1. Introduction

Oxygen (O) is a reactive and abundant element of immense chemical, biological, and environmental importance, warranting classification as a field of study in its own right similar to organic (carbon) and inorganic (transitional metal) chemistries. Gaseous dioxygen (O\(_2\)), composed of two covalently bonded homonuclear atoms, constitutes 20.95% of Earth’s atmosphere, providing an oxidising environment and thermodynamic driving force for many biological processes. Crucially, the controlled reaction of O\(_2\) in the oxidation of glucose during aerobic respiration allows life on Earth to thrive. However, living organisms must also possess kinetic barriers to slow oxidation reactions with O\(_2\) to survive. The gradual decay of these kinetic barriers to oxidation in the body is responsible for ageing. In humans, reactive oxide species (ROSs): O\(_2^{*+}\), O\(_2^{2-}\), O\(_2^{*-}\), HO\(_2\), H\(_2\)O\(_2\) and OH\(^*\) are produced as by-products of normal cell processes and are used in the immune response of white blood cells. However, excessive production of ROSs adds oxidative stress to cells causing cell and DNA damage (specifically mitochondrial DNA) which is linked to the formation of cancer cells.

Metal–oxygen (M–O\(_2\)) batteries have to deal with many of the same problems faced by biological systems, such as controlling the production and side reaction rate of ROSs. In addition, dioxide ligands are prevalent in organometallic and catalysis chemistry. Therefore, the chemistry of dioxide species’ (O\(_2^{\pm}\), where \(x = -2, -1, 0, +1\)) has warranted detailed scientific study in multiple fields. O\(_2^{\pm}\) are homonuclear molecules that possess covalent O–O bonds and their characteristic symmetric stretching vibrations (\(\nu_{\text{O–O}}\)) are visible spectroscopically. Examining the \(\nu_{\text{O–O}}\) spectral band can provide valuable information about the chemical nature of the O\(_2^{\pm}\). The oxidation state and reactivity of O\(_2^{\pm}\), depends on the number of electrons occupying the highest occupied molecular orbital (HOMO), which are the \(\pi^*(2p_x)\) valence antibonding orbitals (Fig. 1), their spin state(s) and the O–O bond order (\(B_\Omega\), defined as the number of covalent bonds, see eqn (1)). O\(_2^{+}\) cations possess one electron in the \(\pi^*(2p_x)\) antibonding orbital and have a \(B_\Omega\) of 2.3. O–O bonding electrons in O\(_2^{+}\) cations experience high effective nuclear charge from the O nuclei, so have a short O–O bond length (\(B_\Omega\), 1.12 Å) and a high wavenumber \(\nu_{\text{O–O}}\) (1876 cm\(^{-1}\)) value. From O\(_2^{+}\) to O\(_2^{2-}\),
Johnson et al. in particular highlighted the effect of solvent choice in the determination of the ORR pathway via either surface or solution route. Using SERS they showed that low-donor number solvents, such as acetonitrile, lead to a surface reaction resulting in thin, insulating Li2O3 film growth. In contrast, in high-donor number solvents, such dimethyl sulfoxide resulted in preferential Li2O2 particle growth in solution. High donor number solvents are thereby preferable in the context of Li−O2 batteries as Li2O2 solution growth leads to higher capacities.

Creighton et al. first reported the νO−O stretch of O2*− with Raman spectroscopy of KO2 and contaminated Na2O2 in 1964. This was corroborated and expanded upon with other group-1 alkali-metal cation complexes ([M···O2*−], where M = Li, Na, K, Rb or Cs) in the following decade at a variety of different temperatures. Sawyer et al. synthesised and spectroscopically characterised the first O2*− salt with an organic cation, in 1983, by substituting K+ with tetramethylammonium (TMA)+. [TMA*−···O2*−] had a spectrum analogous to KO2, though slightly red shifted (~22 cm−1). Since then, the νO−O of many O2*− complexes [C*−···O2*−] have been detected in a variety of systems and phases. Such as: solid salts, dissolved in organic solvents or generated electrochemically at a SERS active electrode−electrolyte interface.

Understanding the reactivity and bonding environment of the O2*− intermediate is crucial in the development of more stable M−O2 battery electrolytes and in particular, ambiguity around the existence and chemical character of the lithium superoxide (Li2O2) intermediary remains. Thus, a detailed look at the νO−O spectra of species produced during ORR/OERs is warranted to better understand in the reaction mechanisms in non-aqueous electrolytes in order to assist in the fundamental understanding of lithium−oxygen and in more general other M−O2 batteries.

2. Experimental

2.1. Calculation of ion properties

Ionic volumes, ionic surface area, charge distributions and electrostatic potential surface distributions for a number of molecular ions were determined using a simple density functional theory (DFT, B3LYP functional, 6-31G* basis set) in Spartan 16. Ionic volumes and surface areas for atomic cations were calculated using a simple pace-filling calotte (CPK) model.

3. Results and discussion

3.1. Meta-analysis of superoxide’s Raman spectra

To investigate possible trends of O2*− bonding, relevant literature reports of >50 species were collated and examined. See Table S1 (ESI†) for a catalogue of reported [C*−···O2*−] complexes.

