

Electron Affinity

Opposite: measures $X(g) + e^- \rightarrow X^-(g)$

Trend as you would expect: Cl likes gaining an electron more than Na. see Fig. 8.33



Chapter 9: Chemical Bonding

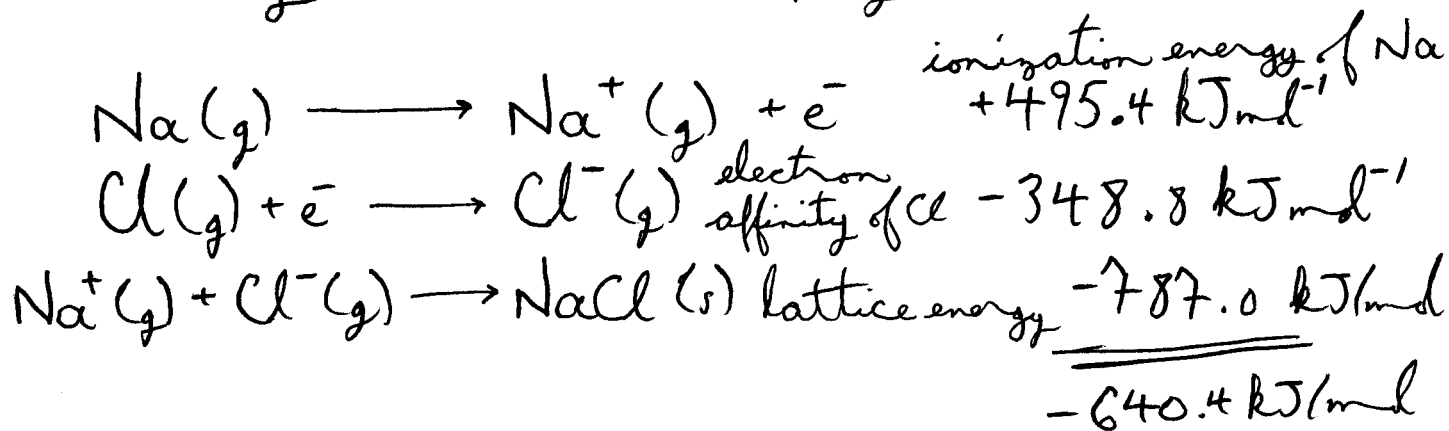
We will look at ionic bonding (salts)
and covalent bonding (molecules)

Ionic compounds: form by a transfer of electrons
the resulting electrostatic attraction between
anions and cations is called an ionic bond.

The fundamental reason some elements combine to
form ionic compounds is that this arrangement lowers
their potential energy the most.

(the process is exothermic \rightarrow gives off heat)

Try to do the bookkeeping:

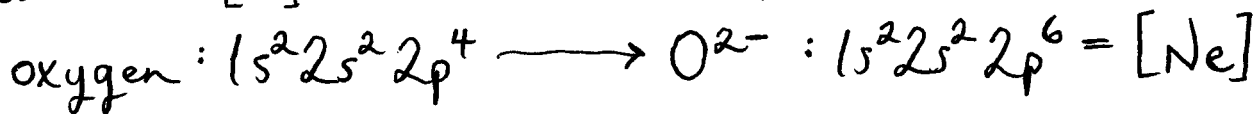
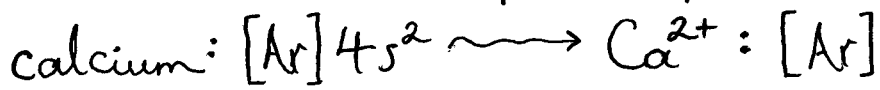
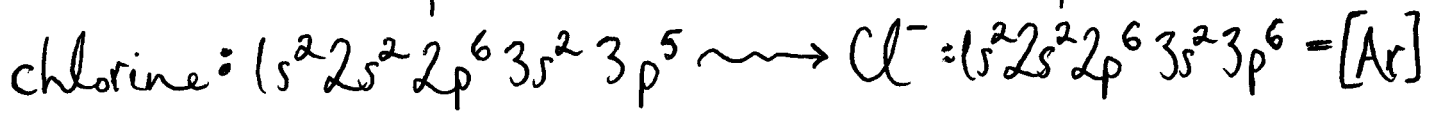
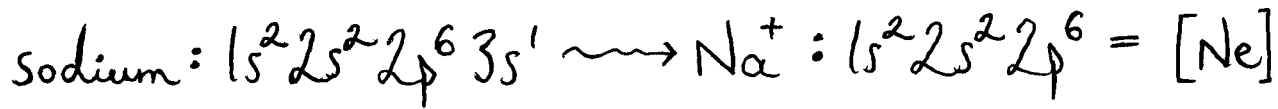


exothermic, potential energy drops

The lattice energy comes from electrostatics. We know from page ③ of my notes that the potential energy between two charges q_1 , q_2 separated by a distance r is proportional to $\frac{q_1 q_2}{r}$. We have to add up all the electrostatic interaction in the salt crystal to get the lattice energy, but very roughly we would expect the lattice energy to be larger for ions with bigger charges or ions that are closer together.

