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Chapter 8

Periodic Relationships Among the Elements

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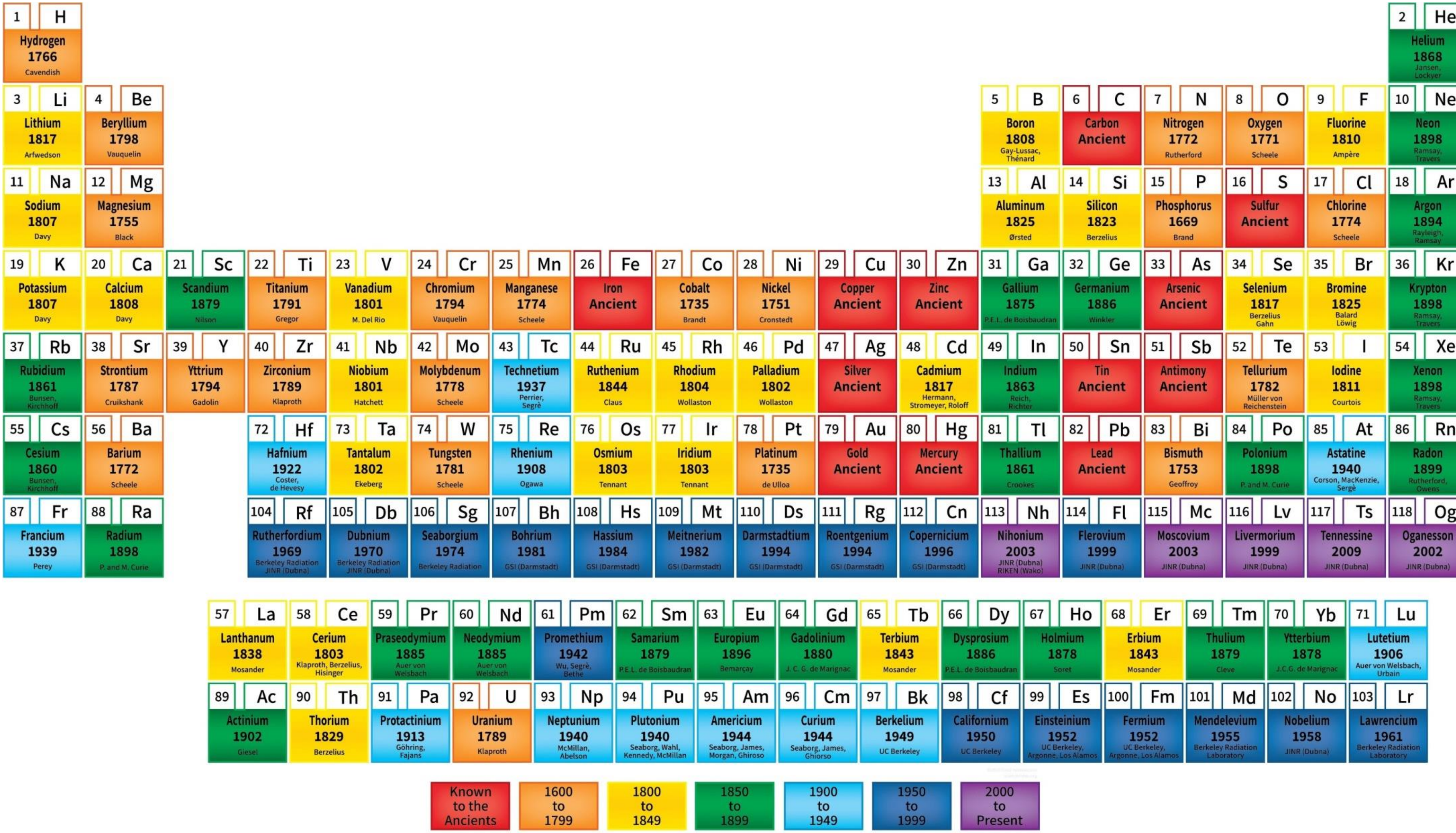
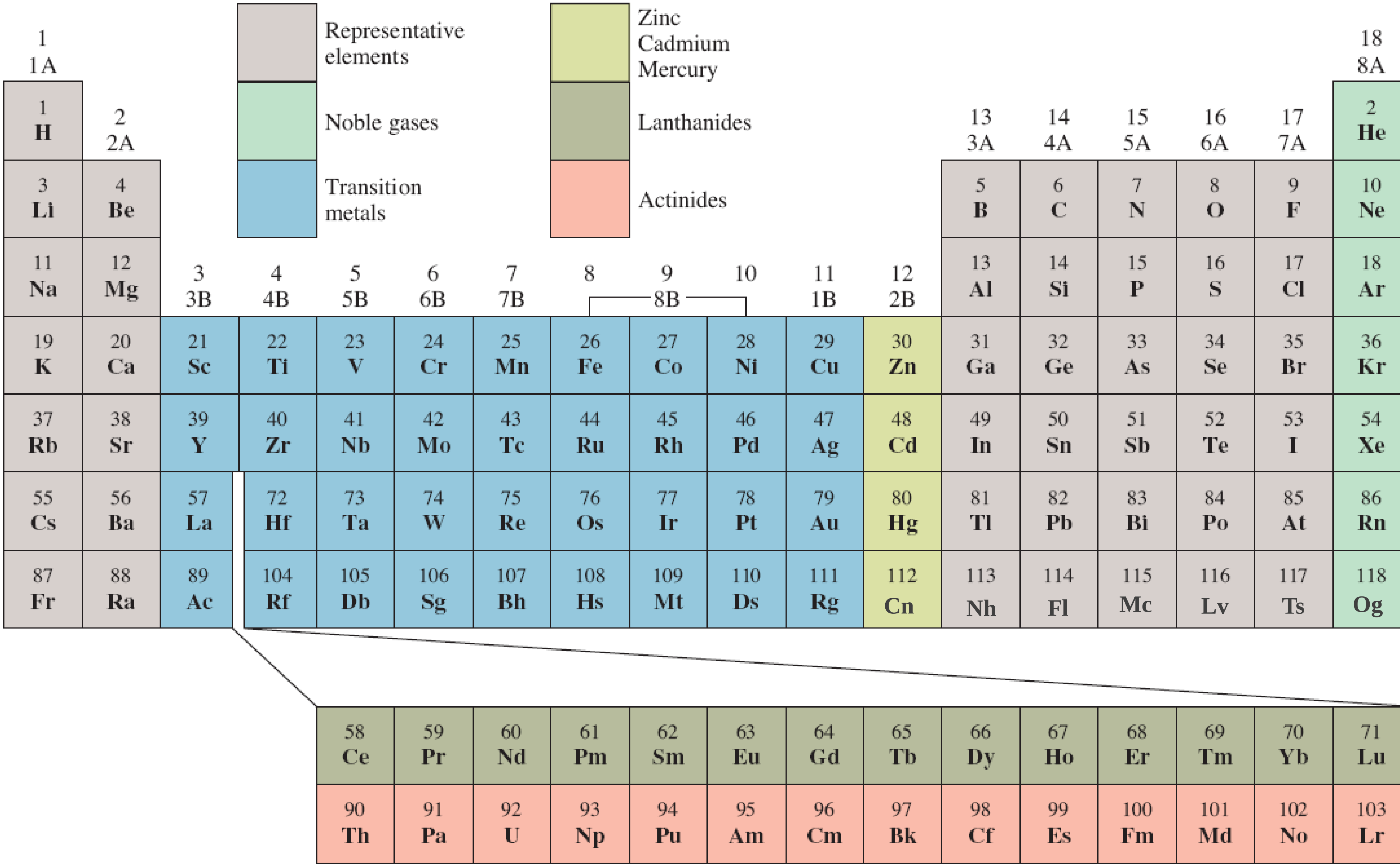
8.2
Periodic classification of the elements

