PERIODIC PROPERTIES & TRENDS

PERIODIC PROPERTIES OF ELEMENTS

- The properties of the elements exhibit trends.
- The trends can be predicted using the periodic table.
- They can be explained by analyzing the electron configurations

A LITTLE HISTORY…

1800's

- Knowledge of chemical processes increases
- Allowed for the isolation of more elements
- Need for organizing and grouping

A LITTLE HISTORY…

1869

Mendeleev (Russia) and Meyer (Germany) published identical papers

- Chemical and physical properties reoccur
- periodically with increasing atomic mass

Mendeleev went beyond just arranging the elements

- Assumed blank spaces were undiscovered elements, for which he predicted the properties of, based on the properties of neighboring elements
- These elements were shortly discovered and were as he described

Mendeleev did have a problem with the order of: Co-Ni and Te-I

A LITTLE HISTORY (CONT.)

1913

Moseley – with x-ray techniques, found atomic number

Modern Periodic Law

There is a periodic repeat of properties when elements are arranged by an increase atomic number

Why is that??

QUANTUM THEORY

- The behaviors of elements repeat because electron arrangements repeat.

Be $1s^2\ 2s^2$

Mg $1s^2\ 2s^2\ 2p^6\ 3s^2$

Ca $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

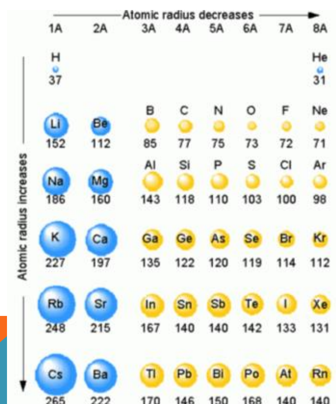
Sr $[Kr]\ 5s^2$

- Elements in the same group (column) have the same number of valence electrons

ATOMIC SIZE (RADIUS)

- **Decreases** left to right across the periodic table
  - As nuclear charge increases, the attraction between the nucleus and electrons increases
- **Increases** top to bottom on the periodic table
  - As the number of energy levels increases, the outermost electrons are farther away

***Also: As the number of energy levels between the outer electrons and the nucleus increases, the attraction between the nucleus and the outer electrons decreases (**shielding effect**)



IONIC SIZE (RADIUS)

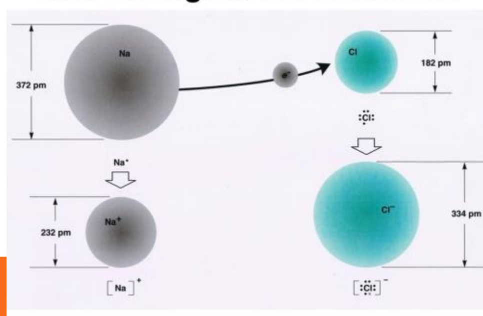
Anions - gain of electrons (generally nonmetal ions)

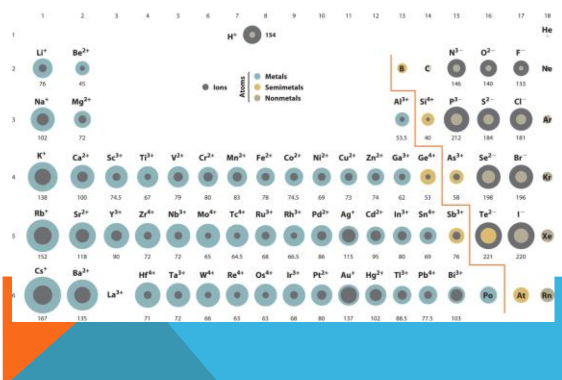
- Are larger than the neutral atom
- The increased number of electrons reduces attractive charge on each

Cations - loss of electrons (generally metal ions)

- Are smaller than neutral
- Usually involves the loss of valence electrons (eliminating a full level) in addition to the excess positive charge drawing the remaining electrons closer

