REVIEW OF ATOMIC IONS

The nature of atom-bound electrons, as we have seen, makes the ionic forms of some atoms much more energetically stable—and thus more common—than the neutral form. For example, we saw that inert gas atoms don't tend to form ions because they have a full **octet** of electrons in their outer shell. Halides (F, Cl, Br, I, At), on the other hand, have one more electron than the next smaller inert gas (e.g. Na compared to Ne), and tend to form +1 ions by losing that "extra" electron.

The periodic table below shows the most common forms of most of the elements. Where two ions are listed, both are observed but the top form is dominant. Notice that while most of the ions follow the trends we'd expect from electron configuration, there are exceptions. These can seem disturbing in a first run through chemistry, but each can be rationalized in terms of an interplay of competing factors that you will learn about as your study of chemistry becomes more sophisticated.

In the atoms section I mentioned that the typical way an atom with more than an octet of electrons loses one (or more) is by coming into contact with one that has the opposite "need."

For example, when sodium metal comes into contact with fluorine gas, the fluorine readily takes an electron from sodium, and both are left in an energetically favorable state. What results is an ion pair of neutral overall charge, $Na^{+}F^{-}$, or just **NaF**.

Atomic Ions
Dominant form on top

1A	2A	3B	4B	5B	6B	7B	8B			1B	2B	3A	4A	5A	6A	7A	0
H^{+}																	He
Li^{+}	Be^{2+}											B	C	N^{3-}	O^{2-}	F^{-}	Ne
Na^{+}	Mg^{2+}											Al^{3+}	Si	P^{3-}	S^{2-}	Cl^{-}	Ar
K^{+}	Ca^{2+}	Sc^{3+}	Ti^{3+} Ti^{4+}	V^{3+} V^{5+}	Cr^{3+} Cr^{2+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^{2+} Cu^{+}	Zn^{2+}	Ga^{3+}	Ge^{4+}	As^{3-}	Se^{2-}	Br^{-}	Kr
Rb^{+}	Sr^{2+}	Y^{3+}	Zr^{4+}	Nb^{5+} Nb^{3+}	Mo^{6+}	Tc^{7+}	Ru^{3+} Ru^{4+}	Rh^{3+}	Pd^{2+} Pd^{4+}	Ag^{+}	Cd^{2+}	In^{3+}	Sn^{4+} Sn^{2+}	Sb^{3+} Sb^{5+}	Te^{2-}	I^{-}	Xe
Cs^{+}	Ba^{2+}	La^{3+}	Hf^{4+}	Ta^{5+}	W^{6+}	Re^{7+}	Os^{4+}	Ir^{4+}	Pt^{4+} Pt^{2+}	Au^{3+} Au^{+}	Hg^{2+} Hg^{+}	Tl^{+} Tl^{3+}	Pb^{2+} Pb^{4+}	Bi^{3+} Bi^{5+}	Po^{2+} Po^{4+}	At^{-}	Rn
Fr^{+}	Ra^{2+}	Ac^{3+}															

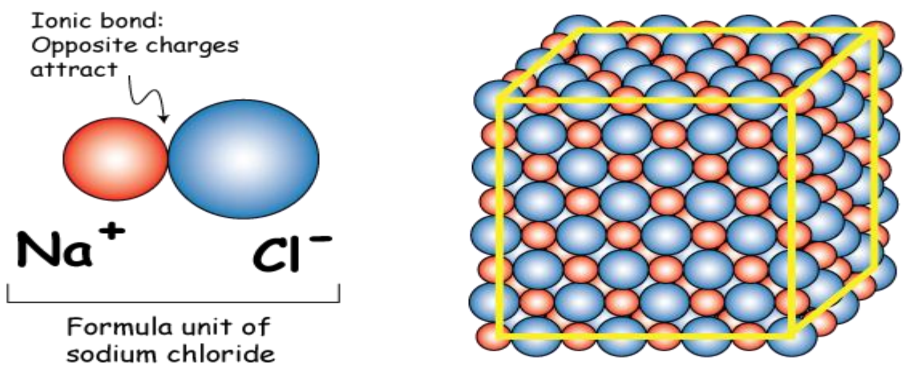
Ionic Compounds

Nature tends to neutralize charges, so ions of opposite charge tend to attract to make ionic compounds. Some examples are NaCl (table salt), KCl, CsF, and RbBr. Each is composed of a positive ion and a negative ion.

We try not to refer to ionic compounds as **molecules** (which we'll get to in a while), because the really never exist in that form.

The abbreviation NaCl, for example, is best thought of as the **formula unit** of sodium chloride (we'll talk about how to name ionic compounds later, too). It's really just the ratio in which the constituent atoms are found in the pure material.

The figure below shows the formula unit of sodium chloride. The relative sizes of the atoms are accurate, although we know they don't have hard edges. On the right is a picture of how sodium chloride formula units can stack together in a cubic configuration to form the well-known salt crystals you spill on the table. The cubic crystal of NaCl is charge-neutral: For every positive ion, there is a negative ion.



Notice that some of our ions are **multiply charged**. For example, the magnesium ion carries a +2 charge. Nature still tends to neutralize this charge, except that now we either need two -1 charged ions or one with a -2 charge. We can form many ionic compounds that satisfy this need. MgO consists of Mg^{2+} and O^{2}-, and $MgCl_2$ is Mg^{2+} and two Cl^{-} ions.

Source: http://www.drcruzan.com/Chemistry_Ionic.html