Tesfaye, Fiseha; Jung, In Ho; Paek, Min Kyu; Moroz, Mykola; Lindberg, Daniel; Hupa, Leena

Thermochemical Data of Selected Phases in the FeO

\[
\begin{align*}
x & \quad \text{–FeSO} \\
4 & \quad \text{–Fe} \\
2 & \quad (\text{SO})^3 \\
4 & \quad \\
3 & \quad \text{System}
\end{align*}
\]

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Thermochemical Data of Selected Phases in the FeO$_x$–FeSO$_4$–Fe$_2$(SO$_4$)$_3$ System

Fischa Tesfaye, In-Ho Jung, Min-Kyu Paek, Mykola Moroz, Daniel Lindberg and Leena Hupa

Abstract Several recent studies have shown the potential of oxy-fuel combustion to reduce NO$_x$(gas) and SO$_2$(gas) emissions. However, the mechanisms through which SO$_2$(gas) reduction takes place has yet to be fully understood. Therefore, the development of oxy-sulfate thermodynamic database for a better understanding and control of SO$_2$(gas) emission during oxy-fuel combustion processes is essential. The focus of this research is on the thermodynamic modelling of the iron oxide–sulfate system with the FactSage 7.2 software package. Thermodynamic properties of selected phases in the FeO$_x$–FeSO$_4$–Fe$_2$(SO$_4$)$_3$ system were critically reviewed, compiled and assessed over a wide temperature range (298–2000 K) to obtain accurate thermodynamic description of the system at different temperatures. New C$_p$ functions, which include the recent experimental data, were optimized. The obtained results are presented and discussed.

Keywords Sulfate · Sulfur dioxide · Decomposition reaction
Thermochemical data

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Introduction

Today, oxy-fuel combustion is considered as one of the major technologies for CO$_2$(gas) capture in power plants. A number of recent studies have also shown its potential to reduce NO$_x$(gas) and SO$_2$(gas) emissions. To control and optimize the emission reduction mechanisms, the development of oxy-sulfate thermodynamic database for the oxy-fuel combustion processes is essential. Therefore, the focus of this research is on the thermodynamic modelling of the FeO$_x$–FeSO$_4$–Fe$_2$(SO$_4$)$_3$ system with the FactSage 7.2 software package as a part of the larger database of the CuO–MgO–Al$_2$O$_3$–SiO$_2$–FeO–Fe$_2$O$_3$–sulfate system.

In the present work, thermodynamic data of FeSO$_4$ and Fe$_2$(SO$_4$)$_3$ were compiled and critically assessed below 2000 K to obtain accurate thermodynamic parameters. All calculations were carried out using the FactSage 7.2 thermochemical software [1], including the pure substances and self-developed databases.

Phase Relations in the Ternary Fe–S–O System

Recently, the relatively well known Fe–O and Fe–S systems were reevaluated by Hidayat et al. [2] and Walder and Pelton [3], respectively. The ternary Fe–S–O system was investigated by Shishin et al. [4]. In this system, only two stable ternary compounds, FeSO$_4$ and Fe$_2$(SO$_4$)$_3$, were reported. At ambient pressure conditions, both ternary phases decompose below 1500 K to form the gas and oxide phases before melting [4]. Chemical structures of the anhydrous ferrous sulfate (FeSO$_4$) and ferric sulfate (Fe$_2$(SO$_4$)$_3$) are illustrated in Fig. 1. Barany and Adam [5], Pankratz and Weller [6] and Majzlan et al. [7] measured the thermodynamic properties of Fe$_2$(SO$_4$)$_3$, and Moore and Kelly [8] measured the low temperature thermodynamic properties of FeSO$_4$. Using the susceptibility measurement technique, Frazer and Brow [9] and Kirfel et al. [10] measured magnetic data ($\mu_B$) for $\alpha$-FeSO$_4$ and $\beta$-FeSO$_4$ to be 4.1 ± 0.4 and 5.44 ± 0.27, respectively.

