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# Reduction of silver ions in molybdates: elucidation of framework acidity as the factor controlling charge balance mechanisms in aqueous zinc-ion electrolyte†

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A percolating network of high electrical conductivity needed to operate electrodes at a fast rate can be formed by *in situ* reduction of  $\text{Ag}^+$  originating from mixed metal oxide lattices, but few studies have elucidated trends in this mechanism as a function of  $\text{Ag}^+$  concentration and structure. Candidates compared for the first time here are spinel  $\text{Ag}_2\text{MoO}_4$ , monoclinic and triclinic  $\text{Ag}_2\text{Mo}_2\text{O}_7$ , and  $\text{Ag}_2\text{Mo}_3\text{O}_{10} \cdot 2\text{H}_2\text{O}$ , which have reduction potentials for  $\text{Ag}^+$  and  $\text{Mo}^{6+}$  strongly decoupled by up to  $\sim 600$  mV in aqueous zinc-ion electrolyte. Under these conditions,  $\text{Ag}^0$  is the first reduction product and a decrease of charge transfer resistance by  $\sim 100\times$  is observed within 2.5% consumption of total  $\text{Ag}^+$  independent of initial structure. However, resistance metrics alone poorly describe materials which are robust to reducing silver with high energy at faster rates. Instead, after accounting for crystallinity and morphology differences, we find that the acidity of the molybdate framework is responsible for a switch in charge balance mechanism from the bulk formation of a mixed  $\text{ZnMoO}_x$  to pseudocapacitive  $\text{Zn}^{2+}$  precipitation, and that this mechanism switch is associated with minimized losses to rate, voltage and capacity yields as carbon/binder free electrodes relative to composites. The location of this acidity cutoff near the pH of the  $\text{ZnSO}_4$  electrolyte may suggest a design principle for future low-carbon electrodes beyond molybdate framework structures.

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## Introduction

Electrodes must increasingly support faster electron transport as the devices they power become more demanding. Towards this goal, the development of materials containing  $\text{Ag}^+$  is pursued for fundamental interest because the reduced product can contain a network of atomically dispersed, conductive  $\text{Ag}^0$  particles that support fast electron transport throughout the bulk.<sup>1,2</sup> Minute amounts of dispersed  $\text{Ag}^+$  are known to significantly improve electronic transport; as example, cathodes of the silver–vanadium–phosphate-oxide type show percolation thresholds with 0.3% volume reduction of  $\text{Ag}^+$  concurrently resulting in  $10^6$  decrease in electrical resistance.<sup>3–5</sup> Silver vanadium oxide (SVO) cathodes remain the dominant material in implantable cardiac defibrillator batteries due to this reductive displacement mechanism.<sup>6</sup>

Given that vanadium oxide cathodes have shown tremendous promise in aqueous zinc-ion batteries (AZIB's)<sup>7–9</sup> recent efforts to incorporate reductive displacement cathodes in this new application have included  $\text{Ag}^0$ -doped  $\text{V}_2\text{O}_5$ ,<sup>10</sup>  $\text{Ag}_{0.4}\text{V}_2\text{O}_5$ ,<sup>11</sup>  $\text{Ag}_{1.2}\text{V}_3\text{O}_8$ ,<sup>12</sup>  $\text{Ag}_2\text{V}_4\text{O}_{11}$ ,<sup>13</sup>  $\text{Ag}_{0.33}\text{V}_2\text{O}_5 @ \text{V}_2\text{O}_5 \cdot n\text{H}_2\text{O}$ ,<sup>14</sup>  $\beta\text{-AgVO}_3$ ,<sup>15</sup> and even Cu analogs in  $\text{CuV}_2\text{O}_6$ ,<sup>16,17</sup>  $\text{Cu}_{0.95}\text{V}_2\text{O}_5$ ,<sup>18</sup>  $\text{Cu}_x\text{V}_2\text{O}_5$ ,<sup>19</sup>  $\text{Cu}_{0.34}\text{V}_2\text{O}_5$  (ref. 20) and  $\text{Cu}_3(\text{OH})_2\text{V}_2\text{O}_7$ .<sup>21,22</sup> However, an early emphasis on reporting individual structural metrics – which we summarize in Table S1† – rather than critical comparisons has led to inconsistencies when viewed broadly that, to our knowledge, have not been addressed. Only one study attempts to directly compare any subset of these structures in AZIBs, finding tunneled  $\text{Ag}_{0.33}\text{V}_2\text{O}_5$  outperformed layered  $\text{Ag}_2\text{V}_4\text{O}_{11}$  and chain-like  $\beta\text{-AgVO}_3$  at cycle life, with the latter exhibiting rapid capacity fade to  $<25$  mA h  $\text{g}^{-1}$  at 1000 mA  $\text{g}^{-1}$  rates in only 25 cycles.<sup>12,23</sup> With reports that chain-like  $\beta\text{-AgVO}_3$  is capable of 95 mA h  $\text{g}^{-1}$  after 1000 cycles at 2000 mA  $\text{g}^{-1}$  rates with 78% retention<sup>15</sup> and layered  $\text{Ag}_2\text{V}_4\text{O}_{11}$  is capable of 93% capacity retention after 6000 cycles at 5 A  $\text{g}^{-1}$  rates,<sup>13</sup> it is clear there exist factors as yet unknown which either caused these materials to perform worse in the comparative study or caused them to improve individually. Among many possibilities,<sup>24</sup> sample crystallinity and morphology may rank highly considering that the

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† Electronic supplementary information (ESI) available: 16 figures and discussion on the synthesis of  $\text{Ag}_2\text{Mo}_3\text{O}_{10} \cdot 2\text{H}_2\text{O}$ . See DOI: 10.1039/d1ra07765a


studies reporting improved metrics from  $\beta$ - $\text{AgVO}_3$  and  $\text{Ag}_2\text{V}_4\text{O}_{11}$  prepared nanowires and the comparative test did not.

We argue that, irrespective of whether this conjecture is true, it is not needed if a new set of testing conditions is widely adopted for reductive displacement cathodes. After all, these are materials which, at their best, enable their own conductivity and should require electrodes with little (or no) additives, especially when compared to manufactured  $\text{LiCoO}_2$  and alkaline  $\text{MnO}_2$  cathodes with  $\sim 86$ – $95\%$  active material.<sup>25–27</sup> In spite of this tremendous advantage, all the aforementioned studies prepared electrodes with as much as 20% carbon and 10% binder. The consequence of this disconnect between the conditions in which SVO's should thrive *vs.* the conditions evaluated is that laudable metrics are reported irrespective of whether the SVO structure is tunneled, layered, or chain-like, and this fact is clearly the major contributor to the above inconsistency. For those in pursuit of structure/function relationships this should not be a satisfactory state of literature, especially considering the presence of  $\text{Ag}^+$  does not even guarantee reductive displacement occurs. For instance, lithiation of  $\text{Ag}_x\text{Mn}_8\text{O}_{16}$  ( $1.0 \leq x \leq 1.8$ ) does not form  $\text{Ag}^0$  immediately, but rather Mn accepts the first  $\sim 4$  electron equivalents.<sup>28,29</sup>

Even with a lower reduction potential,  $\text{V}^{5+}$  is still competitive with  $\text{Ag}^+$  for early electrons in  $\text{Ag}_2\text{VO}_2\text{PO}_4$ ,<sup>1</sup> resulting in  $\text{Ag}^0$  networks which may form differently as a function of rate<sup>30</sup> and other factors,<sup>31</sup> including the identity of the intercalating ion.  $\text{Ag}^0$  appears immediately when  $\text{Ag}_x\text{Mn}_8\text{O}_{16}$  is discharged in sodium batteries,<sup>32</sup> and reoxidation of  $\text{Ag}^0$  which is not known in lithium batteries occurs in aqueous Zn systems.<sup>33</sup> Clearly, there is a need to better understand reductive displacement mechanisms in AZIBs since these new conditions may establish novel paradigms. A promising system to do so would contain a metal with a strongly decoupled reduction potential from  $\text{Ag}^+$ .

