A Synthesis of Systematic Mineralogy

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ABSTRACT

This paper presents several new diagrams that document previously unappreciated patterns and predictability in systematic mineralogy. They illustrate, for example, that sulfide minerals of single cations contain only intermediate to soft cations (cations with at least some outer shell electrons). On the other hand, hard cations (those with no outer shell electrons) only enter into a few sulfide minerals, all of which also contain intermediate cations. Among the simple chloride, fluoride, oxide, and carbide minerals of hard cations, ionic potential (charge/radius) of cations is proportional to that of anions (e.g., from KCl to SiC). Among the simple nitrate, sulfate, carbonate, phosphate, and silicate minerals of hard cations, ionic potential of inter-radical cations is inversely proportional to the ionic potential of the cations in the oxysalt radical (e.g., from KNO₃ to ZrSiO₄). Across both of these sets of minerals, non-simple minerals (H₂O- and OH⁻-bearing minerals or minerals with extra O²⁻) in each group typically involve cations of higher ionic potential than those of the simple minerals. The result is predictability and pattern in the compositions of minerals.

INTRODUCTION

Systematic mineralogy is conventionally treated as a listing of minerals in chemical groups defined by anions or anionic groups (e.g., sulfides, oxides, carbonates, etc.). Both university textbooks (e.g. Klein 2002; Nesse 2000) and advanced reference works (e.g.,
Clark 1993; Gaines et al. 1997) provide little explanation of variation within these groups, except among the silicates, and little attempt is made to show relationships between these different groups.

This paper attempts to rationalize systematic mineralogy by showing patterns and relationships in mineral compositions across the entire range of systematic mineralogy. For example, it answers such questions as "Why are there CaF$_2$, CaO, CaSO$_4$, and CaCO$_3$ minerals, but no CaCl$_2$, Ca$_2$C, or Ca(NO$_3$)$_2$ minerals?" The results are a series of diagrams that contextualize the chemical formulas of minerals.

**FUNDAMENTAL PRINCIPLES**

Three fundamental concepts are essential to understanding this paper. First, cations and anions can be divided into categories ranging from "hard" to "soft" (Pearson 1968; Stumm and Morgan 1996) (Fig. 1). Hard cations have no outer-shell electrons, and so their outer shells are like those of a noble gas. Na$^+$, Mg$^{2+}$, and Al$^{3+}$ are good examples of hard cations. Intermediate cations have some outer-shell electrons, and Fe$^{2+}$ and Fe$^{3+}$ are good examples. Soft cations have many outer-shell electrons and are typically large; Au$^+$ and Hg$^+$ are good examples. Hard anions are those in the second row of the periodic table with full outer shells (F$^-$, O$^{2-}$, N$^{3-}$ and C$^{4-}$), whereas anions become progressively softer or more deformable downwards through the table from S$^{2-}$ to Bi$^{2-}$ (Fig. 1). The significance of this scheme is that hard cations coordinate well with hard anions, whereas soft cations coordinate more extensively with soft anions (Pearson 1968; Stumm and Morgan 1996).

Secondly, both cations and anions can be characterized by their ionic potential. Ionic potential is defined as ionic charge divided by ionic radius, or z/r, and it is a measure of the density of charge of an ion and thus of the strength of the ionic bonds made by that ion. On the periodic table, ionic potential increases across the hard cations from less than 1
for $K^+$ to more than 40 for $N^{5+}$ (Fig. 1). Across the intermediate to soft cations, it increases from less than 1 for $Au^+$ to more than 10 for $Se^{6+}$ (Fig. 1). In contrast, ionic potential varies across the anions only from -0.5 for $I^-$ to -1.5 for $C^{4-}$ (Fig. 1). Ionic radius for any one ion varies with coordination number (e.g., $r = 1.00 \, \text{Å}$ for $Ca^{2+}$ in six-fold coordination vs. $1.06 \, \text{Å}$ in eight-fold coordination, in the data from Shannon (1972) used here). However, the effect on ionic potential (2.00 vs. 1.89 in this example) is small relative to the large range of ionic potential across the various ions. Ionic potential has long been used to characterize cations (e.g., Cartledge, 1928a, 1928b; Goldschmidt, 1937), but this paper will show that ionic potential of both cations and anions is also useful in understanding trends in mineralogy.