The anhydrous iron (III) sulfate (Fe$_2$(SO$_4$)$_3$) crystallizes in two modifications, monoclinic and trigonal. The stability relationship between the two phases is not clear. Only the trigonal polymorph has been reported in nature, under the name mikasaite. Mikasaite is formed by precipitation from hot gases escaping fractures in coal beds [7]. Fe$_2$(SO$_4$)$_3$ is rare as a mineral, but it is an important compound for the development of a thermodynamic database for a number of natural hydrated iron sulfates [11] and oxy-fuel combustion applications. More details on the thermodynamic properties of the phases in the FeO$_x$–FeSO$_4$–Fe$_2$(SO$_4$)$_3$ system are presented and discussed in the subsequent sections.
Thermochemical Data of FeSO₄ and Fe₂(SO₄)₃

**Enthalpy of Formation (ΔH_f°) of FeSO₄**

Thomsen [12] measured the enthalpy of Reaction (1). In the NIST-JANAF thermochemical table [13], this enthalpy of Reaction (1) was used to estimate Δ_f°(298.15 K, FeSO₄) = −924.66 kJ mol⁻¹.

FeCl₂(200H₂O) + H₂SO₄(200H₂O) = FeSO₄(200H₂O) + 2HCl(100H₂O).  \( 1 \)

In the NIST-JANAF thermochemical data [13], Δ_f°(298.15 K, FeSO₄) values from the decomposition pressure of FeSO₄ reported by D’Ans [14], Greulich [15], and Neumann and Heinke [16] were also estimated by the second and third laws with an average value of −925.1 kJ mol⁻¹. Δ_f°(298.15 K, FeSO₄) = −928.85 ± 8.4 kJ mol⁻¹ and Δ_f°(0 K) = −919.33 ± 8.4 kJ mol⁻¹ are the values recommended by Chase [13].

**ΔH_f°(Fe₂(SO₄)₃)**

Based on the data reported for the chemical Reaction (2) by Barany and Adami [5] and using enthalpies of formations of the other components involved in the reaction, Δ_f°(298.15 K, Fe₂(SO₄)₃) = −2583 ± 1.7 kJ mol⁻¹ was estimated and presented in [13].

Fe₂O₃ + 3(H₂SO₄ · 14.85H₂O)(sol.) = Fe₂(SO₄)₃(cr) + 47.56H₂O(l).  \( 2 \)
The equilibrium pressures for decomposition of Fe$_2$(SO$_4$)$_3$ have been determined by several investigators [16–23] at different temperatures via chemical Reactions (3) and (4). The data were critically reviewed by Kellogg [24].

\[
\text{Fe}_2\text{(SO}_4\text{)}_3 = \text{Fe}_2\text{O}_3 + 3\text{SO}_3, \quad (3)
\]

\[
\text{SO}_3\text{(g)} = \text{SO}_2\text{(g)} + 0.5\text{O}_2. \quad (4)
\]

The measured pressure of the chemical equilibrium is the total pressures of \(P(\text{SO}_3)\), \(P(\text{SO}_2)\), and \(P(\text{O}_2)\). In order to calculate the enthalpy change of Reaction (3), the partial pressures of \(\text{SO}_3\text{(g)}\) from the total vapor pressure data at each temperature was evaluated. Based on the derived values for \(P(\text{SO}_3)\), \(\Delta H^\circ_r(298.15\text{ K}, \text{Fe}_2\text{(SO}_4\text{)}_3)\) values were derived using the third law and presented in the NIST-JANAF thermochemical data [13]. These average values of the calculations give \(-2584.164\text{ kJ mol}^{-1}\). This value is higher only by \(-1\text{ kJ mol}^{-1}\) than the value they obtained through the data of Barany and Adami [5].

After experimental studies in the temperature range of 273–395 K, Majzlan et al. [7] determined \(\Delta H^\circ_r(298.15\text{ K}, \text{Fe}_2\text{(SO}_4\text{)}_3) = -2585.2 \pm 4.9\text{ kJ mol}^{-1}\), which is higher by \(-2\text{ kJ mol}^{-1}\) than the value of Barany and Adami [5].

