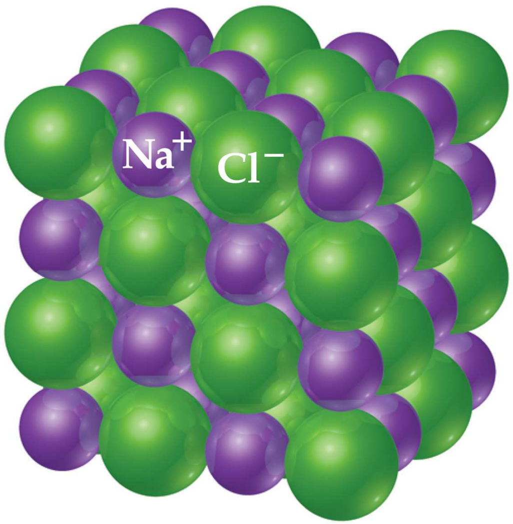
Chapter 8 The Concept of the Chemical Bond

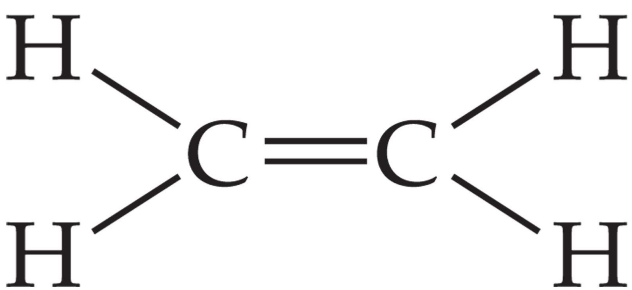
Three basic types of bonds:

Ionic - Electrostatic attraction between ions (NaCl)

Metallic - Metal atoms bonded to each other

Covalent - Sharing of electrons



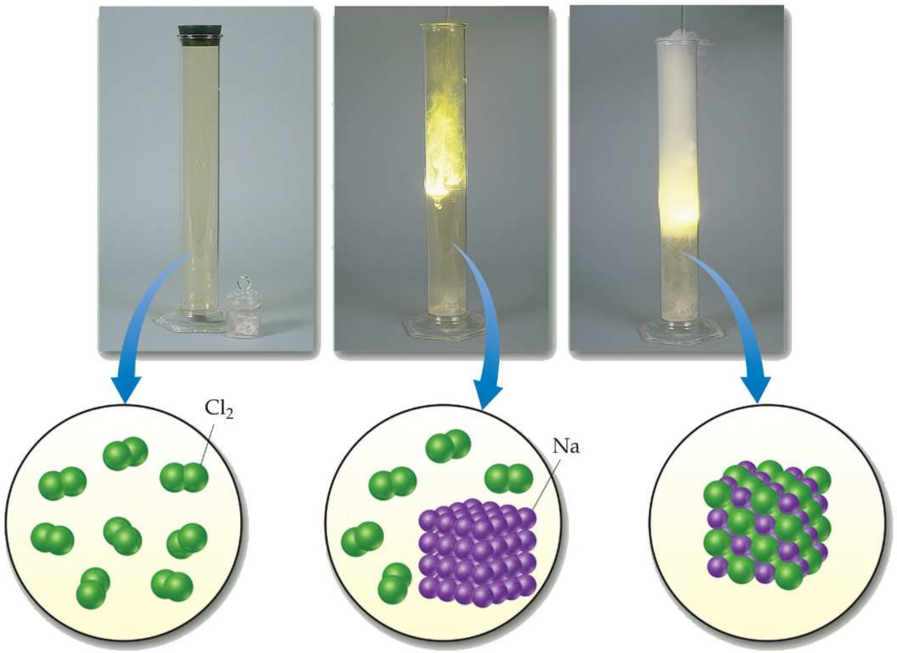


Ionic Bond Energetics

Ionization Energy:
$Na_{(s)} \rightarrow Na^+ + e^-$ $\Delta H = 485$ kJ/mol

Electron Affinity
$Cl + e^- \rightarrow Cl^-$ $\Delta H = -349$ kJ/mol

Lattice Energy
$NaCl \rightarrow Na^+ + Cl^-$ $\Delta H = 788$ kJ/mol



Lattice Energy – Energy gained by forming a lattice

Lattice Energy - Energy required to separate 1 mole of an ionic compound into gaseous ions

