Chapter 8: Periodic properties of the elements

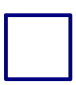
Electron configuration: shows which orbitals are occupied in an atom, and how many electrons they contain

Ground state: lowest energy, most stable state for an atom - has all electrons in the lowest energy possible orbitals

The energy of an electron in an H atom depends only on the principal quantum number, *n*, so in its ground state, the electron in an H atom occupies a ___ orbital.

Electron configuration:

Orbital diagram: figure that organizes electrons into their orbitals. Orbital = box, electron = arrow in box

H: 

Electron spin: direction of an electron's inherent angular momentum - this creates a magnetic field around the electron that either points **up** or **down**.

Spin quantum number, m_s, defines electron spin

- $m_s = +\frac{1}{2}$: up spin
- $m_s = -\frac{1}{2}$: down spin

Multi-electron atoms

The first 3 quantum numbers define an orbital

All 4 quantum numbers define **one electron** in an atom

- ***n*:** size of orbital
- ***ℓ*:** shape of orbital
- $\boldsymbol{m_\ell}$**:** orientation of orbital
- $\boldsymbol{m_s}$: spin of electron in that orbital

<u>Pauli exclusion principle</u>: no two electrons in an atom can have the same four quantum numbers

This means each orbital can hold no more than ___ electrons

He, 2 electrons, ground state electron configuration:

Orbital diagram: □

Be, 4 electrons, configuration:

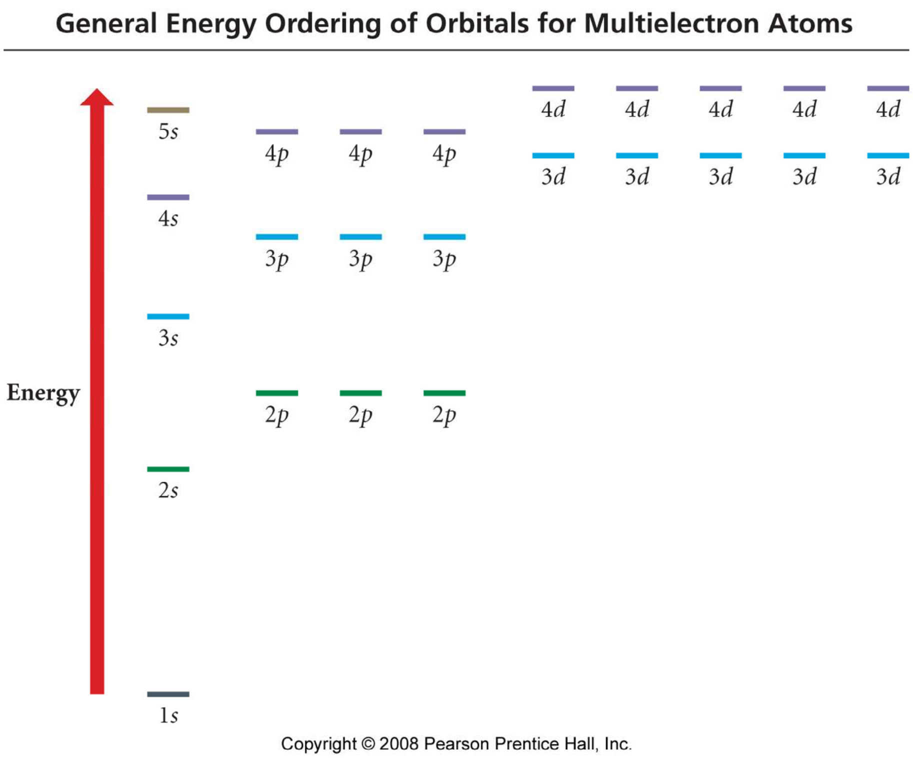
Orbital diagram:

Energy of electrons in multi-electron atoms

In hydrogen, *n* is the only quantum number necessary to calculate the energy of an orbital

In multi-electron atoms, both *n* and ***ℓ*** influence the energy of an orbital

Energies of ***ℓ***: (lowest) **s < p < d < f** (highest)



