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Published in:
Journal of The Electrochemical Society

Link to article, DOI:
10.1149/1.2044197

Publication date:
1995

Document Version
Publisher's PDF, also known as Version of record

Link back to DTU Orbit

Citation (APA):

Technical University of Denmark

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Electrochemical Investigation of the Catalytical Processes in Sulfuric Acid Production

N. J. Bjerum,* I. M. Petrushina, and R. W. Berg

Materials Science Group, Chemistry Departement A, The Technical University of Denmark, DK-2800 Lyngby, Denmark

ABSTRACT

The electrochemical behavior of molten K_2S_2O_7 and its mixtures with V_2O_5 [2-20 mole percent (m/o) V_2O_5] was studied at 440°C in argon, using cyclic voltammetry on a gold electrode. The effect of the addition of sulfate and lithium ions on the electrochemical processes in the molten potassium pyrosulfate was also investigated. The potential window for pure K_2S_2O_7 was estimated as 2.1 V, being limited by the SO_4~2^- oxidation and reduction. The oxidation of SO_4~2^- to oxygen is vs. Ag/AgCl. For all studied compositions, the first reduction stage is a one-electron reaction. The second reduction stage (V(IV) → V(III)), starting at 0.7-0.8 V vs. Ag/AgCl, is reversible for V_2O_5 concentrations lower than 5 m/o and at potential scan rates less than 200 mV/s. For all studied compositions, the first reduction stage is a one-electron reaction. The second reduction stage (V(IV) → V(III)), starting at 0.7-0.8 V, is irreversible and under ohmic control at all studied V_2O_5 concentrations. The presence of Li_2SO_4 causes a noticeable depolarization effect on the V(IV) reduction and the V(IV) oxidation.

**Introduction**

The molten V_2O_5-M_2S_2O_7 system (M is an alkali metal) is used as a catalytic medium for sulfur dioxide oxidation with oxygen or air, and this has been a subject of numerous publications for the last fifty years. However, the mechanism of this catalytic process is not clearly understood, the following problems remaining unsolved: (i) The number of the catalytical reaction steps involved has not been defined; two-step, three-step, and four-step mechanisms have been proposed. (ii) The structure of catalytically active V(IV) species has to be elucidated. Neither proposed mononeric, nor dimeric, trimeric, or tetrameric structure of the dioxovanadium complexes has been convincingly determined. (iii) The first unsolved problem also leads to a question concerning the possibility of the V(V) two-stage reduction (both to V(IV) and V(III) species). The precipitation of solid V(III) compounds in the liquid-gas M_2S_2O_7/V_2O_5-SO_2/O_2/SO_3/N_2 system (where M is Na, K, or Cs) has been proven by chemical and x-ray analysis. Electrochemical investigations in the V_2O_5-K_2S_2O_7 melt could neither prove nor disprove the V(IV) → V(III) reduction reaction. The nature of the promoting effect of alkali ions on the catalytic activity of the V_2O_5-K_2S_2O_7 melt is unknown. It was found in practice that the efficiency increased with the alkali atomic number (or ionic radius) and the presence of sodium cations had a promotive effect at low temperatures (close to the point of catalyst deactivation, i.e., 440°C); and that cesium ions decreased the temperature of deactivation.

Therefore, investigations of the mechanism of catalytical reactions in the molten V_2O_5-M_2S_2O_7 system are still of great significance. By using electrochemical techniques, the oxidation and the reduction parts of the complex chemical reaction can be separated; moreover, some oxidation
and reduction agents (such as O₂, SO₂, and SO₃) can be excluded, thus simplifying the investigation. As a result, useful information for the mechanism could be obtained.