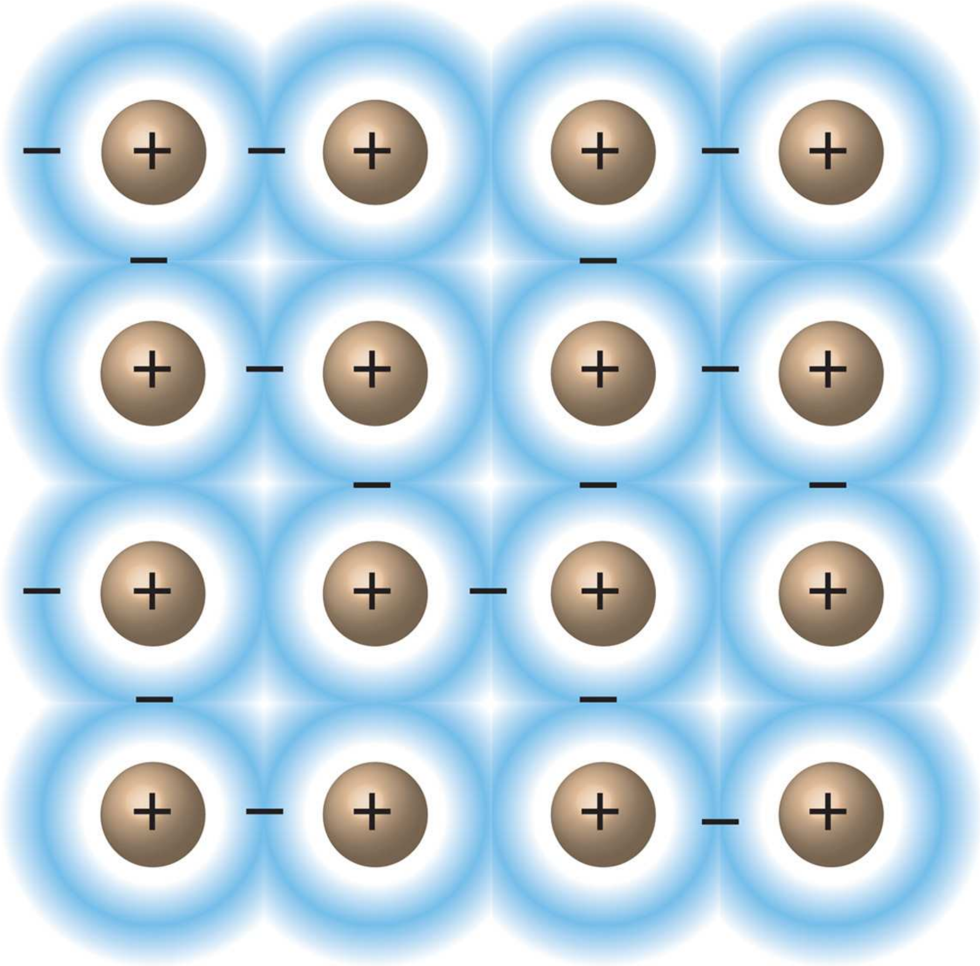
Metallic Bonds

In metals valence electrons are loosely held by every atom

- A 'sea of electrons' can move between each metal atom

- This 'sea of electrons' is responsible for the conduction of heat and electricity in metals

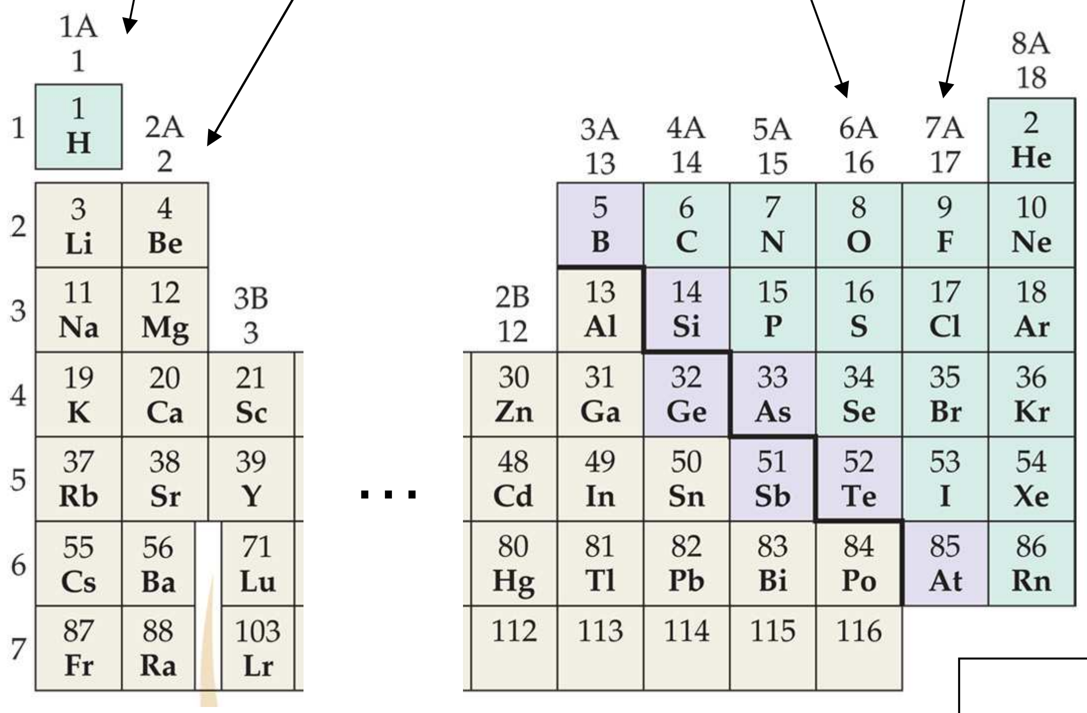
- Different metals have different amount of valence electrons and therefore different properties



