Data obtained from oxidation of these esters with sodium periodate show that the maltose ester is glucopyranosido-4-glucopyranose-1-phosphate and that the xylose ester is D-xylopyranose-1-phosphate. Both esters probably exist in the α-form.

The rate constants for hydrolysis of these esters in 0.376 N hydrochloric acid at 36° and also their dissociation constants were determined.

Maltose-1-phosphate and D-xylose-1-phosphate are not converted to polysaccharide by potato phosphorylase.

BERKELEY, CALIFORNIA RECEIVED JUNE 5, 1946

[CONTRIBUTION FROM THE WELLCOME RESEARCH LABORATORIES]

The Ultraviolet Absorption Spectra of Thiouracils[1]

BY GERTRUDE B. ELION, WALTER S. IDE AND GEORGE H. HITCHINGS

The ultraviolet absorption spectra of pyrimidines have received a good deal of attention in recent years. However, except for the recent paper of Miller, Roblin and Astwood[2] on 2-thiouracil and its oxidation product, no thiopyrimidine spectra have been reported in the literature.

During the present investigation of the absorption spectra of the mono- and dimercapto-analogs of uracil (2,4-dihydroxypyrimidine) and thymine (2,4-dihydroxy-5-methylpyrimidine) it was found that 2-hydroxy-4-mercaptopyrimidine (4-thiouracil) and 2-hydroxy-4-mercapto-5-methylpyrimidine (4-thiothymine) possess absorption bands so far removed from the usual absorption range of pyrimidines as to be unrecognizable as members of that group. It is interesting that the anomalous spectra of these 4-thio-derivatives are paralleled by their lack of anti-thyroid activity. Thus, Astwood[3] has reported that whereas 2-thiouracil and 2,4-dithiouracil have approximately the same antithyroid activity, 4-thiouracil is practically inactive.

Results and Discussion

The ultraviolet absorption spectra of 2-thiouracil, 2,4-dithiouracil, 4-thiouracil, 2-thiothymine, 2,4-dithiothymine and 4-thiothymine at pH values of 1.0, 7.0 and 11.0 are reported here.

A comparison of the spectra of these thio-compounds with the corresponding hydroxy-compounds reveals some expected as well as some unexpected differences. According to the data of Loofbourow, Stimson and Hart,[4] uracil at pH 7 shows an absorption maximum at 2580 Å. with a molecular extinction coefficient (ϵ) of 10,600. As might be anticipated from the increased mass of sulfur, the spectrum of 2-thiouracil at pH 7.0 (Fig. 1) shows a shift in the wave length of maximum absorption to 2740 Å., with an ϵ value only slightly higher than for uracil. The replacement of both oxygens by sulfur, as in 2,4-dithiouracil (Fig. 2) results in a shift of the wave length of maximum absorption still further toward the visible range, and also the introduction of a second, weaker band at 3600 Å. The molecular extinction coefficient of this compound for the main band is much higher than for 2-thiouracil. In contrast to these two compounds, the spectrum of 4-thiouracil (Fig. 3) is quite unusual. The pyrimidine band has been shifted so far toward the visible that in the usual range of pyrimidine absorption (2600–2800 Å.) there occurs a minimum rather than a maximum in the spectral distribution curve. The peak occurs at

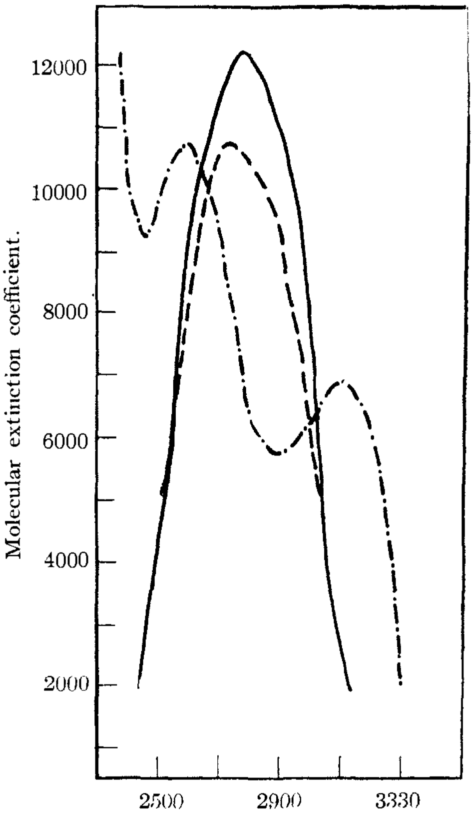

Fig. 1.—Absorption spectra of 2-thiouracil: —— at pH 1.0; - - - at pH 7.0; –·–·– at pH 11.0.