The electrochemical behavior of molten potassium pyrosulfate and its mixtures with potassium sulfate and vanadium pentoxide was studied by Durand et al., Comtat et al., Frank and Winnick, Scott et al., and Dojcinovic et al. in different gas atmospheres. From these references, large disagreements can be seen between the experimental data, and the conclusions on the width of the potential window and the nature of the electrochemical processes involved. According to Durand et al., the potential window of the gold electrode in molten potassium pyrosulfate at 430°C was 0.41 V wide for the acidic melt and 0.51 V wide for the basic melt, with a difference as high as 0.52 V between acidic and basic cathodic limits. The anodic limit was assigned to gold electrode dissolution in the acidic melt, and in the basic melt to the S₂O₇²⁻ ion oxidation with oxygen evolution. The cathodic limit was assigned to a reduction of SO₄²⁻ (or SO₃²⁻). These explanations have not been tested experimentally, and the difference in the potential windows of the acidic and basic K₂S₂O₇ melts seems questionable. Frank and Winnick also explained the 1.0 V difference between the potentials of V(V) ↔ V(IV) and V(IV) ↔ V(III) potentials can exist in dimeric species dissolved in the K₂S₂O₇ melt, given the monomeric structures of molten potassium pyrosulfate through the formation of S₂O₇²⁻ solvate complexes.

\[ 2S₂O₇²⁻ \leftrightarrow S₂O₅²⁻ + SO₄²⁻ \]  

The electrochemical characteristics of molten alkali sulfates are also of interest. Alkali sulfates have been studied more extensively and carefully than the pyrosulfates. 18-22 The electrochemical behavior of solutions of V₂O₅ in molten potassium pyrosulfate was investigated by Durand et al. by using gold electrodes. Frank and Winnick have studied the electrochemistry of V₂O₅ solutions in both acidic and basic molten potassium pyrosulfate. Oxidation of SO₄²⁻ was observed at an anodic limit of 0.8-0.9 V vs. Ag/Ag⁺, whereas S₂O₇²⁻ reduction was seen as the cathodic limiting reaction at 0.25 V. However, the experiments of Frank and Winnick were done at 370-400°C, i.e., at temperatures lower than the melting point of dry K₂S₂O₇ (i.e., 418.52°C). This indicates the presence of water in the molten potassium pyrosulfate, and therefore a possible water effect on the results of the experiments should be taken into account. Thus Comtat et al. observed the same anodic limit for the basic electrolyte at a platinum and a gold electrode. Frank and Winnick also observed an approximately 0.8 V wide potential window at a gold electrode in N₂, air, SO₂/air, SO₃/O₂, nitrogen, and air atmospheres, in both acidic and basic molten potassium pyrosulfate. In short, additional electrochemical measurements should be made under dry conditions and in an inert gas atmosphere (to exclude possible oxygen effect on the potential window of molten potassium pyrosulfate). The present study was conducted to provide useful electrochemical kinetic data for molten potassium pyrosulfate solutions with a wide concentration range of V₂O₅, and to have a good understanding of the catalytic mechanism.

### Experimental

Pure and dry K₂S₂O₇ was obtained by thermal decomposition of K₂S₂O₇ (Merck, Pro Analysis, maximum 0.001% N) as described earlier. K₂SO₄ (Merck, Suprapur) and Li₂SO₄ (Sigma, 99%) were dried at 500°C overnight. V₂O₅ from Cerac (99.9% pure) and Ag₂SO₄ from Heraeus (99.9% pure) were used without further purification. All sample preparation and handling were performed in an argon-filled glove box (Vacuum Atmospheres Inc.), with a measured oxygen and water concentration less than 10 ppm.

Cyclic voltammetry measurements were performed in a hermetically closed three-electrode quartz cell with an epoxy resin-joined and water-cooled stainless steel head (Fig. 1). The reference electrode was a silver wire placed in a Pyrex cylindrical chamber with a Pyrex frit bottom. A melt of potassium pyrosulfate saturated with silver and potassium sulfates was used as the electrolyte in the reference silver electrode chamber. This reference system has been proven to have a stable potential in the studied temperature range.