When the Elements Were Discovered





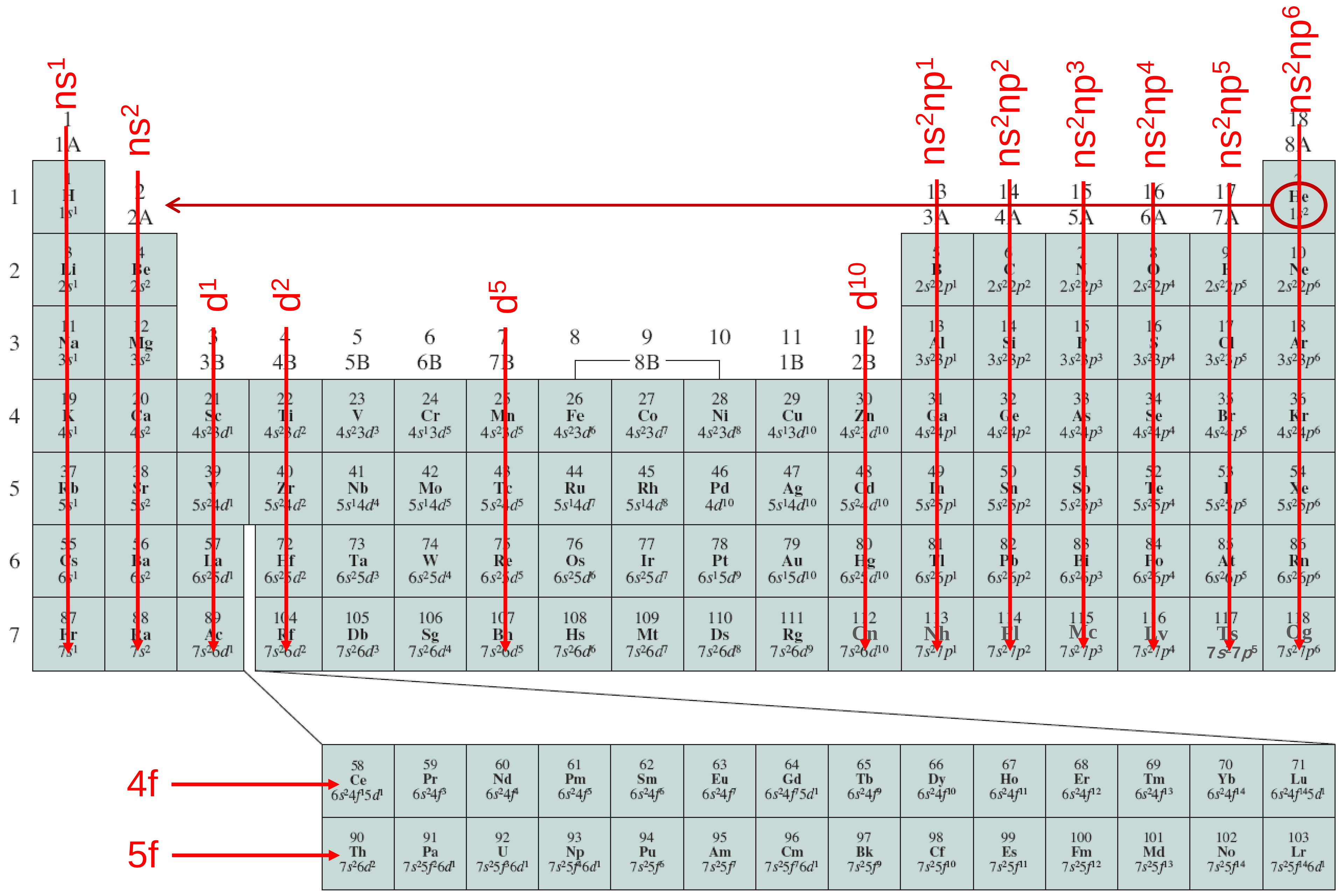
Classification of the elements.

Ground State Electron Configurations of the Elements

	1 (1A)	2 (2A)	3 (3B)	4 (4B)	5 (5B)	6 (6B)	7 (7B)	8 (8B)	9 (8B)	10 (8B)	11 (1B)	12 (2B)	13 (3A)	14 (4A)	15 (5A)	16 (6A)	17 (7A)	18 (8A)
1	1 **H** $1s^1$																	2 **He** $1s^2$
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** $2s^22p^6$
3	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** $3s^23p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^23d^1$	22 **Ti** $4s^23d^2$	23 **V** $4s^23d^3$	24 **Cr** $4s^13d^5$	25 **Mn** $4s^23d^5$	26 **Fe** $4s^23d^6$	27 **Co** $4s^23d^7$	28 **Ni** $4s^23d^8$	29 **Cu** $4s^13d^{10}$	30 **Zn** $4s^23d^{10}$	31 **Ga** $4s^24p^1$	32 **Ge** $4s^24p^2$	33 **As** $4s^24p^3$	34 **Se** $4s^24p^4$	35 **Br** $4s^24p^5$	36 **Kr** $4s^24p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^24d^1$	40 **Zr** $5s^24d^2$	41 **Nb** $5s^14d^4$	42 **Mo** $5s^14d^5$	43 **Tc** $5s^24d^5$	44 **Ru** $5s^14d^7$	45 **Rh** $5s^14d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^14d^{10}$	48 **Cd** $5s^24d^{10}$	49 **In** $5s^25p^1$	50 **Sn** $5s^25p^2$	51 **Sb** $5s^25p^3$	52 **Te** $5s^25p^4$	53 **I** $5s^25p^5$	54 **Xe** $5s^25p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57 **La** $6s^25d^1$	72 **Hf** $6s^25d^2$	73 **Ta** $6s^25d^3$	74 **W** $6s^25d^4$	75 **Re** $6s^25d^5$	76 **Os** $6s^25d^6$	77 **Ir** $6s^25d^7$	78 **Pt** $6s^15d^9$	79 **Au** $6s^15d^{10}$	80 **Hg** $6s^25d^{10}$	81 **Tl** $6s^26p^1$	82 **Pb** $6s^26p^2$	83 **Bi** $6s^26p^3$	84 **Po** $6s^26p^4$	85 **At** $6s^26p^5$	86 **Rn** $6s^26p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89 **Ac** $7s^26d^1$	104 **Rf** $7s^26d^2$	105 **Db** $7s^26d^3$	106 **Sg** $7s^26d^4$	107 **Bh** $7s^26d^5$	108 **Hs** $7s^26d^6$	109 **Mt** $7s^26d^7$	110 **Ds** $7s^26d^8$	111 **Rg** $7s^26d^9$	112 **Cn** $7s^26d^{10}$	113 **Nh** $7s^27p^1$	114 **Fl** $7s^27p^2$	115 **Mc** $7s^27p^3$	116 **Lv** $7s^27p^4$	117 **Ts** $7s^27p^5$	118 **Og** $7s^27p^6$