Both of these concepts are incorporated in the periodic table used in this paper, which is a simplified version of the Earth Scientist's Periodic Table of the Elements and Their Ions (Railsback 2003, 2004). This new table arranges chemical entities by charge, rather than by elemental (uncharged) condition, and as a corollary it shows some elements multiple times (e.g., sulfur as $S^{6+}$, $S^{4+}$, $S^{0}$, and $S^{2-}$ from left to right across the table). This allows the table to separate the hard and intermediate-to-soft cations graphically and to use contours across the table showing variation in ionic potential of both cations and anions (Fig. 1).

Thirdly, all definitions of a "mineral" include the stipulation that minerals be "naturally occurring" (Klein 2002; Nesse 2000) or have "formed as the result of geological processes" (Nickel 1995). Occurrence in nature requires that, for a substance to be a mineral, it must have a melting temperature above at least the coldest earth-surface conditions, it must be sufficiently insoluble to not dissolve after condensation of atmospheric water vapor, and it must be sufficiently non-reactive that it survives long enough to be sampled and identified. Minerals are thus, among all the inorganic chemicals that can be synthesized, the subset that have sufficient chemical stability to survive in
nature. The goal of this paper is to show that existence within this subset can be predicted in terms of hardness and softness of ions and in terms of ionic potential.

**THE "-IDE" MINERALS**

One of the most fundamental observations about systematic mineralogy is that cations forming sulfides, bromides, tellurides, iodides, and bismuthides are almost entirely intermediate to soft cations, rather than hard cations (Fig. 1). Of the 56 sulfide minerals of single cations, all contain intermediate or soft cations rather than hard cations (Fig. 1). Of all 136 sulfide minerals listed by Clark (1993) and Mandarino and Black (2004), only 13 contain any hard cations at all, and then typically only as minor constituents of rare minerals, as in rasvumite (KFe₂S₃). Among the intermediate to soft cations, those of lesser ionic potential, such as Ag⁺, Au⁺, and Hg⁺, coordinate with softer anions like Br⁻, I⁻, and Bi²⁻. Thus the "softness" of the intermediate to soft cations is generally correlative with the softness of the anions with which they combine to form minerals (Fig. 1). Another reflection of the propensity of the cations that bond with soft anions to not bond to hard anions such as O²⁻ is that only 4 of the 350 sulfides, selenides, tellurides, arsenides, and bismuthides listed by Clark (1993) and Mandarino and Black (2004) are hydrous, and none contain F⁻.

In contrast to the behavior of intermediate to soft cations, hard cations form minerals almost exclusively with the hard anions high in the periodic table (Cl⁻, F⁻, O²⁻, and C⁴⁻, or with the O²⁻ of oxysalt radicals) (Figure 1). Cl⁻, intermediate between the hard and soft anions, combines to make minerals with cations of all sorts, from hard cations like Na⁺ and K⁺ in halite and sylvite to intermediate cations like Mn²⁺ in scacchite to soft cations like Hg⁺ in calomel (Fig. 1).
CHLORIDES, FLUORIDES, OXIDES, AND CARBIDES OF HARD CATIONS

Both the properties of minerals and the existence of minerals vary regularly across the portion of the periodic table containing hard cations (left side of Figure 1). For example, melting temperature, hardness, and bulk modulus of oxides are greatest, and solubility is least, in a swath from Be\(^{2+}\) to Zr\(^{4+}\) (Fig. 2). Simple oxide minerals form with cations in the same swath (Fig. 2). This variability can be rationalized in terms of ionic potential. Cations of low ionic potential, such as K\(^+\), Na\(^+\), and Ba\(^{2+}\), have such diffuse positive charge that they only make weak bonds with O\(^{2-}\), and they fail to shield the charge of the anions (O\(^{2-}\)) from each other. Cations of intermediate ionic potential, such as Al\(^{3+}\) to Ti\(^{4+}\), have sufficiently dense positive charge to make strong bonds to O\(^{2-}\) and to shield anions' charge from each other, so that stable three-dimensional cation-oxygen frameworks (i.e., oxide minerals) can form. Cations of high ionic potential, such as C\(^{4+}\), N\(^{5+}\), and S\(^{6+}\), have such focused charge that they make very strong cation-oxygen bonds, but their charge is so dense that the anions cannot shield one positive charge from another, so that the resulting cation-cation repulsion precludes formation of oxide minerals.

