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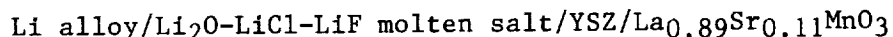
THE ELECTROCHEMICAL GENERATION OF USEFUL CHEMICAL SPECIES FROM LUNAR MATERIALS\*

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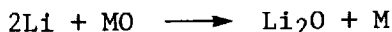
Electrochemical cells have been fabricated for the simultaneous generation of oxygen and lithium from a Li<sub>2</sub>O containing molten salt (Li<sub>2</sub>O-LiCl-LiF). The cell utilizes an oxygen vacancy conducting solid electrolyte, yttria-stabilized zirconia (YSZ), to effect separation between oxygen evolving and lithium reduction half-cell reactions. The cell, which operates at 700-850°C, possesses rapid electrode kinetics at the lithium-alloy electrode with exchange current density (i<sub>0</sub>) values being >60mA/cm<sup>2</sup>. When used in the electrolytic mode, lithium produced at the negative electrode would be continuously removed from the cell for later use (under lunar conditions) as an easily storable reducing agent (compared to H<sub>2</sub>) for the chemical refining of lunar ores. Because of the high reversibility of this electrochemical system, it has also formed the basis for the lithium oxygen secondary battery system which possesses the highest theoretical energy density yet investigated.

INTRODUCTION

The strategy being pursued for lunar ore refining is based upon electrochemical cells possessing the general configuration:



In practical electrolytic cells, lithium produced at the negative electrode would, after removal from the cell, be available for lunar ore refining via the general chemical reaction:

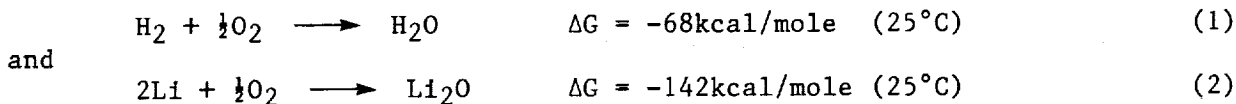


where MO represents a lunar ore. Emphasis to this time has been on Fe<sub>2</sub>O<sub>3</sub>, TiO<sub>2</sub> and the lunar ore ilmenite (FeTiO<sub>3</sub>), all of which have been shown chemically reducible by Li to give metals. The resulting Li<sub>2</sub>O reaction product could then be removed from the solid-state reaction mixture by sublimation and reintroduced into the negative electrode compartment of the electrolytic cell to be electrolyzed again. Hence, this electrochemical approach provides a convenient route for the simultaneous generation of both metals and oxygen from lunar materials on the Moon's surface. It has previously been suggested that oxygen might be extracted from ilmenite by its initial chemical reduction by hydrogen, initially transported from Earth (ref. 1-3), and presumably requiring cryogenic storage on the Moon's surface. Other approaches discussed for the chemical reduction of lunar ores have included carbothermic reduction to give the desired metal (ref. 4,5). Previous work by others has also investigated the direct high temperature electrochemistry of simulated lunar materials (molten

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silicates), using platinum electrodes. Electrolysis of these melts demonstrated simultaneous oxygen evolution at the anode (ref. 6-10) and deposition of an impure metal-silicon alloy slag at the cathode (ref. 11-14). Nevertheless, this preliminary work showed feasibility for high temperature molten salt electrochemistry as an attractive approach for oxygen evolution and metal reduction from lunar type ores. In the approach being discussed here, oxygen evolution and lunar ore reduction are separated into distinct electrochemical and chemical steps. This strategy minimizes the opportunity for lunar ore originating impurities from entering the electrochemical cell and degrading performance.

The reducing power of lithium is significantly greater than that of hydrogen. This fact can be summarized by the two general reactions:

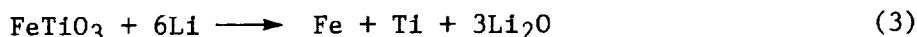


The corresponding free energy of reaction between lithium and candidate metal oxides of interest at 1000°K are summarized in Table 1.

