

# Electronegativity and Bonding

The modern definition of electronegativity is due to Linus Pauling. It is:

The power of an atom in a molecule to attract electrons to itself.

Pauling was able to develop a numerical scale of electronegativities. Below are ten common elements with their values. Your textbook should have a more complete list.

F	4.0	C	2.5
O	3.5	S	2.5
Cl	3.0	H	2.1
N	3.0	Na	0.9
Br	2.8	K	0.8

When you examine a periodic table, you will find that (excluding the noble gases) the electronegativity values tend to increase as you go to the right and up. The reverse statement is that the values tend to decrease going down and to the left. This pattern will help when you are asked to put several bonds in order from most to least ionic without using the values themselves.

Electronegativity values are useful in determining if a bond is to be classified as nonpolar covalent, polar covalent or ionic.

What you should do is look only at the two atoms in a given bond. Calculate the difference between their electronegativity values. Only the absolute difference is important.

**I. Nonpolar Covalent:** This type of bond occurs when there is **equal** sharing (between the two atoms) of the electrons in the bond. Molecules such as  $\text{Cl}_2$ ,  $\text{H}_2$  and  $\text{F}_2$  are the usual examples.

Textbooks typically use a maximum difference of 0.2 - 0.5 to indicate nonpolar covalent. Since textbooks vary, make sure to check with your teacher for the value he/she wants. The ChemTeam will use 0.5.

One interesting example molecule is  $\text{CS}_2$ . This molecule has nonpolar bonds. Sometimes a teacher will only use diatomics as examples in lecture and then spring  $\text{CS}_2$  as a test question. Since the electronegativities of C and S are both 2.5, you have a nonpolar bond.

**II. Polar Covalent:** This type of bond occurs when there is **unequal** sharing (between the two atoms) of the electrons in the bond. Molecules such as  $\text{NH}_3$  and  $\text{H}_2\text{O}$  are the usual examples.

The typical rule is that bonds with an electronegativity difference less than 1.6 are considered polar. (Some textbooks or web sites use 1.7.) Obviously there is a wide range in bond polarity, with the difference in a C-Cl bond being 0.5 -- considered just barely polar -- to the difference the H-O bonds in water being 1.4 and in H-F the difference is 1.9. This last example is about as polar as a bond can get.

**III. Ionic:** This type of bond occurs when there is **complete** transfer (between the two atoms) of the electrons in the bond. Substances such as  $\text{NaCl}$  and  $\text{MgCl}_2$  are the usual examples.

The rule is that when the electronegativity difference is greater than 2.0, the bond is considered ionic.

So, let's review the rules:

1. If the electronegativity difference (usually called  $\Delta EN$ ) is less than 0.5, then the bond is nonpolar covalent.
2. If the  $\Delta EN$  is between 0.5 and 1.6, the bond is considered polar covalent
3. If the  $\Delta EN$  is greater than 2.0, then the bond is ionic.

That, of course, leaves us with a problem. What about the gap between 1.6 and 2.0? So, rule #4 is:

4. If the  $\Delta EN$  is between 1.6 and 2.0 and if a metal is involved, then the bond is considered ionic. If only nonmetals are involved, the bond is considered polar covalent.

Here is an example: Sodium bromide (formula = NaBr;  $EN_{Na} = 0.9$ ,  $EN_{Br} = 2.8$ ) has a  $\Delta EN = 1.9$ . Hydrogen fluoride (formula = HF;  $EN_H = 2.1$ ,  $EN_F = 4.0$ ) has the same  $\Delta EN$ . We use rule #4 to decide that NaBr has ionic bonds and that HF has a polar covalent bond in each HF molecule.

As you might expect, NaBr and HF are very different substances. NaBr exhibits the classic "[lattice structure](#)" of ionic substances whereas [HF](#) is a gas a room temperature.

**Periodic Table of the Elements**  
Electronegativity

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About Chemistry

1A	2A		3A-8B										9A																						
1 H 2.20	2 He no data												10 Ne no data																						
3 Li 0.98	4 Be 1.57	1.0 1.3 1.6 1.9 2.2 2.5 2.8 3.1										5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	11 Na 0.93	12 Mg 1.31	13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar no data											
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr 3.00	37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe 2.6
55 Cs 0.79	56 Ba 0.89	57-71 Lanthanides	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.20	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.62	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2	86 Rn no data	87 Fr 0.7	88 Ra 0.89	89-103 Actinides															

\*\*\* Elements > 104 exist only for very short half-lives and the data is unknown.\*\*\*

Lanthanides	57 La 1.10	58 Ce 1.12	59 Pr 1.13	60 Nd 1.14	61 Pm 1.13	62 Sm 1.17	63 Eu 1.2	64 Gd 1.2	65 Tb 1.2	66 Dy 1.22	67 Ho 1.23	68 Er 1.24	69 Tm 1.25	70 Yb 1.1	71 Lu 1.27
Actinides	89 Ac 1.1	90 Th 1.3	91 Pa 1.5	92 U 1.38	93 Np 1.36	94 Pu 1.28	95 Am 1.3	96 Cm 1.3	97 Bk 1.3	98 Cf 1.3	99 Es 1.3	100 Fm 1.3	101 Md 1.3	102 No 1.3	103 Lr no data