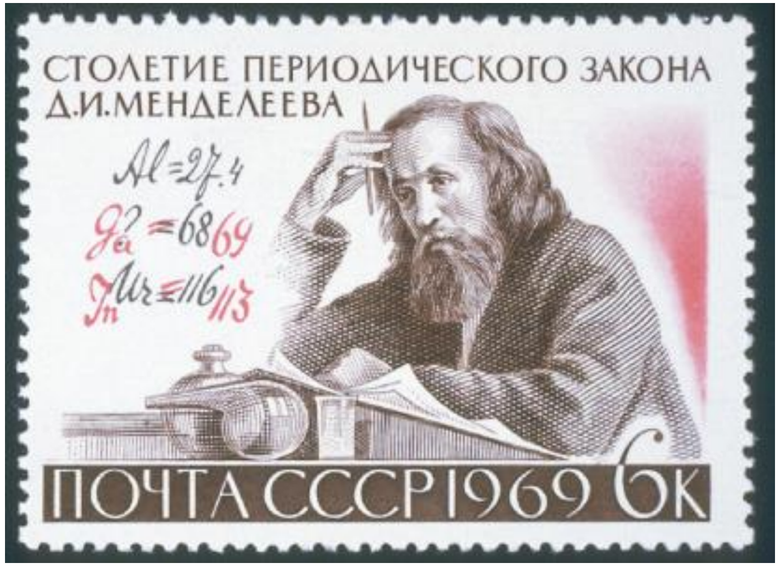
Atoms Lecture 5: The Periodic Table

Learning Objective	Openstax 2e Chapter
Modern Periodic Table	6.5
Size of Atoms	6.5
Effective Nuclear Charge	6.5
Ionization Energy	6.5
Electron Affinity	6.5



Dimitri Mendeleev (1834-1907)

Suggested Practice Problems

Chapter 6 Exercises – Questions: 67, 69, 71, 73, 75, 77, 81, 83, 85

Answers can be found in the Chapter 6 Answer Key

Mendeleev’s Periodic Law (1869)

When elements are arranged in the order of increasing atomic mass, certain properties occur periodically.

Reihen	Gruppe I. — R^2O	Gruppe II. — RO	Gruppe III. — R^2O^3	Gruppe IV. RH^4 RO^2	Gruppe V. RH^3 R^2O^5	Gruppe VI. RH^2 RO^3	Gruppe VII. RH R^2O^7	Gruppe VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	– = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	– = 68	– = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	– = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	–	–	–	– – – –
9	(–)	–	–	–	–	–	–	
10	–	–	?Er = 178	?La = 180	Ta = 182	W = 184	–	Os = 195, Ir = 197, Pt = 198, Au = 199
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208			
12	–	–	–	Th = 231	–	U = 240		

In Mendeleev’s table (1871), similar elements fall in vertical groups, and the properties of the elements (e.g., molar volume) change gradually from top to bottom in the group. He left blank spaces for undiscovered elements.

Modern Periodic Table

Similar properties recur periodically when elements are arranged according to increasing atomic number.

Period	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 H 1.008																	2 He 4.0026
2	3 Li 6.94	4 Be 9.0122											5 B 10.81	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
3	11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.085	15 P 30.974	16 S 32.06	17 Cl 35.45	18 Ar 39.948
4	19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.38	31 Ga 69.723	32 Ge 72.630	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.798
5	37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.96	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57–71 La–Lu	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)	89–103 Ac–Lr	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Ds (271)	111 Rg (272)	112 Cn		114 Fl		116 Lv		

*Lanthanide series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	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 173.05	71 Lu 174.97
†Actinide series	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)	103 Lr (262)

The basis of the periodic table is the electron configurations of elements (provided by quantum mechanics).

Modern Periodic Table

Similar properties recur periodically when elements are arranged according to increasing atomic number.

1	2															3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																															2 He 4.0026
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55 Cs 132.91	56 Ba 137.33	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm 144.91	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 173.05	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 208.98	85 At 209.99	86 Rn 222.02
87 Fr 223.02	88 Ra 226.03	89 Ac 227.03	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu 244.06	95 Am 243.06	96 Cm 247.07	97 Bk 247.07	98 Cf 251.08	99 Es 252.08	100 Fm 257.10	101 Md 258.10	102 No 259.10	103 Lr 262.11	104 Rf 267.12	105 Db 270.13	106 Sg 269.13	107 Bh 270.13	108 Hs 269.13	109 Mt 278.16	110 Ds 281.17	111 Rg 281.17	112 Cn 285.18	113 Nh 286.18	114 Fl 289.19	115 Mc 289.20	116 Lv 293.20	117 Ts 293.21	118 Og 294.21

