(1) Today

Attendance

Review Syllabus

Sections 1.1 – 1.3
atomic structure
electrons, valence vs core electrons

Reviewing Periodic Trends

Section 1.4
Introduction to Chemical Bonding Theories
octet rule etc

Next Class (2)

Section 1.4
Introduction to Chemical Bonding Theories
octet rule etc

Sections 1.5-1.10
Valence Bond Theory

(3) Second Class from Today

Sections 1.5-1.10
Valence Bond Theory

Skipping Section 1.11 for now
An introduction to Molecular Orbital Theory

Sections 1.12
Drawing Chemical Structures

Third Class from Today (4)

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

Sections 2.4 – 2.6
Resonance/Electron Delocalization

Appropriate Problems from McMurry Chap 1

Section 1.3 Problems 1-1 and 1-3

Section 1.4 Problems 1-4 through 1-7, 1-24 through 1-28 1-33

Section 1.5 - 1.10 Problems 1-8, 1-9, 1-10, 1-11, 1-12, 1-13, 1-14, 1-34 – 1-40, 1-48, 1-49, 1-50 (part a is asking how are they similar electronically), 1-56

Section 1.12 Problems 1-15 through 1-21, 1-41 through 1-44, 1-52, 1-53, 1-54, 1-55

Challenging Problems 1-45, 1-46, 1-47, 1-51

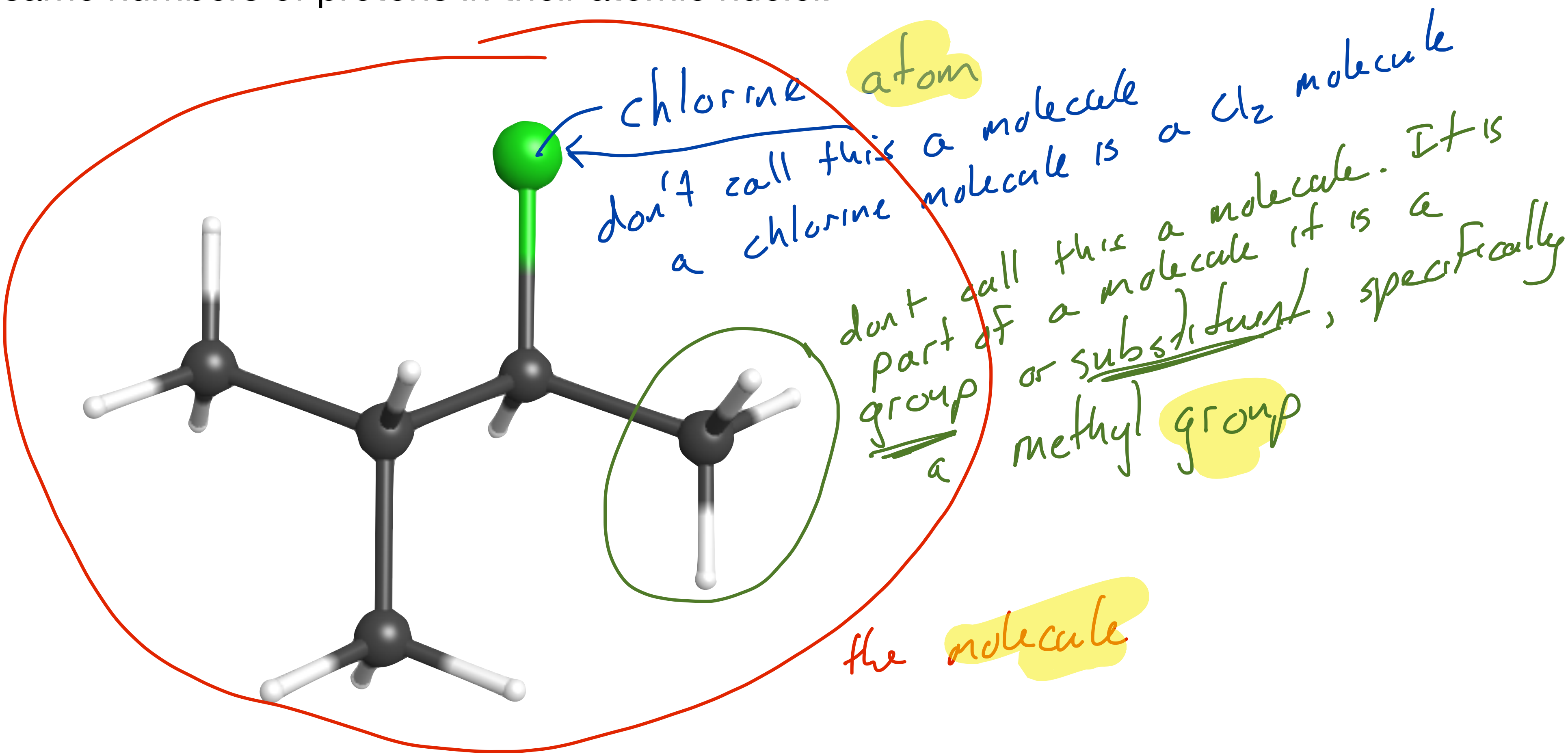
Atoms, Elements, Molecules, and Substituents or Groups