**Heat Capacity (\(C_p\)) and Entropy (\(S^\circ\))**

**FeSO$_4$**

The low-temperature \(C_p\) of the anhydrous FeSO$_4$ in the temperature range of 53–294.6 K was determined by Moore and Kelly [8] experimentally. Based on standard entropy at 50.12 K, they calculated, \(S^\circ(\text{FeSO}_4, 298.15\text{ K}) = 107.57 \pm 0.84\text{ J K}^{-1}\text{ mol}^{-1}\). Since the report by Moore and Kelley [8] did not mention the magnetic entropy contribution, an attempt was made by Knacke et al. [25] to add magnetic entropy \((S^\text{mag})\) to \(S^\circ(\text{FeSO}_4, 298.15\text{ K})\) and reevaluate the decomposition pressure data. They used the theoretical value of magnetic entropy \((S^\text{mag})\) for FeSO$_4$ \(R\ln(2S_{\text{spin}})\), where \(S_{\text{spin}}\) is the magnetic spin. Since the electrons in the iron ions in FeSO$_4$ are in a high-spin configuration, they may have used \(S_{\text{spin}} = \frac{5}{2}\) which implies \(S^\text{mag} = R\ln5 = 13.4\text{ J K}^{-1}\text{ mol}^{-1}\). Therefore, the value \(S^\circ(298.15\text{ K}) = (107.57 + 13.4)\text{ J K}^{-1}\text{ mol}^{-1}\) was adopted for FeSO$_4$. The \(C_p\) above 294.9 K was estimated by comparison with those for MnSO$_4$. In the calculations, the high-temperature \(C_p(\text{MnSO}_4), 870.3–1082.3\text{ K, were taken from the experiments of Southard and Shomate [26].}\)

**Fe$_2$(SO$_4$)$_3$**

The heat capacities were established by comparison with those for FeSO$_4$, assuming their average specific heats, \(J\text{ K}^{-1}\text{ g}^{-1}\), to be the same. The value of \(S^\circ(298.15\text{ K})\) was estimated so that the second and third law \(\Delta H^\circ_r(298.15\text{ K})\) values, derived from
decomposition pressure data, are in reasonable agreement. The value $S'(298.15 \text{ K}) = 120.957 \pm 1.3 \text{ J K}^{-1} \text{ mol}^{-1}$ was recommended in the NIST-JANAF thermochemical table [13].

Majzlan et al. [7] studied thermodynamic properties of the monoclinic Fe$_2$(SO$_4$)$_3$ by acid solution, adiabatic, and semi-adiabatic calorimetry methods. They reported $\Delta H'_{m}(298.15 \text{ K}, \text{ Fe}_2(\text{SO}_4)_3) = -2585.2 \pm 4.9 \text{ kJ mol}^{-1}$ obtained by an appropriate thermochemical cycle with enthalpies of solution of monoclinic Fe$_2$(SO$_4$)$_3$, \(\alpha\)-MgSO$_4$, \(\gamma\)-FeOOH, H$_2$O, and MgO in 5NHCl and at 298 K. They also compiled $C_p$ values from 0.5 to 400 K and represented the data in the temperature range of 273–395 K by Eq. (5).

$$C_p/\text{J K}^{-1} \text{ mol}^{-1} = 213 + 0.312 \cdot (T) - 2.959 \cdot 10^6 \cdot (T)^{-2} \quad (273–395\text{ K})$$ (5)

**Melting and Decomposition Temperatures**

The decomposition temperature 1451 K for Fe$_2$(SO$_4$)$_3$ at which the total pressure of the gaseous products equals 1 atm, was calculated by [13] through the graphical extrapolation of the decomposition pressure data measured by Warner and Ingraham [27]. By a similar method, they also calculated $T_{\text{decomp}}(\text{FeSO}_4) = 944 \text{ K}$ from the data of Greulich [15]. Reactions (3) and (6) are the decomposition reactions for which the vapor pressure experimental data are available.

$$\text{FeSO}_4 \rightleftharpoons \text{FeO} + \text{SO}_3.$$ (6)

The thermal decomposition of FeSO$_4$·6H$_2$O was studied by Masset al. [28] applying the mass spectroscopy coupled with DTA/TG thermal analysis technique under inert atmosphere. After the dehydration, they reported to have observed two decomposition steps for FeSO$_4$ between 798 and 983 K according to the following Reactions (7) and (8):

$$6\text{FeSO}_4 \rightarrow \text{Fe}_2(\text{SO}_4)_3 + 2\text{Fe}_2\text{O}_3 + 3\text{SO}_2 (798–923 \text{ K})$$ (7)

$$\text{Fe}_2(\text{SO}_4)_3 \rightarrow 2\text{Fe}_2\text{O}_3 + 3\text{SO}_2 + 1.5\text{SO}_2 (898–983 \text{ K})$$ (8)

$$\text{Fe}_2(\text{SO}_4)_3 \rightarrow 2\text{FeSO}_4 + \text{SO}_2 + \text{O}_2.$$ (9)

$$\text{FeSO}_4 \rightarrow 0.5\text{Fe}_2\text{O}_3 + \text{SO}_2 + 0.25\text{SO}_2.$$ (10)