A gold wire sealed in a Pyrex tube served as the working electrode. A gold–wire spiral served as counter-electrode. It has been shown that gold is stable in both acidic and basic K₂S₂O₇ melts, and therefore gold was used as the electrode material in previous electrochemical investigations.

After the cell was filled with electrolyte and hermetically closed in an argon glove box, it was placed in a vertical alumina–block furnace, with temperature control within ±0.1°C. The components of the melt could be mixed by automatic rocking of the furnace. The temperature of the melt was measured by a calibrated Chromel-Alumel thermocouple in a Pyrex case placed inside the electrochemical cell.
Voltammetric measurements were carried out with a Schlumberger Model SI1286 electrochemical interface, and a Graphtec Model WX3000-1 X-Y Recorder.

Results and Discussion

Voltammetric Measurements in the K$_2$S$_2$O$_7$ and K$_2$S$_2$O$_7$-K$_2$SO$_4$ Melts at 440°C

A characteristic cyclic voltammogram obtained at a gold electrode in molten K$_2$S$_2$O$_7$ in argon is shown in Fig. 2. From Fig. 2, it can be seen that the potential window of the potassium pyrosulfate is limited by a gradual reduction at -1.1 V and a steep oxidation at 1.0 V. In addition, four minor waves can be seen inside the electrochemical window: three reduction waves (A, B, and D) and one oxidation wave (C). The processes corresponding to the minor waves could affect the potential window of molten potassium pyrosulfate, and therefore their identification is important.

The dependence of the A, C, and D peak parameters on the potential scan rate ($V_s$) is shown in Fig. 3-5. All the peak currents depend linearly on the square root of the potential scan rate, and all the peak potentials are linearly dependent on the logarithm of the potential scan rate. It should be noticed that wave A becomes suppressed during cycling (Fig. 6).

The effect of melt acidity was also studied, by adding potassium sulfate to the potassium pyrosulfate. A comparison of the voltammetric characteristics of molten K$_2$S$_2$O$_7$ and K$_2$S$_2$O$_7$-K$_2$SO$_4$ systems (saturated with K$_2$SO$_4$ according to the data of Ref. 21) is made in Fig. 7. It is seen that in basic (SO$_4^{2-}$-rich) melts, the peak potential of reaction C moves in the positive direction and becomes less sharp. A new anodic reaction (E) appears at approximately 0.9 V. For the reactions C, D, and E in the basic melt, peak currents depend linearly on the square root of the potential scan rate, and peak potentials are independent of the potential scan rate in the 25-200 mV/s region (for E at the 50-100 mV/s region) (Fig. 8 and 9). For reaction A, the peak current and peak potential are linearly dependent on the square root and logarithm of the potential scan rate, respectively (Fig. 8 and 9).

According to Nicholson and Shain, Eq. 2 describes the dependence of the peak potential of the irreversible process on the potential scan rate

$$E_p = E_0 + \frac{RT}{\alpha n F} \left[ 0.78 - \ln k_s + 1/2 \ln \frac{D a n F V_s}{RT} \right]$$

where $E_p$ is the peak potential, $E_0$ is the standard potential, $k_s$ is the specific rate constant, $\alpha$ is the transfer coefficient; $n$ is the number of electrons taking part in the irreversible electrochemical reaction; $D$ is the diffusion coefficient; $R$ is the gas constant; $F$ is the Faraday constant; $T$ is absolute temperature.

The linear dependence of the peak potential on the logarithm of the potential scan rate is an indication of irreversibility of the electron transfer stage of the electrochemical reaction. In the case of a reversible reaction, the peak potential should be independent of the potential scan rate. It is seen from Fig. 4-6 that the electrochemical reactions A, C, and D are irreversible. Therefore we can use the Nicholson and Shain equation (Eq. 3), to calculate the value of $\alpha n$ for the studied electrochemical reactions

$$E_p - E_{p/2} = -1.857 \frac{RT}{\alpha n F}$$

($E_p$ and $E_{p/2}$ are peak and half-peak potentials, respectively).

