
- [24] L Gomes, E. Pereira, B. de Castro, J. Chem. Soc., Dalton Trans. (2000) 1373.
- [25] G.G. Mohamed, Z.H. Abd El-Wahab, J. Thermal Anal. 73 (2003) 347.
- [26] M.S. Masoud, A.M. Hindawy, A.S. Soayed, Trans. Met. Chem. 16 (1991) 372.
- [27] N.K. Gaur, R. Sharma, R.S. Sindhu, J. Indian Chem. 78 (2001) 26.
- [28] N. Mondal, D.K. Dey, S. Mitra, K.M. Abul Malik, Polyhedron 19 (2000) 2707.

- [29] M.F. Iskander, T.E. Khalil, R. Werner, W. Hasses, I. Svoboda, H. Fuess, Polyhedron 19 (2000) 949.
- [30] P.S. Reddy, K.H. Reddy, Polyhedron 19 (2000) 1687.
- [31] M.F. Iskander, T.E. Khalil, R. Werner, W. Hasse, I. Svoboda, H. Fuess, Polyhedron 19 (2000) 1181.
- [32] British Pharmacopia, Her Majesty's Stationery Office, London, UK, 1998, p. 1260.
- [33] N.T. Abdel-Ghani, O. Esherif, Thermochim. Acta 156 (1989) 69.
- [34] A.W. Coats, J.P. Redfern, Nature 20 (1964) 68.