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## Electrocatalysis of Hydrogen Peroxide Reactions on Perovskite Oxides: Experiment versus Kinetic Modeling

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### Abstract

Hydrogen peroxide has been identified as a stable intermediate of the electrochemical oxygen reduction reaction on various electrodes including metal, metal oxide and carbon materials. In this article we study the hydrogen peroxide oxidation and reduction reactions in alkaline medium using a rotating disc electrode (RDE) method on oxides of the perovskite family ( $\text{LaCoO}_3$ ,  $\text{LaMnO}_3$  and  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$ ) which are considered as promising electrocatalytic materials for the cathode of liquid and solid alkaline fuel cells. The experimental findings, such as the higher activity of Mn- compared to Co-perovskites, the shape of RDE curves, and the influence of the  $\text{H}_2\text{O}_2$  concentration, are rationalized with the help of a microkinetic model.

*Key Words:* Hydrogen peroxide reduction reaction (HPRR); hydrogen peroxide oxidation reaction (HPOR); perovskite oxides; kinetic modeling; reaction mechanism; rotating disc electrode (RDE)

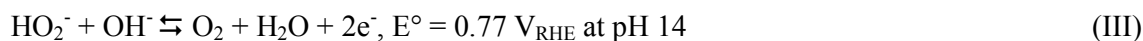
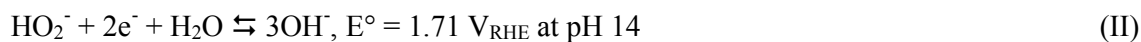
## 1. Introduction

Hydrogen peroxide has been identified as a stable intermediate of the electrochemical oxygen reduction reaction (ORR) on various electrodes including metal, metal oxide and carbon materials. In this article we study the hydrogen peroxide oxidation and reduction reactions (HPOR/HPRR) in alkaline medium on oxides of the perovskite family which are considered as promising electrocatalytic materials for the cathode of liquid and solid alkaline fuel cells. While several researchers studied the catalytic hydrogen peroxide decomposition kinetics on perovskite oxides, very few published results can be found for the study of the electrochemical HPOR/HPRR on perovskite oxides [1-3], and these are unfortunately not sufficiently detailed.

Considering the value of  $pK_A$  of  $H_2O_2$  (11.7 [4]), in alkaline media it transforms into  $HO_2^-$ :



Four electrode reactions can theoretically occur in the presence of  $HO_2^-$ :  $HO_2^-$  reduction into  $OH^-$  (II),  $HO_2^-$  oxidation into  $O_2$  (III), and the inverse reactions,  $OH^-$  oxidation into  $HO_2^-$  (-II) and  $O_2$  reduction into  $HO_2^-$  (-III).



Given the standard potential of  $HO_2^-/OH^-$  couple, the  $OH^-$  oxidation into  $HO_2^-$  (step -II) is unlikely in the potential range of interest for a cathode of an alkaline fuel cell, and will be neglected in this work. Thus, the current-potential curve over a perovskite electrode in a  $H_2O_2$  containing electrolyte is essentially determined by reactions II, III, and -III.

In this work, the HPOR and HPRR are studied using the rotating disc electrode (RDE) method on two oxides of the perovskite family,  $LaCoO_3$  and  $La_{0.8}Sr_{0.2}MnO_3$ , which demonstrated distinct behavior in the ORR (see our recent work in Ref. [5]), as well as on  $LaMnO_3$ . The activity of the oxide materials is compared with that of Pt/C, a reference material in fuel cell electrocatalysis, and with carbon. The latter is an essential component of transition metal oxide-based cathodes improving their electronic conductivity, and thus the degree of the catalyst utilization. To better understand the mechanism of the HPOR and HPRR, and the ORR, on perovskite oxides, a kinetic model is proposed, and the experimental RDE curves are compared with the simulated ones. The study of the influence of the  $H_2O_2$

concentration on the HPOR and HPRR allows one to shed light on the reaction kinetics and helps to validate the model assumptions.

## **2. Experimental**

LaCoO<sub>3</sub>, LaMnO<sub>3</sub> and La<sub>0.8</sub>Sr<sub>0.2</sub>MnO<sub>3</sub> were synthesized at the Moscow State University using polyacrylamide gel with a soft chemistry method, which offers materials with higher surface areas than conventional high temperature approaches. For the synthesis procedure the reader is referred to Ref. [5]. Specific surface areas of LaCoO<sub>3</sub>, LaMnO<sub>3</sub> and La<sub>0.8</sub>Sr<sub>0.2</sub>MnO<sub>3</sub> measured by BET were 10, 14 and 17 m<sup>2</sup> g<sup>-1</sup>, respectively. Oxides were mixed with carbon and deposited on a glassy carbon RDE (0.07 cm<sup>2</sup> geometric area, Autolab) in the form of a thin layer [5]. Carbon of the Sibunit family with the BET surface area of 65.7 m<sup>2</sup> g<sup>-1</sup> was chosen for its high purity and high electron conductivity [6]. In order to improve the adhesion of the perovskite and carbon particles onto the GC support, an alkaline ionomer (AS-4 from Tokuyama Company) was deposited on top of the dry catalyst layer to form a thin film (the ionomer loading was 0.3 mg cm<sup>-2</sup>). For further details on the electrode preparation the reader is referred to Ref. [5]. For all oxide/carbon loadings utilized in this work the catalytic layer thickness was smaller than the diffusion layer thickness.