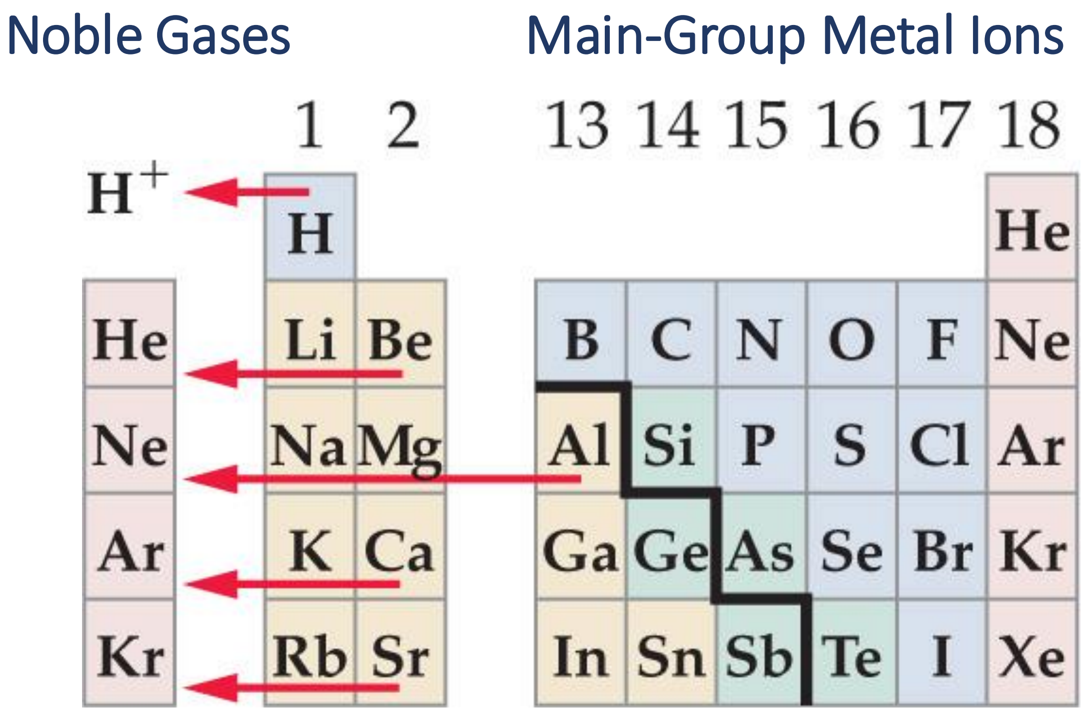
Periodic Table

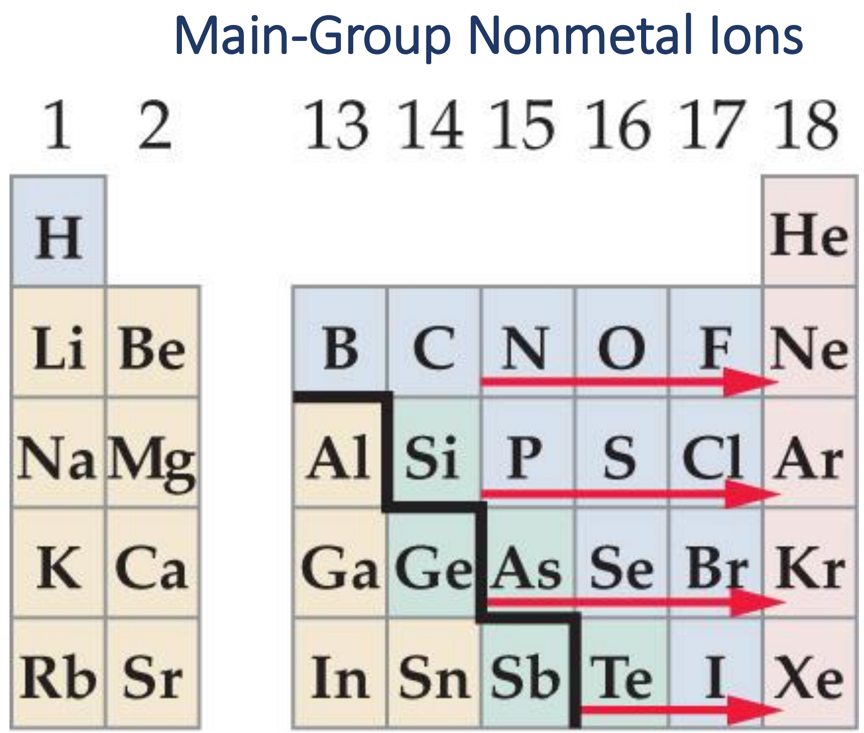
www.webelements.com

Long format (32-column) of the periodic table. This is the standard form of the periodic split between the Groups 2 and 3 to allow the f-block elements to be inserted into the resulting space.

Metals, Nonmetals, and Their Ions

Atoms of elements in the same group have similar electron configurations.



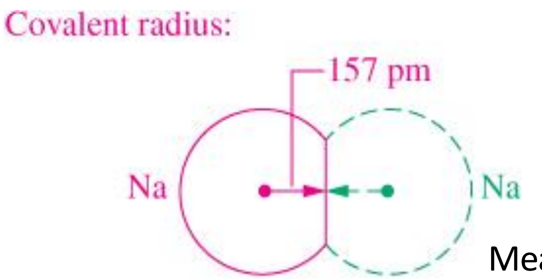


- Metals tend to lose electrons to attain noble gas electron configurations as cations.

- Nonmetals tend to gain electrons to attain noble-gas electron configurations as anions.

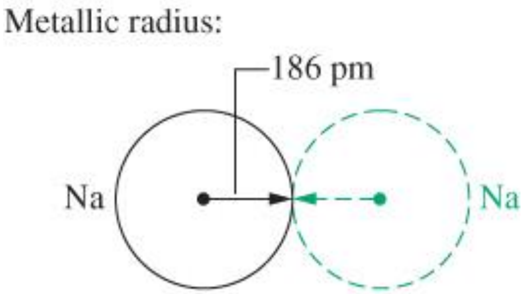
Size of Atoms

Unfortunately, atomic radius is hard to define because an atom has no precise outer boundary.



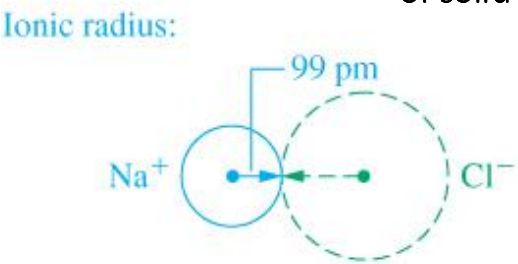
Measured for an isolated Na_2 cluster

- **Covalent radius** is one half of the distance between the nuclei of two identical atoms joined by a single covalent bond.



Measured for a piece of solid Na metal

- **Metallic radius** is one-half the distance between the nuclei of two atoms in contact in the crystalline solid metal.
- **van der Waals radius** is similar to metallic radius except is for solid samples of noble gases.



- **Ionic radius** is based on the distance between the nuclei of ions joined by an ionic bond.
(properly apportioned by assigning $r(O_2^{2-})$ = 140 pm)

Figure 1: Covalent, metallic, and ionic radius.

Size of Atoms

Atomic radius decreases from left to right through a row and increases from top to bottom through a column.