Analogous but shifted patterns can be seen in the melting temperatures of chlorides, fluorides, nitrides, and carbides (Fig. 3). Chlorides of Ra\(^{2+}\), Ba\(^{2+}\), and La\(^{3+}\) have the highest melting temperatures of chlorides, whereas fluorides of La\(^{3+}\) and Sc\(^{3+}\) have the highest melting temperatures of fluorides, oxides of Th\(^{4+}\) and Mg\(^{2+}\) have the highest melting temperatures among oxides, nitrides of Hf\(^{4+}\) and Al\(^{3+}\) have the highest melting temperatures among nitrides, and carbides of Zr\(^{4+}\) and Ti\(^{4+}\) have the highest melting temperatures among carbides. The highs in melting temperature thus progress through ions of increasing ionic potential, from Ba\(^{2+}\) to Ti\(^{4+}\), as one progresses through anions of more negative ionic potential, from Cl\(^-\) to C\(^{4-}\) (Fig. 3).

The progression in melting temperatures observed from chlorides to carbides is mirrored by the existence of minerals in each group (Fig. 4). Simple chlorides form only
with cations of lowest ionic potential (K+ and Na+), fluorides form with cations of low ionic potential (from K+ to Mg2+ and La3+), oxides form with cations of intermediate ionic potential (e.g., Al3+ and Ti4+), and the lone carbide mineral of a hard cation forms with a cation of relatively high ionic potential, Si4+. The general pattern is thus a progression from cations of low ionic potential combining with anions of less negative ionic potential to form minerals, to cations of high ionic potential combining with anions of more negative ionic potential to form minerals. The result is predictability of the chemistry of the chloride, fluoride, oxide, and carbide minerals (Fig. 4).

NITRATES, SULFATES, CARBONATES, PHOSPHATES, AND SILICATES OF HARD CATIONS

The overall pattern seen in Figure 4 can also be seen in the compositions of minerals formed by oxysalt anions, the "-ate" minerals (Fig. 5). Simple nitrates (nitrates with no H+, no H2O, no OH−, and no extra O2−) form only with cations of lowest ionic potential (K+, Na+, and Ba2+, as in NaNO3). Simple sulfates form only with cations of slightly higher ionic potential, from K+ to Ca2+. Simple carbonates form with cations of still higher ionic potential, from Na+ to Mg2+. Simple phosphates form with cations of yet higher ionic potential, from Na+ to Al3+. Simple silicates form with cations of even higher ionic potential, from Na+ to Zr4+ (as in ZrSiO4) (Fig. 5). Thus as one progresses from the cation of highest ionic potential to form an oxysalt radical, N5+ in nitrate, through this sequence to the cation of lowest ionic potential to form an oxysalt radical, Si4+ in silicate, one progresses from inter-radical cations of low ionic potential (e.g., K+) to high ionic potential (e.g., Zr4+). This progression can be rationalized as the result of repulsion of all but the weakest cations by the dense positive charge of N5+, and by the progressive acceptance of more densely charged inter-radical cations (i.e., from K+ to Mg2+ to Al3+ to...
Zr$^{4+}$) by radical-forming cations of lesser ionic potential (i.e., from N$^{5+}$ to S$^{6+}$ to C$^{4+}$ to P$^{5+}$ to Si$^{4+}$) (Fig. 5).

Figure 5 shows that there is also predictability in the compositions of non-simple minerals of oxysalt anions. Fields for H$_2$O-bearing minerals, OH$^-$-bearing minerals, and minerals with extra O$^{2-}$ progress to the right from the fields for simple minerals (Fig. 5). This can be rationalized in terms of ionic potential and cation-cation repulsion, in that incorporation of cations of progressively higher ionic potential to the right in these diagrams is allowed first by neutral groups (H$_2$Os) that simply shield cations from each other and then by anionic groups (OH$^-$ and extra O$^{2-}$) that help cancel positive charge. An alternate rationalization in terms of aqueous speciation would explain this trend as resulting from failure of cations of higher ionic potential to dehydrate or dissociate from ligands before entering mineral structures. The result is the same: increasing incorporation of H$_2$O or anions with incorporation of cations of higher ionic potential (Fig. 5).

Minerals far to the right in the diagrams (i.e., minerals incorporating cations of high ionic potential) also commonly incorporate cations of low ionic potential (e.g., Na$^+$) that can be viewed as diluting the ionic potential of the entire population of cations in a mineral. For example, mongolite (Ca$_4$Nb$_6$Si$_5$O$_{24}$(OH)$_{10,n}$H$_2$O) is a silicate in which the high ionic potential of the principal cation, Nb$^{5+}$, is diluted by that of a cation of lower ionic potential (Ca$^{2+}$), is shielded by H$_2$O molecules, and is countered by the negative charge of both ten OH$^-$ groups and four O$^{2-}$ in excess of those dictated by the stoichiometry of SiO$_4^{4-}$ tetrahedra.

