Today

Sections 1.1 – 1.4
- Periodic Trends
- Different ways of representing molecules

Section 1.6
- An introduction to Molecular Orbital (MO) Theory

Next Class

Section 1.6
An Introduction to MO Theory

Sections 1.7-1.15
Valence Bond Theory

Introduction to Mastering Chemistry for Organic Chemistry is open and due by 11:59 pm on 9/18

Homework Chapter 1 is open and due by 11:59 pm on 9/21

a better estimate
Z_{eff} for C is 3.25
for F is 5.20

valence e^- 4e⁻ 7e⁻
Z_{eff} +4 +7
core e^- 2 e^- 2e⁻
+6 +9

effective nuclear charge is what valence e^-'s experience as the nuclear charge

eneg increases from left to right because nuclear charge increases but the # of core e^-'s stays the same.

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 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

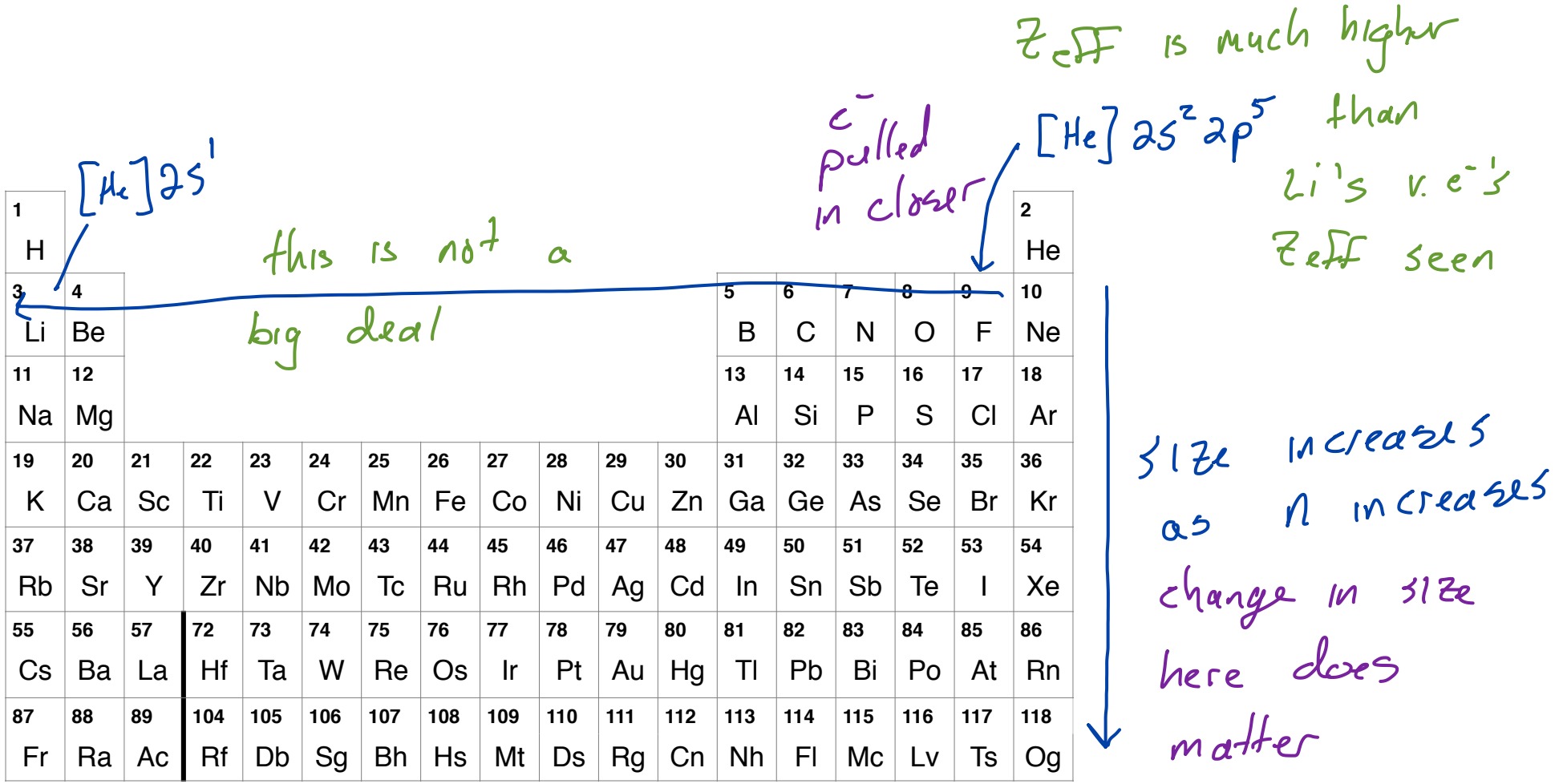
because V e^-'s move farther out from the nucleus as we go down the table (as the principal quantum # increases) attraction decreases so eneg decreases