The $\alpha n$ values obtained for A, C, and D peaks using Eq. 3 at different potential scan rates are given in Table 1. By assuming as usual that $\alpha$ equals 0.5, the numbers of electrons participating in A, C, and D processes are estimated to be 3 (the calculated $n_e$ are 3.27, 3.27, 2.70, and 2.76 at 200, 300, 400, and 500 mV/s, respectively), 3 (3.27, 3.27, 3.27, and 2.87), and 2 (2.87, 1.91, 1.76, and 1.76), respectively.

The number of electrons participating in the reversible electrochemical reaction can be calculated using Eq. 4.
Fig. 3. Cyclic voltammograms measured with gold electrodes in molten K$_2$S$_2$O$_7$ at 440°C and curve 1, 100; curve 2, 200; curve 3, 300; curve 4, 400; curve 5, 500 mV/s.

The $n$ values for the reversible (at least at low potential scan rates) electrochemical reactions, C and E, are given in Table I. For reaction C, these data are in agreement with a three-electron process obtained in the acidic melt ($n = 3.04, 3.43$, and $3.04$ at $50, 75$, and $100$ mV/s, respectively). Process E is evidently a two-electron reaction ($n = 2.28$ and $2.36$ for $50$ and $100$ mV/s, respectively).

In the investigated electrolyte system a three-electron process is possible only when gold participates in the reaction. Therefore, from our data and the results of Durand et al., wave C can be ascribed to the electro-oxidation of gold according to reactions Eq. 5A and 5B

For acidic melt

$$\text{Au + } 2n\text{S}_2\text{O}_7^{2-} \rightarrow \text{Au(SO}_4^{2-})_{3-2n} + n\text{S}_2\text{O}_7^{2-} + 3\text{e}^- \quad [5A]$$

For basic melt

$$\text{Au + nSO}_2^- \rightarrow \text{Au(SO}_4^{2-})_{3-2n} + 3\text{e}^- \quad [5B]$$

The kinetics of oxidation is probably dependent on the chemical reaction forming the gold sulfate complex (so-called EC-mechanism). According to the theoretical investigation of Nicholson and Shain, the following slow chemical reaction will be able to overrule the reversible (i.e., quick) electrochemical reaction from the oxidation-reduction equilibrium. In the case of a fast following chemical reaction an anodic voltammogram for the reversible electrochemical reaction will usually be obtained but without the opposite cathodic process. However, even in this case, we still might observe a voltammetric picture of an EC-mechanism with slow chemical reaction at high potential scan rates. In fact, as seen in Fig. 7 and 9, the electrochemical reaction C is irreversible in the acidic melt (Eq. 5A) and reversible in the basic melt, where the concentration of SO$_4^{2-}$ ions is sufficiently high (Eq. 5B). The related cathodic wave is absent for 5B, and at the potential scan rates higher than $100$ V/s this process becomes irreversible.
The reaction related to wave A, corresponding to gold participation, is eliminated during cycling. Evidently it is a surface gold–oxide reduction, which can be described by the following equation

$$1/2 \text{AuO}_3 + 3/2 \text{SO}_4^{2-} + 3e^- \rightarrow \text{Au} + 3\text{SO}_3^{2-}$$  \[6\]

The voltammetric wave E appears only in basic melts (containing an excess of sulfate). Wave E is caused by a quasi-reversible two-electron reaction (the reversibility depends on the potential scan rate). It is known that sulfate participates in both the anodic and the cathodic limiting processes in the (Li, K)SO₄ eutectic melt at 625°C. The anodic limiting process was the oxidation of the sulfate–ion to sulfur trioxide and oxygen gas. The cathodic limiting process was the reduction of sulfate to sulfite, possibly followed by additional steps leading to the formation of sulfur, sulfide, or both, depending on the conditions of electrolysis. Equilibrium potentials were obtained for the anodic limiting processes, which occur at about +0.9 V vs. Ag/Ag⁺. This is in agreement with Frank and Wim尼克's data for molten K₂S₂O₇. For the platinum electrode, the reduction potential of the O₂/SO₃ couple was between +0.046 and +0.029 V vs. Ag/Ag⁺, depending on the treatment of platinum. Burrows and Hills have shown that the potential-determining reaction involves one or more types of oxides of platinum. Apparently no proper oxygen/oxide electrode has been achieved in this melt. Salzano and Newman have ascertained that sulfur dioxide and trioxide were electrochemically active in the alkali sulfate melts. Boxall and Johnson observed a Nernstian behavior with a gold electrode in a molten K₂S₂O₇.}