The  $\text{Ag}_2\text{O}$ – $\text{MoO}_3$  phase diagram contains numerous viable candidates for this including spinel  $\text{Ag}_2\text{MoO}_4$ , layered  $\text{Ag}_6\text{Mo}_{10}\text{O}_{33}$ , orthorhombic  $\text{Ag}_2\text{Mo}_3\text{O}_{10}$ , and two  $\text{Ag}_2\text{Mo}_2\text{O}_7$  polymorphs.<sup>34–36</sup>  $\text{Ag}_6\text{Mo}_{10}\text{O}_{33}$  was reported as a cathode once in 1978 (ref. 37) and more recently has appeared as a topic of study in supercapacitor application.<sup>38</sup> The electrochemistry of monoclinic and triclinic polymorphs of  $\text{Ag}_2\text{Mo}_2\text{O}_7$  has recently appeared in separate studies, where the former is capable of  $825 \text{ mA h g}^{-1}$  at  $100 \text{ mA g}^{-1}$  rates in lithium ion batteries,<sup>39</sup> and the latter is capable of  $190 \text{ mA h g}^{-1}$  at  $20 \text{ mA g}^{-1}$  rates in sodium ion batteries.<sup>40</sup> Collectively, these reports represent the entire body of research on the electrochemistry of silver molybdenum oxides (SMOs) at present knowledge.

Here, we find a subset of four SMO's ( $\text{Ag}_2\text{MoO}_4$ , monoclinic and triclinic  $\text{Ag}_2\text{Mo}_2\text{O}_7$  and  $\text{Ag}_2\text{Mo}_3\text{O}_{10} \cdot 2\text{H}_2\text{O}$ ) which all reduce in absence of additives to form  $\text{Ag}^0$  early in discharge (0.05 electrons) yielding electrodes with lowered resistance by  $\sim 100\times$  and within an order of magnitude of each other. Even with this property, the electrodes show different capabilities to deliver voltage, capacity, and energy as a function of rate; hence, resistance is not a descriptor of materials which would thrive under low-carbon conditions. We solve for the underlying factors behind these differences and propose a descriptor to optimize Ag reductive displacement in other materials. Further,

we show that when evaluated in electrodes with carbon and binder additives all SMOs achieved similar levels of performance, reinforcing the notion that composite electrodes prevent the prediction of structures which would fail in low-carbon electrodes. Our results contribute to understanding where and how  $\text{Ag}^+$  should be predistributed in electrodes targeting this mechanism in low-carbon applications and support future testing of carbon-free electrodes as part of battery material evaluation.

## Experimental

### Syntheses

All chemicals were reagent grade, purchased commercially, and used/stored at ambient conditions without further purification, except for  $\text{CH}_3\text{NH}_3\text{Cl}$  which was stored in an  $\text{N}_2$  glovebox. The water used in all syntheses, electrolytes and solutions was ultrapure  $18 \text{ M}\Omega$  resistivity.

#### $\beta$ - $\text{Ag}_2\text{MoO}_4$ (1Mo)

$\text{AgNO}_3$  (0.680 g, 4 mmol of Aldrich Lot MKCG5845) dissolved in 15 mL of water was added dropwise to a solution of  $\text{Na}_2\text{MoO}_4 \cdot 2\text{H}_2\text{O}$  (0.484 g, 2.4 mmol of Strem Lot 26713700) dissolved in 15 mL of water. The slight excess of  $\text{MoO}_4^{2-}$  was utilized to ensure the system maintained an alkaline pH, under which  $\text{Ag}_2\text{MoO}_4$  is the preferred product. On initial mixture, fine particulates are formed within a cloudy suspension with appearance of milk. Applying  $\sim 50^\circ\text{C}$  heat for 15 minutes, the milk color disappears, and the suspension consists of filterable particulates in clear solution. The phase pure product can be isolated any time after this point. In this study, we performed an overnight reflux, collecting by filtration; 0.50–0.55 g,  $\sim 73\%$ . The material can be dried at  $100^\circ\text{C}$  safely; we dried for 1 h.

#### Monoclinic $\text{Ag}_2\text{Mo}_2\text{O}_7$ (m-2Mo)

Reproduced from Saito *et al.*<sup>41</sup>  $\text{AgNO}_3$  (0.488 g, 2.9 mmol) dissolved in 20 mL of water was added to 0.226 g (1.6 mmol) of  $\text{MoO}_3$  suspended in 20 mL of water. The suspension was refluxed for 5 h, and the resulting solid collected by filtration and washed with water. The material can be dried at  $100^\circ\text{C}$  safely; we dried for 30 minutes. We were unsuccessful at preparing this polymorph by refluxing  $\text{AgNO}_3$  with ammonium heptamolybdate or acidified  $\text{MoO}_4^{2-}$  due to the presence of triclinic  $\text{Ag}_2\text{Mo}_2\text{O}_7$  impurity, which appears in reactions with prolonged temperatures or times; our results agree with other reports.<sup>42,43</sup>

#### Triclinic $\text{Ag}_2\text{Mo}_2\text{O}_7$ (t-2Mo)

Monoclinic  $\text{Ag}_2\text{Mo}_2\text{O}_7$  (200 mg) was heated to  $450^\circ\text{C}$  for 5 h in air in a steel crucible and cooled to room temperature ambiently in 98% yield. The conversion mechanism appears to involve sample melting and subsequent interaction with the container before recrystallizing, as the use of new porcelain crucibles yielded samples with evenly dispersed, 4% Al dopant impurity (which were not studied further). The use of steel eliminates the impurity yet imparts a 2% loss of sample as a tarnish.



**Ag<sub>2</sub>Mo<sub>3</sub>O<sub>10</sub>·2H<sub>2</sub>O via (CH<sub>3</sub>NH<sub>3</sub>)<sub>2</sub>Mo<sub>7</sub>O<sub>22</sub>**

We were not successful at isolating this material by refluxing AgNO<sub>3</sub> with Mo<sup>6+</sup> below pH 4, due to the persistence of kinetically favorable **m-2Mo**. We slightly modified the procedure of Dessapt *et al.*<sup>44</sup> to better avoid the appearance of an adventitious hexagonal MoO<sub>3</sub> impurity (see ESI† for details and discussion). Solutions of 1.01 g (15 mmol) of CH<sub>3</sub>NH<sub>3</sub>Cl in 5 mL of water and 1.21 g of Na<sub>2</sub>MoO<sub>4</sub> (6 mmol) in 5 mL of water were mixed and the resultant solution was syringe filtered. 4 M HCl was added dropwise until pH 1.5 was obtained in close correlation with the appearance of a sustained white suspension. The suspension was heated in a Teflon Parr reactor, 120 °C, 6 h. The product white powder (CH<sub>3</sub>NH<sub>3</sub>)<sub>2</sub>Mo<sub>7</sub>O<sub>22</sub> obtained in yield quantitative in Mo was washed with water and acetone. The reader is referred to ESI† for characterization of (CH<sub>3</sub>NH<sub>3</sub>)<sub>2</sub>Mo<sub>7</sub>O<sub>22</sub> and related discussion. Subsequently, AgNO<sub>3</sub> (0.235 g) was dissolved in 25 mL of 0.1 M HNO<sub>3</sub>, and (CH<sub>3</sub>NH<sub>3</sub>)<sub>2</sub>Mo<sub>7</sub>O<sub>22</sub> (0.300 g) were added. The suspension was refluxed overnight (~16 h) with stirring, and the yellow product was washed with water and ethanol and dried in air.

