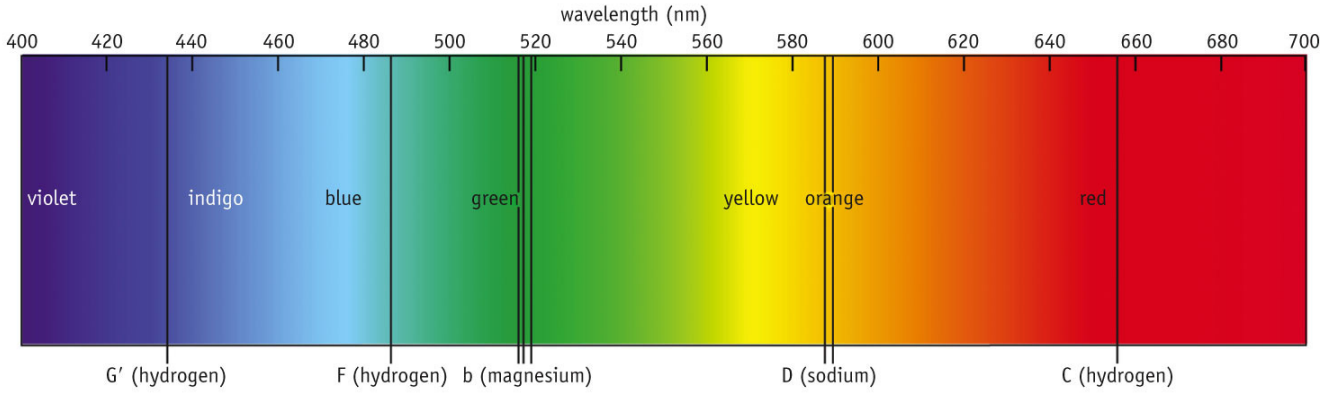
Chem 1045 Lecture Notes

Chemistry & Chemical Reactivity Kotz/Treichel/Townsend, 8th Ed

These Notes are to SUPPLEMENT the Text, They do NOT Replace reading the Text Material. Additional material that is in the Text will be on your tests! To get the most information, READ THE CHAPTER prior to the Lecture, bring in these lecture notes and make comments on these notes. These notes alone are NOT enough to pass any test! The author is not responsible for errors in these notes.

Chapter 7 The Structure of Atoms and Periodic Trends

Chemistry of the Sun **Fraunhofer Lines** are the dark lines present in the visible spectrum of the sun due to absorption of sunlight by elements in the outer layers of the sun.



Sapphires are aluminum oxide crystals with traces of iron and titanium

Rubies are aluminum oxide crystals with traces of Cr^{3+}

Pauli Exclusion Principal: no more than two electrons can occupy the same orbital, and if there are two electrons in the same orbital, they must have opposite spins.

Note: The quantum numbers n , l , m_l , m_s are further discussed and used below. Different electrons in an element must have different quantum numbers; else they are the same electron.

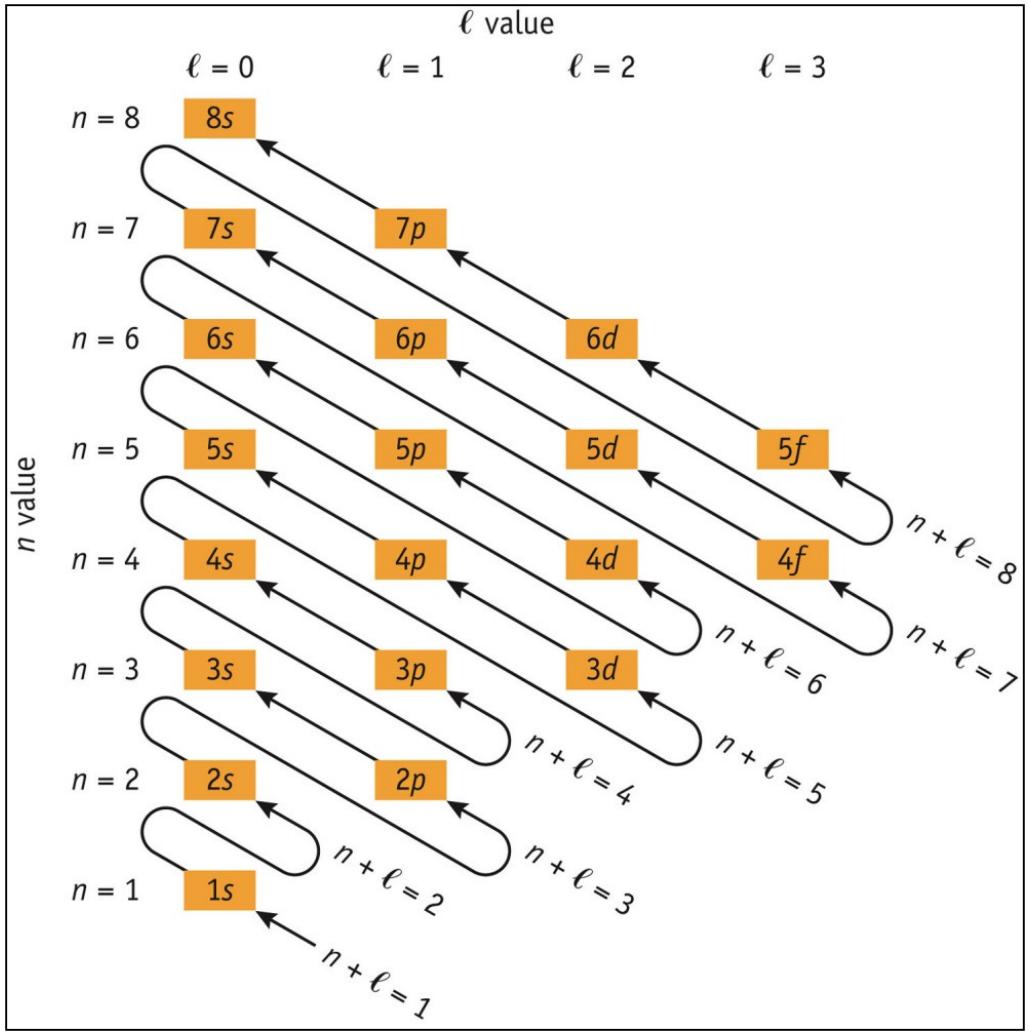
Electron Shell (n)	Subshells Available	Orbitals Available ($2\ell + 1$)	Number of Electrons Possible in Subshell $[2(2\ell + 1)]$	Maximum Electrons Possible for nth Shell ($2n^2$)
1	*s*	1	2	2
2	*s*	1	2	8
	p	3	6	
3	*s*	1	2	18
	p	3	6	
	d	5	10	
4	*s*	1	2	32
	p	3	6	
	d	5	10	
	f	7	14	
5	*s*	1	2	50
	p	3	6	
	d	5	10	
	f	7	14	
	*g**	9	18	
6	*s*	1	2	72
	p	3	6	
	d	5	10	
	*f**	7	14	
	*g**	9	18	
	*h**	11	22	

