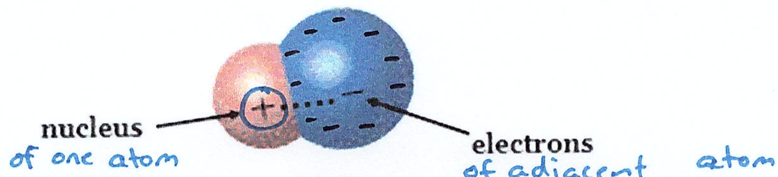
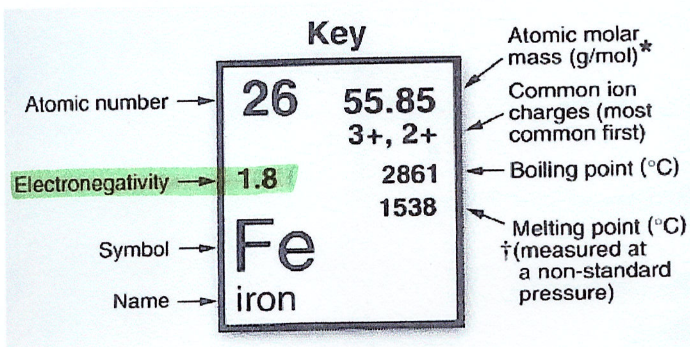
Bond Type and Electronegativity

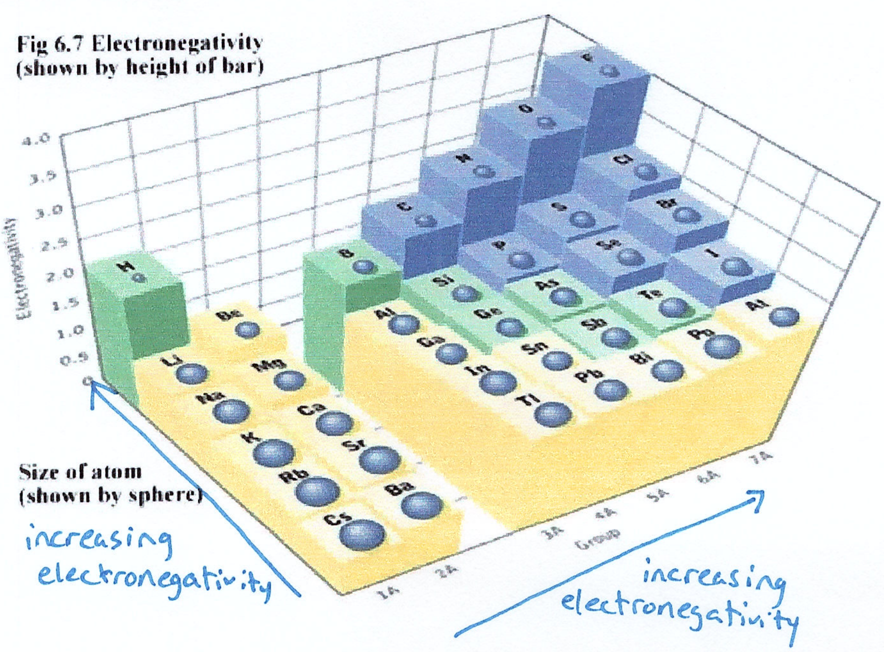
- The **electronegativity** of an element is the relative measure of the ability of the atom to attract electrons in a chemical bond
    - The attraction is between the positively charged protons in the nucleus of an atom and the valence electrons in the adjacent atom it is bonding with
    - Electronegativity can be thought of as the strength of attraction between the protons of an atom and the valence electrons of an adjacent atom



- The values for electronegativities are indicated in your periodic table



- Higher values of electronegativity indicate a greater attraction for electrons
    - The highest value for electronegativity is 4.0
    - The electronegativities for elements increase as you move from left to right across the period table and also increase as you move from the bottom to the top



- Electronegativities correlate with our knowledge of bonding
- Metals have a low electronegativity (metals do not strongly attract electrons) which is consistent with the concept that metals will lose electrons to form positive ions
  - Non-metals have a high electronegativity (non-metals strongly attract electrons) which is consistent with the concept that non-metals will gain electrons to form negative ions
  - The oppositely charged ions attract to each other to form ***ionic bonds***

ionic bond = electron transfer

- Covalent bonds (electrons being shared between atoms) exist in molecular compounds (compounds with only non-metals)

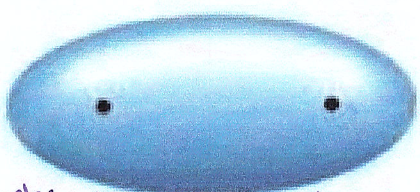
covalent bond = shared electrons

H — H (2.2, 2.2 electronegativity)

- If two bonded non-metal atoms have the same electronegativity, like $H_{2(g)}$, the two nuclei of the atoms will attract the electrons with exactly the same strength, thus the electrons are shared equally between the two atoms. This is referred to as a ***nonpolar covalent bond***.

