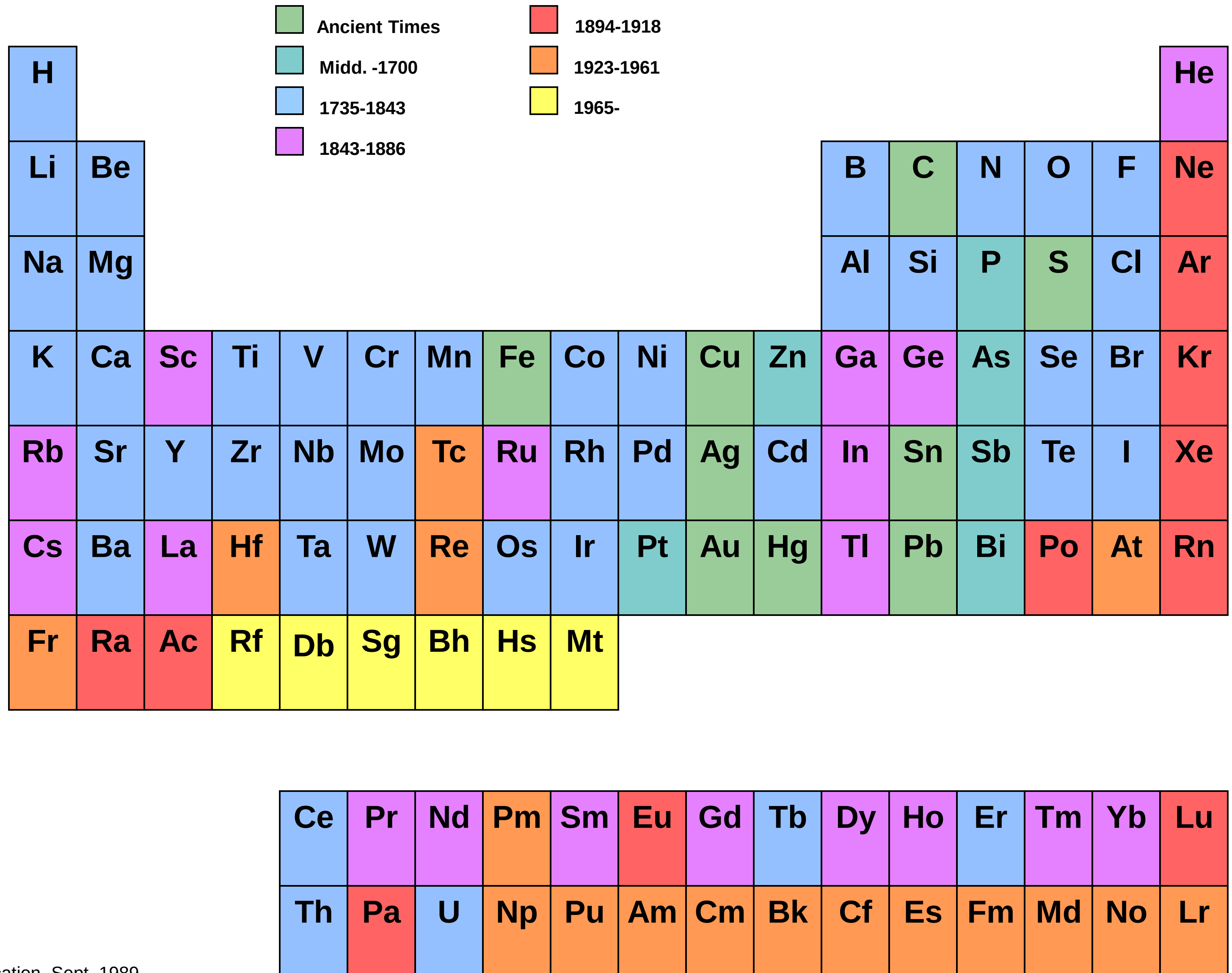
Periodic Trends

History of Periodic Table

- 1869: Dmitri Mendeleev organized the periodic table based on atomic weights
  - "Father of the Periodic Table"

- 1913: Henry Moseley rearranged the periodic table based on the positive charges in the nucleus
  - Lead to the periodic law: the states that a periodic pattern appears in the physical and chemical properties of the element when they are arranged in order of increasing atomic number

Discovering the Periodic Table



Journal of Chemical Education, Sept. 1989

Periodic Trends

- A pattern where elemental characteristics change predictably as you go across a period or down a group
- You need to know the trends in:
  - Metals v. Nonmetals
  - Atomic Radius (Atomic Size)
  - Ionization Energy
  - Electronegativity
  - Ion Formation

Groups of Elements

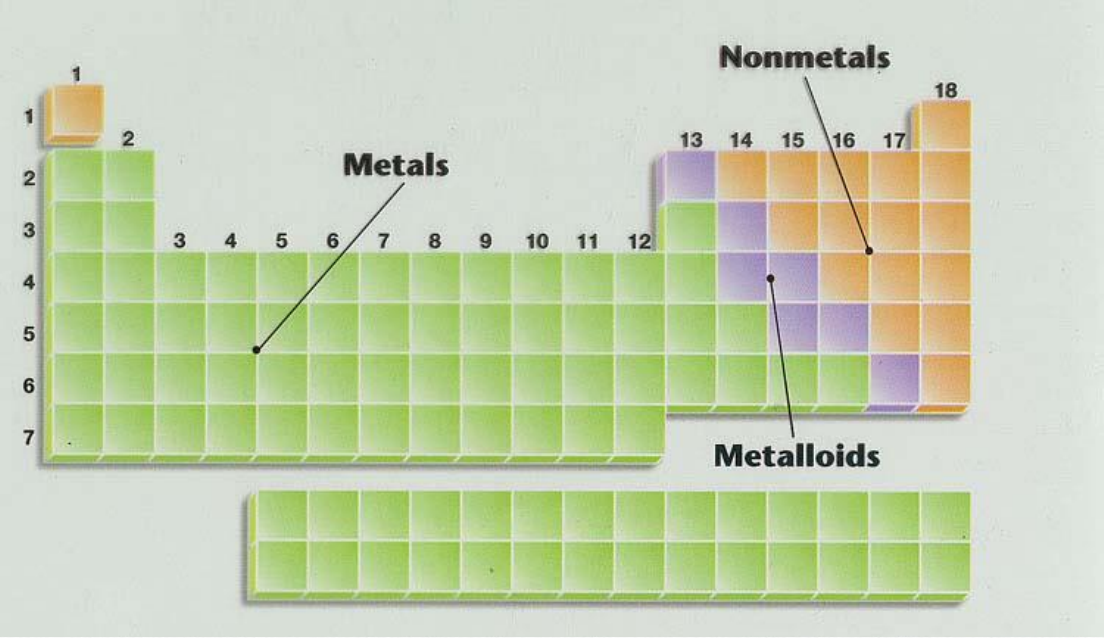
- 1A Alkali metals
- 2A Alkali earth metals
- Transition metals
- 3A Boron group
- 4A Carbon group
- Inner transition metals
- 5A Nitrogen group
- 6A Oxygen group
- 7A Halogens
- 8A Noble gases
- Hydrogen