Pt/C electrodes were prepared by drop-casting a suspension containing Pt/C (40 wt. % Pt on carbon black, Alfa Aesar) to obtain 20 μg cm<sup>-2</sup> Pt loading. The roughness factor of Pt on the electrode estimated using the coulometry of the hydrogen underpotential deposition was equal to 6.

Electrochemical measurements were performed at 25°C in a three electrode cell whose parts in contact with the electrolyte were made out of Teflon. The electrolyte was 1 M NaOH prepared from extra pure NaOH solution (50 wt. % solution in water, Acros Organics) and ultrapure water (Purelab: 18.2 MΩ cm, < 3 ppb TOC). The counter electrode was a platinum wire and the reference electrode was a Hg/HgO/ 1 M NaOH electrode (IJ Cambria Scientific). In what follows the electrode potentials are given in the RHE (reversible hydrogen electrode) scale. RDE curves were IR-corrected by using the value of the electrolyte resistance (15 Ohm) determined from the high frequency part of the electrochemical impedance spectra (measured in the 1 Hz to 100 kHz range). Electrochemical measurements were performed using Autolab potentiostat with an analog scan generator at the scan rate was 10 mV s<sup>-1</sup>. H<sub>2</sub>O<sub>2</sub> solutions were prepared from 30 wt% solution in water (SupraPur, Merck) titrated with standardized KMnO<sub>4</sub>.

## **3. Results and Discussion**

Figure 1a shows positive scans of RDE voltammograms in N<sub>2</sub>-purged 1 M NaOH containing 0.84 mM H<sub>2</sub>O<sub>2</sub> for LaCoO<sub>3</sub>/carbon, LaMnO<sub>3</sub>/carbon and La<sub>0.8</sub>Sr<sub>0.2</sub>MnO<sub>3</sub>/carbon composite electrodes, as well as for Pt/C and Sibunit carbon. Such H<sub>2</sub>O<sub>2</sub> concentration was chosen since it corresponds to the O<sub>2</sub> concentration in O<sub>2</sub> saturated electrolyte. For Pt/C in agreement with the literature data [7-10] the voltammogram is reversible, suggesting fast hydrogen peroxide oxidation and reduction reactions, and the anodic and cathodic current plateaus correspond to the diffusion limited HPOR and HPRR, as confirmed by the Levich-Koutecky (LK) plots in Figures 1b and 1c. For Sibunit carbon, the voltammogram is strongly irreversible, pointing to a very slow HPOR and an even slower HPRR (Figure 1, grey curve). For oxide/carbon composites, both the anodic and the cathodic waves have smaller slopes compared to Pt, pointing out at slower electrocatalysis of hydrogen peroxide reactions on studied perovskite oxides. Furthermore, the mixed potential values, where the sum of currents of all electrode reactions (reactions II/-II and III/-III) is equal to zero, for perovskite oxides are shifted positive compared to those of Pt/C, suggesting a slower HO<sub>2</sub><sup>-</sup> oxidation, notably for LaCoO<sub>3</sub>. In the absence of carbon, perovskite oxides show slower hydrogen peroxide oxidation/reduction kinetics (Figure S1). However, contrary to what has recently been demonstrated for the ORR on oxides of the perovskite family, where carbon plays a dual role, (i) improving the interparticle contact, and also (ii) participating in the ORR electrocatalysis, the enhancement of the rates of HPOR and HPRR is predominantly due to the first factor (improvement of the electronic conductivity of the catalytic layer). This is due to the low catalytic activity of carbon in HO<sub>2</sub><sup>-</sup> reactions (*cf.* Figure 1a, grey curve). Note also that for LaMnO<sub>3</sub> the effect of the carbon addition on the RDE curves is stronger compared to that for the La<sub>0.8</sub>Sr<sub>0.2</sub>MnO<sub>3</sub>. This can be attributed to a higher intrinsic electron conductivity of the doped compared for the undoped Mn-based perovskite (for further details see Supplementary information).

The Mn-based LaMnO<sub>3</sub> and La<sub>0.8</sub>Sr<sub>0.2</sub>MnO<sub>3</sub> perovskites display close mixed potentials (similar trends can be found in the literature for La<sub>1-x</sub>Sr<sub>x</sub>MnO<sub>3</sub> [1]), and very similar activities for HO<sub>2</sub><sup>-</sup> reduction/oxidation, once sufficient amount of carbon is present in the catalytic layer (Figure 1a). Moreover, their HO<sub>2</sub><sup>-</sup> reduction and oxidation currents approach the diffusion plateau reached by Pt/C electrode at low and high potentials, respectively (note however that at 1.2 V the intercepts of the LK plots are non-zero). This suggests that on Mn perovskite oxides at high overpotentials the HPOR and HPRR are diffusion-limited. For LaCoO<sub>3</sub> electrodes, the currents do not reach the plateaus of Pt/C, due to a slower HO<sub>2</sub><sup>-</sup> reduction and