H 1 electron in the 1s orbital $n = 1, l = 0, m_l = 0, m_s = +1/2$ ↑ 1s

He 2 electrons in the 1s orbital $n = 1, l = 0, m_l = 0, m_s = +1/2$ ↑↓ 1s
$n = 1, l = 0, m_l = 0, m_s = -1/2$

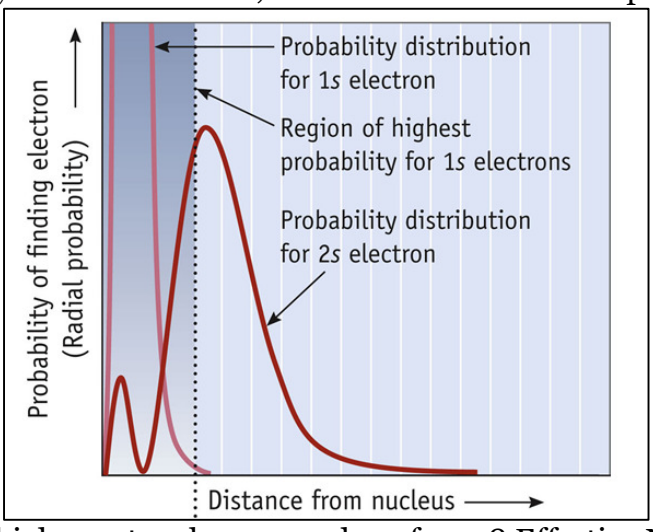
	l	
s subshell can hold 2 electrons	0	↑↓
p subshell can hold 6 electrons	-1 0 +1	↑↓ ↑↓ ↑↓
d subshell can hold 10 electrons	-2 -1 0 +1 +2	↑↓ ↑↓ ↑↓ ↑↓ ↑↓
f subshell can hold 14 electrons	-3 -2 -1 0 +1 +2	↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓

We fill the electrons in order of shells, s then p then d then f. But, there are some cases where an s shell will fill prior to a d or f due to energy content. The Aufbau principal is used to fill the electron shells as we build the periodic table. It is based on subshell energies.



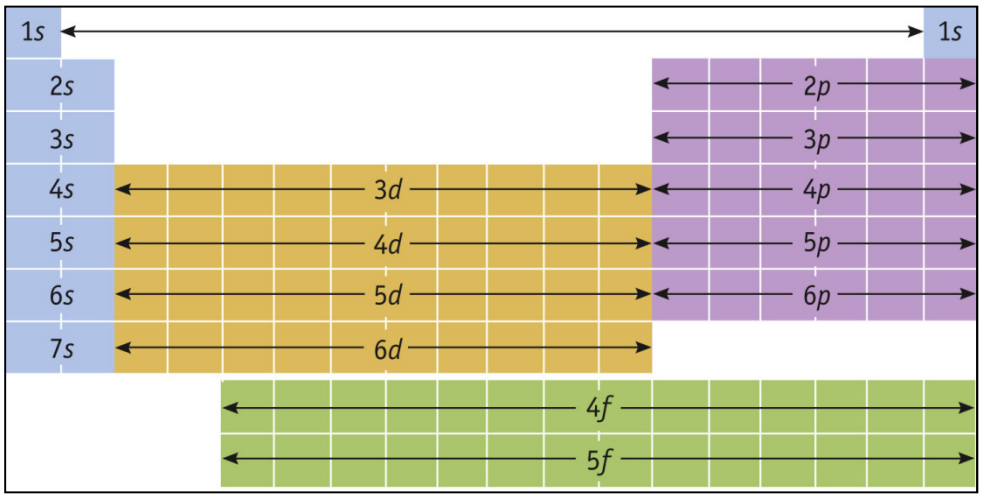
Effective Nuclear Charge Z* is the net charge experience by a particular electron in a multi-electron atom resulting from a balance of attractive force of the nucleus and the repulsive force of other electrons.

Lithium, atomic number 3 has $1s^2\ 2s^1$ electron configuration. The $2s^1$ at a distance from the nucleus will see a net +1 charge. The $1s^2$ will block two of the +1 nuclear charges. Since the $2s^1$ can penetrate some of the area occupied by the $1s^2$ electrons, the $2s^1$ will see more of a positive charge.



