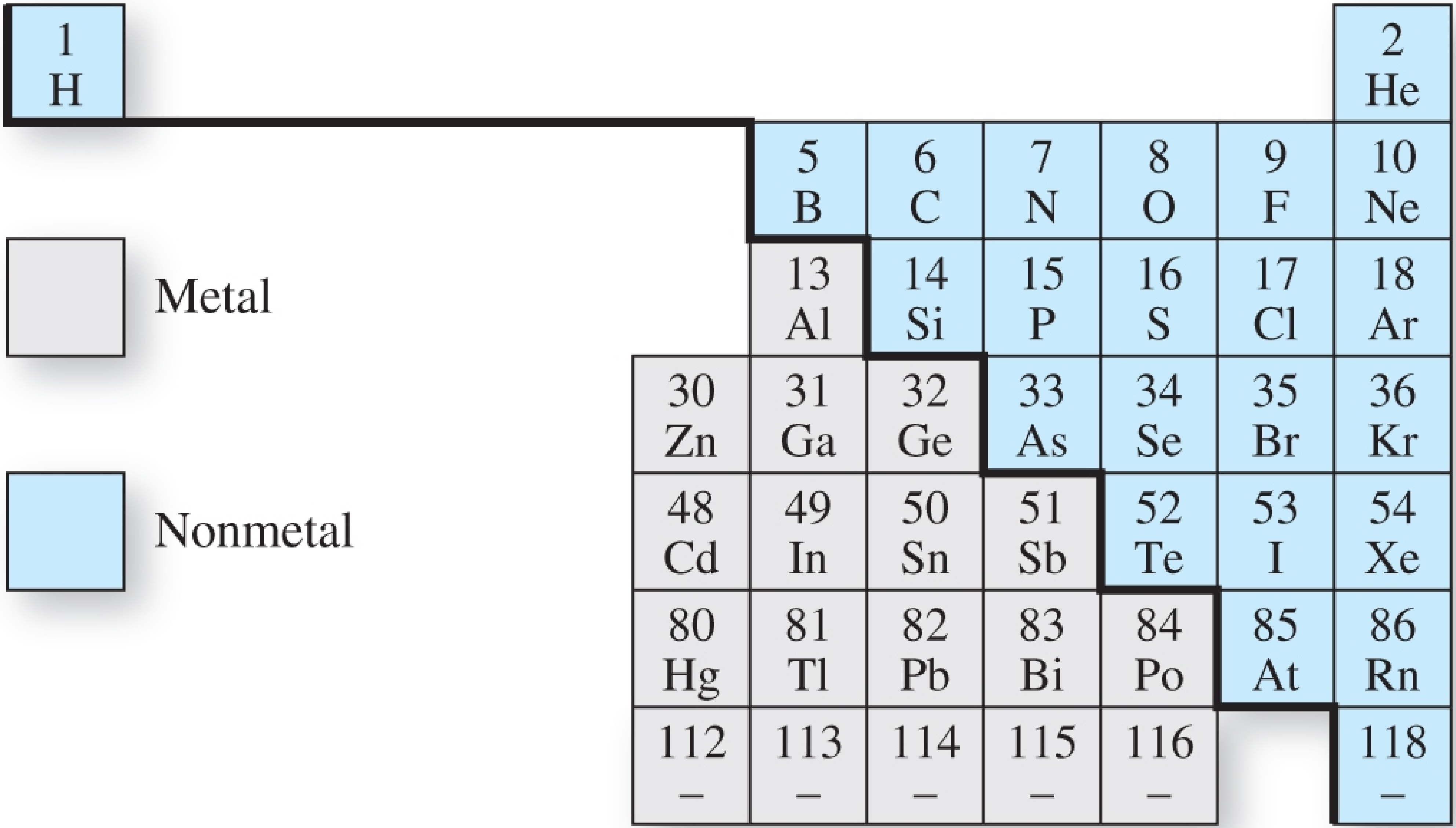
Chapter 6

Periodic Table

- Most elements are metals
  - Metals are shiny, malleable, ductile, and good conductors of heat and electricity
  - Most metals are solid at room temperature

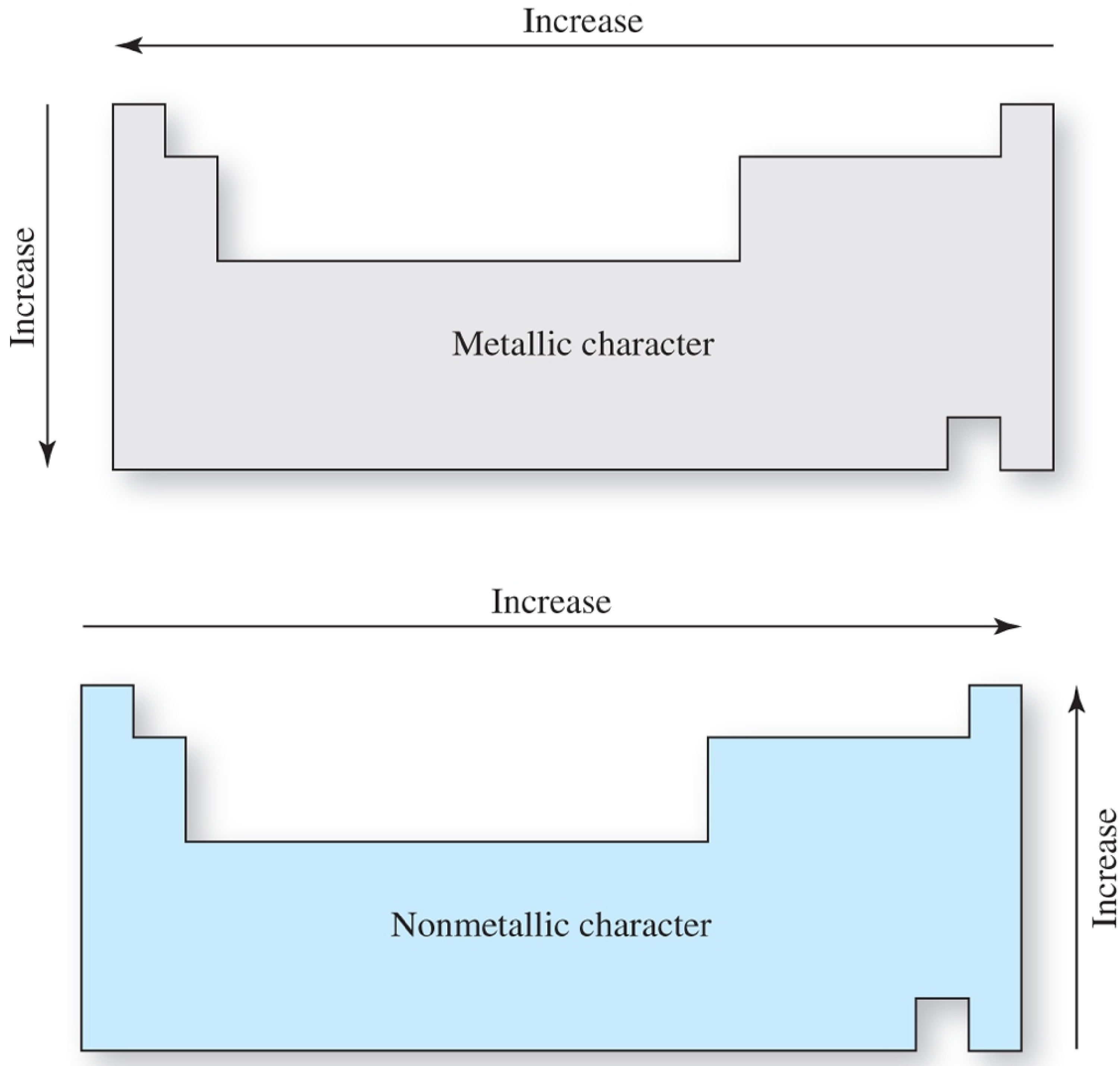
- Non-metals in upper right corner, plus H
  - Non-metals are poor conductors of heat and electricity
  - They are gaseous or brittle solids at room temperature

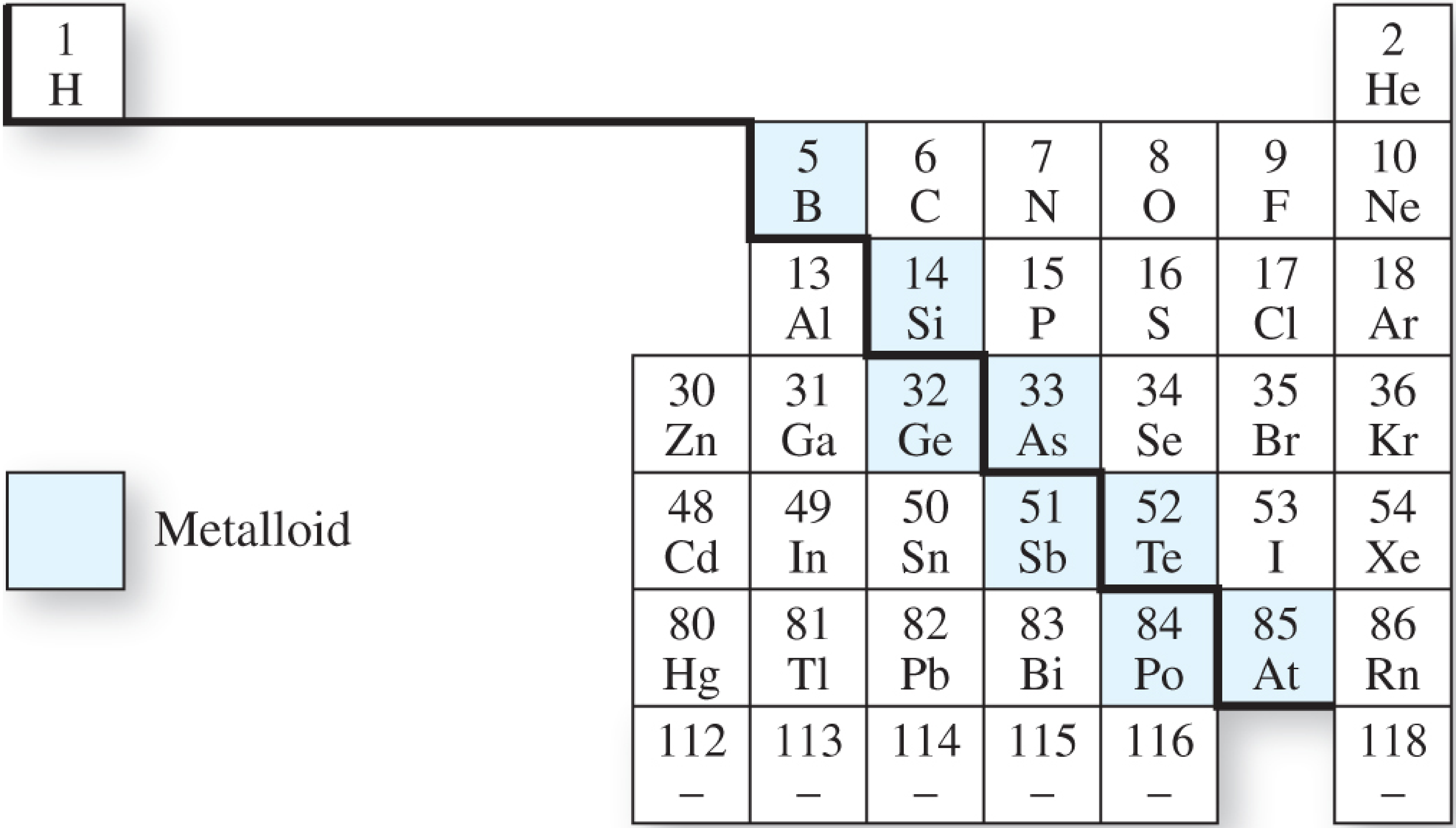
- Metalloids – diagonal between metals and non-metals
  - Have metallic and non-metallic properties