The electrochemical behavior of V₂O₅ in molten K₂S₂O₇ is typical for the preceding slow chemical reaction mechanism. Therefore, wave B can be ascribed to the electrochemical reduction of solvated SO₃⁻, according to Eq. 9

$$3\text{SO}_3^{2-} + 2e^- \rightarrow \text{SO}_2 + 3\text{SO}_4^{2-}$$  \[9\]

The electroreduction stability limit (-1.1 V) of the potassium pyrosulfate melt should be the reduction of the pyrosulfate anion, probably according to Eq. 10

$$2\text{SO}_4^{2-} + 2e^- \rightarrow \text{SO}_2 + 3\text{SO}_4^{2-}$$  \[10\]

The electrochemical behavior of VO₂ in molten potassium pyrosulfate at a gold electrode was studied in a wide concentration range. The voltammetric curve obtained for 2 m/o of VO₂ is shown in Fig. 10. It is seen that two new cathodic waves (F and G) and one anodic wave (H, which obviously is an oxidation of the first reduction product) appear in the presence of VO₂. The dependence of the F, G, and H wave–parameters on the potential scan rate is shown in Fig. 11. The F and H peak currents depend linearly on the square root of the potential scan rate, and the peak potentials are virtually independent on the potential scan rate (Fig. 15 and 16). For a melt with 2 m/o VO₂, the ratio of the first reduction (F) and the oxidation (H) peak currents is close to unity (Table II). The voltammetric characteristics of the K₂S₂O₇–VO₂ system of formal VO₂ concentrations up to 20 m/o are shown in Fig. 12–16 Table II. The F peak current is linearly dependent on the formal VO₂ concentration. Starting from 5 m/o of VO₂, the F and H peak potentials depend linearly on the logarithm of the potential scan rate. The ratio between F and H peak currents decreases with the concentration of VO₂, having a minimum at 5 m/o. The reduction wave G, instead of a peak, shows a linear current-potential dependence for high potential scan rates. A new oxidation wave, I, corresponding to the reduction wave G, becomes noticeable at 10 m/o of VO₂.

Hence, the first reduction–oxidation process (waves F and H) seems to be reversible below 5 m/o of VO₂, but irreversible at higher concentrations. It can be seen from Table II that at all studied concentrations both peaks F and H are one-electron reactions. It means that these waves are caused by the V(V) → (IV) and V(IV) → V(V) reactions, and that only one vanadium atom participates in these processes. They may be described by Eq. 11

$$\text{VO}_2^{2+} + \text{SO}_4^{2-} + e^- \leftrightarrow \text{VO}_2\text{SO}_4\text{S}_2\text{O}_7^{2-} + \text{SO}_3^{2-}$$  \[11A\]

$$\text{VO}_2\text{SO}_4\text{S}_2\text{O}_7^{2-} + \text{SO}_3^{2-} + e^-$$  \[11B\]

The change in kinetics of the reaction given by Eq. 11 at V₂O₅ concentrations higher than 5 m/o could be explained by an increase in degree of dimerization or of polymerization of the V(IV) species, proceeding via a mechanism of electrochemical desorption, in which a slow charge-transfer step is usually involved.

