

Electron Configurations continued:

Electrons in the outermost shell are called *valence electrons*.

- It is the *valence electrons* determine an atom's chemical properties.
- Electrons in the inner shells are inner electrons or *core electrons*.
- Regions in periodic table are designated as the *s-block*, *p-block*, *d-block* and the *f-block*.

Main-group elements

s block

p block

Transition elements

d block

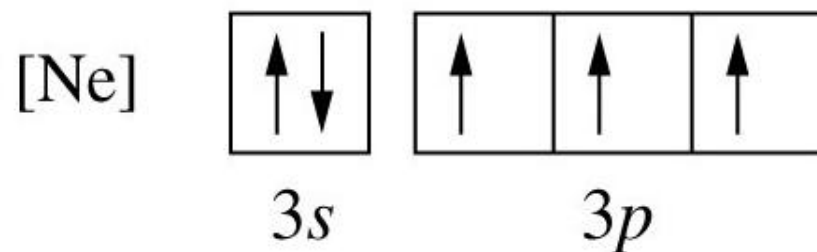
Inner-transition elements

f block

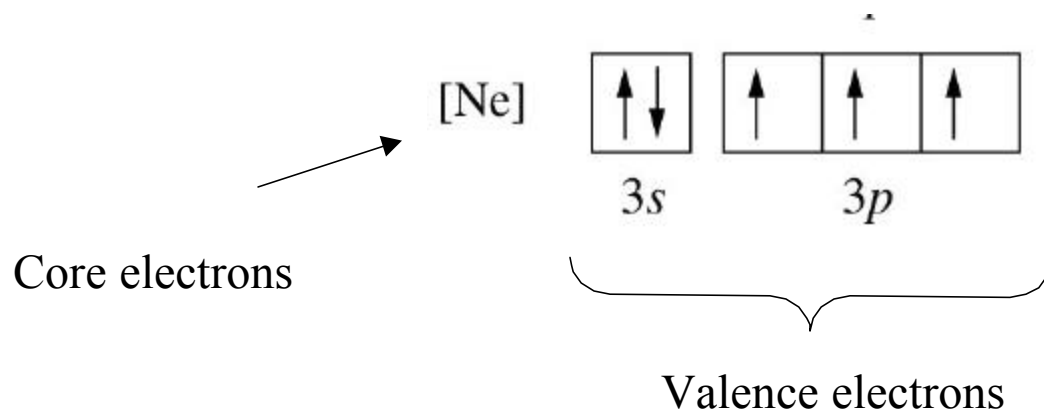
1	Main-group elements																18	
1s																	2s	
H																	He	
3	4	Transition elements										13	14	15	16	17	18	
2s	2p											3s	3p	3d	3p	3p	3p	3p
Li	Be											B	C	N	O	F	Ne	
11	12	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	
3s	3d	4s	4p	4d	4d	4d	4d	4d	4d	4d	4d	4p	4p	4p	4p	4p	4p	
Na	Mg	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Al	Si	P	S	Cl	Ar	
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	
4s	4d	4d	4d	4d	4d	4d	4d	4d	4d	4d	4d	4p	4p	4p	4p	4p	4p	
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54	
5s	5d	5d	5d	5d	5d	5d	5d	5d	5d	5d	5d	5p	5p	5p	5p	5p	5p	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86	
6s	6d	6d	6d	6d	6d	6d	6d	6d	6d	6d	6d	6p	6p	6p	6p	6p	6p	
Cs	Ba	La*	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
87	88	89	104	105	106	107	108	109	110	111	112	81	82	83	84	85	86	
7s	7d	7d	7d	7d	7d	7d	7d	7d	7d	7d	7d	6p	6p	6p	6p	6p	6p	
Fr	Ra	Ac†	Rf	Db	Sg	Bh	Hs	Mt				Tl	Pb	Bi	Po	At	Rn	
Inner-transition elements																		
<i>f</i> block																		
58	59	60	61	62	63	64	65	66	67	68	69	70	71					
4f	4f	4f	4f	4f	4f	4f	4f	4f	4f	4f	4f	4f	4f	4f				
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu					
90	91	92	93	94	95	96	97	98	99	100	101	102	103					
5f	5f	5f	5f	5f	5f	5f	5f	5f	5f	5f	5f	5f	5f	5f				
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr					

What is the *electron configuration* of the element as shown on the right?