oxidation kinetics. The comparison of the current slope near the mixed potential for various perovskites (Figure 1a) confirms that  $\text{LaCoO}_3$  is less catalytically active in  $\text{HO}_2^-$  reactions than either  $\text{LaMnO}_3$  or  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$ . Since the mixed potential of  $\text{LaCoO}_3$  is positively shifted compared to Mn perovskite oxides, we conclude that the HPOR is slower than the HPRR on the cobalt perovskite. The cathodic branch of the RDE voltammograms for  $\text{LaCoO}_3$  shows two slopes (Figure 1a) suggesting that at least two steps are involved in the  $\text{HO}_2^-$  reduction reaction. In what follows this hypothesis will be corroborated with the help of a mathematical model.

The influence of the  $\text{H}_2\text{O}_2$  concentration on the HPOR and HPRR kinetics is shown in Figure 2a for  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$ . While the absolute values of the limiting anodic and cathodic currents increase proportionally to the  $\text{H}_2\text{O}_2$  concentration, the mixed potential is displaced towards the negative values (see close-up in Figure 2c). The same behavior was observed for  $\text{LaCoO}_3$  (not shown). This indicates that the perovskite oxide-catalyzed  $\text{HO}_2^-$  reduction and the  $\text{HO}_2^-$  oxidation reactions have different concentration dependences.

In order to better understand the experimental findings, a microkinetic model was developed. Note that the mathematical model serves to understand the influence of the experimental parameters (potential and concentration), the differences between Mn and Co perovskites, and to verify the consistency of the proposed reaction mechanism with the experimental data, rather than to accurately determine the values of the rate constants. The elementary steps for the  $\text{HO}_2^-$  and  $\text{O}_2$  adsorption and reaction on perovskite oxides were inspired by the experimental studies of Goodenough [11] and Suntivich et al. [12] and the density functional theory (DFT) calculations of Wang and Cheng [13], and adapted in order to reproduce the experimental findings of this work.

Since the oxide surface in an alkaline electrolyte is supposed to be covered by  $\text{OH}_{\text{ad}}$  [11, 14], displacement of  $\text{OH}_{\text{ad}}$  by either  $\text{O}_2$  or  $\text{HO}_2^-$  is likely to be the first step in the electrocatalysis of  $\text{O}_2$  [11] and  $\text{H}_2\text{O}_2$  reactions.



The adsorbed  $\text{HO}_{2,\text{ad}}$  may be oxidized into oxygen in the reverse step of reaction (2):



The adsorbed oxygen molecule resulting from the backward step of reaction (2) may desorb in the reverse step of reaction (3). In agreement with the literature [11, 13], we suppose that upon adsorption oxygen molecules displace hydroxo species on the oxide surface:



The adsorbed  $\text{HO}_{2,\text{ad}}$  can also undergo reduction. To account for the negative shift of the mixed potential with the increase of the  $\text{H}_2\text{O}_2$  concentration, the HPOR and the HPRR must have different concentration dependence. In order to reproduce this, we assume that the  $\text{HO}_{2,\text{ad}}$  reduction occurs in a sequence of a chemical and an electrochemical steps:



Note that steps (1) – (5) also account for the catalytic decomposition of  $\text{H}_2\text{O}_2$  occurring under open circuit conditions.

Considering both the experimental and the literature data we suppose that  $\text{O}_{\text{ad}}$  species are adsorbed on transition metal (B) cations in the high oxidation state  $\text{B}^{(\text{m}+1)+}$  ( $\text{B}=\text{O}$ ), while  $\text{OH}_{\text{ad}}$  species - on B-cations in the low oxidation state  $\text{B}^{\text{m}+}$  ( $\text{B}-\text{OH}$ ). According to the DFT calculations of Wang and Cheng [13], these intermediates ( $\text{OH}_{\text{ad}}$  and  $\text{O}_{\text{ad}}$ ) are indeed strongly adsorbed on the surface of perovskite oxides.

Assuming Langmuir-type adsorption isotherms and Butler-Volmer type of the electrochemical kinetics, the reaction rates may be expressed as follows:

$$v_1 = k_1 c_{\text{HO}_2}^{\text{S}} (1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}}) - k_{-1} \theta_{\text{HO}_2} \quad (6)$$

$$v_2 = k_2 \theta_{\text{O}_2} \exp\left(\frac{-(1-\alpha)FE}{RT}\right) - k_{-2} \theta_{\text{HO}_2} \exp\left(\frac{\alpha FE}{RT}\right) \quad (7)$$

$$v_3 = k_3 c_{\text{O}_2}^{\text{S}} (1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}}) \exp\left(\frac{-(1-\alpha)FE}{RT}\right) - k_{-3} \theta_{\text{O}_2} \exp\left(\frac{\alpha FE}{RT}\right) \quad (8)$$

$$v_4 = k_4 \theta_{\text{HO}_2} (1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}}) \quad (9)$$

$$v_5 = k_5 \theta_{\text{O}} \exp\left(\frac{-\alpha FE}{RT}\right) - k_{-5} (1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}}) \exp\left(\frac{(1-\alpha)FE}{RT}\right) \quad (10)$$