Raman spectral band positions reflect molecular energy levels in a vibrating bond and are well known to be influenced by (1) the mass of the atoms in the bond, (2) the bond force constant and (3) lattice/symmetry effects. For [C*−···O2*−],
counter cations (or strongly coordinating solvents) attract O–O bond electron density at different rates which also influences the force constant of the vibration, resulting in \( \nu_{\text{O-O}} \) band shifts, explained further below.\(^{61} \)

The molecular mass-to-charge ratio \( (M/Q) \), units: mol g\(^{-1}\) C\(^{-1}\) of an ion can dictate its Columbic attractive strength towards counter-ions. Generally, ions with low \( M/Q \) values are stronger Coulombic attractors than ions with high values. In a \([\text{C}^+ \cdots \text{O}_2^\text{−}]\) complex, where interactions between \( \text{O}_2^\text{−} \) and the counter-cation (\( C^+ \)) are weak: electron density is more concentrated on the \( \text{O}_2^\text{−} \) anion, the effective nuclear charge experienced by O-O valence bonding electrons is lower and the bond is longer when \( \text{O}_2^\text{−} \) is strongly coordinated. This effect is visible spectroscopically, with weakly coordinated \( \text{O}_2^\text{−} \) complexes having lower wavenumber \( \nu_{\text{O-O}} \) values (e.g. tetrabutylammonium \( [\text{TBA}^+ \cdots \text{O}_2^\text{−}] \): <1115 cm\(^{-1}\))\(^{15,18,52,66} \) and strongly coordinated \( \text{O}_2^\text{−} \) complexes having higher \( \nu_{\text{O-O}} \) values (e.g. \( [\text{Na}^+ \cdots \text{O}_2^\text{−}] \): \( \geq 1155 \text{ cm}^{-1}\))\(^{15,17,47-52,62} \) The strength of the attraction between the coordinating cation and \( \text{O}_2^\text{−} \) effects valence bond electron density and the O–O bond force constant, causing the observed difference in \( \nu_{\text{O-O}} \) wavenumbers.

For monovalent \([\text{C}^+ \cdots \text{O}_2^\text{−}]\) complexes (independent of phase), plotting reported \( \nu_{\text{O-O}} \) values against the \( M/Q \) value of the coordinating cation (Fig. 2) shows an inverse trend \((R^2 = 0.757)\). In Fig. 2, two further sub-trends are apparent for: (1) single atom cations (shown in blue) and (2) molecular cations (shown in red). Single atom cations such as \( \text{Rb}^+ \) (85.5 mol g\(^{-1}\) C\(^{-1}\)) and \( \text{Cs}^+ \) (132.9 mol g\(^{-1}\) C\(^{-1}\)) have similar \( M/Q \) values as \( \text{TMA}^+ \) (74 mol g\(^{-1}\) C\(^{-1}\)) and tetraethylammonium (TEA\(^+\)) (130.0 mol g\(^{-1}\) C\(^{-1}\)), respectively, but have higher wavenumber \( \nu_{\text{O-O}} \) spectral bands \((> 15 \text{ cm}^{-1})\). This difference in the \( \nu_{\text{O-O}} \) spectral bands can be accounted for by considering the size difference in these cations. Single atom alkali-metal cations are spatially smaller allowing them interact more closely with \( \text{O}_2^\text{−} \) valence bond electron density than large molecular organic cations. Therefore, as a general rule: strong \( \text{O}_2^\text{−} \) coordination has a higher wavenumber \( \nu_{\text{O-O}} \) spectral band.

### 3.2. Ionic charge dispersion

The immediate coordinating environments ability to interact with \( \text{O}_2^\text{−} \) also provides interesting information about the reactivity and basicity of \( \text{O}_2^\text{−} \) in the complex which is reflected in the \( \nu_{\text{O-O}} \) value. From Fig. 2, \([\text{C}^+ \cdots \text{O}_2^\text{−}]\) containing high \( M/Q \) cations (e.g. \( \text{TBA}^+ \)) have low \( \nu_{\text{O-O}} \) values indicative of a more ionic ‘freer’ \( \text{O}_2^\text{−} \) species. In these complexes, \( \text{O}_2^\text{−} \) is a stronger Lewis base and can be expected to more readily donate electron density. Inversely, low \( M/Q \) cation containing \([\text{C}^+ \cdots \text{O}_2^\text{−}]\) complexes such as \((\text{H}^+)\) have stronger \( \text{C}^+\)-to-\( \text{O}_2^\text{−} \) interactions with slight covalent character, resulting in higher wavenumber \( \nu_{\text{O-O}} \) bands indicative of a weaker Lewis base \( \text{O}_2^\text{−} \) species. It was found that for cations coordinating \( \text{O}_2^\text{−} \) species, the: size, mass and charge influence the strength of coordination and in turn the O–O bond force constant which causes detectable changes in \( \nu_{\text{O-O}} \) value. Thus, the physical parameters of one ion can indirectly influence the bond force constant and wavenumber of bonds present in the counter-ion. These values for ions can be simply quantified in terms of the \( M_i \) ionic volume \((i)\) and the charge on the ion \((q)\). In \( \text{O}_2^\text{−} \) complexes; changes in the \( M_i \) and \( i \) of the coordinating counter-ions were found to be inversely proportional to changes in \( \nu_{\text{O-O}} \) whilst changes in the charge of the counter-ion were proportional quantified by the term. Considering the effect the size of the cation has on the \( \nu_{\text{O-O}} \) of \( \text{O}_2^\text{−} \) complexes, this can be quantified by either the ionic volume \((i)\) or the available ionic surface area \((i_{\text{a.s.}})\) measured in Å\(^2\) and Å\(^3\), respectively. \( i \) and \( i_{\text{a.s.}} \) values are known for most alkali-metal cations and can be calculated computationally using simple CPK or DFT models for larger molecular cations.