Assuming (i) the potential difference between F and G reactions, (ii) the electrochemical behavior of the V(IV) cathodic materials, and (iii) the chemical identification of the products as the products of the reduction of V(IV) under the conditions of commercial sulfuric acid production, we propose a V(IV) → V(III) reduction as an explanation for wave G. The shape of the voltammetric curve can be explained by the low solubility of the V(IV) and V(III) complex ions...
pounds at 440°C, and therefore one has an ohmic control of reaction C.

Starting from 10 m/o of V₂O₅, a new oxidation wave (wave I) appears on the voltammetric curves. This wave can be related to a V(III) → V(IV) reaction considering the reversible performance of the VO₂ secondary battery cathodes. Starting from 10 m/o of V₂O₅, a new oxidation wave (wave I) appears on the voltammetric curves. This wave can be related to a V(III) → V(IV) reaction considering the reversible performance of the VO₂ secondary battery cathodes.

Figures 14-16 (and Table II) demonstrate the voltammetric effect of adding 5.5 m/o Li₂SO₄ to a 10 m/o V₂O₅ solution in a potassium pyrosulfate melt. After addition of lithium sulfate, a white precipitate appeared in the melt. Taking into account the higher solubility of Li₂SO₄ than K₂SO₄ (which is 4.64 m/o at 440°C) in molten K₂S₂O₇, and a possible exchange equilibrium (Eq. 12), in the reciprocal system

\[
\text{Li}_2\text{SO}_4 + \text{K}_2\text{S}_2\text{O}_7 \rightarrow \text{Li}_2\text{S}_2\text{O}_7 + \text{K}_2\text{SO}_4 \quad [12]
\]

we assume that all the added lithium sulfate is practically present in the form of lithium pyrosulfate in the K₂S₂O₇ melt.

It can be seen that the Li⁺ ion has a promoting effect on both oxidation and reduction processes, with the depolarization being almost three times higher for the V(V) → V(IV) reduction than for the V(IV) → V(V) oxidation. The ratio between the anodic and the cathodic peak currents is closer to unity in the presence of Li₂SO₄ in the studied electrolytes. It should also be noted that saturation of a K₂S₂O₇ melt by potassium sulfate has only a marginal effect on the V(V) → V(IV) process.

The thermodynamical characteristics of the electrochemical reactions of Eq. 11A and 11B obviously depend on the alkali cation nature: the smaller the cation radius the higher its polarization effect on the complex vanadium par-
particles. This could cause the weakening of the V-O bond and facilitate the vanadium reduction.

The alkali-ion activation effect could qualitatively be divided into two types: "large amount additive" activation when the additive considerably changes the physicochemical properties of the bulk melt (i.e., Cs⁺); and "small amount additive" activation when the bulk physicochemical properties are not significantly changed (Na⁺, Li⁺).

For understanding of the "large amount" promoting effect, at least two factors should be taken into account.

Table I. Peak and half-peak potential difference, $E_p - E_{p/2}$, and the number of participating electrons, $n$, for the irreversible A, C, and D reactions and the number of participating electrons, $n$, for the reversible C and E reactions obtained in molten K$_2$S$_2$O$_7$ and K$_2$S$_2$O$_7$-K$_2$SO$_4$ systems at 440°C.