Cl_2 molecule

Cl

A diversion into the language of chemistry…

“In chemistry, an element is a pure substance consisting only of atoms that all have the same numbers of protons in their atomic nuclei.”[1]

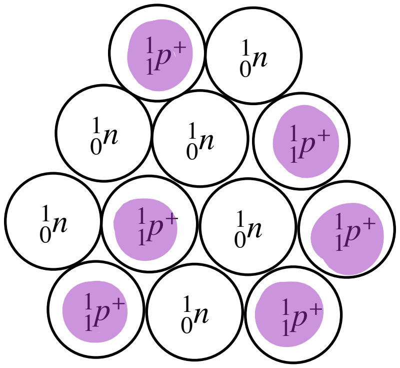


[1] https://en.wikipedia.org/wiki/Chemical_element accessed September 3, 2021

Three isotopes of C are represented below

most of the

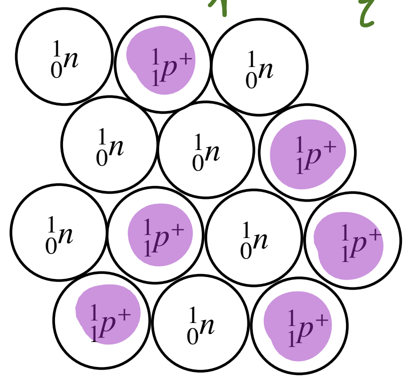
spin ½ nuclei



6 protons
6 neutrons

carbon on Earth

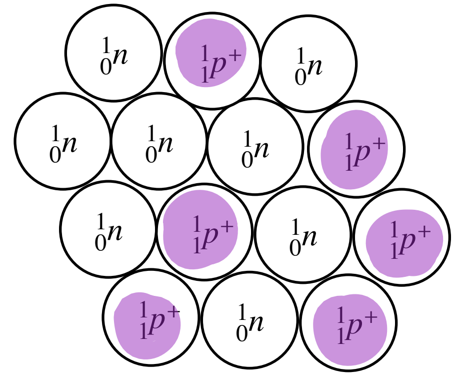
^{12}C



6 protons
7 neutrons

tiny little magnets

^{13}C



6 protons
8 neutrons

radioactive
radiocarbon dating

^{14}C

Chemistry is controled by the # of protons in nucleus
so all of these isotopes do the same chemistry

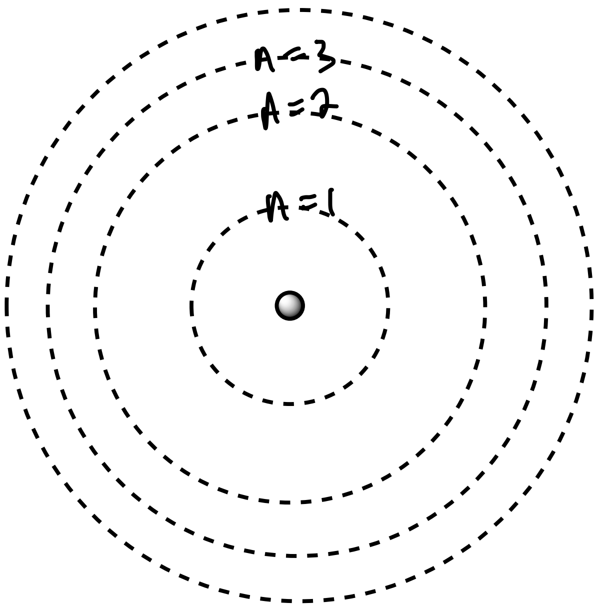
Differences? ... Heavy ones react more slowly
Nuclear properties are different

Remember the structure of an atom

And Where Are the Electrons Again?

