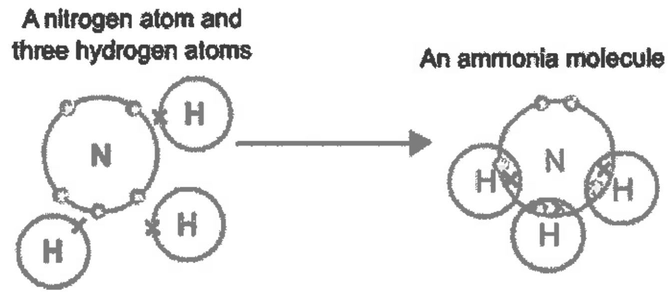
Types of Chemical Bonds

1. Covalent

- sharing of valence electrons to create a stable octet
  - hydrogen forms stable configurations when it shares two electrons (called the duet rule)



- atoms with similar ionization energies usually form covalent bonds (usually nonmetals)

- the shared electrons are considered to belong to both atoms at the same time and holds the atoms together to form a molecule

- pairs of electrons that do not participate in chemical bonds are called lone electron pairs

- more than 1 valence pair may be shared between 2 atoms :O=O: ← each has access to 8
  - 2 valence pairs shared = double bond e.g. O_2 12e⁻
  - 3 valence pair shared = triple bond e.g. N_2 10e⁻ :N≡N:

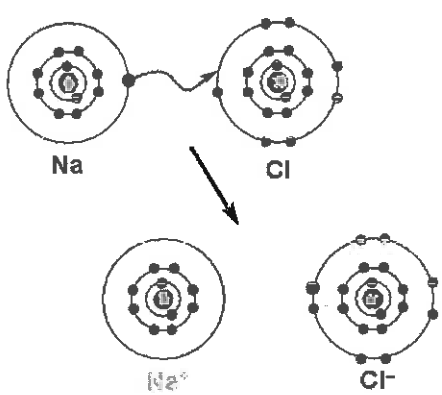
2. Ionic

- transfer of valence electron(s) from a metal to a nonmetal to form an ionic compound

- metals want to lose valence e^- to form cations Na + energy ---> [Na^+] + 1e-

- nonmetals want to gain valence e^- to form anions Cl + 1e- ---> [Cl^-] + energy

For example:



Na = lose 1 e^- to have 10 e^-
e^- configuration similar to Ne
$1s^2 2s^2 2p^6$

Cl = gain 1 e^- to have 18 e^-
e^- configuration similar to Ar
$1s^2 2s^2 2p^6 3s^2 3p^6$

* Note: Compounds can contain both ionic and covalent bonds if they contain a polyatomic ion.

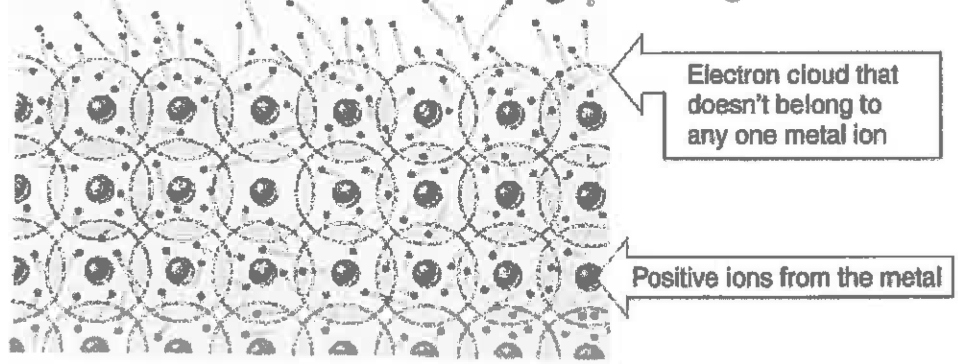
For example: $Ca(OH)_2$ The bond between oxygen and hydrogen in the OH^- ion is covalent while the bond between Ca^+ and OH^- is ionic.