1 H
2 He
Metal
Nonmetal
5 B
6 C
7 N
8 O
9 F
10 Ne
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
112 –
113 –
114 –
115 –
116 –
118 –

Periodic Table

- Elements to the left on a row and down on a column tend to have more metallic character.
- Some elements can not be clearly defined as a metal or a non-metal
- These are Metalloids, and lie on the diagonal between metals and non-metals
  - They have some metallic and some non-metallic properties (such as semiconductors)


Increase
Increase
Metallic character
Increase
Nonmetallic character
Increase


1 H
2 He
5 B
6 C
7 N
8 O
9 F
10 Ne
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
112 –
113 –
114 –
115 –
116 –
118 –
Metalloid

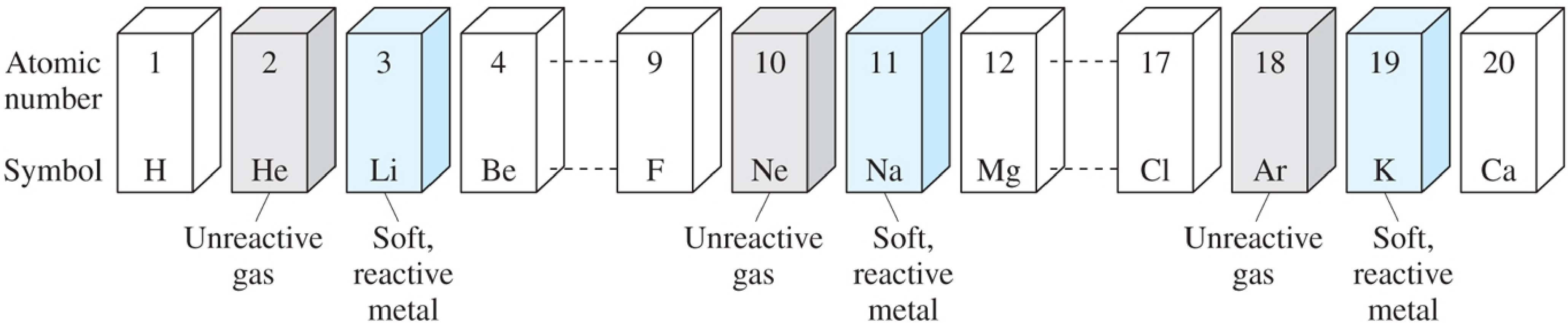
Periodic Table

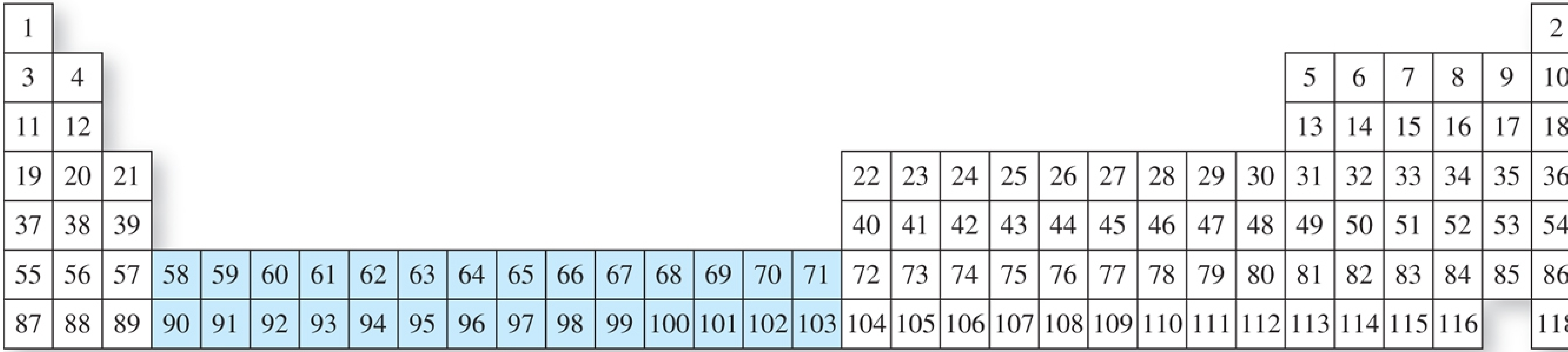
- Elements arranged by atomic #, starting with hydrogen in the upper left corner.

- Elements arranged by properties
  - Elements in a column tend to have similar physical and chemical properties

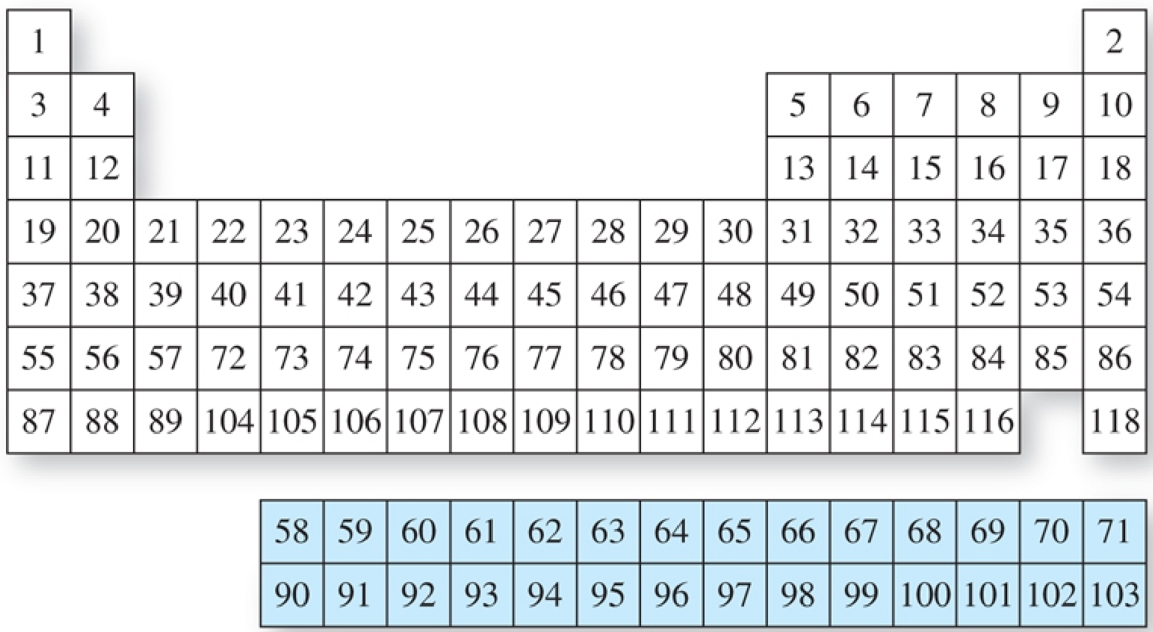
- Rows are called Periods
  - The top row is period 1, next is period 2, ….

- Columns are called Groups
  - 1-18 (or 1A-8A, 1B-8B)