Sections 1.1 – 1.3

Bohr



First person to model the atom with quantized energy levels starting from basic particle physics; e.g.,

E = KE + PE   or $E = 1/2\ mv^2 + Ze^2/r$

centripetal force = mv^2/r

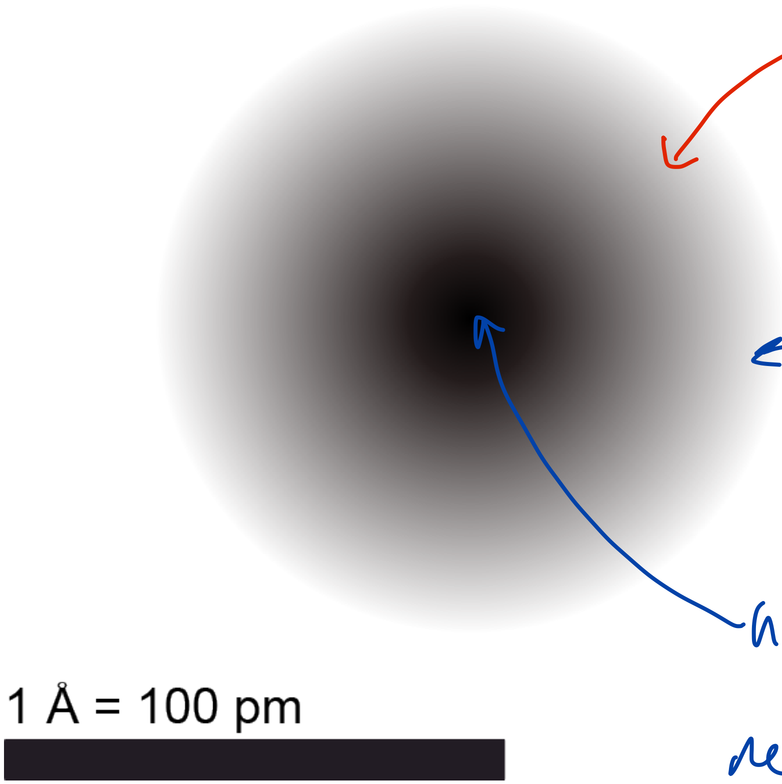
force of attraction between charged particles = Ze^2/r^2

Only worked for atoms with one electron.

Also it is physically impossible for electrons to orbit a nucleus like the Moon orbits the Earth…. The electrons would radiate energy and crash into the nucleus.

And Where Are the Electrons Again?

Wave/Quantum Mechanical Model



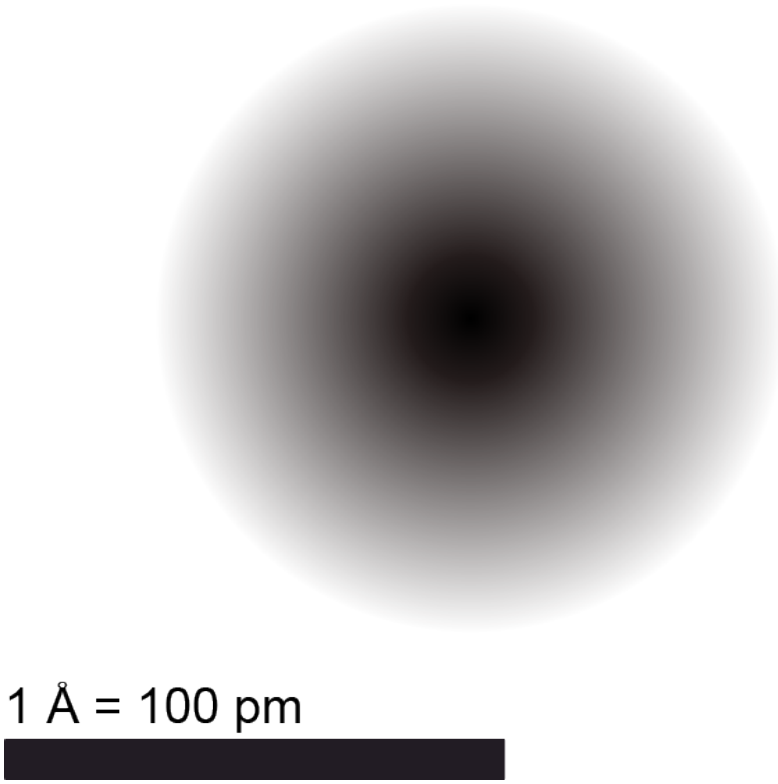
e^-'s are somewhere in here.