Sodium is a solid, not a gas, at room temp. & pressure, so our bookkeeping is not correct: see page 356.

Electronic configuration in salts: ions tend to acquire a noble gas config.



octet rule: atoms like to have 8 outer shell electrons

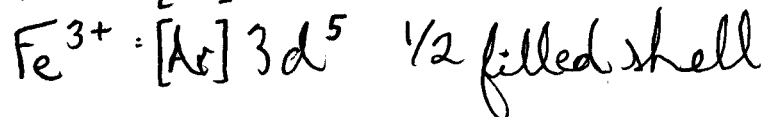
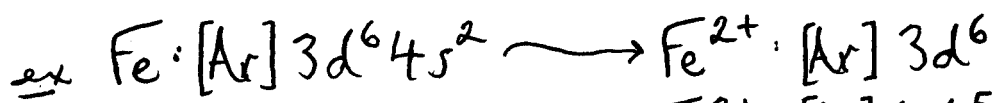
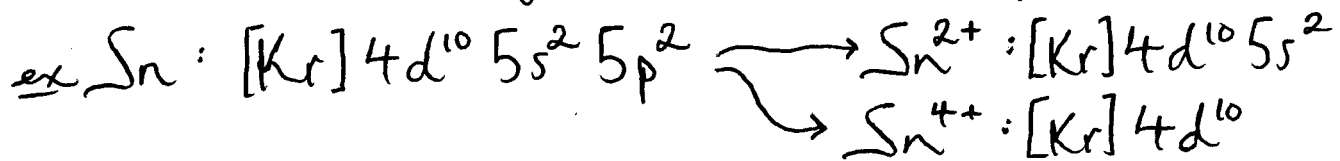
Exceptions

- atoms that form the [He] core ex Li, Be.

- transition + post-transition metals

a) electrons removed from outer shell (largest n)

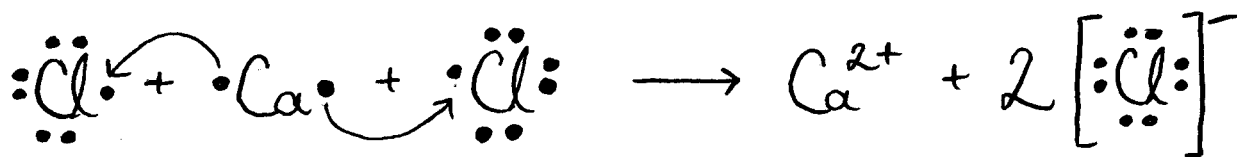
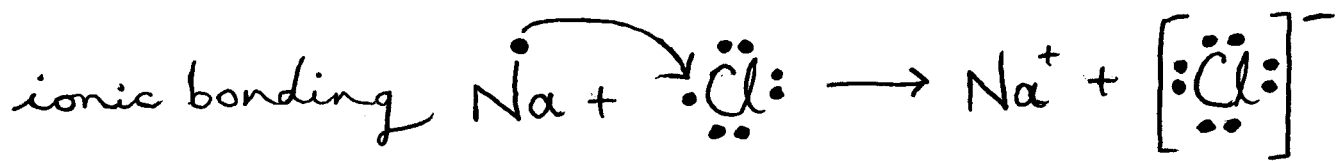
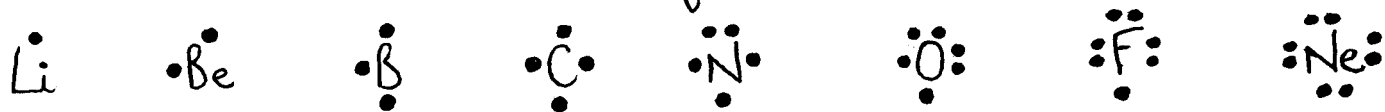
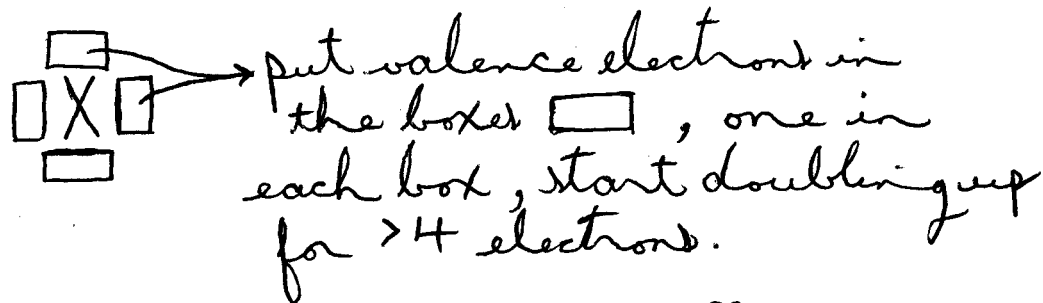
b) for same n , removed from f , then d , then p , s .