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

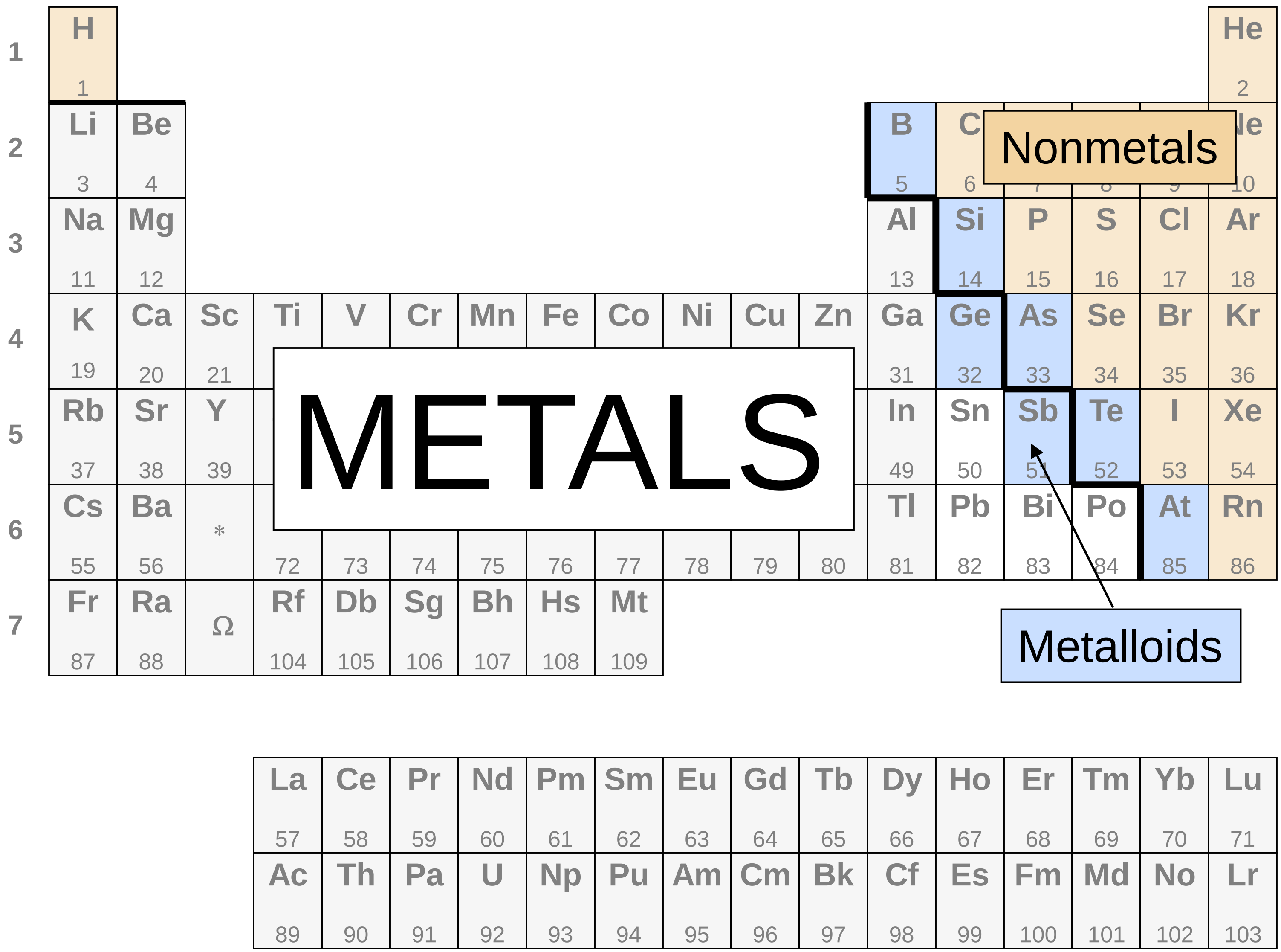
*	La 57	Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71
Ω	Ac 89	Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103

Metals vs. Nonmetals

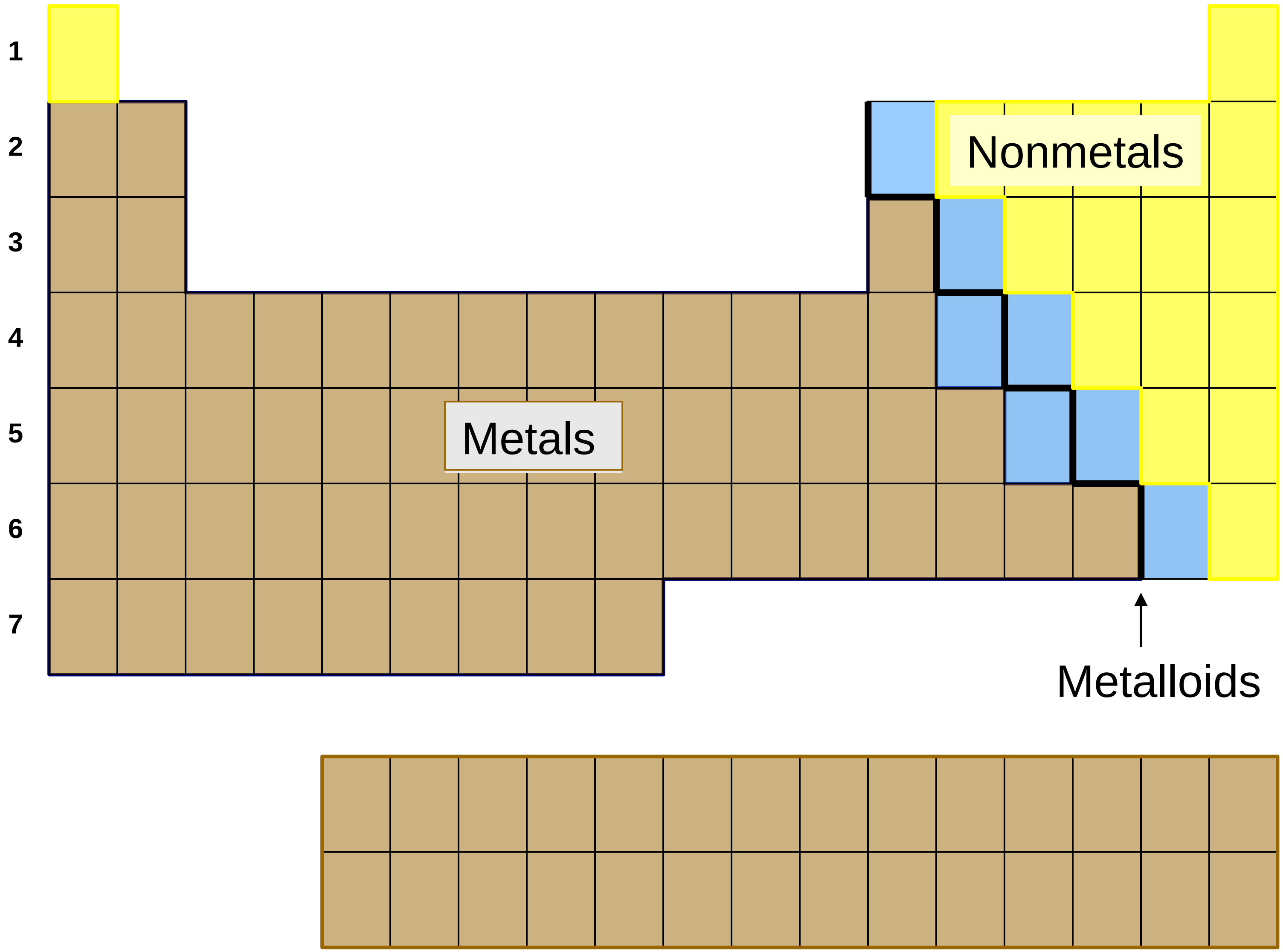
- Mainly divided into metals and nonmetals
- Metals: On the left-hand side (left of stair-step line)
- Non-metals: On the upper right-hand side
- Metalloids: On the stair-step line