Which electrons are the *valence electrons* and which are the *core electrons*?



Answers:



Electron Configurations of ions:

When an atom gains or loses electrons it becomes an ion.

Loss of electrons: **Cation**

Gain of electrons: **Anion**

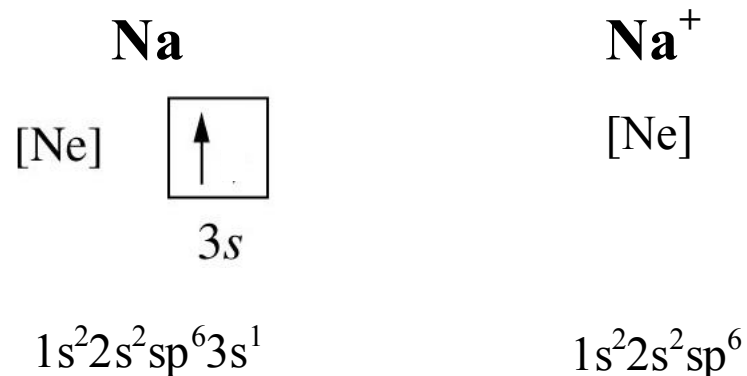
(+)

(-)

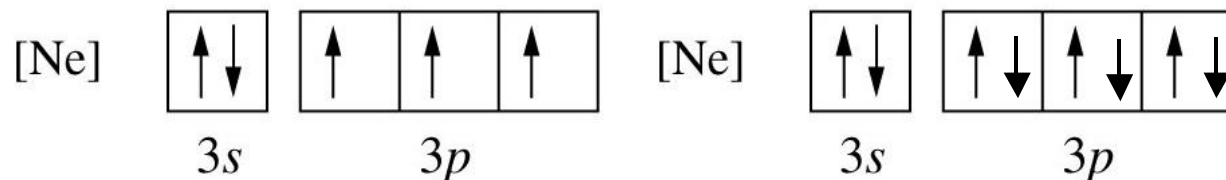
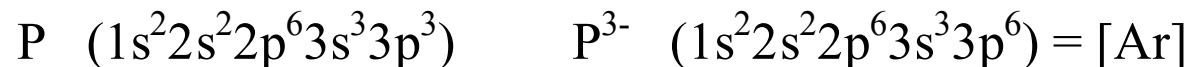
Atoms of group 1, 2 and the non-metals will lose or gain electrons to attain the electron configuration of a *Nobel gas* (ns^2np^6).

Example:

Sodium forming sodium ion:



All group 1 cations (+1) will have the electron configuration of the previous Noble gas.

The formation of the phosphide ion:

All group 5 (15) anions (-3) will have the electron configuration of the next Nobles gas.