+2 cations common for transition metals.

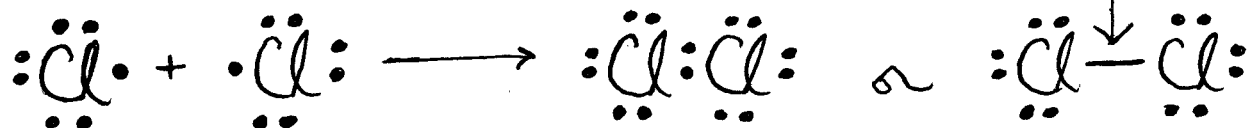
Lewis electron-dot structures

- keeps track of valence electrons (electrons outside largest noble gas core)
- useful for bonding structure.

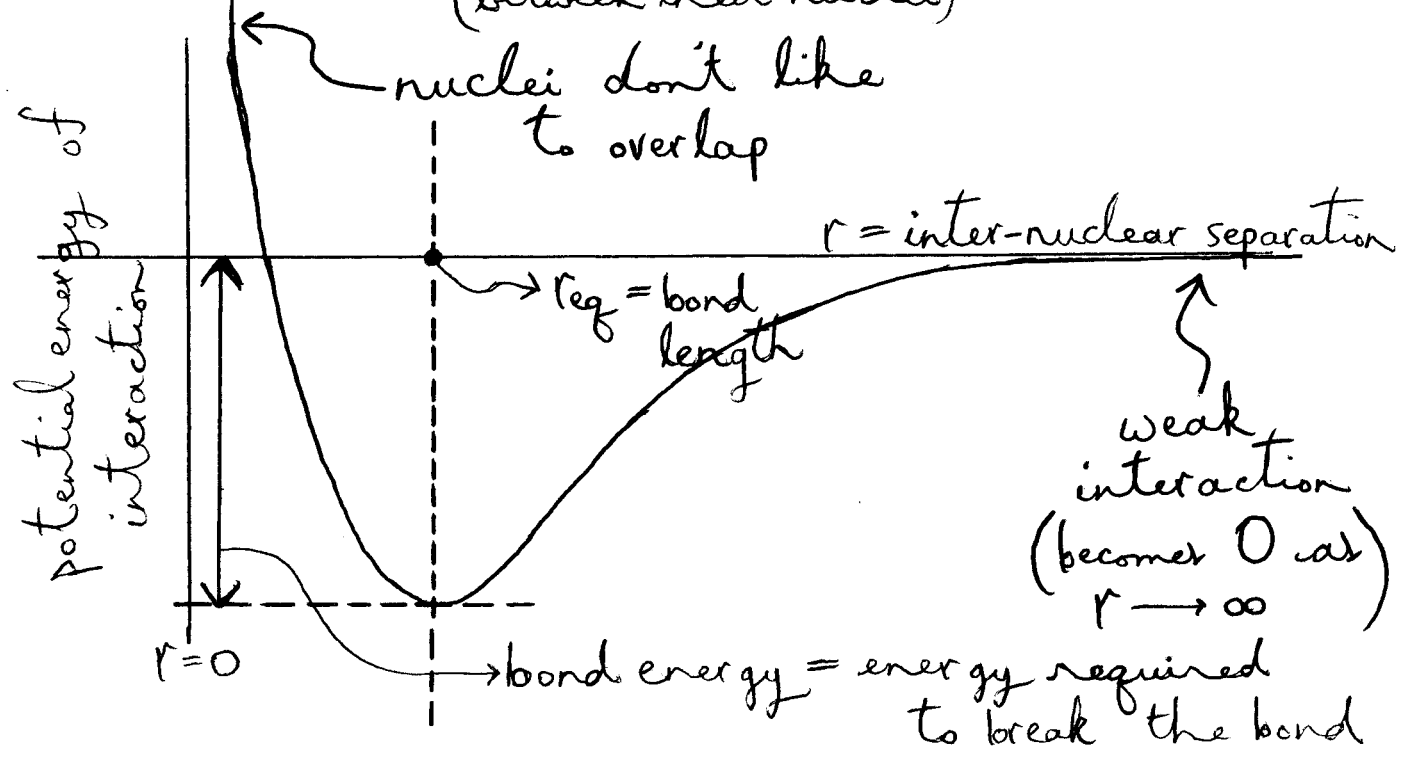


ionic bonding \rightarrow transfer of electrons

covalent bonding \rightarrow sharing of electrons



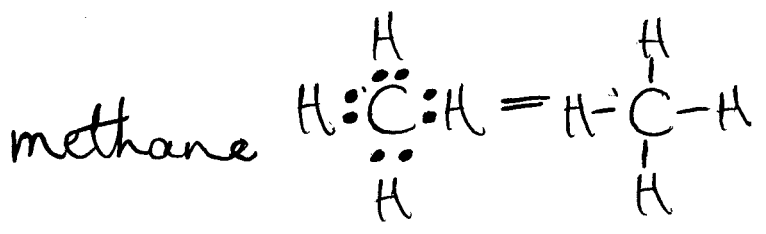
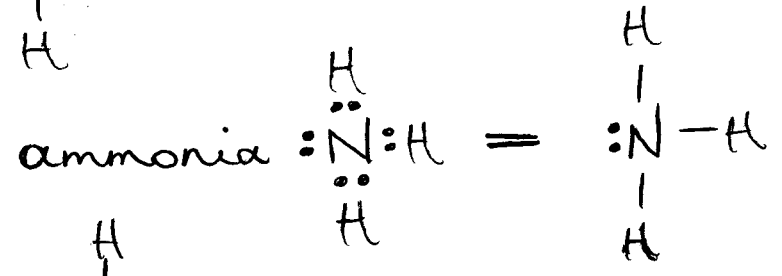
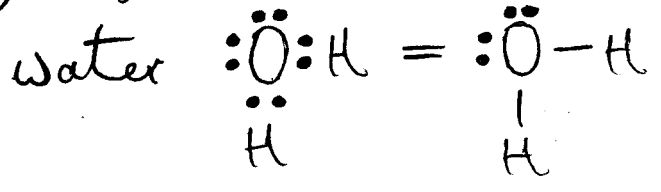