lower probability further out

high probability of finding e^- near the nucleus

Image derived from https://en.wikipedia.org/wiki/Atom#/media/File:Helium_atom_QM.svg

Wave/Quantum Mechanical Model



Bohr had 1 quantum number.

the electron is in the n = 1 or 2 or 3 or 4… shell

Quantum Mechanics requires four quantum numbers to describe an electron: n, l, m_l, and m_s.

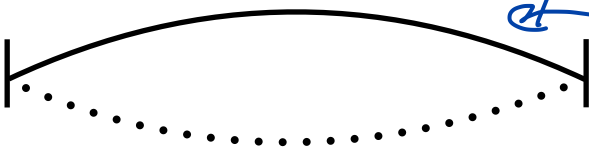
n is the principal energy level
l describes the shape of the orbital
m_l describes the orientation of the orbital and
m_s indicated the spin of the electron.
$+\frac{1}{2}$ or $-\frac{1}{2}$

Further, as ***n*** gets larger more orbital shapes become available and as more shapes become available more orientations become possible

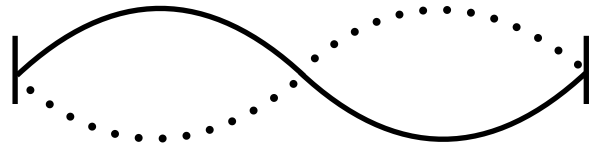
Image derived from 1https://en.wikipedia.org/wiki/Atom#/media/File:Helium_atom_QM.svg

Wave/Quantum Mechanical Model

standing waves

n = 1 

lower E
lower frequency

n = 2 

higher E
higher frequency

The H atom's only electron:

n = 1, l = 0, m_l = 0, and m_s = $^1/_2$

He's two e^-'s:

all e^-'s require four numbers to describe its energy

n = 1, l = 0, m_l = 0, and m_s = $^1/_2$
n = 1, l = 0, m_l = 0, and m_s = -$^1/_2$

l – tells you the shape
m_l – which way the orbital points
n – tells you how many orbitals will be available