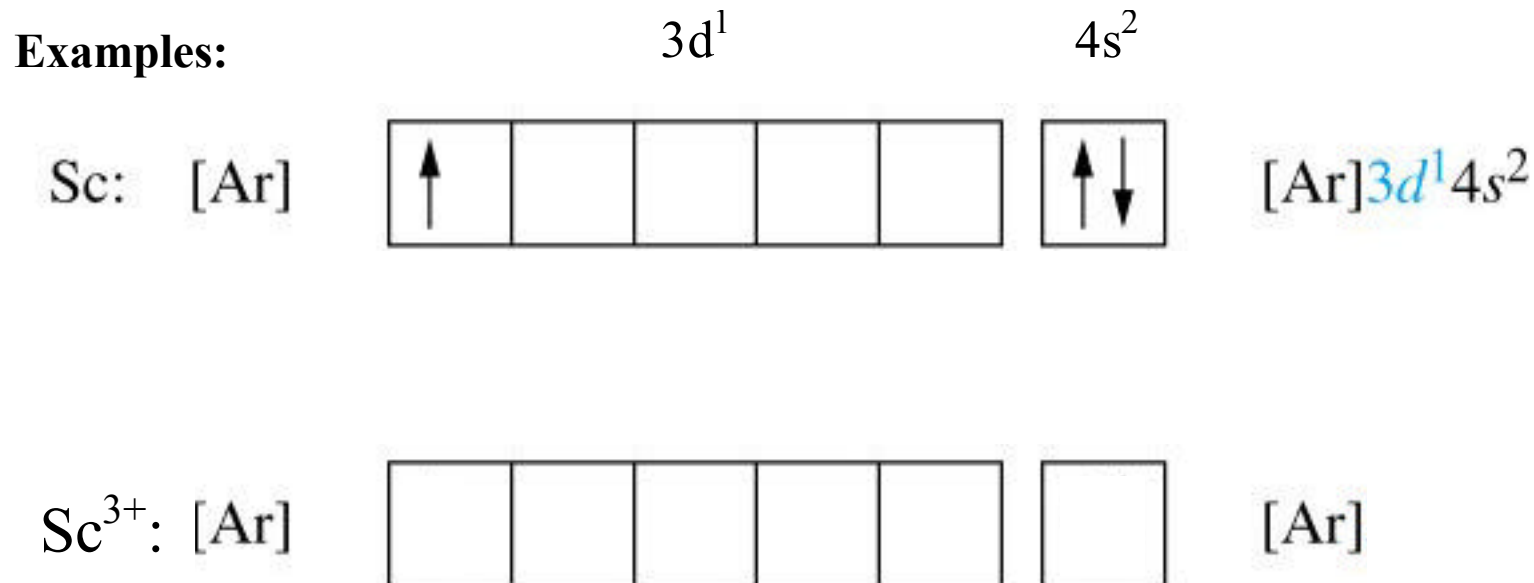
Electron configurations of Transition Metal Ions:

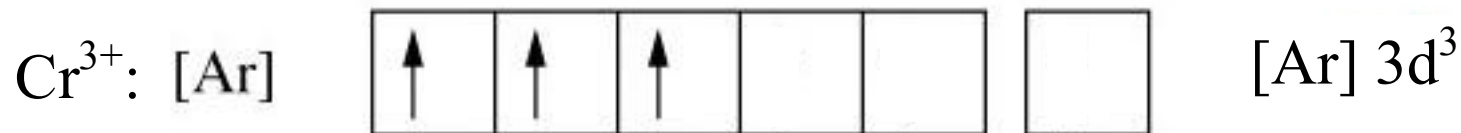
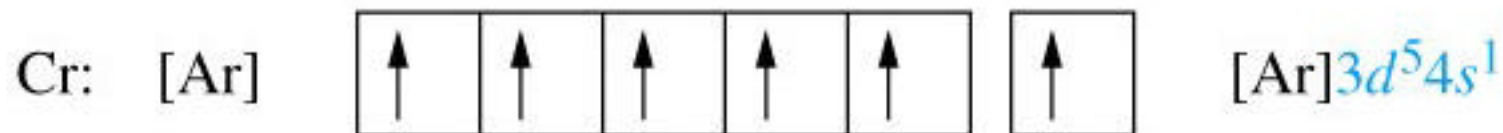
The charge states on most transition metals cannot be equated to the electron configurations of the Noble gases. You will see why in chem. 2C (ch. 24 and 25)

The electron configurations of transitions metal cations must be determined from the charge state.