where  $c_{\text{HO}_2}^{\text{S}}$ ,  $c_{\text{O}_2}^{\text{S}}$  are the concentrations in front of the electrode at  $x=0$  ( $x$  is the distance along the normal to the electrode surface) and  $\theta_{\text{HO}_2}$ ,  $\theta_{\text{O}_2}$ ,  $\theta_{\text{O}}$  the coverages of adsorbed intermediates, with  $\theta_{\text{OH}} = 1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}}$ .  $\alpha$  is the charge transfer coefficient,  $F=96486 \text{ C mol}^{-1}$  is the Faraday constant,  $R=8.314 \text{ J K}^{-1} \text{ mol}^{-1}$  the ideal gas constant and  $T=300 \text{ K}$  the absolute temperature.

Under stationary conditions, this reaction scheme can be cast into a set of equations determining the concentrations of  $\text{HO}_2^-$  and  $\text{O}_2$  in front of the electrode surface, and the coverages of adsorbed intermediates,  $\text{HO}_{2,\text{ad}}$ ,  $\text{O}_{2,\text{ad}}$ ,  $\text{O}_{\text{ad}}$ , and  $\text{OH}_{\text{ad}}$  on perovskite sites:

$$D_{O_2} \left( \frac{\partial c_{O_2}}{\partial x} \right) = -\frac{D_{O_2} c_{O_2}^S}{\delta_{O_2}} = \Gamma_{geo} v_3 \quad (11)$$

$$D_{HO_2^-} \left( \frac{\partial c_{HO_2^-}}{\partial x} \right) = \frac{D_{HO_2^-} (c_{HO_2^-}^b - c_{HO_2^-}^S)}{\delta_{HO_2^-}} = \Gamma_{geo} v_1 \quad (12)$$

$$\frac{d\theta_{O_2}}{dt} = v_3 - v_2 = 0 \quad (13)$$

$$\frac{d\theta_{HO_2}}{dt} = v_1 + v_2 - v_4 = 0 \quad (14)$$

$$\frac{d\theta_O}{dt} = 2v_4 - v_5 = 0 \quad (15)$$

where  $\delta_i$  is the diffusion layer thickness,  $c_{HO_2^-}^b$ , the concentration of  $HO_2^-$  in the bulk of the electrolyte, and  $\Gamma_{geo}$  stands for the number of active sites per geometric area of the electrode. The overall current density is given by equation (16).

$$j = -F\Gamma_{geo}(v_2 + v_3 + v_5) \quad (16)$$

A linear concentration profile for  $O_2$  and  $HO_2^-$  is assumed in front of the electrode surface. The electrode thickness is assumed to be sufficiently small to keep this linear concentration profile. The diffusion coefficients of  $O_2$  and  $HO_2^-$  in 1 M NaOH were determined as  $D_{O_2} = 1.5 \cdot 10^{-5} \text{ cm}^2 \text{ s}^{-1}$  and  $D_{HO_2^-} = 0.8 \cdot 10^{-5} \text{ cm}^2 \text{ s}^{-1}$ , respectively, from the corresponding diffusion limited current plateaus for Pt/C electrodes. Thus, the diffusion layer thickness at 900 rpm is  $\delta_{O_2} = 19 \text{ }\mu\text{m}$  for  $O_2$  and  $\delta_{HO_2^-} = 15 \text{ }\mu\text{m}$  for  $HO_2^-$ .

Assuming that B cations act as active sites, the number of active sites per geometric surface area  $\Gamma_{geo}$  was calculated from the catalyst loading and the number of B sites per unit of the real surface area, which was estimated from the crystalline structure of perovskite oxides as  $4.14 \cdot 10^{-10} \text{ mol cm}^{-2}_{\text{oxide}}$ . The values of the reaction rate constants were adjusted in order to reproduce the main features of the experimental RDE voltammograms for HPOR and HPRR, as well as for the ORR [5], and are given in Table 1. Symmetry factors were all taken equal to 0.5, which is realistic for low and moderate overvoltages considered in this work. The adsorption/desorption rates of  $HO_2^-$  on perovskite sites directly depend on the values of  $k_1$  and  $k_{-1}$  which have to be taken sufficiently high for adsorption and low for desorption to account for the significant activity of perovskites in the HPRR.  $k_2$  and  $k_3$  were adjusted to reproduce both the onset potential and the Tafel slopes for the ORR on perovskite oxides [5], while the rate constants of the reverse reactions  $k_{-2}$ ,  $k_{-3}$  were chosen to account for the standard potential of the  $O_2/HO_2^-$  at 0.77 V vs. RHE. Finally, the values of the mixed potential for the



HPRR/HPOR and the potential of the  $B^{(m+1)+} / B^{(m)+}$  redox peaks (see CVs in Ref. [5]) were used to adjust the values of  $k_4$ ,  $k_5$  and  $k_{-5}$ .