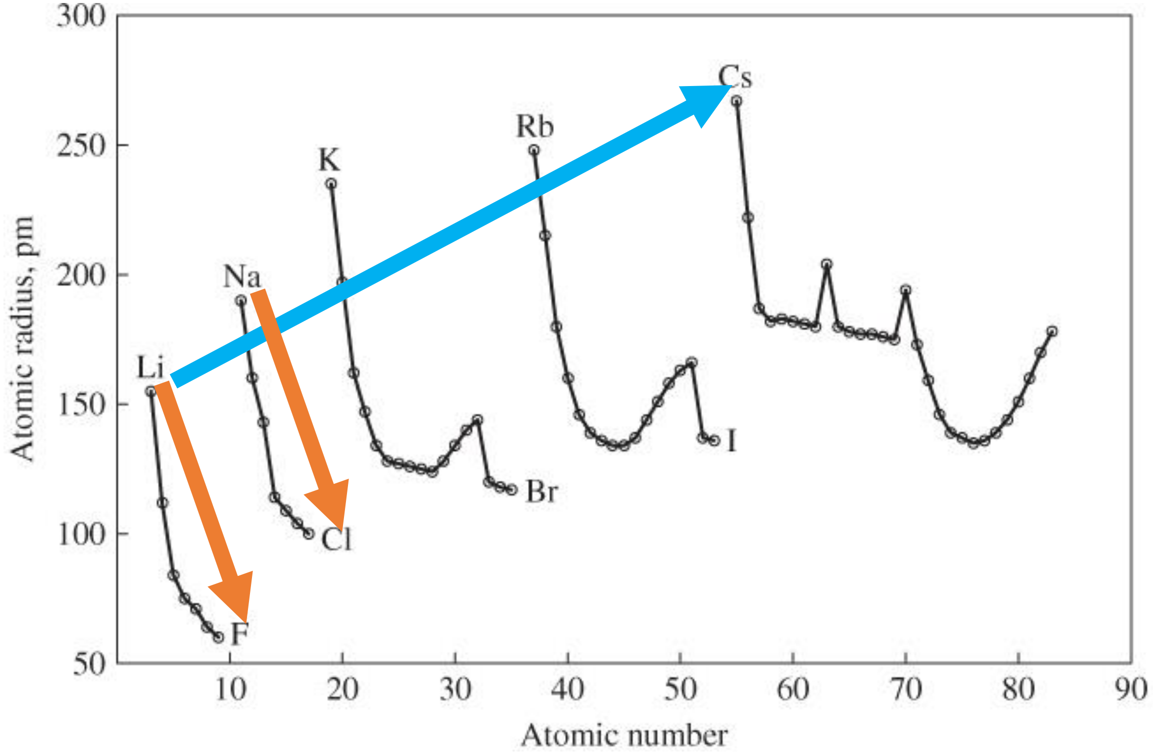
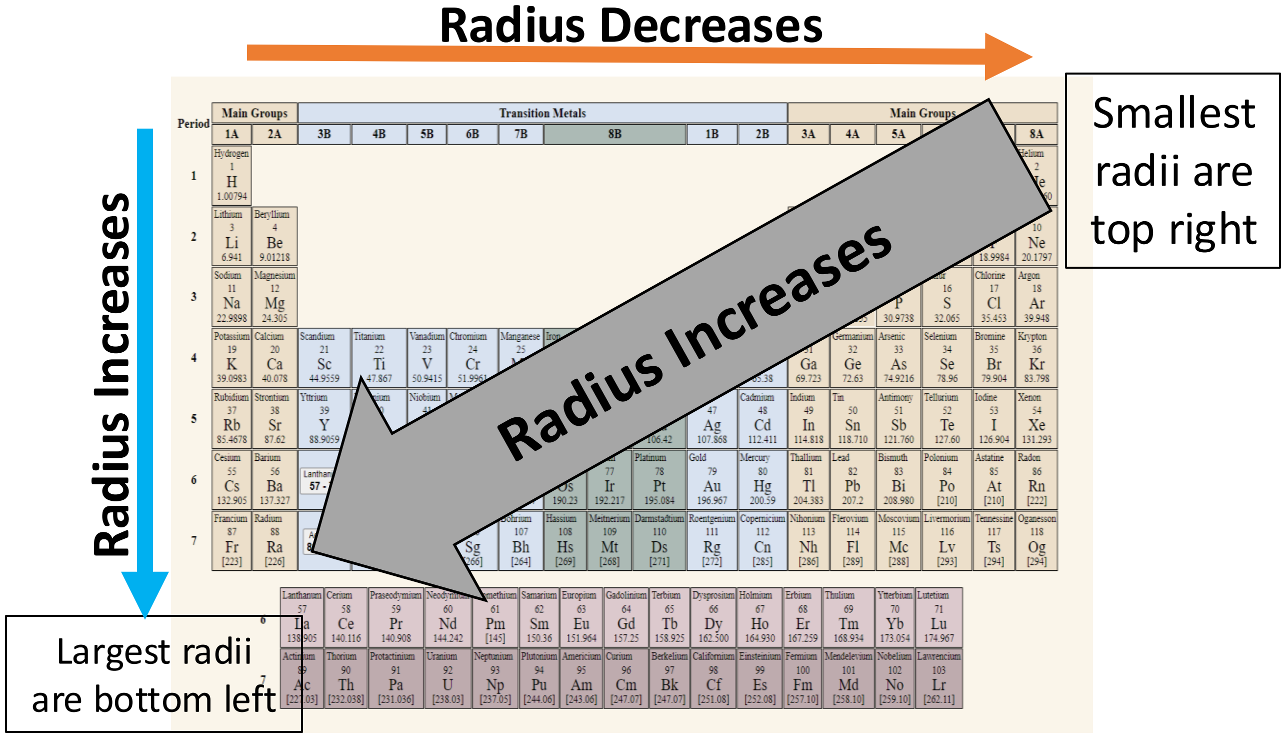
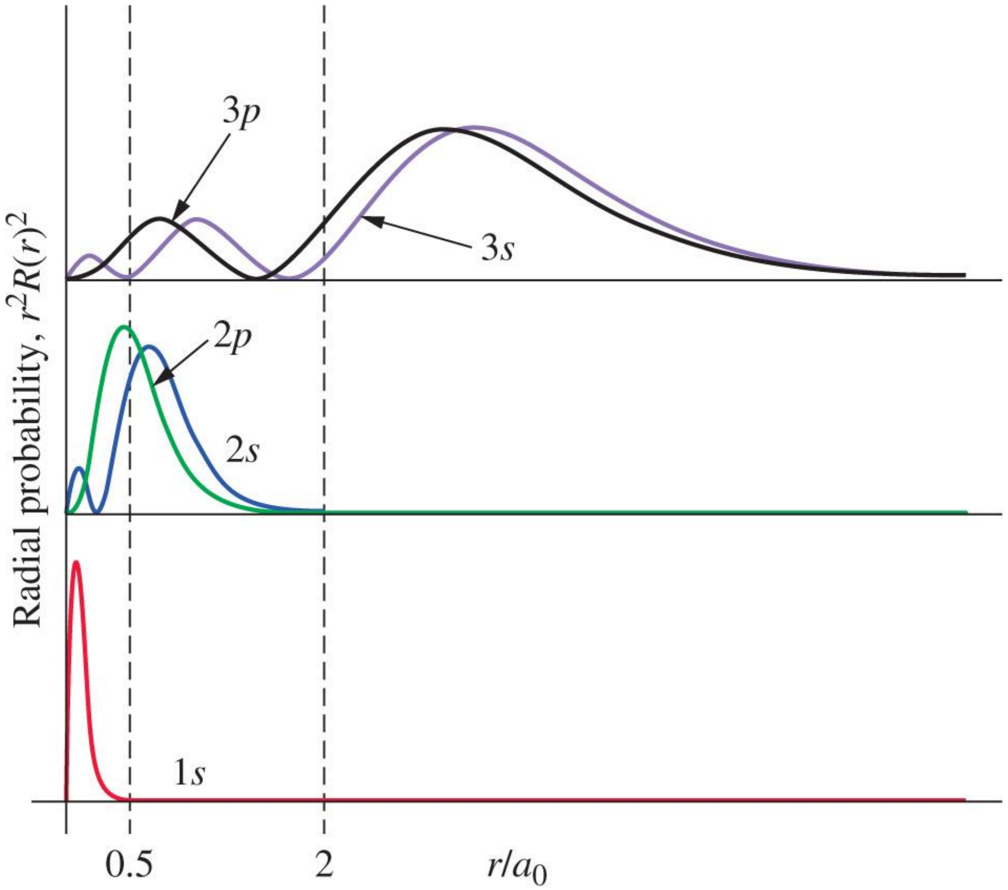


Figure 2: Metallic radii for metals and covalent radii for nonmetals.