How do atoms 'decide' how many electrons to gain / lose? – OCTET RULE

Lose 1 e^- Lose 2 e^-

Gain 2 e^- Gain 1 e^- NO CHANGE



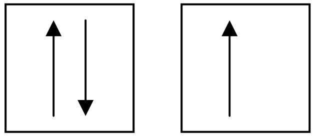
O : [He] $2s^2 2p^4$
F : [He] $2s^2 2p^5$
Ne : [He] $2s^2 2p^6$ → 8 e^- STABLE

Na : [Ne] $3s^1$ Mg : [Ne] $3s^2$
K : [Ar] $4s^1$ Ca : [Ar] $4s^2$

Octet Rule:
8 valence electrons = stable

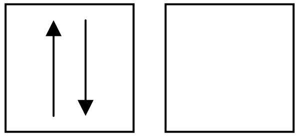
Full Valence Orbitals and Half-full Valence Orbitals = STABLE

Li : $1s^2\ 2s^1$



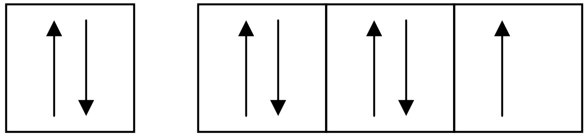
$1s^2$ $2s^1$

Li^+ : $1s^2$



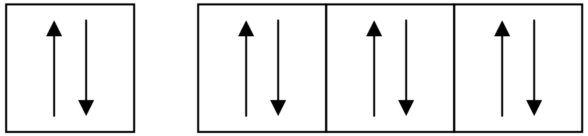
$1s^2$

Cl : [Ne] $2s^2\ 2p^5$



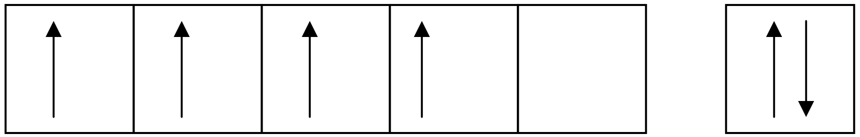
$2s^2$ $2p^5$

Cl^- : [Ne] $2s^2\ 2p^6$



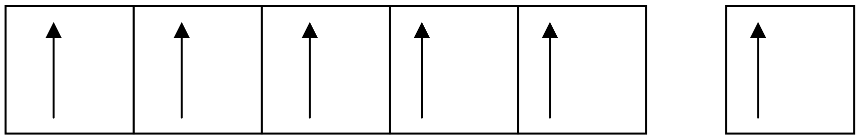
$2s^2$ $2p^6$

Cr : [Ar] $3d^4\ 4s^2$

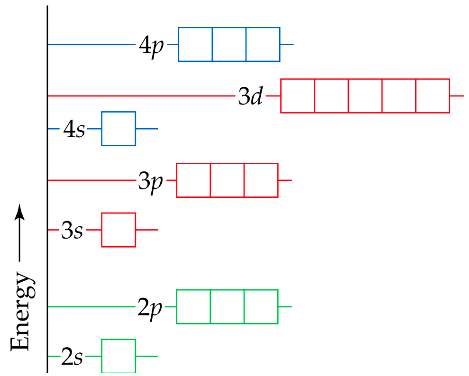


$3d^4$ $4s^2$

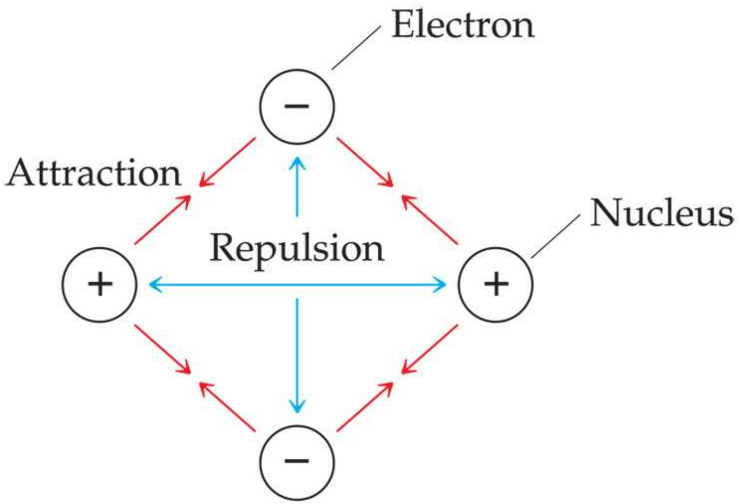
Cr : [Ar] $3d^5\ 4s^1$