58 **Ce** $6s^24f^15d^1$	59 **Pr** $6s^24f^3$	60 **Nd** $6s^24f^4$	61 **Pm** $6s^24f^5$	62 **Sm** $6s^24f^6$	63 **Eu** $6s^24f^7$	64 **Gd** $6s^24f^75d^1$	65 **Tb** $6s^24f^9$	66 **Dy** $6s^24f^{10}$	67 **Ho** $6s^24f^{11}$	68 **Er** $6s^24f^{12}$	69 **Tm** $6s^24f^{13}$	70 **Yb** $6s^24f^{14}$	71 **Lu** $6s^24f^{14}5d^1$
90 **Th** $7s^26d^2$	91 **Pa** $7s^25f^26d^1$	92 **U** $7s^25f^36d^1$	93 **Np** $7s^25f^46d^1$	94 **Pu** $7s^25f^6$	95 **Am** $7s^25f^7$	96 **Cm** $7s^25f^76d^1$	97 **Bk** $7s^25f^9$	98 **Cf** $7s^25f^{10}$	99 **Es** $7s^25f^{11}$	100 **Fm** $7s^25f^{12}$	101 **Md** $7s^25f^{13}$	102 **No** $7s^25f^{14}$	103 **Lr** $7s^25f^{14}6d^1$

The ground-state electron configurations of the elements. For simplicity, only the configurations of the outer electrons are shown.



ns1
ns2
ns2np1
ns2np2
ns2np3
ns2np4
ns2np5
ns2np6
d1
d2
d5
d10
4f
5f
1 1A
2 2A
13 3A
14 4A
15 5A
16 6A
17 7A
18 8A
3 3B
4 4B
5 5B
6 6B
7 7B
8 9 10 8B
11 1B
12 2B
1 2 3 4 5 6 7
1 H 1s1
2 He 1s2
3 Li 2s1
4 Be 2s2
5 B 2s22p1
6 C 2s22p2
7 N 2s22p3
8 O 2s22p4
9 F 2s22p5
10 Ne 2s22p6
11 Na 3s1
12 Mg 3s2
13 Al 3s23p1
14 Si 3s23p2
15 P 3s23p3
16 S 3s23p4
17 Cl 3s23p5
18 Ar 3s23p6
19 K 4s1
20 Ca 4s2
21 Sc 4s23d1
22 Ti 4s23d2
23 V 4s23d3
24 Cr 4s13d5
25 Mn 4s23d5
26 Fe 4s23d6
27 Co 4s23d7
28 Ni 4s23d8
29 Cu 4s13d10
30 Zn 4s23d10
31 Ga 4s24p1
32 Ge 4s24p2
33 As 4s24p3
34 Se 4s24p4
35 Br 4s24p5
36 Kr 4s24p6
37 Rb 5s1
38 Sr 5s2
39 Y 5s24d1
40 Zr 5s24d2
41 Nb 5s14d4
42 Mo 5s14d5
43 Tc 5s24d5
44 Ru 5s14d7
45 Rh 5s14d8
46 Pd 4d10
47 Ag 5s14d10
48 Cd 5s24d10
49 In 5s25p1
50 Sn 5s25p2
51 Sb 5s25p3
52 Te 5s25p4
53 I 5s25p5
54 Xe 5s25p6
55 Cs 6s1
56 Ba 6s2
57 La 6s25d1
72 Hf 6s25d2
73 Ta 6s25d3
74 W 6s25d4
75 Re 6s25d5
76 Os 6s25d6
77 Ir 6s25d7
78 Pt 6s15d9
79 Au 6s15d10
80 Hg 6s25d10
81 Tl 6s26p1
82 Pb 6s26p2
83 Bi 6s26p3
84 Po 6s26p4
85 At 6s26p5
86 Rn 6s26p6
87 Fr 7s1
88 Ra 7s2
89 Ac 7s26d1
104 Rf 7s26d2
105 Db 7s26d3
106 Sg 7s26d4
107 Bh 7s26d5
108 Hs 7s26d6
109 Mt 7s26d7
110 Ds 7s26d8
111 Rg 7s26d9
112 Cn 7s26d10
113 Nh 7s27p1
114 Fl 7s27p2
115 Mc 7s27p3
116 Lv 7s27p4
117 Ts 7s27p5
118 Og 7s27p6
58 Ce 6s24f15d1
59 Pr 6s24f3
60 Nd 6s24f4
61 Pm 6s24f5
62 Sm 6s24f6
63 Eu 6s24f7
64 Gd 6s24f75d1
65 Tb 6s24f9
66 Dy 6s24f10
67 Ho 6s24f11
68 Er 6s24f12
69 Tm 6s24f13
70 Yb 6s24f14
71 Lu 6s24f145d1
90 Th 7s26d2
91 Pa 7s25f26d1
92 U 7s25f36d1
93 Np 7s25f46d1
94 Pu 7s25f6
95 Am 7s25f7
96 Cm 7s25f76d1
97 Bk 7s25f9
98 Cf 7s25f10
99 Es 7s25f11
100 Fm 7s25f12
101 Md 7s25f13
102 No 7s25f14
103 Lr 7s25f146d1

Electron Configurations of Cations and Anions

Ions Derived from Representative Elements

Atoms lose or gain electrons so that cation or anions have a noble-gas outer electron configuration.

In the formation of a cation from the atom of a representative element, one or more electrons are removed from the highest occupied *n* shell.

Na: [Ne] $3s^1$	Na^+: [Ne]
Ca: [Ar] $4s^2$	Ca^{2+}: [Ar]
Al: [Ne] $3s^2$ $3p^1$	Al^{3+}: [Ne]

H: $1s^1$	H^-: $1s^2$ or [He]
F: $1s^2 2s^2 2p^5$	F^-: $1s^2 2s^2 2p^6$ or [Ne]
O: $1s^2 2s^2 2p^4$	O^{2-}: $1s^2 2s^2 2p^6$ or [Ne]
N: $1s^2 2s^2 2p^3$	N^{3-}: $1s^2 2s^2 2p^6$ or [Ne]

In the formation of an anion, one or more electrons are added to the highest partially filled *n* shell.