Effective Nuclear Charge

The core electrons of an atom shield or screen the valence electrons from experiencing the full nuclear charge.



Note: a_0 : Bohr radius = 52.9 pm

Figure 3: Radial distribution functions for aluminum.

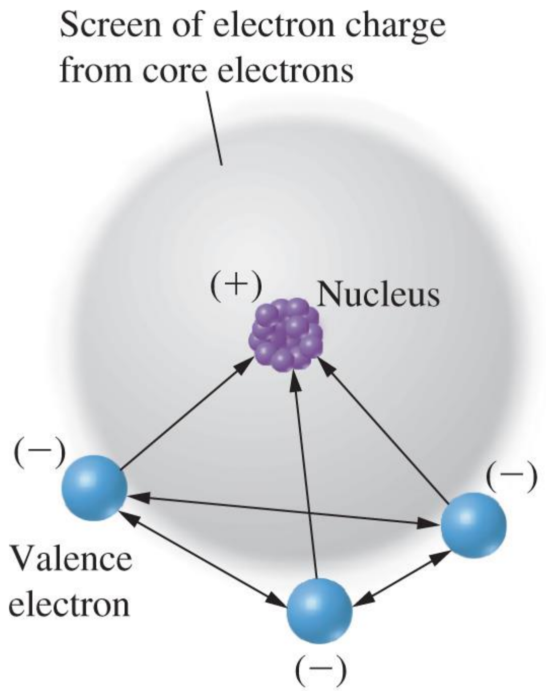


Figure 4: The shielding effect and effective nuclear charge, Z_{eff}

$$F = -\frac{Z_{eff}e^2}{r^2}$$

Effective Nuclear Charge

The core electrons of an atom shield or screen the valence electrons from experiencing the full nuclear charge.

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

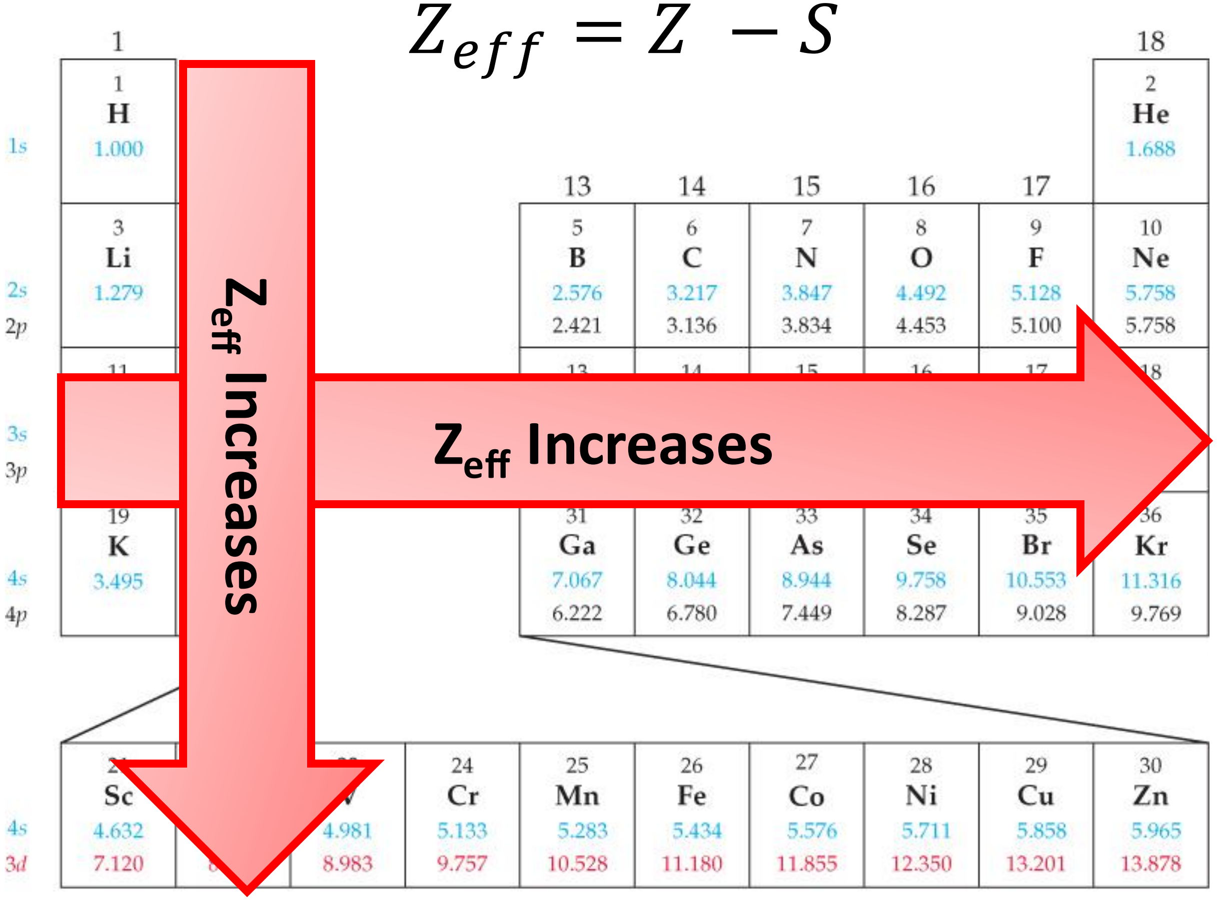


Figure 5: Effective nuclear charges. Z_{eff}, of valence electrons.

