Because binary ionic compounds are confined mainly to group 1 and group 2 elements on the one hand and group VI and VII elements on the other, we find that they consist mainly of ions having an electronic structure which is the same as that of a noble gas. In calcium fluoride, for example, the calcium atom has lost two electrons in order to achieve the electronic structure of argon, and thus has a charge of +2:

\[
1s^22s^22p^63s^23p^64s^2 \rightarrow 1s^22s^22p^63s^23p^6 + 2e^-
\]

By contrast, a fluorine atom needs to acquire but one electron in order to achieve a neon structure. The resulting fluoride ion has a charge of –1:

\[
1s^22s^22p^5 + e^- \rightarrow 1s^22s^22p^6
\]

The outermost shell of each of these ions has the electron configuration \(ns^2np^6\), where \(n\) is 3 for \(Ca^{2+}\) and 2 for \(F^-\). Such an \(ns^2np^6\) noble-gas electron configuration is encountered quite often. It is called an octet because it contains eight electrons. In a crystal of calcium fluoride, the \(Ca^{2+}\) and \(F^-\) ions are packed together in the lattice shown below. Careful study of the diagram shows that each \(F^-\) ion is surrounded by four \(Ca^{2+}\) ions, while each \(Ca^{2+}\) ion has eight \(F^-\) ions as nearest neighbors.

Thus there must be twice as many \(F^-\) ions as \(Ca^{2+}\) ions in the entire crystal lattice. Only a small portion of the lattice is shown, but if it were extended indefinitely in all directions, you could verify the ratio of two \(F^-\) for every \(Ca^{2+}\). This ratio makes sense if you consider that two \(F^-\) ions (each with a –1 charge) are needed to balance the +2 charge of each \(Ca^{2+}\) ion, making the net charge on the crystal zero. The formula for calcium fluoride is thus \(CaF_2\).
Newcomers to chemistry often have difficulty in deciding what the formula of an ionic compound will be. A convenient method for doing this is to regard the compound as being formed from its atoms and to use Lewis diagrams. The octet rule can then be applied. Each atom must lose or gain electrons in order to achieve an octet. Furthermore, all electrons lost by one kind of atom must be gained by the other.

An exception to the octet rule occurs in the case of the three ions having the He $1s^2$ structure, that is, $\text{H}^-$, $\text{Li}^+$, and $\text{Be}^{2+}$. In these cases two rather than eight electrons are needed in the outermost shell to comply with the rule.

Example: Ionic Formula

Find the formula of the ionic compound formed from O and Al.

**Solution:** We first write down Lewis diagrams for each atom involved:

\[ \text{Al:} \quad \text{and} \quad \text{O} \]

We now see that each O atom needs 2 electrons to make up an octet, while each Al atom has 3 electrons to donate. In order that the same number of electrons would be donated as accepted, we need 2 Al atoms ($2 \times 3e^- \text{donated}$) and 3 O atoms ($3 \times 2e^- \text{accepted}$). The whole process is then

\[ \begin{array}{c}
\text{Al:} & \quad \text{O} \quad \rightarrow \quad \text{Al}^{3+} & \quad \left[ \cdot \cdot \cdot \cdot \right]^{2-} \\
+ \quad \text{O} \quad \rightarrow \quad & \quad \left[ \cdot \cdot \cdot \cdot \right]^{2-} \\
\text{Al:} & \quad \text{O} \quad \rightarrow \quad \text{Al}^{3+} & \quad \left[ \cdot \cdot \cdot \cdot \right]^{2-}
\end{array} \]

The resultant oxide consists of aluminum ions, $\text{Al}^{3+}$, and oxide ions, $\text{O}^{2-}$, in the ratio of 2:3. The formula is $\text{Al}_2\text{O}_3$. 

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