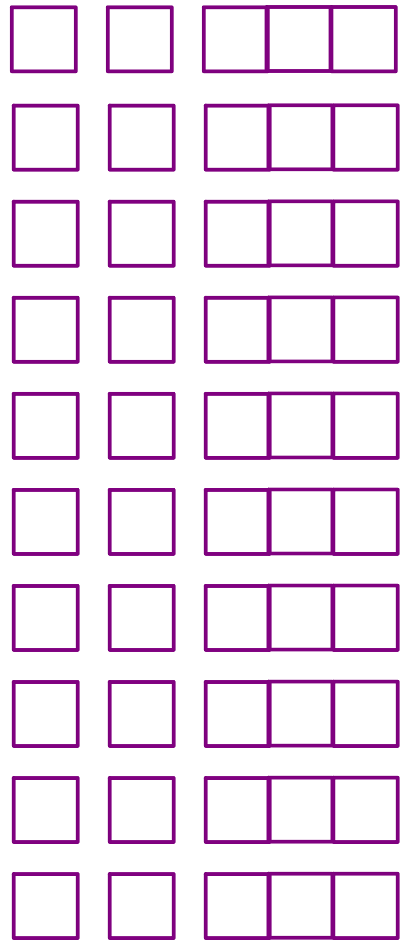
B (5 electrons) - configuration:

Orbital diagram:

Also, since energies get closer together as *n* increases, a **4s** orbital is lower in energy than a **3d** orbital

Electron configurations

H
He
Li
Be
B
C
N
O
F
Ne



<u>Hund's rule</u>: Electrons fill orbitals with equal energy **singly** first with parallel spins.

<u>Inner electrons</u>: A full noble gas electron configuration inside an atom - represented with noble gas symbol in brackets to make **<u>abbreviated electron configuration.</u>**

Periodic table and valence electrons

Outer Electron Configurations of Elements 1–18

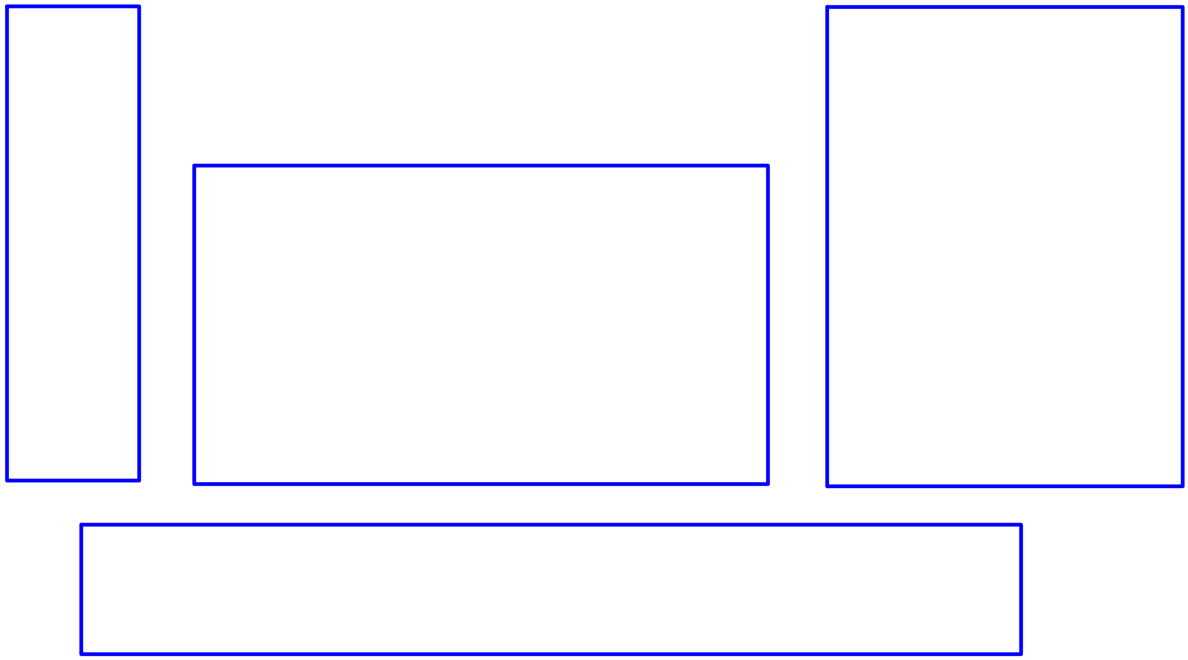
1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$

Valence electrons: number of electrons in outermost principal energy level (*n*) (plus outermost d electrons for transition elements)

Core electrons: inner electrons plus filled d or f sublevels

Periodic table and filling order

The ***reason*** the periodic table has its shape is because of the orbitals occupied in those elements.