Figure 3 shows simulated RDE voltammograms for  $\text{LaCoO}_3$  and  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  oxides. With the rate constants listed in Table 1, an almost reversible voltammogram is obtained for  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  (Figure 3b). By choosing smaller rate constants for steps 2, 3 and 5, we are able to reproduce a slower HPOR on  $\text{LaCoO}_3$  (*cf.* Figure 3a and Figure 3b) and a positive shift of the mixed potential compared to  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$ . Furthermore, the assumption of the  $\text{HO}_{2,\text{ad}}$  reduction as a sequence of two elementary steps allows us to reproduce the observed change of the slope of the HPRR on  $\text{LaCoO}_3$ . Figure 3 also shows the splitting of the RDE voltammograms into a sum of three main contributions for two  $\text{HO}_2^-$  concentrations. The anodic branch corresponds to the  $\text{HO}_2^-$  oxidation into  $\text{O}_2$  (step III), while the cathodic branch consists of two contributions: the  $\text{HO}_2^-$  reduction into  $\text{OH}^-$  (step II) and the  $\text{O}_2$  reduction (this oxygen is formed in the anodic branch) into  $\text{HO}_2^-$  (step -III). One may note that the differences between  $\text{LaCoO}_3$  and  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  are partly due to the lower ORR activity of  $\text{LaCoO}_3$  (*cf.* blue curves in Figure 3a and 3b). For experimental ORR data on  $\text{LaCoO}_3$  and  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  the reader is referred to our earlier publication [5].

The influence of the  $\text{H}_2\text{O}_2$  concentration can be traced from simulated voltammograms in Figure 3b and Figure 2b,d. Increasing the concentration directly affects the reaction rate of the adsorption/desorption of  $\text{HO}_2$  (reaction 1) and therefore the site coverage by  $\text{HO}_{2,\text{ad}}$  species ( $\theta_{\text{HO}_2}$ ). In the anodic direction, an increase of  $\theta_{\text{HO}_2}$  causes an increase in the rate (step -2) of formation of  $\text{O}_{2,\text{ad}}$  from  $\text{HO}_{2,\text{ad}}$  (*cf.* dashed over solid green lines in Figure 3b). On the other hand, in the cathodic branch, an increase of the  $\text{HO}_2^-$  concentration does not lead to a strong acceleration of reaction 4, since its rate depends on both  $\theta_{\text{HO}_2}$  and  $\theta_{\text{OH}} = (1 - \theta_{\text{O}_2} - \theta_{\text{HO}_2} - \theta_{\text{O}})$  (*cf.* dashed over solid magenta lines in Figure 3b). A stronger concentration dependence of the  $\text{HO}_2^-$  oxidation compared to the reduction counterpart, and the potential dependence of the former result in the negative shift of the mixed potential with the  $\text{HO}_2^-$  concentration (Figure 3b, black curves, and Figure 2b,d).

The deconvolution of the whole current into individual contributions shows that the kinetic currents of the  $\text{HO}_2^-$  reduction/oxidation cannot be directly obtained from the total currents, contrary to what has been suggested in the literature [15].

The main difference between the simulated and the experimental curves is the decrease of the  $\text{HO}_2^-$  oxidation current at high potentials for the former. It is caused by the oxidation of  $B^{m+}$  cations (which in our model are required for the adsorption of  $\text{HO}_2^-$ ) into  $B^{(m+1)+}$  cations at

high potentials. Various possible explanations may be offered in order to account for the absence of such a current drop in the experiment, among these (i) surface heterogeneity resulting in a wide potential distribution of the  $B^{m+}/B^{(m+1)+}$  red-ox transitions, (ii) adsorption of  $HO_2^-$  on  $B^{(m+1)+}$  sites as well, (iii) carbon contribution to the HPOR at high electrode potentials (see grey curve in Figure 1).

#### **4. Conclusions**

In this work, it is demonstrated that the  $LaCoO_3$ ,  $LaMnO_3$  and  $La_{0.8}Sr_{0.2}MnO_3$  perovskite oxides, when mixed with carbon, show significant activity towards hydrogen peroxide reduction/oxidation reactions, with  $LaCoO_3$  being less active than Mn perovskites. The kinetics of the hydrogen peroxide oxidation/reduction reactions on perovskite oxides can be reasonably described with a simple kinetic model. Modeling helps to rationalize the observed differences between Co and Mn perovskite oxides in  $H_2O_2$  electrocatalysis, which are largely due to the lower electrocatalytic activity of Co perovskite for the  $H_2O_2$  oxidation and  $O_2$  reduction. Furthermore, it allows one to explain the  $H_2O_2$  concentration dependence (including the displacement of the mixed potential with the concentration) by the multistep mechanism of the hydrogen peroxide reduction. The conclusions obtained have far-reaching consequences for the understanding of the ORR electrocatalysis on oxides of the perovskite family and will be further corroborated in our forthcoming publication.