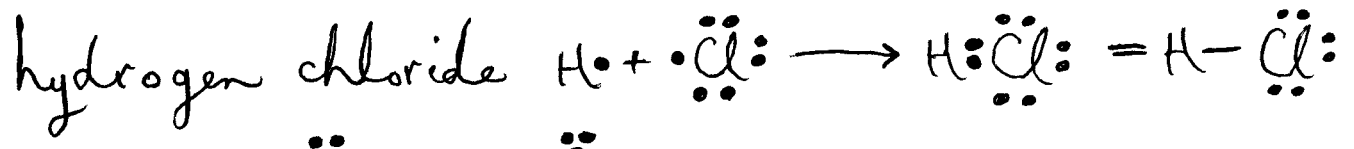
Look at the potential energy between 2 Cl atoms as a function of the distance between them. (between their nuclei)



The decrease of potential when two Cl atoms come together leads to a more stable situation. (again assuming the excess kinetic energy is dissipated into the environment.)

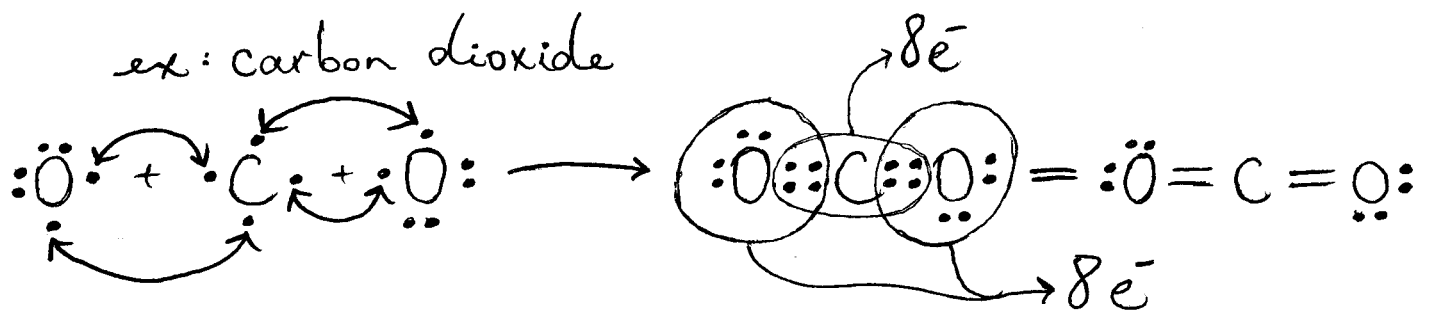
Whether a covalent or an ionic bond forms depends on which one leads to the most stable situation -- which one lowers the potential energy more