In contrast to the fields of non-simple minerals to the right of the fields for simple minerals in Figure 5, fields for H$^+$-bearing minerals consistently fall to the left. In this respect, the bicarbonates, biphosphates, and bisulfates can be viewed as minerals in which cations of low ionic potential provide sufficiently little density of positive charge that H$^+$
ions of undissociated HCO$_3^-$, HPO$_4^{2-}$, and HSO$_4^-$ are tolerated in minerals, opposite the considerations described in the paragraph above.

**IONIC RADIUS AND RADIUS RATIO ARE NOT SUFFICIENT TO EXPLAIN EXISTENCE AND NON-EXISTENCE OF MINERALS.**

A skeptic might suggest that the patterns shown in Figures 4 and 5 need not be explained in terms of ionic potential but might be explained simply as a matter of radius ratio and coordination. For example, the conventional wisdom regarding the non-existence of LiCl, LiNO$_3$, and Li$_2$SO$_4$ as minerals is that Li$^+$ is too small relative to the cations in the stable comparison group to form stable structures (in this example, the comparison group would be Na and K chlorides, nitrates, and sulfates). However, there are in fact synthetic crystalline LiCl, LiNO$_3$, and Li$_2$SO$_4$ phases (Lide, 2003), so the non-existence of these compounds as minerals is not for lack of a possible structure. Likewise, the conventional wisdom regarding the non-existence of B$_2$O$_3$ as a mineral is that B$^{3+}$ is too small compared to Al$^{3+}$ to form a stable structure like that in corundum. However, there is in fact synthetic crystalline hexagonal B$_2$O$_3$ solid (Lide, 2003), again demonstrating that non-existence as a mineral is not for lack of a possible structure. For the example of LiCl, LiNO$_3$, and Li$_2$SO$_4$, Figure 6 shows that these non-mineral crystalline compounds are more soluble than the corresponding mineral-forming compounds, and Figure 2 likewise shows that non-mineral crystalline B$_2$O$_3$ is much more soluble than Al$_2$O$_3$. This greater solubility is clearly not for lack of a possible crystalline structure. It instead probably results from a characteristic of that structure, such as greater cation-cation repulsion (e.g., Li$^+$-Li$^+$, Li$^+$-N$^{5+}$, Li$^+$-S$^{6+}$, and B$^{3+}$-B$^{3+}$) between cations of higher ionic potential than that found in the mineral-forming compounds of cations of lower ionic potential.

This is of course not to say that cation size, radius ratios, bond lengths, and coordination are unimportant to mineralogy. For example, simple carbonates of divalent
cations occur in two structural groups, with smaller cations forming rhombohedral carbonates with six-fold coordination, and larger cations forming orthorhombic carbonates with nine-fold coordination (Fig. 7). Coordination of three O²⁻ anions around C⁴⁺ in these carbonate minerals is likewise dictated by radius ratio. At the ends of the ranges of cation radius for each group, solubility is high (Fig. 7), illustrating the importance of bond length in determining whether a cation forms a mineral within a particular isostructural group. However, the existence of synthetic solid BeCO₃ (Wiberg, 1971, p. 614) again disproves the conventional wisdom that the non-existence of BeCO₃ as a mineral is because Be²⁺ is too small, relative to the cations of rhombohedral and orthorhombic carbonates, to form stable structures. The trends shown in this paper (Figs. 4 and 5) would instead suggest that Be²⁺-C⁴⁺ repulsion is the more likely reason that BeCO₃ is sufficiently unstable that it is not known as a mineral.

**POLYHEDRAL LINKAGE**

The concept of cation-cation repulsion between oxysalt radicals used to explain trends in Figures 2 and 5 is also useful in understanding the extent of polyhedral linkage in different groups of oxysalt minerals (Fig. 8). For example, in minerals of oxysalt radicals of cations of highest ionic potential (N⁵⁺, S⁶⁺, C⁴⁺, and P⁵⁺), no mineral structures involve polyhedral linkage (there is no sharing of oxygens by NO₃⁻, SO₄²⁻, CO₃²⁻, and PO₄³⁻ radicals). As a result, all minerals in these groups of minerals have chemical formulas that preserve the identity of the radical - for example, all nitrates have chemical formulas involving "NO₃", all sulfates have formulas involving "SO₄", and so on (Fig. 8).