Metals and Nonmetals



Metals, Nonmetals, & Metalloids



Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 349

Properties of Metals, Nonmetals, and Metalloids

METALS

malleable, lustrous, ductile, good conductors of heat and electricity

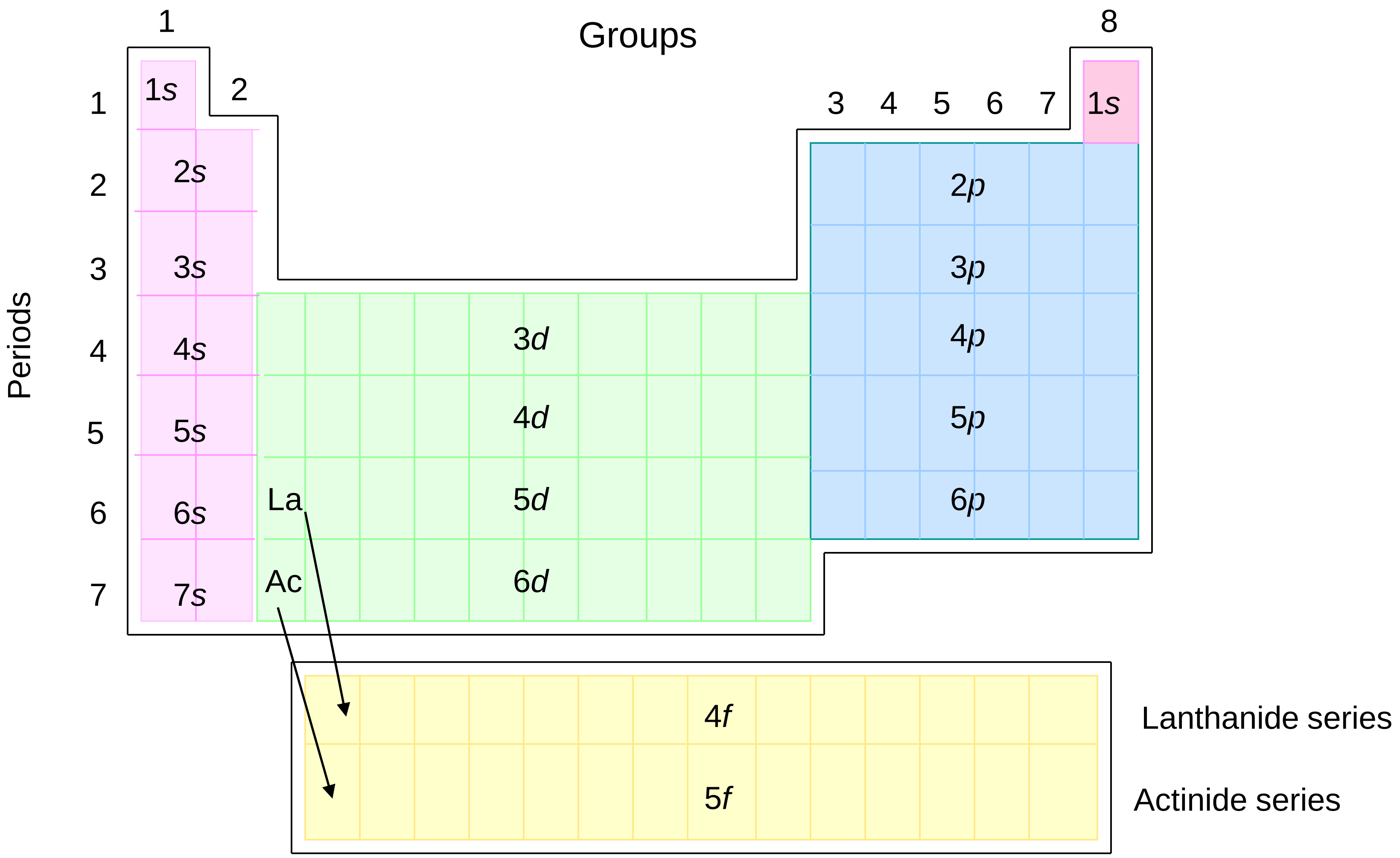
NONMETALS

gases or brittle solids at room temperature, poor conductors of heat and electricity (insulators)

METALLOIDS (Semi-metals)