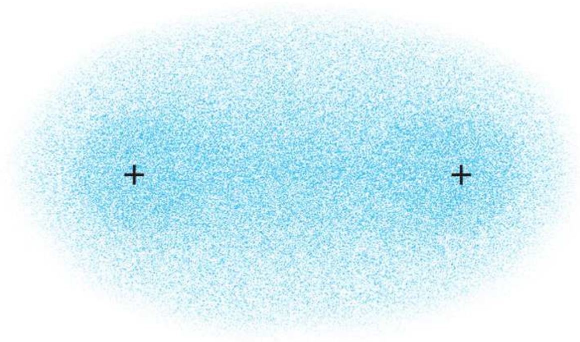
$3d^5$ $4s^1$



Covalent Bond – Atoms ‘glued’ together with electrons



Atoms share valence electrons to complete their octet



Repulsions between nuclei

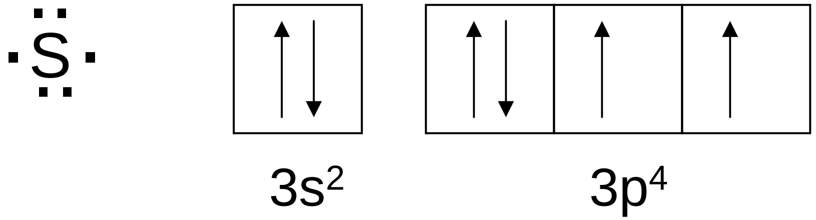
Repulsions between electrons

Attractions between electrons and nuclei

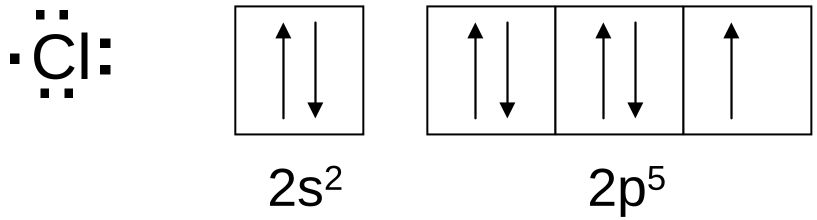
Lewis Symbols

- Show valence electrons and enable predictions about bond formation

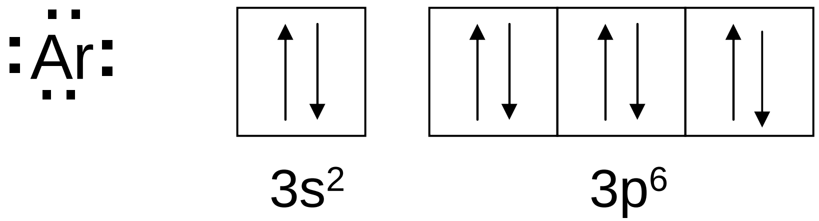
Sulfur : [Ne] $3s^2$ $3p^4$



Chlorine : [Ne] $2s^2$ $2p^5$

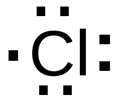


Argon : [Ne] $3s^2$ $3p^6$

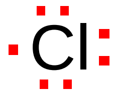


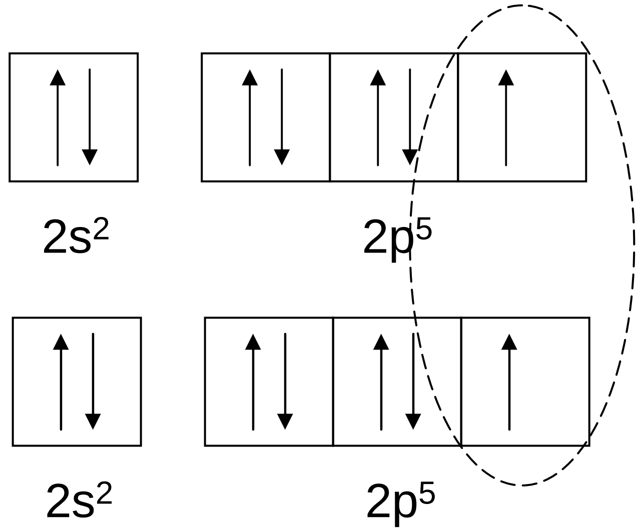
Covalent Bond – Lewis Symbols

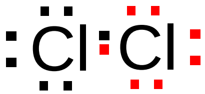
Chlorine : [Ne] $2s^2$ $2p^5$

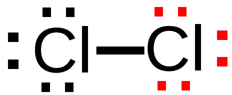


Chlorine : [Ne] $2s^2$ $2p^5$









Draw the Lewis Symbol for HBr, H_2O, NH_3, and CH_4 :