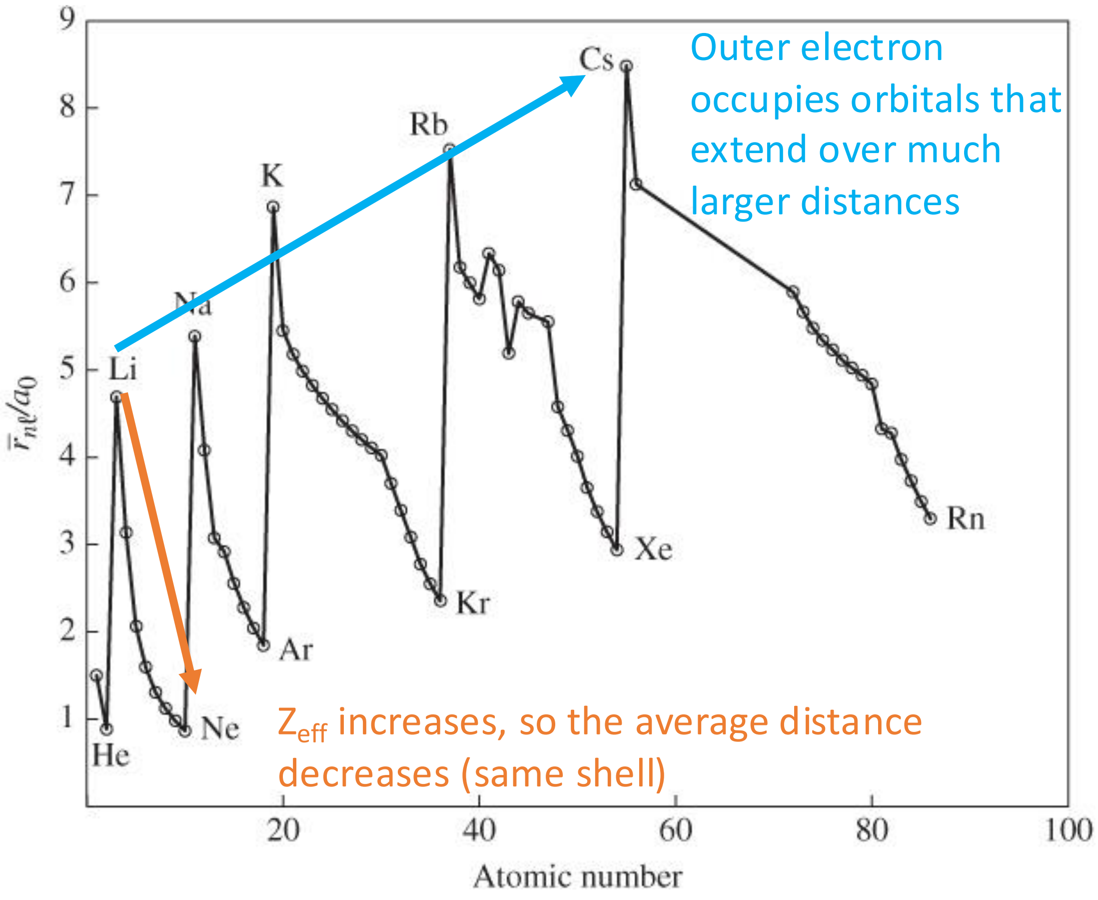
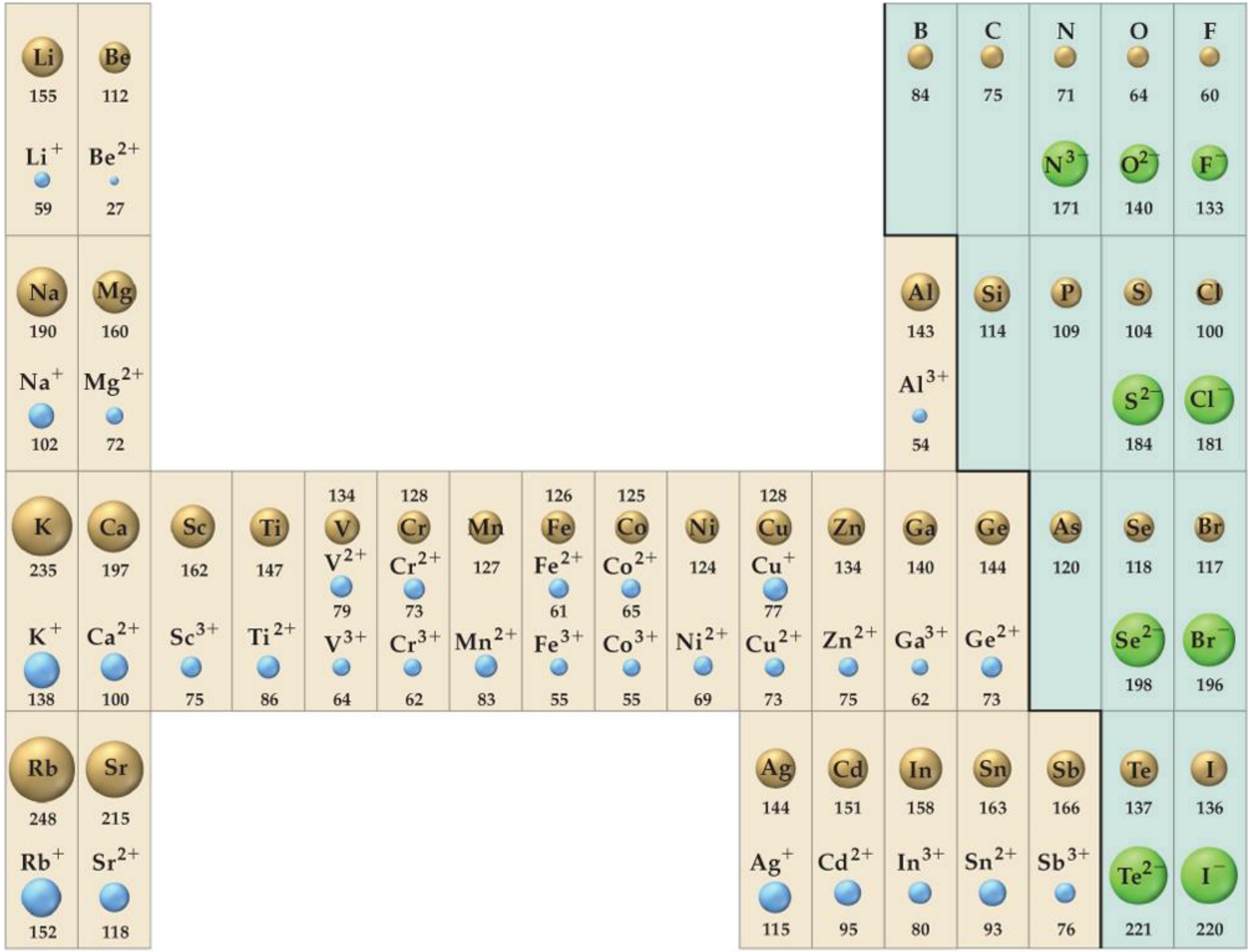


Figure 6: The average distance from the nucleus for the least strongly bound electron

Size of Ions

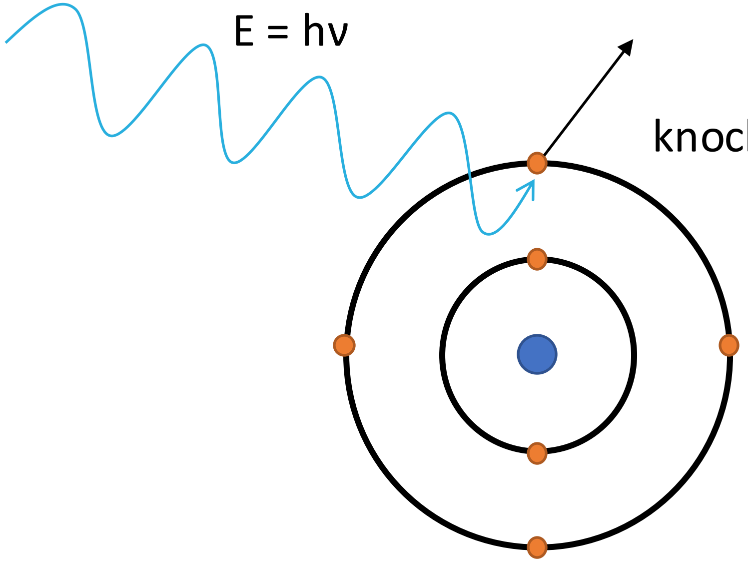
- **<u>Cations</u>** are smaller than the atoms from which they are formed.
- For **<u>isoelectronic cations</u>**, the more positive the ionic charge, the smaller the ionic radius.
- **<u>Anions</u>** are larger than the atoms from which they are formed.
- For **<u>isoelectronic anions</u>**, the more negative the charge, the larger the ionic radius.