The Li $2s^1$ will experience a higher net + charge, a value of +1.28 Effective Nuclear Charge. The effective nuclear charge experienced by electrons in a multi-electron atom is **s > p > d > f.**

7.3 Electron Configuration of Atoms



Electrons are assigned to orbital's in order of increasing n.

Z	Element	Configuration	Z	Element	Configuration	Z	Element	Configuration
1	H	$1s^1$	37	Rb	$[Kr]5s^1$	74	W	$[Xe]4f^{14}5d^46s^2$
2	He	$1s^2$	38	Sr	$[Kr]5s^2$	75	Re	$[Xe]4f^{14}5d^56s^2$
3	Li	$[He]2s^1$	39	Y	$[Kr]4d^15s^2$	76	Os	$[Xe]4f^{14}5d^66s^2$
4	Be	$[He]2s^2$	40	Zr	$[Kr]4d^25s^2$	77	Ir	$[Xe]4f^{14}5d^76s^2$
5	B	$[He]2s^22p^1$	41	Nb	$[Kr]4d^45s^1$	78	Pt	$[Xe]4f^{14}5d^96s^1$
6	C	$[He]2s^22p^2$	42	Mo	$[Kr]4d^55s^1$	79	Au	$[Xe]4f^{14}5d^{10}6s^1$
7	N	$[He]2s^22p^3$	43	Tc	$[Kr]4d^55s^2$	80	Hg	$[Xe]4f^{14}5d^{10}6s^2$
8	O	$[He]2s^22p^4$	44	Ru	$[Kr]4d^75s^1$	81	Tl	$[Xe]4f^{14}5d^{10}6s^26p^1$
9	F	$[He]2s^22p^5$	45	Rh	$[Kr]4d^85s^1$	82	Pb	$[Xe]4f^{14}5d^{10}6s^26p^2$
10	Ne	$[He]2s^22p^6$	46	Pd	$[Kr]4d^{10}$	83	Bi	$[Xe]4f^{14}5d^{10}6s^26p^3$
11	Na	$[Ne]3s^1$	47	Ag	$[Kr]4d^{10}5s^1$	84	Po	$[Xe]4f^{14}5d^{10}6s^26p^4$
12	Mg	$[Ne]3s^2$	48	Cd	$[Kr]4d^{10}5s^2$	85	At	$[Xe]4f^{14}5d^{10}6s^26p^5$
13	Al	$[Ne]3s^23p^1$	49	In	$[Kr]4d^{10}5s^25p^1$	86	Rn	$[Xe]4f^{14}5d^{10}6s^26p^6$
14	Si	$[Ne]3s^23p^2$	50	Sn	$[Kr]4d^{10}5s^25p^2$	87	Fr	$[Rn]7s^1$
15	P	$[Ne]3s^23p^3$	51	Sb	$[Kr]4d^{10}5s^25p^3$	88	Ra	$[Rn]7s^2$
16	S	$[Ne]3s^23p^4$	52	Te	$[Kr]4d^{10}5s^25p^4$	89	Ac	$[Rn]6d^17s^2$
17	Cl	$[Ne]3s^23p^5$	53	I	$[Kr]4d^{10}5s^25p^5$	90	Th	$[Rn]6d^27s^2$
18	Ar	$[Ne]3s^23p^6$	54	Xe	$[Kr]4d^{10}5s^25p^6$	91	Pa	$[Rn]5f^26d^17s^2$
19	K	$[Ar]4s^1$	55	Cs	$[Xe]6s^1$	92	U	$[Rn]5f^36d^17s^2$
20	Ca	$[Ar]4s^2$	56	Ba	$[Xe]6s^2$	93	Np	$[Rn]5f^46d^17s^2$
21	Sc	$[Ar]3d^14s^2$	57	La	$[Xe]5d^16s^2$	94	Pu	$[Rn]5f^67s^2$
22	Ti	$[Ar]3d^24s^2$	58	Ce	$[Xe]4f^15d^16s^2$	95	Am	$[Rn]5f^77s^2$
23	V	$[Ar]3d^34s^2$	59	Pr	$[Xe]4f^36s^2$	96	Cm	$[Rn]5f^76d^17s^2$
24	Cr	$[Ar]3d^54s^1$	60	Nd	$[Xe]4f^46s^2$	97	Bk	$[Rn]5f^97s^2$
25	Mn	$[Ar]3d^54s^2$	61	Pm	$[Xe]4f^56s^2$	98	Ct	$[Rn]5f^{10}7s^2$
26	Fe	$[Ar]3d^64s^2$	62	Sm	$[Xe]4f^66s^2$	99	Es	$[Rn]5f^{11}7s^2$
27	Co	$[Ar]3d^74s^2$	63	Eu	$[Xe]4f^76s^2$	100	Fm	$[Rn]5f^{12}7s^2$
28	Ni	$[Ar]3d^84s^2$	64	Gd	$[Xe]4f^75d^16s^2$	101	Md	$[Rn]5f^{13}7s^2$
29	Cu	$[Ar]3d^{10}4s^1$	65	Tb	$[Xe]4f^96s^2$	102	No	$[Rn]5f^{14}7s^2$
30	Zn	$[Ar]3d^{10}4s^2$	66	Dy	$[Xe]4f^{10}6s^2$	103	Lr	$[Rn]5f^{14}6d^17s^2$
31	Ga	$[Ar]3d^{10}4s^24p^1$	67	Ho	$[Xe]4f^{11}6s^2$	104	Rf	$[Rn]5f^{14}6d^27s^2$
32	Ge	$[Ar]3d^{10}4s^24p^2$	68	Er	$[Xe]4f^{12}6s^2$	105	Db	$[Rn]5f^{14}6d^37s^2$
33	As	$[Ar]3d^{10}4s^24p^3$	69	Tm	$[Xe]4f^{13}6s^2$	106	Sg	$[Rn]5f^{14}6d^47s^2$
34	Se	$[Ar]3d^{10}4s^24p^4$	70	Yb	$[Xe]4f^{14}6s^2$	107	Bh	$[Rn]5f^{14}6d^57s^2$
35	Br	$[Ar]3d^{10}4s^24p^5$	71	Lu	$[Xe]4f^{14}5d^16s^2$	108	Hs	$[Rn]5f^{14}6d^67s^2$
36	Kr	$[Ar]3d^{10}4s^24p^6$	72	Hf	$[Xe]4f^{14}5d^26s^2$	109	Mt	$[Rn]5f^{14}6d^77s^2$
			73	Ta	$[Xe]4f^{14}5d^36s^2$			

