Electronic Configurations

↳ provide a method for describing the distribution of electrons in an atom

Intro: We know the ~~energies~~ energetics of electrons in an atom can be understood by applying Coulomb's Law. We are now going to take that concept of electrons in quantized energy shells and describe the electronic structure of the atom using electron configurations

First Ionization Energies ($^{KJ}/_{mol}$)

H	He	Li	Be	B	C	N	O	F	Ne	Na
1312	2372	520	900	801	1087	1402	1314	1681	2081	496

exception 1 (Be → B) exception 2 (N → O)

Purple Bars:

He → Li 3rd e⁻ in Li requires less energy to remove!

(*) Using Coulomb's Law the distance from the nucleus for that 3rd e⁻ had to increase

↳ starts to point to e⁻ being in shells

Ne → Na 11th e⁻ in Na requires less energy to remove!

↳ 11th e⁻ in Na must be farther from nucleus than 10th e⁻ in Ne

(*) AS WE INCREASE THE NUMBER OF ENERGY SHELLS THEY GET FARTHE FROM NUCLEUS

↳ results in decreased attraction to nucleus

Li → Ne [Elements in Same Period; VE in same energy shell; Shielding is constant; Nuclear charge is increasing]

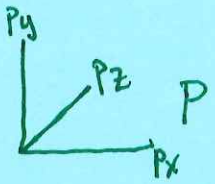
↳ leads to general increase in IE Accross the period

Shielding - occurs because core electrons are generally closer to nucleus than valence electrons so core e^- shield valence e^- from nucleus and Add to electron - electron Repulsion

exception 1: Be → B.

S orbital e^- tend to be closer to nucleus than p orbital e^-

↳ Coulombic interactions predict p orbital e^- less tightly held and so should require less energy to remove

Orbitals		Max e^-
s		$2e^-$
p	 	$6e^-$
d	5 orbitals	$10\ e^-$
f	7 orbitals	$14\ e^-$

exception 2: N → O

All e^- are in p orbitals but there are 3 physically defined regions where e^- reside.

Oxygen is when 1st e^- pairing occurs in p region which increases electron - electron repulsion

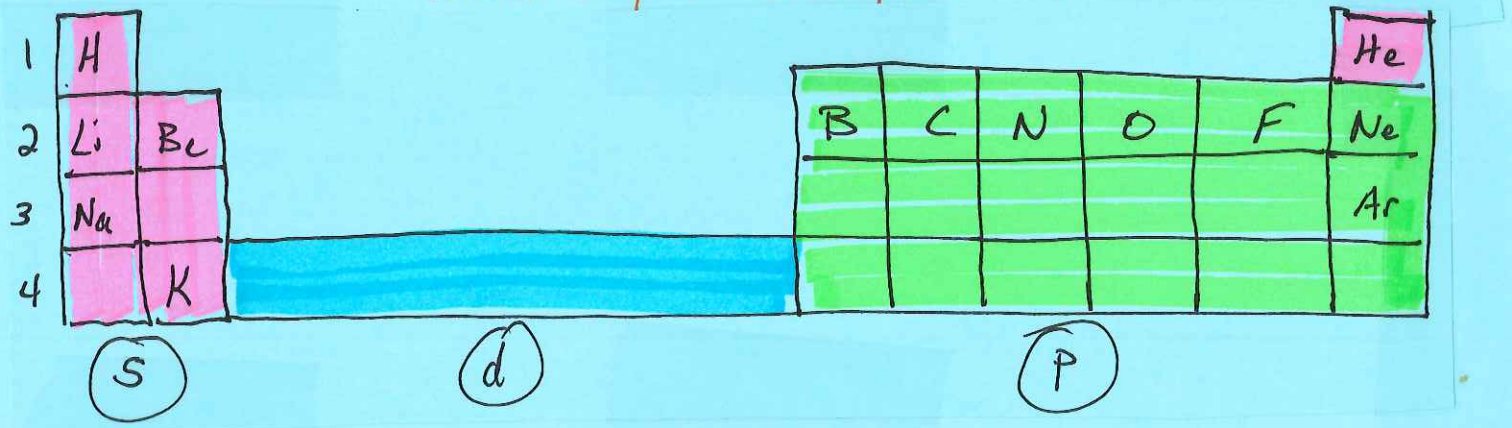
↳ increasing repulsion means less energy needed to overcome attraction to nucleus

Rules for writing electron Configurations:

① Aufbau Principle - e⁻ added to atom will fill lowest energy shell / subshell available

② Pauli's exclusion Principle - orbitals can hold 2e⁻ max that must have opposite spin (+ or −). Since no two e⁻ may have the same principle quantum numbers

③ Hunds Rule - degenerate orbitals are filled one e⁻ at a time before any e⁻ are paired in orbitals



Full → exponents denote e⁻

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

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

Orbital Notation ← arrows denote e⁻

Na ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑ (2p) ↑ (3s)

K ↑↓ (1s) ↑↓ (2s) ↑↓ ↑↓ ↑ (2p) ↑↓ (3s) ↑↑↑ (3p) ↑↓ (4s)

Noble Gas ← core e⁻ summarized with noble gas

Na [Ne] $3s^1$

K [Ar] $4s^2$

Paramagnetic - has unpaired e⁻ and will interact with a magnetic field

diamagnetic - NO unpaired e⁻, 2 e⁻ in every orbital filled. Will not interact with magnetic field