In minerals of oxysalt radicals of cations of slightly lower ionic potential (Cr⁶⁺, Mo⁶⁺, and W⁶⁺), most mineral structures involve no polyhedral linkage, but a small proportion of minerals exhibit limited sharing of oxygens between polyhedra. Chemical formulas reflect this, in that most chromates have formulas in which "CrO₄" appears, but a
few have formulas like $K_2Cr_2O_7$ reflecting the sharing of an oxygen between paired chromate groups, analogous to the tetrahedral linkage in sorosilicates (Fig. 8).

Finally, in minerals of oxysalt radicals of cations of still lower ionic potential ($B^{3+}$, $Si^{4+}$, $V^{5+}$, $Nb^{5+}$, and $Ta^{5+}$), most mineral structures involve polyhedral linkage (Fig. 8).

The most familiar example is the silicates, where only the nesosilicates have no polyhedral linkage, and where polyhedral linkage progresses from sharing of just one oxygen between pairs of tetrahedra (as in the chromates) to sharing of all oxygens in frameworks (in the tectosilicates). The resulting chemical formulas obscure the identity of $SiO_4^{4-}$ groups, in that they end with "$SiO_3$", "$Si_2O_5$", and even "$SiO_2$". This extent of polyhedral linkage in the silicates is of course well known, but it is also characteristic of the borates, vanadates, niobates, and tantalates because of the similar ionic potential of $B^{3+}$, $V^{5+}$, $Nb^{5+}$, and $Ta^{5+}$ (Fig. 8).

In this context, the nesosilicates can be viewed as the only silicate analogs of the phosphates, carbonates, sulfates, and nitrates, in which no polyhedral linkage occurs. The position of a nesosilicate, olivine, at the high-temperature end of Bowen's Reaction Series can thus be rationalized as the result of a silicate structure minimizing tetrahedral repulsion, like those of the phosphates et al.. On the other hand, silicates lower in the discontinuous series involve progressively more tetrahedral linkage, suffer more tetrahedral repulsion, and have lower melting temperatures.

**CONCLUDING REMARKS**

Hardness and ionic potential of both cations and anions are useful in explaining existence and non-existence of possible chemical substances as minerals, and in explaining variation in mineral properties. These concepts have long been used to characterize cations, but their application to anions is also useful in explaining trends and patterns in mineralogy. For example, sulfide minerals of single cations only form with intermediate to
soft cations (Fig. 1). Conversely, hard cations only enter into a few sulfide minerals, all of which also contain intermediate cations. The dearth of hard cations in sulfide minerals means that most sulfide minerals are anhydrous. Chloride, an anion intermediate between hard and soft anions, is distinct in forming minerals with hard cations, intermediate cations, and soft cations from Na\(^+\) to Ag\(^+\) (Fig. 1).

Among the simple chloride, fluoride, oxide, and carbide minerals of hard cations, ionic potential of the cation is proportional to that of the anion. Non-simple (H\(_2\)O- and OH\(^-\)-bearing minerals) in each group typically involve cations of higher ionic potential than those of the simple minerals (Fig. 4). Among the simple nitrate, sulfate, carbonate, phosphate, and silicate minerals of hard cations, ionic potential of the cation (e.g., Ca\(^{2+}\) in calcite) is inversely proportional to the ionic potential of the cation in the oxysalt radical (e.g., C\(^{4+}\) in calcite). Non-simple minerals (H\(_2\)O- and OH\(^-\)-bearing minerals and minerals with extra O\(^{2-}\)) in each group typically involve cations of higher ionic potential than those of the simple minerals. Such minerals also commonly incorporate other cations of low ionic potential (Fig. 5). Polyhedral linkage is increasingly common in oxysalt mineral groups in which the ionic potential of the cation in the oxysalt radical (e.g., C\(^{4+}\) in calcite) is lower (Fig. 8).