Figure 7: A comparison of some atomic and ionic radii in picometers (pm)



Ionization Energy

The quantity of energy a gaseous atom must absorb to be able to expel an electron.



Example: 1^{st} and 2^{nd} Ionization of Magnesium

$$Mg(g) \rightarrow Mg^{+}(g) + e^{-} \qquad I_1 = 738 \text{ kJ}$$

$$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-} \qquad I_2 = 1451 \text{ kJ}$$

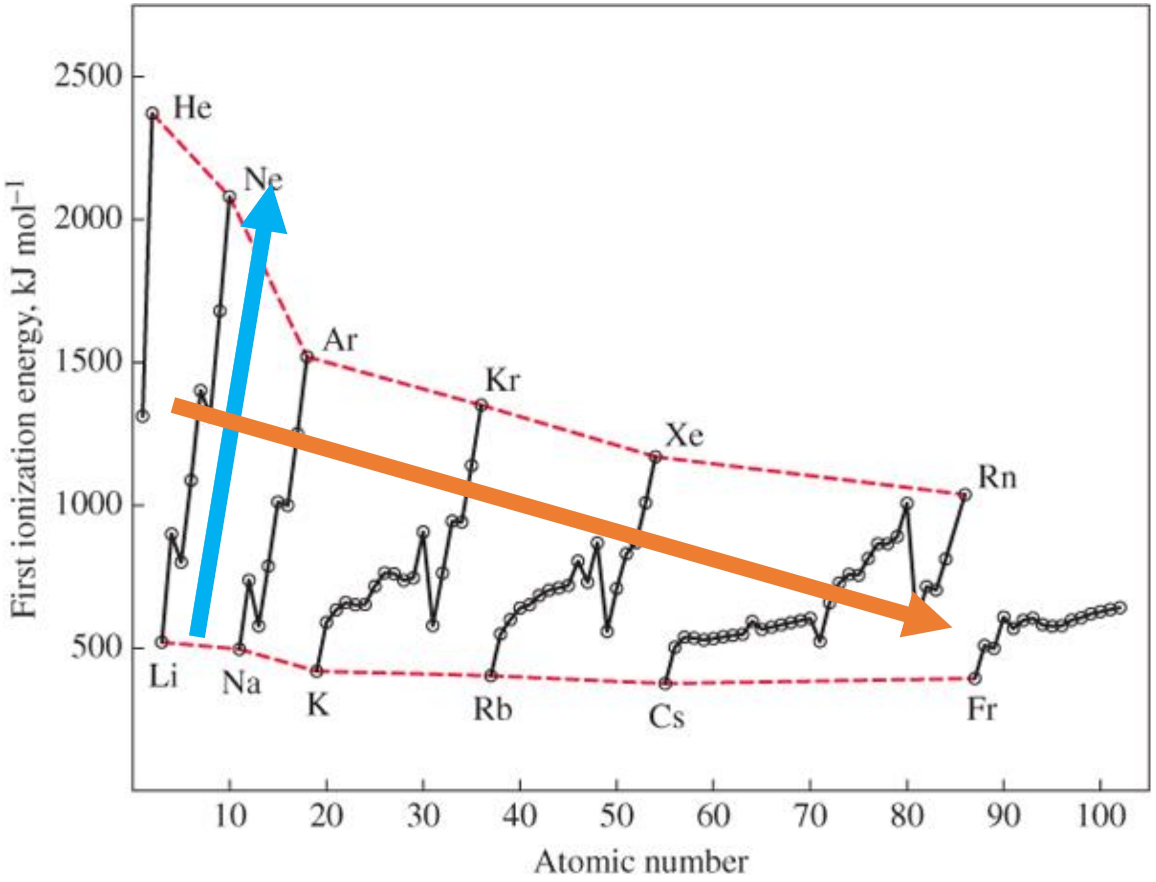


Figure 7: Ionization energy for each element.

With relatively few exceptions, ionization energies increase from left to right across a period and decrease from top to bottom within a group.

Ionization Energy

Ionization energies generally increase across each period (row)
- breaks in the pattern can be understood by examining electron configurations

Figure 8: First ionization energies of the 3rd-row elements.