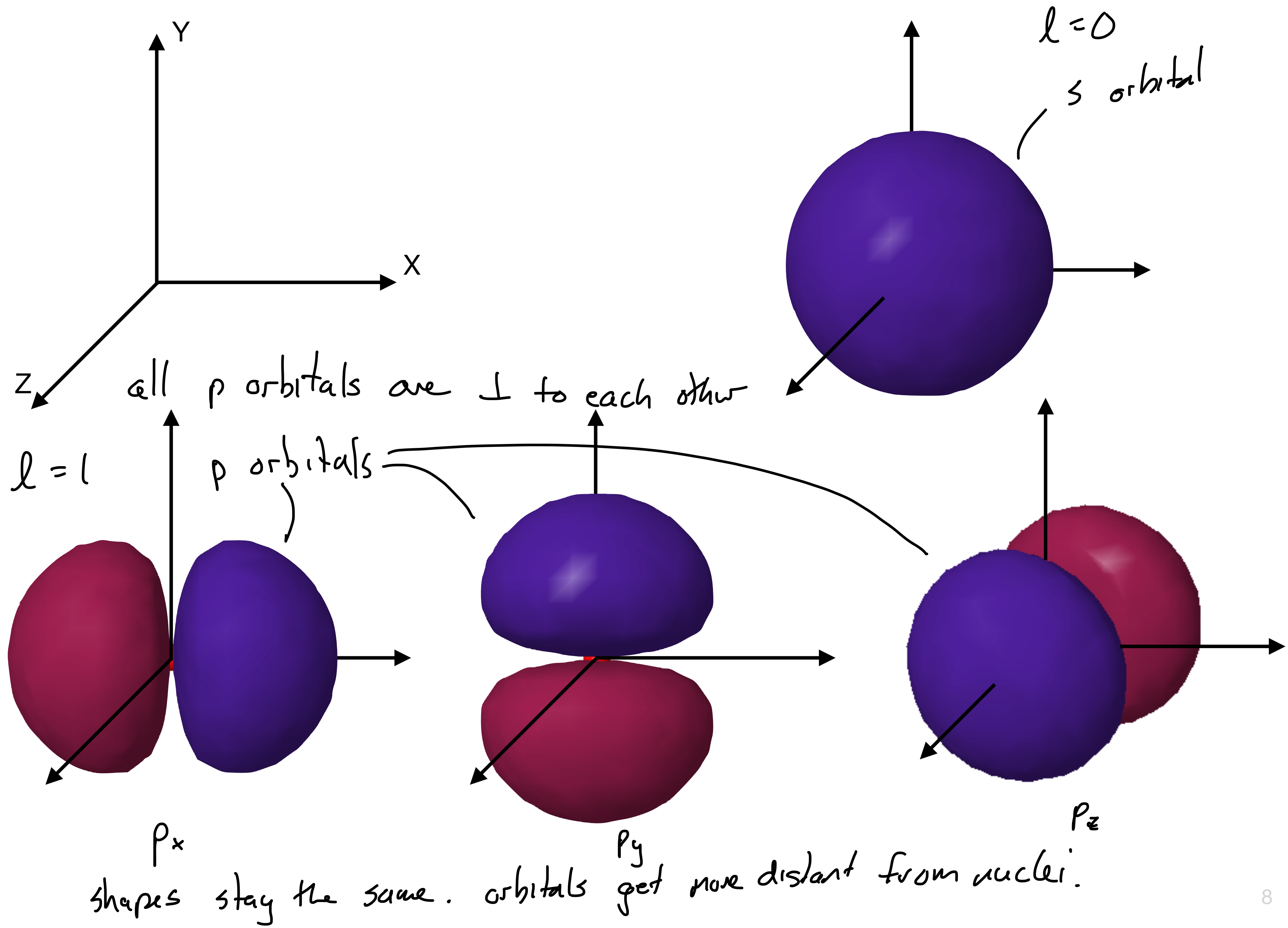
B's five e^-'s:

n = 1, l = 0, m_l = 0, and m_s = $^1/_2$ — l = 0 = s, n = 1 } 1s
n = 1, l = 0, m_l = 0, and m_s = -$^1/_2$
n = 2, l = 0, m_l = 0, and m_s = $^1/_2$ — l = 0, n = 2 => 2s
n = 2, l = 0, m_l = 0, and m_s = -$^1/_2$
n = 2, l = 1, m_l = 1, and m_s = $^1/_2$ — l = 1 = p => 2p

3 p orbitals per shell

Models

And Where Are the Electrons Again?