- When *transition metals* ionize, they lose the s-electrons first.
- This may seem odd as it is opposite to the filling order.
- From a stability argument, the removal of the s-electrons first results in lower energy configuration when we take into account ionization energies of s vs. p and d electrons.



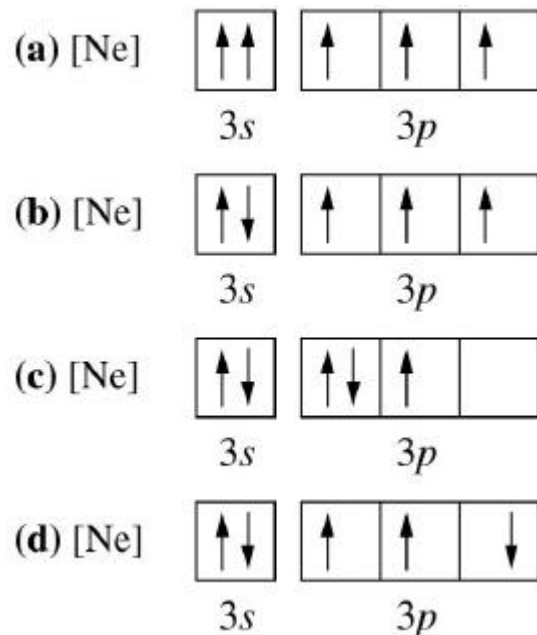


Electron Configurations: *Excited States*

Which of the electron configurations to the right corresponds to phosphorous?

Answer: c

The rest are excited states of phosphorous.

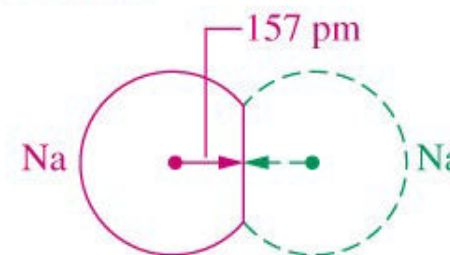


The Periodicity of Atomic Properties: Trends across the periodic table (Ch.10)

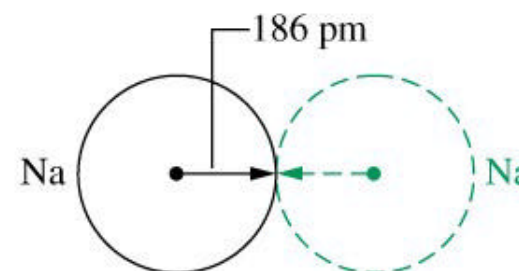
Half the distance between neighboring atoms defines an *atomic radius*.

- In covalent compounds (molecules) the *covalent radius* is half the distance between the nuclei of two identical atoms held together by a single covalent bond.
- The *metallic radius* in a metal is half the distance between nuclei of two atoms in contact in a crystalline solid metal.
- An *ionic radius* is derived from the distance between nuclei of a cation and anion pair in an ionic bond.

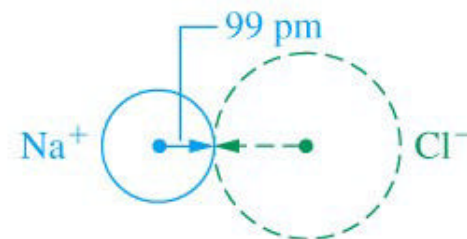
Covalent radius:



Metallic radius:



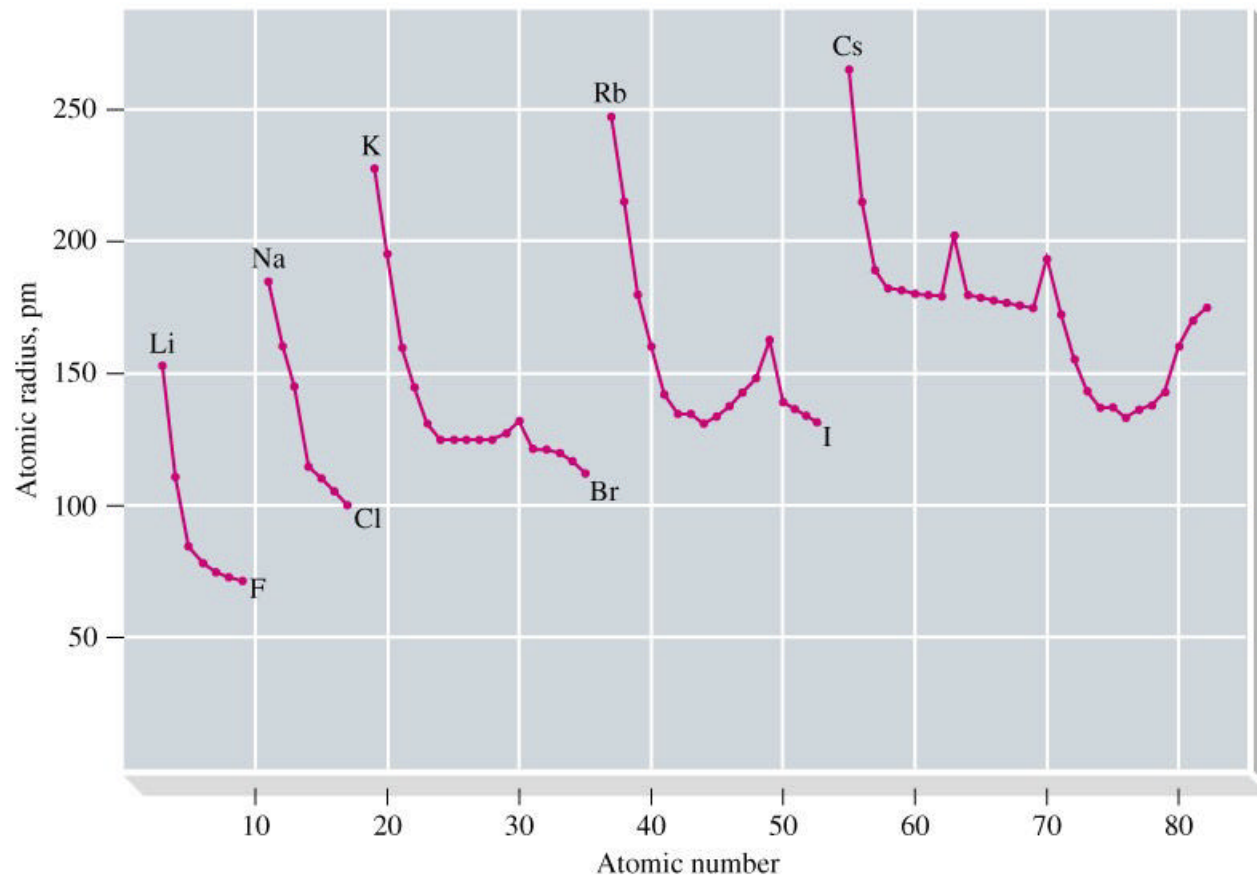
Ionic radius:



The Periodicity of Atomic Radii

The *atomic radius* of atoms decreases going from left-to-right across a period.

- Across a period, the effect of shielding decreases which allows the valence electron to be held tighter.
- Atomic radius increases down a group. Each step down involves adding a new shell, which increases size.