Electron Configuration using Orbital and spdf notation:

H	1 electron in the 1s orbital	1 proton, 1 e⁻	↑	$1s^1$
He	2 electrons in the 1s orbital	2 proton, 2 e⁻	↑↓	$1s^2$
			1s	

$1s^1$ - The first 1 is the electron shell (n), s is the orbital type (l), and upper shift 1 or 2 is the number of electrons in that orbital.

Period 1

H	$1s^1$	↑		
He	$1s^2$	↑↓		

Period 2 is based on Helium [He]

Li	$1s^2$	$2s^1$		↑↓	↑				[He] $2s^1$
Be	$1s^2$	$2s^2$		↑↓	↑↓	_	_	_	[He] $2s^2$
B	$1s^2$	$2s^2$	$2p^1$	↑↓	↑↓	↑	_	_	[He] $2s^2\ 2p^1$
B	$1s^2$	$2s^2$	$2p^1$	↑↓	↑	↑	↑	_	[He] $2s^1\ 2p^2$
C*2	$1s^2$	$2s^2$	$2p^2$	↑↓	↑↓	↑	↑	_	[He] $2s^2\ 2p^2$
C*2	$1s^2$	$2s^2$	$2p^2$	↑↓	↑	↑	↑	↑	[He] $2s^1\ 2p^3$
N	$1s^2$	$2s^2$	$2p^3$	↑↓	↑↓	↑	↑	↑	[He] $2s^2\ 2p^3$
O	$1s^2$	$2s^2$	$2p^4$	↑↓	↑↓	↑↓	↑	↑	[He] $2s^2\ 2p^4$
F	$1s^2$	$2s^2$	$2p^5$	↑↓	↑↓	↑↓	↑↓	↑	[He] $2s^2\ 2p^5$
Ne*1	$1s^2$	$2s^2$	$2p^6$	↑↓	↑↓	↑↓	↑↓	↑↓	[He] $2s^2\ 2p^6$

Period 3 is based on Neon [Ne]

Na	$1s^2$	$2s^2$	$2p^6$	$3s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	[Ne] $3s^1$
					1s	**2s**	**$2p_x$**	**$2p_y$**	**$2p_z$**	**3s**	**NGC*1**

Neon has all of the electrons in the 2 orbital filled, it is called the **Noble Gas Configuration***1
When starting a new electron shell, we can start with the noble gas configuration of the previous.
These electrons, expressed as the noble gas configuration are the **core electrons**

***2 Hund's rule** states the most stable arrangements of electrons is that with the maximum number of unpaired electrons, all with the same spin.

Electrons beyond the core electrons are the **valence electrons**.
The **valence electrons** are the ones that determine the **chemical properties** of an element.

Ne	$1s^2$	$2s^2$	$2p^6$		↑↓	↑↓	↑↓	↑↓	↑↓		[He] $2s^2\ 2p^6$

Period 3

Na	$1s^2$	$2s^2$	$2p^6$	$3s^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	_	_	_	[Ne] $3s^1$
Mg	$1s^2$	$2s^2$	$2p^6$	$3s^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	_	_	_	[Ne] $3s^2$
Al	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	_	_	[Ne] $3s^2\ 3p^1$
Al	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	_	[Ne] $3s^1\ 3p^2$
Si*2	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	_	[Ne] $3s^2\ 3p^2$
Si*2	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	[Ne] $3s^1\ 3p^3$
P	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^3$	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	[Ne] $3s^2\ 3p^3$

S	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^4$	↑↓	↑↓	↑↓	↑↓	↑↓		↑↓	↑↓	↑	↑	[Ne] $3s^2$ $3p^4$
Cl	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$	↑↓	↑↓	↑↓	↑↓	↑↓		↑↓	↑↓	↑↓	↑	[Ne] $3s^2$ $3p^5$
Ar	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	↑↓	↑↓	↑↓	↑↓	↑↓		↑↓	↑↓	↑↓	↑↓	[Ne] $3s^2$ $3p^6$
						1s	**2s**	**$2p_x$**	**$2p_y$**	**$2p_z$**		**3s**	**$3p_x$**	**$3p_y$**	**$3p_z$**	**NGC*1**

