Supplemental notes for Lecture 8

Lattice energies (or melting points) of ionic solids

As we mentioned in class, we use Coulomb’s law to determine relative strengths (energies) of ionic solids. We cannot measure lattice energy directly, but we can measure the melting points of ionic solids, which follow the same trend in relative strength. But remember: lattice energy (an energy, J units) ≠ melting point (a temperature, °C units)!

From Coulomb’s law, we know that $E \propto \frac{Q_1 Q_2}{d}$. That is, $E \uparrow$ as $Q_1 Q_2 \uparrow$, and $E \uparrow$ as $d \downarrow$.

To determine the relative lattice energies (or melting points), we first look at the charges ($Q_1$ and $Q_2$) to coarsely separate the ionic solids, and then use distance, based on our understanding of the periodic trends of ion sizes, to distinguish relative strength within each subclass. Let’s look at an example.

Arrange the following in order of increasing melting point (or lattice energy):

NaCl  MgS  CaS  LiCl

First, we separate based on charge.

<table>
<thead>
<tr>
<th>$Q_1 = +1$; $Q_1 = -1$</th>
<th>$Q_1 = +2$; $Q_1 = -2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td>MgS</td>
</tr>
<tr>
<td>LiCl</td>
<td>CaS</td>
</tr>
</tbody>
</table>

So we know that

\[
\text{NaCl} \quad < \quad \text{MgS}
\]

\[
\text{LiCl} \quad < \quad \text{CaS}
\]

based simply on the charges. Now we have to distinguish between NaCl and LiCl, and between MgS and CaS. We do this based on ion sizes. Because Cl\(^-\) is constant between the NaCl and LiCl, we only consider the Na\(^+\) and Li\(^+\) sizes. Li\(^+\) is smaller than Na\(^+\). (Review your periodic trends to understand why this is the case.) Therefore, we rank NaCl < LiCl with respect to the relative melting point (or lattice energy). We make a similar argument for determining that CaS < MgS.

Thus, we find that NaCl < LiCl < CaS < MgS.