TABLE 1

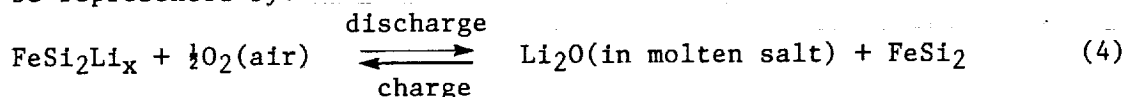
	$\Delta G(\text{kcal/mole})$
$4\text{Li} + \text{SiO}_2 \rightarrow 2\text{Li}_2\text{O} + \text{Si}$	-28.0
$2\text{Li} + \text{FeO} \rightarrow \text{Li}_2\text{O} + \text{Fe}$	-53.9
$4\text{Li} + \text{TiO}_2 \rightarrow 2\text{Li}_2\text{O} + \text{Ti}$	-15.6
$6\text{Li} + \text{Cr}_2\text{O}_3 \rightarrow 3\text{Li}_2\text{O} + 2\text{Cr}$	-87.9

We estimate the free energy change associated with ilmenite reduction by lithium at 1000°K, corresponding to the overall reaction:



to be  $\Delta G = -62.2\text{kcal/mole}$ . Hence, lithium would appear to be a sufficiently strong reducing agent for many of the transition metal oxides anticipated to be present on the Moon's surface.

The electrochemistry occurring in these cells possesses high reversibility and consequently has given us the opportunity to study the lithium/oxygen secondary battery. The overall electrochemistry occurring in this cell upon discharge/charge cycling can be represented by:



with overall capacity being dictated by the concentration of  $\text{Li}_2\text{O}$  that may be incorporated within the negative electrode compartment of this electrochemical cell. Although theoretical energy densities for a given battery system are often a poor indicator of the final practical energy density that might be achieved, it is of interest to note that the energy density for electroactive materials in the lithium-oxygen system calculates to 4082Wh/kg compared to 2211Wh/kg for the lithium-chlorine cell; the highest energy density secondary system investigated to this time. The significant technical observations to be discussed which make this an attractive secondary battery system are i) both electrodes possess high electrochemical reversibility,

ii) the anodic reaction upon cell charge involves the exclusive evolution of oxygen, and iii) a significant concentration of  $\text{Li}_2\text{O}$  can be incorporated into the  $\text{Li}_2\text{O}$ - $\text{LiCl}$ - $\text{LiF}$  ternary melt. A comparison of the lithium-oxygen secondary cell with other major conventions and advanced systems is shown in Table 2.

TABLE 2  
Major secondary batteries for electrical energy storage in comparison to the lithium oxygen system

Battery System	Negative	Positive	OCP(V)	Theoretical Energy Density for Electroactive Matls. (Wh/kg)
Lead-Acid	Pb	$\text{PbO}_2$	2.0	238
Nickel-Cadmium	Cd	$\text{NiOOH}$	1.2	148
Nickel-Zinc	Zn	$\text{NiOOH}$	1.7	448
Zinc-Chlorine	Zn	$\text{Cl}_2$	2.12	829
Lithium-Iron Sulfide	$\text{Li(Al)}$	$\text{FeS}$	1.2	626
Sodium-Sulfur	Na	S	2.1	760
Lithium-Chlorine	Li	$\text{Cl}_2$	3.53	2211
Lithium-Oxygen	Li	$\text{O}_2$	2.3	4082