**Electrochemistry materials and methods**

All tests were recorded on a Biologic VSP potentiostat. Two electrode cells were constructed inside MTI split cells, with Zn foil anode, 2 M ZnSO<sub>4</sub> electrolyte, and Whatman glass fiber filter paper separator. For carbon and binder free working electrodes, silver molybdate powders were pressed between two steel foil

sheets at 20 MPa into pre-massed steel gauze (Type 304, 0.001 in wire). Where applicable and mentioned in the text, composite electrodes were prepared using a slurry of silver molybdate : carbon black : PVDF binder prepared in NMP solvent with composition 8 : 1 : 1 by mass and doctor bladed onto steel foil current collector.  $\frac{1}{2}$ " circles were punched from this coating and used as electrodes.

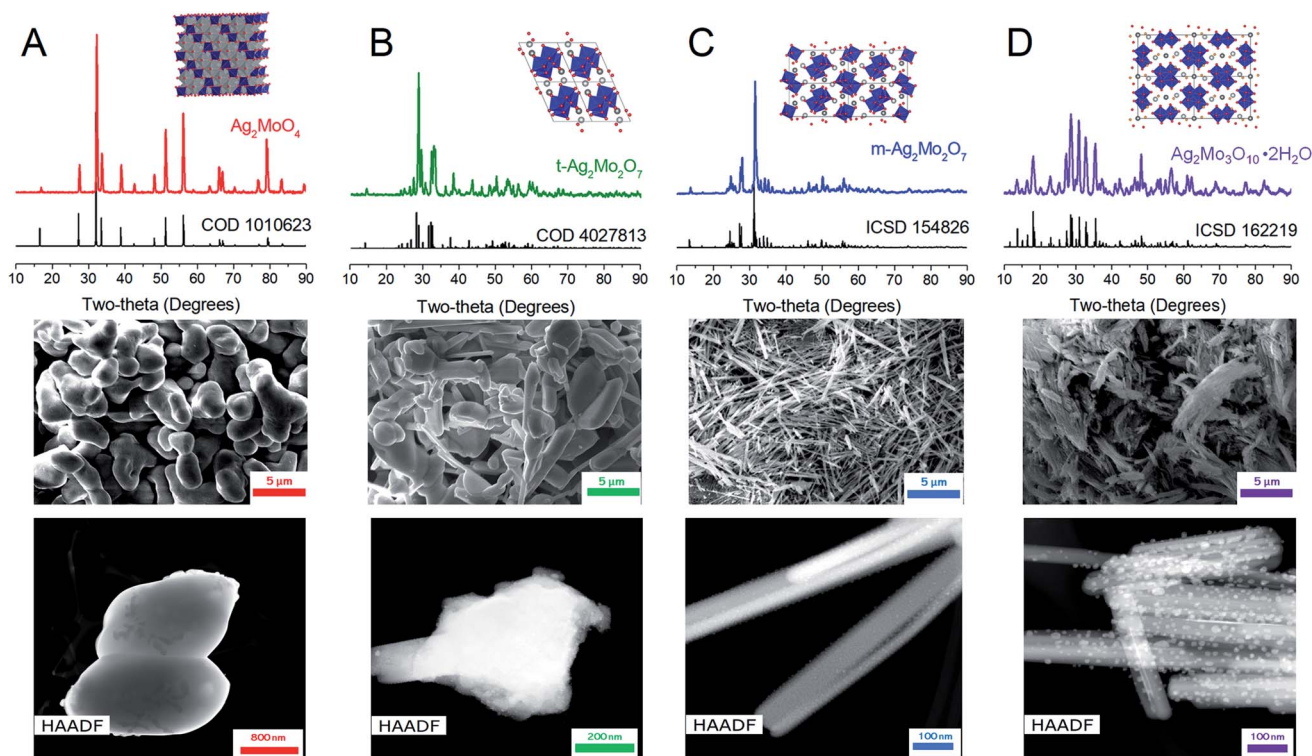
Data for galvanostatic intermittent titration technique (GITT) plots were obtained following 360 second pulses at 1e<sup>-</sup>/2 h rates for each material. To maintain a well-defined surface area, working electrodes in these tests were powders pressed evenly onto precut squares of graphene paper. Ionic diffusion coefficients were calculated by:<sup>45</sup>

$$D = \frac{4}{\pi\tau} \left( \frac{nV_m}{S} \right)^2 \left( \frac{\Delta E_s}{\Delta E_t} \right)^2$$

Here,  $\tau$  = pulse length in seconds,  $n$  is the moles of material per electrode,  $V_m$  is the molar volume for each polymorph (59.7 cm<sup>3</sup> mol<sup>-1</sup> **1Mo**, 89.6 cm<sup>3</sup> mol<sup>-1</sup> **m-2Mo**, 88.4 cm<sup>3</sup> mol<sup>-1</sup> **t-2Mo** and 147 cm<sup>3</sup> mol<sup>-1</sup> **3Mo**),  $S$  is the electrode surface area (in cm<sup>2</sup>),  $\Delta E_s$  is the difference in equilibrated voltage before and after the pulse, and  $\Delta E_t$  is the difference in voltage between the end of the pulse and the equilibrated potential.

**Other materials and methods**

PXRD data were recorded on a Rigaku Miniflex with Cu source. SEM data for Fig. 1 were recorded on a JEOL JSM 6610LV,



**Fig. 1** PXRD, SEM and TEM data for the materials in this work. Experimental powder diffraction patterns (colored traces) are labelled as (A) **1Mo**, (B) **t-2Mo**, (C) **m-2Mo** and (D) **3Mo** over structural references (black traces). In the crystal structure representations blue polyhedra represent Mo and where applicable Ag<sup>+</sup> is represented by two shades of silver, differentiated by symmetry. SEM microscopy images were taken at 5000× magnification.



whereas data for reduced electrodes were taken with a S-Hitachi-4700-II SEM. Raman data were recorded on a Thermo-fisher DXR3xi Raman microscope with 532 nm laser excitation. TEM data were recorded on an FEI Talos TEM with an FEI Tecnai F20ST S/TEM. For characterization of reduced electrodes, samples were washed with 0.1 M acetic acid and water following battery disassembly. The freestanding gauze – and if applicable a steel foil support – was inserted directly into SEM and XRD for respective measurements, whereas the sample was scratched from the mesh onto a lacey carbon grid for TEM analysis.

## Results

### Syntheses and characterization

For this study disilver molybdate ( $\text{Ag}_2\text{MoO}_4$ , **1Mo**) both known polymorphs of disilver dimolybdate ( $\text{Ag}_2\text{Mo}_2\text{O}_7$ , **2Mo**) and disilver trimolybdate dihydrate ( $\text{Ag}_2\text{Mo}_3\text{O}_{10} \cdot 2\text{H}_2\text{O}$ , **3Mo**) were prepared. Literature syntheses were adapted for **1Mo**, **3Mo** and monoclinic **m-2Mo** (see Experimental). For triclinic **t-2Mo**, **m-2Mo** was heated for 5 hours at 450 °C in air; this synthesis improves upon prior reports which have prepared this polymorph *via* solid-state sintering of  $\text{Ag}_2\text{MoO}_4$  or  $\text{Ag}_2\text{O}$  with  $\text{MoO}_3$  requiring 2–4 days.<sup>40,46</sup>

Fig. 1 depicts experimental X-ray diffraction patterns demonstrating phase purity and microscopy characterization of sample morphology. Two of the samples have nanorod morphology (**m-2Mo** and **3Mo**), and the latter exhibits a greater degree of bundling. From the Scherrer equation, the crystallite sizes were 27 and 23 nm respectively. **T-2Mo** and **1Mo** consist of larger particulates; the former with a diverse range from rods to plates, the latter with micron sized, dense potato shapes. These crystallites were 43 and 32 nm, respectively.