---

(1) Presented before the Division of Organic Chemistry at the 109th meeting of the American Chemical Society, Atlantic City, N. J., April, 1946.

(2) Miller, Roblin and Astwood, THIS JOURNAL, **67**, 2201 (1945).

(3) Astwood, Bissell and Hughes, *Endocrinology*, **37**, 456 (1945).

(4) Loofbourow, Stimson and Hart, THIS JOURNAL, **65**, 148 (1943).

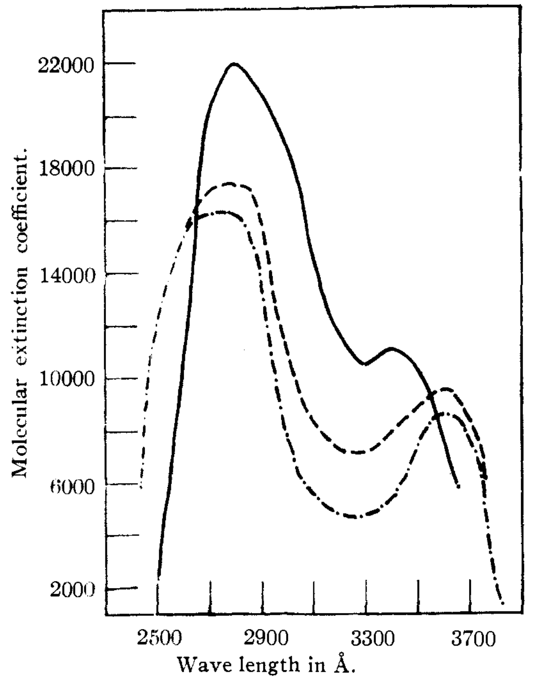

Fig. 2.—Absorption spectra of 2,4-dithiouracil: —— at *p*H 1.0; - - - at *p*H 7.0; -·-·- at *p*H 11.0.

3270 Å. and its ε value is almost as high as for 2,4-dithiouracil.

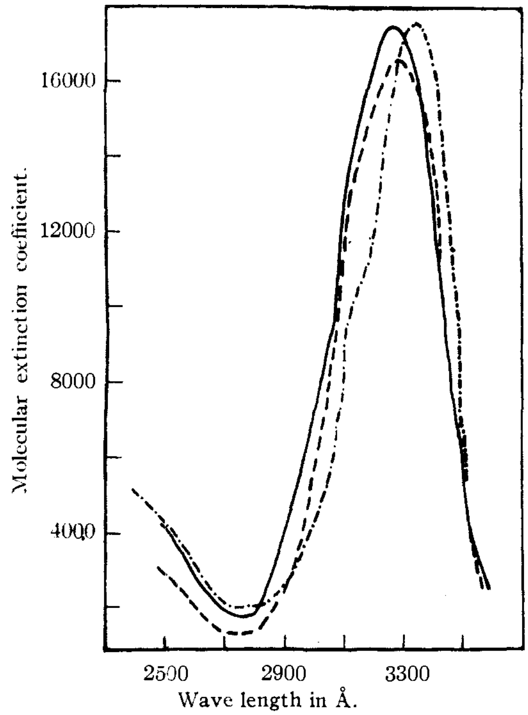

Fig. 3.—Absorption spectra of 4-thiouracil: —— at *p*H 1.0; - - - at *p*H 7.0; -·-·- at *p*H 11.0.

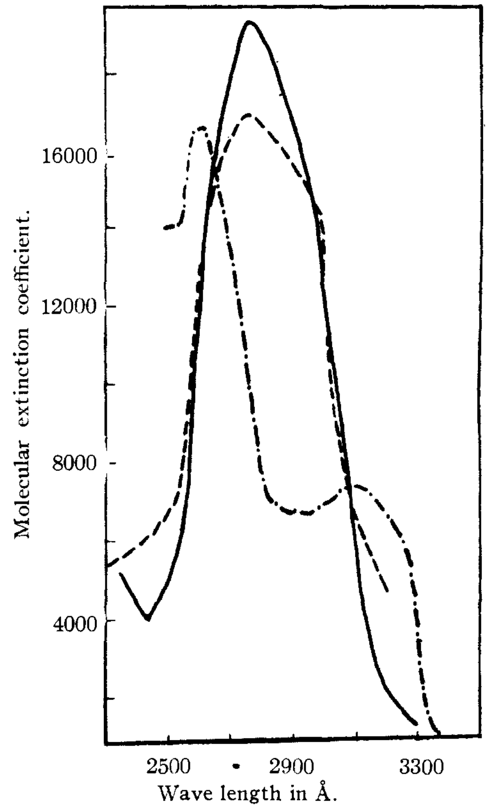

Fig. 4.—Absorption spectra of 2-thiothymine: —— at *p*H 1.0; - - - at *p*H 7.0; -·-·- at *p*H 11.0.