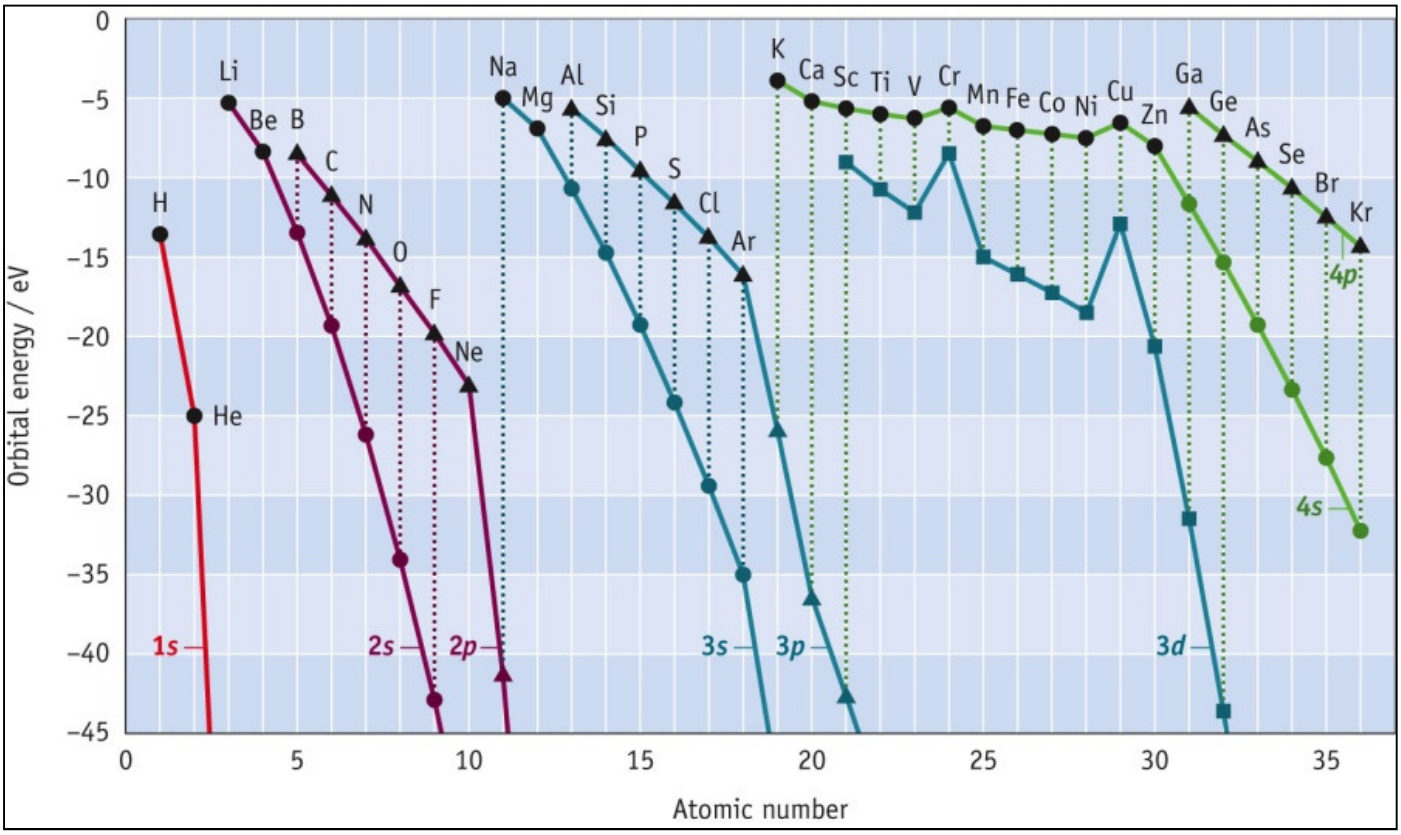
Period 4 will first fill the 4s shell for K and Ca. It will then skip the 4p subshell and start filling the 4d subshell (See Aufbau Principle) for the transition elements and then the 4p.

Filling the **4d subshell,** the elements are called the **Transition Elements**
Filling the **4f subshell**, the elements are called the **Lanthanides**
Filling the **5f subshell**, the elements are called the **Actinides**

Period 4

K	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^1$			↑										
Ca	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$			↑↓								[Ar] $4s^1$		
Sc	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^1$		↑↓	↑							[Ar] $4s^2$ $3d^1$		
Ti	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^2$		↑↓	↑	↑						[Ar] $4s^2$ $3d^2$		
V	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^3$		↑↓	↑	↑	↑					[Ar] $4s^2$ $3d^3$		
Cr	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^4$		↑↓	↑	↑	↑	↑				[Ar] $4s^2$ $3d^4$		
Cr	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^4$		↑	↑	↑	↑	↑	↑			[Ar] $4s^1$ $3d^5$		
Mn	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^5$		↑↓	↑	↑	↑	↑	↑			[Ar] $4s^2$ $3d^5$		
Fe	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^6$		↑↓	↑↓	↑	↑	↑	↑			[Ar] $4s^2$ $3d^6$		
Co	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^7$		↑↓	↑↓	↑↓	↑	↑	↑			[Ar] $4s^2$ $3d^7$		
Ni	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^8$		↑↓	↑↓	↑↓	↑↓	↑	↑			[Ar] $4s^2$ $3d^8$		
Cu	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^9$		↑↓	↑↓	↑↓	↑↓	↑↓	↑			[Ar] $4s^2$ $3d^9$		
Cu	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^9$		↑	↑↓	↑↓	↑↓	↑↓	↑↓			[Ar] $4s^1$ $3d^{10}$		
Zn	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$		↑↓	↑↓	↑↓	↑↓	↑↓	↑↓			[Ar] $4s^2$ $3d^{10}$		
Ga	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^1$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑			[] $4s^2$ $3d^{10}$ $4p^1$	
Ge	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^2$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑		[] $4s^2$ $3d^{10}$ $4p^1$	
As	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^3$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	[] $4s^2$ $3d^{10}$ $4p^1$	
Se	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^4$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	[] $4s^2$ $3d^{10}$ $4p^1$	
Br	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^5$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	[] $4s^2$ $3d^{10}$ $4p^1$	
Kr	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^6$	$4s^2$	$3d^{10}$	$4p^6$	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	[] $4s^2$ $3d^{10}$ $4p^1$	
									4s	$3d^1$	$3d^2$	$3d^3$	$3d^4$	$3d^5$	$4p^x$	$4p^y$	$4p^z$		

Orbital Energies, Z* and Electron Configuration



The above plot of atomic number vs orbital energies shows:

- Going left to right across a period, the atomic orbital decreases in size
- The Orbital Energy decreases

7.4 Electron Configuration of Ions

Chemistry deals with the chemistry of Anions and Cations. Electrons are always removed first from the electron shell of the highest n.