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Table 1 : Rate constants used in the mathematical model to simulate the experimental curves of  $\text{LaCoO}_3$  and  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  electrodes

Rate constant	Units	Value	
		$\text{LaCoO}_3$	$\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$
$k_1$	$\text{cm}^3 \text{mol}^{-1} \text{s}^{-1}$	$5 \cdot 10^7$	$5 \cdot 10^7$
$k_{-1}$	$\text{s}^{-1}$	50	50
$k_2$	$\text{s}^{-1}$	$5.1 \cdot 10^8$	$5.1 \cdot 10^9$
$k_{-2}$	$\text{s}^{-1}$	$4.9 \cdot 10^{-8}$	$4.9 \cdot 10^{-7}$
$k_3$	$\text{cm}^3 \text{mol}^{-1} \text{s}^{-1}$	$1.6 \cdot 10^{11}$	$1.6 \cdot 10^{12}$
$k_{-3}$	$\text{s}^{-1}$	$1.6 \cdot 10^{-8}$	$1.6 \cdot 10^{-7}$
$k_4$	$\text{s}^{-1}$	40	40
$k_5$	$\text{s}^{-1}$	$6 \cdot 10^7$	$1.2 \cdot 10^8$
$k_{-5}$	$\text{s}^{-1}$	$2.9 \cdot 10^{-9}$	$5.8 \cdot 10^{-9}$