3. Metallic

like Rice Krispies
- puffed rice are metal nuclei
- marshmallows are the e⁻ holding everything together

- metal atoms share a "sea of electrons"
- electrons can "float" freely between atoms; allows metals to conduct electricity well

due to ↓ valence e⁻ and empty orbitals



Predicting Bond Type

Bonding is not usually purely ionic or covalent, but somewhere in between. The difference in electronegativity (ΔE) of the atoms in a bond can help us identify the bond type.

Electronegativity is an atom's ability to attract the shared paired electrons to itself in a covalent bond (a measure of the affinity for electrons by element).

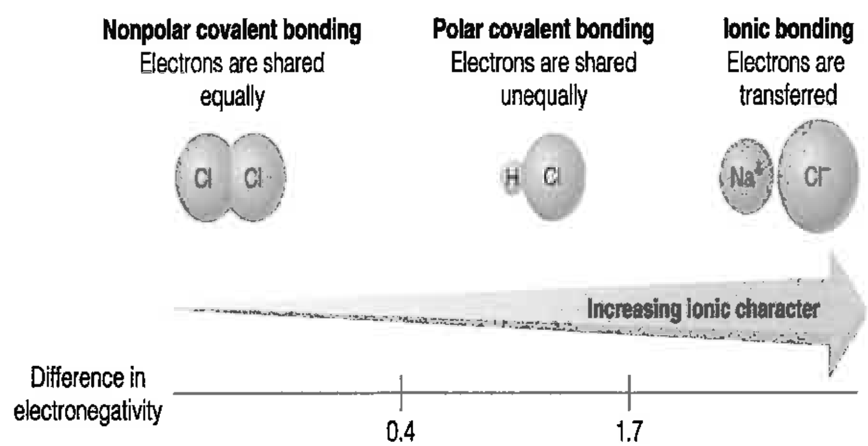
Increasing →

Decreasing ↓

1 H 2.1																
3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0
11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0
19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.9	28 Ni 1.9	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8
37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.7	49 In 1.7	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5
55 Cs 0.7	56 Ba 0.9	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.9	83 Bi 1.9	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1														

Electronegativities of the Elements

- Metals have low electronegativity values.
- Nonmetals have high electronegativity values.
- Down a group, electronegativity decreases
- Across a row, electronegativity increases



1. Pure (Nonpolar) Covalent Bonding ($\Delta E < 0.4$)

- equal or near equal <u>sharing</u> of electrons between atoms

$\Delta E = |\quad|$ or higher # first − second E

For example: F_2 — B_4Si_3

$\Delta E = F - F$
$= 4 - 4$
$= 0$

$\Delta E = B - Si$
$= 2 - 1.8$
$= 0.2$

2. Polar Covalent Bonding ($\Delta E = 0.5 – 1.7$)

- unequal sharing of electrons between atoms, meaning the electron pair spends more time near one atom than the other
- results in a <u>partial charge</u> on each atom

For example: HCl

$\Delta E = Cl - H$
$= 3 - 2.1$
$= 0.9$ ($>0.4 \therefore$ unequal sharing)

- indicates Cl has greater affinity for electrons
  - the shared electron pair spend more time here than near H
  - this separation of positive and negative charges is called a <u>dipole</u>
  - to show the partial charge, an arrow pointing in the direction of the slightly negative atom is used

$$\underset{\delta^+}{H} - \underset{\delta^-}{Cl}$$

Other Examples: H_2O — CH_4

$\Delta E = 3.5 - 2.1$
$= 1.4$ ($>0.4 \therefore$ unequal sharing + a dipole exists)

$\Delta E = 2.5 - 2.1$
$= 0.4$ ($\leq 0.4 \therefore$ no dipole)

3. Ionic Bonding ($\Delta E > 1.7$)

- electron(s) transferred