Cations and Anions Of Representative Elements

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	+2											+3		-3	-2	-1	
1	1 H $1s^1$																	2 He $1s^2$
2	3 Li $2s^1$	4 Be $2s^2$											5 B $2s^2 2p^1$	6 C $2s^2 2p^2$	7 N $2s^2 2p^3$	8 O $2s^2 2p^4$	9 F $2s^2 2p^5$	10 Ne $2s^2 2p^6$
3	11 Na $3s^1$	12 Mg $3s^2$											13 Al $3s^2 3p^1$	14 Si $3s^2 3p^2$	15 P $3s^2 3p^3$	16 S $3s^2 3p^4$	17 Cl $3s^2 3p^5$	18 Ar $3s^2 3p^6$
4	19 K $4s^1$	20 Ca $4s^2$	21 Sc $4s^2 3d^1$	22 Ti $4s^2 3d^2$	23 V $4s^2 3d^3$	24 Cr $4s^1 3d^5$	25 Mn $4s^2 3d^5$	26 Fe $4s^2 3d^6$	27 Co $4s^2 3d^7$	28 Ni $4s^2 3d^8$	29 Cu $4s^1 3d^{10}$	30 Zn $4s^2 3d^{10}$	31 Ga $4s^2 4p^1$	32 Ge $4s^2 4p^2$	33 As $4s^2 4p^3$	34 Se $4s^2 4p^4$	35 Br $4s^2 4p^5$	36 Kr $4s^2 4p^6$
5	37 Rb $5s^1$	38 Sr $5s^2$	39 Y $5s^2 4d^1$	40 Zr $5s^2 4d^2$	41 Nb $5s^1 4d^4$	42 Mo $5s^1 4d^5$	43 Tc $5s^2 4d^5$	44 Ru $5s^1 4d^7$	45 Rh $5s^1 4d^8$	46 Pd $4d^{10}$	47 Ag $5s^1 4d^{10}$	48 Cd $5s^2 4d^{10}$	49 In $5s^2 5p^1$	50 Sn $5s^2 5p^2$	51 Sb $5s^2 5p^3$	52 Te $5s^2 5p^4$	53 I $5s^2 5p^5$	54 Xe $5s^2 5p^6$
6	55 Cs $6s^1$	56 Ba $6s^2$	57 La $6s^2 5d^1$	72 Hf $6s^2 5d^2$	73 Ta $6s^2 5d^3$	74 W $6s^2 5d^4$	75 Re $6s^2 5d^5$	76 Os $6s^2 5d^6$	77 Ir $6s^2 5d^7$	78 Pt $6s^1 5d^9$	79 Au $6s^1 5d^{10}$	80 Hg $6s^2 5d^{10}$	81 Tl $6s^2 6p^1$	82 Pb $6s^2 6p^2$	83 Bi $6s^2 6p^3$	84 Po $6s^2 6p^4$	85 At $6s^2 6p^5$	86 Rn $6s^2 6p^6$
7	87 Fr $7s^1$	88 Ra $7s^2$	89 Ac $7s^2 6d^1$	104 Rf $7s^2 6d^2$	105 Db $7s^2 6d^3$	106 Sg $7s^2 6d^4$	107 Bh $7s^2 6d^5$	108 Hs $7s^2 6d^6$	109 Mt $7s^2 6d^7$	110 Ds $7s^2 6d^8$	111 Rg $7s^2 6d^9$	112 Cn $7s^2 6d^{10}$	113 Nh $7s^2 7p^1$	114 Fl $7s^2 7p^2$	115 Mc $7s^2 7p^3$	116 Lv $7s^2 7p^4$	117 Ts $7s^2 7p^5$	118 Og $7s^2 7p^6$

58 Ce $6s^2 4f^1 5d^1$	59 Pr $6s^2 4f^3$	60 Nd $6s^2 4f^4$	61 Pm $6s^2 4f^5$	62 Sm $6s^2 4f^6$	63 Eu $6s^2 4f^7$	64 Gd $6s^2 4f^7 5d^1$	65 Tb $6s^2 4f^9$	66 Dy $6s^2 4f^{10}$	67 Ho $6s^2 4f^{11}$	68 Er $6s^2 4f^{12}$	69 Tm $6s^2 4f^{13}$	70 Yb $6s^2 4f^{14}$	71 Lu $6s^2 4f^{14} 5d^1$
90 Th $7s^2 6d^2$	91 Pa $7s^2 5f^2 6d^1$	92 U $7s^2 5f^3 6d^1$	93 Np $7s^2 5f^4 6d^1$	94 Pu $7s^2 5f^6$	95 Am $7s^2 5f^7$	96 Cm $7s^2 5f^7 6d^1$	97 Bk $7s^2 5f^9$	98 Cf $7s^2 5f^{10}$	99 Es $7s^2 5f^{11}$	100 Fm $7s^2 5f^{12}$	101 Md $7s^2 5f^{13}$	102 No $7s^2 5f^{14}$	103 Lr $7s^2 5f^{14} 6d^1$

Isoelectronic atoms have the same number of electrons, and hence the same ground-state electron configuration.

For example:

H^- and He are isoelectronic.

For example:

Na^+: [Ne]

Al^{3+}: [Ne]

F^-: $1s^2 2s^2 2p^6$ or [Ne]

O^{2-}: $1s^2 2s^2 2p^6$ or [Ne]

N^{3-}: $1s^2 2s^2 2p^6$ or [Ne]

Na^+, Al^{3+}, F^-, O^{2-}, and N^{3-} are all isoelectronic with Ne.

Electron Configurations of Cations and Anions

Cations Derived from Transition Metals

When a cation is formed from an atom of a transition metal, electrons are always removed first from the ***ns*** orbital and then from the $\mathbf{(}\boldsymbol{n}\,\mathbf{-1)}\boldsymbol{d}$ orbitals.

For example:

Fe: $[Ar]4s^2 3d^6$

Fe^{2+}: $[Ar]4s^0 3d^6$ or $[Ar]3d^6$

Fe^{3+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

For example:

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

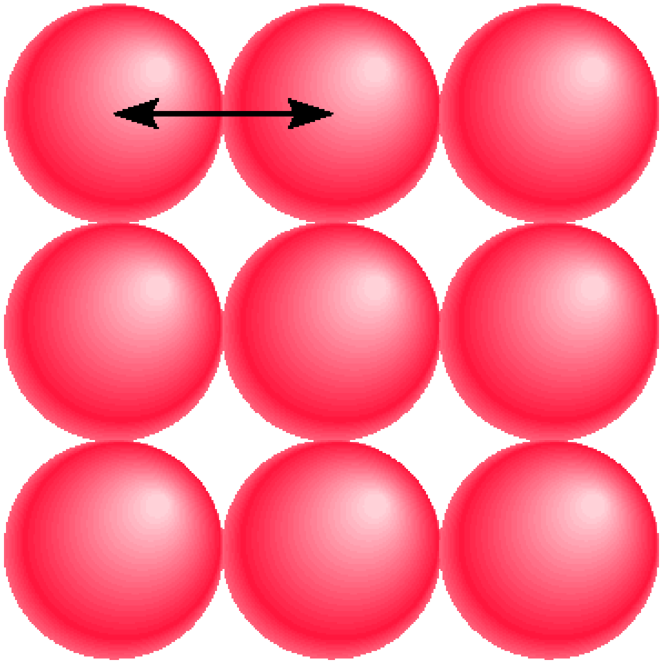
Mn^{2+}: $[Ar]4s^0 3d^5$ or $[Ar]3d^5$

Mn^{3+}: $[Ar]4s^0 3d^4$ or $[Ar]3d^4$

Mn^{4+}: $[Ar]4s^0 3d^3$ or $[Ar]3d^3$

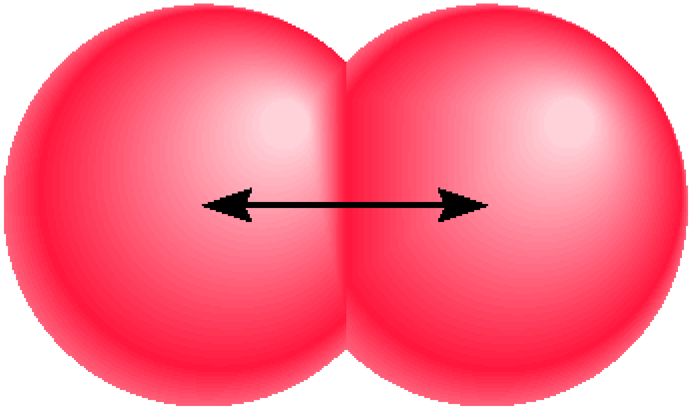
8.3
Periodic variation in physical properties

Atomic Radii



metallic radius

In metals such as beryllium, the atomic radius is defined as one-half the distance between the centers of two adjacent atoms.