Multiple Bonds

O_2

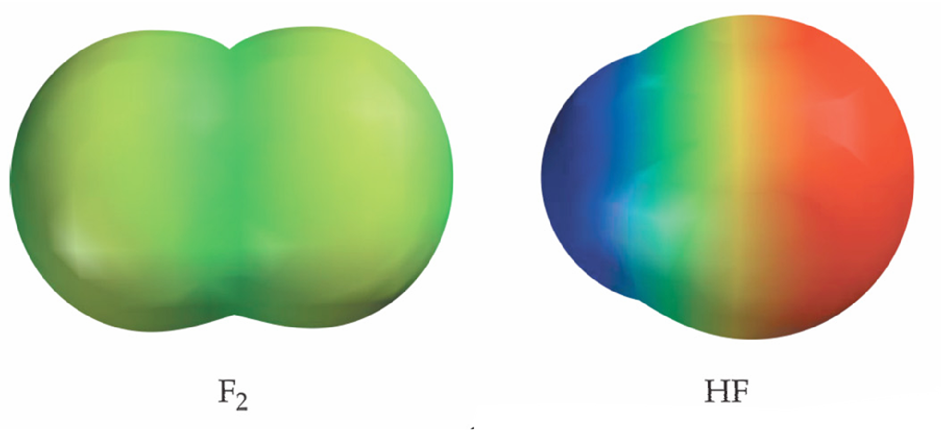
Double Bond

N_2

Triple Bond

Bond length: triple < double < single

Bond Polarity



Although atoms often form compounds by sharing electrons, the electrons are not always shared equally.

Fluorine pulls harder on the electrons it shares with hydrogen than hydrogen does.
Therefore, the fluorine end of the molecule has more electron density than the hydrogen end.

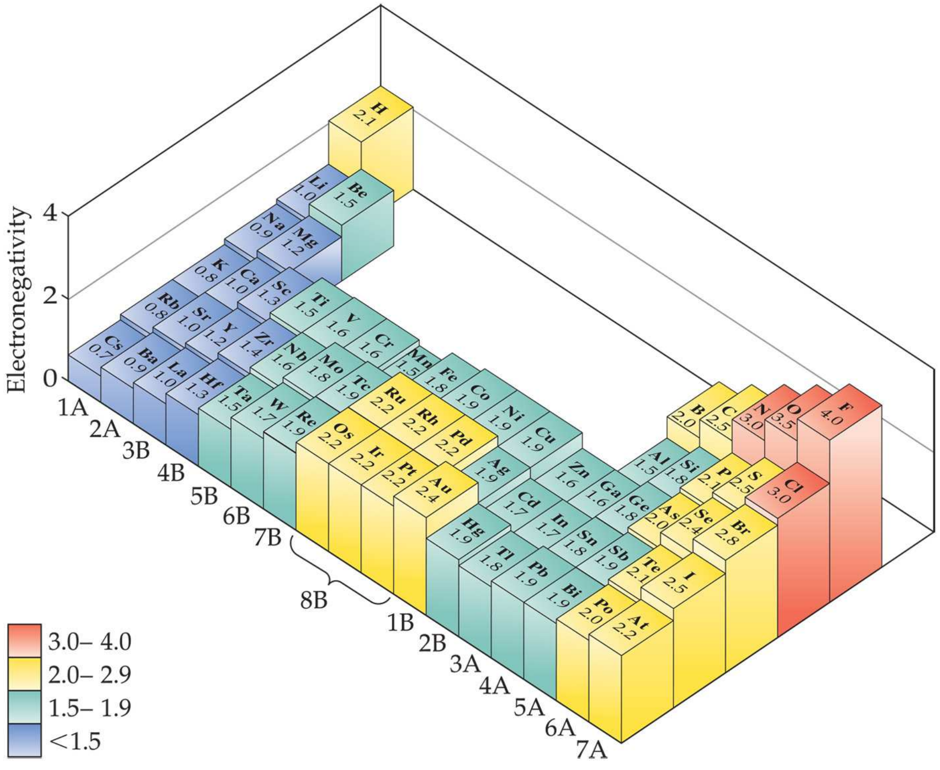
When two atoms share electrons unequally, a bond dipole results