Remember periodic trends Z_{eff} = 5.2 & 6.1

F^- Cl^-

The Periodic Table Is Your Friend: Size

Sections 1.1 – 1.3



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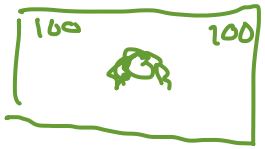
Remember periodic trends

high energy electrons are reactive

concentrated electrons



$F^{\ominus}$

weakly basic

e⁻'s near to higher positive charge

strong base

which ⊖ is more stable or less attractive to H⊕

low energy electrons are less reactive

diffuse electrons

$Cl^{\ominus}$

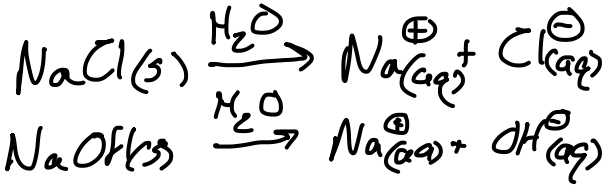
not a base

e⁻'s near to lower positive charge

weak base

same(ish) distance more ⊕ nucleus

less reactive, more stable



In general,

$$\text{energy} \propto \frac{(q_1)(q_2)}{r}$$

The Periodic Table Is Your Friend

Sections 1.1 – 1.3

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

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

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

-3 -2 -1

non metals

metals

+1 +2

$Na \rightarrow Na^+ + 1e^-$

$F + 1e^- \rightarrow F^-$

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

