We can actually calculate the % abundance of each isotope IF given info about each isotope

Bromine has two naturally occurring isotopes, ^{79}Br + ^{81}Br, which have atomic masses of 78.9183 amu and 80.9163 amu, respectively.

Determine the % abundance of each isotope.

This is a weighted distribution:

mass of 1st isotope

Average = %A (A) + % B (B) + ...

comes from periodic table

this is what we need to calculate

mass of 2nd isotope

$79.904 = x(78.9183) + y(80.9163)$

ALSO: the % Abund must add up to 100%

$x + y = 1$

$y = 1 - x$

$79.904 = 78.9183x + 80.9163(1-x)$

$79.904 = 78.9183x + 80.9163 - 80.9163x$

$-1.0123 = -1.998x$

$x = 0.5067$ so ^{79}Br has a 50.67% natural abundance

$y = 1 - 0.5067$

$y = 0.4933$

^{81}Br = 49.33% Abundant

other examples:

Copper has two naturally occurring isotopes: ^{63}Cu and ^{65}Cu

If a sample of Cu is 30.83 % ^{65}Cu (64.9278 amu), calculate the molar mass and abundance of ^{63}Cu.

$100\% = 30.83\% + x$

$x = 69.17\% = 0.6917$

$63.546 = 0.6917(M_{63}) + 0.3083(64.9278)$

from Periodic table

$63.546 = 0.6917 M_{63} + 20.02$

$43.53 = 0.6917 M_{63}$

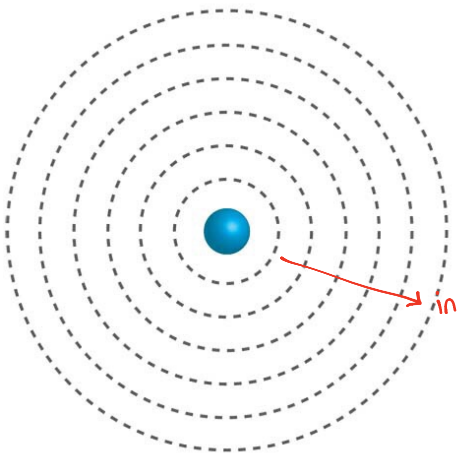
$M_{63} = 62.93$ amu

For atoms that have 1 e^-, it's very easy to predict how that e^- will interact with an atom. However, things get more complicated when more than 1 e^- is present.

First thing to consider ⟶ shells have e^- capacity

Inner shells have small volumes

↓ volume = ↓ e^- capacity

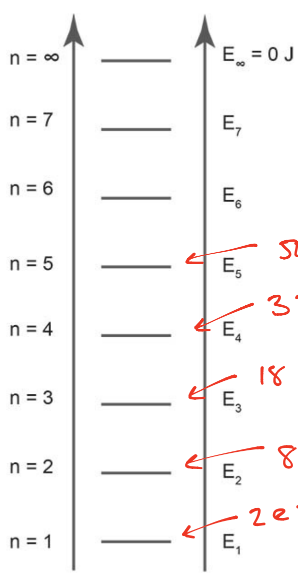


increasing e^- capacity

Shell capacity = $2n^2$

Shell	capacity
1	2
2	8
3	18
4	32
5	50

Now lets think about this in the context of our energy levels



50 e^-

32 e^-

18 e^-

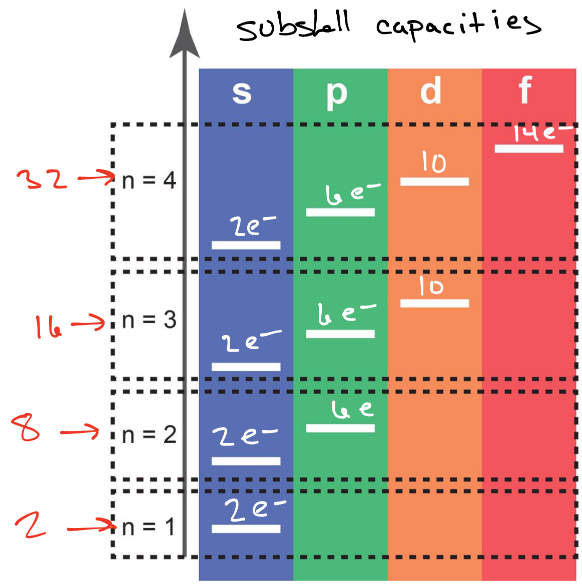
8 e^-

2 e^-

But this image is very misleading!