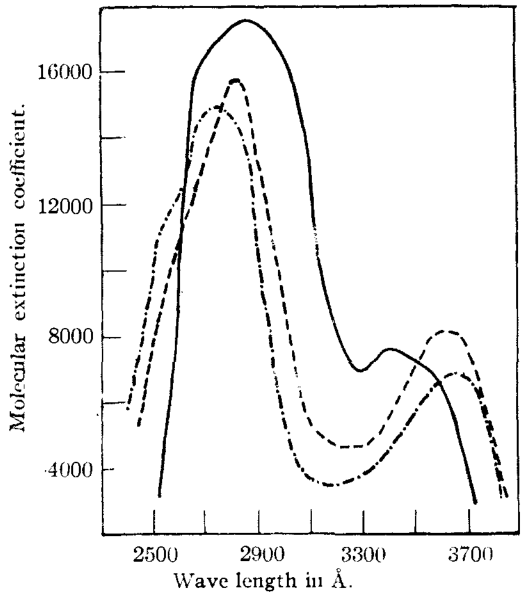

Fig. 5.—Absorption spectra of 2,4-dithiothymine: —— at *p*H 1.0; - - - at *p*H 7.0; -·-·- at *p*H 11.0.

The absorption spectra of the three thiothymines show very much the same relationship to the spectrum of thymine, as can be seen by a comparison of Figs. 4, 5 and 6 with the data for thymine reported by Stimson and Reuter,[5] $\lambda_{max.} = 2620$ Å., $\epsilon = 7800$. It will be noted,

---

(5) Stimson and Reuter, THIS JOURNAL, **67**, 847 (1945).

however, that 2-thiothymine shows an unusually high extinction coefficient, even though thymine itself has a lower ϵ than uracil. As before, the 4-thio derivative shows very little absorption in the usual pyrimidine range, although there is a slight rise in that region.

The effect of changes in *p*H value on the spectra of these thio compounds was studied. For comparison, figures for the changes in the spectra of uracil and thymine with *p*H are given.

TABLE I

	*p*H	$\lambda_{max.}$ in Å.	ϵ
Uracil[4]	3	2580	11,000
	7	2580	10,600
	11	2850	7,700
Thymine[5]	0.1 *N* HCl	2620	7,800
	0.1 *N* NaOH	2900	5,000

It has been suggested[4] that the shift of 270 Å. for uracil between *p*H 7 and 11 is due to a lactam–lactim tautomerism, and that uracil exists mainly in the diketo form in the acid range, and in both mono-enol forms at *p*H 11.

2-Thiouracil and 2-thiothymine (Figs. 1, 4), like the corresponding hydroxy derivatives, show little change in absorption between *p*H 1 and 7; at *p*H 11, however, they show not one broad band but two distinct bands, the more intense band being displaced toward the ultraviolet, the lesser band toward the visible.

The change with *p*H of the 2,4-dithio compounds (Figs. 2, 5) is not as pronounced, since two bands already exist in the acid range. As the *p*H increases from 1 to 11, the two bands become more widely separated, the second band being shifted about 200 Å. toward the visible.

The 4-thio derivatives (Figs. 3, 6), in addition to their anomalous absorption range, show much less change with *p*H than the other compounds. If one assumes that the changes in the spectra of uracil and thymine are due to an enolic shift between *p*H 7 and 11, then it is possible that the changes in 2-thiouracil and, to a lesser extent, in 2,4-dithiouracil, are due to the same causes. The spectra of 4-thiouracil and 4-thiothymine vary so little with *p*H, however, that one is inclined to suspect that one tautomeric form predominates over the entire *p*H range. In this connection it might be mentioned that 1,3-dimethyluracil, which is incapable of lactam–lactim tautomerism, still shows a variation of 70 Å. between its absorption bands in acid and alkaline solution.[4]

It is evident that the question of the tautomerism of these compounds cannot be adequately answered on the basis of their ultraviolet absorption spectra alone. A survey of the literature reveals the variety of interpretations which may be put upon spectral data. For example, Loofbourow, Stimson and Hart[4] have suggested the diketo form for uracil in the acid range because of the similarity of its spectrum with that of 1,3-dimethyluracil, which can exist only in the diketo form. Austin,[6] on the other hand, has reported that whereas uracil and 3-methyluracil have the same absorption maxima, the bands of 1-methyluracil and 1,3-dimethyluracil are displaced 60 Å. toward the visible, and that this difference is significant indication that uracil exists in the 2-hydroxy-4-keto- form rather than the diketo form.