Periodic Table of the Elements

Period	1 IA	2 IIA		3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 H 1.008																		2 He 4.003
2	3 Li 6.939	4 Be 9.012												5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31												13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08		21 Sc 44.96	22 Ti 47.90	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62		39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.96	43 *Tc* (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.4	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57-70 *	71 Lu 174.97	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89-102 **	103 *Lr* (257)	104 *Rf* (261)	105 *Db* (262)	106 *Sg* (271)	107 *Bh* (272)	108 *Hs* (270)	109 *Mt* (276)	110 *Ds* (281)	111 *Rg* (280)	112 *Uub* (285)	113 *Uut* (284)	114 *Uuq* (289)	115 *Uup* (288)	116 *Uuh* (292)	117	118 *Uuo* (294)

*	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 *Pm* (147)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04
**	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 *Np* (237)	94 *Pu* (244)	95 *Am* (243)	96 *Cm* (247)	97 *Bk* (247)	98 *Cf* (251)	99 *Es* (252)	100 *Fm* (257)	101 *Md* (258)	102 *No* (259)

Reference: http://www.webelements.com

Ca:

Sc:

Sometimes, electron configurations are rearranged into order of increasing *n* (to group valence electrons better)

Se:

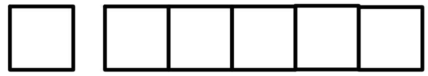
Abbreviated configurations:

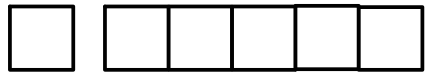
Po:

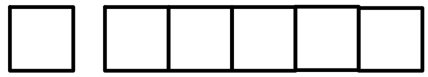
Bh:

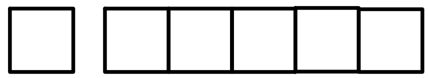
Electron configurations and magnetic properties

4s has lower energy than 3d, but they are still close.

Mn: 

Zn: 

Cr: 

Cu: 

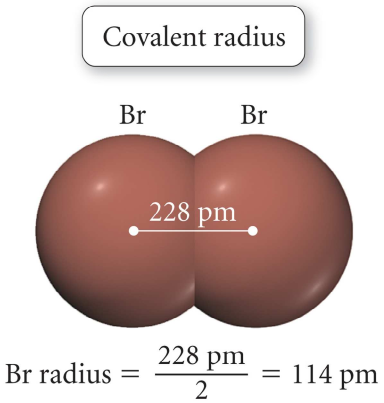
(You should be aware of the conditions behind these anomalies and be able to explain it if it occurs elsewhere, but do not memorize every exception on the periodic table!)

Magnetic properties

Unpaired electrons in the orbital diagram will make the element **paramagnetic** (weakly attracted to magnetic field)

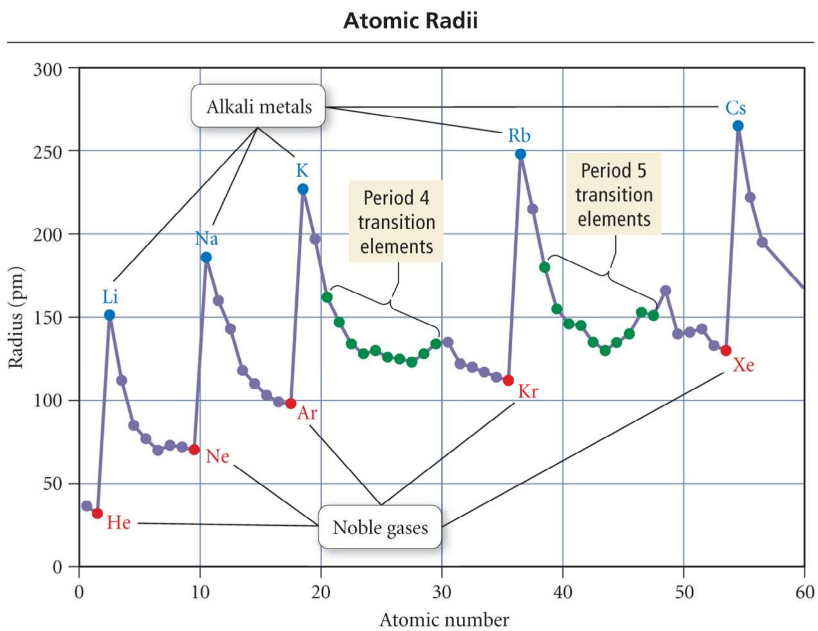
If all electrons are paired, the element is **diamagnetic** (not attracted by magnetic field)