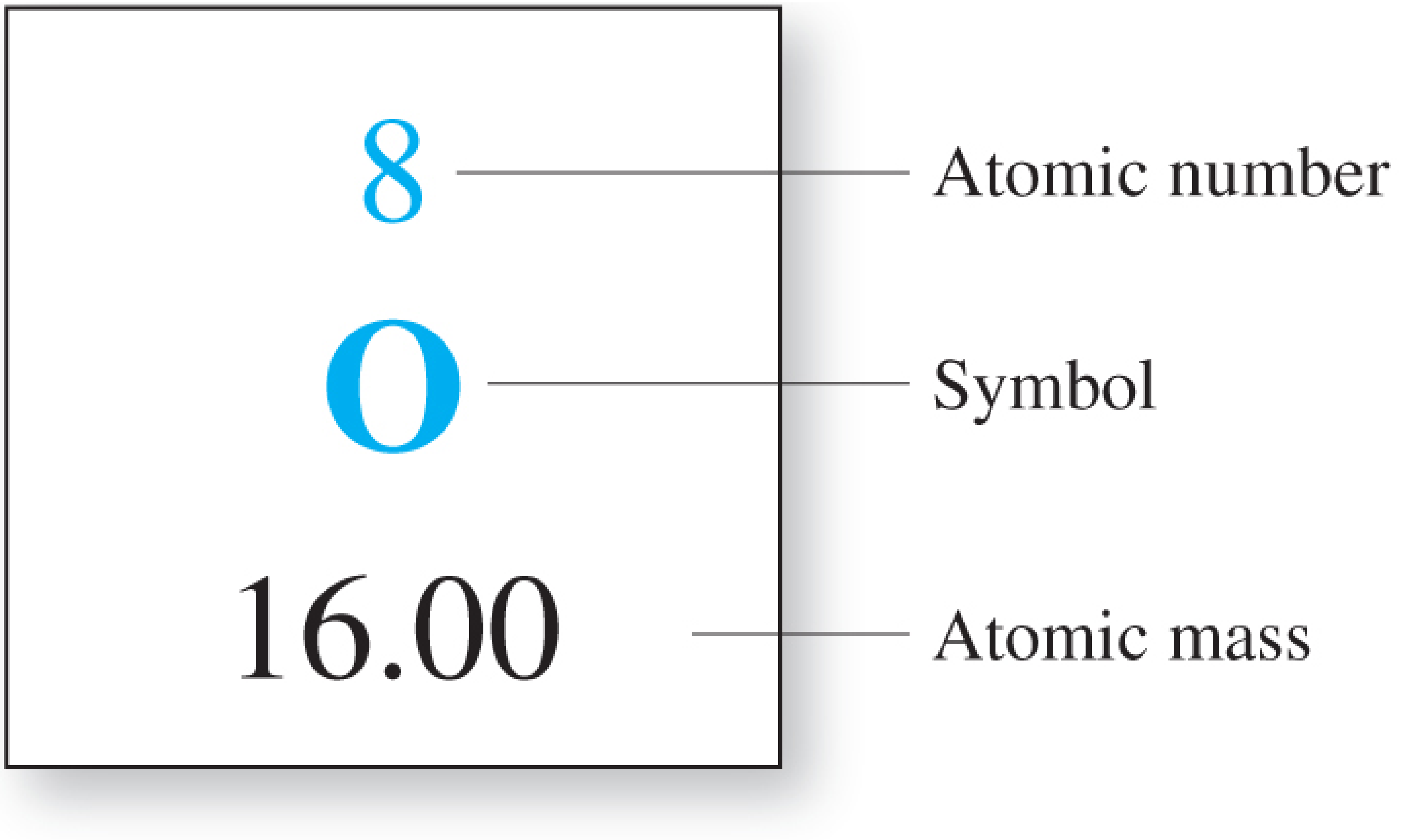

Atomic number
Symbol
1
H
2
He
3
Li
4
Be
9
F
10
Ne
11
Na
12
Mg
17
Cl
18
Ar
19
K
20
Ca
Unreactive gas
Soft, reactive metal
Unreactive gas
Soft, reactive metal
Unreactive gas
Soft, reactive metal


1 2
3 4 5 6 7 8 9 10
11 12 13 14 15 16 17 18
19 20 21 22 23 24 25 26 27 28 29 30 31 32 33 34 35 36
37 38 39 40 41 42 43 44 45 46 47 48 49 50 51 52 53 54
55 56 57 58 59 60 61 62 63 64 65 66 67 68 69 70 71 72 73 74 75 76 77 78 79 80 81 82 83 84 85 86
87 88 89 90 91 92 93 94 95 96 97 98 99 100 101 102 103 104 105 106 107 108 109 110 111 112 113 114 115 116 118

(a)


1 2
3 4 5 6 7 8 9 10
11 12 13 14 15 16 17 18
19 20 21 22 23 24 25 26 27 28 29 30 31 32 33 34 35 36
37 38 39 40 41 42 43 44 45 46 47 48 49 50 51 52 53 54
55 56 57 72 73 74 75 76 77 78 79 80 81 82 83 84 85 86
87 88 89 104 105 106 107 108 109 110 111 112 113 114 115 116 118
58 59 60 61 62 63 64 65 66 67 68 69 70 71
90 91 92 93 94 95 96 97 98 99 100 101 102 103

(b)


8
O
16.00
Atomic number
Symbol
Atomic mass

1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group ←	9 Group VIIIB	10 Group →	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
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3 **Li** 6.94	4 **Be** 9.01											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
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87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)	113 **–** (284)	114 **–** (289)	115 **–** (288)	116 **–** (293)		118 **–** (294)

Metals ← → Non-metals

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
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)

- Group 1: Alkali metals (except hydrogen)
  - Most reactive metals

- Group 2: Alkaline earth metals
  - Very reactive metals

- Groups 3-12: Transition metals (elements)

- Group 17: Halogens
  - Most reactive non-metals

- Group 18: Noble gases
  - Almost never react

Electrons

- Fundamental subatomic particle
- Found outside of nucleus in orbitals or 'electron clouds'
- Moves rapidly and defines the radius of the atom
- Occupies only discrete, quantized energy states in the atom

Chemical bonding depends on the configuration of electrons in the atoms.

From quantum mechanics, there are four quantum numbers that describe each electron in an atom:

- One tells the overall energy level (shell)
- One describes sub levels in each main level
- One indicates the number of orbitals in each sub level
- And the last tells the number of electrons in each orbital

Energy Shells (or Levels)

- 3D regions around the nucleus where electrons may be found
- Shells are distinguished by quantum numbers: $n = 1, 2, 3, 4, \ldots$
- All electrons in a particular shell have similar energies (shells may overlap slightly in energy)

- Each Shell has a maximum # of electrons it can hold
- Lower # shells:
  - Are physically smaller
  - Have lower energy electrons
  - Are filled up before higher energy shells

Each shell has at least one sub-shell (or sub level)

Sub-shells are also labeled by quantum numbers:

l = 0, 1, 2, 3, 4…. (n-1)

but letters are usually used to avoid confusion:

l = s, p, d, f, g, ….

1st shell – only one sub-shell – s

2nd shell – only two sub-shells – s, p

3rd shell – only three sub-shell – s, p, d

4th shell – only four sub-shell – s, p, d, f

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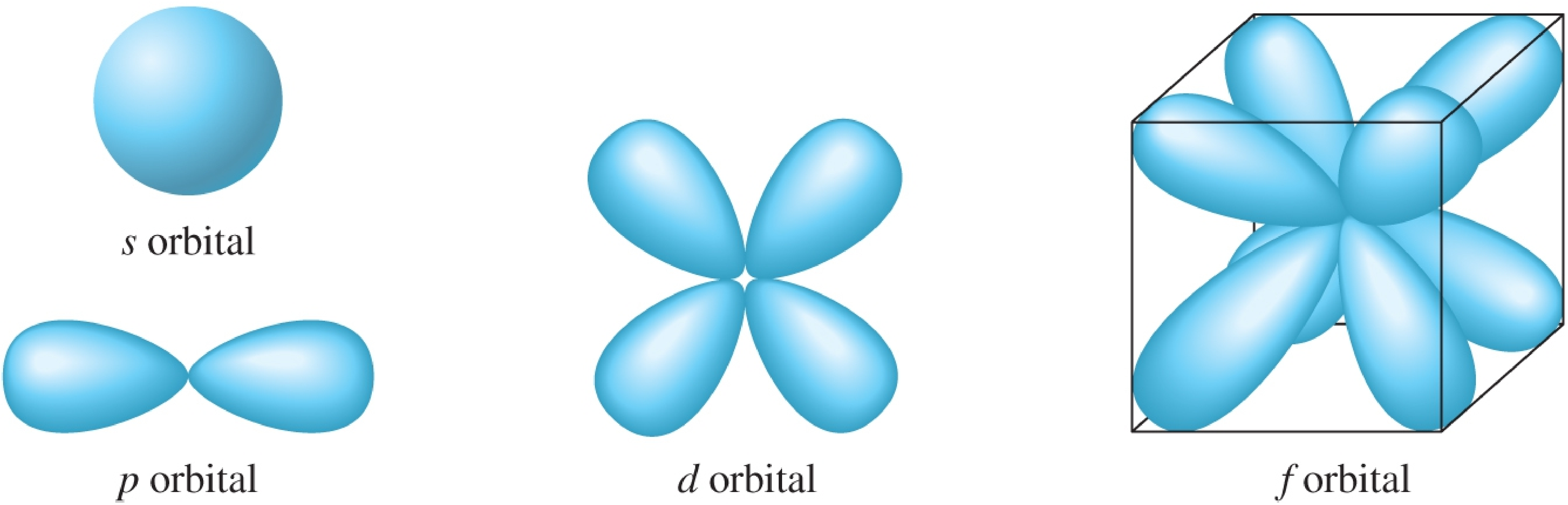
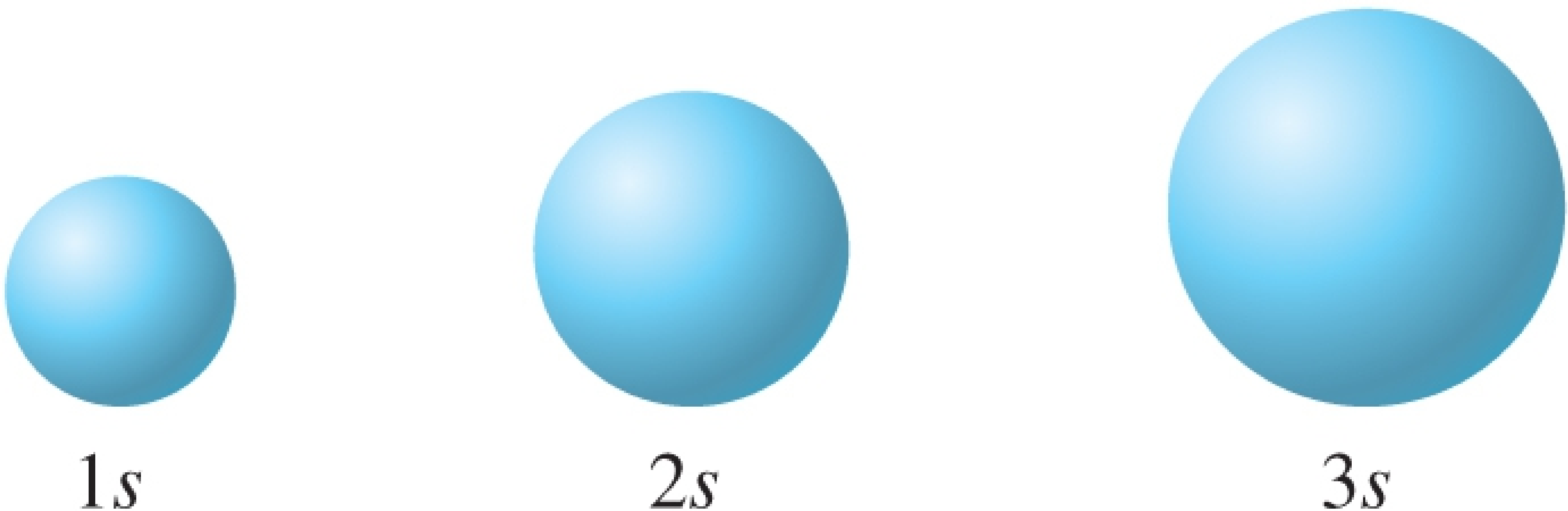

s orbital
p orbital
d orbital
f orbital


1s
2s
3s

Each sub-shell has a fixed number of orbitals (from 3rd quantum number):

s – 1 orbital (spherically shaped)

