

Electron Configurations & The Periodic Table

Define:

Quantum numbers: Numbers that specify a property of an orbital or an e^-
 $n = \text{Principal} \Rightarrow \text{pos. integer } (1, 2, 3, \dots)$ (energy, size) $m_s = \text{Spin} \Rightarrow +\frac{1}{2} \text{ or } -\frac{1}{2}$
 $l = \text{angular momentum} \Rightarrow \text{integer from } 0 \text{ to } n-1$ (orbital shape #)
 $m_l = \text{magnetic} \Rightarrow \text{integer from } -l \text{ to } +l$ (orientation of orbital) (e^- spin direction)

exclusion principle: No two e^- in an atom can have the same set of 4 quantum numbers. This arises from the fact that an orbital has a max occupancy of $2e^-$ & their spins are paired.

shielding: The ability of other e^- , especially inner ones, to lessen the nuclear attraction for an outer e^- .

effective nuclear charge: $Z_{\text{eff}} \rightarrow$ The nuclear charge an e^- actually experiences as a result of shielding effects due to the presence of other e^-

electron penetration: The process by which an outer e^- moves through the region occupied by core e^- to spend part of its time closer to the nucleus. Penetration increases average Z_{eff} for that e^- .

aufbau principle: from German "auf bauen" (to build up). Conceptual basis of a process of building up atoms by adding one proton (and one or more neutrons) at a time to the nucleus & one e^- around it to obtain the ground state e^- configurations of the elements.

electron configuration: The distribution of e^- within the orbitals of atoms of an element. Also, the notation for the distribution.

orbital diagram: A depiction of e^- number & spin in an atom's orbitals by means of arrows in a series of small boxes or lines.

Hund's rule: When orbitals of equal energy are available, the e^- configuration of lowest energy has the max number of unpaired e^- w/ parallel spins. I.e. e^- will fill single before they spin pair.

Rank the orbital sublevels in order of lowest energy to highest energy.

$$s < p < d < f$$

What is the correlation between l value and sublevel energy?

For a given n value, the lower the l value, the lower the sublevel energy. I.e. $s \rightarrow l=0$

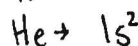
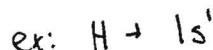
Writing electron configurations - two common ways to designate the orbital and its electrons:

1. e^- configuration. Written $n l^{\#}$ where

$n = \text{principal energy level}$

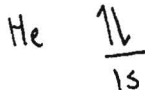
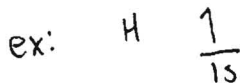
$l = \text{letter designation of sublevel}$

$\# = \text{number of } e^- \text{ in the sublevel}$



2. Orbital diagram

Box or lines for each orbital in an energy level, grouped by sublevel, w/ arrow indicating spin.



Convention s: 1. Draw spin up arrows first ($m_s = +\frac{1}{2}$)

2. fill left to right

(fill $m_l = -1$ before $m_l = 0$ for, say, an $l=1$ p orbital)

Relating quantum numbers and electron configurations.

For the first 10 elements of the periodic table, give Z , the values for all 4 quantum numbers of the *last* electron added to the element, and its orbital diagram.

* Note: Some people find it easier to write the configuration first, then the orbital diagram, others prefer to diagram then write configuration. It is up to you which way you find makes most sense.

He: $Z=2$
 $1s^2$
 $\uparrow\downarrow$
 $1s$
 $n=1, l=0$
 $m_l=0, m_s=-\frac{1}{2}$

H: $Z=1$ $1s^1$ $\frac{1}{1s}$ $n=1, l=0, m_l=0, m_s=+\frac{1}{2}$

* Energy of sublevels increases from left to right

Li: $Z=3$ $1s^2 2s^1$ $\frac{\uparrow\downarrow}{1s}$ $\frac{1}{2s}$ --- --- --- $2p$
 $n=2, l=0, m_l=0, m_s=+\frac{1}{2}$

Be: $Z=4$ $1s^2 2s^2$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ --- --- --- $2p$ $n=2, l=0, m_l=0, m_s=-\frac{1}{2}$

B: $Z=5$ $1s^2 2s^2 2p^1$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{1}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=-1, m_s=+\frac{1}{2}$

C: $Z=6$ $1s^2 2s^2 2p^2$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{1}{2p}$ $\frac{1}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=0, m_s=+\frac{1}{2}$

N: $Z=7$ $1s^2 2s^2 2p^3$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{1}{2p}$ $\frac{1}{2p}$ $\frac{1}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=+1, m_s=+\frac{1}{2}$

O: $Z=8$ $1s^2 2s^2 2p^4$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{1}{2p}$ $\frac{1}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=-1, m_s=-\frac{1}{2}$

F: $Z=9$ $1s^2 2s^2 2p^5$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{1}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=0, m_s=-\frac{1}{2}$

Ne: $Z=10$ $1s^2 2s^2 2p^6$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ --- --- --- $2p$ $n=2, l=1, m_l=+1, m_s=-\frac{1}{2}$

Looking back to fluorine, write the four quantum numbers associated with its 3rd electron and its 8th electron.

3rd e⁻ → 1 of 2s ⇒ $n=2, l=0, m_l=0, m_s=+\frac{1}{2}$

8th e⁻ → 1 of 1st 2p ⇒ $n=2, l=1, m_l=-1, m_s=-\frac{1}{2}$

Period 3

Write the electron configurations and orbital diagrams for

Na: $Z=11$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{1}{3s}$ --- --- --- $3p$
 $1s^2 2s^2 2p^6 3s^1$

Mg: $Z=12$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{3s}$ --- --- --- $3p$
 $1s^2 2s^2 2p^6 3s^2$

Al: $Z=13$ $\frac{\uparrow\downarrow}{1s}$ $\frac{\uparrow\downarrow}{2s}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{2p}$ $\frac{\uparrow\downarrow}{3s}$ $\frac{1}{3p}$ --- --- --- $3p$
 $1s^2 2s^2 2p^6 3s^2 3p^1$

What do all three of these already have in common?