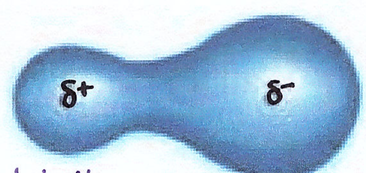
H — Cl (2.2, 3.2)

- If two bonded non-metal atoms have different electronegativity, like $HCl_{(g)}$, the sharing of electrons is no longer equal
  - The element with the higher electronegativity pulls the electrons closer to itself
  - This results in one end of the bond having a slightly negative charge (δ-) and the other end of the bond having a slightly positive charge (δ+)
- Bonds that have unequal sharing of electrons are called **polar covalent bonds**
  - The word "*polar*" just means that there is an overall charge separation (ie. one end is positive and the other end is negative)
  - Since these bonds have a negative "pole"/end and a positive "pole"/end, they are also called **bond dipoles**



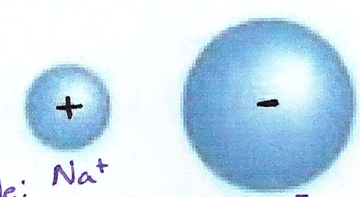
example: H —— H

Nonpolar covalent bond



example: H —— Cl

Polar covalent bond



example: Na^+ Cl^-

Ionic bond

* The difference in the electronegativities of the elements involved in the bond determines the bond type
  * The **electronegativity difference (ΔEN)** is calculated by subtracting the smaller electronegativity from the larger electronegativity

memorize! * general guideline

Electronegativity Difference Between Two Atoms (ΔEN)	Type of Bond Between Atoms	Description of the Electrons in the Bond
greater or equal to 1.7	mostly ionic	Transfer of electrons between metal and non-metal to form ions
between 0 and 1.7	polar covalent	Electrons shared unequally between unlike atoms
equal to 0	non-polar covalent	Electrons shared equally between identical non-metal atoms

EXAMPLES: For each pair of elements,

a. calculate the electronegativity difference
b. identify the bond type
c. draw dipole bond arrows if applicable

$[Na]^+ [Cl]^-$

Na—Cl

$\Delta EN = EN_{Cl} - EN_{Na}$
$\Delta EN = 3.2 - 0.9$
$\Delta EN = 2.3$ ∴ mostly ionic

arrow head points toward "–" end/pole; dipole bond arrow (0.8); "+" end/pole

$\delta-$ **N—H** $\delta+$

$\Delta EN = EN_N - EN_H$
$\Delta EN = 3.0 - 2.2$
$\Delta EN = 0.8$ ∴ polar covalent

P—H

$\Delta EN = EN_P - EN_H$
$\Delta EN = 2.2 - 2.2$
$\Delta EN = 0$ ∴ non-polar covalent

(2.0)

$\delta+$ **B—F** $\delta-$

$\Delta EN = EN_F - EN_B$
$\Delta EN = 4.0 - 2.0$
$\Delta EN = 2.0$

* can't be ionic b/c only non-metal ∴ polar covalent

* Bond classification is not simple and has no clear distinction between ionic and covalent bonding
  * Bonding needs to be thought in terms of a continuum with electron transfer at one end of the continuum, polar covalent in the middle, and nonpolar covalent at the other end of the continuum

*****Now try Practice Problems #1-4*****

3.3 — 1.7 — 0 (ΔEN)

mostly ionic — polar covalent — non-polar covalent

Practice Problems

1. For each pair of elements,
   i. calculate the electronegativity difference
   ii. identify the bond type
   iii. draw dipole bond arrows if applicable

a. N – N

$\Delta EN = 3.0 - 3.0$

$\Delta EN = 0$

∴ non-polar covalent

b. $\delta+$ S – O $\delta-$ (dipole arrow → 0.8)

$\Delta EN = 3.4 - 2.6$

$\Delta EN = 0.8$

∴ polar covalent

c. K – Br

$[K]^+[Br]^-$

$\Delta EN = 3.0 - 0.8$

$\Delta EN = 2.2$

∴ mostly ionic

d. $\delta-$ C – H $\delta+$ (dipole arrow ← 0.4)

$\Delta EN = 2.6 - 2.2$

$\Delta EN = 0.4$

∴ polar covalent

e. $\delta+$ Si – Cl $\delta-$ (dipole arrow → 1.3)

$\Delta EN = 3.2 - 1.9$

$\Delta EN = 1.3$

∴ polar covalent

f. $\delta+$ P – F $\delta-$ (dipole arrow → 1.8)

$\Delta EN = 4.0 - 2.2$

$\Delta EN = 1.8$

∴ polar covalent
b/c 2 non-metals!

g. Cl – Cl

$\Delta EN = 3.2 - 3.2$

$\Delta EN = 0$

∴ non-polar covalent

h. $\delta-$ O – H $\delta+$ (dipole arrow ← 1.2)

$\Delta EN = 3.4 - 2.2$

$\Delta EN = 1.2$

∴ polar covalent

i. $\delta+$ N – O $\delta-$ (dipole arrow → 0.4)

$\Delta EN = 3.4 - 3.0$

$\Delta EN = 0.4$

∴ polar covalent

2. What types of atoms form bonds that are:

   a. mostly ionic

   metal & a non-metal

   b. polar covalent

   2 non-metals with different electronegativities

   c. non-polar covalent

   2 non-metals with equal electronegativities

3. Arrange the elements in sets (a) through (d) in order of increasing attraction for electrons involved in a bond.

   a. C and O $\Delta EN = 0.8$

   b. Mn and O $\Delta EN = 1.8$

   c. N and Br $\Delta EN = 0$

   d. Ca and I $\Delta EN = 1.7$

   c, a, d, b

4. Describe the nature of the chemical bond that would form between the following pairs of elements.

   a. Mg and Cl metal & non-metal $\Delta EN = 1.9$

   ∴ mostly ionic

   b. Se and O 2 non-metals $\Delta EN = 0.8$

   ∴ polar covalent

   c. As and H 2 non-metals $\Delta EN = 0$

   ∴ non-polar covalent

   d. Si and F 2 non-metals $\Delta EN = 2.1$

   ∴ polar covalent