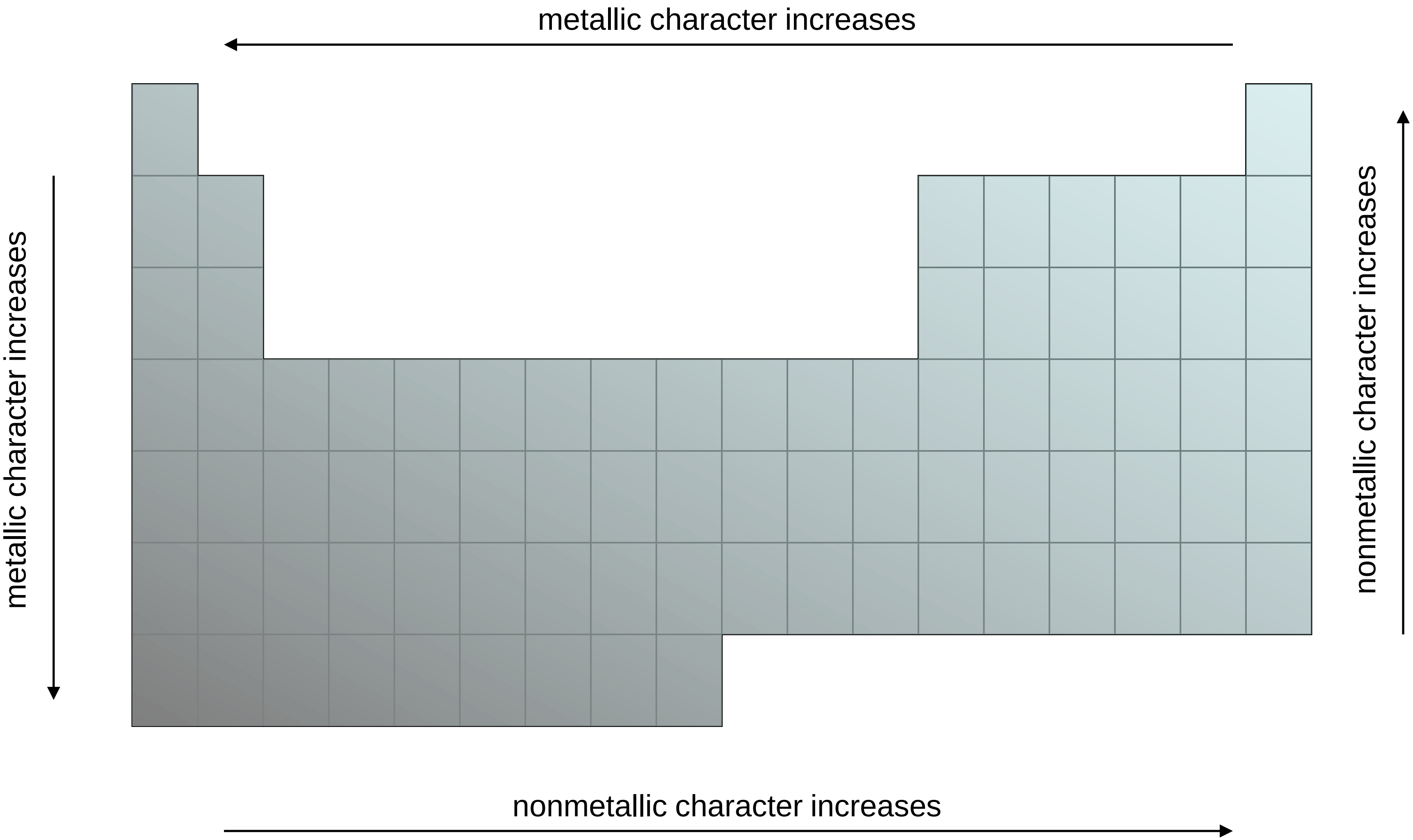
dull, brittle, semi-conductors (used in computer chips)

Orbitals Being Filled



Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 345

Electron Filling in Periodic Table



Melting Points

Mg 650

Symbol
Melting point °C

> 3000 °C

2000 - 3000 °C

1	H -259.2																	He 0.126 -269.7
2	Li 180.5	Be 1283											B 2027	C 4100	N -210.1	O -218.8	F -219.6	Ne -248.6
3	Na 98	Mg 650											Al 660	Si 1423	P 44.2	S 119	Cl -101	Ar -189.6
4	K 63.2	Ca 850	Sc 1423	Ti 1677	V 1917	Cr 1900	Mn 1244	Fe 1539	Co 1495	Ni 1455	Cu 1083	Zn 420	Ga 29.78	Ge 960	As 817	Se 217.4	Br -7.2	Kr -157.2
5	Rb 38.8	Sr 770	Y 1500	Zr 1852	Nb 2487	Mo 2610	Tc 2127	Ru 2427	Rh 1966	Pd 1550	Ag 961	Cd 321	In 156.2	Sn 231.9	Sb 630.5	Te 450	I 113.6	Xe -111.9
6	Cs 28.6	Ba 710	La 920	Hf 2222	Ta 2997	W 3380	Re 3180	Os 2727	Ir 2454	Pt 1769	Au 1063	Hg -38.9	Tl 303.6	Pb 327.4	Bi 271.3	Po 254	At	Rn -71

Ralph A. Burns, Fundamentals of Chemistry , 1999, page 1999

Densities of Elements

1	**H** 0.071																	**He** 0.126
2	**Li** 0.53	**Be** 1.8											**B** 2.5	**C** 2.26	**N** 0.81	**O** 1.14	**F** 1.11	**Ne** 1.204
3	**Na** 0.97	**Mg** 1.74											**Al** 2.70	**Si** 2.4	**P** 1.82w	**S** 2.07	**Cl** 1.557	**Ar** 1.402
4	**K** 0.86	**Ca** 1.55	**Sc** (2.5)	**Ti** 4.5	**V** 5.96	**Cr** 7.1	**Mn** 7.4	**Fe** 7.86	**Co** 8.9	**Ni** 8.90	**Cu** 8.92	**Zn** 7.14	**Ga** 5.91	**Ge** 5.36	**As** 5,7	**Se** 4.7	**Br** 3.119	**Kr** 2.6
5	**Rb** 1.53	**Sr** 2.6	**Y** 5.51	**Zr** 6.4	**Nb** 8.4	**Mo** 10.2	**Tc** 11.5	**Ru** 12.5	**Rh** 12.5	**Pd** 12.0	**Ag** 10.5	**Cd** 8.6	**In** 7.3	**Sn** 7.3	**Sb** 6.7	**Te** 6.1	**I** 4.93	**Xe** 3.06
6	**Cs** 1.90	**Ba** 3.5	**La** 6.7	**Hf** 13.1	**Ta** 16.6	**W** 19.3	**Re** 21.4	**Os** 22.48	**Ir** 22.4	**Pt** 21.45	**Au** 19.3	**Hg** 13.55	**Tl** 11.85	**Pb** 11.34	**Bi** 9.8	**Po** 9.4	**At** ---	**Rn** 4.4