#### RESULTS AND DISCUSSION

Previous work performed in our laboratory (ref. 15) showed that the ternary molten salt system  $\text{Li}_2\text{O}$ - $\text{LiF}$ - $\text{LiCl}$  is an attractive candidate electrolyte for this electrochemical cell. The utility of this molten salt electrolyte was dependent, however, upon gaining some insight into the amount of  $\text{Li}_2\text{O}$  that could be accommodated in the  $\text{LiF}$ - $\text{LiCl}$  (70<sup>m</sup>/o) molten salt and its impact upon the resulting solid-liquidus curve. This was performed by progressively introducing  $\text{Li}_2\text{O}$  in 2-5<sup>m</sup>/o increments into a  $\text{LiF}$ - $\text{LiCl}$  (70<sup>m</sup>/o) binary mixture in a conventional conductivity type cell. After each  $\text{Li}_2\text{O}$  addition, the ternary molten salt mixture was initially heated to 550°C, at which temperature it became molten. The resulting molten salt ionic conductivity values were measured between stainless steel electrodes (1cm<sup>2</sup>) placed 1.5cm apart during slow cell cooling so as to minimize interelectrode thermal gradients. The molten salt freezing point was detected by a dramatic decrease in measured ionic conductivity. The resulting solid-liquidus curve obtained for this ternary system is shown in Figure 1, which indicates that for  $\text{Li}_2\text{O}$  contents above  $\approx 18^m$ /o the molten salt goes into a two-phase region consisting of  $\text{Li}_2\text{O}$  (18<sup>m</sup>/o),  $\text{LiF}$  (24.6<sup>m</sup>/o),  $\text{LiCl}$  (57.4<sup>m</sup>/o) in direct contact with solid  $\text{Li}_2\text{O}$ . The important point here is that excess  $\text{Li}_2\text{O}$  could be introduced into the electrolyte cell and would not promote melt freezing.

Both half and full cell electrochemical measurements have been performed on this cell. Coulometric lithium deposition has been performed on both iron and iron silicide ( $\text{FeSi}_2$ ) negative electrode substrates. Iron silicide was of interest since lithium deposition results in formation of a series of distinct electrochemically reversible lithium compounds up to the composition  $\text{FeSi}_2\text{Li}_{10}$ . This has not only permitted us to conveniently store lithium in the negative electrode compartment for the reversible lithium oxygen storage battery, but also acts as a substrate for wetting lithium, facilitating its later removal from the negative electrode compartment of the electrolytic cell. Preliminary current-potential characteristics of this electrode were performed in electrolytic cells possessing the general configuration shown in Figure 2. The charging current-potential curves for this cell (Figure 3) clearly show voltage plateaus associated with each ternary lithium compound. Electrode ki-

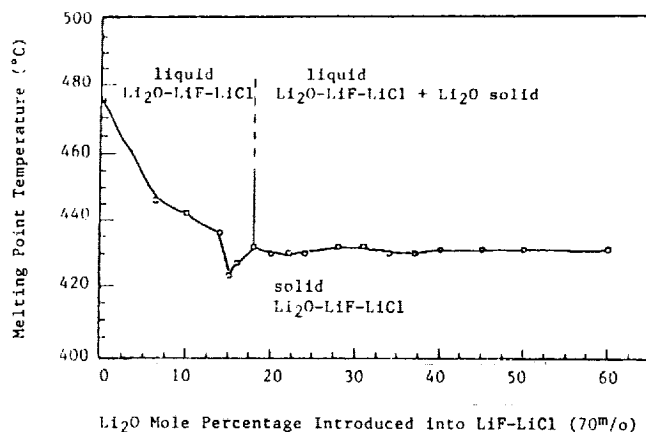


Figure 1. Solid-liquidus curve for  $\text{Li}_2\text{O-LiF-LiCl}$  molten salt as a function of  $\text{Li}_2\text{O}$ .

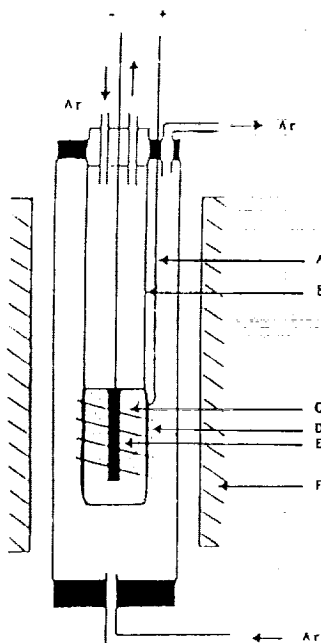


Figure 2. Schematic drawing of cell configuration used for the simultaneous electrolytic generation of lithium and oxygen. A) anode current collector; B) calcia or yttria stabilized zirconia; C) molten salt; D)  $\text{La}_{0.89}\text{Sr}_{0.11}\text{MnO}_3$  anode; E) stainless steel or  $\text{FeSi}_2$  cathode; F) furnace.

