Announcements

Monday, November 30, 2009

Exam 3 is next Monday, Dec 7 covering 6.7-6.8, and chapters 7-9 (through Wednesday's lecture).

MasteringChemistry due dates (all at 11:59 pm)

- Ch 8: Wed, Dec 2
- Ch 9: Fri, Dec 4

Final exam information (including a list of topics to study) is posted to the webpage under handouts.

- Exam is comprehensive
- 70 multiple choice questions, 110 minutes
- **NO programmable calculators** - buy a non-programmable scientific calculator now if you have not yet
- Wed Dec 16 1:30 pm - 3:30 pm

Ionization energy and Electron affinity

$2\ Li(s) + 2\ H_2O(l) \rightarrow 2\ LiOH(aq) + H_2(g)$ (Li^+, OH^-)

oxidations

$Li \rightarrow Li^+ + 1e^-$ $IE_1 = 520$ kJ/mol

$Na \rightarrow Na^+ + 1e^-$ $IE_1 = 496$ kJ/mol

$K \rightarrow K^+ + 1e^-$ $IE_1 = 419$ kJ/mol

K is the strongest reducing agent of these

Reactions of the Alkali Metals with Water

Lithium Sodium Potassium

Electron affinity (EA): energy change associated with an electron being added to a neutral atom.

$Li + 1e^- \rightarrow Li^-$ EA = -60 kJ/mol

$O + 1e^- \rightarrow O^-$ EA = -141 kJ/mol

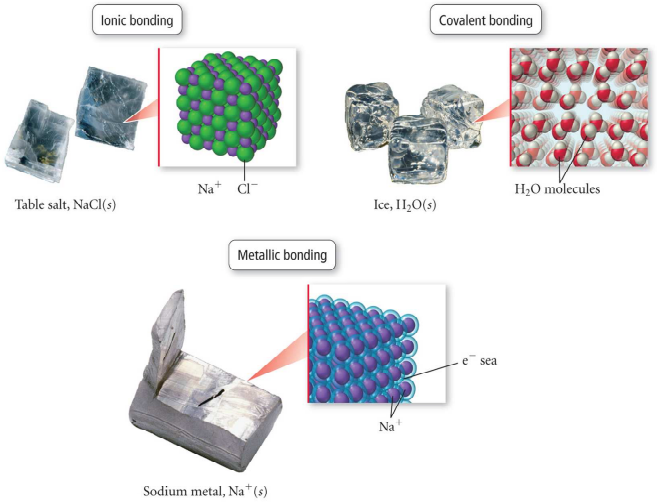
$F + 1e^- \rightarrow F^-$ EA = -328 kJ/mol

particularly stable ion

same e^- as [Ne]

Three types of chemical bonding:

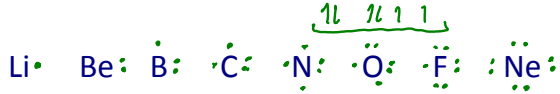
Bond	Atoms	Electron behavior
Ionic	Metal + nonmetal	Electrons transferred
Covalent	Nonmetal + nonmetal	Electrons shared
Metallic	Metal + metal	Electrons pooled



Coulomb's law: $E = k\frac{q_1 q_2}{r}$ (two charges; r ← dist between 2 charges)

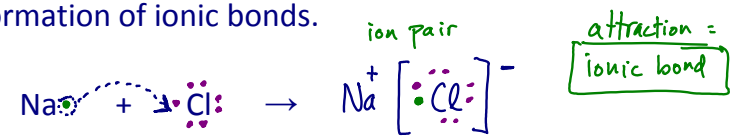
- when **opposite** charges are brought closer, potential energy becomes more negative ($\to -\infty$) E = negative
- when **like** charges are brought closer, potential energy increases ($\to \infty$) E = positive

Recall that we can use dots to show valence electrons - these are called **Lewis electron-dot structures**:



Ionic bond: attraction of two oppositely-charged ions (recall Coulomb's law)

Lewis structures can be used as a simple way to show formation of ionic bonds.

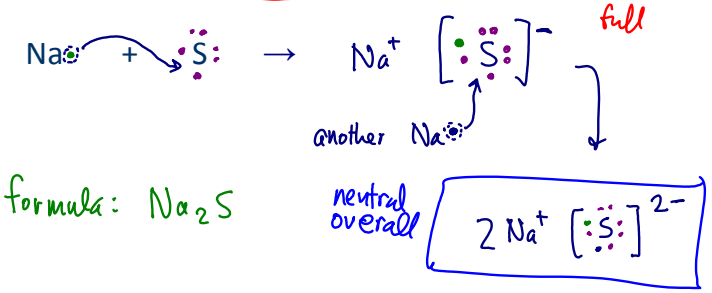


Octet rule: main group (s or p block) atoms or ions tend to be stable when they have 8 valence electrons

full s + p

Na = $1s^2 2s^2 2p^6 3s^1$ $Na^+ = 1s^2 2s^2 2p^6$ (full)

Cl = $1s^2 2s^2 2p^6 3s^2 3p^5$ $Cl^- = 1s^2 2s^2 2p^6 3s^2 3p^6$ (full)



Lattice energy

Actual formation of an ionic compound is usually very exothermic:

