

Ch. 7 NOTES ~ Completing the Model of the Atom

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics*.

7.1 Notes

I. Developing a Model of Atomic Structure

A. Review of atomic theory development

- 1) ancient Greece: *Democritus of Abdera*
 - *proposed the existence of atoms (atoma)*
 - *went against Aristotle's idea of continuous matter (hyle)*
- 2) 1803: *Dalton's atomic theory*

DALTON'S ATOMIC THEORY
1) All elements are composed of submicroscopic, indivisible particles called atoms. (He didn't know about subatomic particles and how to split an atom.)
2) Atoms of the same element are identical. (He didn't know about isotopes.) Atoms of different elements are different.
3) Atoms of elements can physically mix or form compounds by chemically combining in whole-number ratios. (Law of Multiple Proportions)
4) Chemical reactions involve the separation, joining, or rearranging of atoms. Atoms of an element are never changed into atoms of another element in a chemical reaction. (He didn't know about nuclear reactions.)

- 3) 1897: *Thomson's plum pudding model*
 - *electrons embedded in a positive ball*
- 4) 1909-1911: *Rutherford: gold foil experiment*
 - *the atom is mostly space*
- 5) 1913: *Bohr's planetary model of the atom*
 - *nucleus in the center, electrons in orbit around it*
- 6) 1935+ : *modern 3-D electron-cloud model; probability model*

B. review of valence electrons and Lewis structures (electron dot diagrams)

- 1) **valence** = *outermost electron shell or level*
- 2) valence electrons are exposed and *available for bonding*
- 3) *atoms will try to become like Noble Gases*
(*they don't change what's in the nucleus, but they change the valence shell*)
 - a) *gain electrons to become anions (ionic bonding) OR*
 - b) *lose electrons to become cations (ionic bonding) OR*
 - c) *share electrons with other atoms (covalent bonding)*

UPDATED ELECTRON INFORMATION TABLE:

GROUP #	# VAL. e-	# DOTS	STATUS	ACTION	IONIC CHARGE
IA (1)	1	1	1 / 8	lose 1	+ 1
IIA (2)	2	2	2 / 8	lose 2	+ 2
IIIA (13)	3	3	3 / 8	lose 3	+ 3
IVA (14)	4	4	4 / 8 , half full	lose, gain, or neither	varies ("mixed")
VA (15)	5	5	5 / 8	gain 3	- 3
VIA (16)	6	6	6 / 8	gain 2	- 2
VIIA (17)	7	7	7 / 8	gain 1	- 1
VIIIA (18)	8	8	8 / 8, full	nothing	0: does not form ions

“Charge Chant”: +1 +2 +3 mixed -3 -2 -1 0
 +2 in the middle, unless they tell you otherwise

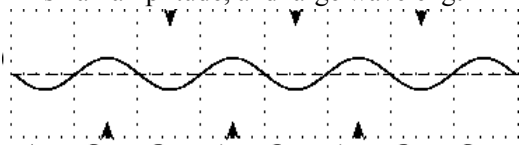
C. review of **em (electromagnetic) radiation**

1) em radiation includes

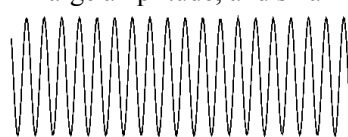
RADIO WAVES, MICROWAVES, INFRARED (IR) RADIATION, VISIBLE LIGHT
 (“ROYGBIV”: red, orange, yellow, green, blue, indigo, violet), *ULTRAVIOLET (uv), X-RAYS, GAMMA RAYS*

2) review of wave dimensions

Slow wave with low energy, low frequency, small amplitude, and large wavelength



Fast wave with high energy, high frequency, large amplitude, and small wavelength



3) review **WAVE-PARTICLE DUALITY OF NATURE:**

waves can act as particles AND particles can act as waves

II. Distribution of Electrons in Energy Levels

A. **atomic orbitals**

- 1) orbitals are *areas of high probability (~95%) of finding electrons*
- 2) **Heisenberg Uncertainty Principle**—*it is not possible to know both the exact position and velocity of an object simultaneously*
 - a) if you measure the position, you have to stop it (technically), so you cannot find the velocity or momentum
 - b) if you measure the velocity, it is moving, so you cannot pinpoint one position to measure