netics for this electrode were performed using a molten salt of composition  $\text{Li}_2\text{O}$  (6.4<sup>m/o</sup>),  $\text{LiF}$  (28.1<sup>m/o</sup>),  $\text{LiCl}$  (65.5<sup>m/o</sup>) using  $\text{FeSi}_2\text{Li}_x$  for the working and  $\text{SiLi}_4$  for counter and reference electrodes. For voltage plateaus corresponding to the respective ternary alloys  $\text{FeSi}_2\text{Li}_6$ ,  $\text{FeSi}_2\text{Li}_8$  and  $\text{FeSi}_2\text{Li}_{10}$  cyclic voltammetry scans (50mV/s) were performed  $\pm 200\text{mV}$  from its initial open circuit potential (OCP) with respect to the  $\text{SiLi}_4$  reference. A representative current-overpotential curve is shown in Figure 4, which was analyzed by use of the Allen-Hickling relationship:

$$\log \frac{i}{e^{\alpha n F / RT} - 1} = \log i_0 - \frac{\alpha n F}{2.303 RT} \eta \quad (5)$$

where  $i_0$  was the exchange current density from which a direct estimation of relative electrode kinetics were determined,  $\eta$  was the electrode overpotential,  $\alpha$  the transfer coefficient and the other symbols had their usual significance (ref. 16). Exchange current densities ( $i_0$ ) were obtained from the intercept by plotting  $\eta$  versus  $\log \frac{i}{e^{\alpha n F / RT} - 1}$ . The dependency of  $i_0$  upon both cell temperature and lithium alloy composition is summarized in Table 3. These current-voltage measurements clearly showed the presence of rapid electrode kinetics at the negative electrode suggesting that this Faradaic process will not be the rate limiting process in the finally developed electrochemical cell.

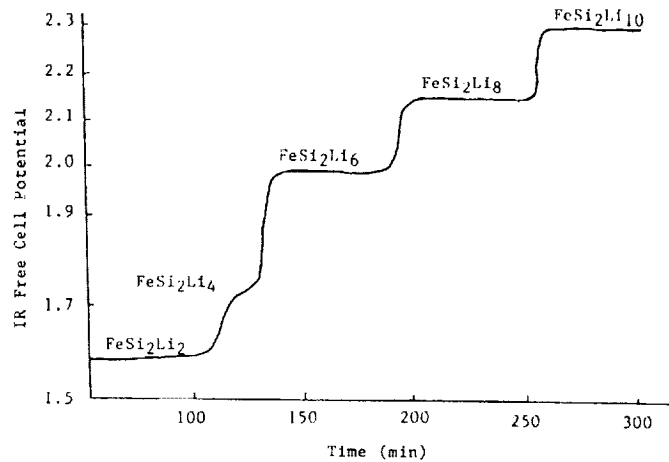


Figure 3. IR free charge curve for  $\text{FeSi}_2$  in 28.5<sup>m/o</sup> LiF, 66.5<sup>m/o</sup> LiCl and 5<sup>m/o</sup>  $\text{Li}_2\text{O}$  at 650°C vs  $\text{La}_{0.89}\text{Sr}_{0.11}\text{MnO}_3/\text{Pt}$  (air). Current density 10mA/cm<sup>2</sup>.

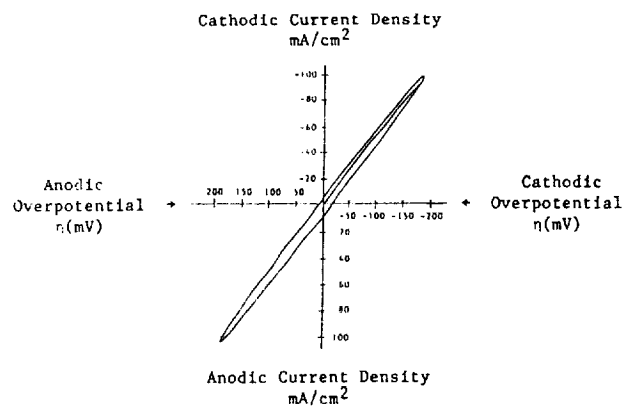


Figure 4. Current-overpotential curve for the cell  $\text{FeSi}_2\text{Li}_{10}/\text{Li}_2\text{O}-\text{LiF}-\text{LiCl}/\text{SiLi}_4$  at 800°C. Initial open circuit potential -71mV with respect to  $\text{SiLi}_4$  reference. Scan rate 50mV/sec.