The Periodic Table Is Your Friend

arrangement is based on Quantum mechanics

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

n=1
n=2
n=3
n=4

H e⁻ config is $1s^1$

Li is $1s^2 2s^1$

Be is $1s^2 2s^2$

He $1s^2$

C $1s^2 2s^2 2p^2$

P $1s^2 2s^2 2p^6 3s^2 3p^3$

p block

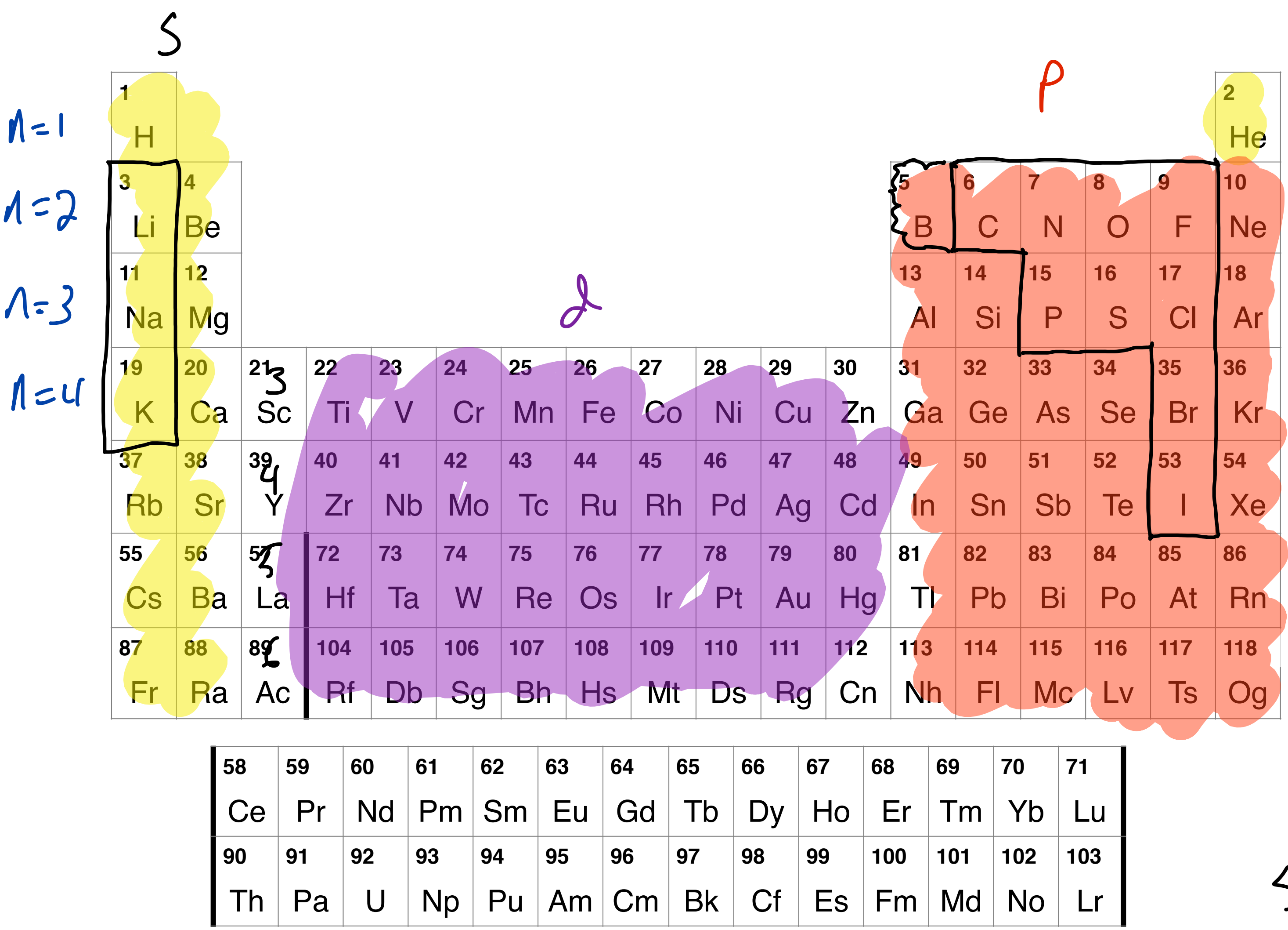
p block – last e⁻'s go in p orbitals

d block

f block

Remember how electrons are distributed, electron configuration
Remember the importance of valence electrons/the valence shell

Example Electron configurations



Si

electron config and # of valence e⁻'s?