Gallagher et al. [29] studied the thermal decomposition of FeSO$_4$·xH$_2$O using the conventional thermogravimetry, differential thermal analysis, and evolved gas analysis techniques. In an oxidizing atmosphere, they observed that FeSO$_4$ oxidizes as

$$2\text{FeSO}_4 + 0.5\text{O}_2 \rightarrow \text{Fe}_2\text{O}(\text{SO}_4)_2.$$ (11)
Fe₂O(SO₄)₂ → Fe₂O₃ + 2SO₂. (873–948K)  \hspace{1cm} (12)

In the inert atmosphere

2FeSO₄ → Fe₂SO₄ + SO₂. (748–848 K)  \hspace{1cm} (13)
Fe₂O₂SO₄ → Fe₂O₃ + SO₂ + 0.5SO₂. (823–948K)  \hspace{1cm} (14)

Other proposed decomposition reactions under N₂(gas) protective atmosphere are for Reaction (7), 891 K [30] (heating rate 20 K/min) and 960 K [31], and for Reaction (3), 951 K [30] (heating rate 20 K/min) and 1028 K [31].

A comparative summary of the decomposition temperatures of FeSO₄ and Fe₂(SO₄)₃ determined by different researchers and methods are compiled in Table 1.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>T_{trans}/K</th>
<th>T_{decomp}/K</th>
<th>Ref.</th>
<th>Remark</th>
</tr>
</thead>
<tbody>
<tr>
<td>α-FeSO₄ → β-FeSO₄</td>
<td>623</td>
<td>–</td>
<td>[10]</td>
<td>High-pressure phase transition (at P_{tot} ≈ 15 kbar)</td>
</tr>
<tr>
<td>Reaction (3): Fe₂(SO₄)₃ → Fe₂O₃ + 3SO₃</td>
<td>1461</td>
<td>[13]</td>
<td>Determined by the gas pressure measurements of [27]</td>
<td></td>
</tr>
<tr>
<td>–</td>
<td>1017</td>
<td>This work</td>
<td></td>
<td></td>
</tr>
<tr>
<td>–</td>
<td>951</td>
<td>[30]</td>
<td>Heating rate 20 K min⁻¹</td>
<td></td>
</tr>
<tr>
<td>–</td>
<td>1028</td>
<td>[31]</td>
<td>–</td>
<td></td>
</tr>
<tr>
<td>Reaction (7): 6FeSO₄ → 2Fe₂(SO₄)₃ + 2Fe₂O₃ + 3SO₂</td>
<td>798–923</td>
<td>891</td>
<td>[28]</td>
<td>Inert atmosphere</td>
</tr>
<tr>
<td>–</td>
<td>960</td>
<td>[30]</td>
<td>Heating rate 20 K min⁻¹</td>
<td></td>
</tr>
<tr>
<td>Reaction (8): Fe₂(SO₄)₃ → Fe₂O₃ + 5SO₂ + 1.5O₂</td>
<td>898–983</td>
<td>[28]</td>
<td>Inert atmosphere</td>
<td></td>
</tr>
<tr>
<td>Reaction (11): 2FeSO₄ + 0.5SO₂ → Fe₂O₅(SO₄)₂</td>
<td>870–948</td>
<td>873–948</td>
<td>[29]</td>
<td>Oxidizing atmosphere</td>
</tr>
<tr>
<td>Reaction (12): Fe₂O₅(SO₄)₂ → Fe₂O₃ + 2SO₂</td>
<td>–</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Reaction (13): 2FeSO₄ → Fe₂O₃SO₄ + SO₂</td>
<td>748–848</td>
<td>[29]</td>
<td>Inert atmosphere</td>
<td></td>
</tr>
<tr>
<td>Reaction (14): Fe₂O₃SO₄ → Fe₂O₃ + SO₂</td>
<td>823–948</td>
<td>–</td>
<td>–</td>
<td></td>
</tr>
</tbody>
</table>
Thermochemical Data of Selected Phases in the...