TABLE 3

Cell Temperature °C	Measured Exchange Current Density $i_0$ (mA/cm <sup>2</sup> )		
	$\text{FeSi}_2\text{Li}_{10}$	$\text{FeSi}_2\text{Li}_8$	$\text{FeSi}_2\text{Li}_6$
650	49.7	50.6	43.7
700	54.5	54.7	50.9
750	56.0	57.2	55.6
800	64.5	65.7	58.2

In fully developed electrolytic cells we must facilitate the convenient removal of faradaically deposited lithium at the negative electrode. One of the major obstacles to achieve this is related to the high surface tension of Li at 700°C (320 dynes/cm) which compares to Hg (400 dynes/cm) at room temperature. Deposited lithium resident in a porous electrode structure, in the absence of wetting, can as a consequence prove difficult to physically remove via a proximate orifice. We have found, however, that both  $\text{SiLi}_x$  and  $\text{FeSi}_2\text{Li}_x$  can act as effective substrate sites for wetting faradaically deposited lithium, thereby providing a strategy for the subsequent continuous removal of lithium from the electrolytic cell. A schematic of the experimental arrangement used for comparing approaches for lithium removal is shown in Figure 5. Lithium could be conveniently removed from the cell by application of  $\approx 5$ psi into the lithium chamber C, thereby enabling its transfer into glass trap D. In the event of lithium/molten salt mixtures being formed in the electrochemical cell, we have found that effective segregation into two distinct phases can be achieved if the stainless steel transfer tube between the electrolytic cell and lithium storage compartment is initially silicided to form a thin coating of  $\text{FeSi}_2$ . Subsequent reaction of this thin coating with lithium to give  $\text{FeSi}_2\text{Li}_x$  will provide the necessary wetting characteristic necessary for promoting preferential lithium migration from the electrolytic cell.

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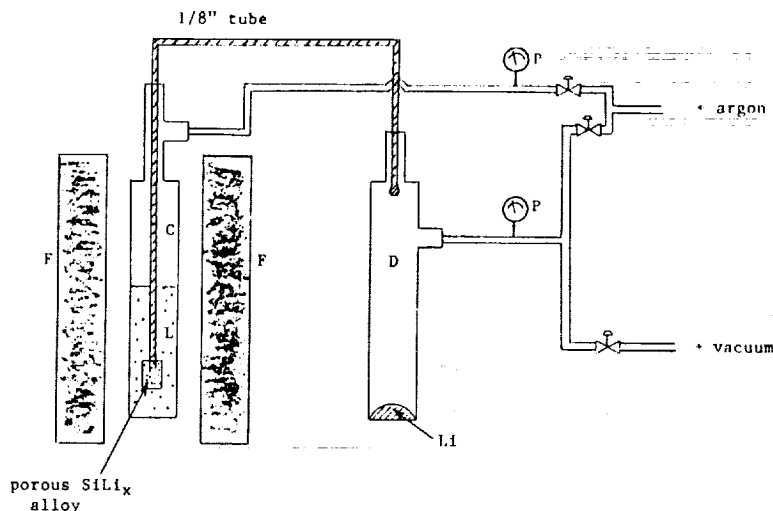


Figure 5. Schematic diagram for simulated lithium removal process.  
C) lithium containing chamber; D) lithium collecting chamber;  
F) furnace; L) lithium; P) pressure gauge