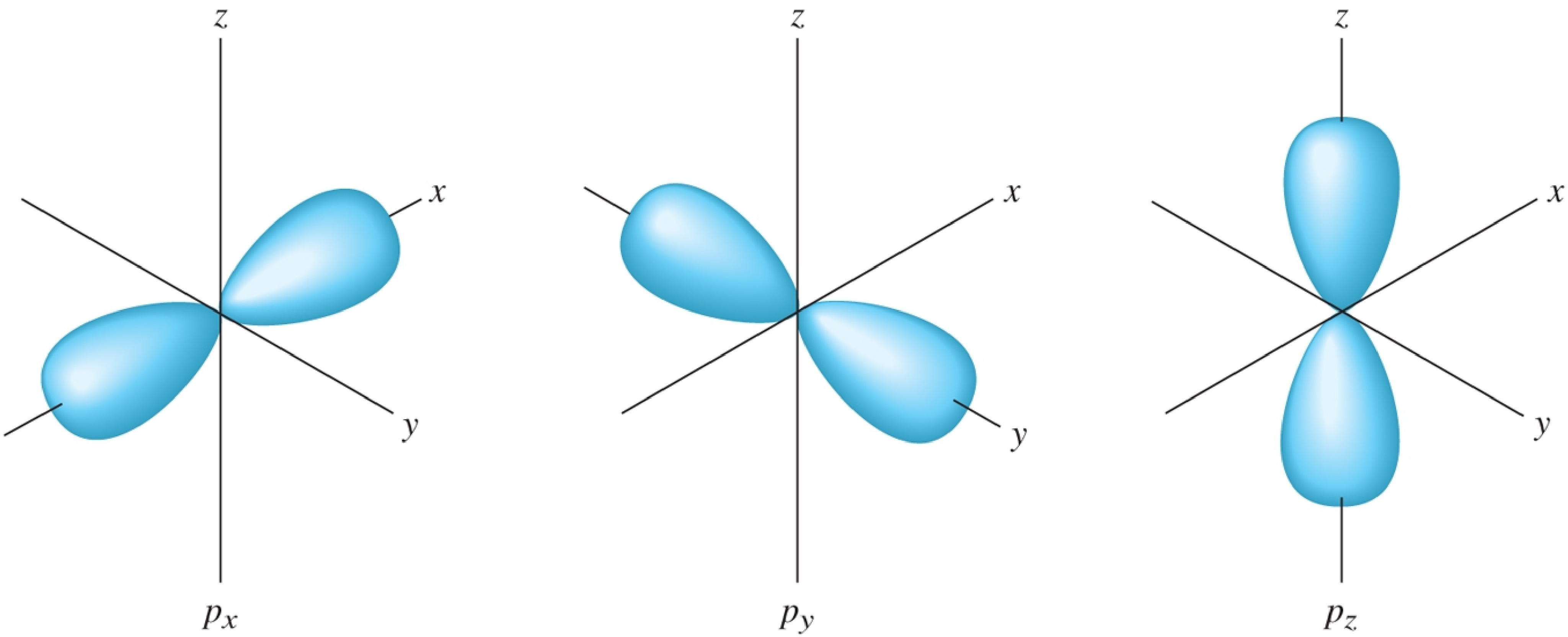
p – 3 orbitals (lobed shaped, along each axis)

d – 5 orbitals

f – 7 orbitals

g – 9 orbitals

...


z
x
y
z
x
y
z
x
y
p_x
p_y
p_z

Each orbital can hold a maximum of two electrons (from 4th quantum number)

- Electrons in an orbital must have opposite spins
  - "Spin up" or "spin down" (+1/2 or -1/2)

- This determines the maximum # of electrons a sub-shell and shell can hold:
  - s – 1 orbital – 2 electrons
  - p – 3 orbitals – 6 electrons
  - d – 5 orbitals – 10 electrons
  - f – 7 orbitals – 14 electrons

shell	sub-shells	orbitals per subshell	electrons per sub-shell	maximum possible electrons
1	s	1	2	2
2	s p	1 3	2 6	8
3	s p d	1 3 5	2 6 10	18
4	s p d f	1 3 5 7	2 6 10 14	32

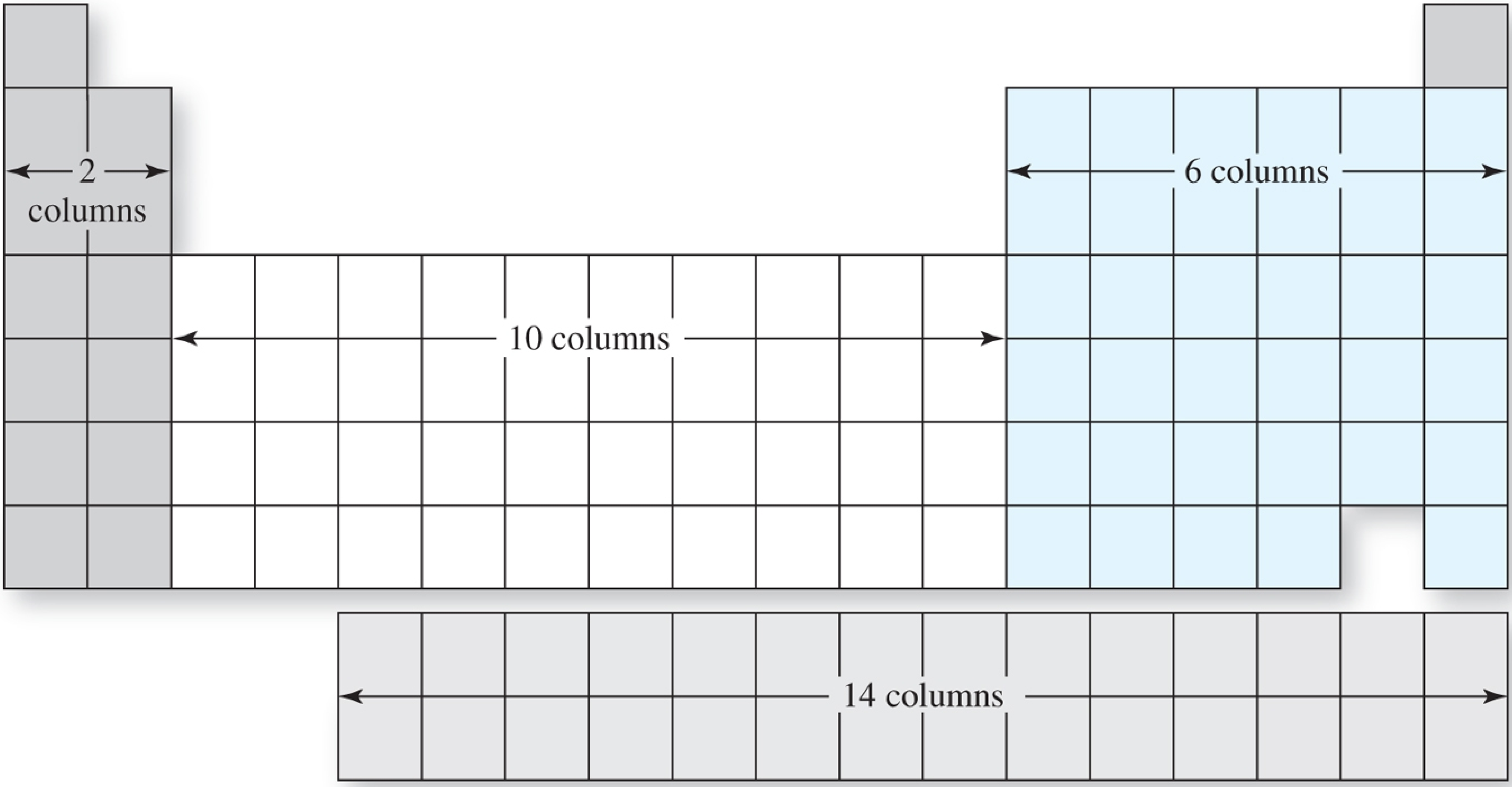
- Electrons want to occupy the lowest available energy <u>state</u>.,

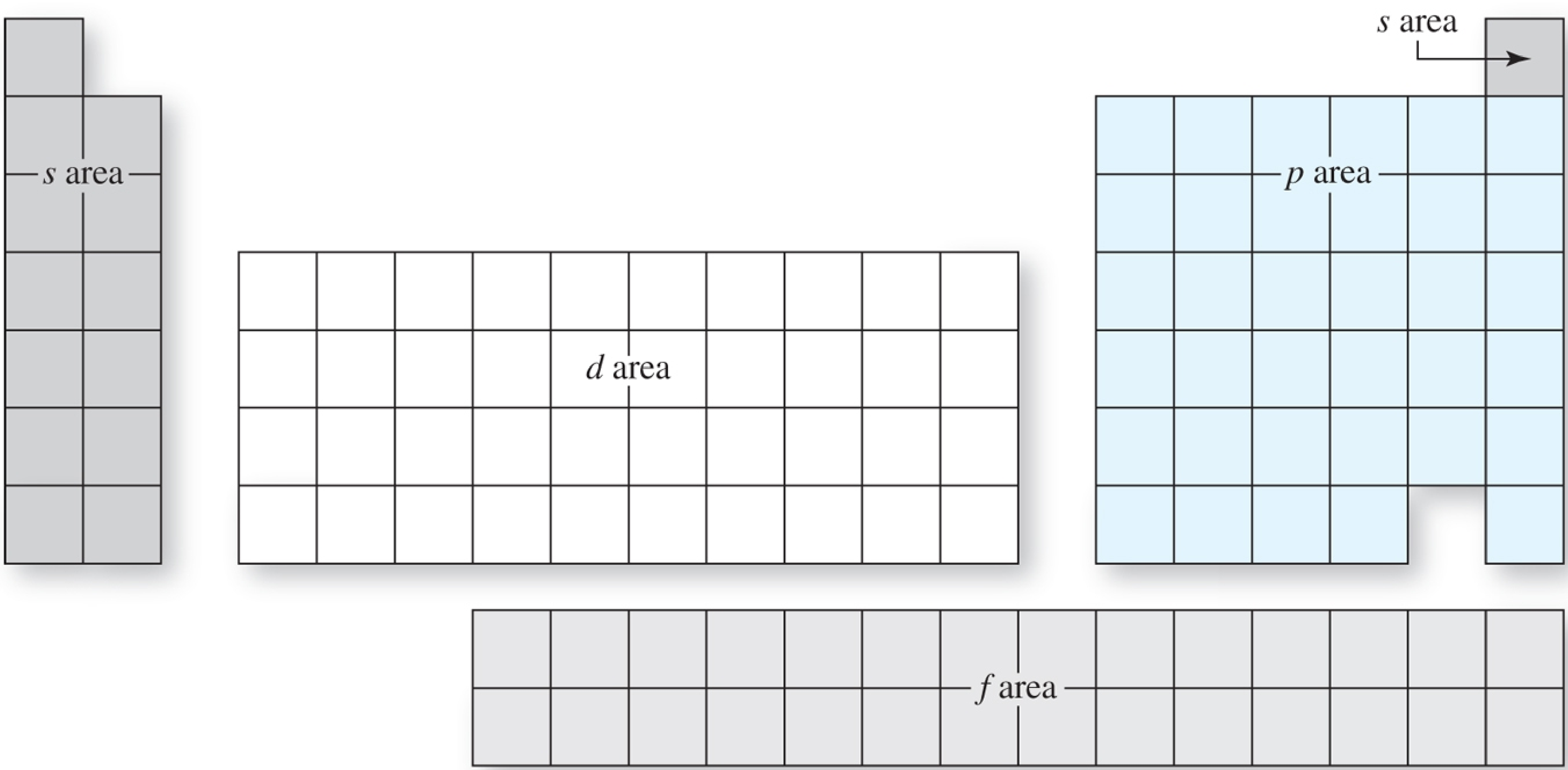
- Some shells overlap, so the lowest <u>state</u> is not always in the lowest <u>shell (level)</u>.