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

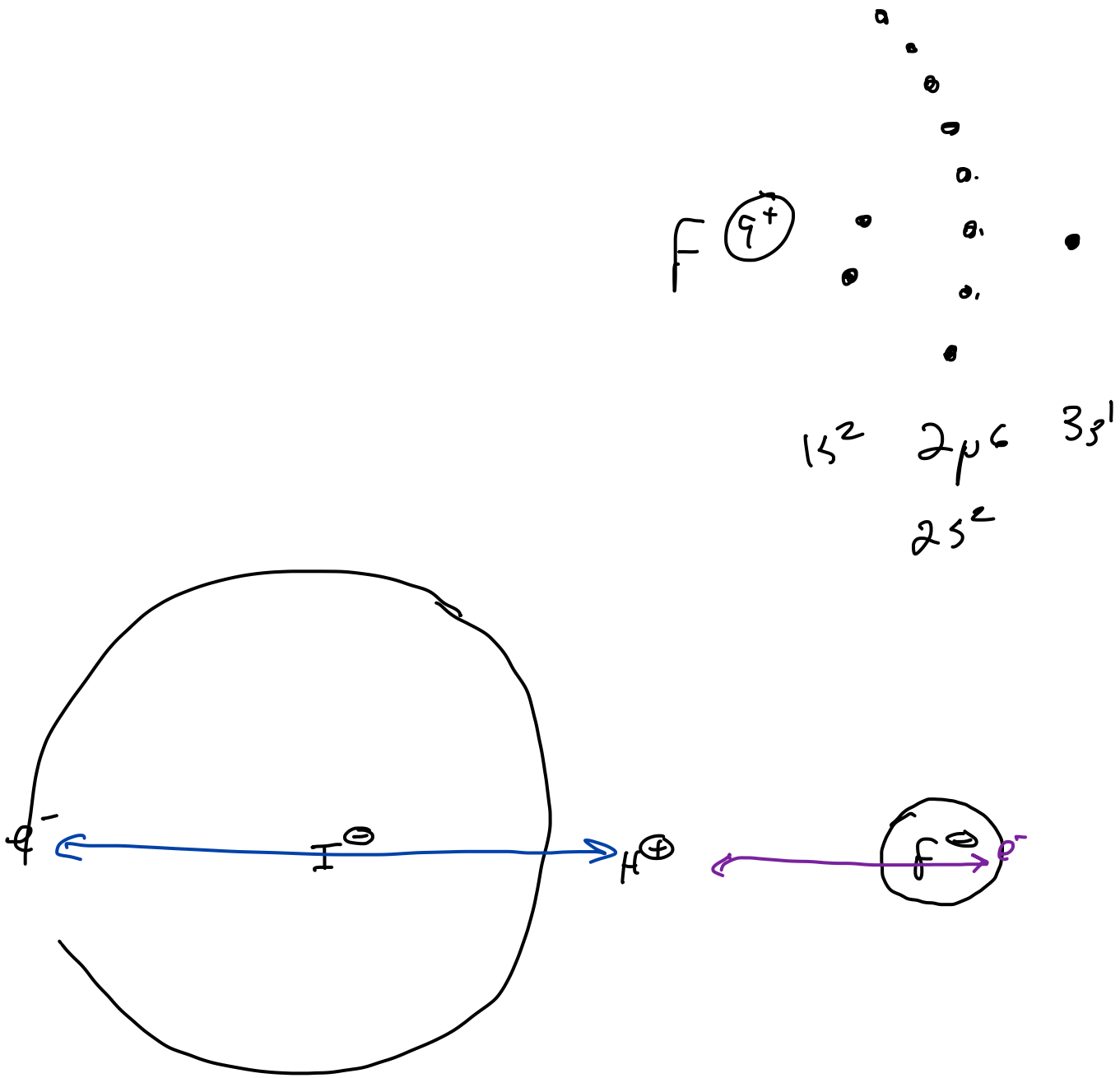
F^{2-} $1s^2\ 2s^2 2p^6\ 3s^1$

past the valence shell

metals tend to lose e⁻'s when they react

non-metals tend to gain e⁻'s when they react

Identify metals and non-metals


F
9+
1s2
2μc
3s1
2s2
e-
I⊖
H⊕
F⊖
e-

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$H + 1e^- \rightarrow H^-$ $1s^2$ 1 bond H-H

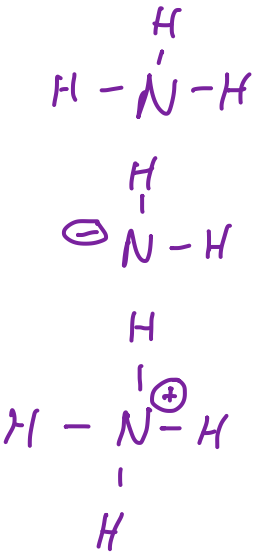
$O + 2e^- \rightarrow O^{2-}$ 2 bonds H-O-H

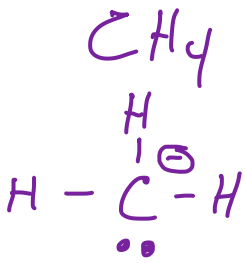
1 bond $^{\ominus}O-H$

3 bonds $H-\ddot{O}^{\oplus}-H$ (H)

$H-O-H + H^{\oplus}$

$N + 3e^- \rightarrow N^{3-}$ 3 bonds





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

no 5 2e⁻ bonds to C

Ionic Interactions, Polar Bonds, and Nonpolar Bonds

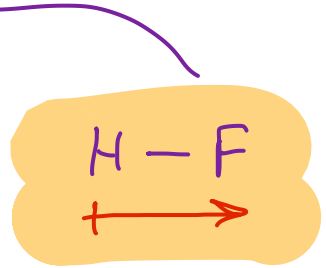
Sections 1.1 – 1.3

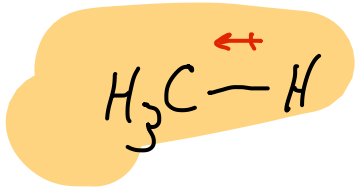
metal nonmetal

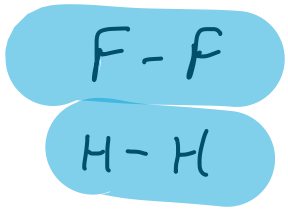
polyatomic ion

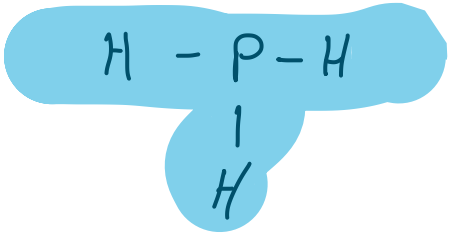
polyatomic ion

$NH_4^{\oplus}$ $OH^{\ominus}$









Chemists use different drawings to place emphasis on different aspects of a molecule.