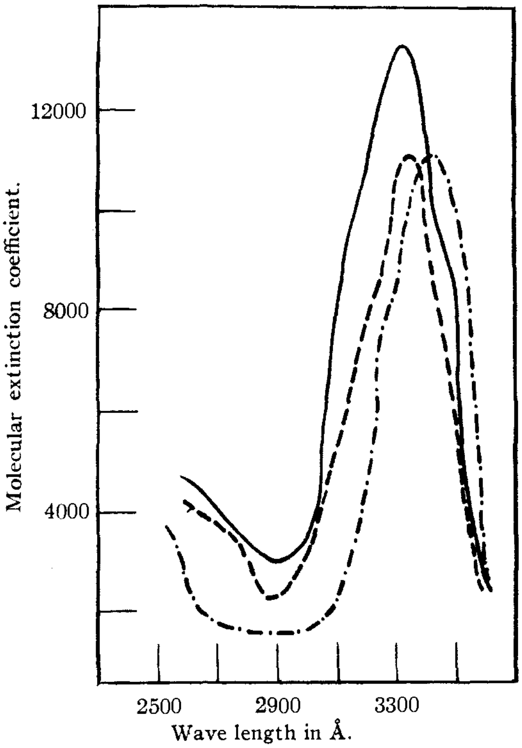

Fig. 6.—Absorption spectra of 4-thiothymine: —— at *p*H 1.0; - - - at *p*H 7.0; –·–·– at *p*H 11.0.

To establish conclusively the predominance of one or the other tautomeric form by means of ultraviolet absorption spectra, it would be necessary to have available compounds, representing the various tautomeric modifications, in which the labile hydrogen atoms have been replaced by alkyl groups.

Experimental

Materials.—All the compounds investigated were prepared in this Laboratory. Methods for the synthesis of these thiopyrimidines were reported by Wheeler and associates.[7,8]

Method.—The absorption spectra were determined using a Bausch & Lomb spectrograph, cuvette 1 cm., concentration 5–10 mg./liter. A Sørensen glycine–sodium hydroxide buffer, a Sørensen phosphate buffer[9] and 0.1 *N* hydrochloric acid were used to obtain solutions of *p*H values 11, 7 and 1, respectively.

---

(6) Austin, THIS JOURNAL, **56**, 2141 (1934).

(7) Wheeler and Liddle, *Am. Chem. J.*, **40**, 547 (1908).

(8) Wheeler and McFarland, *ibid.*, **43**, 19 (1910).

(9) W. M. Clark, "The Determination of Hydrogen Ions," Williams and Wilkins Co., Baltimore, Md., 1938, pp. 206, 210.

Summary

1. The ultraviolet absorption spectra of 2-thiouracil, 2,4-dithiouracil, 4-thiouracil, 2-thiothymine, 2,4-dithiothymine and 4-thiothymine are given at *p*H values 1, 7 and 11.

2. Comparison is made between the spectra of these thio derivatives with the spectra of uracil and thymine and the anomalous position of the absorption bands of 4-thiouracil and 4-thiothymine is discussed.

TUCKAHOE, N. Y. RECEIVED JUNE 19, 1946

[CONTRIBUTION FROM THE PACIFIC EXPERIMENT STATION, BUREAU OF MINES, UNITED STATES DEPARTMENT OF THE INTERIOR]

Low-Temperature Heat Capacities and High-Temperature Heat Contents of $Al_2O_3·3H_2O$ and $Al_2O_3·H_2O$[1]

BY C. HOWARD SHOMATE[2] AND OSCAR A. COOK[2]

One of the recent programs of the Pacific Experiment Station of the Bureau of Mines has been the measurement of thermodynamic properties of aluminum compounds of interest in certain processes for extracting alumina from clay and alunite. Several previous papers[3] on this subject have been published. The present paper, the concluding member of this series, presents low-temperature heat-capacity and high-temperature heat-content data for two crystalline hydrates of alumina.