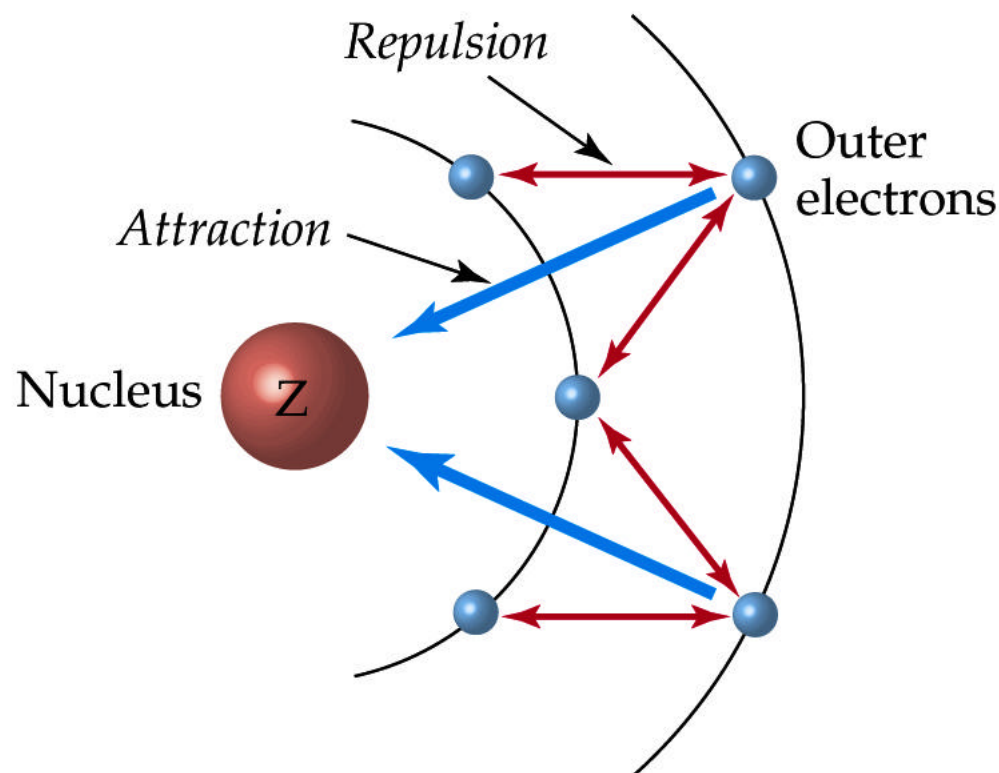
Effective Nuclear Charge:

The periodic trend of atomic radius is an example of the effect of shielding.

Shielding reduces the charge Z that the valence electrons feel to Z_{eff} , which is called the ***effective nuclear charge***.

The innermost electrons (***core***) screen or “***shield***” the nucleus from the view of the outer (***valence***) electrons.

The valence electrons in effect see a charge (Z_{eff}) that is less than the actual charge on the nucleus, Z .



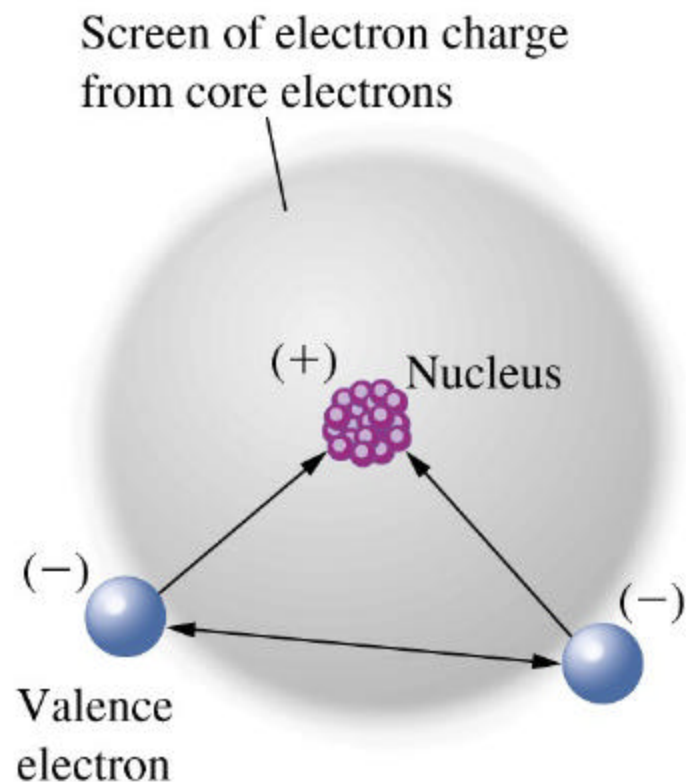
The interaction of the electrons reduces the attractive forces between the nucleus and the electrons.