Na: [$1s^2\ 2s^2\ 2p^6\ 3s^1$] → Na^+ [$1s^2\ 2s^2\ 2p^6$] + 1 e-

Ge: [Ar] $3d^{10}\ 4s^2\ 4p^2$ → Ge^{2+} [Ar] $3d^{10}\ 4s^2$ + 2 e-

Note in these Transition elements, the ns electrons are lost first:

Ti: [Ar] $3d^{10}\ 4s^2$ → Ge^{2+} [Ar] $3d^{10}$ + 2 e-

Fe → Fe^{2+} Fe: [Ar] $3d^6\ 4s^2$ → Fe^{2+} [Ar] $3d^6$ + 2 e-

Fe^{2+} → Fe^{3+} Fe^{2+} [Ar] $3d^6$ → Fe^{3+} [Ar] $3d^5$ + 2 e-

Example 7.4 Give the electron configuration for Cu, Cu^+, Cu^{2+} and determine which is paramagnetic? (has unpaired electrons and are attracted to a magnet)

Cu $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑ [Ar] $4s^2\ 3d^9$

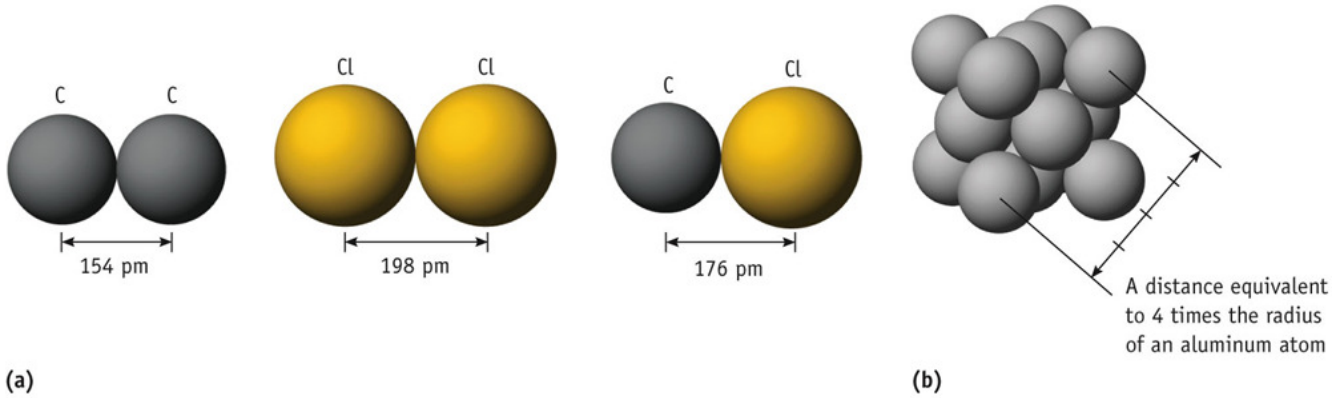
Cu^+ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ _ [Ar] $4s^2$

Cu^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^9$ ↑↓ ↑↓ ↑↓ ↑↓ ↑ _ [Ar] $4s^2\ 3d^9$

Cu^{2+} has one unpaired electron, so it is paramagnetic.

7.5 Atomic Properties and the Periodic Trend

Atomic Size: The atomic size of an element is usually determined by ½ the distance between two elements together (Cl_2, Crystal Structures, etc). They are not determined from distances of ionic compounds!