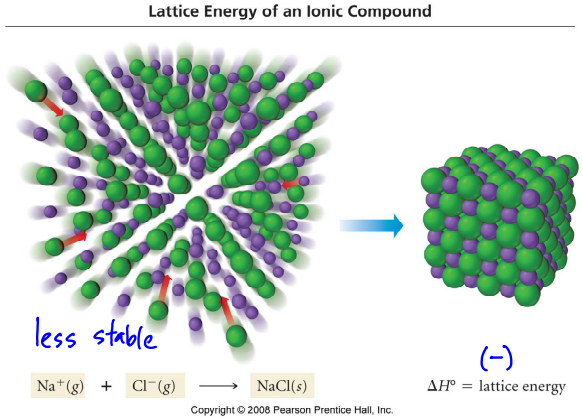
$$Na(s) + \tfrac{1}{2}Cl_2(g) \longrightarrow NaCl(s);\ \Delta H_f^o = -410\ kJ/mol$$

exothermic

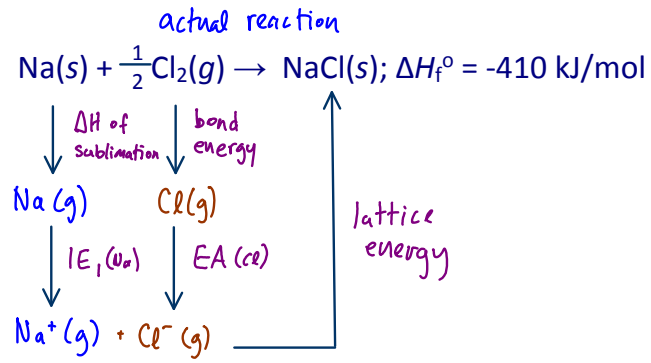
…but when we add up the ionization energy of Na and the electron affinity of Cl, that's actually endothermic!

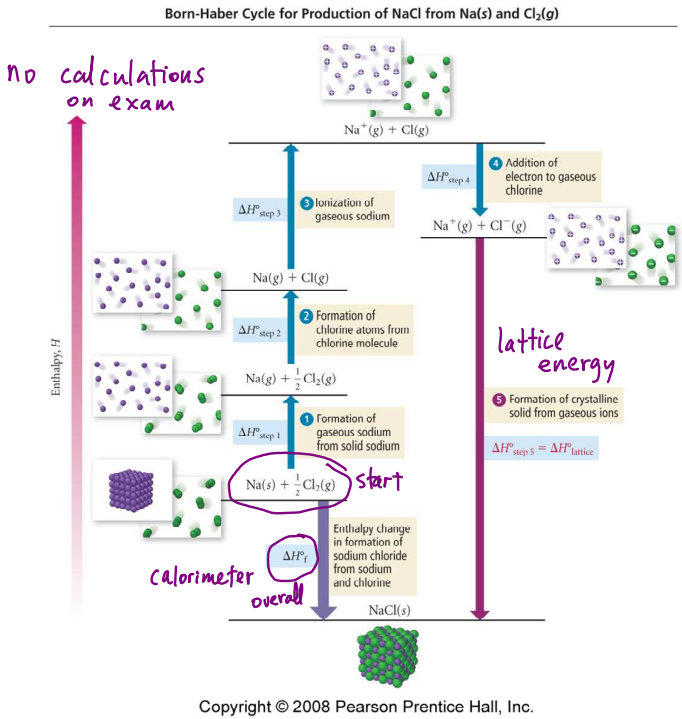
$Na \longrightarrow Na^+ + e^-$; $IE_1 = +496$ kJ/mol
$Cl + e^- \longrightarrow Cl^-$; $EA = -349$ kJ/mol
+147 kJ/mol

… so the release of energy does not come from formation of the ions! It actually primarily comes from the **lattice energy**, the energy associated with forming a crystalline ionic lattice from separated, gaseous ions.



Born-Haber Cycle: Calculating lattice energy





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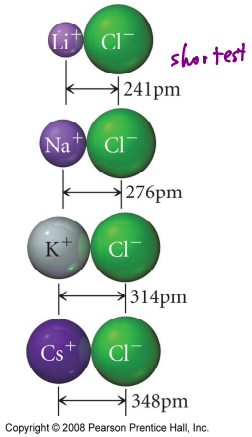
Periodic trends in lattice energy



shortest

Ion size:

Compound	Lattice energy
LiCl	-834 kJ/mol
NaCl	-787
KCl	-701
CsCl	-657

get less energy out b/c ions are not as close together

Ion charge: larger effect than ion size

Compound	Lattice energy	Melting point
NaF	-910 kJ/mol	993 °C
CaO	-3414 kJ/mol	2572 °C

$q_1 q_2$: (+1)(-1), (+2)(-2); (+) (−), (2+)(2−)

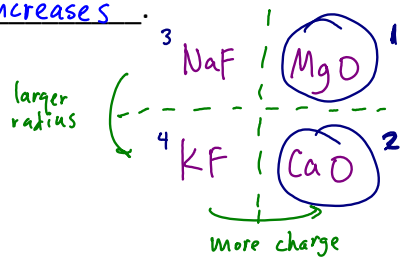
4x charge

Coulomb's law: $E = k\frac{(q_1 q_2)}{r}$

most favorable:
- big charges
- small ions

most E given away

As lattice energy becomes more negative, the ions become more difficult to separate, and melting point increases.



Lattice energy

1 Most negative > 4 least negative