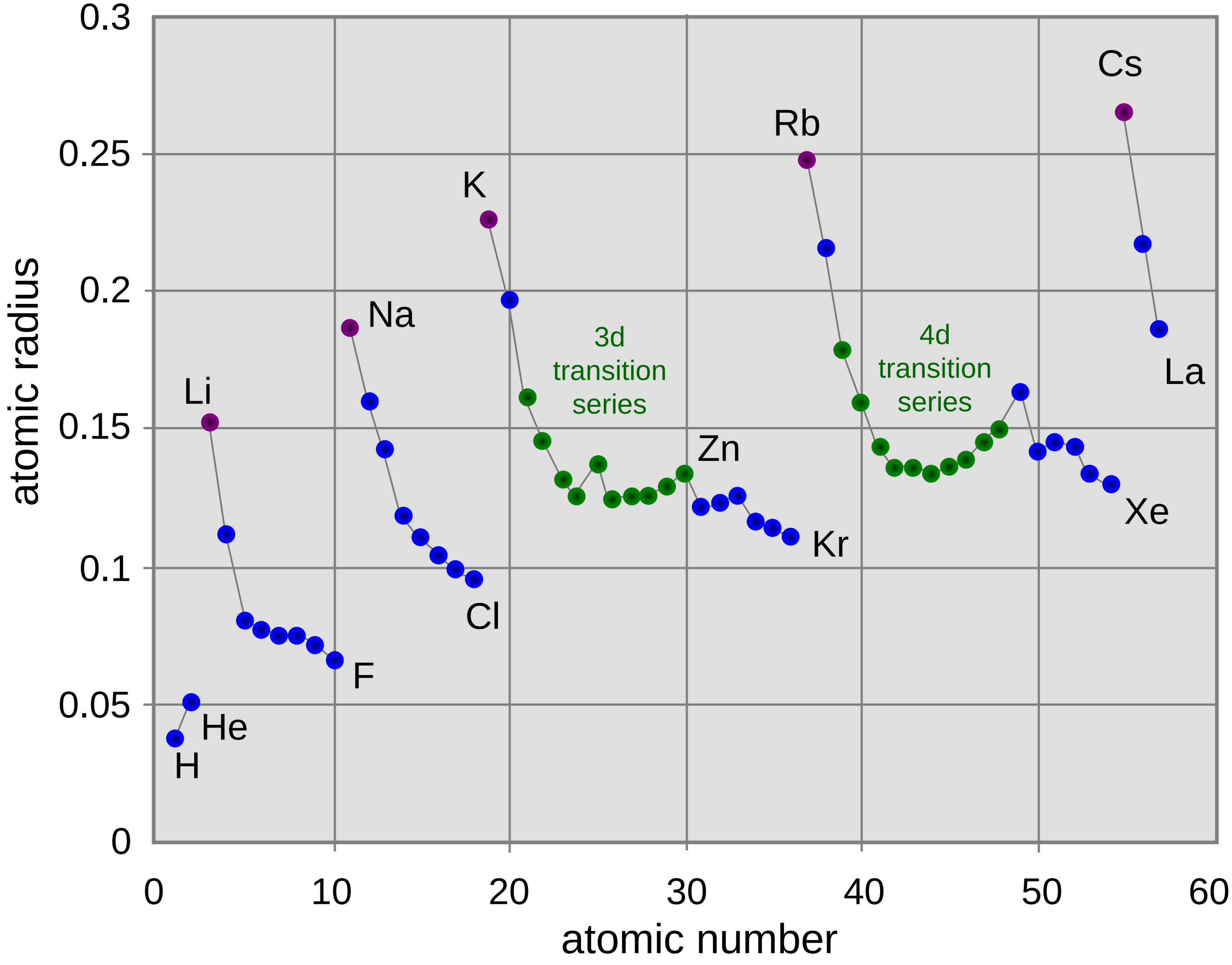
8.0 – 11.9 g/cm^3 12.0 – 17.9 g/cm^3 > 18.0 g/cm^3

Mg	Symbol
1.74	Density in g/cm^{3}C, for gases, in g/L

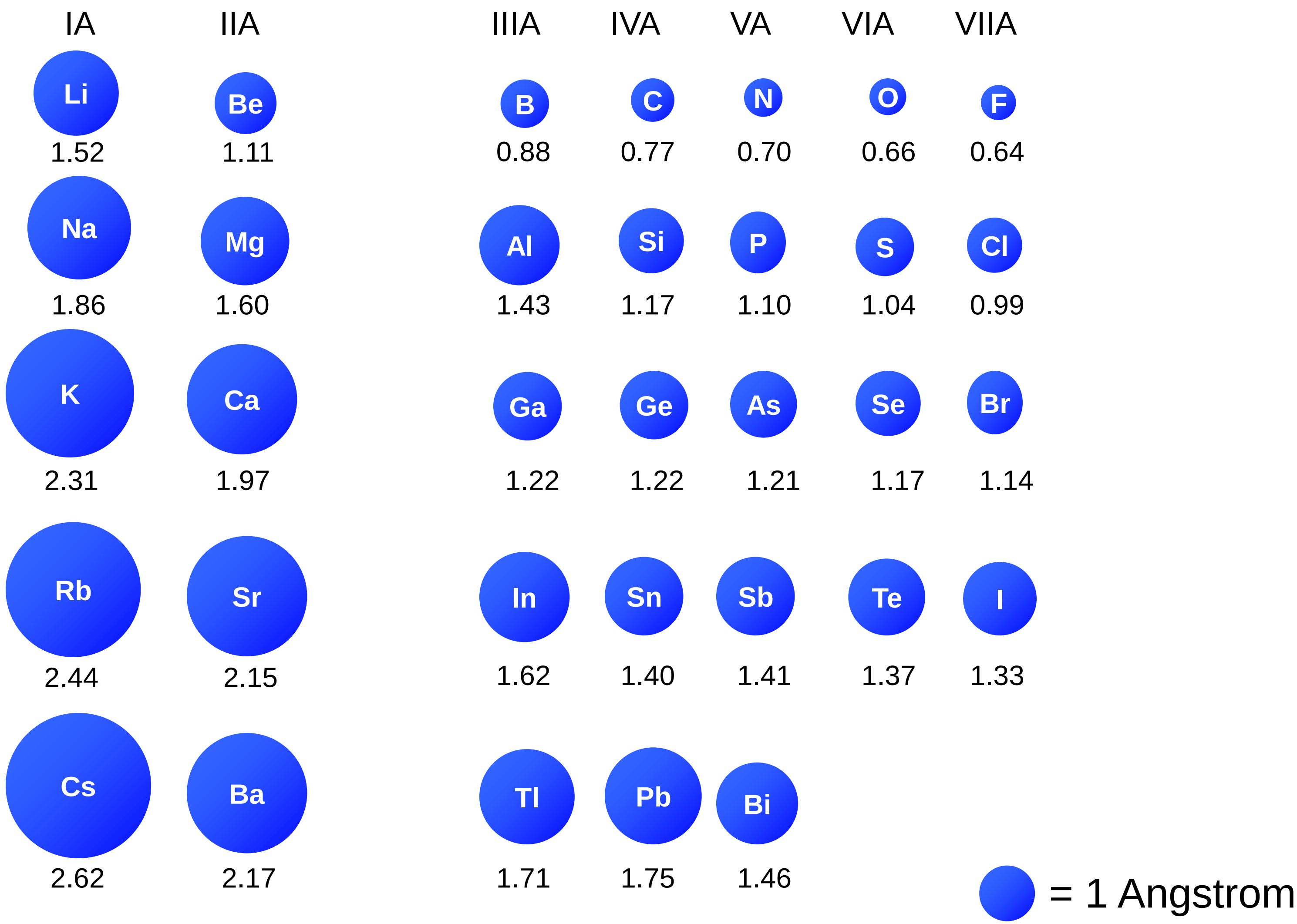
Atomic Radius

- The distance from the nucleus to the outer egde of the electron cloud
- **Increases** as you go **down** a group
  - Example: Calcium atoms are larger than beryllium atoms
- **Decreases** as you go **across** a period
  - Example: Fluorine atoms are smaller than Oxygen atoms
- Why?
  - Down groups, you have more electron shells
  - Across periods, you have the same number of electron shells, but more protons (positive). This pulls in the electrons.