**Gibbs Energy Data**

The solid gas reactions below 1100 K were studied using the EMF technique by Skeaff and Espelund [32], Hsieh and Chang [33], Musbah and Chan [34], Kobe and Couch [35], Rosenqvist and Hofseth [36], Schaefer [37], in a continuously flowing gas atmosphere, using the vapor pressure measurements technique by Warner and Ingraham [27] and using the DTA technique by Alcock et al. [40]. Gibbs energies of reactions determined by the different methods are summarized in Table 2. These EMF experiments may not correspond to equilibrium with the gas phase, which can contain other gaseous species such as SO₃, SO₂ and O₂, but rather to a fixed potential of SO₂; i.e., the other gaseous species do not form fast enough to significantly change the composition of the flowing SO₂. If the equilibrium with the gas phase were attained, it would be almost pure SO₂ over the range of $P_{O_2}$ from $10^{-4}$ to $10^{-14}$ atm. At higher oxygen potentials, amounts of SO₃ and O₂ become substantial, while at lower $P_{O_2}$ the partial pressure of S₂ starts to increase. The composition of the gas flowing out of the EMF furnaces was estimated by Shishin et al. [4] which indicated that large amounts of condensed phases in EMF cells would have to react in order to produce the gas phase of equilibrium composition, which is not what happened in the experiments.

Jacob and Iyengar [39] studied the $\text{Fe}_2\text{O}_4 + \text{Fe}_2(\text{SO}_4)_3$ phase equilibria by the EMF method. They measured EMF values between 800 and 1000 K. Based on their results, they have concluded that oxysulfates do not form below 1100 K in the Fe–S–O

<table>
<thead>
<tr>
<th>Equilibrium reaction</th>
<th>$\Delta G^\circ$ (kJ mol⁻¹)</th>
<th>T/K</th>
<th>Ref.</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Fe}_2(\text{SO}_4)_3 \rightleftharpoons \text{Fe}_2\text{O}_3 + 3\text{SO}_2 + 1.5\text{O}_2$</td>
<td>970.05–0.724 · $T$</td>
<td>673–1073</td>
<td>[44]</td>
</tr>
<tr>
<td></td>
<td>772.32–0.724 · $T$</td>
<td>920–1020</td>
<td>[32]</td>
</tr>
<tr>
<td></td>
<td>870.235–0.8245 · $T$ (±1.4)</td>
<td>800–1000</td>
<td>[39]</td>
</tr>
<tr>
<td></td>
<td>704.61–0.697 · $T$</td>
<td>878–955</td>
<td>[40]</td>
</tr>
<tr>
<td></td>
<td>802.86–0.762 · $T$</td>
<td>906–995</td>
<td>[27]</td>
</tr>
<tr>
<td></td>
<td>753.098–0.744 · $T$</td>
<td>900–1000</td>
<td>[27] *</td>
</tr>
<tr>
<td></td>
<td>730.5–0.688 · $T$</td>
<td>800–900</td>
<td>[43]</td>
</tr>
<tr>
<td>$\text{Fe}_2(\text{SO}_4)_3 \rightleftharpoons \text{Fe}_2\text{O}_3 + 3\text{SO}_3$</td>
<td>576.895–0.546.1 · $T$ (±0.5)</td>
<td>800–1000</td>
<td>[39]</td>
</tr>
<tr>
<td></td>
<td>557.42–0.558 · $T$</td>
<td>900–1000</td>
<td>[27] *</td>
</tr>
<tr>
<td>$\text{Fe}_2(\text{SO}_4)_3 \rightleftharpoons 2\text{FeSO}_4 + \text{SO}_2 + \text{O}_2$</td>
<td>395.64–0.352 · $T$</td>
<td>703–904</td>
<td>[32]</td>
</tr>
<tr>
<td>$\text{Fe}_2\text{O}_3 + 0.5\text{Fe}_2\text{O}_3 + \text{SO}_2 + 1/4\text{O}_2$</td>
<td>258.74–0.202 · $T$</td>
<td>773–903</td>
<td>[44]</td>
</tr>
<tr>
<td>$\text{Fe}_2\text{O}_3 \rightleftharpoons 0.5\text{Fe}_2\text{O}_3 + \text{SO}_2 + 1/4\text{O}_2$</td>
<td>203.47–0.202 · $T$</td>
<td>779–900</td>
<td>[32]</td>
</tr>
</tbody>
</table>

*Re-calculated from the experimental data of Warner and Ingraham [27]
The determined Gibbs energies of reactions involving the gases SO₃, SO₂, O₂ are compiled in Table 2.

