gave consistent and reliable results on the sodium nitrate solubility curve was our improved microtechnique. Hence points obtained by this technique only were used in the plotting of this curve. For the sake of comparison, Early and Lowry's solubility curves for this system are included as dotted lines.

Figure 1 shows the solubility curves of ammonium nitrate to be slightly higher than Early and Lowry's[1] curves for the corresponding parts. The solubility curve of sodium nitrate is most interesting as it runs considerably above that of these authors. Although the curve was obtained by the method of least squares using only the data from our improved microtechnique, actually we determined seventeen other points by the other methods previously mentioned in this paper which, although scattered, support the data actually used rather well. Adding the distances above the curve and parallel to the temperature axis together and subtracting the sum of the distances below the curve from these and then dividing by the number of points, we find that the average deviation per point is 0.05° under the curve.

Summary

The system ammonium nitrate–sodium nitrate has been redetermined over the range 0–35% sodium nitrate. The solubility curves for the two forms of ammonium nitrate are linear while that for sodium nitrate is distinctly parabolic. An improved microtechnique has been developed that gives good results except near the eutectic point where all melting point methods are at their worst.

The eutectic temperature was found to be 121.2 ± 0.2° and the transition temperature to be 126.2 ± 0.2° by the cooling curve method which is most reliable for these temperatures. Using these points and applying the method of least squares to the curves, the eutectic composition was found to be 21.0 ± 0.1% sodium nitrate and the transition composition to be 18.0 ± 0.1%.

The method of purification of the sodium nitrate, the length of time of drying the materials and the size of the sample appear to have no effect on the temperature of disappearance of the last crystal on heating the melt.

BOSTON, MASS. RECEIVED DECEMBER 28, 1943

[CONTRIBUTION NO. 495 FROM THE DEPARTMENT OF CHEMISTRY OF THE UNIVERSITY OF PITTSBURGH]

The Heat of Acetylation of α- and β-D-Glucose from the Heats of Combustion of the Pentaacetates[1]

BY THEODORE H. CLARKE AND G. STEGEMAN

In the investigation of the thermal properties of sugars which is being carried out in this Laboratory, some interest developed in the determination of the heats of combustion of α- and β-D-glucose pentaacetate, wherein the isomeric structural differences involve a much heavier radical than is present in the case of α- and β-D-glucose.

In this paper are presented the results of the combustion measurements, the specific heats of the compounds at 25°, and an approximate value of the density (used to correct the weight to the vacuum basis). From these and other data the heats of formation and the heats of acetylation were calculated.

A general description of the calorimetric equipment used in these determinations has been reported.[2] The energy equivalent of the calorimeter was determined by the use of benzoic acid obtained from the National Bureau of Standards. Sample No. 39e having a certified heat of combustion of 26428.4 international joules[3] or 6317.60 conventional cal. per gram mass *in vacuo*, was employed under the following conditions which are recommended.

The reaction is referred to 25°; the oxygen pressure should be 30 atmospheres; the weight of the sample should be 3 g. per liter of bomb volume; the bomb is to contain 3 g. of water per liter of bomb volume.

These conditions were adhered to with sufficient exactness so that no additional corrections were necessary.

The results of six determinations produced a value: $C_{\mathrm{Sys}} = 2772.95 \pm 0.2$ cal./°C. (precision error, ± 0.009%). The value was redetermined upon the completion of the measurements and no change was detected. The bomb has a volume of 0.516 liter and contained 1.5 ml. of water and oxygen at an initial pressure of 30 atmospheres. The initial temperature of the calorimeter was adjusted to 22.6 to 22.7° and the average temperature rise was 3.15°. The temperature rise in degrees was determined from the resistance *vs.* time readings as outlined by Dickinson.[4] Calculations of the heats of combustion were made following the method outlined by Dickinson which gives the heat of combustion at the final temperature reached in the combustion process. The necessary specific heats were employed to correct the values to 25°.