High angle annular dark field scanning transmission electron microscopy (HAADF/STEM) images of all samples revealed the existence of surface particles. These were as small as 2 nm in both **2Mo**, 5–15 nm in **3Mo** and, in **1Mo**, the particles were growing as large as 50 nm in real time during the measurement. To our knowledge, all prior studies pursuing transmission electron microscopy characterization on SMOs have reported sample instability and reduction of  $\text{Ag}^+$  to  $\text{Ag}^0$  in the beam.<sup>41,47–51</sup> EDS spectroscopy confirmed that the surface particles were pure Ag; data for the most prominent samples **1Mo** and **3Mo** are provided as Fig. S1.† Given that no  $\text{Ag}^0$  peaks exist in any XRD data, we consider these results consistent with beam damage. For the images in Fig. 1, EDS verified the compositional Ag : Mo ratios expected for each sample (Fig. S2†) and SEM/EDS mapping confirmed the elements were evenly distributed on the micron scale (Fig. S3†). One rationale for the beam damage results could be the  $\text{Ag}^+-\text{Ag}^+$  distances of each SMO, when compared to the product  $\text{Ag}^0$  crystals containing atoms 2.89 Å apart. The **2Mo** samples which exhibited very small  $\text{Ag}^0$  particles do not have  $\text{Ag}^+-\text{Ag}^+$  distances shorter than 3.36 Å. By contrast, **3Mo** contains one of the shortest  $\text{Ag}^+-\text{Ag}^+$  distances of any SMO in this study (3.17 Å), and in **1Mo** all  $\text{Ag}^+-\text{Ag}^+$  distances are 3.27 Å. Hence,  $\text{Ag}^0$  appears to grow more prominently in the TEM beam when the ions do not have to migrate as far.

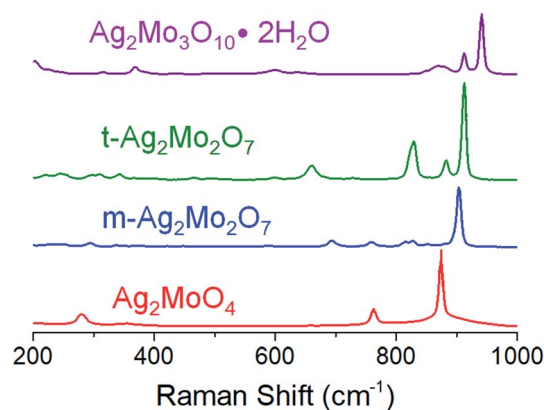


Fig. 2 Raman spectra of the materials in this study; see text for peak values.

Fig. 2 depicts Raman spectroscopy characterization ( $\lambda = 532$  nm). **1Mo** exhibits peaks at 873, 762, 354 and 279  $\text{cm}^{-1}$ , which are in agreement within  $\pm 1$   $\text{cm}^{-1}$  to the known  $A_{1g}$ ,  $2T_{2g}$  and  $E_g$  stretches respectively.<sup>52,53</sup> **m-2Mo** has a major peak at 905  $\text{cm}^{-1}$  with minor bands at 853, 828, 817, 760 and 694  $\text{cm}^{-1}$ , which are in agreement within  $\pm 2$   $\text{cm}^{-1}$  to other data.<sup>43</sup> We are not aware of other Raman spectroscopy data for **t-2Mo**, which our results find a major peak at 913  $\text{cm}^{-1}$  with minor bands at 864, 830, 662, 344 and 311  $\text{cm}^{-1}$ . Finally, **3Mo** gives a major peak at 941  $\text{cm}^{-1}$  with minor bands at 912, 869, 636, 598, 368, 314 and 202  $\text{cm}^{-1}$ , which agrees with literature.<sup>44</sup> Of particular emphasis from this dataset: larger values for the major peak at *ca.* 900  $\text{cm}^{-1}$  indicate stronger Mo–O bonds,<sup>54</sup> and by extension more acidic oxos. Our Raman data therefore confirms the expected trend  $\text{MoO}_4^{2-} < \text{Mo}_2\text{O}_7^{2-} < \text{Mo}_3\text{O}_{10}^{2-}$  for increasing acidity of the molybdate frameworks.<sup>43,55</sup>

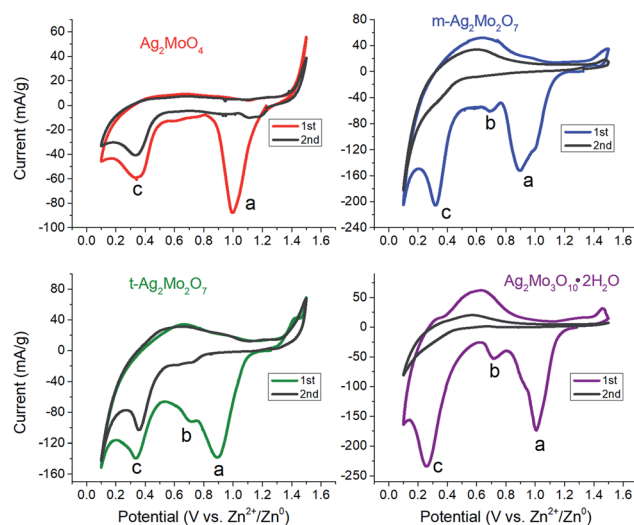


Fig. 3 Cyclic voltammetry at 0.1  $\text{mV s}^{-1}$  for each of the materials in this study. Conditions: 2 M  $\text{ZnSO}_4$  electrolyte, Zn foil anode and reference.



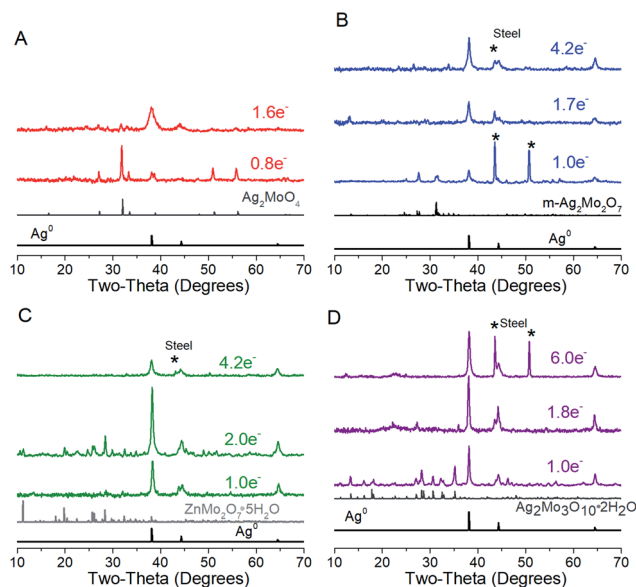


Fig. 4 XRD characterization of electrodes reduced at 5 mA g<sup>-1</sup>; data are also representative up to 40 mA g<sup>-1</sup>. (A) Ag<sub>2</sub>MoO<sub>4</sub>, (B) m-Ag<sub>2</sub>Mo<sub>2</sub>O<sub>7</sub>, (C) t-Ag<sub>2</sub>Mo<sub>2</sub>O<sub>7</sub>, and (D) Ag<sub>2</sub>Mo<sub>3</sub>O<sub>10</sub>·2H<sub>2</sub>O.