Combining the high-temperature change in enthalpy of reaction Fe₂(SO₄)₃ = Fe₂O₃ + 3SO₂(g) + 1.5O₂(g) reported by Jacob and Iyengar [39] with the pure substances enthalpy data of FactSage 7.2 for SO₂(g) and Fe₂O₃, we have calculated ΔH°^f(Fe₂(SO₄)₃) = −2586.55 kJ mol⁻¹. This value is higher only by −1.35 kJ mol⁻¹ than that of Majzlan et al. [7]. Warner and Ingraham [27] have also determined the enthalpy of decomposition of Fe₂(SO₄)₃ from the vapor pressure measurements to be ΔH°^d,decomp. = 566.51 kJ mol⁻¹.

**Results and Discussion**

The C_p(SO₃) functions obtained through the reactions MeO + SO₃ = MeSO₄ (Me = Ca, Mg, Mn) were calculated by applying the Neumann–Kopp’s rule described in Alcock et al. [41]. With the obtained C_p values, the corresponding C_p(FeSO₄) values were calculated according to Reaction (6). The results are illustrated in Fig. 2.

The heat capacities of FeSO₄ were calculated between 273 and 400 K by comparison with those experimental data reported for Fe₂(SO₄)₃, assuming their average

![Fig. 2 C_p versus T diagram of FeSO₄](image-url)
specific heats, $c_p$ (J K$^{-1}$ g$^{-1}$), to be the same. We have also calculated the experimental $C_p$(FeSO$_4$) values reported by Moore and Kelly [8] for the temperature range 273–295 K. By combining these all $C_p$ values with the $C_p$(FeSO$_4$) obtained through the derived $C_p$(SO$_4$) from the reaction MnSO$_4$ = MnO + SO$_3$ using the data of Barin Sadakane et al. [42], optimal $C_p$(FeSO$_4$) was calculated at each temperature, in the temperature range of 273–2000 K. By using the normal format for $C_p$ polynomials, we fitted all the data together in two temperature ranges, 273–500 K (Eq. 15) and 500–2000 K (Eq. 16). Likewise, we have calculated $C_p$(Fe(SO$_4$)$_3$) and the results are illustrated in Fig. 3. The optimized $C_p$ polynomials for Fe(SO$_4$)$_3$ in the temperature range 273–2000 K is expressed by Eq. (17).

$$
C_p$(FeSO$_4$), 273–500 K) = 118.67 + (0.047) \cdot T - (24.903) \cdot 10^5 \cdot T^{-2}
- (5.943) \cdot 10^{-6} \cdot T^2, \quad (15)
$$

$$
C_p$(FeSO$_4$, 500–2000 K) = 108.213 + (0.0612) \cdot T - (12.842) \cdot 10^5 \cdot T^{-2}
- (10.980) \cdot 10^{-6} \cdot T^2, \quad (16)
$$

$$
C_p$(Fe$_2$(SO$_4$)$_3$, 273–2000 K) = 305.363 + (0.136) \cdot T - (62.0743) \cdot 10^5 \cdot T^{-2}
- (20.7454) \cdot 10^{-6} \cdot T^2. \quad (17)
$$
As shown in Fig. 2, at high-temperatures, the $C_p$ values are remarkably higher than the values currently used in the FactSage 7.2 database, which is based on the NIST-JANAF thermochemical data [13].

Recently, Majzlan et al. [7] determined the change in enthalpies of formations of Fe$_2$(SO$_4$)$_3$ in the temperature range of 273–395 K. This value is higher by about $-2.2$ kJ mol$^{-1}$ than the value recommended by NIST-JANAF [13]. In this work, we adopted the value $\Delta H_f^\circ(298.15$ K, Fe$_2$(SO$_4$)$_3$) = $(2585.2 \pm 4.9)$ kJ mol$^{-1}$ reported by Majzlan et al. [7], which was obtained through rigorous experimental studies.

