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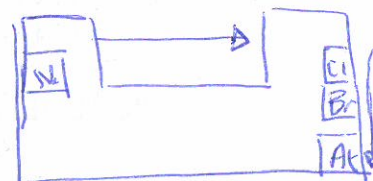
Chemistry 11 ELECTRONEGATIVITY + COVALENT BONDING

ELECTRONEGATIVITY- the tendency of an atom to attract electrons from a neighbouring atom.

The following exercises are intended to help you get a better understanding of what trends in electronegativity occur in the periodic table:

Example 1. Use the atoms Na and Cl for the following questions:

- Which atom has a larger atomic radius? Na
- Which atom has the stronger attraction to the outer electrons on a neighbouring atom, based only on the atomic radius? Cl
- Which atom has the greater nuclear charge? Cl
- Which atom can attract electrons from an adjacent atom most strongly, based on both size and nuclear charge? Cl
- Summarize the above by completing this statement:
 IN GENERAL, when going from LEFT to RIGHT across the periodic table the electronegativity of the atoms will INCREASE.



Example 2. Compare Br and At

- Which atom has a larger atomic radius? At
- Which atom has a stronger attraction to the outer electrons of another atom? Br
- Summarize the above by filling in the statement below:
 IN GENERAL, when going DOWN a family of the periodic table, the electronegativity of the atoms will DECREASES.

Example 3.

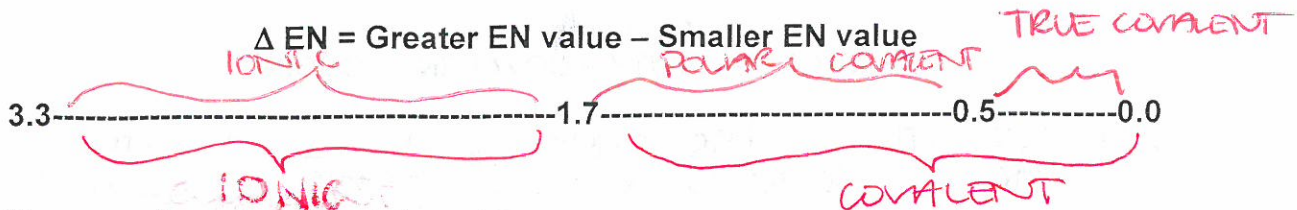
- Ignoring the NOBLE GASES, which atom is the most ELECTRONEGATIVE? Fluorine
- Ignoring the NOBLE GASES, which atom is the least ELECTRONEGATIVE? Francium
- Which is more electronegative: Cs or Be?
- Which is more electronegative: Sn or F?

TABLE OF ELECTRONEGATIVITIES:

H 2.2											B 2.0	C 2.5	N 3.0	O 3.5	F 3.9	
Li 1.0	Be 1.5											Al 1.5	Si 1.8	P 2.1	S 2.4	Cl 2.8
Na 0.9	Mg 1.2											Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.7
K 0.9	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.2
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.5	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Cs 0.7	Ba 0.9	La-Lu 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.4	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9															

Source: L. Pauling, The Nature of the Chemical Bond and the Structure of Molecules and Crystals, Cornell University Press, Ithaca, New York, 1960. Used by permission of the publisher.

You can use the table of electronegativities and the following formula predict BOND TYPES:



There are three possible outcomes:

1. ΔEN is between 0.0 and 0.5 = "COVALENT" BOND
2. ΔEN is between 0.5 and 1.7 = POLAR "COVALENT" BOND
3. ΔEN is between 1.7 and 3.3 = IONIC BOND

IN GENERAL, compounds with a ΔEN value ≥ 1.7 is IONIC while a ΔEN value < 1.7 is COVALENT.

Example 4. Use the above formula to predict the bond type for the following compounds:

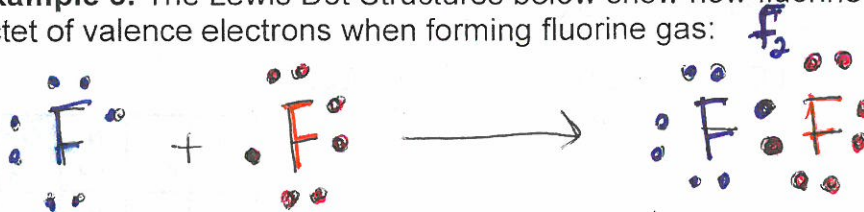
- a. $\text{HCl} = 2.8 - 2.2 = 0.6 \rightarrow$ ^(POLAR) COVALENT
- b. $\text{H}_2\text{O} = 3.5 - 2.2 = 1.3 \rightarrow$ (POLAR) COVALENT
- c. $\text{Cl}_2 = 2.8 - 2.8 = 0.0 \rightarrow$ COVALENT
- d. $\text{NH}_3 = 3.0 - 2.2 = 0.8 \rightarrow$ ^(POLAR) COVALENT

COVALENT BONDING: a chemical bond that involves the "SHARING" of electrons.

A covalent bond is formed when two atoms having less than full shells of electrons are able to share one or more of their electrons with each other in order to attain full electron shells.

OCTET RULE: ATOMS in groups 14 to 17 of the P.T. tend to form covalent bonds so as to have 8 valence e's!

Example 5. The Lewis Dot Structures below show how fluorine atoms gain an octet of valence electrons when forming fluorine gas:



IN GENERAL, covalent bonds are formed when a NON-METAL combines with another NON-METAL !!!!!

SEATWORK/HOMEWORK:

1. READ pgs 70 -74
2. Do Section Review QUESTIONS 3-6 pg 74
3. Ionic, Covalent + London Forces Ex 79, 80, 81 + 83