Materials

The tri-hydrate of alumina was prepared in this Laboratory by A. E. Salo[4] by dissolving aluminum wire in 0.2 *N* potassium hydroxide. The precipitated crystals of $Al_2O_3·3H_2O$ were washed with 4 *N* hydrochloric acid to remove iron resulting from the original impurity in the aluminum wire, washed finally with distilled water, and dried at 140°. Analysis was made by fusion with sodium bisulfate and precipitation by the 8-hydroxyquinoline method. Analysis gave 65.62% aluminum oxide (theoretical 65.35%). The sample used in the low-temperature heat capacity measurements weighed 86.36 g., and the high-temperature heat content sample weighed 4.129 g.

The mono-hydrate of alumina was prepared by heating the tri-hydrate at 220° for three days. The aluminum oxide content was found to be 3% too high, so the required amount of water was added and the mixture stabilized by heating at 80° for sixteen hours in an evacuated flask. Since the product was not readily fusible in sodium bisulfate nor satisfactorily soluble in any ordinary reagent, analysis was made by igniting to aluminum oxide in a platinum crucible over an oxygen–natural gas flame. Analysis gave 85.00% aluminum oxide (theoretical 84.90%). A 92.30-g. sample was used in the low-temperature measurements, and a 4.228-g. sample for the measurements at high temperatures.

X-Ray examinations showed the tri-hydrate to have the structure of gibbsite ($Al_2O_3·3H_2O$), and the mono-hydrate a structure similar to bayerite ($Al_2O_3·3H_2O$), differing from the latter principally in the intensities of the lines.

Low-Temperature Heat Capacities

The method and apparatus used in the low temperature heat-capacity measurements were described previously.[5,6] The experimental results, expressed in defined calories (1 calorie = 4.1833 int. joules),[7] are listed in Table I and shown graphically in Fig. 1. The values of the heat capacities at 298.16°K., read from a smooth curve through the experimental points, are also included in Table I. The molecular weights are in accordance with the 1941 International Atomic Weights.

TABLE I

MOLAL HEAT CAPACITIES

$Al_2O_3·3H_2O$ Mol. wt., 155.99		$Al_2O_3·H_2O$ Mol. wt., 119.96	
T, °K.	C_p, cal./deg.	*T*, °K.	C_p, cal./deg.
52.8	3.213	52.7	2.179
56.9	3.708	56.9	2.528
61.1	4.259	61.0	2.914
65.2	4.816	65.2	3.307
69.8	5.506	69.3	3.733
74.0	6.164	73.8	4.216
78.4	6.854	78.3	4.709
85.3	8.002	84.8	5.441
94.8	9.677	94.7	6.630
104.7	11.54	104.5	7.865
115.3	13.57	115.1	9.228
125.3	15.51	125.3	10.58
135.1	17.47	135.0	11.92
145.4	19.50	145.6	13.31
155.3	21.39	155.4	14.62
165.5	23.38	165.5	15.98
175.7	25.35	175.4	17.33
185.7	27.15	185.6	18.66
196.0	28.97	195.8	20.00
205.9	30.72	205.9	21.32
216.2	32.55	216.2	22.57
226.2	34.09	226.2	23.74
235.8	35.55	235.9	24.83
246.2	37.19	246.2	25.99
256.2	38.72	256.4	27.14
266.3	40.17	266.1	28.26
276.3	41.63	276.2	29.27
286.3	42.94	286.2	30.17
296.5	44.27	296.4	31.13
(298.16)	(44.49)	(298.16)	(31.37)

---

[1] Published by permssion of the Director, Bureau of Mines, U. S. Department of the Interior. Not copyrighted.

[2] Chemist, Pacific Experiment Station, Bureau of Mines.

[3] (a) Shomate and Naylor, THIS JOURNAL, **67**, 72 (1945); (b) Young, *ibid.*, **67**, 257 (1945); (c) Shomate, *ibid.*, **67**, 765 (1945); (d) Young, *ibid.*, **67**, 851 (1945); (e) Shomate, *ibid.*, **67**, 1096 (1945).

[4] Metallurgist, Pacific Experiment Station, Bureau of Mines.

[5] Kelley, THIS JOURNAL, **63**, 1137 (1941).

[6] Shomate and Kelley, *ibid.*, **66**, 1490 (1944).

[7] Mueller and Rossini, *Am. J. Physics*, **12**, 1 (1944).