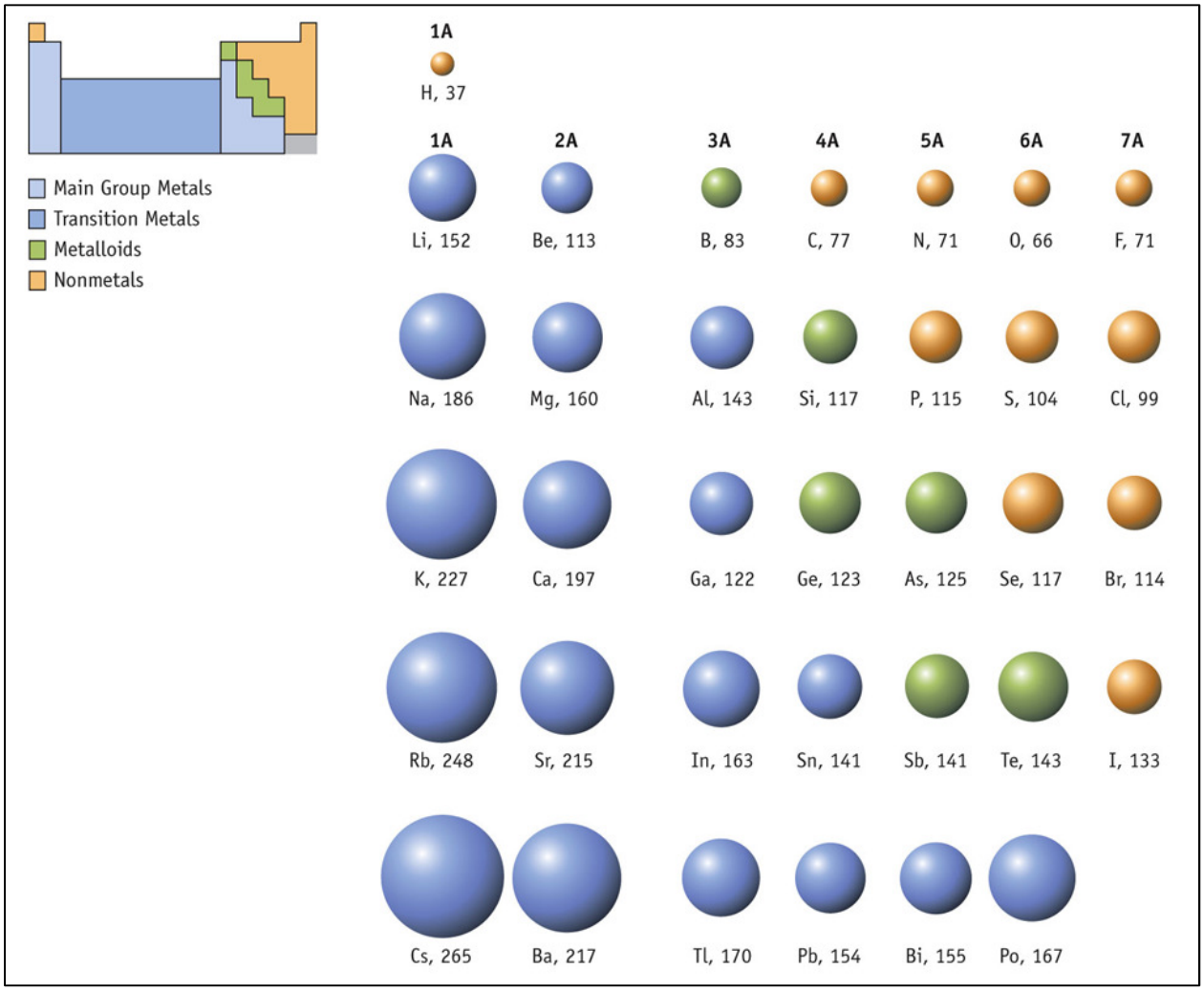


The size of an atom is determined by the outermost electrons
For the main group elements:

The atomic radius increases going down a group
The atomic radius decreases going across a period (L -> R).
The effective nuclear charge, Z*, increases going across a period.

For the Transition metal elements:

Going L -> R, the radii initially decrease, then change very little, then a small increase in size



Ionization Energy (IE) is the energy required to remove an electron from an atom in the gas phase

Atom $_{\text{ground state (g)}}$ → Atom $^{*}{}_{(g)}$ + e- ΔU = Ionization Energy, IE

First IE $Mg_{(g)}$ → $Mg^{+}{}_{(g)}$ + e- IE1 = 738 kJ/mole
$1s^2\ 2s^2\ 2p^6\ 3s^2$ $1s^2\ 2s^2\ 2p^6\ 3s^1$

Second IE $Mg^{+}{}_{(g)}$ → $Mg^{2+}{}_{(g)}$ + e- IE1 = 1451 kJ/mole
$1s^2\ 2s^2\ 2p^6\ 3s^2$ $1s^2\ 2s^2\ 2p^6$

Third IE $Mg^{2+}{}_{(g)}$ → $Mg^{3+}{}_{(g)}$ + e- IE1 = 7732 kJ/mole
$1s^2\ 2s^2\ 2p^6$ $1s^2\ 2s^2\ 2p^5$

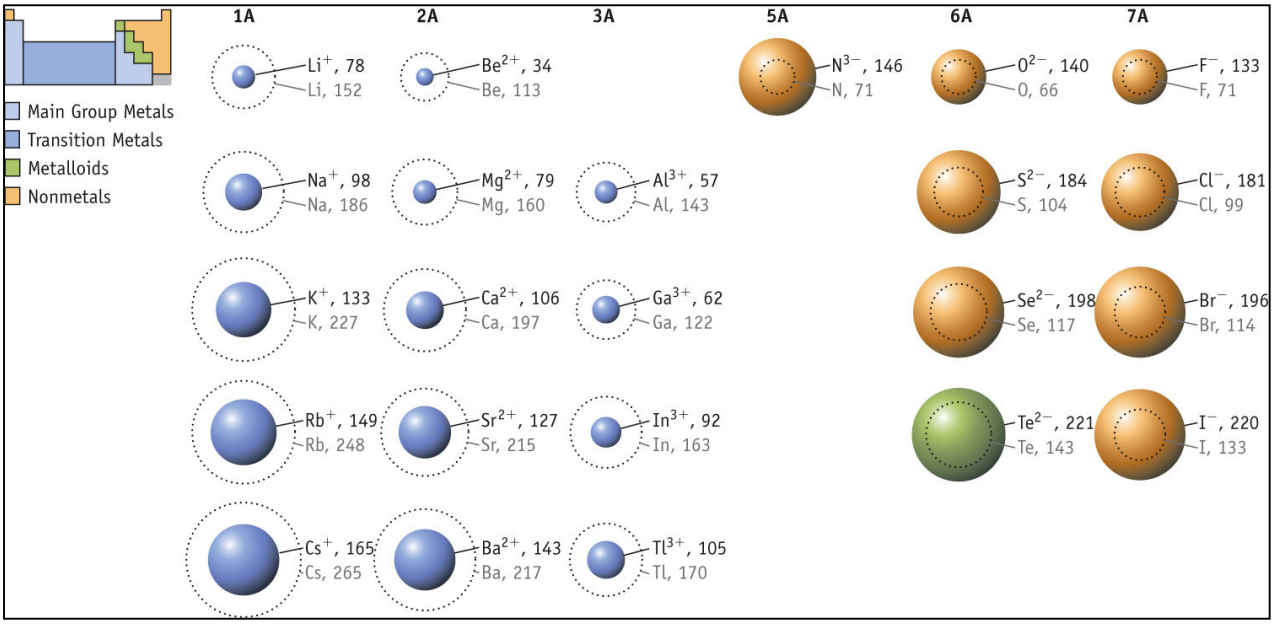
Each subsequent e- removed after the first requires more energy because an e- is being removed from a more positive ion.