The extent of this effect is a function of the screening ability of the core electrons:

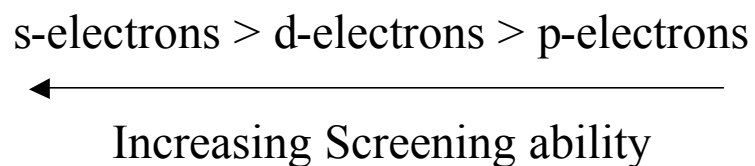
$$Z_{\text{eff}} = Z - S$$

Z = atomic number

S = screening constant



The trend is:



This indicates that Z_{eff} increases from left-to-right across a period.

As one moves down a group, Z_{eff} decreases as screening increases.

The over all trend is:

Z_{eff} Increases →

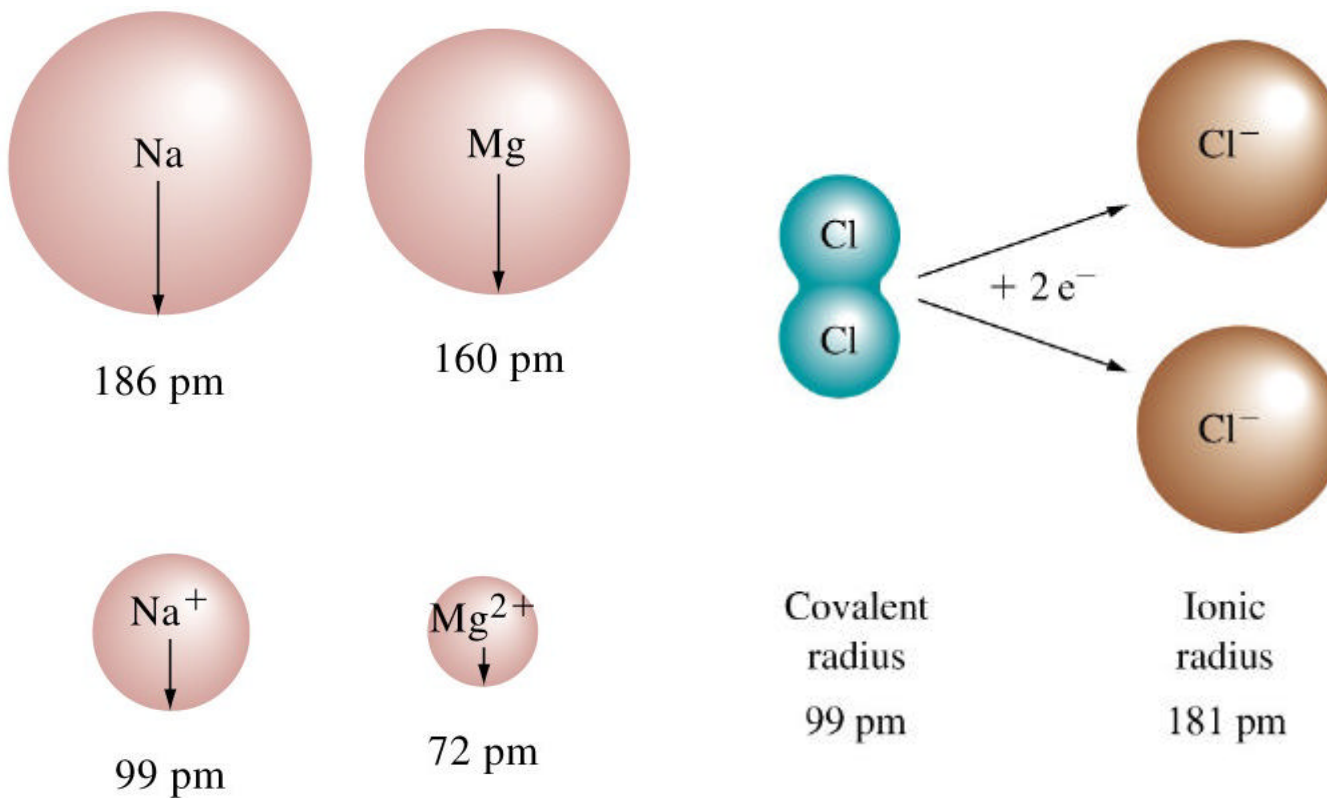
Z_{eff} Decreases ↓

1 1A												18 8A					
1 H 1.00794	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	18 2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.98976	12 Mg 24.3050	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B			11 1B	12 2B	13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc 98	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	*La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.385	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	*Ac 227.028	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (264)	108 Hs (265)	109 Mt (266)	110 (267)	111 (268)	112 (269)	114 (285)	116 (289)	118 (289)	119 (290)	120 (291)	118 (293)
*Lanthanide series			58 Ce 140.115	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967	
†Actinide series			90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	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)	103 Lr (260)	

This explains the trend in atomic radius: Across, as Z_{eff} gets larger, the valence electrons are held more tightly. Down a group, the reverse is seen.

