(3) **Today**

Reviewing Periodic Trends

Section 1.4
Introduction to Chemical Bonding Theories
octet rule etc

Sections 1.5-1.10
Valence Bond Theory

Next Class (4)

Skipping Section 1.11 for now
An introduction to Molecular Orbital Theory

Sections 1.12
Drawing Chemical Structures

(5) **Second Class from Today**

Sections 1.12
Drawing Chemical Structures

Third Class from Today (6)

Sections 2.1 - 2.4
Polar Covalent Bonds, Formal Charges,
Resonance/Electron Delocalization

Sections 2.4 – 2.6
Resonance/Electron Delocalization

Lab starts today, Monday, September 8

Goggles and lab coats will be required starting Monday, September 22.

Use the periodic table to determine electron configurations ✓

Use the periodic table to determine the number of valence electrons ✓

Use the periodic table to identify metals and non-metals ✓

Use the periodic table to remember trends in size

Use the periodic table to remember trends in electronegativity

Use the periodic table to predict likely charges of ions

Use the periodic table to predict likely bond formation

Use trends in size, electron configuration, and nuclear charge to explain electronegativity trend

Introduce Valence Bond Theory (hybridization)

Different Ways of Representing Chemicals

The Periodic Table Is Your Friend: Metals tend to gain and nonmetals tend to lose electrons

+1	+2			-3	-2	-1	
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	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

58 Ce	68 Er	69 Tm	70 Yb	71 Lu
90 Th	100 Fm	101 Md	102 No	103 Lr

+11 Na $\underbrace{1s^2\ 2s^2\ 2p^6}_{\text{shielding } e^-}$ $3s^1$ (+2.2)

+6.85 Na^+ $1s^2$ } $2s^2\ 2p^6\ 3s^0$

Na^{2+} $1s^2$ } $2s^2\ 2p^5\ 3s^0$

these e^-'s experience a much higher effective nuclear charge (+6.85) and it is just too strongly attracted to the nucleus to be removed during typical chemical reactions

+5.2

F $1s^2\ 2s^2\ 2p^5$

F^- $1s^2\ 2s^2 2p^6$

the 6th e^- experiences are + charge of ~+5.2

lots of attraction

F^{2-} $1s^2\ 2s^2 2p^6$ } $3s^1$

the new e^- experiences a +0.2 charge ... not enough attraction to keep e^- on F^{2-}

shielding core } valence e^-

Why

Na 10.7, 6.85, 2.20; F 8.7, 5.2, 0.2

The Periodic Table Is Your Friend: Size

Review

Which is larger... Li or F? F has a +9 nucleus & its e^-'s are strongly attracted to the nucleus

F's e^-'s are pulled in closer to the nucleus than Li's

Li out electron experiences +2.4

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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 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Size increases from right to left

small change in size

effective nuclear charge is more important

Size increases from top to bottom

putting e^-'s in more distant shells is a significant change in size

Remember periodic trends

The Periodic Table Is Your Friend: Electronegativity

A measure of an atom's ability to attract e^-'s that are being shared with another atom.

Which would be better at attracting e^-'s in a bond F or Li

Li: e^- e^- F

+2.6 +5.2

high Z_{eff}

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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
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weird stuff happens here

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
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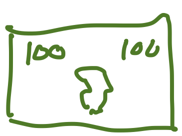
increase from left to right and from bottom to top

Remember periodic trends

Why does electronegativity or the size of the atom matter? Review

High energy electrons are reactive

concentrated e^-'s experiencing a low ⊕ charge nucleus



concentrated money attractive

concentrated ⊖ charge attractive

low energy electrons are less reactive

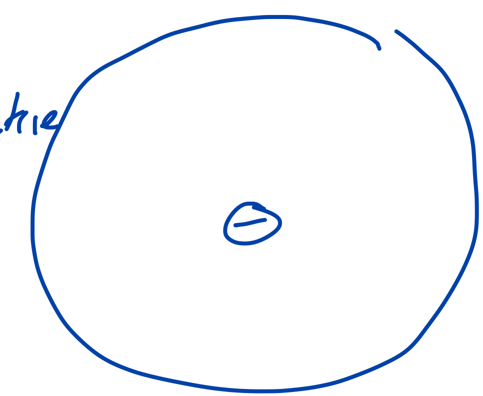
Stabilize e^- by getting them close to a high + charge

stabilize e^- by putting them in large orbitals

difuse money not as attractive

10,000 ¢

diffuse ⊖ charge not as attractive



Most common # of bonds formed by an atom is equal to the # of e^-'s the atom acquires when it becomes an anion

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3 Li	4 Be			5 B	6 C	7 N	8 O	9 F	10 Ne
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Halogens tend to form 1 bond