As we have previously discussed, lithium is a strong reducing agent which should in principle, be able to reduce simulated lunar ores to the corresponding metal. This has been examined by investigating direct reaction between lithium and the respective metal oxides  $\text{TiO}_2$ ,  $\text{Fe}_2\text{O}_3$  and  $\text{FeTiO}_3$ . Here, thermal analyses were performed

under both argon and vacuum ( $\sim 100\mu$ ). Lithium and the metal oxide of interest were initially placed into a small boron nitride crucible located inside a 3/4" Swagelock union used as the reaction chamber. A chromel alumel thermocouple was placed in direct contact with the reaction mixture. Because of the anticipated high reducing ability of lithium, only 0.02g was used for each reaction with an excess of metal oxide being present. The reaction chamber was heated at 5°C/min. The inception of solid-state reaction was manifested by a rapid temperature increase in all cases investigated, indicating the reaction to be highly exothermic and irreversible. Thermal analyses for respectively  $\text{Fe}_2\text{O}_3$  and  $\text{TiO}_2$  are shown in Figure 6 A and B. The reaction product between Li and  $\text{Fe}_2\text{O}_3$  was clearly shown to be elemental iron and could be conveniently removed with a magnet. Analysis of the reaction mixture between Li and  $\text{TiO}_2$  was determined by initially dissolving with 37% w/o HCl. Here, unreacted  $\text{TiO}_2$  was insoluble, where Ti became dissolved as  $\text{TiCl}_3$  to give a dark blue solution. Quantitative determination of elemental Ti was then performed spectrophotometrically. Based upon the initial Li present in the reaction mixture, approximately 80% participated in promoting  $\text{TiO}_2$  reduction to Ti. Temperatures experimentally observed for the inception of these reactions are summarized in Table 4 together with the corresponding calculated heat of reaction at each respective temperature.

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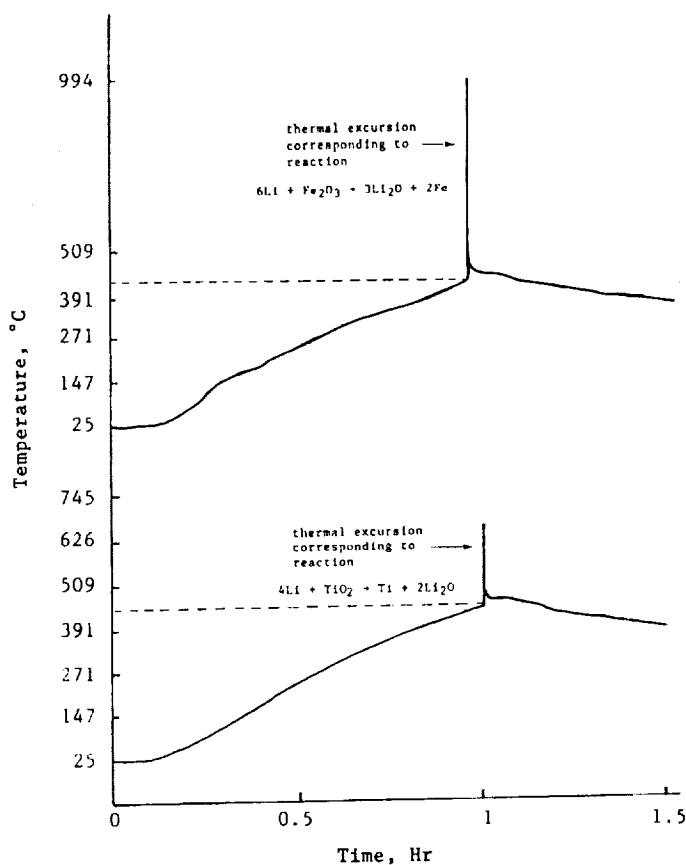


Figure 6. Thermal analysis for detection of solid-state reaction temperature between Li and  $\text{Fe}_2\text{O}_3$  (excess) A, and  $\text{TiO}_2$  (excess) B, under argon.

This electrochemical technology, as we have already discussed, is also compatible for a high energy secondary battery. Initial work (ref. 17) towards this goal focused upon electrochemical cell configurations analogous to that shown in Figure 2

TABLE 4

Experimentally observed temperatures and heat of reaction for the inception of reaction between lithium and the metal oxide of interest.