The \( \nu_{\text{O-O}} \) of a \([\text{C}^+ \cdots \text{O}_2^\text{−}]\) complex changes proportionally with the charge \((q)\) on the cation and inversely proportionally with the molecular mass \((M_i)\) and \( i \) (and \( i_{\text{a.s.}} \)) of the coordinating cation (eqn (2)). Assuming a uniform charge distribution over the ion: multiplying the \( M_i Q^{-1} \) and a spatial component \((i \text{ or } i_{\text{a.s.}})\) gives a parameter (eqn (3)), hence the named the ‘ionic charge dispersion’ \((J)_i\).

\[
\nu_{\text{O-O}} \propto \frac{Q^C}{M_i^C \cdot \frac{1}{j}} \quad \text{or} \quad \frac{1}{i_{\text{a.s.}}}
\]

(2)

\[
J_i^C = \frac{(i^C M_i^C)}{Q^C}
\]

(3)

\[
\nu_{\text{O-O}} \propto \frac{1}{J_i^C}
\]

(4)
where: $\nu_{O-O} = O_2^{-*}$ stretch (cm$^{-1}$), $M_c^i$ = molecular mass of coordinating cation, $Q^i_c$ = charge on cation, $\ell_c^i$ = ionic volume of cation (Å$^3$).

The $\mathcal{K}$ value can be used to estimate the Coulombic attractive strength of an ion and the effect it will have on counter-ions. The Cyrillic symbol for $Z\bar{\mathcal{K}}$ is used for ionic charge dispersion, subscripts $i_v$ and $i_{s.a.}$ denote the use of ionic volume or ionic surface area, respectively, and superscripts $A$ and $C$ denote anion and cation, respectively. $\mathcal{K}$ is an analogue of the charge density of an ion ($M_iQ^{-1}i_v$). However, charge density does not account for the proportional relationship between: $M$, $i_v$ and the Columbic attractive strength of the ion. Thus, $\mathcal{K}$ has been derived as a simple value to describe these ionic properties. $\mathcal{K}$ is a measure the ‘bulkiness’ of the counter-ion, where, the higher the $\mathcal{K}$ value the bulkier the ion is.

Excluding outliers, plotting the reported $\nu_{O-O}$ values of $[[C\cdots O_2^*]]$ complexes from Table 1 against the calculated $\mathcal{K}^i_{C}$ (units: mol g$^{-1}$ Å$^3$ C$^{-1}$) or $\mathcal{K}^i_{s.a.}$ (units: mol g$^{-1}$ Å$^3$ C$^{-1}$) of each coordinating cation shows a clear inverse exponential relationship (Fig. 3 and Fig. S1, respectively, ESI†). $\mathcal{K}$ provides a much better fit than using the $M_c^i$ or $\ell_c^i$ of the coordinating cation alone. The spatial components $i_v^C$ or $i_{s.a.}^C$ are used to calculate $\mathcal{K}$ but it appears to be unimportant which is used as they both change proportionally with one another between different cations. However, $\ell_c^i$ gives a slightly better fit with a higher $R^2$ value of 0.992 (given the larger number of independent reports using different phases, systems and detection equipment spanning over >50 years this is a fairly good fit) compared with 0.942 using $i_{s.a.}^C$ (Fig. S1, ESI†). Plotting on a logarithmic scale shows this correlation more clearly (Fig. 3). All reports of LiO$_2$ except for one$^{36}$ were excluded, this is discussed in Section 3.6.

When detecting the $\nu_{O-O}$ of $O_2^{*+}$, the derived equation for the line of best-fit (eqn (5)) from Fig. 3 can be used to help estimate its ionic character and the coordination strength of the environment in terms of $\mathcal{K}^i_{C}$. With this knowledge, $O_2^{*+}$ can be used as a diagnostic molecule to probe ion and even electrolyte interactions spectroscopically in environments where the coordination strength is unknown (e.g. novel electrolytes).

This relationship between $\nu_{O-O}$ and $\mathcal{K}^i_{C}$ in Fig. 3 was found to hold in most cases. However, it can be manipulated by changing the symmetry, steric hindrance and charge of the coordinating cation, as well as the solvents Gutmann acceptor/donor numbers$^{63}$ and the potential at an electrode surface, where $O_2^{*+}$ is generated electrochemically (to be discussed elsewhere).

$$\nu_{O-O} = -7.327\log(\mathcal{K}^i_{C}) + 1192.9 \quad (5)$$

Considering these findings, hydrogen superoxide (HO$_2$) has a high $\nu_{O-O}$ value (>1165 cm$^{-1}$)$^{26,29}$ and $O_2^{*+}$ can be considered