### Electrochemistry mechanism: reduction of Ag<sup>+</sup>

Cyclic voltammograms of SMO powders as cathodes in 2-electrode AZIBs are provided as Fig. 3. All SMO's exhibit a cathodic peak 'a' located at 0.90 (2Mo) through 1.00 V (1Mo, 3Mo), which is assigned to the reduction of Ag<sup>+</sup> based on several lines of evidence. (1) It is the most plausible thermodynamic match based on reduction potentials measured of Ag<sup>+</sup> and Mo<sup>6+</sup> standards (Fig. S4†), which occurred across 1.0–1.2 V and 0.20–0.38 V ranges vs. Zn<sup>2+</sup>/Zn<sup>0</sup> respectively. (2) In all SMO's no more than 2 electron equivalents (ee) of charge per mole can be passed at peak 'a'; this precisely matches the amount of charge needed to reduce both formula silvers quantitatively. (3) XRD (Fig. 4) identifies Ag<sup>0</sup> in samples reduced by 0.8–1.0ee at the peak 'a' potential. (4) The Scherrer equation, when applied to the main peak at 38° 2θ, calculates the Ag<sup>0</sup> crystallite size as increasing up to 2ee, beyond which there is no significant difference (Fig. S5†). This is consistent with the Ag<sup>0</sup> network growing in this region to form crystals that ultimately exhibited a correlation of increasing size with decreasing [Ag<sup>+</sup>]<sub>0</sub>, with 6 nm in 1Mo, 17 nm for m-2Mo, 18 nm for t-2Mo and 28 nm for 3Mo. (5) Impedance data (discussed later) shows all SMO electrodes become ~100× more conductive within 0.05ee, consistent with an expected result from forming Ag<sup>0</sup> *in situ*.

The peak current in CV's is known to be the combination of surface capacitive effects – which are proportional to scan rate  $\nu$  – and bulk diffusion effects – which are proportional to  $\nu^{1/2}$ . Hence, for the overall dependence of current on scan rate *via* the equation:<sup>56</sup>

$$i = a\nu^b \quad (1)$$

A plot of log(*i*) vs. log( $\nu$ ) should yield a line with slope *b* between 0.5 (pure diffusion) and 1 (pure capacitance). Applied

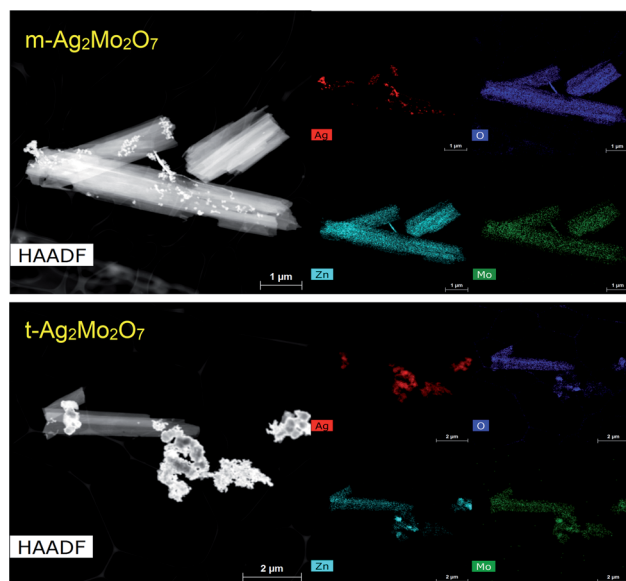
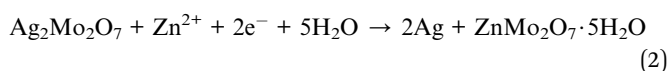


Fig. 5 HAADF/STEM and EDS maps for m-2Mo (top) and t-2Mo (bottom) samples reduced by 2ee at 1e<sup>-</sup>/2 h, showing Zn displacing Ag from the initial structures.

to peak 'a', the *b* values measured for the 2Mo and 1Mo samples (0.55–0.67, Fig. S6†) were all internally consistent with evidence of Ag<sup>+</sup> reduction limited by diffusion in the bulk. However, the *b* value for 3Mo (0.88) instead suggests a more surface-limited, pseudocapacitive mechanism. In support of these electrochemistry assignments, XRD identified ZnMo<sub>2</sub>O<sub>7</sub>·5H<sub>2</sub>O (ICSD 245617) as the other crystalline product after 2ee reduction of t-2Mo (Fig. 4C), thereby proving a full bulk reductive displacement mechanism occurring *via*:



The ZnMo<sub>2</sub>O<sub>7</sub>·5H<sub>2</sub>O is less crystalline than the starting t-2Mo (25 vs. 43 nm, respectively). Attempts to detect a ZnMo<sub>2</sub>O<sub>7</sub> phase from reduction of less-crystalline m-2Mo were unsuccessful but HAADF/STEM characterization (Fig. 5) provides evidence further confirming a closely related mechanism. In these images both 2ee reduced 2Mo samples were found to contain particles analyzing as pure Ag in addition to sheets clearly characterized by EDS as the zinc-molybdenum oxide

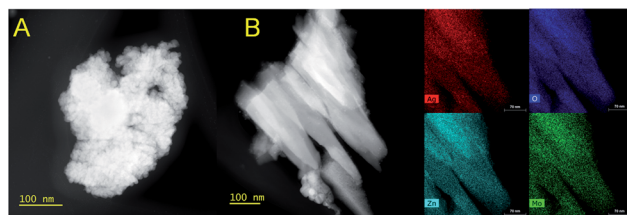
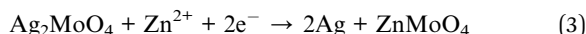


Fig. 6 HAADF/STEM and EDS maps for 1Mo reduced by 1.6ee at 1e<sup>-</sup>/2 h. (A) The morphology of a pure Ag aggregate. (B) Sheets of ZnMoO<sub>x</sub> with residual silver existing as ultrafine particles; the yellow box was analyzed by EDS.



(Fig. 5, S7 and S8†). SEM-EDS mapping on the micron scale (Fig. S9†), further confirms the primarily bulk mechanism for **m-2Mo**.

Reduced **1Mo** electrodes consisted of two types of particles: aggregates of pure nanosized Ag (Fig. 6A), and larger sheets with Ag : Zn : Mo of 1 : 2 : 2 (Fig. 6B). The sheets were covered in ultrafine, ~2 nm particles resembling the Ag in the aggregates (Fig. S10†), though EDS mapping does not assign the total Ag concentration detected to solely those particles. The close spatial association of Zn with Mo, combined with clear evidence of Ag separate from Mo seen also by SEM (Fig. S11†), is also consistent with a bulk reduction displacement mechanism of:



We repeatedly detected  $\text{Ag}^0$  as the only crystalline product from reduction of **3Mo**. Microscopy images of reduced **3Mo** electrodes are shown in Fig. 7 and S12† and were consistent with the absence of a bulk mechanism by electrochemistry. In this data, the major samples detected were sheets with high concentrations of Zn + O, with Mo existing in low concentrations and in distributions not matching Zn. The Zn + O sheets were either in direct contact or closely associated with large domains of Ag, which confirmed the XRD data detecting increased  $\text{Ag}^0$  crystallinity. These results overall do not provide evidence for substitution of Zn into the  $\text{Mo}_3\text{O}_{10}^{2-}$  framework. Rather, the data suggests the existence of a silver-coated basic zinc salt (e.g.,  $\text{Zn}_4\text{SO}_4(\text{OH})_6$ ), which is a known product from Zn/ $\text{MoO}_3$  and numerous other Zn electrochemistries<sup>57</sup> as resulting from a mechanism in which the pH of the electrolyte increases. We had some success detecting a pH increase in  $\text{ZnCl}_2$  electrolyte, which allowed the identification of simonkoellite  $\text{Zn}_5(\text{OH})_8\text{Cl}_2 \cdot \text{H}_2\text{O}$  on the anode (Fig. S13†) by PXRD.

The close association of  $\text{Ag}^0$  with the basic zinc salt observed both by TEM and SEM (Fig. S14†) suggests a **3Mo** reduction mechanism which increases the electrolyte's pH locally near where  $\text{Ag}^0$  forms. Here, we propose:

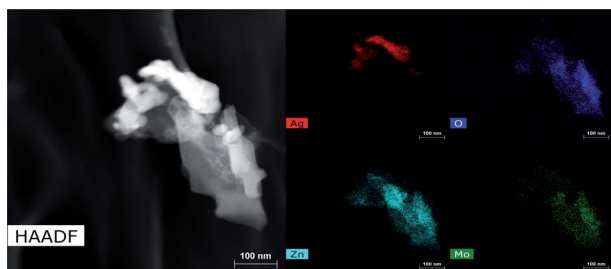


Fig. 7 HAADF/STEM and EDS maps for **3Mo** reduced by 2e<sup>-</sup> at 1e<sup>-</sup>/2 h, showing regions of Zn + O mixed with Ag. The Zn : Mo ratio is 2.3 : 1.