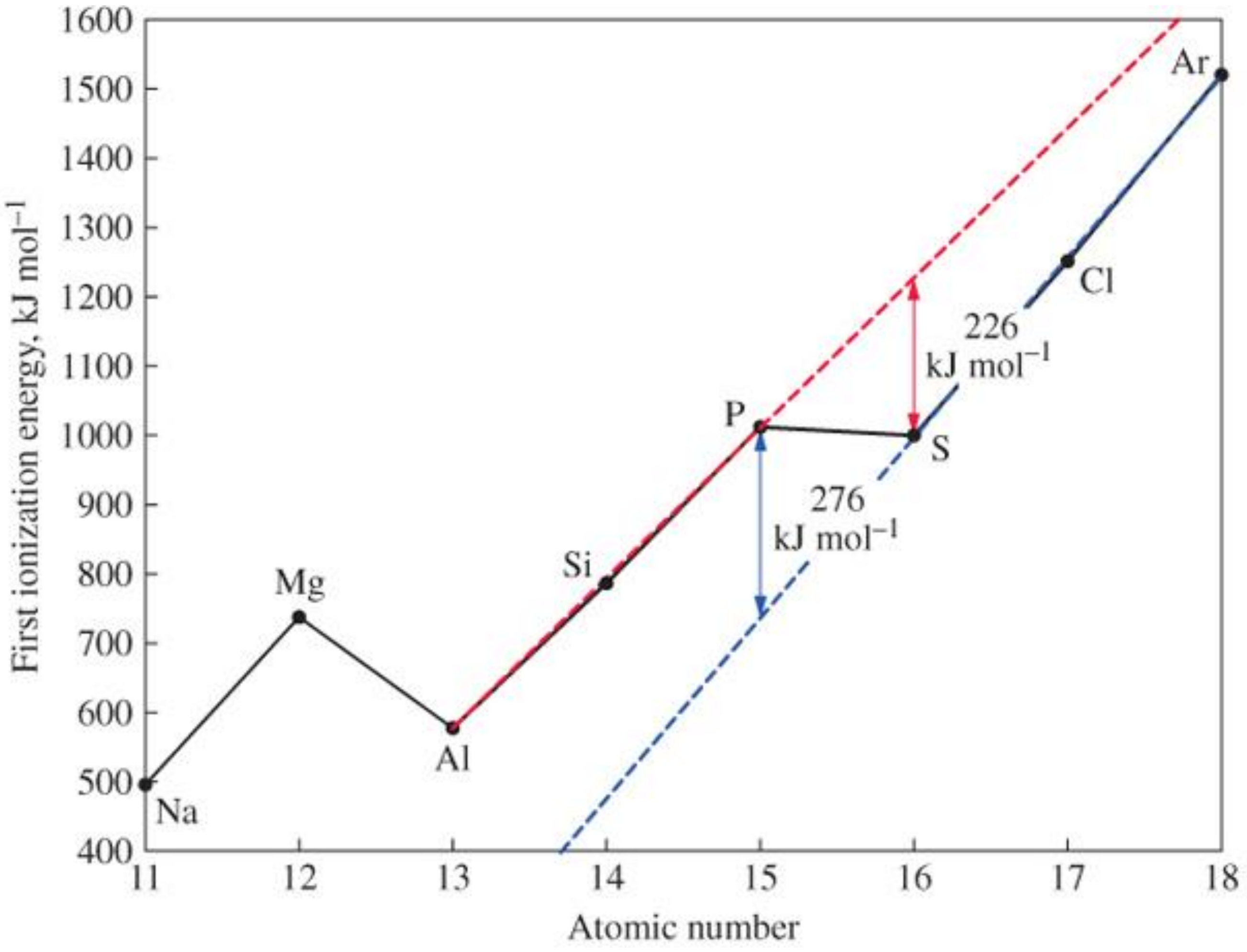
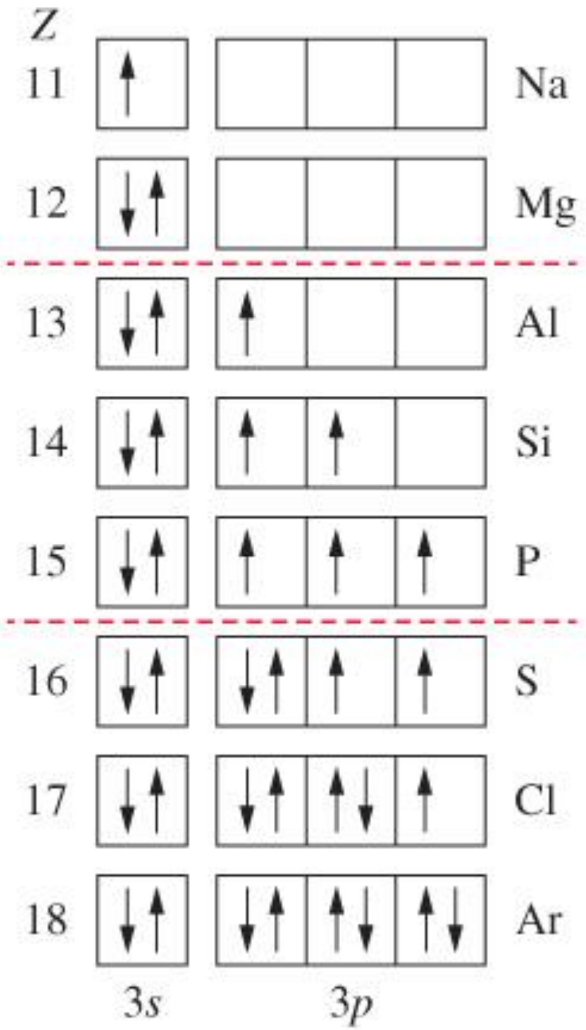


Figure 9: Valence electron configurations.



Removing 3s

Removing unpaired 3p

Removing paired 3p

Electron Affinity

The energy change that occurs when an atom in the gas phase gains an electron.

$$X(g) + e^- \longrightarrow X^-(g)$$

1	2	13	14	15	16	17	18
H −72.8							He >0
Li −59.6	Be >0	B −26.7	C −121.8	N +7	O −141.0	F −328.0	Ne >0
Na −52.9	Mg >0	Al −42.5	Si −133.6	P −72	S −200.4	Cl −349.0	Ar >0
K −48.4	Ca −2.37	Ga −28.9	Ge −119.0	As −78	Se −195.0	Br −324.6	Kr >0
Rb −46.9	Sr −5.03	In −28.9	Sn −107.3	Sb −103.2	Te −190.2	I −295.2	Xe >0
Cs −45.5	Ba −13.95	Tl −19.2	Pb −35.1	Bi −91.2	Po −186	At −270	Rn >0

Figure 10: Electron affinities (kJ/mol) of main-group elements.

The First Electron Affinity of O is negative (more stable) due to the strong electron affinity of the neutral O atom.

$O(g) + e^- \rightarrow O^-(g)$ EA = −141 kJ

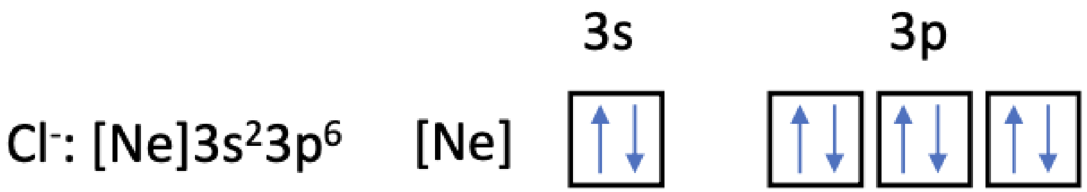
The Second Electron Affinity is positive due to the strong repulsive force between the second electron and the negative O ion.

$O^-(g) + e^- \rightarrow O^{2-}(g)$ EA = +744 kJ

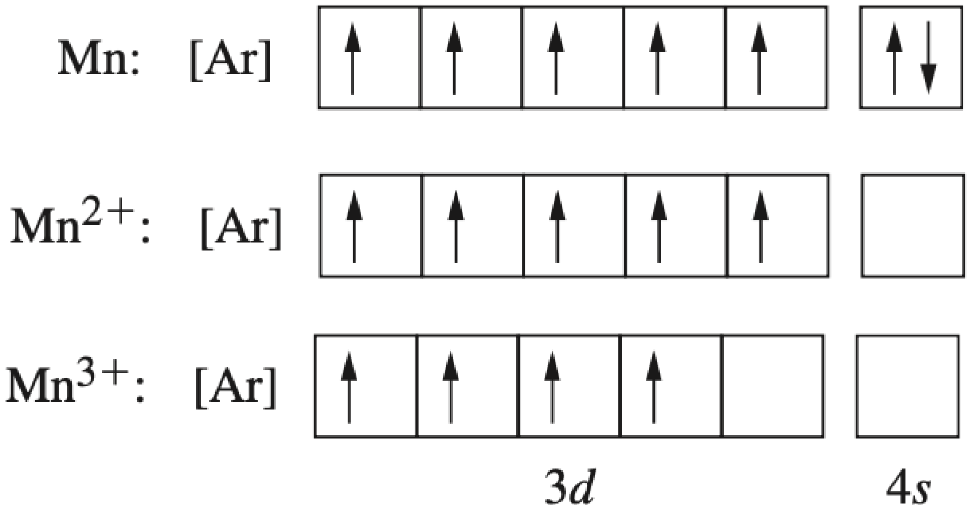
Magnetic Properties

A key property related to electron configurations of atoms and ions is their behaviour in a magnetic field.