Look at single bonds : involves 2 electrons

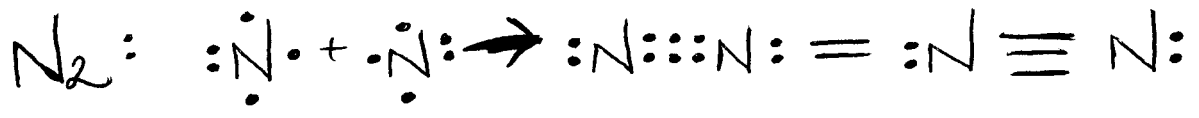


Double bond : involves 4 electrons

ex: carbon dioxide



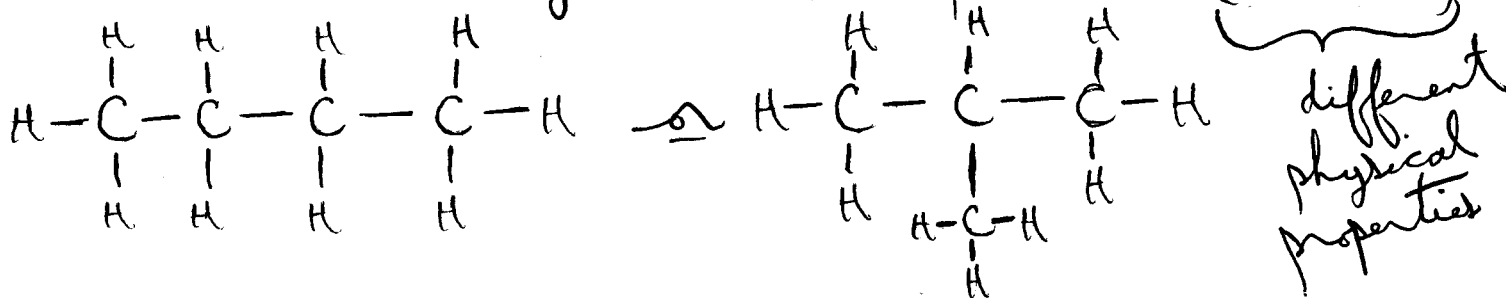
Triple bond : involves 6 electrons



Organic compounds involve all kinds of covalent bonds with C, H, O, P, ..., N, ...

You will study this next year if you take organic chemistry, so we only take a brief look at these compounds now.

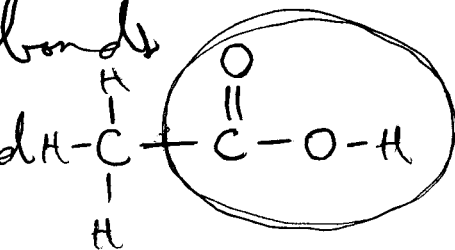
C₄H₁₀ : all single bonds, two possibilities (2 isomers)



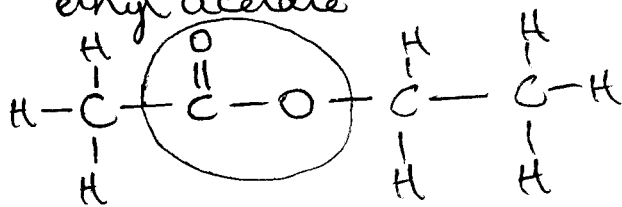
oxygen can have double or single bonds

ex carboxylic acids

acetic acid (vinegar)



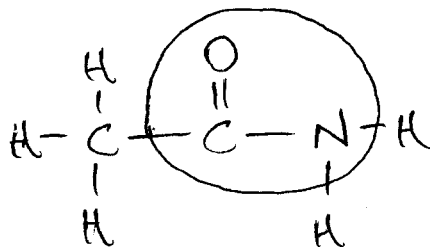
ex esters ethyl acetate



fruity smell: used as artificial flavors and in perfumes

nitrogen is common as well:

ex amides acetamide



Electronegativity

In HCl, the chlorine nucleus attracts the electrons in the bond more strongly than the H nucleus

We indicate this by writing $\text{H} - \ddot{\text{Cl}}:$
 $\delta^+ \quad \delta^-$

The delta symbol δ indicates a partial charge (a complete transfer of charge would be an ionic bond).

The result is a polar bond with a dipole moment μ

$$\mu = q \cdot r, \quad q = |\delta^+| = |\delta^-|, \quad r = \text{distance between atoms} = r_{\text{eq.}}$$

Why is this important? Affects solubility and spectroscopy among other things. ex polar molecules tend to be soluble in polar solvents, while non-polar molecules tend to be soluble in non-polar solvents.

Electronegativity measures how strongly an atom competes for the electrons in a chemical bond.

A bond between atoms of equal E.N. is purely covalent

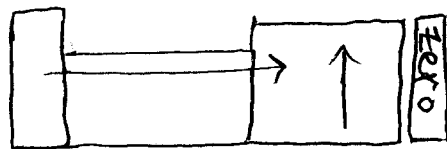
A bond between atoms of very different E.N. is ionic because one atom grabs the electrons for itself.