## Electrochemistry mechanism, subsequent steps

We observed two reduction steps: a minor one at 0.7 V and a major one at ~0.3 V (labeled as (b) and (c) respectively) in CV's of both **2Mo** and **3Mo** following the consumption of their  $\text{Ag}^+$  ions (Fig. 3).

Given that aqueous  $\text{Mo}^{6+}$  standards reduce at 0.3 V (Fig. S4†), and that peak 'b' cannot account for quantitative reduction of  $\text{Mo}^{6+}$  in any sample, peak 'c' is assigned to the reduction of  $\text{Mo}^{6+}$ .

The assignment of peak 'b' occurring for a fraction of the  $\text{Mo}^{6+}$  is developed by considering that it is most prominent in **t-2Mo** and absent from the CV of **1Mo**. The only redox-active component in the **t-2Mo** electrode capable of accounting for peak 'b' is  $\text{ZnMo}_2\text{O}_7 \cdot 5\text{H}_2\text{O}$ , which is a layered structure of  $\text{Mo}_2\text{O}_7^{2-}$  chains linked by  $\text{Zn}^{2+}$  cations.<sup>58</sup> Other layered  $\text{MoO}_3$  AZIB cathodes are well-known to exhibit peaks at 0.7 V and ~0.3 V.<sup>57,59</sup> By contrast, the proposed stoichiometry of the phase in the **1Mo** sample is  $\text{ZnMoO}_4$ , for which all known crystalline polymorphs<sup>60,61</sup> are not layered. Therefore, peak 'b' is proposed to indicate the existence of layered domains in the  $\text{Mo}^{6+}$  oxide produced following reduction of each sample's silver ions.

Prior literature has shown the duration of each potential step at 0.7 V and 0.3 V for layered  $\text{Mo}^{6+}$  cathodes is ascribed to the increasing effects of  $\text{H}_3\text{O}^+$  rather than  $\text{Zn}^{2+}$  as the charge carrier.<sup>57,59</sup> We utilized the galvanostatic intermittent titration technique (GITT) to characterize peaks 'b'-'c' since  $\text{Zn}^{2+}$  diffusion into  $\text{MoO}_3$  ( $10^{-10}$  to  $10^{-12}$   $\text{cm}^2 \text{s}^{-1}$ )<sup>59,62</sup> is known to be slower than  $\text{H}^+$  ( $10^{-8}$  to  $10^{-9}$   $\text{cm}^2 \text{s}^{-1}$ ).<sup>59</sup> For all SMO's, we measured ion diffusion coefficients within the range of  $10^{-11}$  to  $10^{-12}$   $\text{cm}^2 \text{s}^{-1}$  while on the  $\text{Ag}^+$  reduction plateau, which is where either Zn diffusion or growth is confirmed by microscopy (Fig. S15†). In direct correlation with the drop to lower potentials for peaks 'b'-'c', the ionic diffusion coefficients increased by ca. 2 orders of magnitude. Therefore,  $\text{Zn}^{2+}$  clearly charge balances the reduction of  $2\text{Ag}^+$  and, once complete,  $\text{H}^+$  is proposed to charge balance the reduction of the resulting  $\text{ZnMoO}_x$  or  $\text{MoO}_x$ .

Under our conditions the SMO's exhibited poor cyclability by CV (Fig. 3) and most other methods tested by our group. Both **2Mo** and **3Mo** show an anodic curve at 0.6 V, assigned to the reoxidation of a fraction of the reduced Mo, which is also observed for aqueous  $\text{Mo}^{6+}$  standards. In all SMO's the anodic peak for  $\text{Ag}^0$  initiates at 1.5 V but to prevent a significant loss of the electronic conductive network we bounded CVs at that potential. Subsequent scans showed no appreciable current for the reduction of  $\text{Ag}^+$ . Our findings are largely in agreement with the consensus that  $\text{MoO}_3$  in aqueous batteries exhibits poor capacity retention upon cycling,<sup>63</sup> with evidence so far suggesting that this is due to the dissolution of reduced Mo which is known from aqueous electrochemistry of Mo oxide passive layers.<sup>64</sup> Numerous reports have indicated generally underwhelming metrics for aqueous  $\text{MoO}_3$  battery or capacitor electrochemistry unless polypyrrole composites, gel electrolytes, or water-in-salt electrolytes are utilized to stop dissolution;<sup>57,63,65-69</sup> these strategies lie outside the scope of this study.





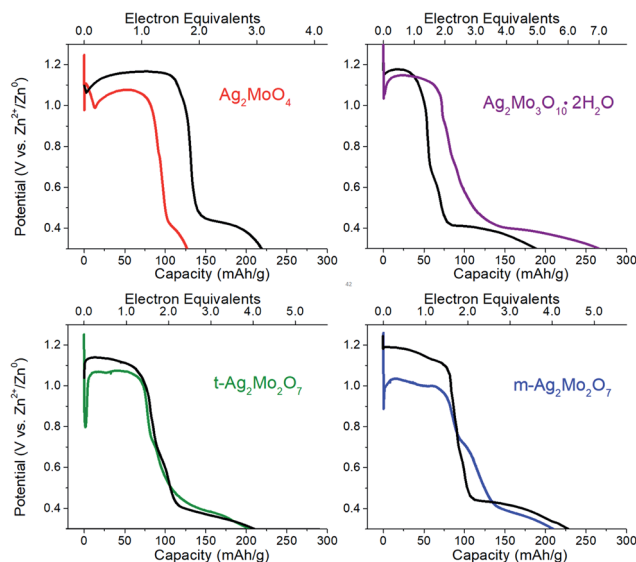


Fig. 8 Representative galvanostatic discharge at 40 mA g<sup>-1</sup> rates for the materials in this study. Working electrodes are the pure powder pressed into steel mesh (colored traces: red **1Mo**, green **t-2Mo**, blue **m-2Mo** and purple **3Mo**) or a composite electrode with 80/10/10 active material/carbon/PVDF composition (black traces). Conditions: 2 M ZnSO<sub>4</sub> electrolyte, Zn foil anode and reference.

### Metrics for Ag<sup>+</sup> reduction

We compared the electrochemical performance of SMO's as stand-alone pellets (Fig. 8, colored traces) vs. composite electrodes containing carbon predistributed (Fig. 8, black traces) at 40 mA g<sup>-1</sup> rates. Notably, the stand-alone pellets all exhibit an

initial voltage drop followed by a recovery – a property well known in SVO cathodes and analogs<sup>31,70,71</sup> as the initial polarization being overcome by the *in situ* generation of a conductivity change provided by the Ag<sup>0</sup> network. We consider the following metrics:

(1) **Operating voltage plateau.** With a carbon network able to lower interparticle resistance, the initial voltage drop/recovery is mitigated in all cases. The resulting reduction of silver is sustained at 1.17 ± 0.02 V, crucially irrespective of the SMO structure. However, without carbon only **3Mo** is capable of attaining >1.1 V for the duration of the plateau; by contrast, **1Mo** and both **2Mo** are at least 50 mV lower.

(2) **Capacity at high voltage.** For all SMO's with carbon the 2e<sup>-</sup> reduction level occurs as the voltage falls from the 1.17 V plateau. While both **2Mo** and **3Mo** reach this boundary without carbon, **1Mo** does not discharge by 2e<sup>-</sup> without carbon.