C $1s^2$ $2s^2 2p^2$

Si $1s^2$ $2s^2 2p^6$ $3s^2 3p^2$

valence e⁻

short hand is noble gas + valence e⁻

C e⁻ config [He] $2s^2 2p^2$

Si e⁻ config [Ne] $3s^2 3p^2$

# of valence e⁻'s in O   2+4=6 valence e⁻

e⁻ config [He] $2s^2 2p^4$

or the # of valence e⁻'s can be determined by counting accross the table (skip the d's)

when talking about bonding + reactivity, we concentrate on valence e⁻ config + principal energy level (the n #)

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

Use the periodic table to predict likely bond formation

Introduce Valence Bond Theory (hybridization)

The Periodic Table Is Your Friend

Review

carbon will share e^-
these others can too
can form ions

+1 | +2 | ? | -4 | -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	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

metal

tend to lose e^-'s when reacting with non metals

non metal tend to gain or share e^-

much easier to remove 1 e^-

3^+ Li $1s^2$ | $2s^1$ ← this e^- is attracted to a +1 charge

+3 -2 | +1

core e^-'s

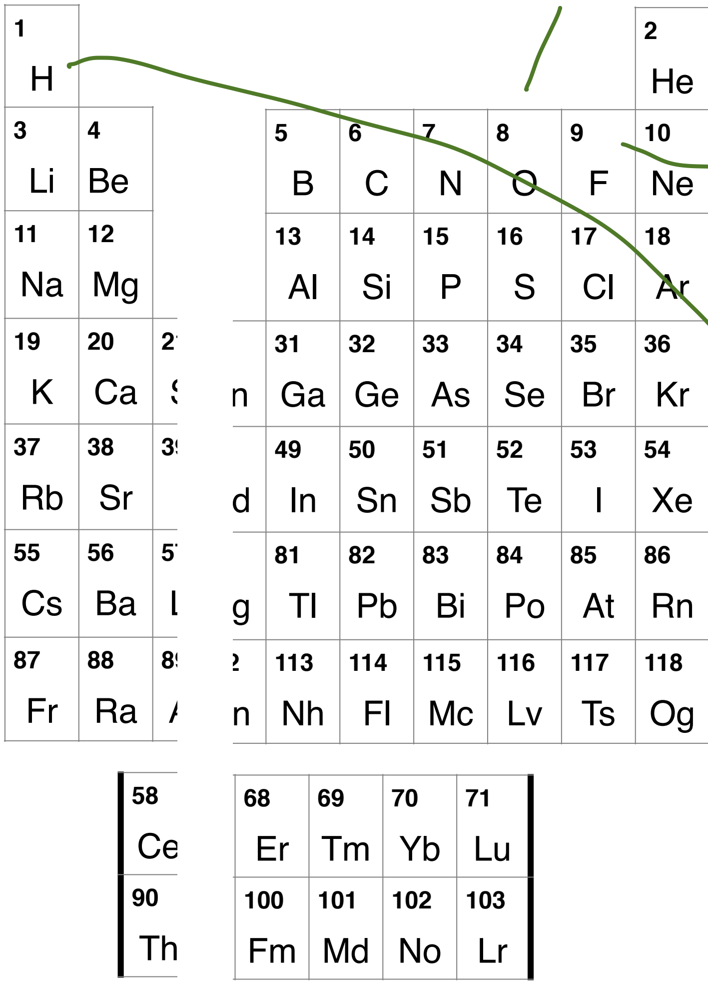
9^+ F $1s^2$ | $2s^2 2p^5$ ← each of these e^-'s is attracted to a +7 charge

+9 -2 | +7

core e^-'s

too hard to remove e^-

Identify metals and non-metals



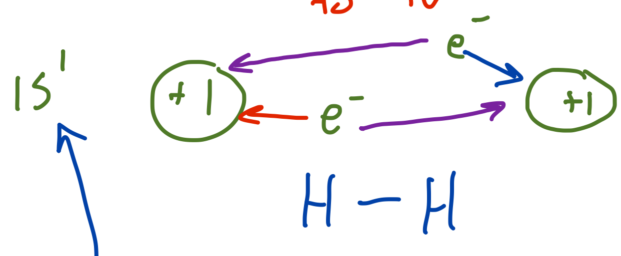
how many bonds can O make? room for 2 e^-'s ... O tends to make 2 bonds

how many bonds can F make? room for 1 e^-. F tends to make 1 bond

how many bonds can H make? an H atom's 1s orbital has room for 1 more e^-

H•

both e^-'s get to be attracted to the +1 nucleus



room for 1 more e^-

mutual attraction for a pair of e^-'s ... covalent bond

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

room for 1 e^- means room to form 1 covalent bond