\begin{table}[h]
\centering
\caption{Table of calculated $i_v^C$, $i_{s.a.}^C$, $\mathcal{K}^i_{C}$ and $\mathcal{K}^i_{s.a.}$ values for some common cations and IL cations. Cations with smaller $\mathcal{K}$ values are expected to be stronger Lewis acids.}
\begin{tabular}{|c|c|c|c|c|c|c|}
\hline
Cation & $M_i$ (mol g$^{-1}$) & $i_v^C$ (Å$^3$) & $i_{s.a.}^C$ (Å$^3$) & $Q$ (C) & $\mathcal{K}^i_{C}$ (g mol$^{-1}$ Å$^3$ C$^{-1}$) & $\mathcal{K}^i_{s.a.}$ (g mol$^{-1}$ Å$^3$ C$^{-1}$) \\
\hline
H$^+$ & 1.0 & 7.2 & 18.1 & 1 & 7 & 18 \\
Li$^+$ & 6.9 & 5.6 & 10.2 & 1 & 39 & 71 \\
Na$^+$ & 23.0 & 10.1 & 19.3 & 1 & 233 & 443 \\
Ca$^{2+}$ & 40.1 & 15.0 & 29.4 & 2 & 301 & 589 \\
K$^+$ & 39.1 & 21.1 & 37.0 & 1 & 827 & 1446 \\
Bi$^{3+}$ & 209.0 & 31.85 & 48.60 & 3 & 2219 & 3386 \\
RB$^+$ & 85.5 & 26.6 & 43.1 & 1 & 2272 & 3681 \\
Cs$^+$ & 132.9 & 35.2 & 51.9 & 1 & 4678 & 6903 \\
TMA$^+$ & 74.1 & 105.2 & 128.4 & 1 & 7799 & 9523 \\
TMP$^+$ & 91.7 & 115.2 & 141.8 & 1 & 10497 & 12923 \\
TES$^+$ & 119.3 & 148.5 & 109.8 & 1 & 17704 & 13093 \\
TEA$^+$ & 130.3 & 195.4 & 177.5 & 1 & 23451 & 23115 \\
Pyrr$_{14}$ & 142.3 & 183.2 & 129.4 & 1 & 26600 & 18408 \\
TPA$^+$ & 180.4 & 250.4 & 273.8 & 1 & 46662 & 51030 \\
Di-I$^+$ & 296.4 & 358.0 & 383.4 & 2 & 53062 & 50154 \\
TBA$^+$ & 242.5 & 324.0 & 355.0 & 1 & 78566 & 86065 \\
N$_{1888}$ & 368.7 & 489.0 & 540.1 & 1 & 180316 & 199128 \\
\hline
\end{tabular}
\end{table}
to have some covalent character, sharing part of its valence bonding electron with the cation. Whereas, tetrabutylammonium superoxide (TBAO₂) has a low ν₀₋₋ value and can be considered to contain dissociated ions. Fig. 4 depicts this relationship between the cation, O₂⁻⁻ and ν₀₋₋ schematically.

3.3. Ionic charge dispersion of other ions

The $\mathcal{K}$ value estimates the dispersion of a formal charge on an ion by assuming the charge is delocalised (i.e. dispersed uniformly across the volume and molecular mass of the ion). $\mathcal{K}$ is a simple value that can be calculated for any ion to approximate its coordination strength and/or propensity to interact with oppositely charged ions. Small $\mathcal{K}$ values indicate ions that are strong: coordinators or Lewis acid/bases, whilst large values represent weak: coordinators or Lewis acid/bases. This can be further explained in terms of hard–soft acid–base (HSAB) theory. Hard acids interact better with hard bases and soft acids with soft bases. Ions with low $\mathcal{K}$ values can be considered hard Lewis acid/bases and ions with high $\mathcal{K}$ values can be considered soft Lewis acid/bases.

Tables 1 and 2 list calculated $\mathcal{K}$ values for some common cations and anions, respectively. The $\mathcal{K}$ value is a clear simplification of the charge on an ion, excluding many important contributions to inter-ion interactions such as: steric hindrance, charge density, the component molecules oxidation states and symmetry etc. However, as discussed, $\mathcal{K}$ values give a good approximation for the influence of the cation on the ν₀₋₋ of O₂⁻⁻ and are likely applicable to other systems too. Furthermore, additional quantifiable values for other ion properties (such as: symmetry etc.) could be added to refine the $\mathcal{K}$ parameter further in future. When selecting ions for their attractive strength $\mathcal{K}$ parameters (and possibly $\mathcal{K}^\alpha$: $\mathcal{K}^\beta$ ratios) could be used to quickly screen some of the near infinite number of ions and combinations that can be selected. Also, $\mathcal{K}$ parameters give a quick and cheap starting point when designing and tailoring novel IL electrolytes or salts. This is a good starting point compared with more precise computational or experimental techniques which are expensive, time consuming and would necessarily come later after first narrowing the field of search.

Based on the $\mathcal{K}_c^\alpha$ value, the calculated Coulombic attractive strength of cations on anions is suggested to be similar for: Na⁺ and Ca²⁺ (233 and 301 g mol⁻¹ Å⁻¹ C⁻¹), Rb⁺ and Bi³⁺ (2272 and 2219 g mol⁻¹ Å⁻¹ C⁻¹) and TEA⁺ and Pyr14⁺ (25 451 and 26 060 g mol⁻¹ Å⁻¹ C⁻¹, respectively).

Similarly, for the $\mathcal{K}_a^\gamma$ of anions, the Coulombic attractive strength of anions on cations is suggested to be similar for: F⁻ and OH⁻ (−276 and −296 g mol⁻¹ Å⁻¹, respectively), Cl⁻ and O₂⁻⁻ (−840 and 738 g mol⁻¹ Å⁻¹, respectively), whilst I⁻ is comparable with dicyanamide (DiCN⁻⁻) and ClO₄⁻⁻ (−4414, −4295 and −5384 g mol⁻¹ Å⁻¹, respectively).

3.4. Dioxygen spectra and important bond parameters

It is of great practical value to relate physical bond parameters with spectral measurements so they can be derived with ease experimentally. To that end, several efforts have been made, however, these methods remain empirical. Key O₂⁻⁻ bond
parameters such as the bond order ($B_0$), bond dissociation enthalpy ($B_1$), bond force constant ($B_2$) and bond length ($B_3$) require various experimental procedures to derive, like X-ray crystallography (XRC). Bond parameter values for most $O_2^x$ species exist and can be correlated well with empirically measured $\nu_{O-O}$ values. Table 3 lists these bond parameters for different $O_2^x$ species. Correlating these O–O bond parameters allows for variations in each bond parameter to be estimated by only measuring the $\nu_{O-O}$ of the $O_2^x$ species.