Atomic Radius vs. Atomic Number



Atomic Radii



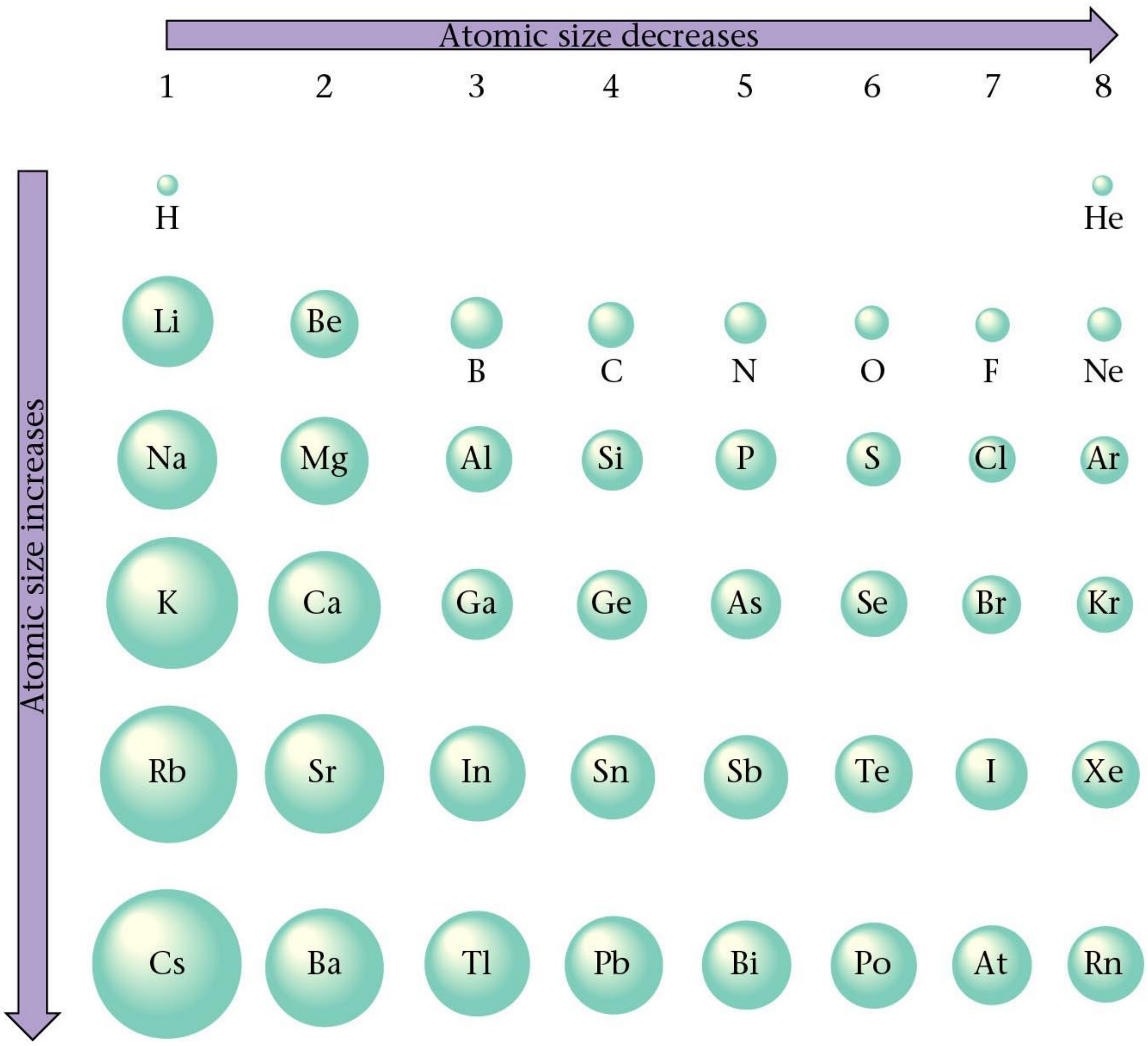
Periodic Trends in Atomic Radii

Atomic Radii of Representative Elements (nm)

1A	2A	3A	4A	5A	6A	7A
Li 0.152	Be 0.111	B 0.088	C 0.077	N 0.070	O 0.066	F 0.064
Na 0.186	Mg 0.160	Al 0.143	Si 0.117	P 0.110	S 0.104	Cl 0.099
K 0.231	Ca 0.197	Ga 0.122	Ge 0.122	As 0.121	Se 0.116	Br 0.115
Rb 0.244	Sr 0.215	In 0.162	Sn 0.14	Sb 0.141	Te 0.137	I 0.133
Cs 0.262	Ba 0.217	Tl 0.171	Pb 0.175	Bi 0.146	Po 0.14	At 0.140

LeMay Jr, Beall, Robblee, Brower, Chemistry Connections to Our Changing World , 1996, page 175