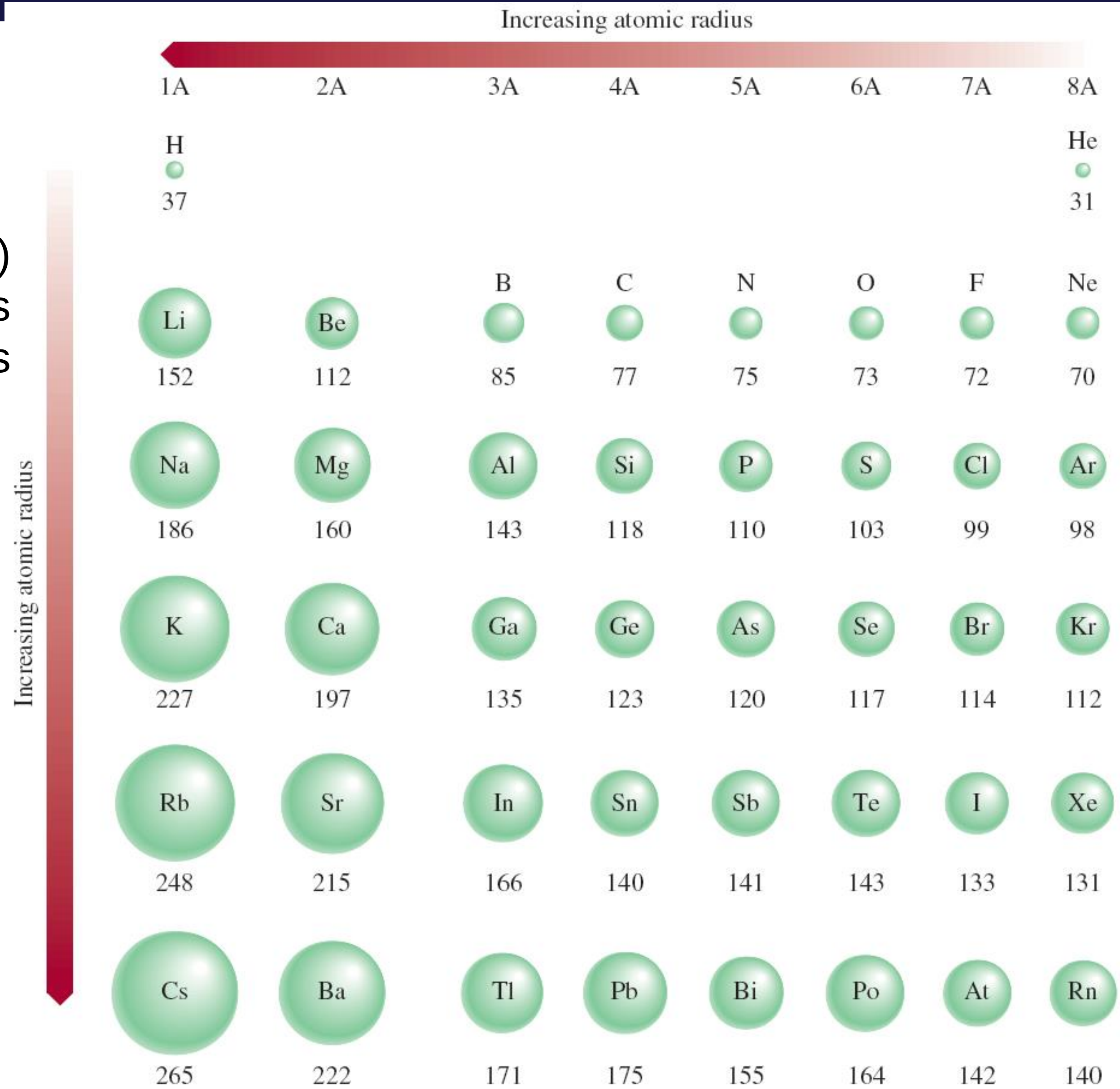


covalent radius

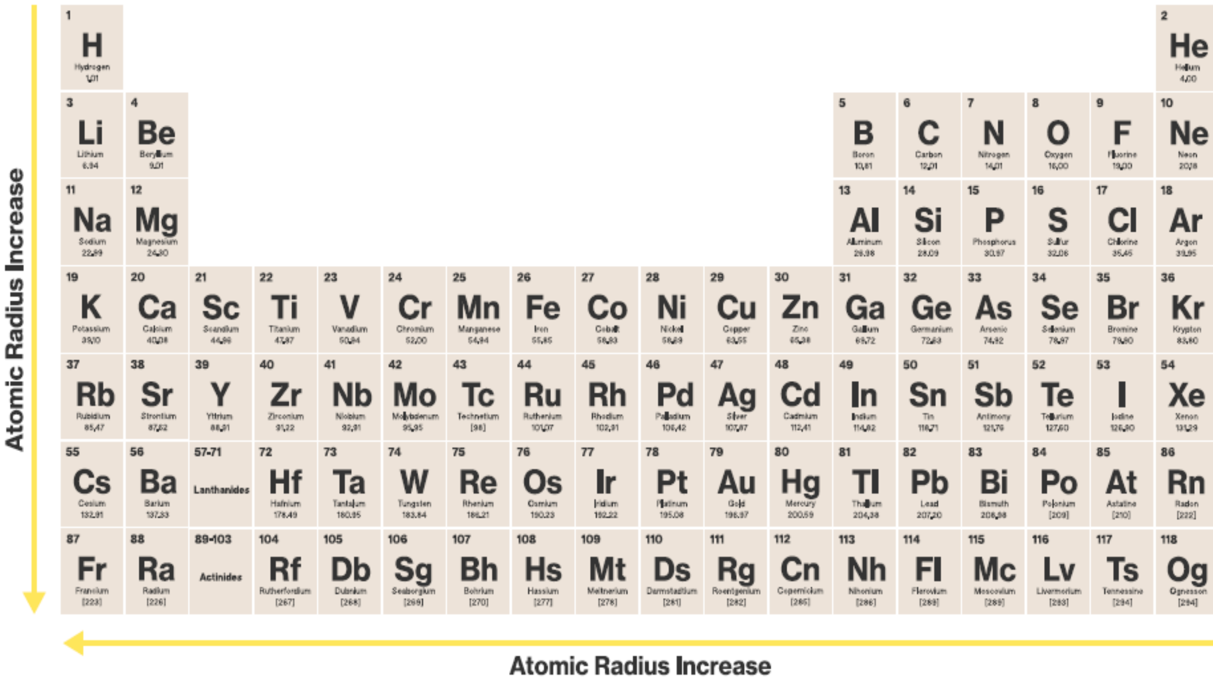
For elements that exist as diatomic molecules, such as iodine, the radius of the atom is defined as one-half the distance between the centers of the atoms in the molecule.

Trends in Atomic Radii

Atomic radii (in picometers) of representative elements according to their positions in the periodic table.

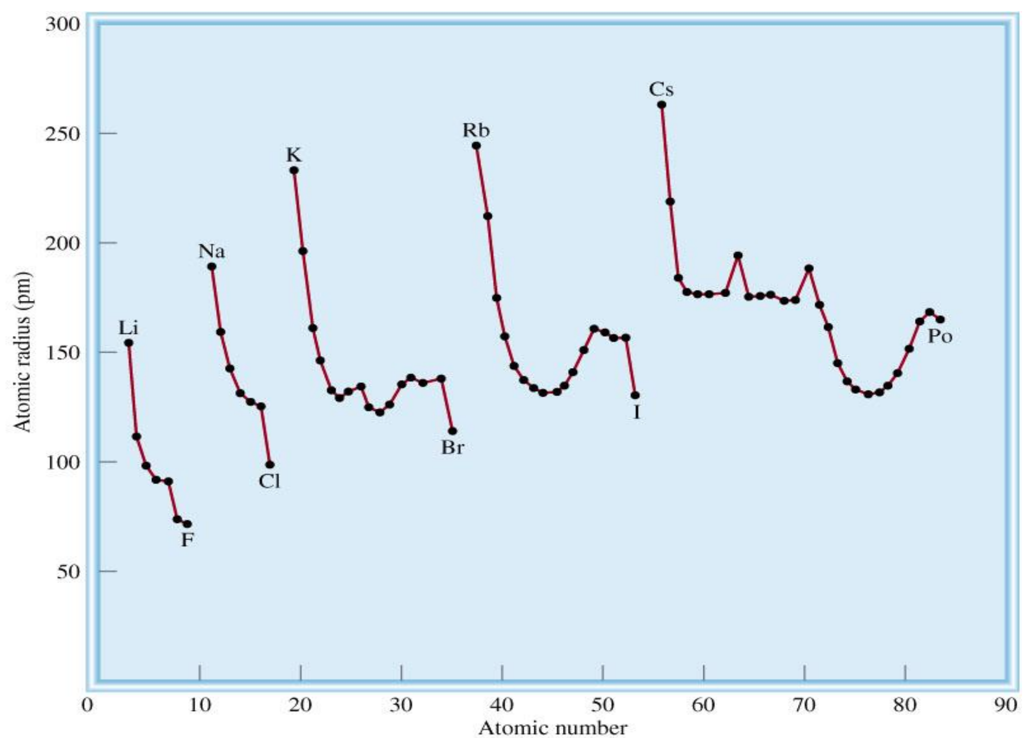


Note that there is no general agreement on the size of atomic radii. We focus only on the trends in atomic radii, not on their precise values.



General Trends in Atomic Radius

Plot of atomic radii (in picometers) of elements against their atomic numbers.



EXAMPLE

Referring to a periodic table, arrange the following atoms in order of increasing atomic radius: P, Si, N.

Solution

N and P are in the same group (Group 5A). Therefore, the radius of N is smaller than that of P (atomic radius increases as we go down a group).
Both Si and P are in the third period, and Si is to the left of P. Therefore, the radius of P is smaller than that of Si (atomic radius decreases as we move from left to right across a period).

Thus, the order of increasing radius is:
N < P < Si

Practice Exercise

Arrange the following atoms in order of decreasing radius: C, Li, Be.

8.4
Ionization energy

Ionization energy

Ionization energy is the minimum energy (kJ/mol) required to remove an electron from a gaseous atom in its ground state.

$$\text{energy} + X(g) \longrightarrow X^{+}(g) + e^{-}$$

$$I_1 + X_{(g)} \longrightarrow X^{+}_{(g)} + e^{-}$$ I_1 first ionization energy

$$I_2 + X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-}$$ I_2 second ionization energy

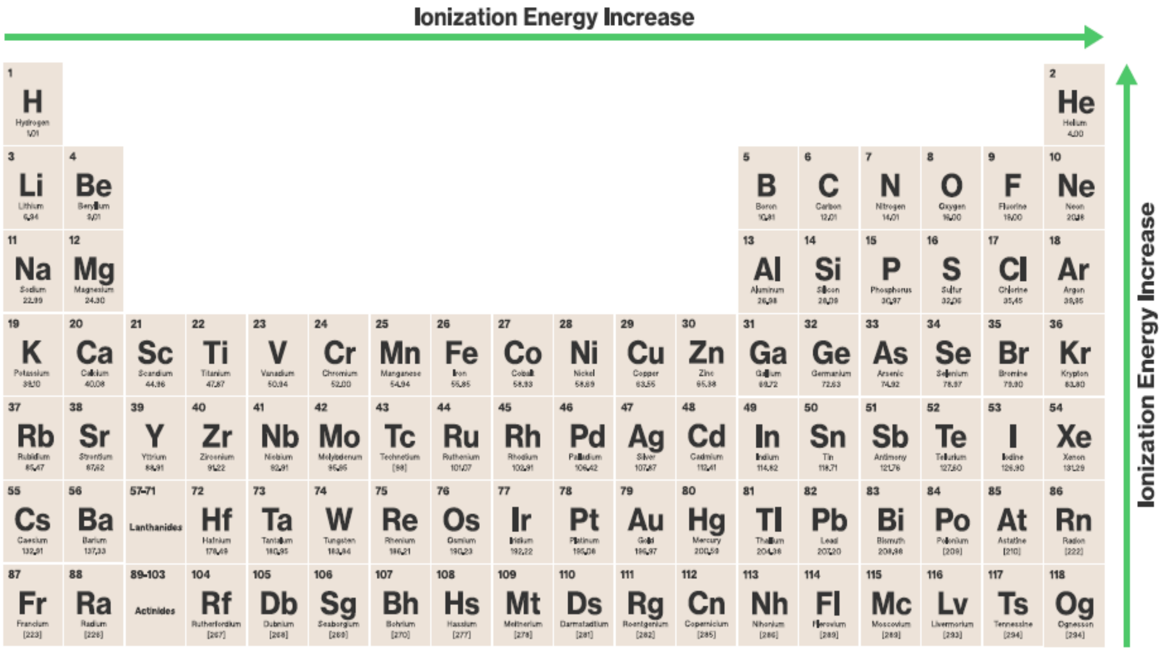
$$I_3 + X^{2+}_{(g)} \longrightarrow X^{3+}_{(g)} + e^{-}$$ I_3 third ionization energy

