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Thermochimica Acta 273 (1996) 11–15

thermochimica
acta

The enthalpy of isomerization of peroxynitrite to nitrate

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Received 10 July 1995; accepted 31 July 1995

Abstract

The enthalpy of isomerization of peroxynitrite [oxoperoxonitrate (1-)] to nitrate (ΔH°_{297}, $O{=}NOO^- \rightarrow NO_3^-$) was measured by solution calorimetry. Oxoperoxonitrate (1-) synthesized from both hydroxylamine and azide yielded enthalpies of isomerization of -163 ± 4 kJ mol^{-1} at 24°C. Since $\Delta_f H(NO)_3^-) = -205$ kJ mol^{-1}, $\Delta_f H(O{=}NOO^-) = -42 \pm 4$ kJ mol^{-1}.

Keywords: Enthalpy of formation; Enthalpy of isomerization; Oxoperoxonitrate (1-); Peroxynitrite; Solution calorimetry

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1. Introduction

Peroxynitrite [oxoperoxonitrate (1-), $O{=}NOO^-$] is a toxic compound which is formed in vivo near activated macrophages and neutrophils from the near diffusion-controlled reaction of superoxide and nitrogen monoxide [1–3]. Oxoperoxonitrate (1-) participates in reactions which are cytotoxic, such as the initiation of lipid peroxidation [4], oxidation of sulfhydryls [5] and methionine [6], and the hydroxylation and nitration of aromatic compounds [7, 8].

Oxoperoxonitrate (1-) is relatively stable in alkaline solutions (pH > 12) [9, 10]. However, peroxynitrous acid [hydrogen oxoperoxonitrate, O=NOOH], which has a pK_a of 6.8 [11], is unstable and decays to nitrate at a rate of 1.3 s^{-1} at 25°C, Eq. (1)

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SSDI 0040-6031(95)02690-8

[11, 12]. At alkaline pH, a much slower reaction occurs, Eq. (2) [9].

$$O{=}NOO^- + H^+ \rightarrow O{=}NOOH \rightarrow NO_3^- + H^+ \quad (1)$$

$$2O{=}NOO^- \rightarrow 2NO_2^- + O_2 \quad (2)$$

In 1962 the heat of isomerization of oxoperoxonitrate(1-) to nitrate was determined as 39 ± 2 kcal mol^{-1} (163 ± 8 kJ mol^{-1}) at 1°C [13]. In this study, the sum of the enthalpy of oxoperoxonitrate(1-) formation *and* of its decay was measured. The plot of temperature versus time was biphasic: the first rapid increase was caused by the formation of oxoperoxonitrate(1-) from nitrite and hydrogen peroxide and the second slow increase was from the isomerization to nitrate. Given the scatter of the data points, it is difficult to judge where the first process ended, and the second process started. In 1994 the enthalpy of the decay of peroxynitrite was measured by DSC, and a value of -37.8 kcal mol^{-1}, or 158 kJ mol^{-1}, was reported [14]. However, at higher temperatures oxoperoxonitrate(1-) may not decay solely to nitrate, but also form nitrite and oxygen, see Eq. (2), above.

Since the Gibbs energy of formation and other thermodynamic data of oxoperoxonitrate(1-) [11] depend on the enthalpy of isomerization, it is important to establish a reliable value for this process.

2. Experimental

Two different methods were used to prepare oxoperoxonitrate(1-) solutions. In the first method, oxoperoxonitrate(1-) was prepared by ozonolysis [15, 16] of an aqueous solution of sodium azide and 0.10 M sodium hydroxide. The concentration of sodium azide and time of ozonation were optimized to obtain solutions containing a high concentration of oxoperoxonitrate(1-) and a minimal amount of by-products. In the second method, oxoperoxonitrate(1-) was prepared by autoxidation of 50 mM hydroxylamine in an aqueous solution containing 0.5 M sodium hydroxide and 100 μM dtpa [17]. Chemicals used were of reagent grade purity. Water was purified by reverse osmosis, deionization, and filtration (Osmonics, Inc.). Oxoperoxonitrate(1-) solutions were freshly prepared prior to each experiment. The concentration of the stock oxoperoxonitrate(1-) solution was determined from its absorbance at 302 nm ($\varepsilon = 1670$ M^{-1} cm^{-1}) [18] with a Beckman DU-7HS spectrophotometer.