B. **principal quantum number** = n

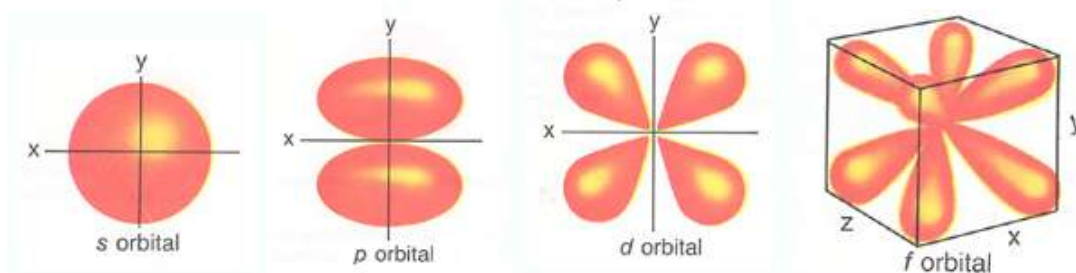
- 1) *electron energy level or “shell” number*
- 2) $n = 1, 2, 3, 4, 5, 6, 7$

C. energy **sublevels**—*divisions within a level*

- 1) many similar energy states grouped together in a level
- 2) different shapes: spherical, dumbbell, cloverleaf
- 3) **sublevels: s, p, d, f**
 (s p d f stand for sharp, principal, diffuse, fundamental)
- 4) *maximum number of e^- in a principal energy level = $2n^2$*

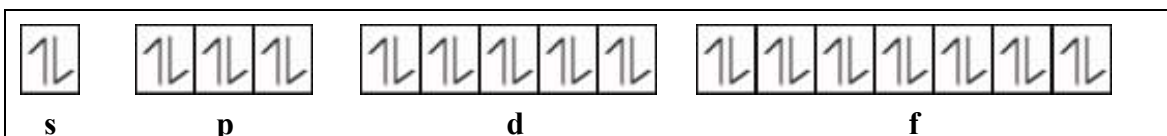
<u>n</u>	<u>SUBLEVELS</u>
n = 1	1s
n = 2	2s, 2p
n = 3	3s, 3p, 3d
n = 4	4s, 4p, 4d, 4f
n = 5	5s, 5p, 5d, 5f
n = 6	6s, 6p, 6d – where the new elements are being placed
n = 7	as of now: 7s

SHAPES OF ORBITALS from <http://boomeria.org>



D. electron capacity of the sublevels

- 1) s can hold 2 e⁻
- 2) p can hold 6 e⁻
- 3) d can hold 10 e⁻
- 4) f can hold 14 e⁻



- E. orbital = one "box;" can also be a line or a small circle
- F. electrons = "arrows"
- G. **antiparallel spin**—electrons in an orbital spin in opposite directions, so as not to repel one another (this is shown as arrows pointing in opposite directions)

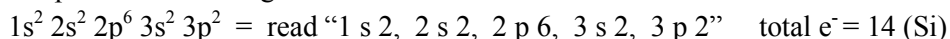
III. Electron Configurations

A. **electron configuration**—specific e⁻ arrangement in orbitals

B. rules

- 1) Aufbau Principle—e⁻ enter low-energy orbitals first
- 2) Pauli Exclusion Principle—no more than two e⁻ per orbital; the e⁻ must have compatible spins (shown by ↑ and ↓)
- 3) Hund's Rule: *degenerate orbitals (orbitals of equal energy)* have one e⁻ filling each orbital before pairs can exist