(3) **Capacity at low voltage.** With Mo<sup>4+</sup> as the lowest oxidation state known in Zn/MoO<sub>3</sub> batteries, plausible upper bound capacities are 4 electrons in **1Mo** (284 mA h g<sup>-1</sup>), 6 electrons in **2Mo** (312 mA h g<sup>-1</sup>) and 8 electrons in **3Mo** (304 mA h g<sup>-1</sup>). Experimentally, we measured **3Mo** (89%, 270 mA h g<sup>-1</sup>) > **m-2Mo** (67%, 210 mA h g<sup>-1</sup>) ≈ **t-2Mo** (65%, 200 mA h g<sup>-1</sup>) ≫ **1Mo** (45%, 130 mA h g<sup>-1</sup>) when ranking the percentage delivered by carbon-free electrodes above 0.3 V. **3Mo** significantly underperforms in the composite electrode suggesting the binder is inhibiting this material's more pseudocapacitive mechanism.

For greater resolution on the formation of the Ag<sup>0</sup> network in stand-alone pellets pulse tests were performed, where 10 × 0.05 electron equivalent intervals were titrated into each SMO; electrodes were relaxed to a stable OCV between pulses. Fig. 9A depicts representative Nyquist plots of SMO electrodes prior to

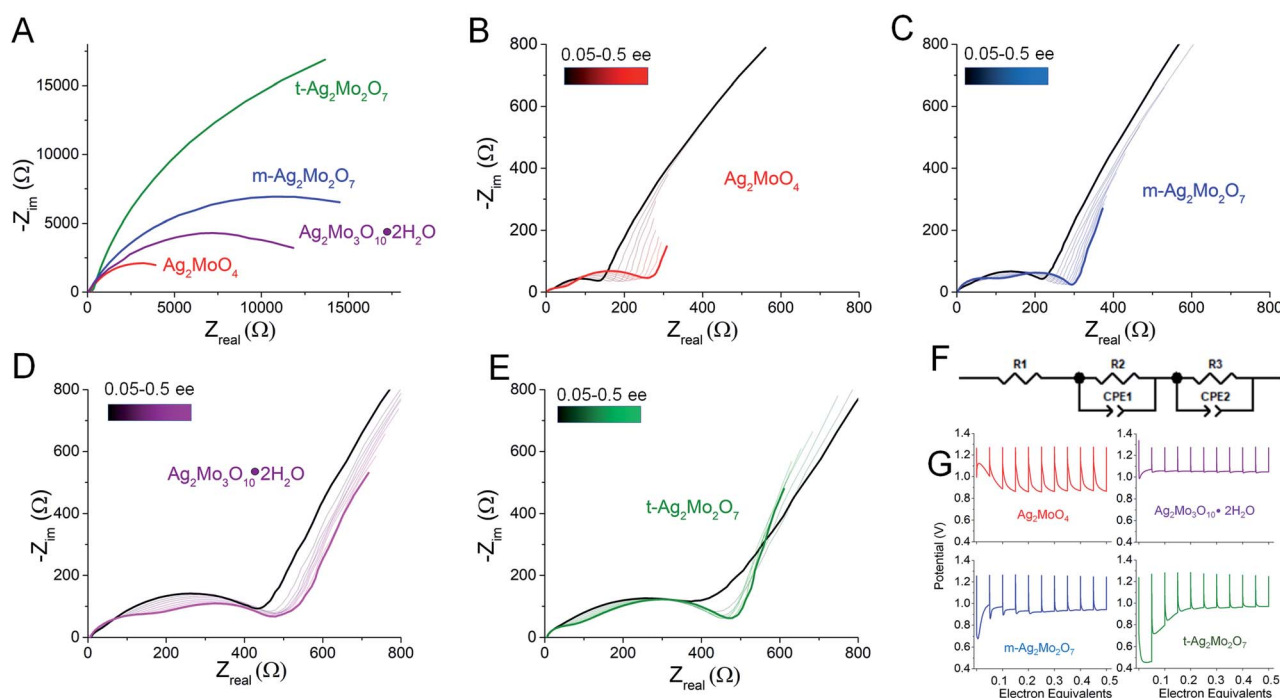


Fig. 9 (A) Representative Nyquist plots of pellet electrodes; (B–E) Nyquist plots following titration of 0.05 electron equivalents into each SMO at 40 mA g<sup>-1</sup>; (F) the circuit used to fit the two semicircles in the impedance data. (G) Potential vs. electron equivalents at 40 mA g<sup>-1</sup> rates.



any reduction. For quantitative fit, the semicircles were modeled *via* a Randles-type circuit with two  $R/C$  parallels, (Fig. 9F), which has been used for other silver metal oxides.<sup>1,3,31,72</sup> In this circuit,  $R_1$  is the solution resistance, the  $R_2/CPE$  parallel models the anode and the  $R_3/CPE$  parallel models the cathode.  $R_3$  is of interest as it represents the interparticle charge transfer resistance,  $R_{ct}$ . Results from fittings quantify the initial  $R_{ct}$  ranging 9–45 k $\Omega$ .

The electrode potentials were monitored during each 0.05 electron pulse (Fig. 9G). Fig. 9B–E depict that, upon relaxation after the first 0.05 electron equivalent reduction, all SMO's exhibit a compressed semicircle indicative of a decreased charge transfer resistance.  $R_{ct}$  is lower than 530  $\Omega$  in all cases, representing a resistance decrease of up to 100 $\times$ , with the largest decrease from the **t-2Mo** electrode.

The first 0.05ee is sufficient to generate enough  $Ag^0$  to recover the voltage in **1Mo**, **3Mo** and **m-2Mo** before the end of pulse duration (Fig. 9G). That **t-2Mo** requires additional  $Ag^0$  is not ascribed to its morphology but rather its initially highest  $R_{ct}$ , which causes a higher polarization to need to overcome. The

data in Fig. 9G clearly show that the initial voltage drop of the electrode is proportional to the initial resistance; we observed similar trends in galvanostatic reduction (Fig. S16†).

In comparison to the 0–0.05ee decrease, subsequent titration of 0.05–0.5ee changed  $R_{ct}$  far less significantly. In particulate samples  $R_{ct}$  gets slightly worse: **1Mo** increased from 141 to 275  $\Omega$  (Fig. 9B), correlating with an electrode which operated at 0.87–0.90 V in each pulse. For **t-2Mo** (Fig. 9E) the increase observed was 390–465  $\Omega$ , but this primarily occurred in early electron increments and ultimately correlated with higher pulse voltages of 0.97 V. **m-2Mo** mostly retained the same  $R_{ct}$  (230–220  $\Omega$ , Fig. 9C) and exhibited a comparable 0.94 V pulse voltage (Fig. 9G). Finally, **3Mo** recovered voltage to the highest plateau at 1.05 V, but this correlated with a relatively high  $R_{ct}$  that was consistently improving with each pulse from 528 to 385  $\Omega$  (Fig. 9D).

Collectively our results show there is no association of low  $R_{ct}$  with operational pulse voltage, which varies for  $Ag^+$  reduction by as much as 140 mV. The clear potential order **1Mo** < **2Mo** < **3Mo** from pulse tests matches that from galvanostatic reduction (point 1). Given that each SMO appears to form a sufficient  $Ag^0$  network to fully recover voltage within 0.2ee, we evaluated the silver networks by two more metrics:

**(4) Rate capability.** We reduced each carbon-free SMO by 0.2 electrons at  $0.5e^- h^{-1}$  rates to form a  $Ag^0$  network. Then, the reduction rate was increased by a factor of 5. Over the next 1.5 electrons – the upper bound for comparison – the resulting electrodes all deliver a fraction of the  $Wh kg^{-1}$  energy relative to when the rate is unmodified. For the sample data shown in Fig. 10, top, this fraction was 84% (**3Mo**) > 73% (**t-2Mo**) > 48% (**m-2Mo**) > 15% (**1Mo**).