Size Change in Ion Formation





Sizes of atoms and their ions in pm

Group 1		Group 2		Group 3		Group 16		Group 17	
Li^+ 90	Li 134	Be^{2+} 59	Be 90	B^{3+} 41	B 82	O 73	O^{2-} 126	F 71	F^- 119
Na^+ 116	Na 154	Mg^{2+} 86	Mg 130	Al^{3+} 68	Al 118	S 102	S^{2-} 170	Cl 99	Cl^- 167
K^+ 152	K 196	Ca^{2+} 114	Ca 174	Ga^{3+} 76	Ga 126	Se 116	Se^{2-} 184	Br 114	Br^- 182
Rb^+ 166	Rb 211	Sr^{2+} 132	Sr 192	In^{3+} 94	In 144	Te 135	Te^{2-} 207	I 133	I^- 206

neutral atoms gray.
cations red. anions blue.

OTHER TRENDS

Periodic Table of the Elements

Density

http://chemistry.about.com
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About Chemistry

Solid or Liquid: g/cm³ at 20°C and 1 atm
Gas: g/liter at 0°C and 1 atm

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

OTHER TRENDS

Periodic Table of the Elements

Boiling Point
°C and 1 atm

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About Chemistry

tp = triple point

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

OTHER TRENDS

Periodic Table of the Elements

Melting Point
°C and 1 atm

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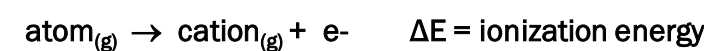
About Chemistry

tp = triple point
sp = sublimation point

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

IONIZATION ENERGY (IE)

The energy required to remove the outermost electron from a gaseous atom

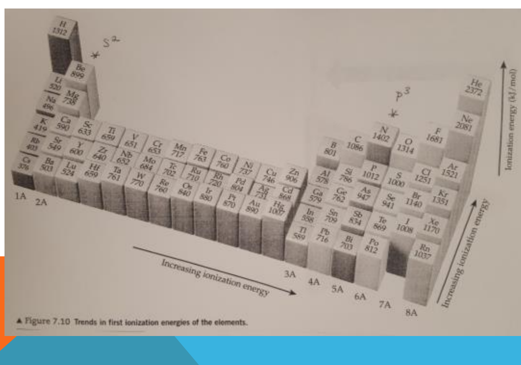


atom$_{(g)}$ $\rightarrow$ cation$_{(g)}$ + e-   ΔE = ionization energy

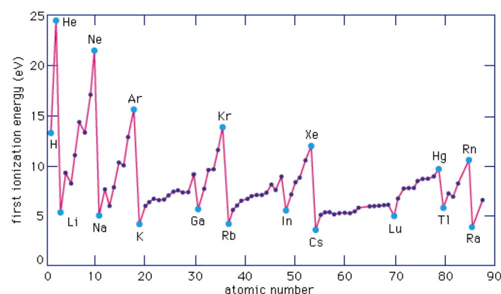
1[st] IE of elements:

- Increases moving left to right across the periodic table
- Decreases moving top to bottom

(The opposite trend of atomic size – is there a connection?)



▲ Figure 7.10 Trends in first ionization energies of the elements.



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IONIZATION ENERGY (IE)

There is a break in the trend at Be and N. Why?

Be - alkaline earth metals s^2 = partial stability

N - nitrogen group p^3 (½ filled sublevel) = partial stability

I_1 (1st ionization energy)
- The energy to remove the outermost electron

I_2 (2nd ionization energy)
- The energy to remove a second (once the 1st one is gone)

I_3 (3rd ionization energy)
- The energy to remove a third (once the first two are gone)

Successive Ionization Energies in Kilojoules per Mole for the Elements in Period 3

General increase →; General decrease ↑

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4560					
Mg	735	1445	7730	Core electrons*			
Al	580	1815	2740	11,600			
Si	780	1575	3220	4350	16,100		
P	1060	1890	2905	4950	6270	21,200	
S	1005	2260	3375	4565	6950	8490	27,000
Cl	1255	2295	3850	5160	6560	9360	11,000
Ar	1527	2665	3945	5770	7230	8780	12,000

*Note the large jump in ionization energy in going from removal of valence electrons to removal of core electrons.