- We can use the periodic table to determine which sub-shells should be filled up first.

- The periodic table also shows how many electrons each sub-shell can hold.

- The periodic table can be split into 4 blocks:
  s, p, d, and f blocks


2
columns
6 columns
10 columns
14 columns



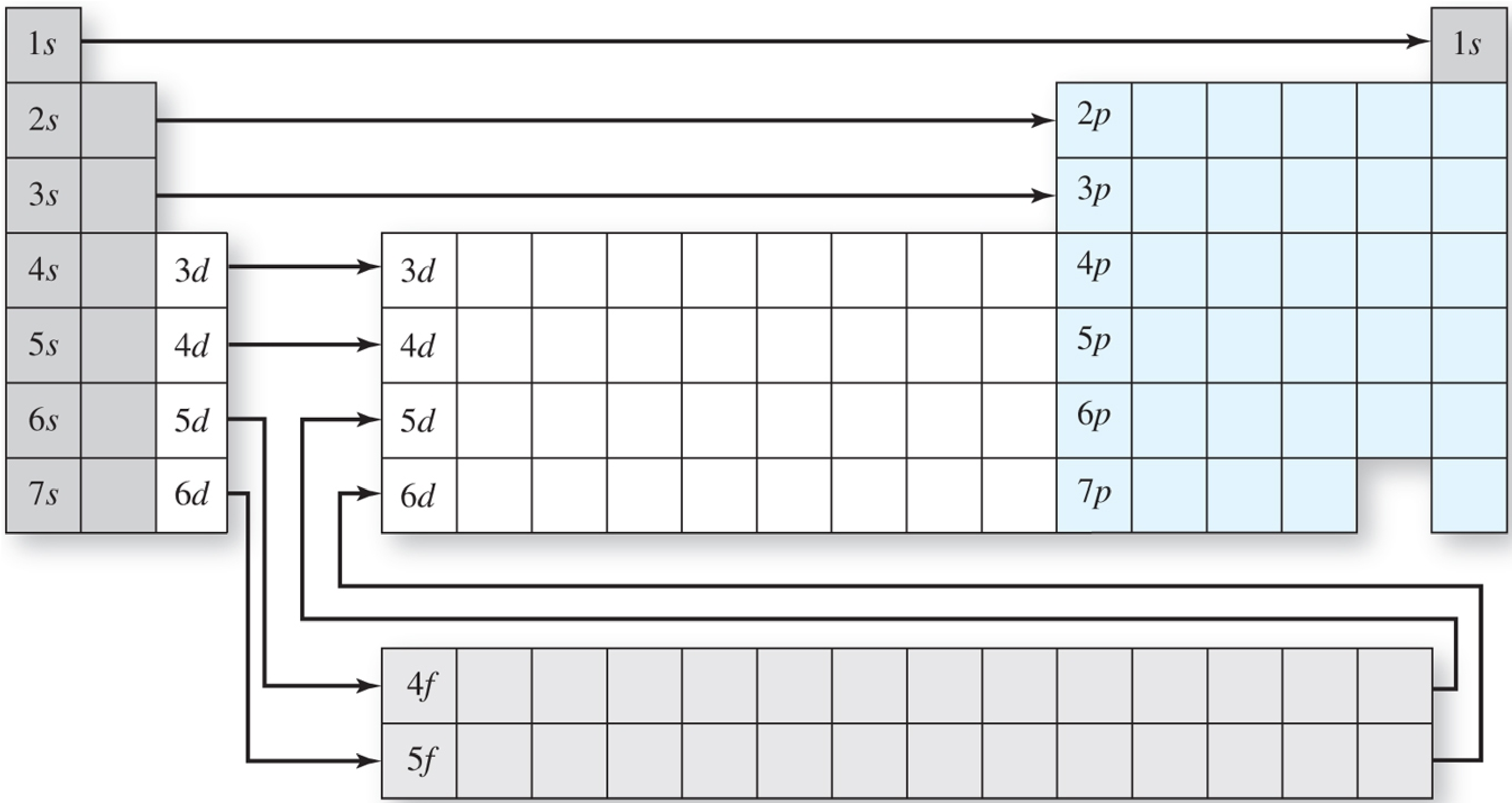
s area
s area
d area
p area
f area

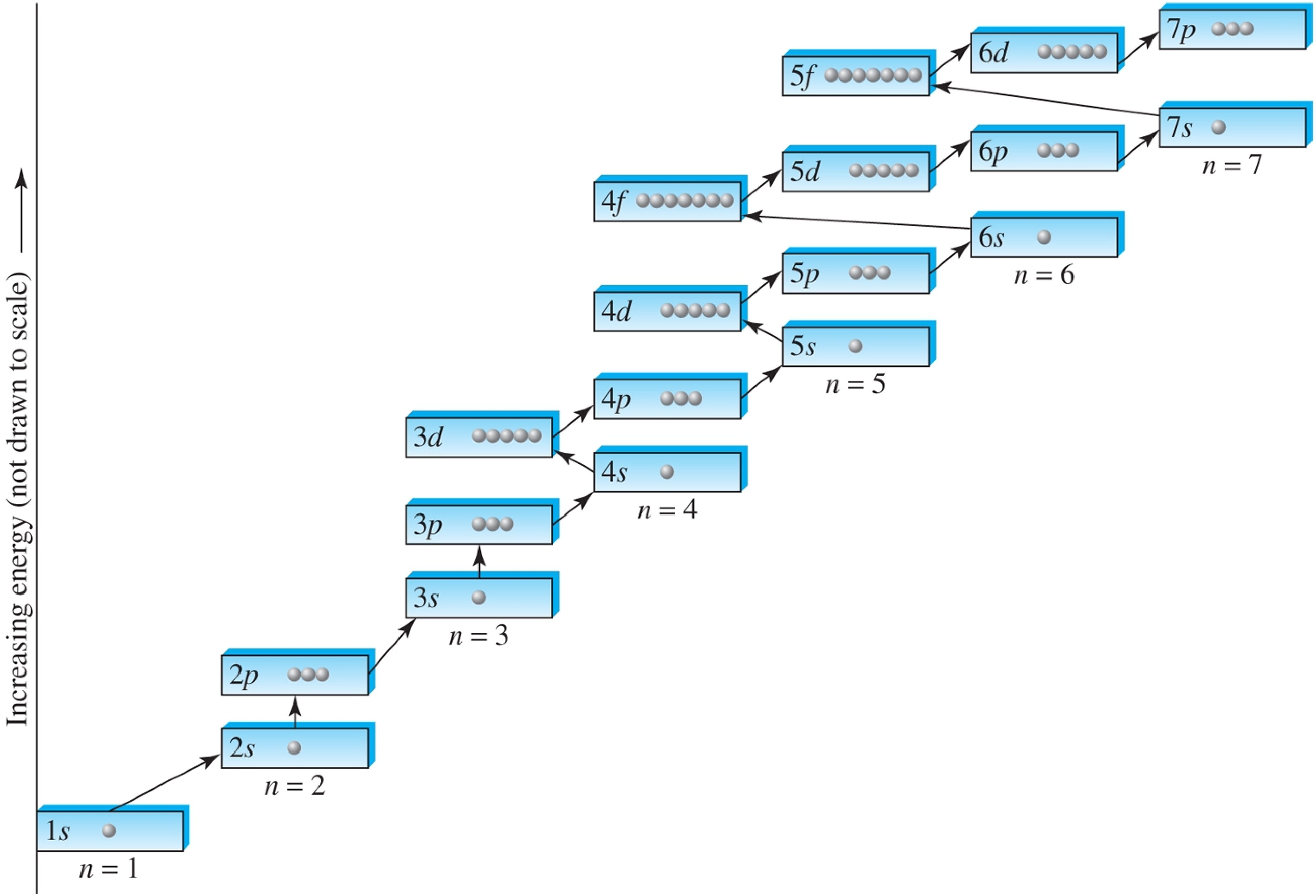
- Each box can hold one electron, so fill up the boxes, starting in the upper left, until all electrons have been used
- The sub-shells are labeled with the energy level (shell) first, then the sub-shell: e.g. 1s, 2p, 4d....
- Superscripts after each sub-shell tell how many electrons are in that sub-shell: e.g. $1s^2$, $3p^4$,....