A bond between atoms of somewhat different E.N. is polar. (covalent).

See Fig 9.6

Trend in E.N. similar to I.E. :

gets stronger as we go from left to right, and bottom to top, of periodic table. (Fig 9.5)

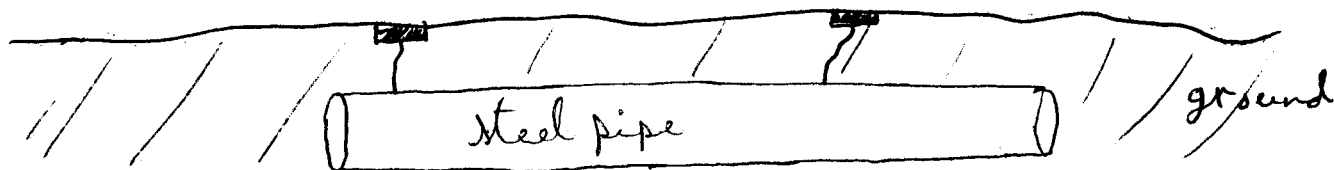


Electronegativity can be related to reactivity.

A metal is reactive if it is easily oxidized (loses e^-) which roughly reflects how low its electronegativity is.

see Fig 9.7. Gold, silver, and copper are not very reactive whereas magnesium is highly reactive.

Industrial importance : magnesium or zinc metal is attached at regular intervals with a wire to underground steel pipes -- prevents corrosion of the pipe, easy to replace.



A non-metal is reactive if it is easily reduced (gains e^-)
→ high electronegativity.

To get reduced, the nonmetal must oxidize something.
So we call F or O oxidizing agents.

Lewis Structures

Simple and gives a lot of chemical insight.

Basic rules (but we will discuss many exceptions)

1.) locate the central atom & decide where the bonds are.

CHAPTER 2 IS USEFUL HERE! + formula

ex H_2SO_4

ex H_3PO_4

ex HNO_3

2.) count all the valence electrons

3.) make all bonds single (put 2 electrons in place)

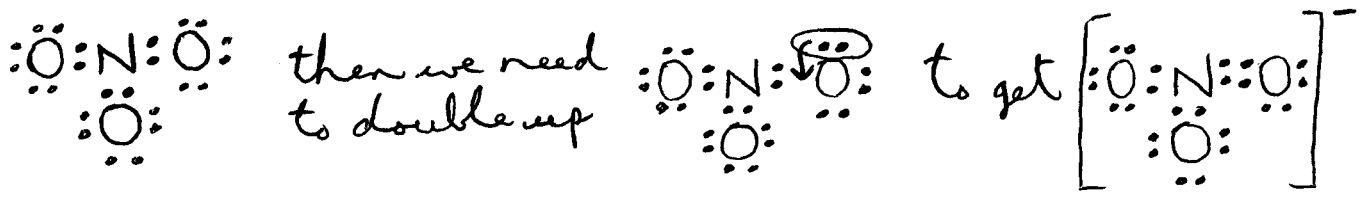
4.) complete the octets of the atoms attached to the central atom by adding pairs of electrons.

5.) place remaining electrons on the central atom

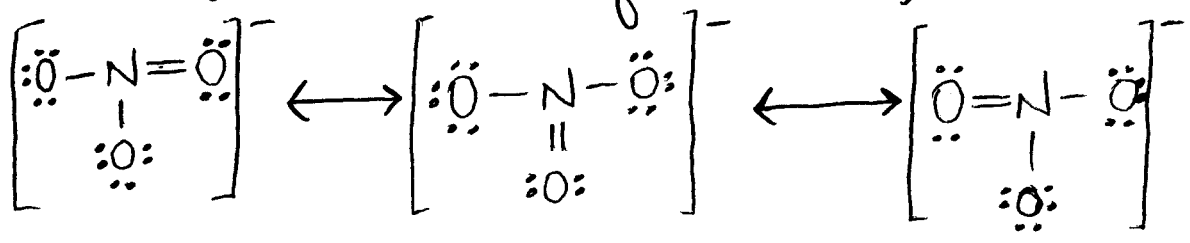
6.) if central atom does not have an octet, form double or triple bonds.

ex NO_3^- nitrate ion

N is the central atom. $5 + 6 + 6 + 6 + 1 = 24$ valence e⁻



but we had a choice, so the correct answer is actually a combination of the choices, called a resonance structure



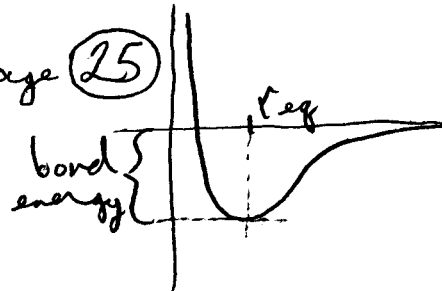
The "extra" bond is shared between the O atoms, roughly $\left[\begin{array}{c} \text{O}=\text{N}=\text{O} \\ \text{||} \\ \text{O} \end{array} \right]^-$