Table 7.5 First, Second, and Third Ionization Energies for the Main Group Elements in Periods 2–4 (kJ/mol)

2nd Period	Li	Be	B	C	N	O	F	Ne
1st	513	899	801	1086	1402	1314	1681	2080
2nd	7298	1757	2427	2352	2856	3388	3374	3952
3rd	11815	14848	3660	4620	4578	5300	6050	6122
3rd Period	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
1st	496	738	577	787	1012	1000	1251	1520
2nd	4562	1451	1817	1577	1903	2251	2297	2665
3rd	6912	7732	2745	3231	2912	3361	3826	3928
4th Period	**K**	**Ca**	**Ga**	**Ge**	**As**	**Se**	**Br**	**Kr**
1st	419	590	579	762	947	941	1140	1351
2nd	3051	1145	1979	1537	1798	2044	2104	2350
3rd	4411	4910	2963	3302	2735	2974	3500	3565

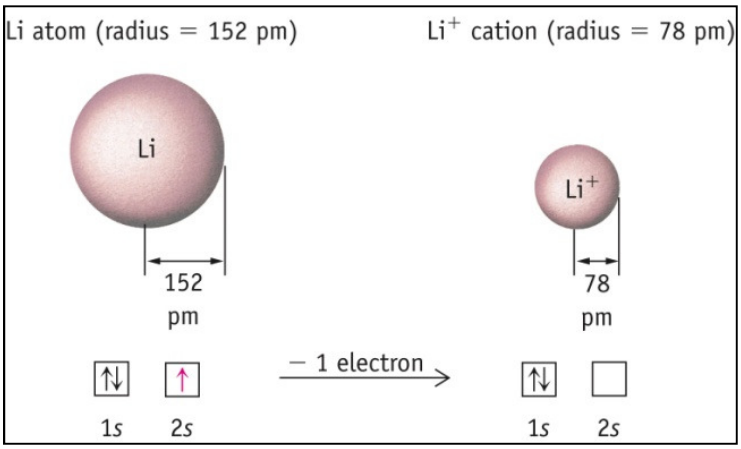
The First IE increases going across a period and down a group
The atomic radius decreases and the IE increases due to the increase in effective nuclear charge

Trends in Ion Sizes

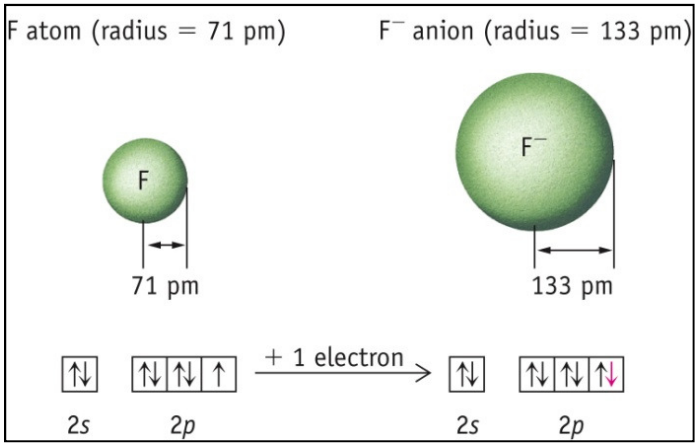


Positive and negative ions increase in size going down a group

When an e- is removed from an atom to form a cation, the size shrinks considerably.

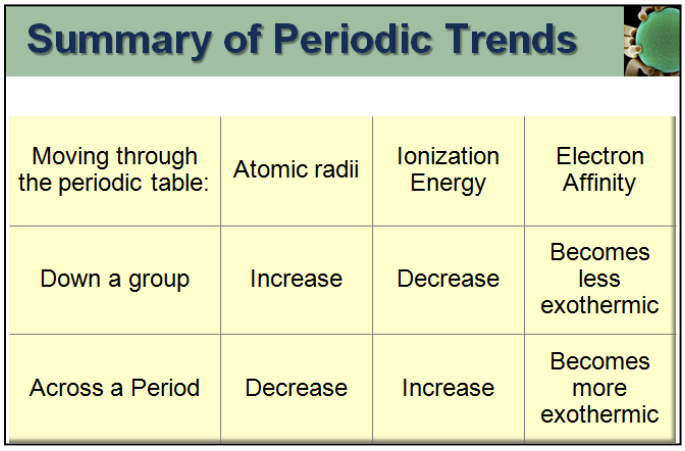


When an e- is added to an atom to form an anion, the size is always larger



Isoelectronic ions have the same number of electrons, but a different number of protons:

Ion	**N^{3-}**	**O^{2-}**	**F^{-}**	**Na^{+}**	**Mg^{2+}**
# of Electrons	10	10	10	10	10
# of Protons	7	8	9	11	12
Ionic Radius (pm)	146	140	133	98	79



The main group elements form Cations with an electron configuration equivalent to that of the **preceding noble gas**

Nonmetals acquire enough electrons to form an anion with the electron configuration of the **next noble gas.**

Students should read: Milestones in the Development of Chemistry…

Radioactivity: Alpha, Beta, Gamma are separated by passing radiation between electrically charged plages.

Thompsons experiment to measure the electron charge-to-mass ratio

Goldstein's Canal Rays

Rutherfords experiment to determine the structure of an aom – hit a gold foil with + particals, detect with a ZnS screen.

Milikan's Oil Drop experiment to determine the charge on an electron