<u>Metal Oxide</u>	<u>Temperature for Inception of Solid-State Reaction, °C</u>	<u>Heat of Reaction kcal/mole, Calc'd at Temperature</u>
TiO <sub>2</sub>	457	-71.6
Fe <sub>2</sub> O <sub>3</sub>	440	-244.8
FeTiO <sub>3</sub> (ilmenite)	432	--
Fe <sub>2</sub> O <sub>3</sub> +TiO <sub>2</sub> (1:1)	428	-157.2

where maintenance of an inert (Ar) atmosphere in the negative electrode component could be conveniently achieved (ref. 18). A preliminary charge-discharge curve for this type cell using a FeSi<sub>2</sub>Li<sub>x</sub> negative electrode is shown in Figure 7. Again, the voltage plateaus observed correspond to the reversible formation of ferro silicon lithium alloys. Furthermore, it was found during cell charge that the volume of oxygen generated was in fact Faradaic, demonstrating that even in direct contact with halide containing molten salt, the solid electrolyte remains an exclusive oxygen anion conductor.

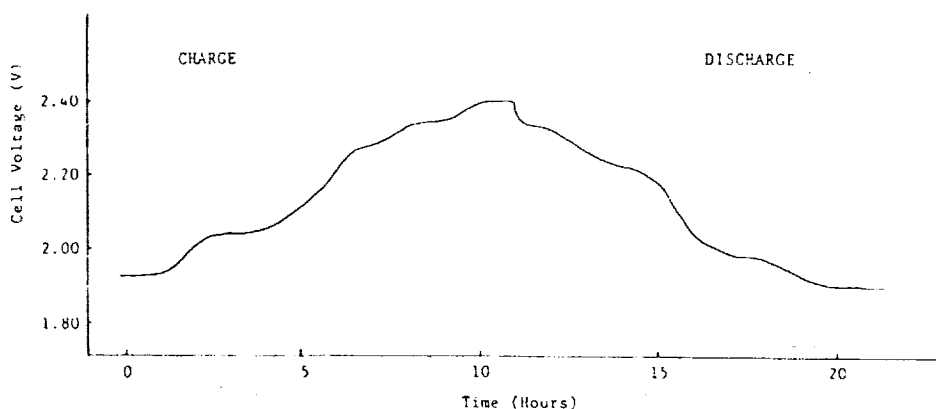


Figure 7. IR-free charge-discharge curve for the cell Li<sub>x</sub>FeSi<sub>2</sub>/(52.5-23.6-23.9)<sup>m/o</sup> LiCl-LiF-Li<sub>2</sub>O/ZrO<sub>2</sub>(5<sup>w/o</sup> CaO)/La<sub>0.89</sub>Sr<sub>0.11</sub>MnO<sub>3</sub>/Pt at 20mA/cm<sup>2</sup> (at negative electrode). Total cell resistance 24Ω. Temperature 650°C.

For practical reversible lithium/oxygen batteries for terrestrial applications, however, we must identify a design which will enable us to operate in the atmosphere. This has been addressed by fabrication of small prototype cells possessing the general design shown in Figure 8. The design is based upon a 5cm high YSZ crucible possessing respective inside and outside diameters of 2.1 and 2.35cm (total volume 19ml). Cells were fabricated in the partially charged state using FeSi<sub>2</sub>Li<sub>8</sub> as the negative electroactive material. Current collection from this electrode was via an aluminized nickel chromium alloy (Ni (76%), Cr (16%), Al (4.5%), Fe (3%), Y (trace), Haynes Alloy 214). The La<sub>0.89</sub>Sr<sub>0.11</sub>MnO<sub>3</sub> oxygen electrode was initially prepared (ref. 19) from a 15<sup>w/o</sup> suspension in polyvinyl alcohol-ethylene glycol of the appropriate stoichiometric metal nitrates. Current collection from this region was via a Pt wire initially tightly coiled in this area. Synthesis of the perovskite electrocatalyst was achieved by heating in the atmosphere to ≈1000°C for 1h. Protection of the negative electrode compartment from direct contact to the atmosphere was achieved using a machined alumina cover (Cotronics, Inc.) whose residual porosity was removed by use of a coating of high density ceramic cement (Sauereisen #8), followed by curing at