When an electron is removed from an atom, the repulsion among the remaining electrons decreases. Because the nuclear charge remains constant, more energy is needed to remove another electron from the positively charged ion. Thus, ionization energies always increase in the following order:

$$I_1 < I_2 < I_3 < \ldots.$$

The ionization energies (kJ/mol) of the first 20 elements

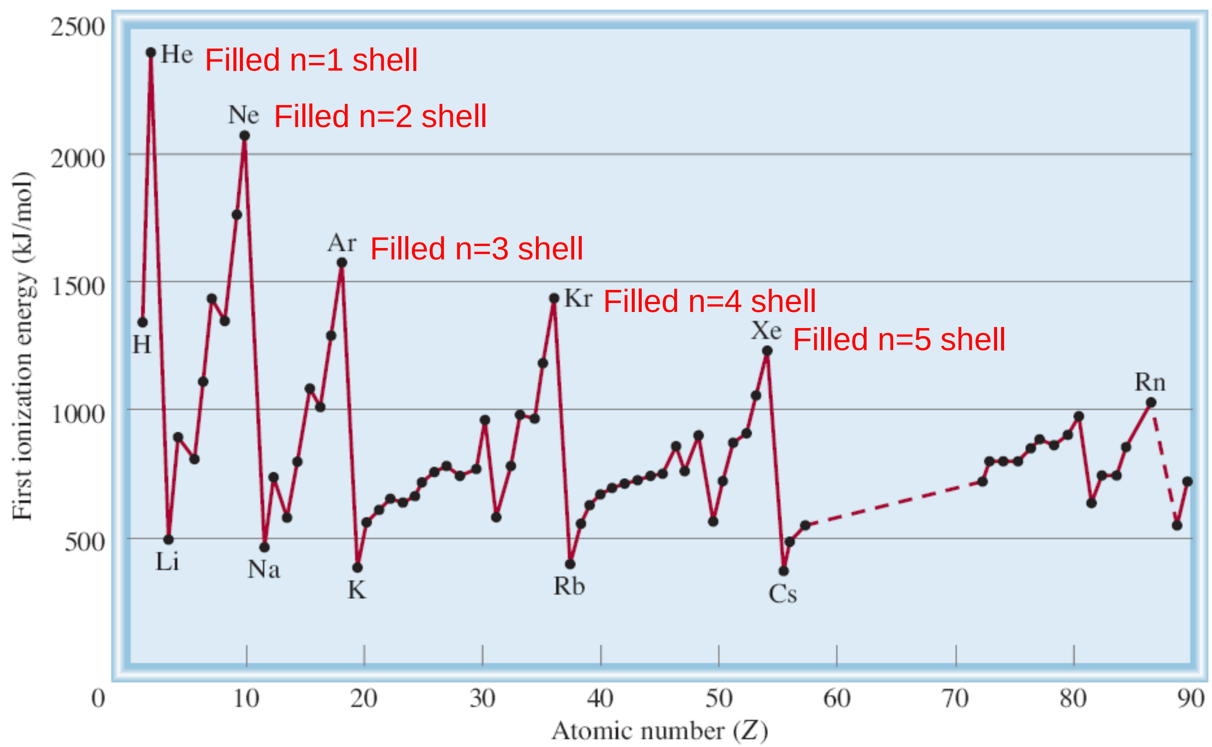
Z	Element	First	Second	Third	Fourth	Fifth	Sixth
1	H	1,312					
2	He	2,373	5,251				
3	Li	520	7,300	11,815			
4	Be	899	1,757	14,850	21,005		
5	B	801	2,430	3,660	25,000	32,820	
6	C	1,086	2,350	4,620	6,220	38,000	47,261
7	N	1,400	2,860	4,580	7,500	9,400	53,000
8	O	1,314	3,390	5,300	7,470	11,000	13,000
9	F	1,680	3,370	6,050	8,400	11,000	15,200
10	Ne	2,080	3,950	6,120	9,370	12,200	15,000
11	Na	495.9	4,560	6,900	9,540	13,400	16,600
12	Mg	738.1	1,450	7,730	10,500	13,600	18,000
13	Al	577.9	1,820	2,750	11,600	14,800	18,400
14	Si	786.3	1,580	3,230	4,360	16,000	20,000
15	P	1,012	1,904	2,910	4,960	6,240	21,000
16	S	999.5	2,250	3,360	4,660	6,990	8,500
17	Cl	1,251	2,297	3,820	5,160	6,540	9,300
18	Ar	1,521	2,666	3,900	5,770	7,240	8,800
19	K	418.7	3,052	4,410	5,900	8,000	9,600
20	Ca	589.5	1,145	4,900	6,500	8,100	11,000



General Trends in First Ionization Energies

The increase in first ionization energy from left to right across a period and from bottom to top in a group for representative elements.

Variation of the first ionization energy with atomic number. Note that the noble gases have high ionization energies, whereas the alkali metals and alkaline earth metals have low ionization energies.



EXAMPLE

(a) Which atom should have a smaller first ionization energy: O or S?
(b) Which atom should have a higher second ionization energy: Li or Be?

Solution

(a) Sulfur should have a smaller first ionization energy.
First ionization energy decreases as we go down a group because the outermost electron is farther away from the nucleus and feels less attraction.

(b) The electron configurations of Li and Be are $1s^2\ 2s^1$ and $1s^2\ 2s^2$, respectively.
Therefore, it should be easier to remove a 2s electron from Be^+ than to remove a 1s electron from Li^+.
Removal of the outermost electron requires less energy if it is shielded by a filled inner shell.

Practice Exercise

(a) Which of the following atoms have a larger 1^{st} ionization energy: N or P?
(b) Which of the following atoms have a smaller 2^{nd} ionization energy: Na or Mg?