C. sample electron configuration:



D. the sum of the superscripts equals the atomic number of the element

E. superscripts are NOT exponents (nothing is being squared, etc.)

F. *** valence configurations will be s OR s and p ***

G. ways to do electron configurations

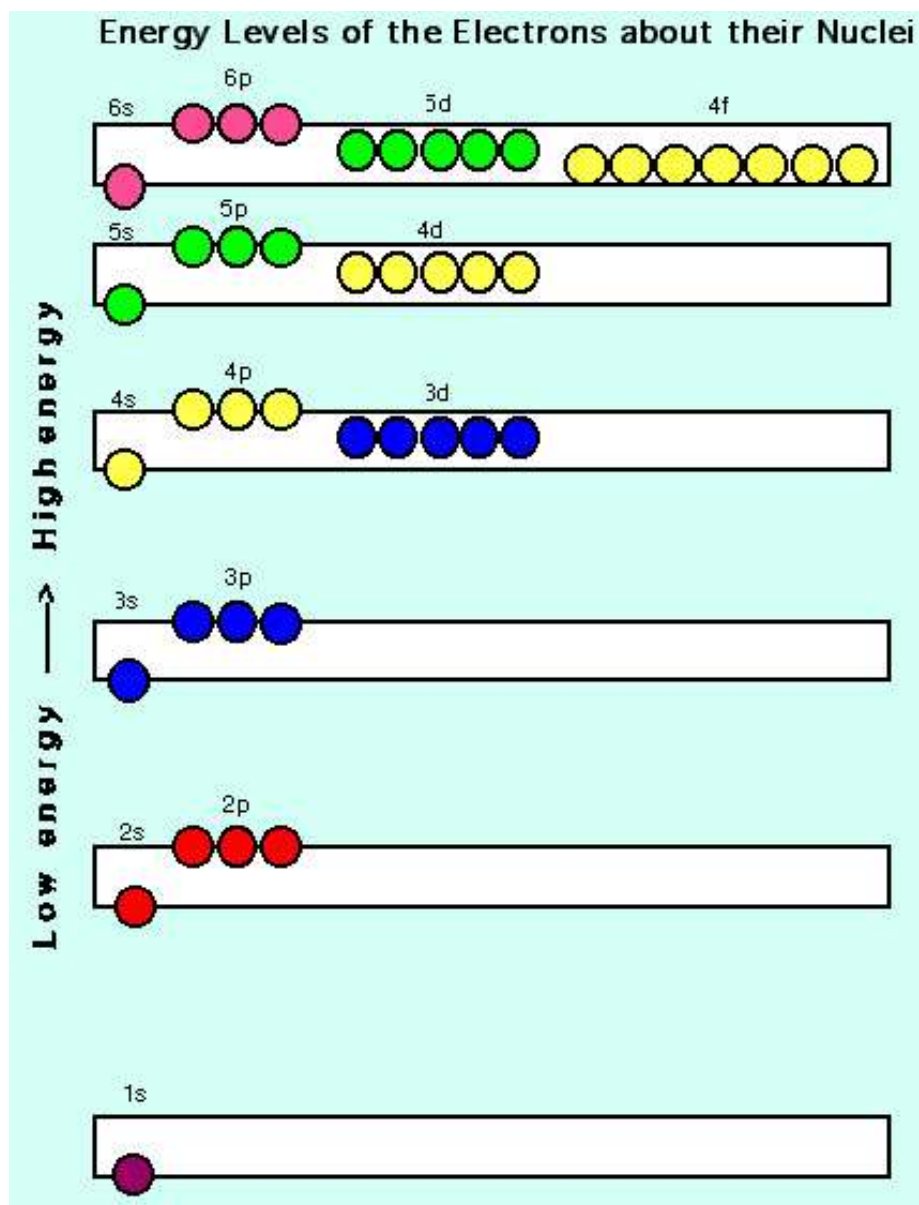
- 1) the Aufbau diagram can be used to determine electron configurations
- 2) the placement of elements in the periodic table shows electron configuration

Be able to do electron configurations using both methods.

H. Examples with the Aufbau diagram

- E1) Write out the electron configuration for nitrogen.
- E2) Write the electron configuration for cadmium.

ANSWERS: E1) $1s^2 2s^2 2p^3$
 E2) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$



7.2 Notes

IV. The Periodic Table and Atomic Structure

A. electron configurations according to location

- 1) Group IA (1) and IIA (2) = end in s
- 2) Group IIIA-VIIIA (13-18) = end in p
- 3) B Groups (Groups 3-12 transition metals) = end in d
- 4) Inner transition metals = end in f
- 5) **Noble gas configuration** = $s^2 p^6$
- 6) **Pseudo-Noble gas configuration** (an exception; see VI. C) = $s^2 p^6 d^{10}$

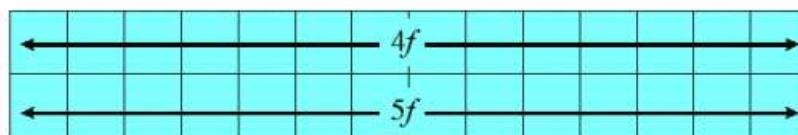
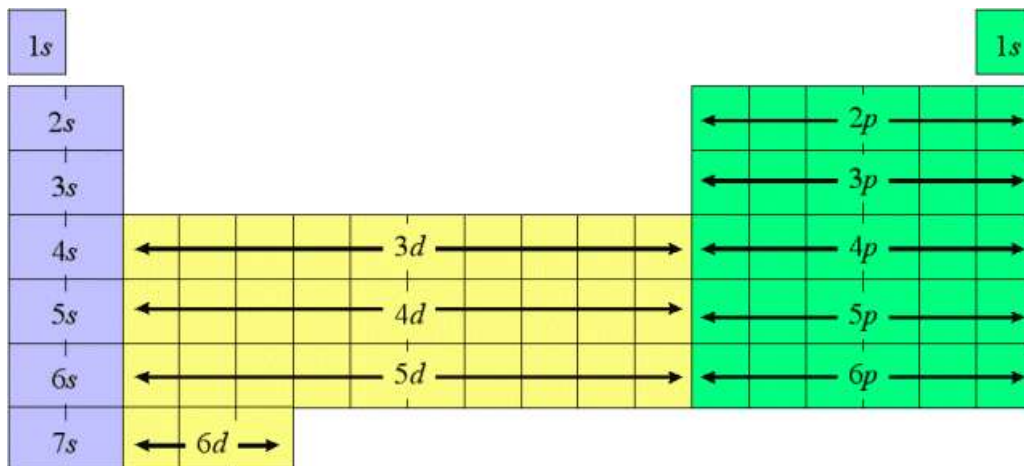
B. Complete vs. valence electron configurations

- 1) *complete*—lists all orbitals (superscripts match the atomic number)
- 2) *valence*—only lists the outer shell (matches the dot diagram)
DOTS: "right, left, up, down, top all the way around."