Considering the relationship between $\nu_{O-O}$ and the coordinating cation discussed in the previous section: $[\rm C^+\cdot O_2^*]$ complexes with strong coordinating cations can be considered to have partial-covalent character, i.e. the coordinating cation abstracts some amount of the $O_2^*$ valence bond electron density. This bond electron density abstraction interaction between the dioxide ion and its coordinator is here on referred to as secondary covalency. For different $O_2^x$ species, the amount of electron density abstracted by the cation can be estimated by considering the $B_0$ of other $O_2^*$ species. A change of 0.5 in the $B_0$ corresponds to an electron being added or removed from the O–O bond. By assuming the $B_0$ is variable, then a change of 0.25 indicates a valence bond electron being shared with the coordination environment, discussed later. A fully dissociated $O_2^*$ species has a $B_0$ of 1 and there is an approximately linear relationship between the $B_0$ and the bond force constant. As the bond force constant dictates the wavenumber of the $\nu_{O-O}$, therefore, calculating the ratio between the $\nu_{O-O}$ of various $O_2^*$ species and the $\nu_{O-O}$ of $O_2^2-$ ($v_{O-O}^{\peroxide}$) generates values that match well with the actual O–O $B_0$ (Table 3) as observed by Livneh et al.9

Due to their high $\Delta H^\circ$ values, $\text{Cs}^+$ (4678 g mol$^{-1}$ Å$^{2}$ C$^{-1}$) and $\text{TBA}^+$ (78656 g mol$^{-1}$ Å$^{2}$ C$^{-1}$) are weak coordinating cations. Therefore, the $B_0$ of $\text{Cs}_2\text{O}_2$ and $\text{TBA}^+\cdot O_2^*$ can be assumed to be 1 and 1.5 and their Raman spectra used as $\nu_{O-O}$ values for dissociated $O_2^2-$ and $O_2^*$, respectively. In reality, both oxides will have some slight covalent character, however, even without considering this, and also the contributions from anharmonicity in the O–O bond vibration (which are relatively small and decrease in magnitude from $O_2^2-$ to $O_2^*$), we see good agreement between the $B_0$ and $\nu_{O-O}^{\peroxide}$ ratios in Table 3. Fitting bond parameters in Fig. 5 with respect to known $B_0$ values allows simple fits (eqn (6) and Table S2, ESI) for estimating $O_2^x$ bond parameters to be derived. The trends are approximately linear, however, more accurate cubic fits where used. Therefore, provided the $\nu_{O-O}$ of a $O_2^x$ is known, it is suggested that all other bond parameters can be estimated.

\[
y = Ax^3 + Bx^2 + Cx + D
\]  

(6)

where $B_0$ = bond order, $n$ = number of bonding electrons, $n^*$ = number of antibonding electrons.

To further qualify that the relationship between the $\nu_{O-O}$ of $O_2^x$ and its bond parameters are approximately linear/cubic: it would be expected that a $O_2^x$ species with secondary covalency (i.e. a $O_2^x$ species with the outer valence electron partially shared with its coordinating environment) would have a $B_0$ between that of two fully dissociated $O_2^*$ species. For example, a $O_2^x$ species with a $B_0$ between that of $O_2^2-$ ($B_0 = 1$) and $O_2^*$ ($B_0 = 1.5$) would be expected to have a $B_0$ of ~1.25, and would have a $\pi^*(2p)$ orbital structure between that of a $O_2^2-$, $\left[\right]$, and $O_2^*$, $\left[\right]$, species. Thus, such a $O_2^x$ species would have a

![Fig. 5](image-url)  

**Fig. 5** Fitted plots of $O_2^x$ bond parameters from Table 3 against the $B_0$. Colour-coordinated axis labels are shown above the graph. Cubic lines of best fits are shown by dashed lines. Fit values shown in Table 4.

<table>
<thead>
<tr>
<th>O-O bond parameters</th>
<th>$O_2^{2-}$</th>
<th>$O_2^*$</th>
<th>$O_2$</th>
<th>$O_2^+$</th>
<th>Ref.</th>
</tr>
</thead>
<tbody>
<tr>
<td>$n^*(2p)$, MO outer valence structure</td>
<td>1</td>
<td>1</td>
<td>1</td>
<td>1</td>
<td>—</td>
</tr>
<tr>
<td>Bond length $B_3$ (Å)</td>
<td>1.49</td>
<td>1.34</td>
<td>1.24</td>
<td>1.12</td>
<td>68</td>
</tr>
<tr>
<td>Bond enthalpy $B_1$ (kJ mol$^{-1}$)</td>
<td>149</td>
<td>360</td>
<td>498</td>
<td>644</td>
<td>68</td>
</tr>
<tr>
<td>Bond force constant $B_2$ (mdyn Å$^{-1}$)</td>
<td>2.56</td>
<td>6.18</td>
<td>11.4</td>
<td>16.3</td>
<td>8, 13 and 69</td>
</tr>
<tr>
<td>$\nu_{O-O}$</td>
<td>743$^a$</td>
<td>1108$^b$</td>
<td>1556$^c$</td>
<td>1876$^c$</td>
<td>9, 15, 53, 58 and 70–72</td>
</tr>
<tr>
<td>Bond order $B_0$</td>
<td>1</td>
<td>1.5</td>
<td>2</td>
<td>2.5</td>
<td>—</td>
</tr>
<tr>
<td>$\nu_{O-O}^{\peroxide}$</td>
<td>1</td>
<td>1.49</td>
<td>2.10</td>
<td>2.53</td>
<td>—</td>
</tr>
</tbody>
</table>