Trends in Ionic Radii

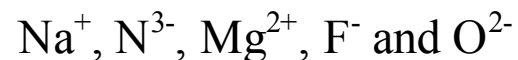
- The radii of cations are smaller than radii of the corresponding neutral atoms. (same Z , less electrons)
- The radii of anions are larger than radii of the corresponding neutral atoms. (same Z , more electrons)



The trends of ionic radii are a function of the charge state of the ion.

Li 152 Li⁺ 59	Be 111 Be²⁺ 27																		B 88	C 77	N 75 N³⁻ 171	O 73 O²⁻ 140	F 71 F⁻ 133			
Na 186 Na⁺ 99	Mg 160 Mg²⁺ 72													Al 143 Al³⁺ 53	Si 117	P 110 P³⁻ 212	S 104 S²⁻ 184	Cl 99 Cl⁻ 181								
K 227 K⁺ 138	Ca 197 Ca²⁺ 100	Sc 161 Sc³⁺ 75	Ti 145 Ti²⁺ 86	V 132 V²⁺ 79 V³⁺ 64	Cr 125 Cr²⁺ 82 Cr³⁺ 62	Mn 124 Mn²⁺ 83	Fe 124 Fe²⁺ 77 Fe³⁺ 65	Co 125 Co²⁺ 75 Co³⁺ 61	Ni 125 Ni²⁺ 70	Cu 128 Cu⁺ 96 Cu²⁺ 73	Zn 133 Zn²⁺ 75	Ga 122 Ga³⁺ 62	Ge 122		As 121		Se 117 Se²⁻ 198	Br 114 Br⁻ 196								
Rb 248 Rb⁺ 149	Sr 215 Sr²⁺ 113									Ag 144 Ag⁺ 115	Cd 149 Cd²⁺ 95	In 163 In³⁺ 79	Sn 141 Sn²⁺ 93	Sb 140 Sb³⁺ 76		Te 137 Te²⁻ 221	I 133 I⁻ 220									

Rank the following ions in order of decreasing size?



ion	# protons	# of electrons	ne/np
Na^+	11	10	0.909
N^{3-}	7	10	1.43
Mg^{2+}	12	10	0.833
F^-	9	10	1.11
O^{2-}	8	10	1.25

- Notice that they all have 10 electrons: They are *isoelectronic* (same electron configuration) as Ne.
- Since N^{3-} has the highest ratio of electrons to protons, it must be the largest.
- Mg^{2+} must then be the smallest.