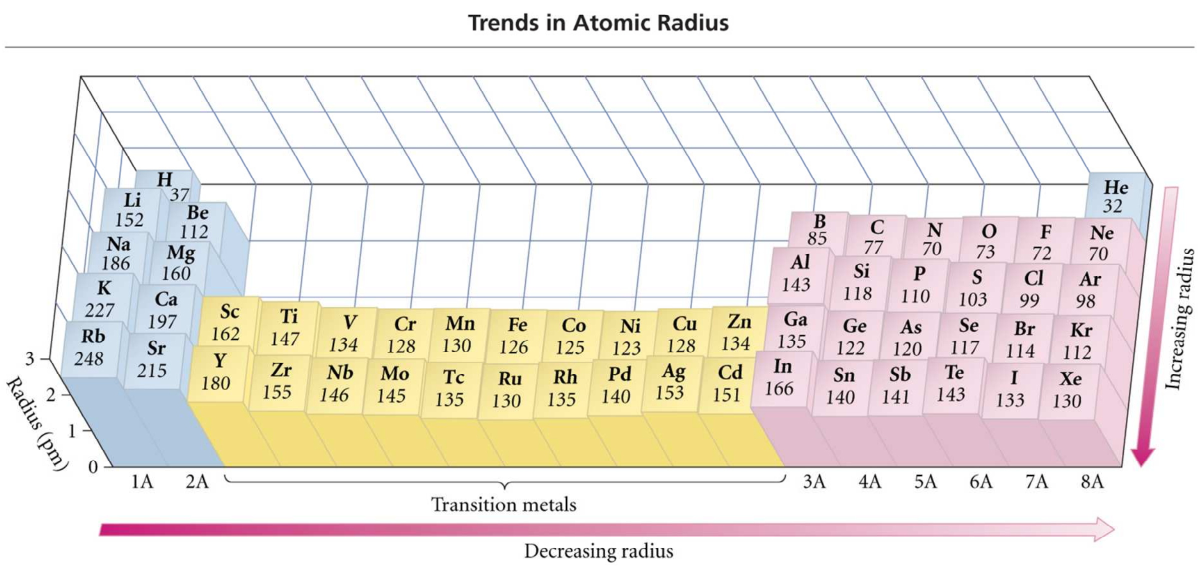
Periodic trends in atomic radius



Atomic radius **increases** going down a column:

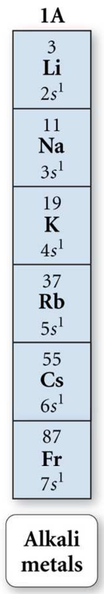
Atomic radius **decreases** going across a period (row):

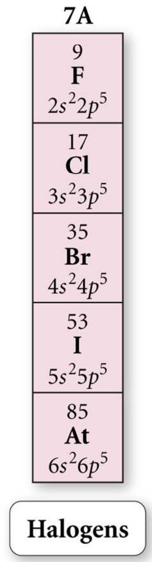




Ions

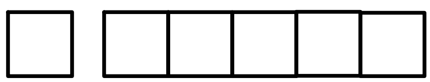
Main group ions:





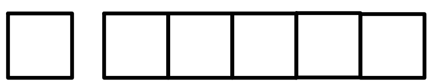
Transition metal ions:

Zn: [Ar] $3d^{10}$ $4s^2$



Experimentally, the Zn^{2+} ion is diamagnetic:

Zn^{2+}:



Transition metals tend to lose the ___ electrons **before** the ___ electrons!

(Writing the configuration in order of increasing *n* makes ion formation easier!)