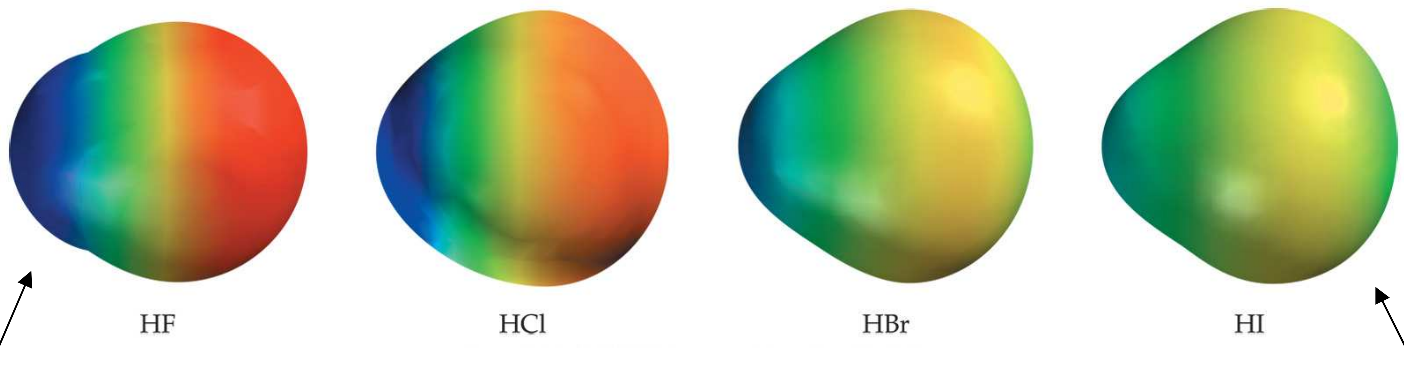
Bond Polarity

The ability of atoms in a molecule to attract electrons to itself.

On the periodic chart, electronegativity increases

- from left to right across a row.
    - from the bottom to the top of a column.





More Polar

Less Polar

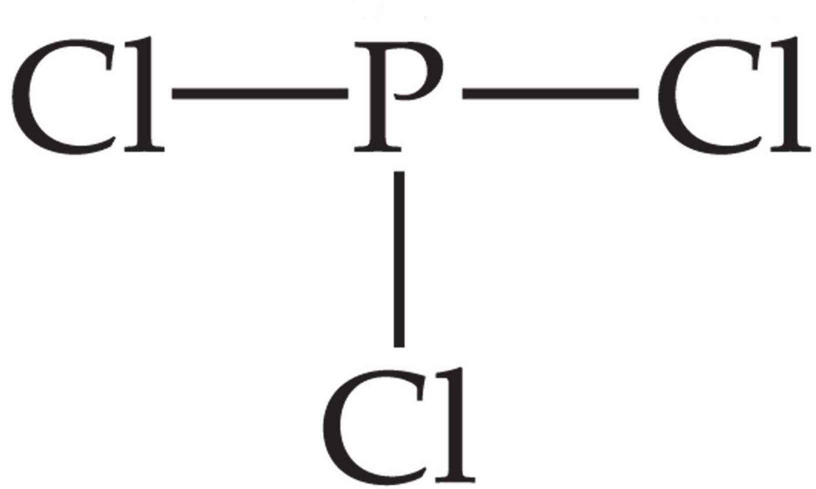
Writing Lewis Structures

PCl_3

$5 + 3(7) = 26$

1. Find the sum of valence electrons of all atoms in the polyatomic ion or molecule.
   - If it is an anion, add one electron for each negative charge.
   - If it is a cation, subtract one electron for each positive charge.

Writing Lewis Structures

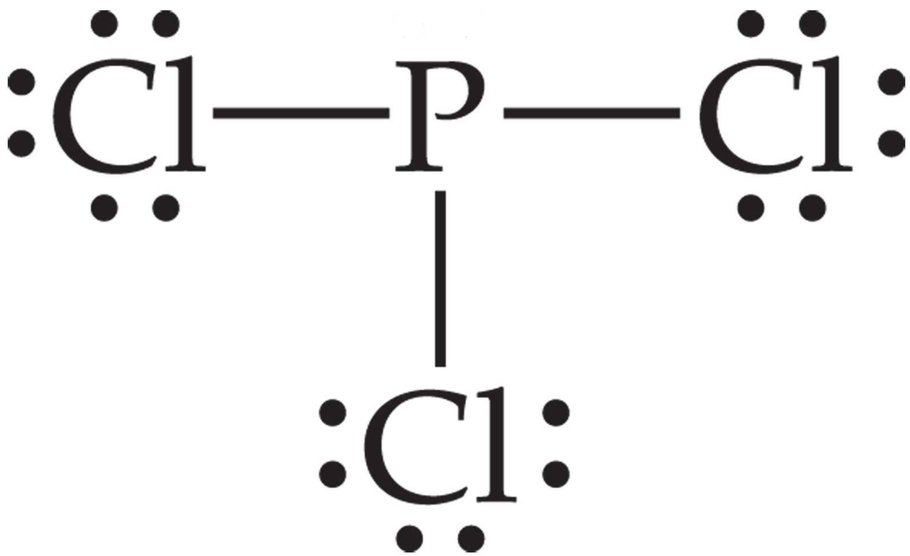


2. The central atom is the *least* electronegative element that isn't hydrogen. Connect the outer atoms to it by single bonds.