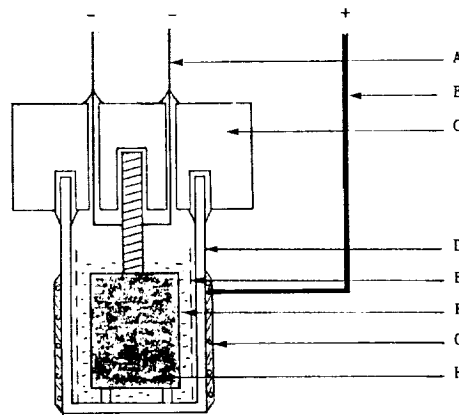


Figure 8. Schematic design of the lithium/oxygen secondary storage cell discussed in the text. A) current collector nickel-chromium (80:20) (-); B) platinum current collector (+) tightly wound around YSZ crucible; C) machined alumina cap; D) YSZ  $O^{2-}$  conducting solid electrolyte; E) zirconia felt separator; F) molten salt containing 66.5<sup>m/o</sup> LiCl, 28.5<sup>m/o</sup> LiF and 5<sup>m/o</sup>  $Li_2O$ ; G)  $La_{0.89}Sr_{0.11}MnO_3$  electrocatalyst; H)  $FeSi_2Li_8$  negative electrode

90°C for one day.

In partially charged cells fabricated here, the negative electrode compartment initially contained a molten salt of composition LiCl (66.5<sup>m/o</sup>), LiF (28.5<sup>m/o</sup>),  $Li_2O$  (5<sup>m/o</sup>). This cell was operated in the atmosphere by initially heating to 600°C, at which temperature the open-circuit potential became 2.2V. Between 650° and 800°C the overall cell resistance decreased from 47Ω to 3.9Ω. This reflected the improved  $O^{2-}$  conductivity for the solid electrolyte at higher temperatures.  $E_a$  was found to be 16kcal/mole, similar to literature values (ref. 20). Current-potential and discharge curves for this cell are shown in Figure 9 (A and B).

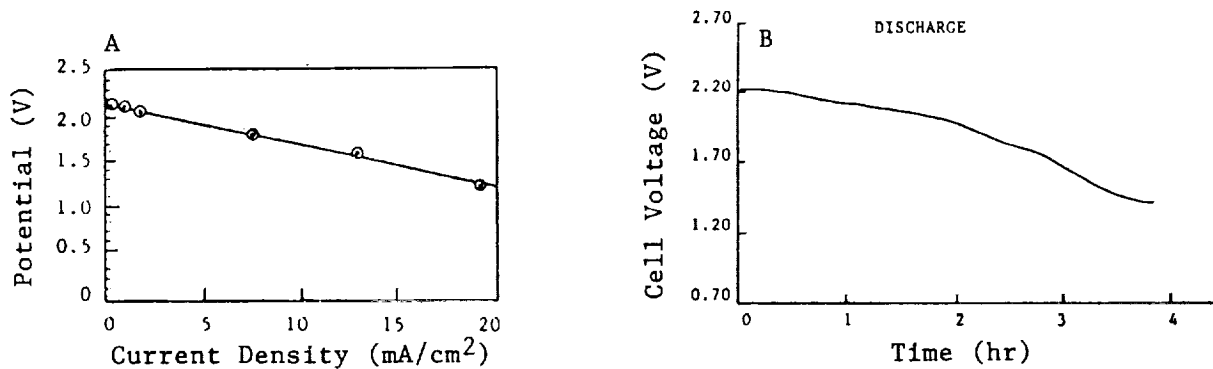


Figure 9. Current-potential and discharge curve for the lithium/oxygen cell at A) 800°C ( $R_{cell} = 3.9\Omega$ ) and B) 850°C at a C/50 rate (5.3mA/cm<sup>2</sup>).

The above discussed results demonstrate that this electrochemical technology is evolving into a practical option for both the generation of chemical species on the

Moon's surface and as a high energy secondary battery for either terrestrial or lunar applications.

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