These rules seem arbitrary. What is the link to reality? What experimental data is there?

bond lengths (r_{eq}) and bond energies -- see page (25)

data shows that as the

bond order increases, the bond between two elements strengthens and shortens

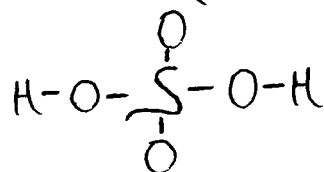


Bond order:

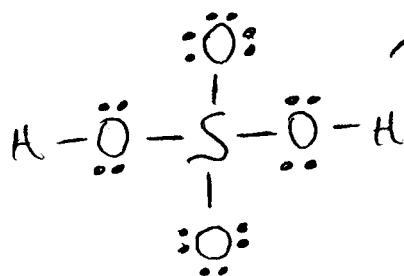
- 1 for a single bond
- 2 for a double bond
- 3 for a triple bond
- 1.33 for the N-O bond in NO_3^-

This kind of data leads to an important rule for Lewis structures. We will introduce this rule to correct a problem with the structure of H_2SO_4 , sulfuric acid.

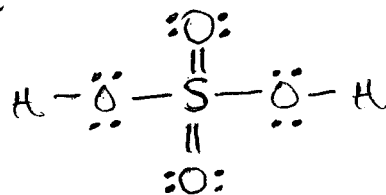
Let's follow our six rules: S is obviously the central atom, and the hydrogen atoms are bonded to oxygen. (we don't see this until we do ch. 5)



12 e^- in the six single bonds, leaves 20 valence e^- these go perfectly on the oxygens to give



But (see Fig 9.9) the S-O bonds are significantly shorter than the S-OH bonds so we draw

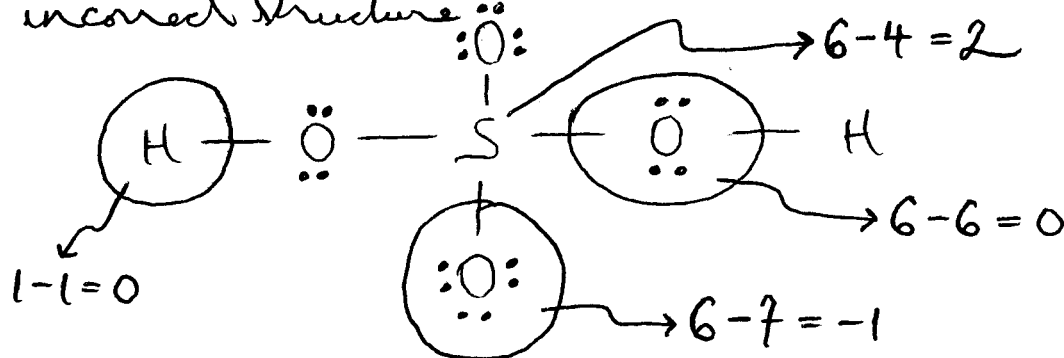


We add a rule to fix things.

- Choose the Lewis structure with the smallest / fewest formal charges.

$$\text{formal charge of an atom} = \left(\begin{array}{l} \# \text{ of valence } e^- \\ \text{in isolated atom} \end{array} \right) - \left[\begin{array}{l} \# e^- \text{ in the Lewis} \\ \text{structure (bonding} \\ e^- \text{ only count } \underline{\text{half}}) \end{array} \right]$$

Analyze incorrect structure

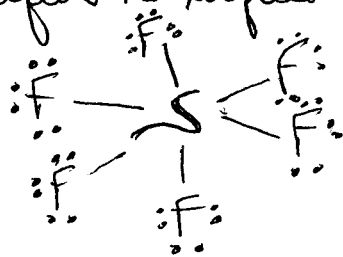


The correct structure assigns every atom a formal charge of 0.

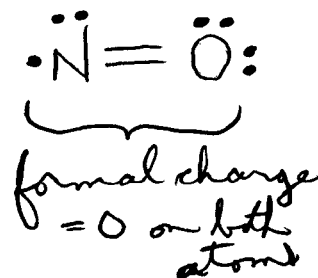
≠ We will finish the chapter now with a few more special cases.

- Sulfur & phosphorus can expand their valence shells by using 3d orbitals.