Keep track of the electrons:

26 – 6 = 20

Writing Lewis Structures

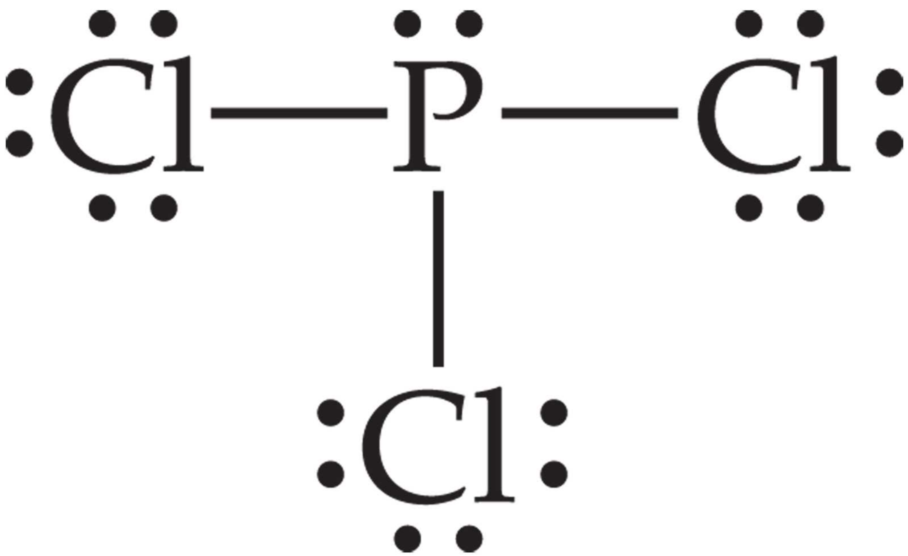


3. Fill the octets of the outer atoms.

Keep track of the electrons:

26 – 6 = 20 – 18 = 2

Writing Lewis Structures



4. Fill the octet of the central atom.

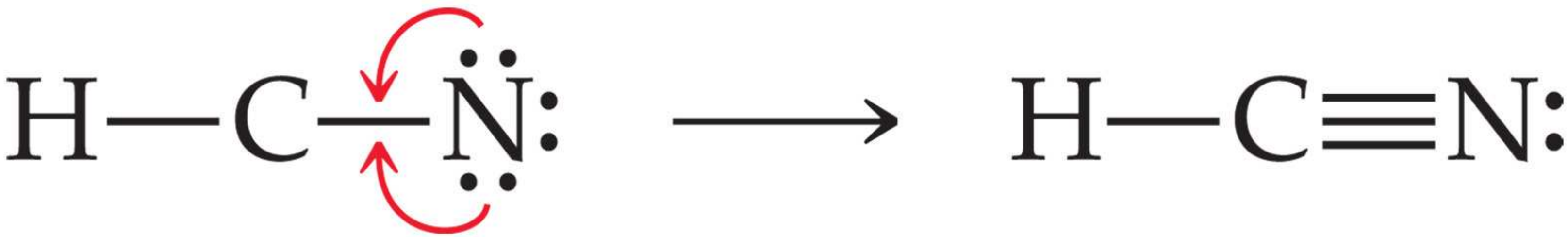
Keep track of the electrons:

26 – 6 = 20 – 18 = 2 – 2 = 0

Writing Lewis Structures

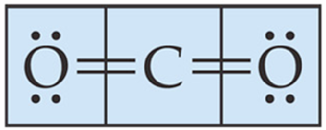
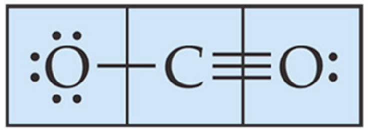
5. If you run out of electrons before the central atom has an octet…

…form multiple bonds until it does.



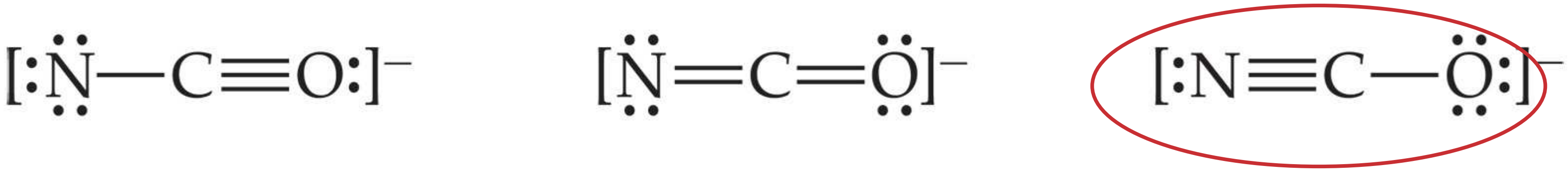
Writing Lewis Structures

- Then assign formal charges.
  - For each atom, count the electrons in lone pairs and half the electrons it shares with other atoms.
  - Subtract that from the number of valence electrons for that atom: The difference is its formal charge.