A Parr® 1451 solution calorimeter was used to measure the heat of isomerization of oxoperoxonitrate(1-). This instrument enables temperature differences of 10^{-3} K to be measured. The temperature of the solutions was brought to 24°C with a water bath prior to each experiment; this considerably reduced the equilibration time in the calorimeter.

3. Results and discussion

Upon addition of acid to an alkaline solution of oxoperoxonitrite(1-), there are, in principle, three reactions in which heat is either generated or consumed: (1) the formation of water from H^+ and OH^-; (2) the protonation of $O{=}NOO^-$; and (3) the isomerization of O=NOOH to NO_3^- and H^+. Since the pK_a of O=NOOH appears to be temperature-independent, the enthalpy of the second process is zero [11]. In order to correct for the first process, oxoperoxonitrate(1-) was made to decay by two different methods. The first type of blank solution was prepared by titration of 5.00 mL of oxoperoxonitrate(1-) solution containing 0.100 M HCl to a pH between 5 and 7. After titration with HCl, the same volume of 0.100 M NaOH was added in order to arrive at the pH of the original solution. The second type of blank solution was prepared by decaying 5.00 ml of oxoperoxonitrate(1-) solution in a water bath at 50°C for 2h, after which no oxoperoxonitrate(1-) could be detected spectroscopically. The alkalinity of these solutions was the same as that of the oxoperoxonitrate(1-)-containing solutions. The blank solutions were transferred to the calorimeter which contained 100 ml of dilute (0.005–0.010 M) hydrochloric acid. Upon equilibration to 24°C, the solutions were mixed and the temperature change was recorded. Next, 5 ml of a oxoperoxonitrate(1-) solution (10–40 mM concentration) and a volume of water equivalent to the volume of hydrochloric acid and sodium hydroxide solutions added to the first type of blank solution were mixed with dilute hydrochloric acid as previously described.

The heat, Q, was calculated using Eq. (3). It was assumed that the specific heats of both the oxoperoxonitrate(1-) and the decayed oxoperoxonitrate(1-) solutions are the same as that of water at 24°C, that is: $c_{H_2O} = 0.99838$ cal g^{-1} °C^{-1} [19].

$$Q = \Delta t[e' + m_{HCl})(c_{HCl}) + (m_{ONOO^-})(c_{ONOO^-}) + (m_{H_2O})(c_{H_2O})] \quad (3)$$

where Q is the energy change in cal, Δt is the temperature change in °C, e' is the energy equivalent of the empty calorimeter in cal °C^{-1}, c_{HCl}, c_{ONOO^-}, and c_{H_2O} are the specific heats of hydrochloric acid, oxoperoxonitrate(1-) and water at 24°C in cal g^{-1} °C^{-1}, and m_{HCl}, m_{ONOO^-}, and m_{H_2O} are the masses in g of hydrochloric acid, the oxoperoxonitrate(1-) solution and water, respectively.

The heats of the oxoperoxonitrate(1-) solutions and the blank solutions determined in these experiments are indicated as $Q_{ONOO^- + OH^-}$ and Q_{OH^-}, respectively. The heat of isomerization of oxoperoxonitrate(1-) to nitrate is then given by Eq. (4).

$$\Delta H^\circ_{297}, (O{=}NOO^- \rightarrow NO_3^-) = -(Q_{O=NOO^- + OH^-} - Q_{OH^-})/n_{O=NOO^-} \quad (4)$$

where ΔH°_{297}, $(O{=}NOO^- \rightarrow NO_3^-)$ is the enthalpy of isomerization of oxoperoxonitrate(1-) to nitrate and $n_{O=NOO^-}$ is the number of moles of oxoperoxonitrate(1-) added to the hydrochloric acid.