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[1] This investigation was made possible by a grant from the Buhl Foundation, Pittsburgh, Pa.

[2] Clarke and Stegeman, THIS JOURNAL, **61**, 1726 (1939).

[3] Jessup, *J. Research Natl. Bur. Standards*, **29**, 247 (1942).

[4] Dickinson, *Bur. Standards Bull.*, **11**, 189–257 (1915).

TABLE I

PURIFICATION, DRYING AND COMBUSTION DATA

Sample	Re-crystn. abs. alc.	Drying agent	Drying t for final 24 hr. at 0.01 mm.	True mass, g.	Δt cor., °C.	Heat evolved by sample, cal.	Cor. to 25°C., cal./g.	$-\Delta U_B/M$ 25°C., cal./g.	Δ
				α-D-Glucose pentaacetate					
1a	6	$CaCl_2$	40	1.96171	3.1567	8647.91	0.26	4408.09	1.49
1a	6	$CaCl_2$	40	1.97049	3.1676	8680.81	.26	4405.14	−1.46
1b	7	$CaCl_2$	40	1.95958	3.1520	8633.19	.23	4405.40	−1.20
2a	5	$CaCl_2$	60	1.95608	3.1484	8620.61	.21	4406.87	0.27
2a	5	$CaCl_2$	60	1.96266	3.1566	8651.03	.21	4407.59	.99
2b	6	Drierite	60	1.95811	3.1446	8628.93	.23	4406.53	− .07
				Mean = 4406.60 ± 0.97 cal./g. (Precision error = ±0.022%)					
				β-D-Glucose pentaacetate					
1a	4	$CaCl_2$	40	1.95218	3.1424	8622.26	0.24	4416.49	−0.63
1a	4	$CaCl_2$	40	1.95933	3.1628	8658.15	.26	4418.67	1.55
1b	5	$CaCl_2$	40	1.96374	3.1656	8678.26	.27	4418.98	1.86
2a	5	$CaCl_2$	60	1.96270	3.1601	8667.17	.24	4415.70	−1.42
2a	5	$CaCl_2$	60	1.96554	3.1646	8680.17	.24	4415.93	−1.19
2b	6	$CaCl_2$	60	1.96794	3.1731	8696.08	.24	4418.63	1.51
2b	6	Drierite	60	1.96290	3.1610	8667.57	.23	4415.46	−1.66
				Mean = 4417.12 ± 1.19 cal./g. (Precision error = ±0.027%)					

$$\text{Heat Evolved by the Sample} = (C_{\text{Sys.}} + C_{\text{Sample}})\Delta t - q_{\text{Fe}} - q_{\text{HNO}_3} - q_{\text{Igniter}}$$
$$(\text{For } \alpha\text{-isomer, sample 1a}) = 8753.37 + 1.86 - 17.60 - 0.58 - 89.14$$

The defined calorie = 4.1833 int. j.[5] is the unit of energy employed. Deviations are recorded in terms of the precision error.[6] The optical rotations were determined at 20° using the proper dichromate filter to give D line illumination. A two-decimeter tube was employed. Melting points, obtained by the use of an Anschütz thermometer, were determined in the usual manner.[7]

Treatment of Samples.—Two samples each of α- and β-D-glucose pentaacetate were obtained from Eastman Kodak Co. and were given the following treatment:

The original samples were dissolved in and recrystallized from four to six times from absolute alcohol. It was the purpose of the treatment to produce a material whose heat of combustion would show no change following a recrystallization. For example, product 1a (Table I) was recrystallized six times. The sample was then divided and half of it was given an additional crystallization from absolute alcohol. This product is designated as 1b. The two products were then dried *in vacuo* over calcium chloride and drierite, respectively, at room temperature until a constant weight was obtained. The time interval was about twenty-four hours between weighings. An additional drying for twenty-four hours at 40° and in the case of a second sample at 60° and 0.01 mm. resulted in no further loss in weight.