- The shell of each s and p sub-shell is the same as the period on the PT, but not so for d and f:

- The simplest way to determine energy shell is:

  - The first s is 1s
  - The first p is 2p
  - The first d is 3d
  - The first f is 4f


1s
1s
2s
2p
3s
3p
4s
3d
3d
4p
5s
4d
4d
5p
6s
5d
5d
6p
7s
6d
6d
7p
4f
5f


Increasing energy (not drawn to scale)
1s
n = 1
2s
2p
n = 2
3s
3p
3d
n = 3
4s
4p
4d
4f
n = 4
5s
5p
5d
5f
n = 5
6s
6p
6d
n = 6
7s
7p
n = 7

- Electronic Configuration
  - Use shells, sub-shells, and superscripts to show where all electrons are positioned
    - Ground state – electrons occupy lowest available states
    - Excited state – electrons are at higher energy states due to added energy
    - There are exceptions to expected config.

1 Group IA	2 Group IIA	3 Group IIIB	4 Group IVB	5 Group VB	6 Group VIB	7 Group VIIB	8 Group ←	9 Group VIIIB	10 Group →	11 Group IB	12 Group IIB	13 Group IIIA	14 Group IVA	15 Group VA	16 Group VIA	17 Group VIIA	18 Group VIIIA
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87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)	113 **–** (284)	114 **–** (289)	115 **–** (288)	116 **–** (293)		118 **–** (294)

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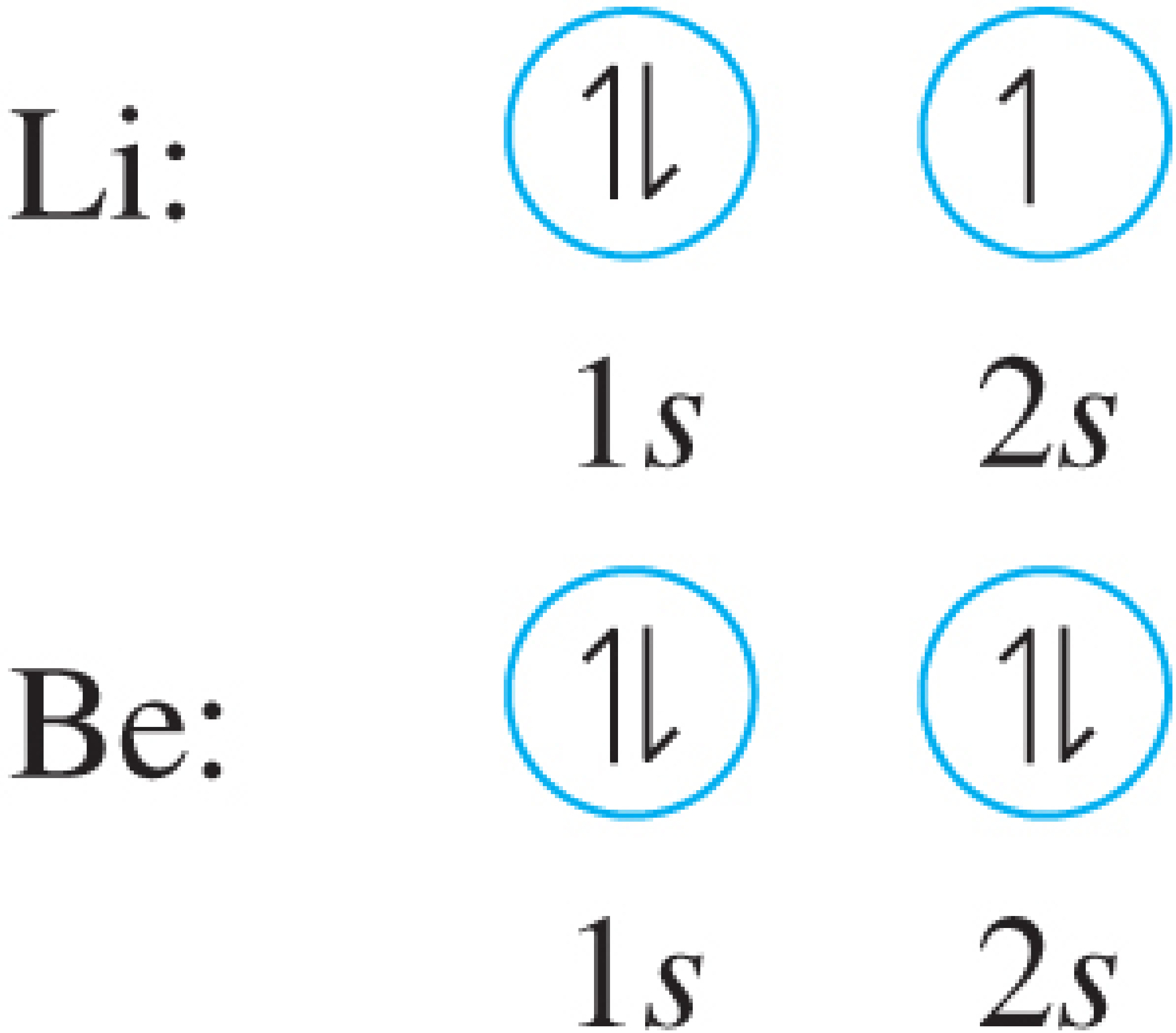
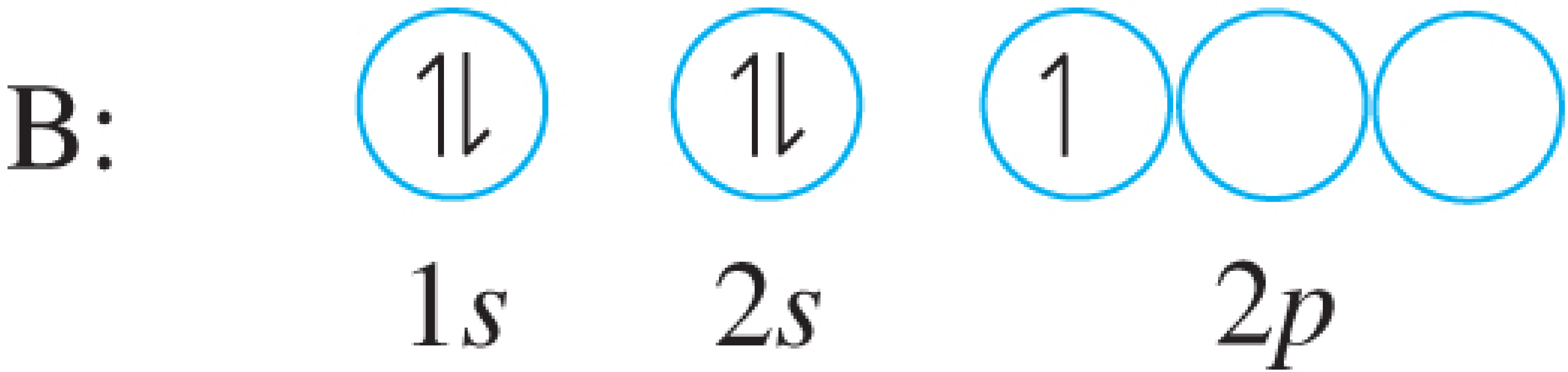
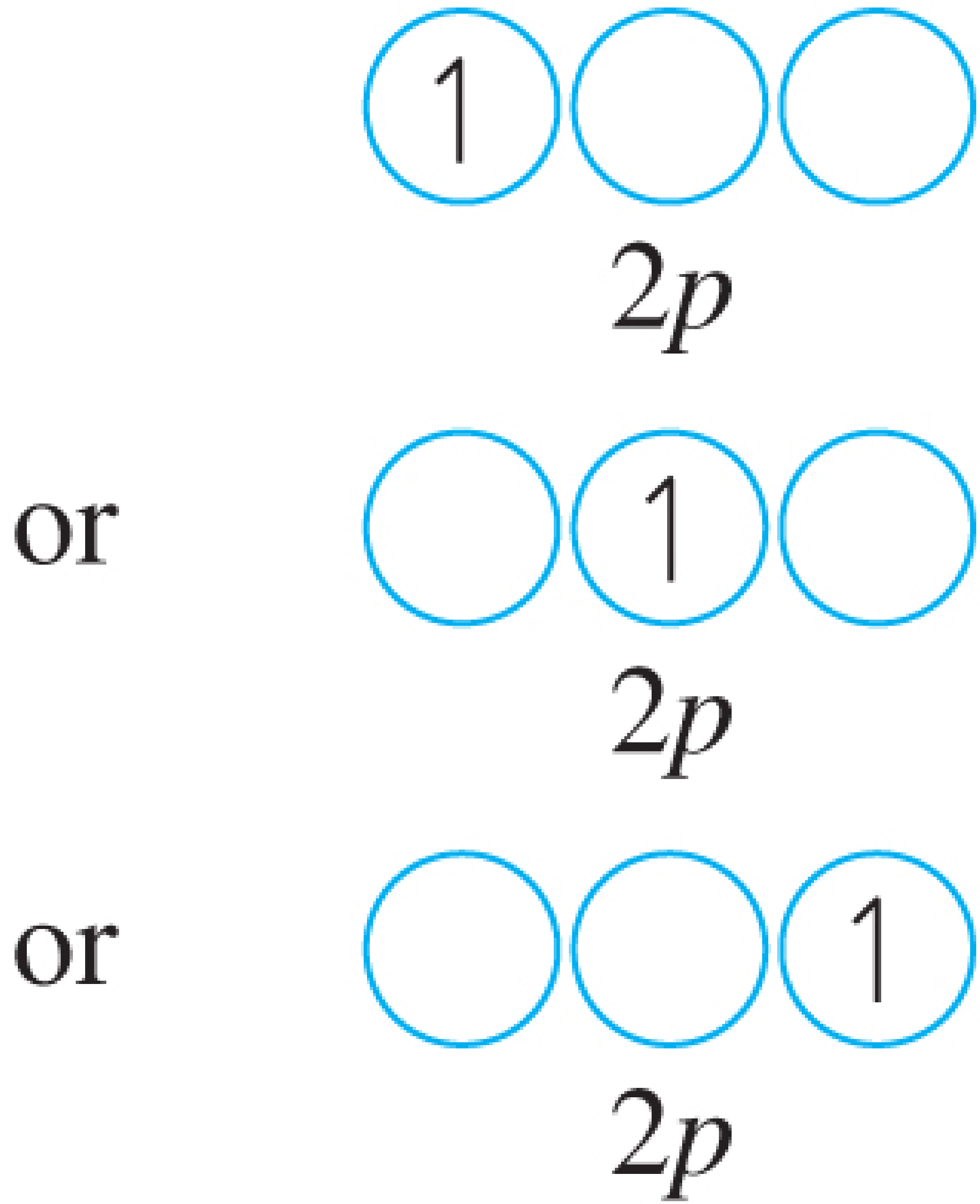
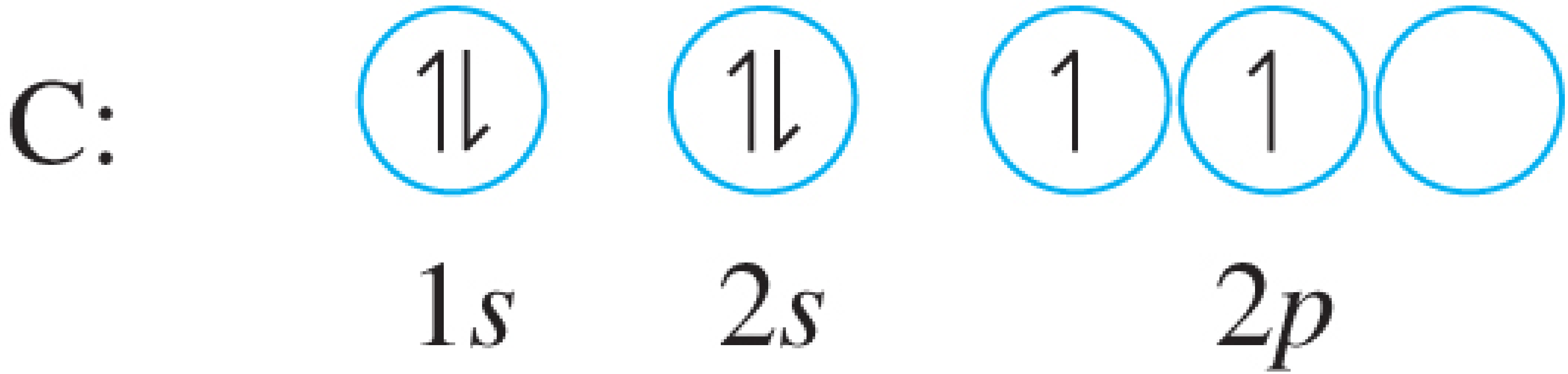
- Noble Gas configuration
  - Shorthand – only shows outer electrons
  - Brackets around Noble Gas represents the electronic configuration of that gas
  - e.g. [Ar] = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