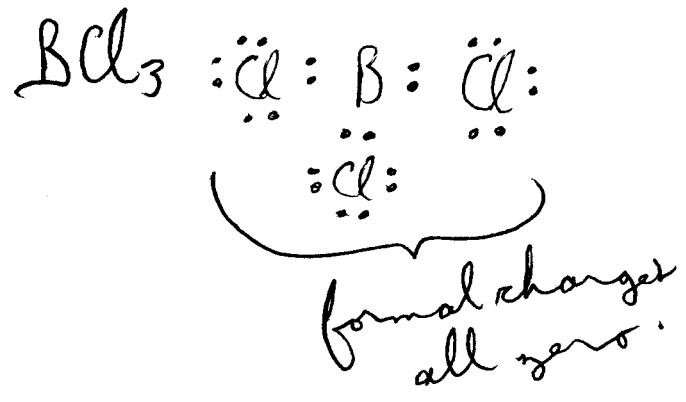
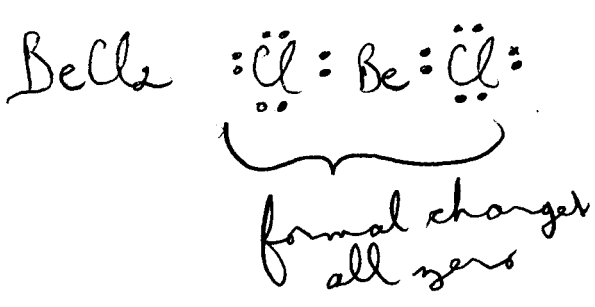
ex SF_6 sulfur hexafluoride $6 + 7 \cdot 6 = 48$ valence e^-



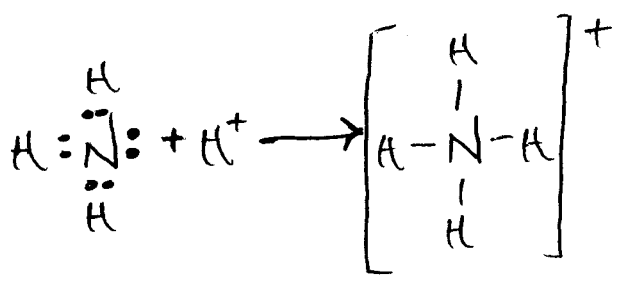
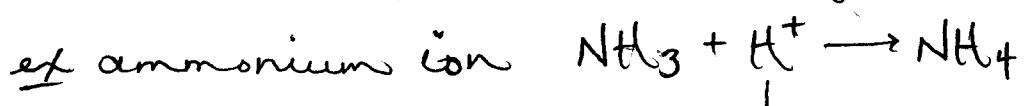
- odd # of valence e^- : NO , NO_2 , ClO_2



- B and Be compounds are electron deficient.

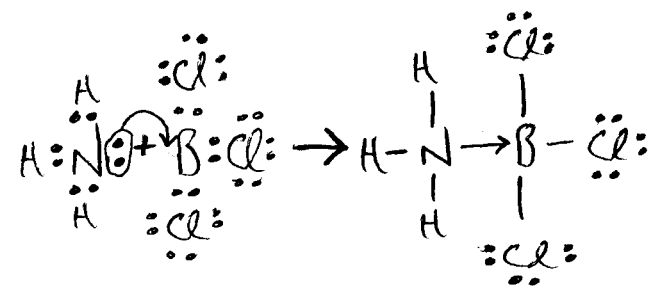


- Coordinate covalent bonds
 both bonding electrons come from the same atom

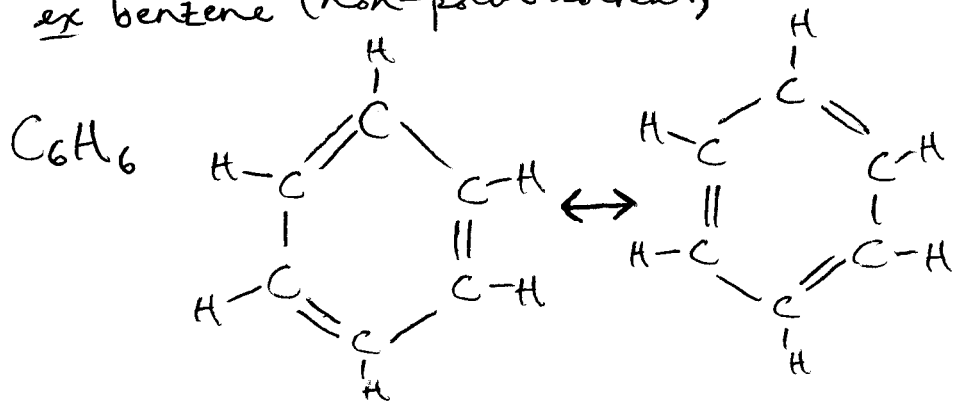


no electrons to contribute!

indicate this features with an arrow



- resonance structures are more stable than one of the choices would suggest
- ex benzene (non-polar solvent)



draw as



to indicate partial double bond character