Relative Size of Atoms

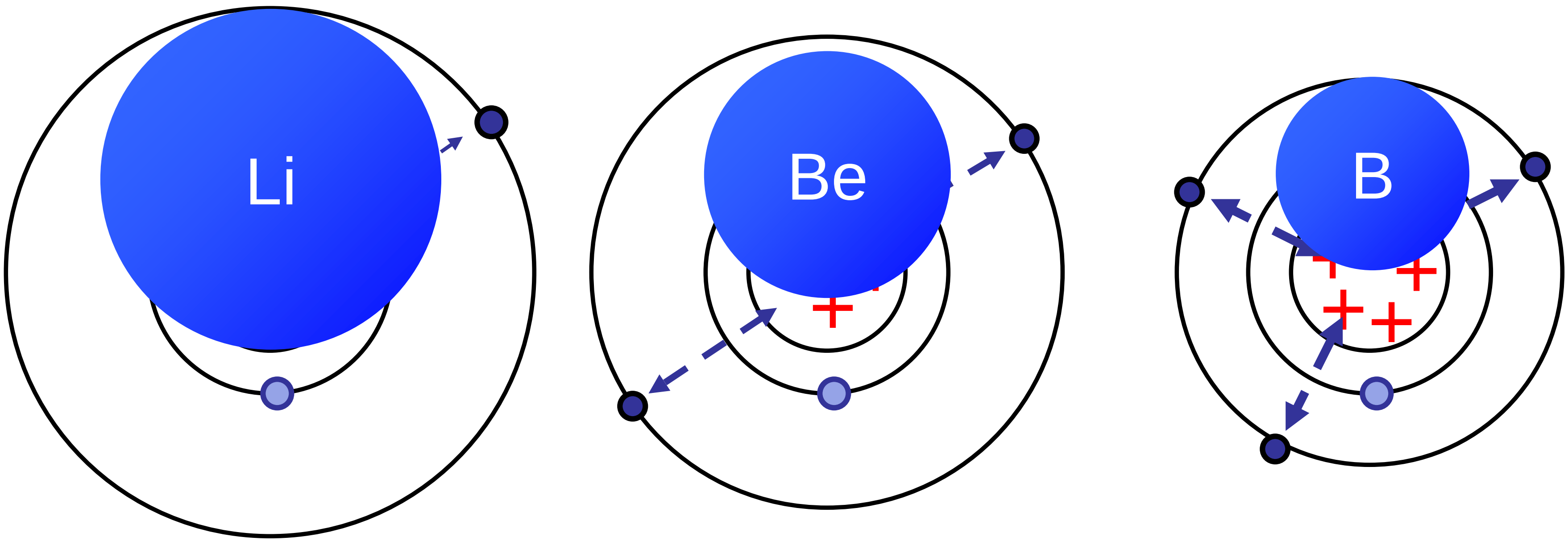


Zumdahl, Zumdahl, DeCoste, World of Chemistry 2002, page 350

Decreasing Atomic Size
Across a Period

- As the attraction between the (+) nucleus and the (–) valence electrons ↑, the atomic size ↓. Greater coulombic attraction.
- From left to right, size decreases because there is an increase in nuclear charge and Effective Nuclear Charge (# protons – # core electrons).
- Each valence electron is pulled by the full ENC

Li $1s^2 2s^1$ (ENC = 1) **Be** $1s^2 2s^2$ (ENC = 2) **B** $1s^2 2s^2 2p^1$ (ENC = 3)

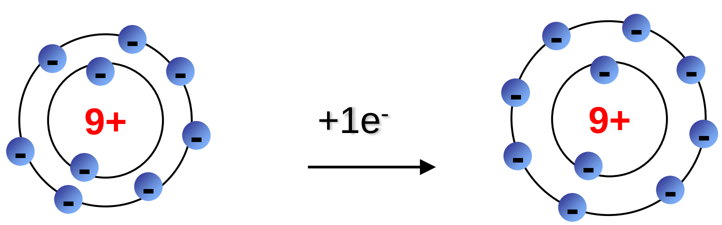


Ionic Radii

- The distance from the nucleus to the edge of the electron cloud of an ion
- **Increases** as you go **down** a group
  - Example: Calcium ions are larger than beryllium ions
- **Decreases** as you go **across** a period
  - Example: Fluorine ions are smaller than Oxygen ions
- Why?
  - Down groups, you have more electron shells
  - Across periods, you have the same number of electron shells, but more protons (positive). This pulls in the electrons.

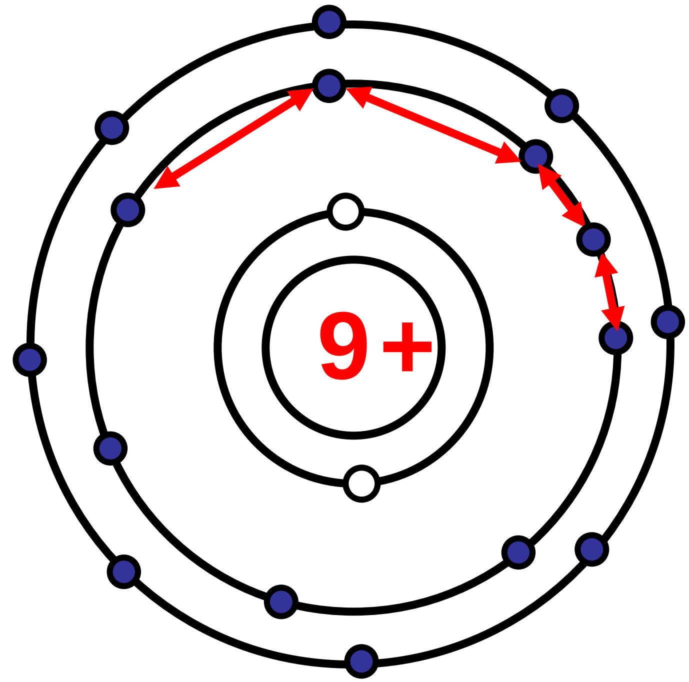
Sizes of ions: electron repulsion

- Valence electrons repel each other.
- When an atom becomes a anion (adds an electron to its valence shell) the repulsion between valence electrons increases without changing ENC
- Thus, F^{-} is larger than F



Fluorine atom
F
$1s^22s^22p^5$

Fluoride ion
F^{1-}
$1s^22s^22p^6$



Atomic Radii

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	Li 1.52	Be 1.11	B 0.88	C 0.77	N 0.70	O 0.66	F 0.64
	Na 1.86	Mg 1.60	Al 1.43	Si 1.17	P 1.10	S 1.04	Cl 0.99
	K 2.31	Ca 1.97	Ga 1.22	Ge 1.22	As 1.21	Se 1.17	Br 1.14
	Rb 2.44	Sr 2.15	In 1.62	Sn 1.40	Sb 1.41	Te 1.37	I 1.33
	Cs 2.62	Ba 2.17	Tl 1.71	Pb 1.75	Bi 1.46		

Ionic Radii

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
	Li^{1+} 0.60	Be^{2+} 0.31			N^{3-} 1.71	O^{2-} 1.40	F^{1-} 1.36
	Na^{1+} 0.95	Mg^{2+} 0.65	Al^{3+} 0.50			S^{2-} 1.84	Cl^{1-} 1.81
	K^{1+} 1.33	Ca^{2+} 0.99	Ga^{3+} 0.62			Se^{2-} 1.98	Br^{1-} 1.85
	Rb^{1+} 1.48	Sr^{2+} 1.13	In^{3+} 0.81			Te^{2-} 2.21	I^{1-} 2.16
	Cs^{1+} 1.69	Ba^{2+} 1.35	Tl^{3+} 0.95				

Cations: smaller than parent atoms

Anions: LARGER than parent atoms

● = 1 Angstrom

Electronegativity

- The ability of an atom to attract electrons towards itself from a covalent chemical bond
- Fluorine is the most electronegative element
- **Decreases** as you **go down** a group
  - Example: Chlorine has a lower electronegativity than Fluorine
- **Increases** as you **go across** a period
  - Example: Fluorine has a higher electronegativity than Oxygen
- Why?
  - Bigger atoms have a harder time pulling electrons in to themselves