They all have $1s^2 2s^2 2p^6$ in their config, all full up to $n=3$, all have e⁻ config of neon.

Is there another way you can write these configurations?

Yes, condensed e⁻ config

Na: [Ne] 3s¹

Mg: [Ne] 3s²

Al: [Ne] 3s² 3p¹

* Caution on exams: if asked for full e⁻ config, you must write out $1s^2 2s^2 2p^6 3s^1$ instead of [Ne] 3s¹ for something like Na. So read questions carefully!

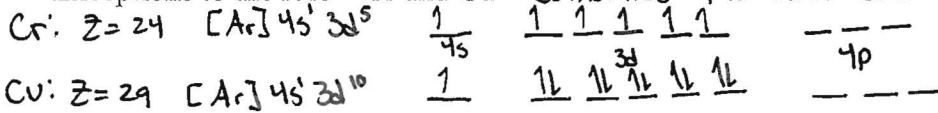
Period 4 and beyond

Which fills first, 3d or 4s? Why?

4s, it is lower energy than 3d due to penetration and shielding effects

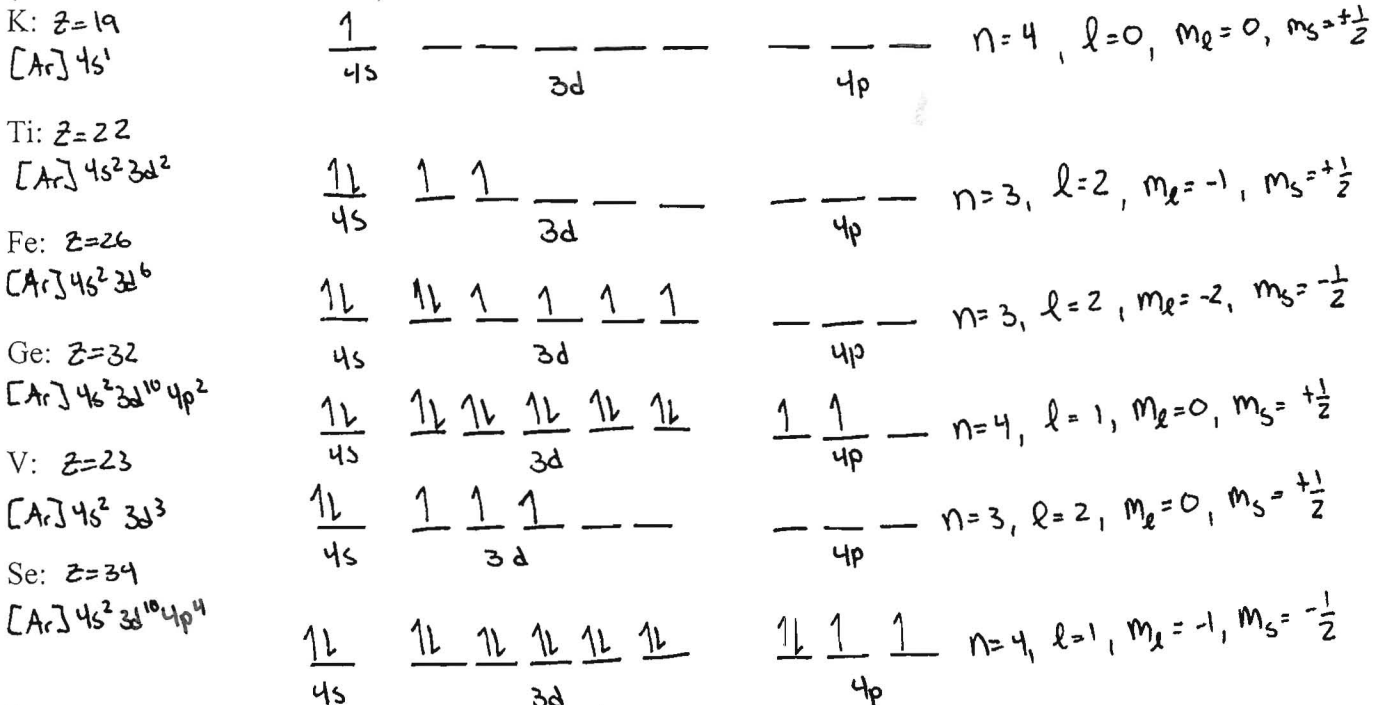
Exceptions to the rule -- Cr and Cu

Sometimes the most stable configuration is unexpected.



* MEMORIZE these

Write the condensed electron configuration, orbital diagram, and quantum numbers (based on last electron added) for



Periodic Trends

Size trends: Size increases to the left and down



Main group trends:

- as $n \uparrow$, probability that outer e^- will spend more time away from nucleus \uparrow , size \uparrow
- as $Z_{eff} \uparrow$, outer e^- are pulled closer to nucleus, size of atoms \downarrow

Transition element trends:

- After size shrinks for the 1st 2 or 3 elements, size then remains relatively constant due to shielding by d electrons which counteracts the usual increase in Z_{eff} .

Ionization energy trends:

- energy (kJ) required for the complete removal of 1 mole e^- from 1 mol gaseous atoms or ions
- \nearrow IE increases to the right & up (inversely proportional to size)

Electron affinity trends:

- Similar to IE, but w/ many exceptions

note that electronegativity also follows a similar trend. F is most electronegative atom. as IE \uparrow , it takes more energy to remove e^- . The more electronegative an atom is, the harder it is to remove e^- . (Though, F does not have the highest IE, so you can see how trends are a general thing)