IONIZATION ENERGY (IE)

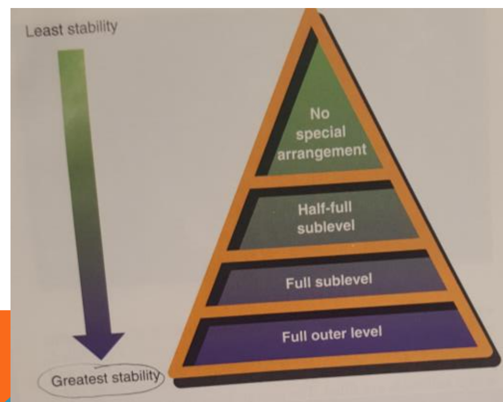
With each additional electron removed, the IE increases.

Note: a BIG increase at some point in the series for each element. Why?

For Na, at I_2

For Mg, at I_3

For P, at I_6



ELECTRON AFFINITY

The energy change that occurs when an electron is added to a gaseous atom

$atom_{(g)}$ + e- → $anion_{(g)}$ ΔE = e- affinity

- Halogens: Largest negative values
- Noble gases: Positive values
- Alkaline earth metals/Nitrogen group elements:
  full s sublevel ½ full p sublevel
- These are either positive or not as negative as would be predicted based on the trend

Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0

METALS

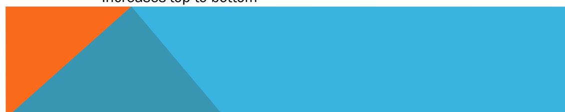
Properties

- Shiny, Malleable, Ductile
- Oxides form basic ionic solids
  - Ex: $Na_2O_{(s)} + H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)}$
- React with acids to form salt and water
  - Ex: $MgO_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_2 + H_2O_{(l)}$
- Low ionization energies

Metallic character

The extent to which an element exhibits these physical/chemical properties

- Decreases left to right across the periodic table
- Increases top to bottom



NONMETALS

Properties

- More diverse in behavior than metals
- Non-lustrous, Poor heat/electrical conductors
- Lower melting point than metals
- Most nonmetal oxides are acidic
  - Ex: $CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$
- React with bases to form salt and water
  - Ex: $CO_{2(g)} + 2\ NaOH_{(aq)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$

METALLOIDS

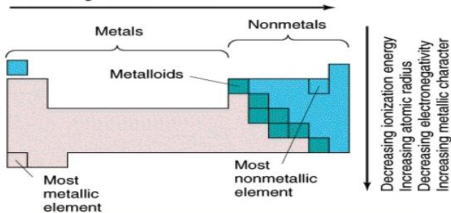
- Properties intermediate between metals and nonmetals
- Useful as semiconductors

Ex: Silicon – metallic luster, but brittle

INCREASING METALLIC CHARACTER

INCREASING METALLIC CHARACTER

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112	113	114				

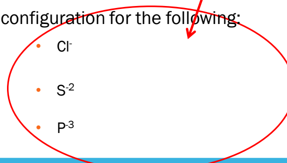


E- CONFIGURATIONS FOR IONS

- Loss or gain of e-
  - Loss of electrons (cation) – positive charge
  - Gain of electrons (anion) – negative charge

ISOELECTRONIC SERIES
- same # of e-
- Same e- configuration

- Write the condensed e- configuration for the following:
  - Cl^-
  - S
  - P
  - Cl^-
  - S^{-2}
  - P^{-3}



NOBLE GASES

- Stable
- Full s & full p sublevels
  - 2 e- + 6 e- = 8 e-
- Don't bond to form compounds (in nature)

MORE PRACTICE W/ ION E- CONFIGURATIONS

- Write the condensed e- configuration for the following:
  - Ca
  - Zn
  - Ga
  - Pb
  - Ca^{+2}
  - Zn^{+2}
  - Ga^{+3}
  - Pb^{+2}
  - Pb^{+4}