Figures 1, 4, and 5 show that, amidst the diverse array of chemical formulas of minerals, we can find patterns and make predictions. For example, Figure 1 shows that we can expect sulfides of intermediate-to-soft cations, and we can conversely expect oxides and fluorides of hard cations, but we are unlikely to work with or discover a sulfide mineral solely of Ca\(^{2+}\) or an oxide of Ag\(^+\). As another example, Figure 4 shows that we can expect to encounter chlorides of cations of low ionic potential, and we might discover a BaCl\(_2\) mineral, but we needn't worry ourselves about BeCl\(_2\) minerals. We likewise should know about the SiC mineral (moissanite) and might someday discover a TiC mineral, but we needn't worry about a Ca\(_2\)C\(^-\) mineral. Similarly, Figure 5 leads us to expect simple nitrates
of cations of low ionic potential, and hydrous phosphates of cations of intermediate ionic potential, but we are unlikely to learn about or discover an Al(NO₃)₃ mineral or a K₅PO₄(OH)₂ mineral. We can thus make predictions about chemical formulas of minerals, if we appreciate the distinction between hard and soft anions and between hard and soft cations, and if we use the ionic potential of cations and anions to understand minerals.

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FIGURE CAPTIONS

Figure 1. Existence of simple carbide, oxide, fluoride, chloride, sulfide, bromide, iodide, and bismuthide minerals across the periodic table. Note the presence of symbols for sulfides, bromides, iodides, and bismuthides among the intermediate to soft cations in the middle of the table, and the absence of such symbols from the hard cations at the left of the table. Minerals are from Clark (1993) and Gaines et al. (1997). Format of the periodic table used here is an abridged and simplified form of the Earth Scientist's Periodic Table of the Elements and Their Ions (Railsback 2003, 2004). This diagram's left side, which shows hard cations, is the field shown in Figures 2 to 5 and 8.

Figure 2. Variation in properties of mineral and non-mineral oxides of hard cations, shown across the periodic table (cf. Fig. 1). Melting temperatures (°C) are from pages 4-38 to 4-96 of Lide (2003). Solubilities, expressed as the log of the activity of dissolved species, were calculated from thermodynamic data from Robie et al. (1979), Robie and Hemingway (1995), Drever (1988), Langmuir (1997), and Stumm and Morgan (1996). Hardness data are from Nickel and Nichols (1991) with data on TiO₂ from Dubrovinskaia et al. (2001). Data on bulk modulus (K_s in GPa) are from Bass (1995), Knittle (1995) and Leger et al. (1993). Circles show the relative size of hard cations across the periodic table.

Figure 3. Variation in melting temperatures of chlorides, fluorides, oxides, nitrides, and carbides of hard cations. The five panels are arranged in order of ionic potential of the anions (see Figure 1); note the overall progression to the right of the shaded fields of highest melting temperatures. Data are from pages 4-38 to 4-96 of Lide (2003).

Figure 4. Chloride, fluoride, oxide, nitride, and carbide minerals of hard cations (there is no nitride mineral of a hard cation). Note the shift of the blue field for simple minerals
(minerals of one cation with no H₂O or OH⁻) to the right as one progresses upward through the diagram. The five panels are arranged in order of ionic potential of the anions (see Figure 1).

Figure 5. Nitrate, sulfate, carbonate, phosphate, and silicate minerals of hard cations. Note the shift of the blue field for simple minerals (minerals of one cation with no H₂O or OH⁻) to the right as one progresses upward through the diagram. The five panels are arranged in order of ionic potential of the cation in the oxysalt radical (see Figure 1). Not all possible minerals are shown for lack of space.

Figure 6. Solubility of chlorides, nitrates, and sulfates of alkali metal cations (A) and alkali earth cations (B) as a function of ionic potential (top of diagram) and ionic radius (bottom of diagram). Filled symbols indicate minerals. Open symbols indicate crystalline solids (Lide, 2003) not known from nature as minerals. Note that, contrary to conventional wisdom, failure of smaller cations to form minerals is not for lack of a possible crystalline structure. Solubility data are from Lide (2003).

Figure 7. Solubility of carbonates of alkali earth cations as a function of ionic potential (top of diagram) and ionic radius (bottom of diagram). Filled symbols indicate minerals, with triangles for rhombohedral minerals and squares for orthorhombic minerals. Open symbol is for synthetic BeCO₃ (Wiberg, 1971), which is not known from nature as a mineral. Four fold-coordination for Be²⁺ in BeCO₃ is inferred from four-fold coordination in BeNO₃ (Wiberg, 1971). Dashed line indicates that solubility of BeCO₃ is not well constrained. Solubility data for minerals are from Robie and Hemingway (1995).
Figure 8. Extent of polyhedral linkage by oxysalt radicals of hard cations across the periodic table. Representative chemical formulas and schematic sketches of structures are shown to right and below.