NOT all e^- in a shell have the same energy!

n	# of "subshells"	s	p	d	f
1	1	✓			
2	2	✓	✓		
3	3	✓	✓	✓	
4	4	✓	✓	✓	✓



Electrons are added to subshells according to energy → low E 1st

Electron configurations: $\underline{n}$ + subshell$^{\# \text{ of } e^-}$.

For example, Hydrogen contains 1 e⁻ is the 1st shell and s subshell

H: $1s^1$ ← # of e⁻

shell subshell

He: $1s^2$

Li: $1s^2 2s^1$

2 e⁻ in 1s subshell

1e⁻ in 2s

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

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

Ca: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Things get a little more complicated after Ca → The 3d subshell is lower energy than 4p!

As: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

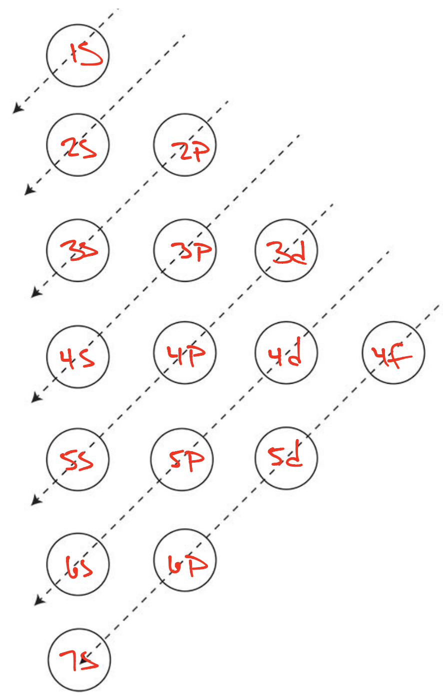
these are NOT valence electrons!

Use the periodic table as a guide.

s block = 2 columns
p block = 6 columns

d block = 10 columns
f block = 14 columns

or follow the diagonal guide



Noble gases all represent full shells (np^6). These are very stable e⁻ configs! as such, we use them to make writing e⁻ configs easier

Ar = $1s^2 2s^2 2p^6 3s^2 3p^6$

In an e⁻ config [Ar] =

So... Mn: [Ar] $4s^2$ $3d^5$

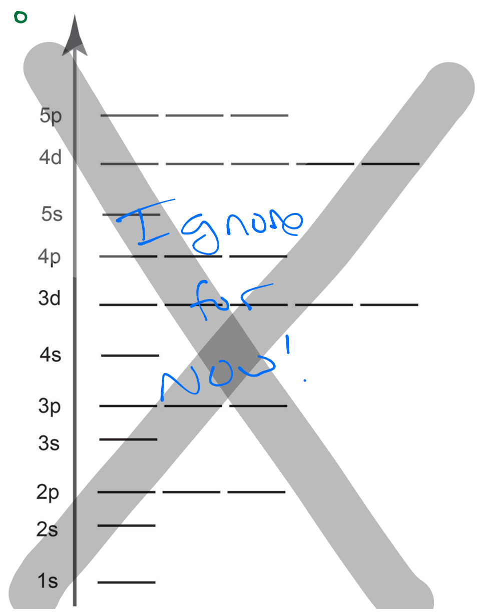
Sn: [Kr] $5s^2 4d^{10} 5p^2$

Pb: [Xe] $6s^2 4f^{14} 5d^{10} 6p^2$

How do e⁻ populate a subshell? Orbital

- these are regions of space, predicted by complex calculations, that it is most likely to find a e⁻.