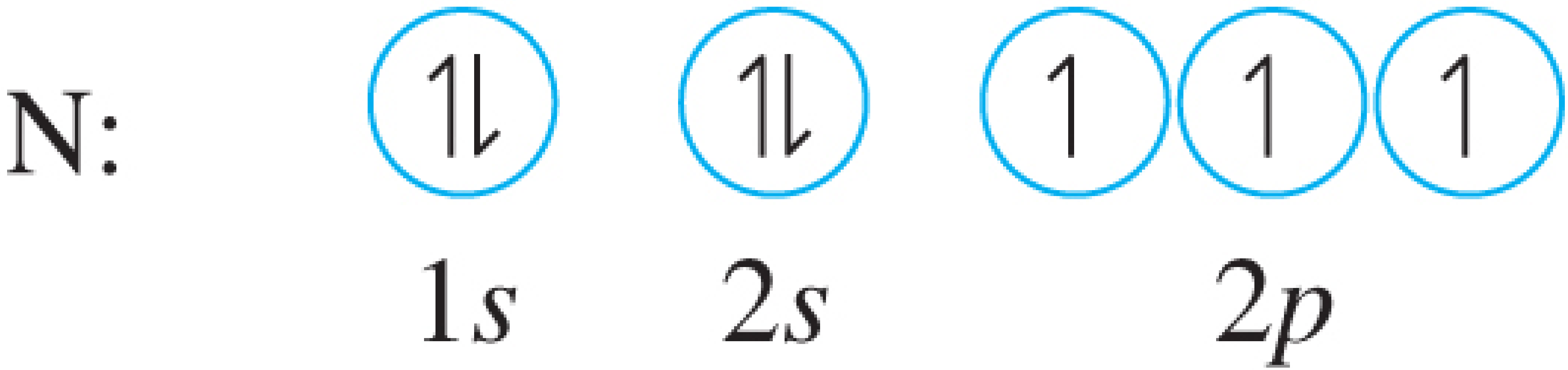
- Orbital notation
  - Use boxes, lines, or circles to show individual orbitals
  - Usually only orbitals of outer sub-shells are shown
  - In a given sub-shell:
    - Place one electron per orbital before pairing
    - All single electrons must have same spin
    - Paired electrons must have opposite spin

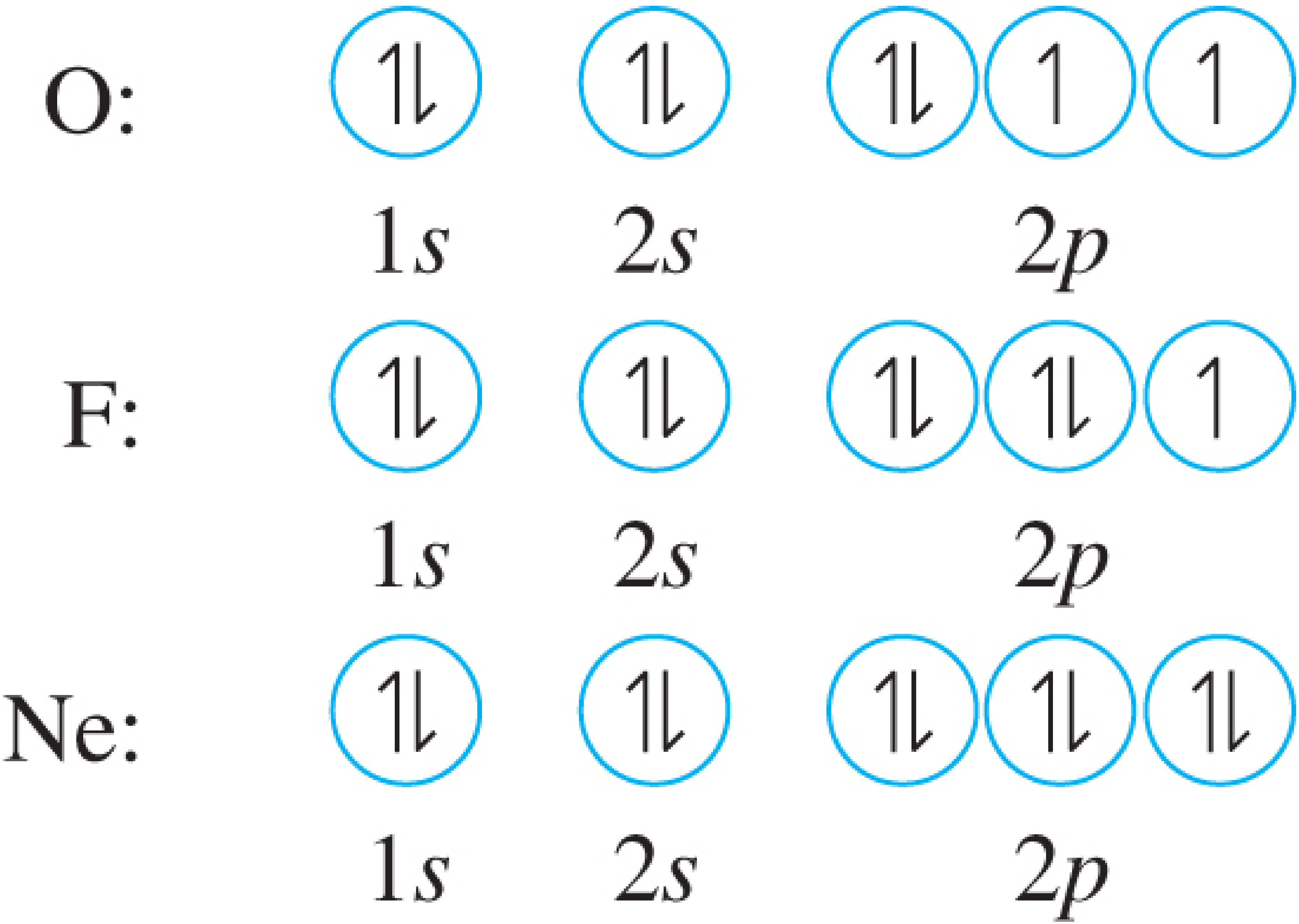

Li:
1s
2s
Be:
1s
2s

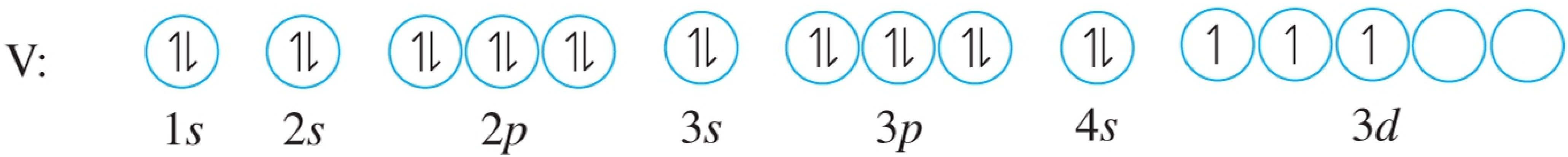

B:
1s
2s
2p


1
2p
or
1
2p
or
1
2p


C:
1s
2s
2p


N:
1s
2s
2p


O:
1s
2s
2p
F:
1s
2s
2p
Ne:
1s
2s
2p


V:
1s
2s
2p
3s
3p
4s
3d

$_{3}Li$: $1s^2 2s^1$

$_{11}Na$: $1s^2 2s^2 2p^6 3s^1$

$_{19}K$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

$_{37}Rb$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

$_{9}F$: $1s^2\,2s^2\,2p^5$

$_{17}Cl$: $1s^2 2s^2 2p^6\,3s^2\,3p^5$

$_{35}Br$: $1s^2 2s^2 2p^6 3s^2 3p^6\,4s^2\,3d^{10}\,4p^5$

$_{53}I$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6\,5s^2\,4d^{10}\,5p^5$

Periodic Trends

- Atomic radii
  - Increases going down a group
    - Due to electrons filling higher (larger) shells
  - Decreases going across a period
    - Due to increasing atomic charge and decreased shielding effects