C. Why are valence configurations always s OR s/p and not normally include d or f?

- 1) *d orbitals* are "step-downs" – they are one level below the valence
- 2) *f orbitals* are "two step-downs" – they are two levels below the valence

ELECTRON CONFIGURATIONS AND THE PERIODIC TABLE (from FSU)



GROUP #	VALENCE #	SAMPLE VALENCE CONFIG.	# DOTS
IA (1)	1	$3s^1$	1
IIA (2)	2	$3s^2$	2
IIIA (13)	3	$3s^2 3p^1$	3
IVA (14)	4	$3s^2 3p^2$	4
VA (15)	5	$3s^2 3p^3$	5
VIA (16)	6	$3s^2 3p^4$	6
VIIA (17)	7	$3s^2 3p^5$	7
VIIIA (18)	8	$3s^2 3p^6$	8

V. *Exceptions* to electron configurations

- A. to *fill* a sublevel = stable
- B. to *half-fill* a sublevel = stable
- C. new unpredicted configurations are observed

Cu, expected: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$

Cu, observed: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

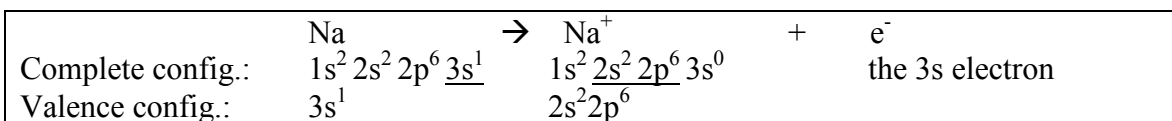
VI. Electron Configurations for Ions

Ions become **isoelectronic** (*having the same configuration as*) with the Noble Gases.

A. **cation** examples (s^2p^6)

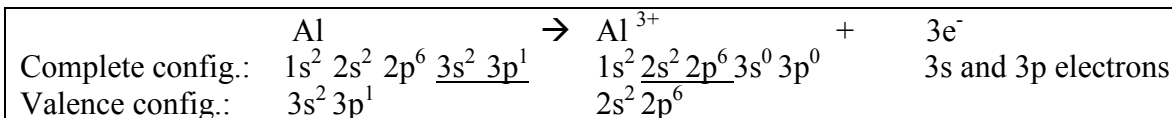
** *Cations lose electrons to become isoelectronic with the previous Noble Gas.* **

- 1) Na has one valence e^- . It forms a 1+ ion, exposing its filled $n = 2$ shell.



The sodium ion has the electron configuration of neon (10 e^-), but it still has a nucleus with 11 protons, keeping it sodium.

- 2) Al has three valence e^- . It forms a 3+ ion, exposing its filled $n = 2$ shell.

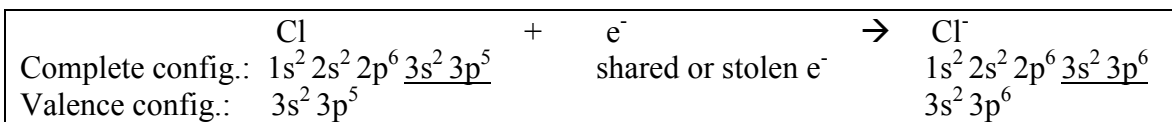


The aluminum ion has the electron configuration of neon (10 e^-), but it still has a nucleus with 13 protons, keeping it aluminum.

B. **anion** examples (s^2p^6)

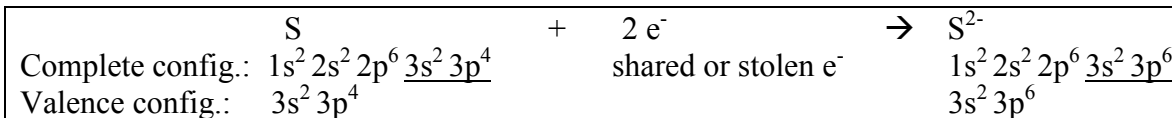
** *Anions become isoelectronic with the Noble Gas at the end of their row.* **

- 1) Cl has seven valence electrons. It forms a 1- ion and has the electron configuration of argon.



The chloride ion has the electron configuration of argon (18 e^-), but it still has a nucleus with 17 protons, keeping it an ion of Cl.

- 2) S has six valence electrons. It forms a 2- ion and has the electron configuration of argon.



The sulfide ion has the electron configuration of argon (18 e^-), but it still has a nucleus with 16 protons, keeping it an ion of S.