						
Valence electrons:	6	4	6	6	4	6
−(Electrons assigned to atom):	6	4	6	7	4	5
Formal charge:	0	0	0	−1	0	+1

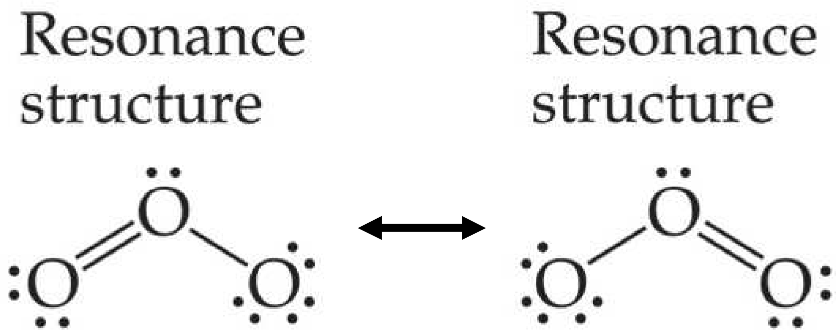
Writing Lewis Structures

- The best Lewis structure…
  - …is the one with the fewest charges.
  - …puts a negative charge on the most electronegative atom.



Resonance

This is the Lewis structure we would draw for ozone, O_3.

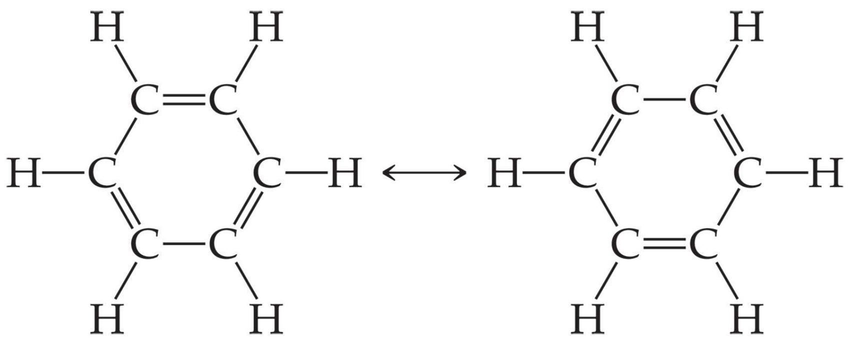


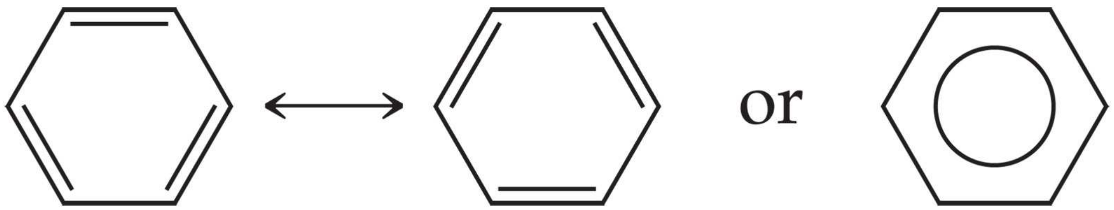
One Lewis structure cannot accurately depict a molecule such as ozone.

We use multiple structures, resonance structures, to describe the molecule.

Resonance Example

Benzene





Exceptions to the Octet Rule

There are three types of ions or molecules that do not follow the octet rule:

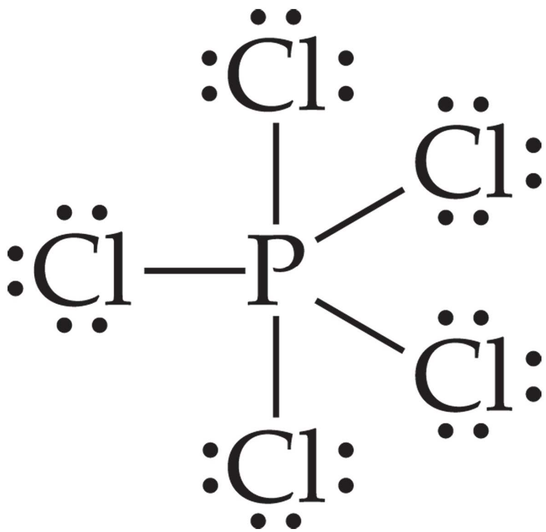
1) Ions or molecules with an odd number of electrons. ← rare

2) Ions or molecules with less than an octet. ← B or Be

3) Ions or molecules with more than eight valence electrons (an expanded octet). ← Common S, P, Cl

Examples of an Expaned Octet (>8 e$^-$)

PCl_5



PO_4^{-3}