Ag^+:

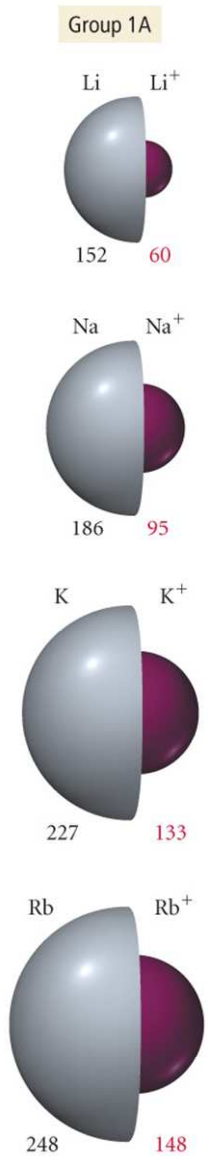
Ionic radius

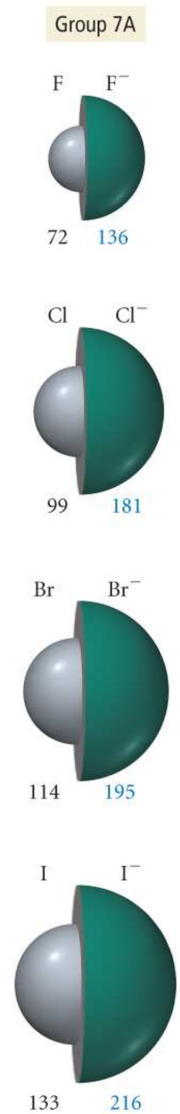
Na: [Ne] $3s^1$

Na^+: [Ne]

Cl: [Ne] $3s^2 3p^5$

Cl^-: [Ne] $3s^2 3p^6$





Ionization energy (IE): energy required to remove 1 electron from an atom or ion

Na $\rightarrow$ IE_1 = 496 kJ/mol

Na^+ $\rightarrow$ IE_2 = 4560 kJ/mol

Trends in first ionization energy (IE_1):

Mg $\rightarrow$ IE_1 = 738 kJ/mol

(additional protons across a period will increase attraction to the electrons)

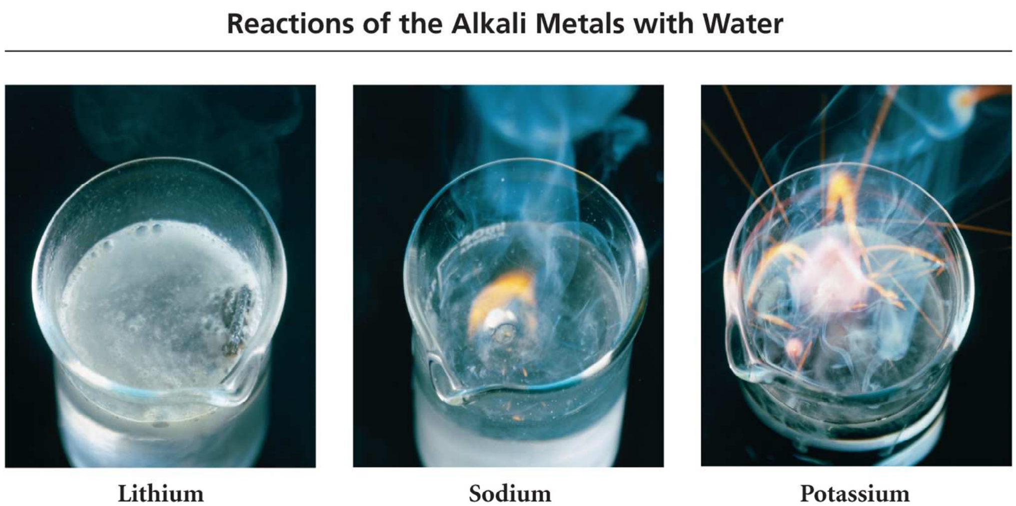
K $\rightarrow$ IE_2 = 419 kJ/mol

(an additional shell shields the valence electrons from the nuclear attraction)

Ionization energy and Electron affinity

$2\ Li(s) + 2\ H_2O(l) \rightarrow 2\ LiOH(aq) + H_2(g)$

$Li \rightarrow Li^+$
$Na \rightarrow Na^+$
$K \rightarrow K^+$



Electron affinity (EA): energy change associated with an electron being added to a neutral atom.

Li EA = -60 kJ/mol

O EA = -141 kJ/mol

F EA = -328 kJ/mol