O + S tend to form 2 bonds

H-O-H $^{\ominus}$O-H can do 1 $H-C^{\ominus}(H)(H)$ can do 3

N + P tend to form 3 bonds

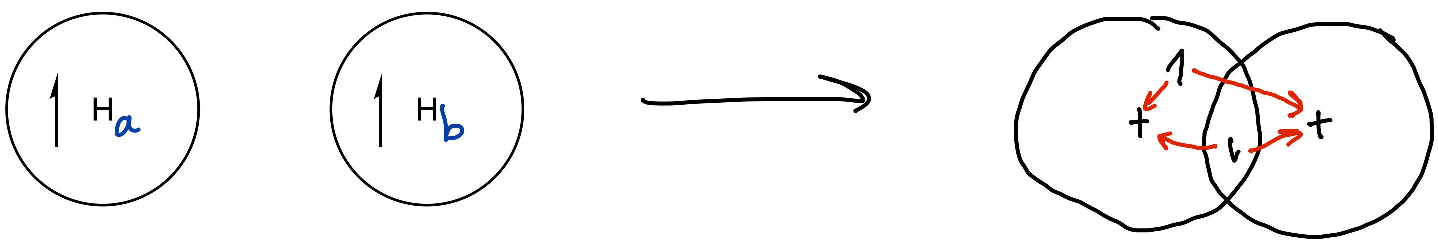
$H-N(H)-H$ $H-N^{\ominus}-H$ can do 2 $H-N^{\oplus}(H)(H)-H$ can do 4

C tends to form 4 bonds

C≡O $^{\ominus}$C≡N very few stable molecules w/3 bonds to C

Predict the number of electrons or bonds needed for an element to form a stable compound

H—H



$1s^1$ $1s^1$

room for 1 more e^-

room for 1 more e^-

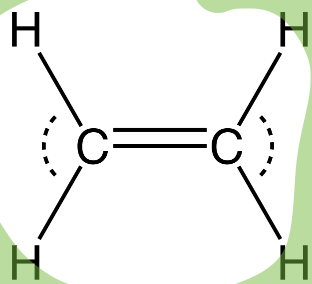
each orbital has room for 1 more e^-, so the e^- from H_a can be attracted to the nucleus of H_b

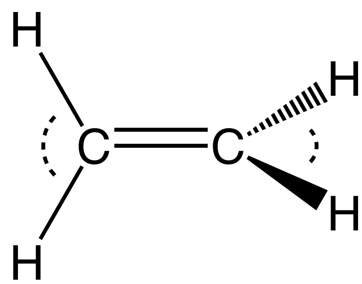
e^-'s made more stable by being attracted to more + charge

nuclei stay next to each other because of their shared attraction to a pair of e^-

Wait, what can we use Valance Bond Theory for?

How do we use the $2s$, $2p_x$, $2p_y$, + $2p_z$ orbitals on C to form bonds





Which one? Both C atoms are trigonal planar

Why is there free rotation around C to C single bonds but not C to C double bonds?

Which bond is stronger?

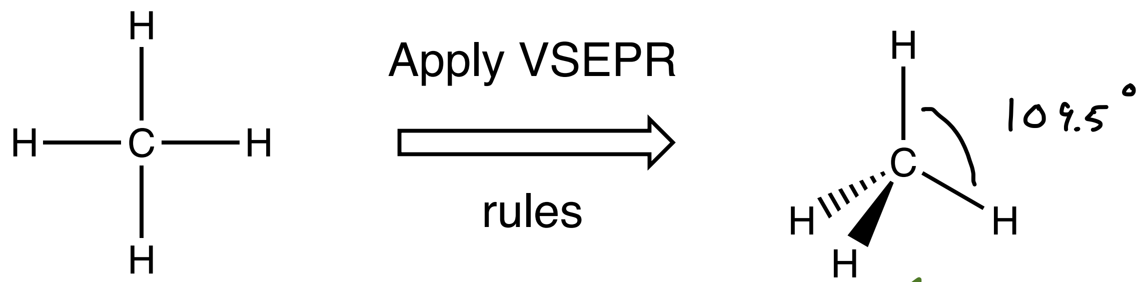
$HC\equiv C-CH_3$

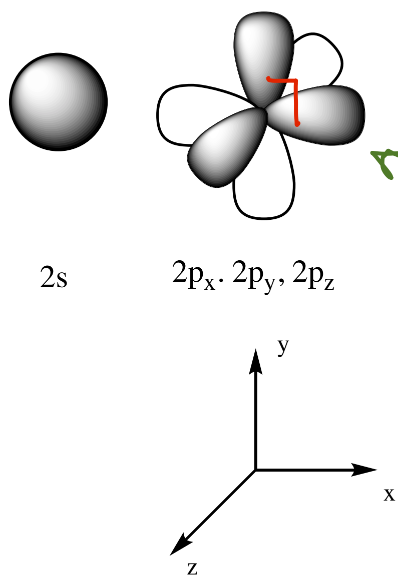
?

$H_3C-CH_2-CH_3$

Valence bond theory can explain/ answer these questions

Explain observations and make predictions based on Valence Bond Theory





how do we get 109.5° bond angles when the 2p orbitals are all perpendicular to each other?