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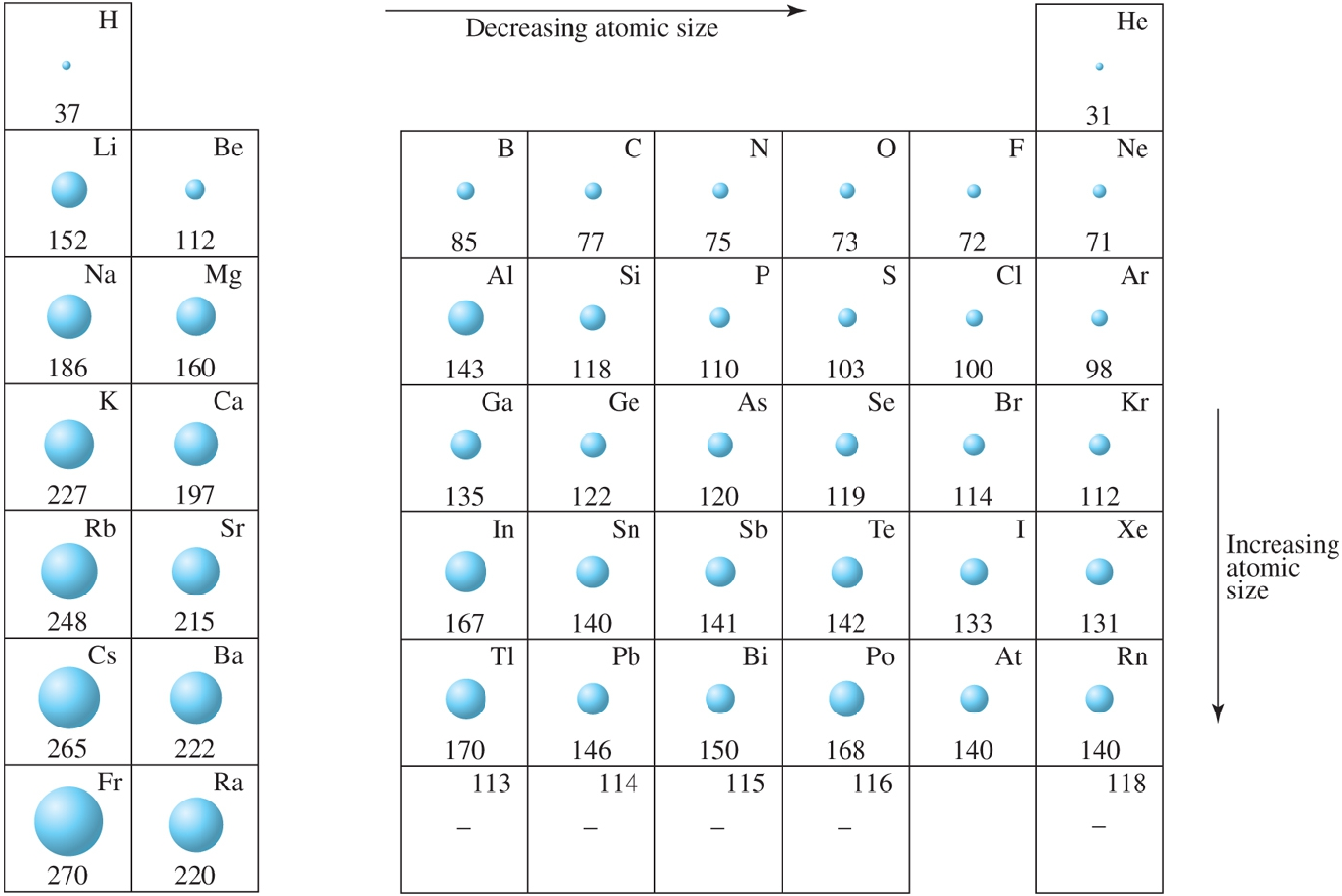
19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.87	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.69	29 **Cu** 63.55	30 **Zn** 65.38	31 **Ga** 69.72	32 **Ge** 72.64	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
37 **Rb** 85.47	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** 106.42	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
55 **Cs** 132.91	56 **Ba** 137.33	57 **La** 138.91	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)
87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (267)	105 **Db** (268)	106 **Sg** (271)	107 **Bh** (272)	108 **Hs** (270)	109 **Mt** (276)	110 **Ds** (281)	111 **Rg** (280)	112 **–** (285)	113 **–** (284)	114 **–** (289)	115 **–** (288)	116 **–** (293)		118 **–** (294)

Metals ← → Non-metals

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
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)


Decreasing atomic size
Increasing atomic size
H 37
He 31
Li 152
Be 112
B 85
C 77
N 75
O 73
F 72
Ne 71
Na 186
Mg 160
Al 143
Si 118
P 110
S 103
Cl 100
Ar 98
K 227
Ca 197
Ga 135
Ge 122
As 120
Se 119
Br 114
Kr 112
Rb 248
Sr 215
In 167
Sn 140
Sb 141
Te 142
I 133
Xe 131
Cs 265
Ba 222
Tl 170
Pb 146
Bi 150
Po 168
At 140
Rn 140
Fr 270
Ra 220
113 –
114 –
115 –
116 –
118 –

Ionization energy – energy needed to remove an outer electron from an atom (creates a positive ion)

e.g. $Li + energy \rightarrow Li^{+} + e^{-}$

- Increases up a group
  - Less shielding for smaller radii, so nucleus has stronger attraction for outer electrons

- Increases across a period
  - Also less shielding for smaller nuclei on the right side of a period

- Larger ion. en. → more difficult to ionize
- Lower ion. en. → more easily forms a positive ion

- Electron affinity – attraction for electrons, forming a negatively charged ion:

e.g. $F + e^- \rightarrow F^-$

- Increases up a group
  - Less shielding for smaller nuclei, so nucleus is more strongly attracted to an 'extra' electron.

- Increases across a period
  - Also less shielding for smaller nuclei...

- Larger magnitude el. aff. - more likely to form an anion (a negatively charged ion)

- Electronegativity – attraction for shared (bonding) electrons

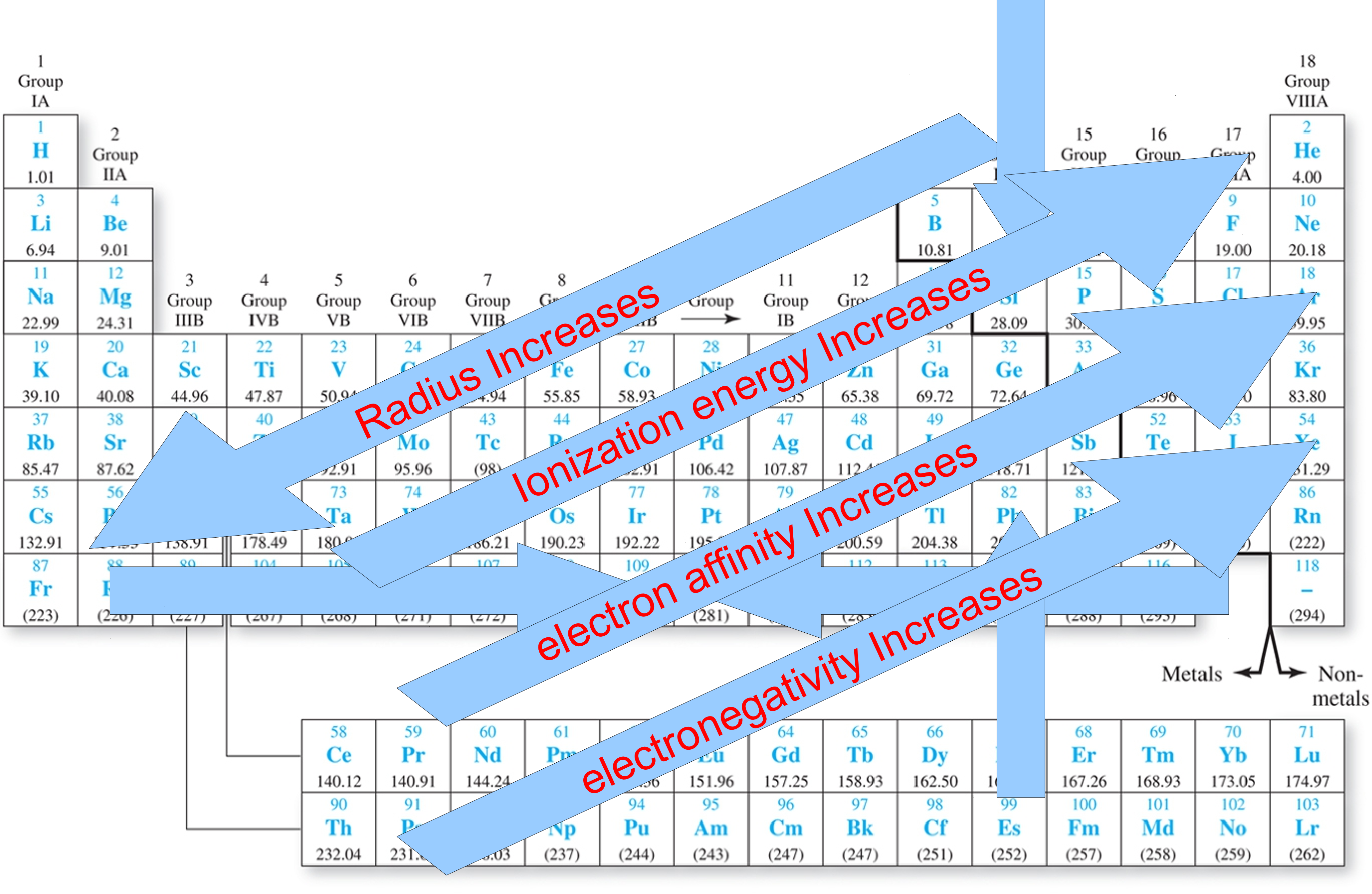
  - very similar to Electron Affinity

  - increases up a group
    - due to less shielding

  - increases across a period
    - also due to less shielding

    e.g.   H—Cl

    the chlorine atom is more electronegative and pulls more strongly on the shared electrons.


Radius Increases
Ionization energy Increases
electron affinity Increases
electronegativity Increases
Metals
Non-metals

Ions

- Cation – positively charged ions
  - Lost one or more electrons
  - Lower ionization energy – easier to form cations
- Anions – negatively charged ions
  - Gained one or more electrons
  - Higher electron affinity – stronger attraction for e^-, so easier to forms anions
- A numerical superscript after the element symbol is used to show the charge:
  Na^+, Ca^{+2}, O^{-2}, Br^{-1}