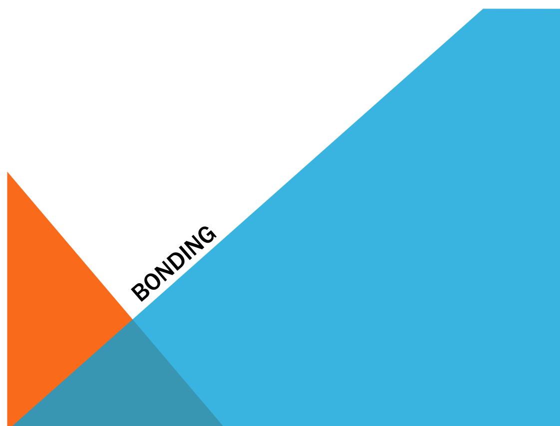
ION FORMATION OF TRANSITION METALS

- Usually can't attain noble gas configuration
- Octet rule has limitations
- Atoms don't usually lose more than 3 e-
- Electrons are lost from highest n level first

Ex: cobalt $[Ar]4s^2 3d^7$

Co^{+2} $[Ar]3d^7$

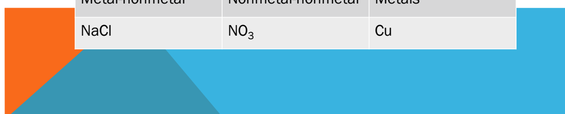
Co^{+3} $[Ar]3d^6$



CHEMICAL BONDS

- forces that hold groups of atoms together to make them function as a unit
- valence electrons are shared or transferred

Ionic	Covalent	Metallic
Transfer e-	Sharing e-	Metal nuclei floating in a sea of e-
Electrostatic forces	Overlap of orbitals	
Metal-nonmetal	Nonmetal-nonmetal	Metals
NaCl	NO_3	Cu



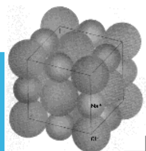
WHY BOND?

- The system can achieve a state of lower energy (become more stable)
- Octet Rule: bonded atoms tend to achieve 8 valence electrons

IONIC COMPOUNDS

Characteristics:
- Atoms held together by ionic bonds
- Consists of closely packed, oppositely charged ions
- Metal + Nonmetal
- Brittle
- Relatively high melting and boiling points
- Soluble in water
- Conduct electricity when molten or dissolved in water

Examples: NaCl, KBr, $CaSO_4$



IONIC BONDS

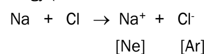
- One atom (metal) **loses** electron(s) ← endothermic
- One atom (nonmetal) **gains** electron(s) ← exothermic
- The endothermic rxn is always larger in magnitude than the exothermic rxn
- Ex: For the reaction Na + ½ Cl_2 → NaCl

Na needs to **lose** an e- → **IE** Na = +496 kJ/mol

Cl needs to **gain** an e- → **e- affinity** Cl = -349 kJ/mol

+147 kJ/mol

It appears that the formation of NaCl from ions does not decrease in energy (ie – is **not more stable**)



$Na + Cl \rightarrow Na^{+} + Cl^{-}$

[Ne] [Ar]

However ... Na + ½ Cl_2 → NaCl ΔH_f° = -410.9 kJ/mol

In actuality....there is a decrease in energy (it **is more stable** as NaCl)

+147 kJ/mol ≠ -410.9 kJ/mol ?!