Where: $a = \nu_{O-O}$ of (weakly coordinated) $\text{Cs}_2\text{O}_2$ salt, $b = \nu_{O-O}$ of $\text{TBAO}_2$ generated in dimethyl sulfoxide (DMSO), $c = \nu_{O-O}$ of gaseous dioxides.
detectable bond vibration and $\nu_{O-O}$ band somewhere between that of a strongly coordinated $O_2^{x-}$ species (e.g. $H_2O_2$: 875 cm$^{-1}$) and a fully-dissociated $O_2^{*+}$ (e.g. a TBA$^{+-}$$cdot$$O_2^{*+}$: $\sim$ 1108 cm$^{-1}$) species. Therefore, $\nu_{O-O}$ values between those of these two species can be considered to have secondary covalent character, with valence bond electron density being shared by the $O_2^{x-}$ anion with its coordinating environment. Such a $O_2^{x-}$ species would be expected to have a partial dipole moment induced by the coordination interaction and the $\nu_{O-O}$ band would be both Raman and IR active. Searching the literature this was indeed found to be the case for the $\nu_{O-O}$ of organic peroxy,$^{24,75}$ transition metal superoxo-complexes,$^{6,76,77}$ and metalloenzymes,$^{78}$ which all had Raman and/or IR values reported in the range between 875–1108 cm$^{-1}$. Also, this trend in $\nu_{O-O}$ would be expected to be the same for $O_2^{x-}$ species with superoxo-oxyl and oxo-dioxygenyl character which was also observed to be the case and is discussed in detail below. To clarify language, the $O_2^{x-}$ valence electron is assumed to be partially shared with the coordinating species and:

- Peroxo-superoxyl refers to a coordinated $O_2^{x-}$ species with a $\pi^*(2p)$ valence orbital structure between $[\|]$ and $[1\|]$, a $\nu_{O-O}$ between 1–1.5 and a $\nu_{O-O}$ value between 875–1108 cm$^{-1}$.
- Superoxo-oxyl refers to a coordinated $O_2^{x-}$ species with a $\pi^*(2p)$ valence orbital structure between $[\|]$ and $[1\|]$, a $\nu_{O-O}$ between 1.5–2 and a $\nu_{O-O}$ value between 1179–1552 cm$^{-1}$.
- Oxo-dioxygenyl refers to a coordinated $O_2^{x-}$ species with a $\pi^*(2p)$ valence orbital between $[\|][1\|]$, a $\nu_{O-O}$ between 2–2.5 and a $\nu_{O-O}$ value between 1552–1825 cm$^{-1}$.

### 3.5. Secondary covalent bonding

All the above discussed $\nu_{O-O}$ spectral bands for various $O_2^{x-}$ species ($>200$) have had their estimated bond properties calculated (using eqn (6) and Table 4) and have been plotted against their calculated $B_0$ (Fig. 6). $O_2^{x-}$ species with $B_0$ values divisible by 0.5 can be considered ionic or neutral whilst those with $B_0$ values in-between can be considered to have some amount of electron donation occurring between $O_2^{x-}$ and its coordinating environment, i.e. secondary covalent bonding. Therefore, the $B_0$ is assumed to be on a spectrum and to vary depending on how strongly coordinated the $O_2^{x-}$ diatom is (and therefore, how much of its valence electron it shares).

Experimental data on O–O bond parameters have been reported throughout the scientific literature for different $O_2^{x-}$ complexes. XRC or neutron diffraction techniques have been used to measure the $B_0$ of many species, therefore, the estimated $B_0$ and $B_L$ values calculated from the $\nu_{O-O}$ value can be cross-compared with experimental results to gauge their accuracy. Where the information was available for some of these cross-comparisons are listed in Table S2 (ESI†) with the difference between the measured and calculated $B_L$ values shown as a percentage (the literature is exhaustive and largely disorganised, i.e. individual bond parameters and $\nu_{O-O}$ values tend to be reported in different chemical, crystallographic, spectroscopic journals, and could warrant further compilation in future). Overall, the experimentally reported $B_L$ and the estimated $B_L$ values match well (generally <4% difference), supporting the correlation well. This relationship is useful for analysing O–O properties from experimental $\nu_{O-O}$ results.

$$\%e^{-}\text{donation} = \frac{(B_{0,e} - B_{0,i})}{0.25} \times 100$$

(7)

where: $B_{0,e}$ = calculated bond order, $B_{0,i}$ = nearest $B_0$ to the $B_{0,e}$ deivable by 0.5 (if the nearest bond order is less than $B_{0,e}$ then the species is covalent in the opposite direction). A change in bond order of 0.5 is equivalent to adding/removing an electron from $O_2^{x-}$. A change in bond order of 0.25 is equivalent to adding/removing 1/2 an electron (i.e. secondary covalent bond). An electron donation value of 100% indicates half the $O_2^{x-}$ valence electron is being shared with the coordinate environment, i.e. a fully covalent interaction.

Considering the change in the estimated $B_0$ value of $O_2^{x-}$ species (calculated using the measured $\nu_{O-O}$ values) between strong and weak coordination, the secondary covalent character (i.e. amount of electron donation/abstraction between $O_2^{x-}$ and its coordinating environment) can be roughly estimated (eqn [7]).

