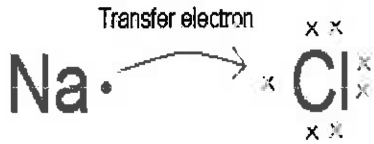
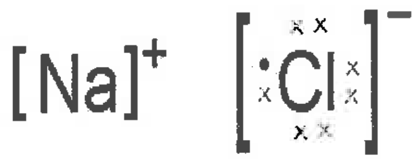
Representing Chemical Bonds Based on Bond Type

Ionic (metal + nonmetal) ($\Delta E > 1.7$) * watch F

Ionic compounds involve electron transfers.



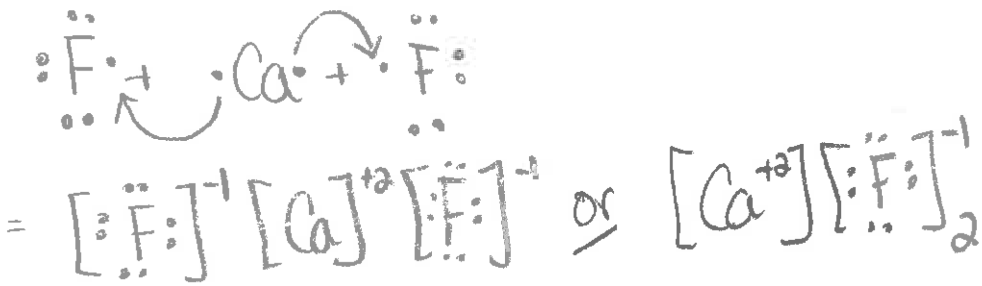
Square brackets [ ] enclose the symbol of the element and the valence electrons, while the charge is shown as a superscript. Subscripts can be included to show how many ions are involved.



Notice the metal has no valence electrons while the non-metal will have achieved a stable octet (period 2 metals will follow duet rule).

happy with 2 e^-

For example: Calcium Fluoride, CaF_2



Covalent (nonmetals) ($\Delta E = 0.5 - 1.7$) * watch F

Given the sharing of molecules, structures can be more difficult to visualize and thus steps for drawing Lewis structures are listed in your textbook on page 197.

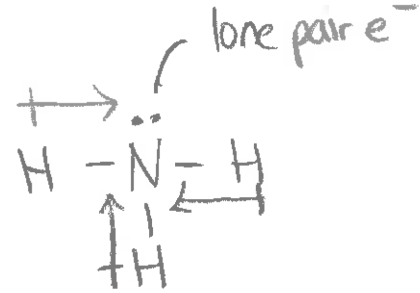
A single bond is shown by a single line and lone pairs are drawn as dots.

The actual shape of the molecule is determined by the number of valence electrons present (but more on that later).

For example: Ammonia, NH_3 $8e^-$

① Highest bonding capacity atom in center
② # of valence e^-
③ add single bonds b/w atoms
④ place remaining e^- as lone pairs
⑤ check octet rule for central atom
- octet/duet for others

⑥ show dipoles if present ($\Delta E = 0.5$–1.7)



For example: **Carbon disulfide**, CS_2 $16e^-$

$\Delta E = 0 \therefore$ no dipole

Don't forget to show the <u>dipole</u> in the case of polar covalent bonding!

Your turn: (a) Na_2O

(<u>ionic</u>

$[Na]^+_2 [:\ddot{O}:]^{-2}$

(b) C_2H_5Cl $20e^-$