- **Diamagnetic** atoms or ions:
  - All *e*⁻ are paired.
  - Weakly repelled by a magnetic field.



- **Paramagnetic** atoms or ions:
  - *Unpaired e* ⁻.
  - Attracted to an external magnetic field. The more unpaired electrons present, the stronger is this attraction.



Polarizability

For an atom in the vicinity of another atom, molecule, or ion, the electron cloud is distorted.

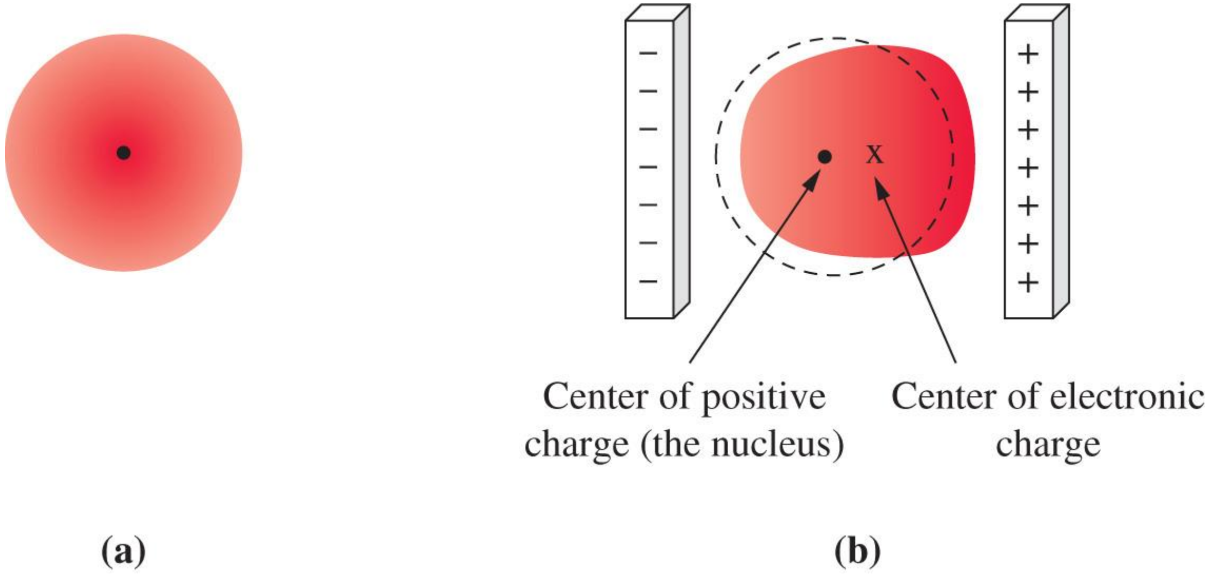


Figure 11:

(a) For an isolated atom, the distribution of electronic charge about the nucleus is spherical.

(b) In an electric field, the distribution of electronic charge is non-spherical, and the centers of positive and negative charge no longer coincide. The atom is said to be polarized.

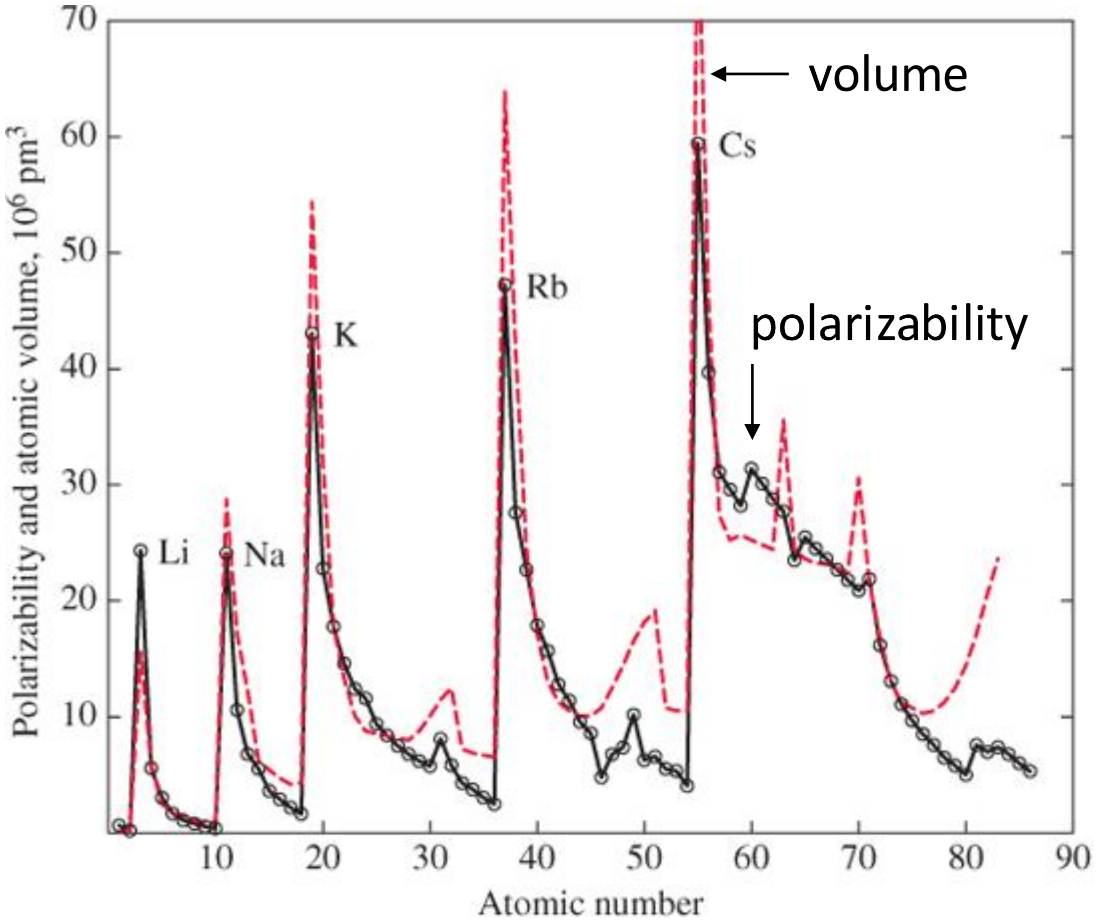


Figure 12: Polarizability of an atom is similar in magnitude to the atomic volume calculated from atomic radii.

Atomic Properties & the Periodic Table

Summary