- Stability of **ionic compounds** comes from the attraction between the positive and negative ions
- When the ions are drawn together, energy is released

Lattice Energy

The energy required to completely separate one mole of a solid ionic compound into gaseous ions

Ex: $NaCl_{(s)} \rightarrow Na^{+1}_{(g)} + Cl^{-1}_{(g)}$ $\Delta H_{lattice}$ = +788 kJ/mol

$\Delta H_{lattice}$ increases as the ionic charges increase and radii decrease

COLOUMB'S LAW

$$E = k\frac{Q_1Q_2}{d}$$

Q_1Q_2 ← charges on ions

d ← distance between nuclei (nm)

k = 2.31×10^{-19} J nm

Crystal Lattice

Attraction between ions is multidirectional

Ex: NaCl each Na^{+1} is surrounded by 6 Cl^{-}

each Cl^{-} is surrounded by 6 Na^{+1}

Different ions pack together differently to produce different crystal shapes

COVALENT COMPOUNDS

Characteristics
- form molecules
- relatively low melting and boiling points
- have bond angles and bond lengths

Examples		
Sugar ($C_6H_{12}O_6$)	Water (H_2O)	Natural gas CH_4 (methane)

COVALENT BONDS

- Formed by two atoms sharing a pair of electrons between them
- Sharing may be equal or may not
  - a. equal sharing = ***nonpolar*** covalent bond
  - b. not equal sharing = ***polar*** covalent bond

- Polar covalent bonds
  - electrons attracted to one of the atoms in the bond more than the other
  - this creates a *dipole*

HOW DO YOU KNOW WHAT TYPE OF BOND WILL FORM BETWEEN TWO ATOMS?

CHART OF ELECTRONEGATIVITY

- each element is assigned an numerical value
- the greater the value, the more that atom attracts shared electrons
- increases left to right on periodic table
- decreases top to bottom

If the difference in e-neg values of the two atoms is...	
less than 0.5	nonpolar covalent
from 0.5 to less than 2.0	polar covalent
2.0 or greater	ionic

CHART OF ELECTRONEGATIVITY

H 2.1																	
Li 1.0	Be 1.5											B 1.5	C 2.5	N 3.0	D 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 3.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
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 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9		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 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9															

LEWIS STRUCTURES & FORMAL CHARGES

LEWIS STRUCTURES (ELECTRON DOT DIAGRAMS)

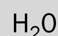
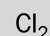
- Symbol of element is surrounded by dots and/or dashes that represent valence electrons

Examples	
Na	Li
Mg	C
S	Ne

LEWIS STRUCTURES FOR MOLECULAR COMPOUNDS

- Count up total number of valence e- of all atoms
- Distribute them in a way so that:
  - hydrogen and alkali metals are surrounded by 2 e-
  - alkaline earth metals are surrounded by 4 e-
  - boron group elements are surrounded by 6 e-
  - carbon group and beyond follow the *Octet Rule* (surround with 8 electrons)
- Dash = 2 e- = single bond

Examples

Examples	
 H_2O	 Cl_2
 N_2	 BeH_2
 SO_4^{-2}	 CO_2

FORMAL CHARGES

- Charges assigned to each atom in a Lewis structure
- Answers the question: which structure is better?

# of valence e- *(all unbonded e-)*

- number of e- in structure *(½ of electrons in its bonds)*

formal charge

Ex: CO_2

EXCEPTIONS TO OCTET RULE

- Atoms with less than an octet (mentioned previously)
  - Alkali Metals, Hydrogen and Helium → 2 e-
  - Alkaline Earth Metals → 4 e-
  - Boron Group elements → 6 e-
- Atoms with more than an octet
  - Ex: SF_6
- Molecules with an odd number of e- (radicals/unstable)

BOND STRENGTH

- Comparing bonds between the same atoms:

Single bond strength < double bond strength < triple bond strength

TABLE 8.4 Average Bond Enthalpies (kJ/mol)

Single Bonds							
C—H	413	N—H	391	O—H	463	F—F	155
C—C	348	N—N	163	O—O	146		
C—N	293	N—O	201	O—F	190	Cl—F	253
C—O	358	N—F	272	O—Cl	203	Cl—Cl	242
C—F	485	N—Cl	200	O—I	234		
C—Cl	328	N—Br	243			Br—F	237
C—Br	276			S—H	339	Br—Cl	218
C—I	240	H—H	436	S—F	327	Br—Br	193
C—S	259	H—F	567	S—Cl	253		
		H—Cl	431	S—Br	218	I—Cl	208
Si—H	323	H—Br	366	S—S	266	I—Br	175
Si—Si	226	H—I	299			I—I	151
Si—C	301						
Si—O	368						
Multiple Bonds							
C=C	614	N=N	418	O_2	495		
C≡C	839	N≡N	941				
C=N	615	N=O	607	S=O	523		
C≡N	891			S=S	418		
C=O	799						
C≡O	1072						