**(5) Formation rate kinetics of  $Ag^0$ .** We initiated reduction of each carbon-free SMO at  $3e^- h^{-1}$ ; a  $6\times$  faster rate than the prior test (Fig. 10, bottom). **T-2Mo**, which needs more  $Ag^0$  to overcome an intrinsically higher  $R_{ct}$ , fails this test most prominently. At the other extreme, **3Mo** delivered 200  $mA h g^{-1}$ , which is 86% of the value from the  $6\times$  slower rate.

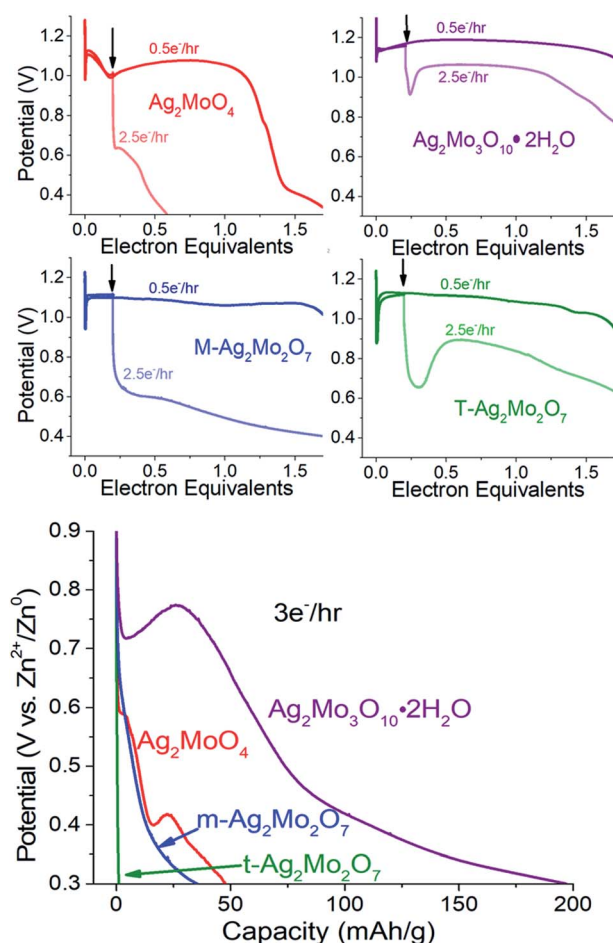


Fig. 10 Top: The SMOs are discharged at  $0.5e^- h^{-1}$  rates until the position of the arrow at 0.2ee. Light curves indicate a rate change to  $2.5e^- h^{-1}$  while the dark curve indicates the remainder of discharge at  $0.5e^- h^{-1}$ . Bottom: Representative galvanostatic discharge at  $3e^- h^{-1}$  for the materials in this study.

## Discussion

In the design of cathodes capable of displacing  $Zn^{2+}$  for  $2Ag^+$  in aqueous conditions, the above data (points 1–5) suggests a trend of **1Mo**  $\ll$  **m-2Mo**  $\leq$  **t-2Mo** for the most optimized initial structure and the “best”  $Ag^0$  network at replicating carbon composites. However, much superior metrics are obtained by **3Mo**, which does not perform  $Zn^{2+}/2Ag^+$  bulk substitution directly nor yields electrodes with lowest  $R_{ct}$ . Such a trend is not addressed by morphology nor crystallinity because both are comparable to the worse-performing **m-2Mo**. Rather, the trend appears easily explained in the context of the acidity of the molybdate framework by Raman spectroscopy (Fig. 2). The consequence of a molybdate framework with greater acidity is that the  $Ag^+$  in **3Mo** have longer  $Ag-O$  bonds 2.44–2.88 Å relative to **m-2Mo** (2.33–2.64 Å), **t-2Mo** (2.27–2.76 Å), and especially **1Mo** (2.31 Å). With the  $Ag^+$  ions most weakly bound to oxygen, the energy penalty for  $Ag-O$  bond breakage is lowered and results in





improved access to  $\text{Ag}^+$  with minimal voltage loss and fast kinetics.

The tradeoff that naturally follows by making a cation easier to displace from a framework is that it becomes harder to put a new cation in. As our measurements show, this tradeoff in molybdates is at first surmountable (because **2Mo** > **1Mo**) but then is avoided entirely through a charge balancing mechanism switch (**3Mo**). Hence, our results appear to demonstrate the existence of a “crossover” acidity below which  $\text{Ag}^+$  can be made more accessible and above which  $\text{Zn}^{2+}$  insertion no longer occurs. In AZIBs this crossover thus occurs between the acidities of  $\text{Mo}_2\text{O}_7^{2-}$  and  $\text{Mo}_3\text{O}_{10}^{2-}$  frameworks. As molecular species, it is intriguing that 10 mM solutions give pH's of 4.58 and 3.52 respectively<sup>55</sup> which appear to straddle the typical ~4.0–4.5 pH of aqueous  $\text{ZnSO}_4$ .

Prior reports on nonaqueous Li/SVO cells do not appear to mention this concept, though there are striking comparisons. Several reports qualitatively correlated materials with higher  $\text{Ag}^+$  content promoting overall higher voltage but with a tradeoff of lower capacity.<sup>24,73,74</sup> This correlation doesn't describe structures, but our Zn/SMO composite electrode results agree with the conceptual trend. Chen *et al.* proposed silver ions in SVO interlayers were more facile to reduce than silver ions as structured octahedra.<sup>75</sup> In our study the silver ions in both **2Mo** and **3Mo** qualify as “interchain”, compared to the structurally rigid  $\text{Ag}^+$  in **1Mo**, but with an observed order of **1Mo** < **2Mo** < **3Mo**, we note this description appears conceptually correct but incomplete. Therefore, our results not only appear consistent with prior trends in nonaqueous electrochemistry but also provide an overall revision.

Our data do not rule out the solvation of  $\text{Ag}^+$  with water molecules unique to **3Mo** as another possible descriptor causing or aiding a change in mechanism. However, a kinetic advantage to weakly-bound  $\text{Ag}^+$  sites has been noted in nonaqueous chemistry recently: purposeful placement of  $\text{Ag}^+$  on the surface of  $\text{Ag}_x\text{Mn}_8\text{O}_{16}$  nanowires improves fast-rate discharge capacity by 2-fold compared to when the  $\text{Ag}^+$  was added to the bulk  $\text{MnO}_x$  tunnel structure.<sup>28,72</sup> We propose, therefore, that this design principle has applications beyond aqueous electrolyte and  $\text{Zn}^{2+}$  ions.

## Conclusions

This investigation was motivated by needs to understand reduction displacement mechanisms under aqueous conditions and to establish the impacts of varying the initial distribution, concentration, and structure of  $\text{Ag}^+$  in mixed metal oxides. We have found across a systematic electrochemical study of four silver molybdates that less than 0.2 electron equivalents (10% of  $\text{Ag}^+$ ) need be reduced to reach locally low values of  $R_{\text{ct}}$  in all cases, and while these values are ultimately comparable in magnitude irrespective of initial structure, they are not predictive of materials which are robust to rate changes nor those which can minimize losses when reducing the carbon content of electrodes. The best descriptor for this is the increase of acidity of molybdate frameworks, which first causes improved displacement of  $\text{Ag}^+$  for  $\text{Zn}^{2+}$  in aqueous electrolyte

before forcing an alternative pseudocapacitive mechanism of charge balance to predominate. This factor contributed more than differences in crystallinity or morphology across the samples.

All these materials and mechanisms deliver comparable voltage for the  $\text{Ag}^+$  reduction when utilized in composite electrodes containing carbon and binder, but at least one case exists (**3Mo**) where additives are in fact detrimental to the mechanism. So, while composite electrodes are necessary to advance the electrochemistry of new  $\text{Zn}^{2+}$  hosts and attain good individual metrics, we have shown here that they potentially can improve the performance of materials with fundamental weaknesses that only research in efforts to increase functional percentages in the next generation of electrochemical energy storage devices will expose.

## Conflicts of interest

There are no conflicts to declare.

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