<table>
<thead>
<tr>
<th>Electrochemical reaction</th>
<th>Potential scan rate, V/s</th>
<th>$E_p - E_{p/2}$, V</th>
<th>$\alpha_n$</th>
<th>$n_{irr}$, n</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>0.200</td>
<td>0.070</td>
<td>1.637</td>
<td>3.27</td>
</tr>
<tr>
<td></td>
<td>0.300</td>
<td>0.085</td>
<td>1.248</td>
<td>2.70</td>
</tr>
<tr>
<td></td>
<td>0.400</td>
<td>0.083</td>
<td>1.380</td>
<td>2.76</td>
</tr>
<tr>
<td></td>
<td>0.500</td>
<td>0.083</td>
<td>1.380</td>
<td>2.76</td>
</tr>
<tr>
<td>C</td>
<td>0.200</td>
<td>0.070</td>
<td>1.637</td>
<td>3.27</td>
</tr>
<tr>
<td></td>
<td>0.300</td>
<td>0.080</td>
<td>1.433</td>
<td>2.87</td>
</tr>
<tr>
<td></td>
<td>0.500</td>
<td>0.080</td>
<td>1.433</td>
<td>2.87</td>
</tr>
<tr>
<td>D</td>
<td>0.100</td>
<td>0.110</td>
<td>1.042</td>
<td>2.08</td>
</tr>
<tr>
<td></td>
<td>0.200</td>
<td>0.120</td>
<td>0.955</td>
<td>1.91</td>
</tr>
<tr>
<td></td>
<td>0.300</td>
<td>0.130</td>
<td>0.878</td>
<td>1.76</td>
</tr>
<tr>
<td></td>
<td>0.400</td>
<td>0.130</td>
<td>0.878</td>
<td>1.76</td>
</tr>
<tr>
<td>C</td>
<td>0.050</td>
<td>0.045</td>
<td>3.04</td>
<td></td>
</tr>
<tr>
<td></td>
<td>0.075</td>
<td>0.045</td>
<td>3.43</td>
<td></td>
</tr>
<tr>
<td></td>
<td>0.100</td>
<td>0.045</td>
<td>3.43</td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>0.050</td>
<td>0.060</td>
<td>2.28</td>
<td></td>
</tr>
<tr>
<td></td>
<td>0.100</td>
<td>0.060</td>
<td>2.28</td>
<td></td>
</tr>
</tbody>
</table>

$a$ $\alpha_n$ is the product of the transfer coefficient and the number of electrons participating in the irreversible A, C, D reactions.

Fig. 9. Peak potential (a) and peak current (b) of waves A, C, and E as function of potential scan rate obtained in molten K$_2$S$_2$O$_7$, saturated with K$_2$SO$_4$, with gold electrodes and at 440°C.

Fig. 10. Cyclic voltammograms measured with gold electrodes in (1) molten K$_2$S$_2$O$_7$ and (2) molten K$_2$S$_2$O$_7$ + 2 m/o V$_2$O$_5$ at 440°C and 100 mV/s.

Fig. 11. Cyclic voltammograms measured with gold electrodes in molten K$_2$S$_2$O$_7$ + 2 m/o V$_2$O$_5$ at 440°C and curve 1, 25; curve 2, 50; curve 3, 100; curve 4, 200 mV/s.

1. Most cesium salts have supercooling and glass-forming tendencies, which are more pronounced for their near-eutectic mixtures with strongly polarizing cation salts (lithium, sodium, or small alkali-earth ions): the more diversity and number of particles in a multicomponent system, the higher the glassforming tendency (so-called confusion principle). Moreover, the solubility of the vanadium species will be higher in the pyrosulfates with weakly polarizing cations, VO$_5^-$ and VO$_{12}^-$ are strong polarizing...
Table II. $\alpha n_a$ for the irreversible and the number of participating electrons, $n$, for the reversible $V(V) \leftrightarrow V(IV)$ reaction as well as $I_p/I_{p}^{a}$ in the molten $K_2S_2O_7-V_2O_5$ system at $440^\circ C$ and 100 mV/S.