Electronegativities

Period	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
1	**H** 2.1																	
2	**Li** 1.0	**Be** 1.5											**B** 2.0	**C** 2.5	**N** 3.0	**O** 3.5	**F** 4.0	
3	**Na** 0.9	**Mg** 1.2											**Al** 1.5	**Si** 1.8	**P** 2.1	**S** 2.5	**Cl** 3.0	
4	**K** 0.8	**Ca** 1.0	**Sc** 1.3	**Ti** 1.5	**V** 1.6	**Cr** 1.6	**Mn** 1.5	**Fe** 1.8	**Co** 1.8	**Ni** 1.8	**Cu** 1.9	**Zn** 1.7	**Ga** 1.6	**Ge** 1.8	**As** 2.0	**Se** 2.4	**Br** 2.8	
5	**Rb** 0.8	**Sr** 1.0	**Y** 1.2	**Zr** 1.4	**Nb** 1.6	**Mo** 1.8	**Tc** 1.9	**Ru** 2.2	**Rh** 2.2	**Pd** 2.2	**Ag** 1.9	**Cd** 1.7	**In** 1.7	**Sn** 1.8	**Sb** 1.9	**Te** 2.1	**I** 2.5	
6	**Cs** 0.7	**Ba** 0.9	**La*** 1.1	**Hf** 1.3	**Ta** 1.5	**W** 1.7	**Re** 1.9	**Os** 2.2	**Ir** 2.2	**Pt** 2.2	**Au** 2.4	**Hg** 1.9	**Tl** 1.8	**Pb** 1.8	**Bi** 1.9	**Po** 2.0	**At** 2.2	
7	**Fr** 0.7	**Ra** 0.9	**Ac**$^{\psi}$ 1.1															

*Lanthanides: 1.1 - 1.3

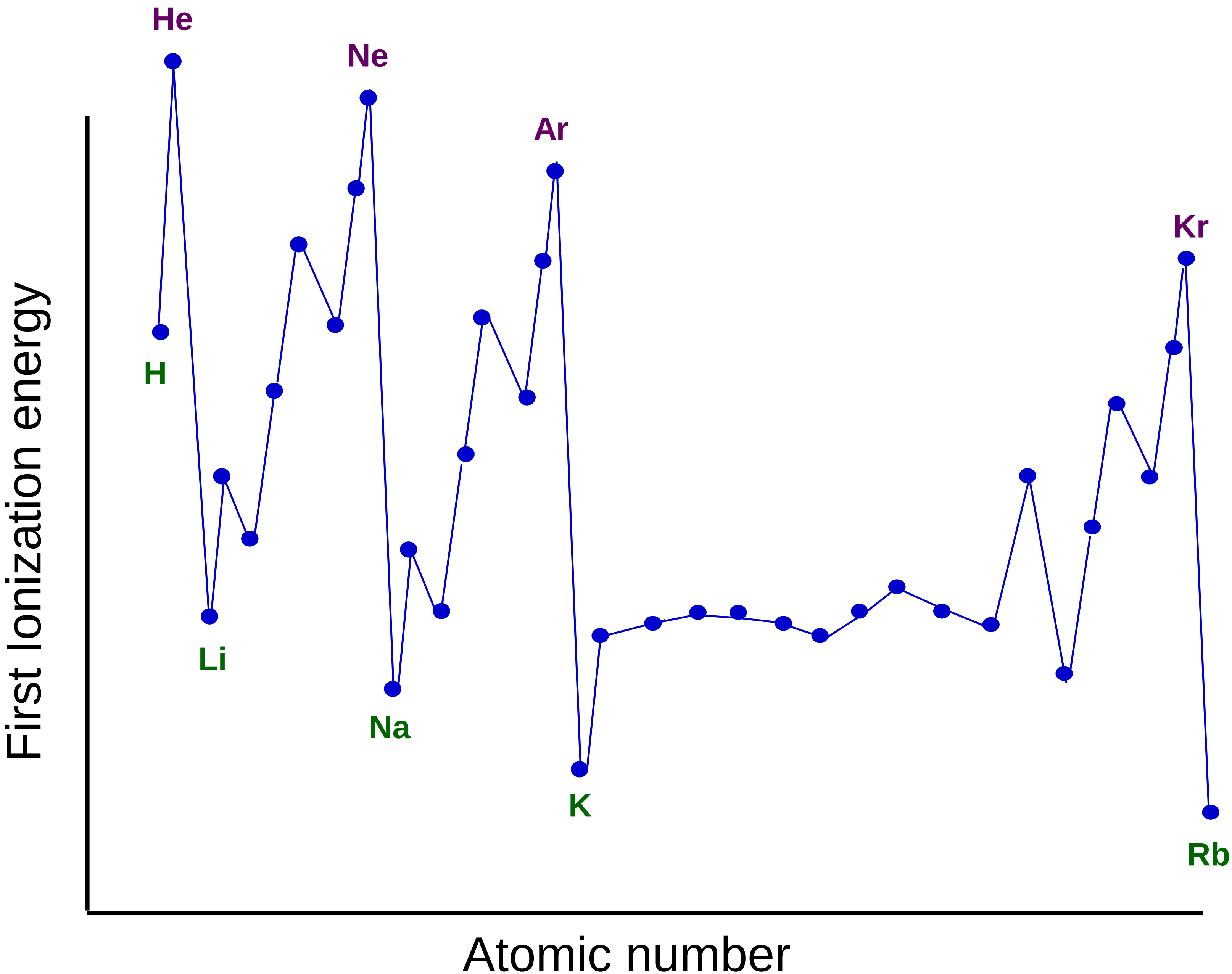
$^{\psi}$Actinides: 1.3 - 1.5

- Below 1.0
- 1.0 - 1.4
- 1.5 - 1.9
- 2.0 - 2.4
- 2.5 - 2.9
- 3.0 - 4.0

Hill, Petrucci, General Chemistry An Integrated Approach 2nd Edition, page 373

Ionization Energy

- The amount of energy needed to remove an electron from an atom
- **Decreases** as you go **down** a group
  - Example: Lithium has a higher ionization energy than sodium
- **Increases** as you go **across** a period
  - Example: Chlorine has a higher ionization energy than phosphorous
- Why?
  - Electrons are easier to remove from large elements
  - Electrons are harder to remove from atoms that almost have their “happy” full shell of 8 electrons


He
Ne
Ar
Kr
H
Li
Na
K
Rb
First Ionization energy
Atomic number

Ionization Energies

- It takes more energy to remove the second electron from an atom than the first, and so on.

- There are two reasons for this trend:

  1. The second electron is being removed from a positively charged species rather than a neutral one, so more energy is required.

  2. Removing the first electron reduces the repulsive forces among the remaining electrons, so the attraction of the remaining electrons to the nucleus is stronger.

- Energy required to remove electrons from a filled core is prohibitively large and simply cannot be achieved in normal chemical reactions.

Factors Affecting Ionization Energy

Nuclear Charge

The larger the nuclear charge, the greater the ionization energy.

Shielding effect

The greater the shielding effect, the less the ionization energy.

Radius

The greater the distance between the nucleus and the outer electrons of an atom, the less the ionization energy.

Sublevel

An electron from a full or half-full sublevel requires additional energy to be removed.

Smoot, Price, Smith, Chemistry A Modern Course 1987, page 189

Summary of Periodic Trends