8.5
Electron affinity

Electron affinity

Electron affinity is the negative of the energy change that occurs when an electron is accepted by an atom in the gaseous state to form an anion.

$$X_{(g)} + e^- \longrightarrow X^-_{(g)}$$

$$F_{(g)} + e^- \longrightarrow X^-_{(g)} \qquad \Delta H = -328 \text{ kJ/mol} \qquad EA = +328 \text{ kJ/mol}$$

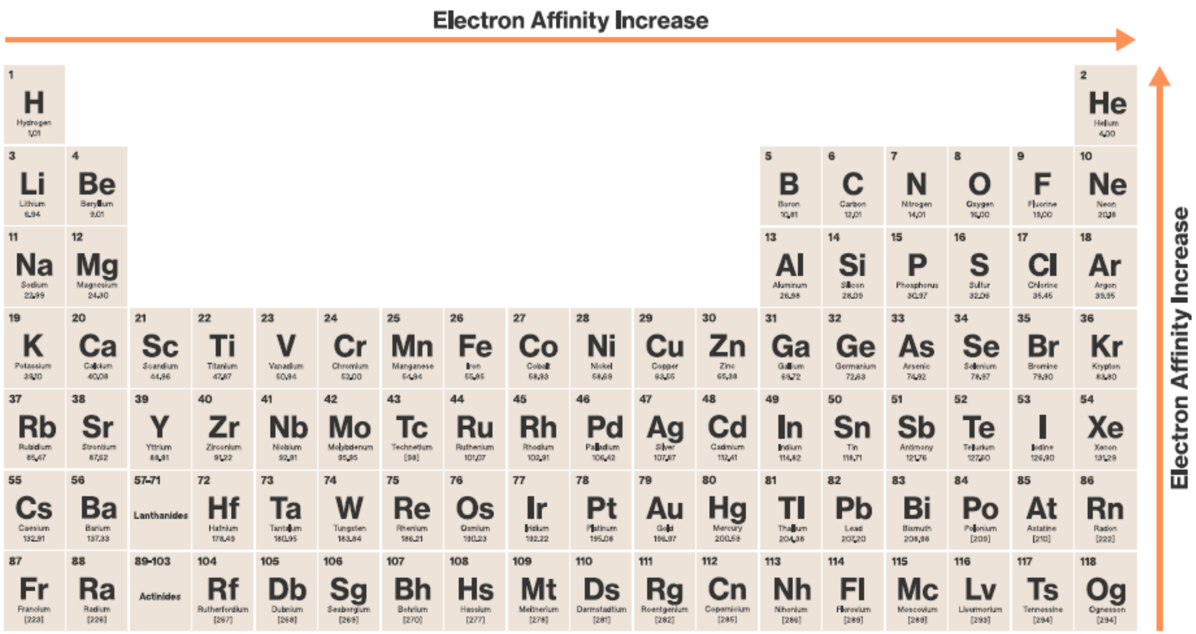
$$O_{(g)} + e^- \longrightarrow O^-_{(g)} \qquad \Delta H = -141 \text{ kJ/mol} \qquad EA = +141 \text{ kJ/mol}$$

The more positive is the electron affinity of an element, the greater is the affinity of an atom of the element to accept an electron.

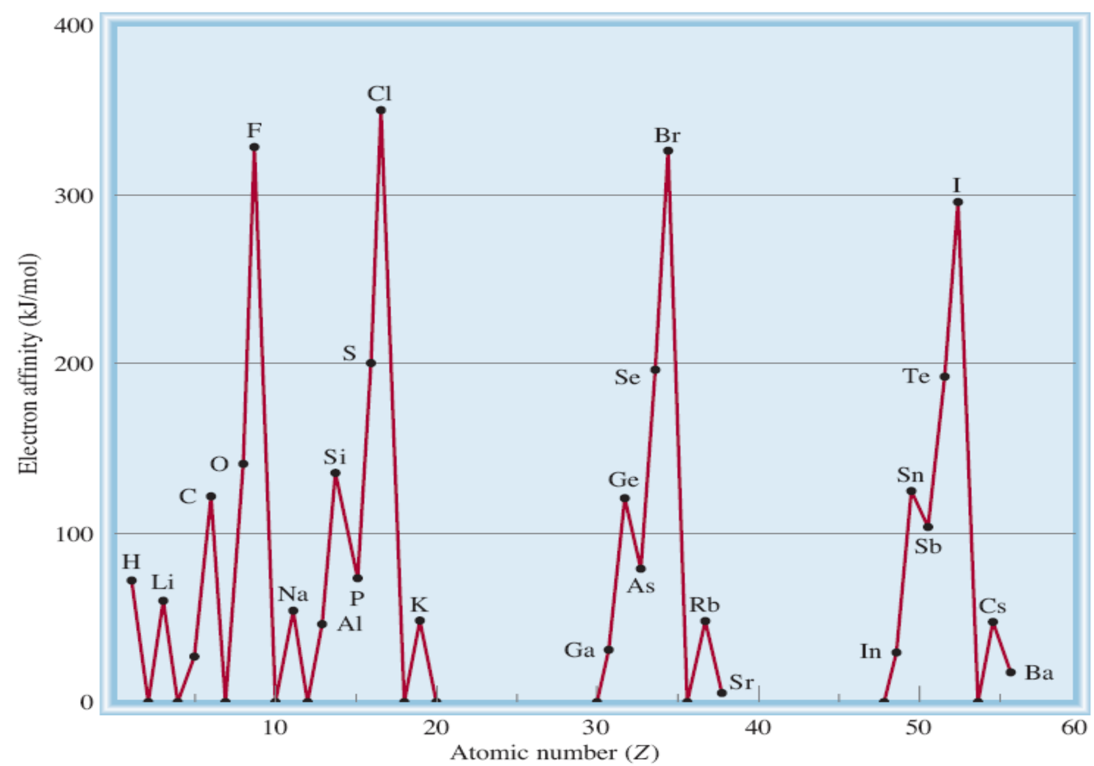
Electron affinities (kJ/mol) of some representative elements and the noble gases

1A	2A	3A	4A	5A	6A	7A	8A
H							He
73							< 0
Li	Be	B	C	N	O	F	Ne
60	≤ 0	27	122	0	141	328	< 0
Na	Mg	Al	Si	P	S	Cl	Ar
53	≤ 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Kr
48	2.4	29	118	77	195	325	< 0
Rb	Sr	In	Sn	Sb	Te	I	Xe
47	4.7	29	121	101	190	295	< 0
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
45	14	30	110	110	?	?	< 0

General Trends in Electron Affinities



Variation of electron affinity with atomic number from hydrogen to barium.



EXAMPLE

Why are the electron affinities of the alkaline earth metals, shown in the Table (above), either negative or small positive values?

Solution

The valence electron configuration of the alkaline earth metals is ns^2, where n is the highest principal quantum number. For the process

$$\underset{ns^2}{M(g)} + e^- \longrightarrow \underset{ns^2np^1}{M^-(g)}$$

where M denotes a member of the Group 2A family, the extra electron must enter the np subshell, which is effectively shielded by the two ns electrons (the ns electrons are more penetrating than the np electrons) and the inner electrons. Consequently, alkaline earth metals have little tendency to pick up an extra electron.

Practice Exercise

Is it likely that Ar will form the anion Ar^- ?


Thank You!