### Table 4 Cubic fit constants from Fig. 5 and eqn (6). In principle, all $O_2^{x-}$ parameters can be estimated by simply measuring $\nu_{O-O}$

<table>
<thead>
<tr>
<th>$x$</th>
<th>$y$</th>
<th>$A$</th>
<th>$B$</th>
<th>$C$</th>
<th>$D$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$B_0$</td>
<td>$\nu_{O-O}$</td>
<td>592</td>
<td>−550</td>
<td>884</td>
<td>−184</td>
</tr>
<tr>
<td>$B_0$</td>
<td>1</td>
<td>−0.0933</td>
<td>0.32</td>
<td>−1.157</td>
<td>2.22</td>
</tr>
<tr>
<td>$B_0$</td>
<td>k</td>
<td>−2.607</td>
<td>15</td>
<td>−17.948</td>
<td>8.15</td>
</tr>
<tr>
<td>$B_0$</td>
<td>H</td>
<td>108</td>
<td>−632</td>
<td>1489</td>
<td>−816</td>
</tr>
</tbody>
</table>
interaction between a O$_2^x$ species and its environment based solely on its experimental $\nu_{O-O}$ bond vibration (see examples below). Thus, a value for the amount of electron donation/abstraction for a O$_2^x$ can be teased out of the $\nu_{O-O}$ literature. One side note, these values are not suggested to be precise; as contributions from anharmonicity are excluded etc., rather they are useful tools for quickly calculating approximate values for O$_2^x$ bond parameters, which confirm and enhance the understanding of its vibrational spectra. This is particularly useful for helping to understand dynamic systems that are in flux such as at an electrode surface where countless thermodynamic, kinetic and charge transfer processes are occurring simultaneously, allowing for a general understanding of reaction processes to be rationalised with respect to O$_2^x$ in terms of its $\nu_{O-O}$. This also enables O$_2^x$ species to be used as diagnostic molecules to probe the coordination strength of its environment by measuring the generated $\nu_{O-O}$ band.

Some examples calculations are given: NaO$_2$ (1156 cm$^{-1}$), TBAO$_2$ (~1108 cm$^{-1}$), CuO$_2$ complex A (1033 cm$^{-1}$), have calculated B$_0$ values of 1.549, 1.5, and 1.409 and electron donation values of 19.5, 0 and $\approx$36.4%, respectively. Values greater than $\pm$50% can be considered to have significant secondary covalent character. In NaO$_2$, Na$^+$ can be thought of as abstracting $\approx$20% of an electron from the outer electron bond density of O$_2^+$, which has slight oxygen character, relative to TBAO$_2$. While in CuO$_2$ complex A, the Cu-centred ligand can be thought of as donating $\approx$36% of an electron to O$_2^-$ giving it slight peroxide character, relative to TBAO$_2$.

Therefore:

\[ \% \text{e}^- \text{donation to Na}^+ = \left( \frac{1.549 - 1.5}{0.25} \right) \times 100 = 19.5\% \quad \text{(NaO}_2) \]

\[ \% \text{e}^- \text{donation to TBA}^+ = \left( \frac{1.5 - 1.5}{0.25} \right) \times 100 = 0\% \quad \text{(TBAO}_2) \]

\[ \% \text{e}^- \text{donation to Cu complex A} = \left( \frac{1.409 - 1.5}{0.25} \right) \times 100 = -36.4\% \quad \text{(Cu : O}_2 \text{ A)} \]

It should be emphasised this method is not meant to be precise, rather to be used as a tool for approximating O-O bond properties, at a glance, using empirically derived $\nu_{O-O}$ values for O$_2^x$ species’ with complex structures or coordinate environments. This approach of deriving bond properties from spectral bond vibration values in different coordination environments could deliver comparable results for other homonuclear bonds and provide similar quick approximations, though this would require further work.

### 3.6. The trouble with lithium superoxide

Meta-analysis of reported Raman bands of LiO$_2$ shows that it sticks out like a sore-thumb in terms of both a wide range of reported values and that these values, are in the main, fall outside the trend of all other measured superoxide species (Fig. 2 and 3).\textsuperscript{28} Raman spectral bands in the O$_2^-$ region (1100–1200 cm$^{-1}$) have been reported as being either chemically stable LiO$_2$ or from polyvinylidene fluoride (PVDF) binder degradation during cycling of non-aqueous Li–O$_2$ cells (Fig. \textsuperscript{7}).\textsuperscript{37,38,79} Fundamental in situ SERS studies also report the presence of LiO$_2$ as an intermediate produced during ORR at the roughened Au (rAu) electrode surface.\textsuperscript{10,11,26} LiO$_2$ has long been believed to be both thermodynamically and kinetically unstable at temperatures above 25 K\textsuperscript{10,50,80–85} and it is still unclear whether this observed intermediary species is a pure LiO$_2$ phase or a partially electrolys/electrolyte/electrode stabilised [Li$^+$····O$_2^-$] complex.\textsuperscript{86,87} This has led to some uncertainty over whether LiO$_2$ can exist as a stable discharge product, reaction intermediate, or at all. In light of the trend in $\nu_{O-O}$ of O$_2^-$ discussed previously, looking at the $\nu_{O-O}$ spectra of these reports can add to the debate.

Variation in reported bands comes down to the following: (1) O$_2^-$ is reacting with PVDF abstracting a proton and forming an alkene bond that causes a concomitant shift in the G-band with a serendipitously similar bond vibration to O$_2^-$ observed in situ.\textsuperscript{(2)} $\delta$ proton sites in the structure of PVDF ([–CH$_2$–CF$_2$–]) binder are stabilising O$_2^-$ and Li$^+$ interactions allowing a stabilised [Li$^+$]$_n$····O$_2^-$ complex to deposit and grow on the cathode/binder. This could explain the accompanying distorted carbon G-band ($\approx$1505 cm$^{-1}$).\textsuperscript{(3)} Due to similarities in the wavenumber of their spectral bands, A Li$_2$O$_2$·2H$_2$O species, similar to the Na$_2$O$_2$·2H$_2$O species ($\approx$1135 cm$^{-1}$)\textsuperscript{46} observed in non-aqueous Na–O$_2$ batteries, is being formed when H$_2$O is present in the electrolyte.\textsuperscript{(4)} A thermodynamically irregular or amorphous LiO$_2$ phase or a weakly coordinated [Li$^+$]$_n$····O$_2^-$ complex is formed.\textsuperscript{(5)} A partially intercalated O$_2^-$ species in the carbon cathode that allows Li$^+$ to bind but inhibits disproportionation to Li$_2$O$_2$ is formed. The interaction between O$_2^-$ and the graphitic layers could produce the observed G-band ($\approx$1505 cm$^{-1}$) shift.