By applying our new database for FeO$_4$–FeSO$_4$–Fe$_2$(SO$_4$)$_3$, $\Delta G_f^\circ(T)$ for an equilibrium reaction Fe$_2$O$_3$ + 3SO$_2$ + 1.5O$_2$ $\rightleftharpoons$ Fe$_2$(SO$_4$)$_3$ was calculated in the temperature range of 760–1000 K, in which all the available experimental literature data fall. The obtained result is illustrated in Fig. 4 together with the selected experimental literature data presented in Table 2. Except for the values reported by Sadakane et al. [43] and Alcock et al. [40], the literature values are in agreement with our calculation.

![Fig. 4 $\Delta G_f^\circ$ versus $T$ diagram for the ferric sulfate formation reaction together with the literature experimental data. The solid line is calculated with the database developed in this work and the dashed line is calculated using the pure substances database in the commercial version of FactSage 7.2 [1]]
Fig. 5  Calculated potential phase diagram of Fe−O2−SO2 at P(SO2) = 1 atm

Potential Phase Diagrams

Based on the results obtained for FeSO4 and Fe2(SO4)3, in this work, and the literature data for the other phases, chemical potential phase diagram has been calculated for the Fe−O2−SO2 system between 580 and 1250 K.

The potential phase diagram of log10 P(O2) versus T shown in Fig. 5 was calculated at P(SO2) = 1 atm and P_{Total} = 2 atm. The results obtained are in good agreement with those of Shishin et al.'s [4], hence they may have also used a total pressure close to 2 atm in their calculations.

By graphical extrapolation of the decomposition pressure data measured by Warner and Ingraham [27], we calculated T_{decomp}(Fe2(SO4)3) = 1017 K at which the total pressure of the gaseous products equals one atmosphere. This value is 434 K lower than the value reported in the NIST-JANAF thermochemical data [13] using the same source of experimental data. This large deviation indicates that there were most likely calculation errors in the NIST-JANAF thermochemical data [13]. Our calculations are in agreement with the results reported by [7, 31] and presented in Table 2.

Summary and Conclusions

The development of an oxy-sulfate thermodynamic database for a better understanding and control of SO2 (gas) emission during oxy-fuel combustion processes is essen-
tial. Therefore, the focus of this research was on the thermodynamic modelling of the iron oxide–sulfate system with the FactSage 7.2 software package. Thermodynamic properties of selected phases in the FeOx–FeSO₄–Fe₂(SO₄)₃ system were critically reviewed, compiled and assessed over a wide temperature range of 298–2000 K to obtain accurate thermodynamic parameters. The obtained results were generally in agreement with the literature values in the low-temperature conditions and deviate in the high-temperature conditions. New high-temperature $C_p$ functions, which include the recent experimental data, were optimized. Regarding the decomposition temperatures of FeSO₄ and Fe₂(SO₄)₃, large deviation were observed among the literature values. This warrants that there is a need for new experiments to accurately determine the decomposition temperatures. According to our assessment of the available literature values, we recommend 944 K for the decomposition of FeSO₄ and 1017 K for the decomposition of Fe₂(SO₄)₃ according to the reactions FeSO₄ $\rightarrow$ FeO + SO₃(g) and Fe₂(SO₄)₃ $\rightarrow$ Fe₂O₃ + 3SO₃(g), respectively, at ambient pressure condition.

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References

12. Thomsen J (1880) Thermochemische untersuchungen. Verlag von Johann Ambrosius Barth, Leipzig
21. Grünzweig M (1913) PhD dissertation, Darmstadt, Germany
22. Blanks RF (1961) PhD dissertation, University of Melbourne, Australia
23. Ingraham TR (1963) Internal Rept. EMT-63-17, Department of Mines and Technical Surveys, Ottawa, Canada
37. Schaefer SC (1980) Electrochemical determination of gibbs energies of formation of manganese sulfide and iron sulfide (Fe3S8), RJ, 8486, Albany Research Center Bureau of Mines, Albany, OR