The values for the optical rotations and melting points were reported by Hudson and Dale.[8] For the α-isomer the value of the optical rotation is listed as 96.7° for a 5.302% solution in benzene and a melting point of 112–113°. For the β-isomer the values listed are 2.2° for a 5.061% solution in benzene and a melting point of 132°. The products prepared in this Laboratory gave rotation values of 96.7° and 2.1°, and melting points of 112.7° and 132.0° for the α and β isomers, respectively.

Data.—The combustion data for the prepared products were obtained in the following manner. The first two determinations indicated as 1a were made on the sample which had received the smaller number of recrystallizations. The third and in one case a fourth determination indicated as 1b were made on the sample which had received an additional crystallization. The combustion data and the number of recrystallizations of each product are listed in Table I.

The derived results and certain other data are presented in Table II. The derived values are $-\Delta U_B$, the isothermal heat of combustion in kcal. per mole under the bomb conditions; $-\Delta U_R$, the heat of combustion in kcal. per mole with the reactants and products at a pressure of one atmosphere; $-\Delta H_R$, the isobaric heat of combustion at one atmosphere; and $-\Delta H^0_{f298.16}$, the heat of formation in kcal. per mole at one atmosphere and constant pressure. It was assumed that the product burned was completely oxidized to carbon dioxide and water. The heats of formation of the pentaacetates were calculated from the heats of combustion employing the values, −68,318.1[9] and −94,029.8[10] cal. per mole

TABLE II

DERIVED RESULTS AND OTHER DATA AT 25°

	α-isomer	β-isomer
Mol. wt.	390.336	390.336
$-\Delta U_B$, kcal./mole	1720.06 ± 0.38	1724.16 ± 0.46
$-\Delta U_R(= -\Delta H_R)$, kcal./mole	1718.62	1722.63
$-\Delta H^0_{f298.16}$, kcal./mole	537.36	533.25
$-d\Delta H/dt$, cal./°C./mole	109	108
C_p, cal./°C./g.	0.301	0.305

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(5) Rossini and Deming, *J. Wash. Acad. Sci.*, **29**, No. 10, Oct. 15 (1939).

(6) Rossini, *Chem. Rev.*, **18**, 252 (1936).

(7) Bates, "Polarimetry, Saccharimetry and the Sugars," Circular of the National Bureau of Standards C440, page 547, or U. S. Pharmacopoeia XI, pp. 455–457.

(8) Hudson and Dale, THIS JOURNAL, **37**, 1264 (1915).

(9) Rossini, *J. Research Natl. Bur. Standards*, **22**, 407 (1939).

(10) Rossini and Jessup, *ibid.*, **21**, 1491 (1938).

for the heats of formation of liquid water and gaseous carbon dioxide.

Discussion of Results

Karrer and Fioroni[11] reported a value for the heat of combustion of glucose pentaacetate. On the basis of the 1942 value for the heat of combustion of benzoic acid and in units of the defined calorie, Karrer and Fioroni's value is 1722.4 kcal. per mole for $-\Delta U_B$. It is not clearly stated that these authors were dealing with α-D-glucose pentaacetate, but this is assumed to be the case. No value has been reported previously for β-D-glucose pentaacetate.

The following equation indicates that the acetylation of glucose with acetic acid is an endothermic reaction for both isomers.

D-glucose (s) + 5 acetic acid (l) = D-glucose pentaacetate (s) + $5H_2O$ (l)
α-isomer, ΔH = 6.32 kcal./mole
β-isomer, ΔH = 8.92 kcal./mole

The use of acetic anhydride, however, produces the acetylation with a considerable evolution of energy.

D-Glucose (s) + 5/2 acetic anhydride (l) = D-glucose pentaacetate (s) + $5/2H_2O$ (l)
α-isomer, ΔH = −30.88 kcal./mole
β-isomer, ΔH = −28.29 kcal./mole

The rather large exothermic value for ΔH in the latter reaction suggests that the use of the acetic anhydride will produce a favorable yield of the pentaacetates since the ΔF value will probably have the same order of magnitude. This is in harmony with experience.