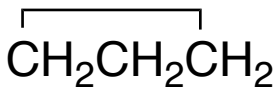
Remember the basics of Lewis Structure (we will practice drawing them as a lab activity)

C_3H_9O $CH_3OCH_2CH_3$ $CH_3CH_2OCH_3$ $CH_3CH_2CH_2OH$ $CH_3CHOHCH_3$

In organic, condensed structures typically start with a C, and everything immediately to the right of the C is connected to that first C. When the the first C is finally connected to the second C, now that atoms right of the second C are connected to second C. In acyclic, molecules atoms to the right of the second C are never connected to the first C.

CH_2CHCH_3

Because bonds are not drawn, condensed structures require the reader to bring some chemical knowledge to their interpretation.


$CH_2CH_2CH_2$

$CH_3CH_2CH(CH_3)_2$ $CH_3CH(OH)CH_2CH_3$ $CH_3(CH_2)_3CH_3$

Parentheses **( )** in structures are typically used for to set off side chains, to indicate a repeating unit, or to indicate multiple groups of the same structure.

$CH_3CH(CH_3)_2$ $CH_3CH(OH)CH_2CH_3$ $CH_3(CH_2)_3CH_3$ $CH_3COCH_2CH_3$

Parentheses **( )** in structures are typically used for to set off side chains, to indicate a repeating unit, or to indicate multiple groups of the same structure.

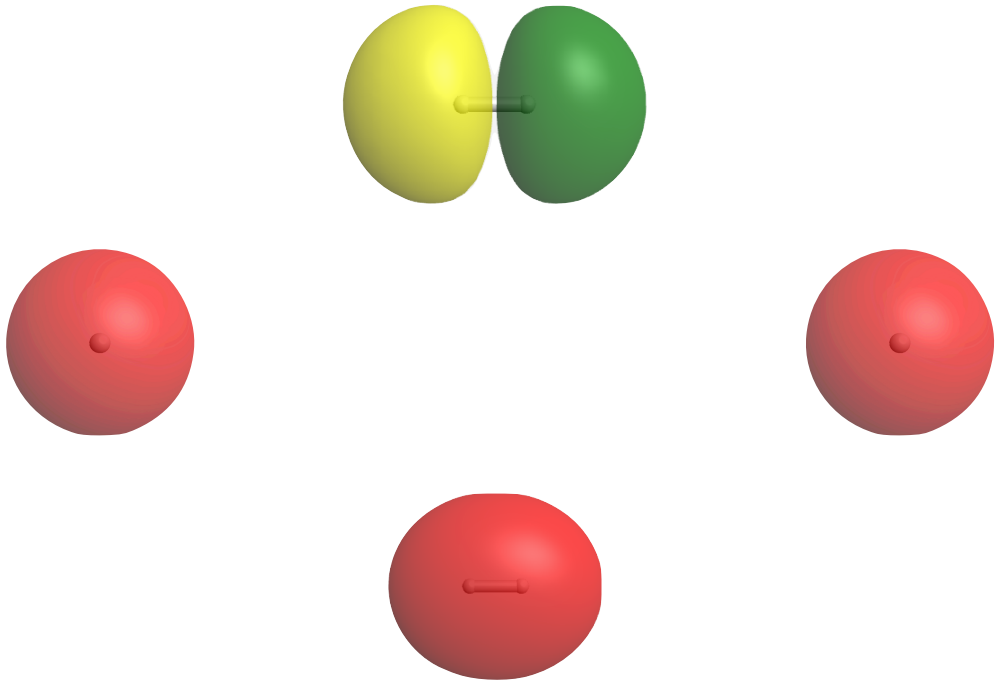
Sometimes chemists omit parentheses when they are not absolutely necessary, and sometimes chemists do things for aesthetic reasons.

Because bonds are not drawn, condensed structures require the reader to bring some chemical knowledge to their interpretation.

Chemists regularly leave out parentheses for small side chains and often change conventions for aesthetic reasons

$CH_3CHOHCH_2CH_3$ $CH_3C(O)CH(CH_3)_2$ CH_3CHO

Different structures serve different purposes, but they represent the same things



Molecules have orbitals just like atoms have orbitals