- each orbital can hold a maximum o

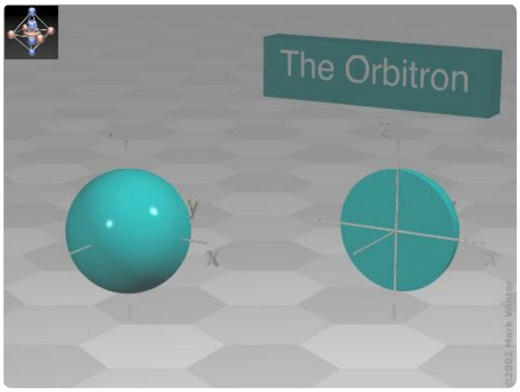


SO... 1 s orbital
3 p orbitals
5 d orbita
7 f orbitals

e⁻ populate orbitals with spin up (↑) or spin

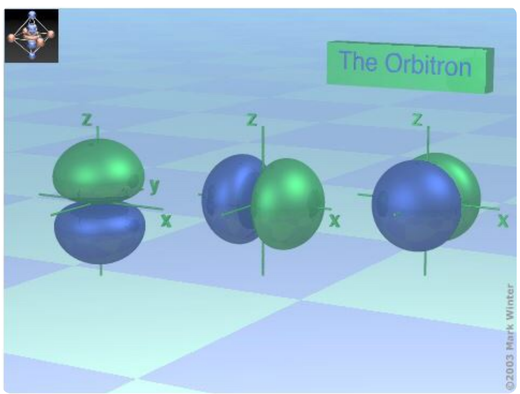
Hunds rule: atoms tend to be more stable IF e⁻ spi maximized. This has two practical applications:

1. orbitals in a subshell are filled 1 at a 1
   - spin pair (↑↓) happens only after all have an electron



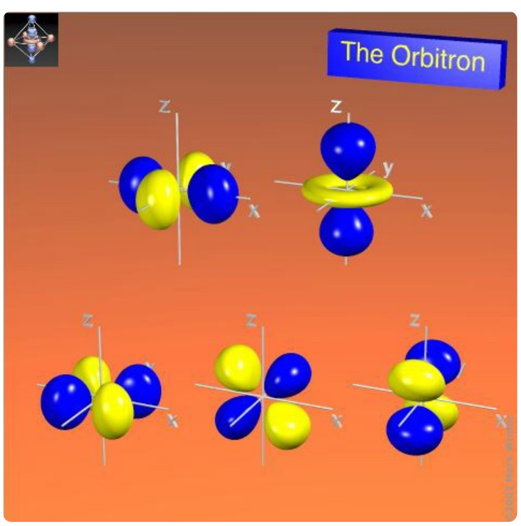
S - orbitals

- only 1 S orbital per shell
- smallest and lowest energy orbital in a shell
- spherical outer shape



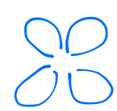
P - orbitals

- 3 different P-orbitals
- dumbell shaped
- do not exist in 1st shell



D- orbitals

- 5 different orbitals
- clover leaf shape (except one)



- do not exist in 1st or 2nd shell

Quantum Numbers

- this is a simple system that has been developed to allow a scientist to identify each e^- in an atom with a unique identifier

n = shell #

l = subshell $s = 0$ $p = 1$ $d = 2$ $f = 3$

m_l = orbital

- rule: $-l \rightarrow +l$ including $\emptyset$

s: $\emptyset$ — only 1 s orbital

p ($l = 1$): -1, 0, 1 — 3 p orbitals

d ($l = 2$): -2, -1, 0, 1, 2 — 5 d orbitals

f ($l = 3$): -3, -2, -1, 0, 1, 2, 3 — 5 f orbitals

m_s = electron spin $\rightarrow$ either $\underbrace{+1/2 \text{ or } -1/2}$ so each orbital can have 2 electrons

Example: What are the two possible sets of quantum #s for an e^- in a 4s orbital?

4s: $n = 4$, $l = 0 \longrightarrow m_l = 0$ (only option)

$m_s = +1/2$ or $-1/2$

answer like this

$[4, \emptyset, \emptyset, 1/2]$ or $[4, 0, 0, -1/2]$

2nd example: 3P electron

$n = 3$, $l = 1$; m_l: -1, 0, 1

m_s: $+1/2$, $-1/2$

$[3, 1, -1, 1/2]$ $[3, 1, -1, -1/2]$

$[3, 1, 0, 1/2]$ $[3, 1, 0, -1/2]$

$[3, 1, -1, 1/2]$ $[3, 1, -1, -1/2]$

6 possibilities

6 electrons in a p subshell