In making the indicated calculations the following thermal values were employed. The values employed have been recalculated by the authors in terms of the defined calorie and based on the 1942 value for the heat of combustion of benzoic acid. This places all values secured by and utilized by the authors on a comparable basis. The heat of combustion of α-D-glucose is given as $-\Delta H_R$ = 669.81 kcal./mole.[12] Employing the values for carbon dioxide and water previously mentioned, the heat of formation was calculated to be −304.28 kcal./mole. The authors used the value obtained by Huffman and Fox for the heat of combustion of β-D-glucose of $-\Delta H_R$ = 671.33 kcal. per mole. Huffman and Fox state that this value is not of the highest precision, but it is probably very near to the true value. While its agreement with a value calculated from solution data is good, Huffman and Fox point out that the same type of irregularities may have complicated both measurements. The calculated heat of formation from the data of Huffman and Fox for β-D-glucose is −302.76 kcal. per mole. The heat of formation of acetic acid was calculated to be −116.20 kcal./mole, from the heat of combustion reported by Schjanberg.[13] The heat of formation of acetic anhydride was calculated to be −149.20 kcal./mole from the heat of combustion by Thomsen.[14]

While the conversion of β- to α-D-glucose is accompanied by a heat evolution of about 1.5 kcal./mole,[12] the conversion of β- to α-D-glucose pentaacetate is accompanied by a heat evolution of about 4.0 kcal./mole. How much of this heat effect is to be attributed to the β- to α-shift about the saccharide carbon cannot be estimated at present on the basis of information now available

Summary

Values for the heats of combustion at 25° and 1 atm. pressure are given for α- and β-D-glucose pentaacetate. The value for the specific heat at 25° is also reported. From these and other data, the heats of formation and heats of acetylation were computed.

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[11] Karrer and Fioroni, *Helv. Chim. Acta*, **6**, 396 (1923).

[12] Huffman and Fox, THIS JOURNAL, **60**, 1403 (1938).

[13] Schjanberg, *Svensk. Kem. Tids*, **44**, 227–231 (1932).

[14] Thomsen, *Z. physik. Chem.*, **52**, 343 (1905).

PITTSBURGH, PA. RECEIVED MAY 14, 1943

[CONTRIBUTION FROM THE CHEMICAL LABORATORY OF NORTHWESTERN UNIVERSITY]

The Mode of Action of Sulfonamides

BY IRVING M. KLOTZ

Since the discovery that *p*-aminobenzoic acid inhibits the action of sulfonamides,[1] attempts have been made to correlate the activity of these drugs with their structure. Some investigators[2] soon observed that the sulfonamide potency seemed to be directly related to the amount of the ionic form of the drug. The work of Cowles[3] indicated that sulfonamide activity increased as the acid *pK* of the drug decreased toward 7, but that further increase in acid strength resulted in diminished activity. A more comprehensive study by Bell and Roblin[4] has shown that bacteriostatic activity is a parabolic type of function of acid strength of the sulfonamide, with maximum potency being obtained, in a solution of *p*H 7, with drugs whose *pK*'s are approximately 6.7.

Inasmuch as current opinion attributes sulfonamide activity to the blocking of an enzyme system,[1] it was considered advisable to examine

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[1] Woods, *Brit. J. Exptl. Path.*, **21**, 74 (1940).

[2] Fox and Rose, *Proc. Soc. Exptl. Biol. Med.*, **50**, 142 (1942); Schmelkes, Wyss, Marks, Ludwig and Strandskov, *ibid.*, **50**, 145 (1942).

[3] Cowles, *Yale J. Biol. Med.*, **14**, 599 (1942).

[4] Bell and Roblin, THIS JOURNAL, **64**, 2905 (1942).