<table>
<thead>
<tr>
<th>m/o $V_2O_5$</th>
<th>Reaction</th>
<th>$n$</th>
<th>$\alpha n_a$</th>
<th>$I_p/I_{p}^{a}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Anodic</td>
<td>0.90</td>
<td></td>
<td>0.90</td>
</tr>
<tr>
<td></td>
<td>Cathodic</td>
<td>0.75</td>
<td></td>
<td>0.90</td>
</tr>
<tr>
<td>3</td>
<td>Anodic</td>
<td>0.67</td>
<td></td>
<td>0.90</td>
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<tr>
<td></td>
<td>Cathodic</td>
<td>0.71</td>
<td></td>
<td>0.90</td>
</tr>
<tr>
<td>5</td>
<td>Anodic</td>
<td>0.76</td>
<td></td>
<td>0.55</td>
</tr>
<tr>
<td></td>
<td>Cathodic</td>
<td>0.53</td>
<td></td>
<td>0.55</td>
</tr>
<tr>
<td>10</td>
<td>Anodic</td>
<td>0.71</td>
<td></td>
<td>0.60</td>
</tr>
<tr>
<td></td>
<td>Cathodic</td>
<td>0.45</td>
<td></td>
<td>0.60</td>
</tr>
<tr>
<td>20</td>
<td>Anodic</td>
<td>0.55</td>
<td></td>
<td>0.68</td>
</tr>
<tr>
<td></td>
<td>Cathodic</td>
<td>0.38</td>
<td></td>
<td>0.68</td>
</tr>
<tr>
<td>10</td>
<td>Anodic</td>
<td>0.55</td>
<td></td>
<td>0.68</td>
</tr>
<tr>
<td>5.5 m/o Li$_2$SO$_4$</td>
<td>Cathodic</td>
<td>0.38</td>
<td></td>
<td>0.79</td>
</tr>
</tbody>
</table>

$\alpha n_a$ is the product of the transfer coefficient and the number of electrons participating in the irreversible reaction.

cations, and tend to form complexes with the solvent. Therefore, the less competition for the solvent anions, the higher solubility of the vanadium compounds.

2. The electroconductivity of the vanadium-containing pyrosulfate melts should have noticeable effect on their catalytical activity. Large alkali cations have less tendency to form clusters. They could have a mobility equal to that for small cations (lithium or sodium), the so called Chemla effect. Moreover, for molten salts with easily polarizable molecular anions, the electroconductivity of cesium or potassium salts increases in the presence of small amounts of lithium (or sodium) salts. This can be explained by the weakening of the Cs$^+$-anion bonding in the presence of highly polarizing cations.

In industrial catalysis, the activation effect of small amounts of sodium sulfate is well known. This effect can be explained as a change in the parameters of the catalytical reaction by highly polarizing Na$^+$ ions in the catalytically active vanadium complexes.

There is no indication in the literature that there would be a beneficial effect of the addition small amounts of Li$^+$ (sulfate or pyrosulfate). Large amounts of lithium pyrosulfate has shown no or a negative promoting effect in a series of alkali metal pyrosulfates. If our explanation of the Na$^+$
promoting effect is correct, and the Li⁺ activation effect on the electrochemical V(IV) ↔ V(II) processes is taken into account, we would expect a promoting effect for diluted Li⁺ solutions, also for industrial catalysts.

Conclusions

The potential window of pure molten K₂S₂O₇ was determined to be 2.1 V, being limited by S₈O₇²⁻ oxidation and reduction. The formation of oxygen in basic melts proceeds reversibly at more negative potentials than in acidic melts where the process is irreversible.

Fig. 15. V(V) → V(IV) and V(IV) → V(V) peak currents as function of potential scan rate obtained in molten K₂S₂O₇-V₂O₅ and K₂S₂O₇-V₂O₅-Li₂SO₄ systems with gold electrodes at 440°C. c, Cathodic; a, anodic.

The V₂O₅ reduction process includes two stages: V(V) → V(IV) and V(IV) → V(III). The first stage is reversible up to a fraction of 5 m/o of V₂O₅, and is a one-electron reaction at all studied concentrations (i.e., up to 20 m/o). Hence, only one vanadium atom is involved in the reaction.

The presence of Li⁺ ions causes a noticeable acceleration of the V(IV) electroreduction.

Acknowledgments

This investigation has been supported by the EEC BRITE-EURAM II Project BRE2 CT93 0447 and the Danish Natural Science Research Council. The authors thank R. Fehrmann for the valuable discussion.

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Manuscript submitted July 7, 1994; revised manuscript received Feb. 26, 1995.