BOND LENGTH

- Comparing bonds between the same atoms:

single bond length > double bond length > triple bond length

TABLE 8.5 Average Bond Lengths for Some Single, Double, and Triple Bonds

Bond	Bond Length (Å)	Bond	Bond Length (Å)
C—C	1.54	N—N	1.47
C=C	1.34	N=N	1.24
C≡C	1.20	N≡N	1.10
C—N	1.43	N—O	1.36
C=N	1.38	N=O	1.22
C≡N	1.16		
		O—O	1.48
C—O	1.43	O=O	1.21
C=O	1.23		
C≡O	1.13		

RESONANCE STRUCTURES

- Two or more equally good Lewis structures for one molecule
  - Ex: NO_3^{-1}

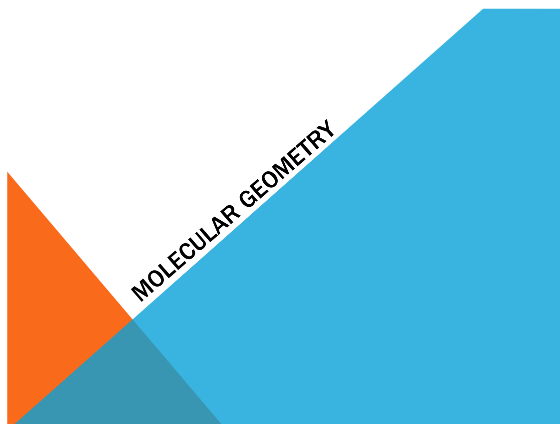
- One short bond (double) and 2 longer bonds (singles)??
- No… three equal length bonds, intermediate of single and double
- Explain?

HOW DO WE EXPLAIN THIS?

- Electrons are *delocalized*
  - double bond is moving from one atom to another
  - leads to an average bond length
  - leads to higher stability
- Ex: Benzene, C_6H_6

METALLIC BOND

- Most metals have 1-3 valence electrons
- Most metals have empty d orbitals
- Valence electrons are loosely held and freely move from one nucleus to another



MOLECULAR GEOMETRY

- Description of the three dimensional shape of a molecule

VSEPR Theory (Valence Shell Electron Pair Repulsion)

- Like charges repel, so electron groups around an atom move as far away from each other as possible

H_2O H – O – H linear? no.

	H	
CH_4	H – C – H	bond angle 90°? no.
	H	

LINEAR

- Includes all molecules with only two atoms

H_2	H – H	
O_2	O – O	No bond angles
HCl	H – Cl	

- Includes some molecules with three of more atoms

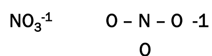
BeH_2	H – Be – H	
		180° bond angle
N_2O	N – O – N	

TRIGONAL PLANAR

BF_3

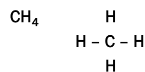
120° bond angle

NO_3^{-1}



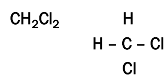
TETRAHEDRON

CH_4



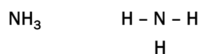
Bond angle 109.5°

CH_2Cl_2



TRIGONAL PYRAMID

NH_3



Bond angle 107°

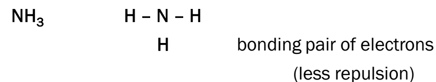
BENT OR ANGULAR

H_2O

Bond angle 105°

nonbonding pair of electrons
(more repulsion)

NH_3



bonding pair of electrons
(less repulsion)

REPULSION ORDER

single bond < double bond < triple bond

$2\ e^-$ $4\ e^-$ $6\ e^-$

trigonal planar 120°