From the derived $\Delta H^\circ$ relationship, the expected $\nu_{O-O}$ for LiO$_2$ was estimated to be 1167 ± 10 cm$^{-1}$, with this range shown via a red box in Fig. 7. Region (A) $\nu_{O-O}$ values would fit with the expected value of a LiO$_2$ species. Only one value
matched this and it was reported by Xia et al. who observed a broad $\nu_{\text{O-O}}$ band (1150–1200 cm$^{-1}$) after discharging a non-aqueous Li–O$_2$ battery.\textsuperscript{26} Region (B) $\nu_{\text{O-O}}$ values would be expected to be partially stabilised [Li$^{i+}$–O$_2$]* complexes (where $x < 1$). Current evidence suggests that for species generated on cathodes with PVDF binder (1) is the most probable cause of these reported ‘$\nu_{\text{O-O}}$’ spectral bands. However, (2) and (5) are other possibilities, but again are not a pure LiO$_2$ product. In addition, during in situ spectroelectrochemical investigations of the interface partially electrolyte stabilised [Li$^{i+}$–O$_2$]* complexes are being detected. Though, the report by Xia et al. suggests that it may still be feasible to produce a pure LiO$_2$ phase during discharge if the conditions are right though the presence of toroid discharge product morphologies (associated with water containing electrolytes)\textsuperscript{89,90} may suggest HO$_2$ was being detected rather than LiO$_2$.

4. Conclusions

Superoxide (O$_2$*) vibrational spectral bands reported in the literature were collated (>50 species) and trends in their coordinating environment were observed and described. The $\mathcal{K}$ parameter based on the mass-to-charge ratio and ionic volume of the coordinating ion was derived, which gave an excellent approximation of the collated results. The $\mathcal{K}$ parameter can be calculated with relative ease for most ions and is analogous to charge density and gives a simple low-cost method of quantifying an ion Coulombic attractive strength over oppositely charged species. It was determined that O$_2$* can be used as a diagnostic molecule to probe the coordination strength of its immediate environment, by observing its vibrational spectrum. Furthermore, dioxygen vibrational spectral bands reported in the literature were collated (>200 species). The trends due to the changes in the coordinating environment were observed, where changes in the O–O vibrational spectral band were shown to be a result of electron abstraction/donation from/into the O–O bond via a ‘secondary covalent’ bonding interaction between the dioxygen species and its coordinating environment. A simple cubic approximation was drawn that enabled the bond order, bond dissociation enthalpy, bond length and bond force constants to be estimated solely using experimentally measured O–O spectral bond vibrations. For the bond length this approximation gave estimated values that matched extremely well with reported values (0.5–5%) measured experimentally using X-ray crystallography. It was shown by estimating the bond order allowed for the level of e$^-$ abstraction by the environment (i.e. the strength of the secondary covalent interaction) can also be estimated for any dioxygen species based solely on its vibrational spectra.

Raman spectrum of LiO$_2$ was analysed and most reported bands were found to be too low to be a pure LiO$_2$ phase, suggesting that coordinate [Li$^{i-}$–O$_2$]* complexes or degradation products are the most likely cause of these spectral reports in the non-aqueous Li–O$_2$ battery literature.

Nomenclature

\begin{itemize}
    \item O$_2^+$ Dioxygen cation
    \item O$_2$ Dioxygen
    \item O$_2$\textsuperscript{-} Superoxide anion
    \item O$_2$\textsuperscript{2-} Peroxide anion
    \item ROS Reactive oxide species
    \item O$_2^n$ Dioxide species, where $x = \text{-2, -1, 0, +1}$
    \item $\nu_{\text{O-O}}$ O-O symmetric stretching vibration
    \item [M$^+$\cdots O$_2$\textsuperscript{-}]\textsuperscript{n} Alkali-metal superoxide complex
    \item [C$^+$\cdots O$_2$\textsuperscript{-}]\textsuperscript{n} Cation coordinated superoxide complex
    \item C\textsuperscript{+} Coordinate cation
    \item $M_iQ^{-1}$ Molecular mass-to-charge ratio
    \item $[(\text{Li}^{i+})\cdots O_2\textsuperscript{-}]$ Superoxide rich lithium coordinate complex
    \item XRC X-ray crystallography
    \item $i_v$ Ionic volume
    \item $i_{\text{sa}}$ Ionic surface area
    \item $B_O$ Bond order
    \item $B_L$ Bond length
    \item $B_k$ Bond force constant
    \item $B_H$ Bond dissociation enthalpy
    \item C Cation
    \item A Anion
    \item $\mathcal{K}$ ‘Ionic charge dispersion’
    \item $q$ No. of bonding electrons
    \item $q^*$ No. of anti-bonding electrons
    \item $Q$ Charge on an ion (C)
\end{itemize}

Conflicts of interest

There are no conflicts to declare.

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Notes and references

7 G. Herzberg, Molecular spectra and molecular structure, vol. 1, 1950.
9 T. Livneh, A. Band and R. Tenne, Raman scattering from the peroxide ion in Cs2O2, J. Raman Spectrosc., 2002, 33, 675–676.
39 Z. Jiang, K. C. Lau, H. Wang, J. Wen, D. J. Miller, J. Lu, F. Kang, B. Li, W. Yang, J. Gao, E. Indacochea, L. A. Curtiss...