### Figure captions

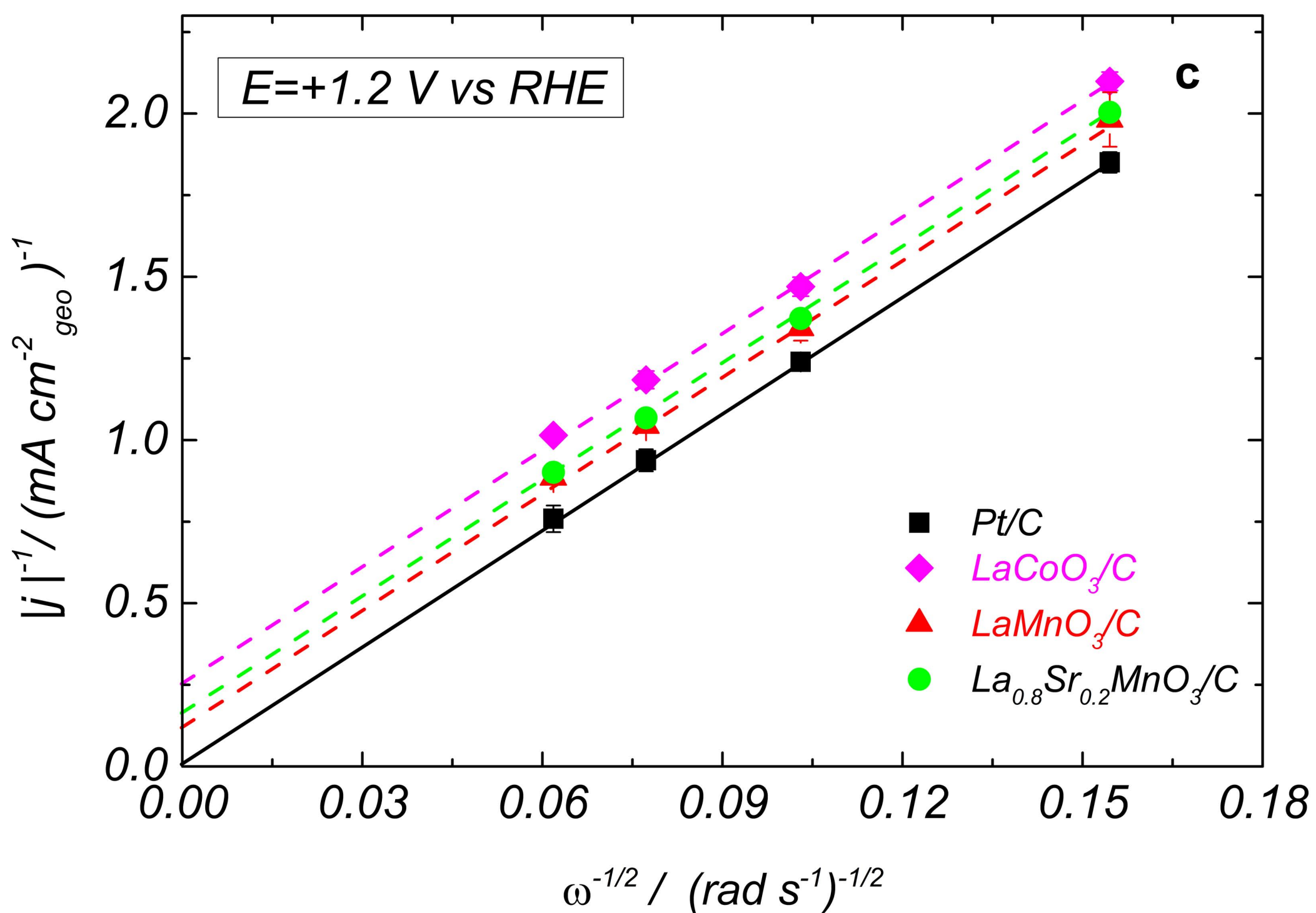
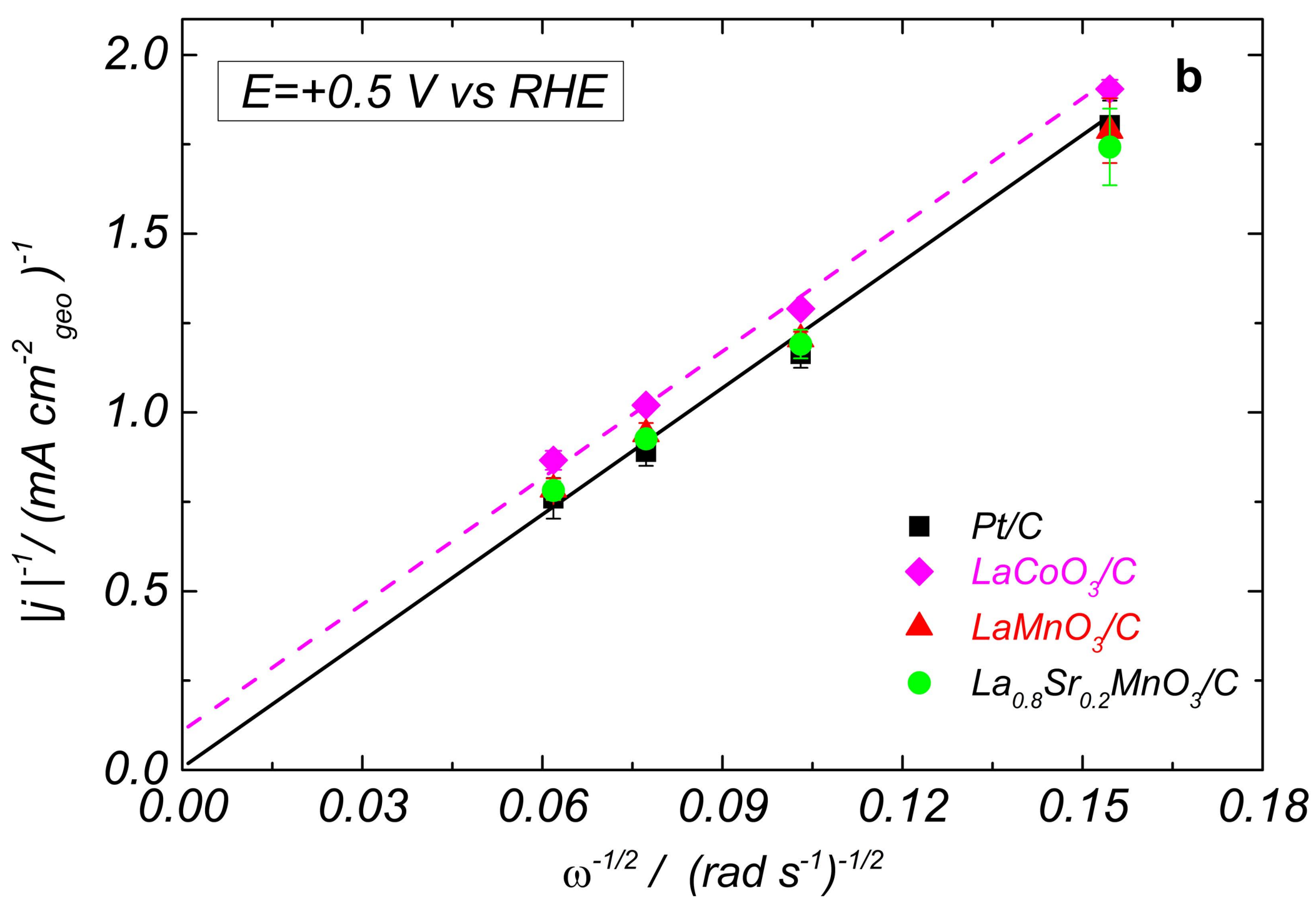
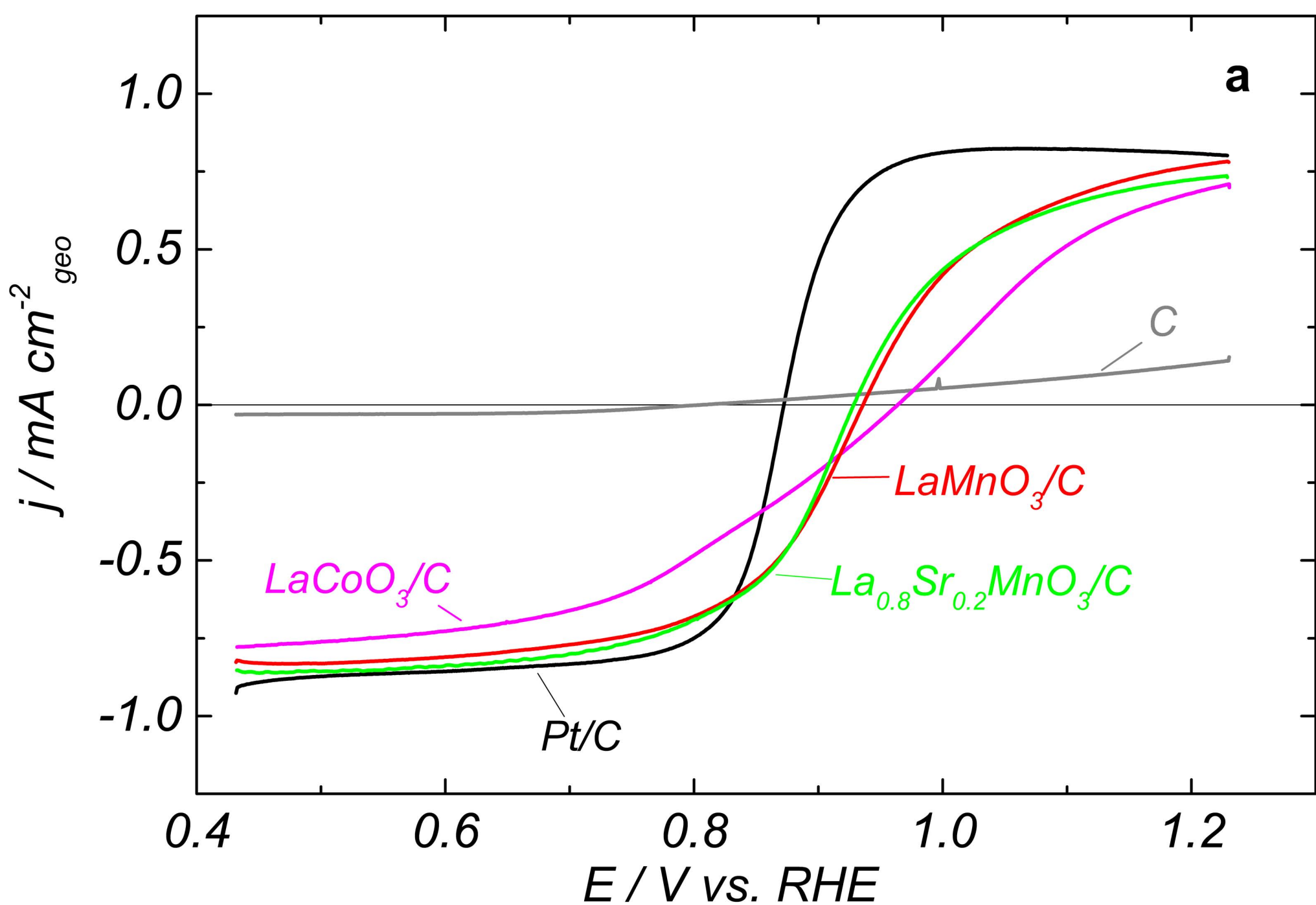
Figure 1 : (a) Positive scans of the RDE voltammograms in  $\text{N}_2$ -purged 1 M NaOH + 0.84 mM  $\text{H}_2\text{O}_2$  at 900 rpm and  $10 \text{ mV s}^{-1}$  and corresponding Levich-Koutecky plots at  $+0.5 V_{\text{RHE}}$  (b) and at  $+1.2 V_{\text{RHE}}$  (c) for GC-supported thin films of  $\text{LaCoO}_3 + \text{C}$  (magenta),  $\text{LaMnO}_3 + \text{C}$  (red),  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3 + \text{C}$  (green), C (grey) and Pt/C (black). Catalyst loadings are  $91 \mu\text{g cm}^{-2}$  perovskite +  $37 \mu\text{g cm}^{-2}$  carbon for oxide electrodes,  $37 \mu\text{g cm}^{-2}$  for carbon electrodes, and  $50 \mu\text{g cm}^{-2}$  for Pt/C electrodes. Currents are normalized to the electrode geometric area and corrected to the background currents measured in the  $\text{N}_2$  atmosphere. Error bars in panels (b) and (c) represent standard deviations from at least two independent repeated measurements.