The results are shown in Fig. 1 for the two different preparations of oxoperoxonitrate(1-). Every point is the average of six determinations. The error bars indicate standard deviations, which for six determinations amounts to the confidence interval at the 95% confidence level. The experimental value for ΔH°_{297}, $(O{=}NOO^- \rightarrow NO_3^-)$ is -161 kJ mol^{-1} (-39 ± 1 kcal mol^{-1}), as determined from the slope of the line in Fig. 1. This result is in complete agreement with the value determined by Ray [13] in 1962 as -39 ± 2 kcal mol^{-1}, and the -37.8 kcal mol^{-1} determined recently by Bohle et al. [14] by differential scanning calorimetry. As expected, the enthalpy of isomerization of oxoperoxonitrate(1-) to nitrate is not affected by the method used to prepare the oxoperoxonitrate(1-) solutions. Given a $\Delta_f H^\circ$ of -205 kJ mol^{-1} for the nitrate anion [20], $\Delta_f H^\circ(O{=}NOO^-) = -42 \pm 4$ kJ mol^{-1}.

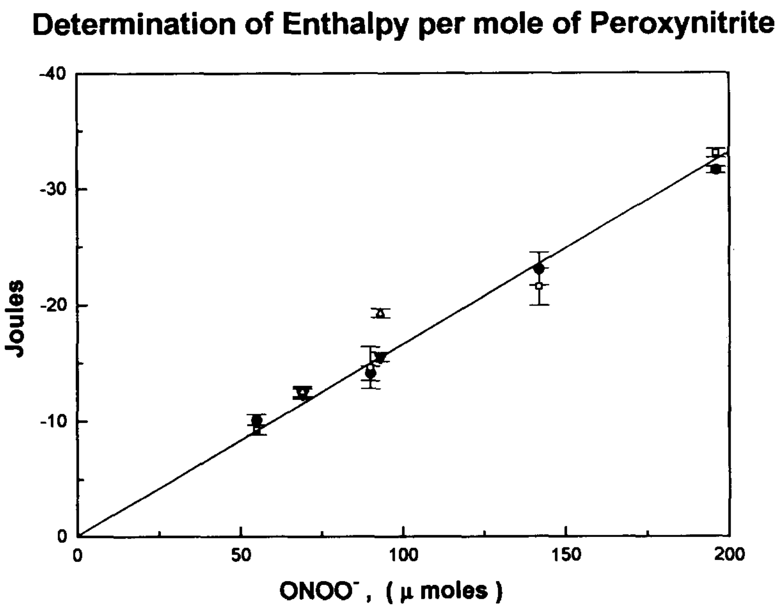

Fig. 1. Determination of the molar enthalpy of isomerization of peroxynitrite. ●, $Q_{\mathrm{ONOO^-+OH^-}}$ from azide — $Q_{\mathrm{OH^-}}$ from blank decayed by acid; □, $Q_{\mathrm{ONOO^-+OH^-}}$ from azide — $Q_{\mathrm{OH^-}}$ from blank decayed at 50 °C; ▼, $Q_{\mathrm{ONOO^-+OH^-}}$ from hydroxylamine — $Q_{\mathrm{OH^-}}$ from blank decayed by acid; Δ, $Q_{\mathrm{ONOO^-+OH^-}}$ from hydroxylamine — $Q_{\mathrm{OH^-}}$ from blank decayed at 50 °C.

Acknowledgement

We thank the Department of Chemistry and Physics of Southeastern Louisiana University for the loan of the Parr® 1451 solution calorimeter. This work was supported by grants from the National Institutes of Health (GM 48829) and the Council for Tobacco Research-USA, Inc.

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