Figure 2 : (a) Experimental positive scans of the RDE voltammograms in  $\text{N}_2$ -purged solutions of 1 M NaOH and various  $\text{H}_2\text{O}_2$  concentrations at 900 rpm and  $10 \text{ mV s}^{-1}$  for GC-supported thin films of  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  ( $91 \mu\text{g cm}^{-2}$ ) + carbon ( $74 \mu\text{g cm}^{-2}$ ). Currents are normalized to the geometric area of the electrode and corrected to the background currents measured in  $\text{N}_2$  atmosphere. (b) Simulated RDE voltammograms. (c) and (d) show close-ups for experimental (c) and simulated (d) voltammograms. For details see the text.

Figure 3 : Simulated RDE voltammograms for (a)  $\text{LaCoO}_3$  and (b)  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$  in 0.84 mM  $\text{H}_2\text{O}_2$  (solid lines) and 1.68 mM  $\text{H}_2\text{O}_2$  (dashed lines) at 900 rpm. Number of sites per geometric surface area is  $3.8 \cdot 10^{-9} \text{ mol cm}^{-2}_{\text{geo}}$  for  $\text{LaCoO}_3$  and  $6.4 \cdot 10^{-9} \text{ mol cm}^{-2}_{\text{geo}}$  for  $\text{La}_{0.8}\text{Sr}_{0.2}\text{MnO}_3$ . Color codes for the individual contributions:  $\text{HO}_2^-$  reduction (magenta),  $\text{O}_2$  reduction (blue),  $\text{HO}_2^-$  oxidation (green), and total current obtained by the addition of the previously mentioned contributions (black). Currents are calculated per geometric area of the electrode.

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## EXPERIMENT

## MODELING

