

Periodic Table of the Elements

1 H Hydrogen 1.01																	2 He Helium 4.00
3 Li Lithium 6.94	4 Be Beryllium 9.01											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.06	17 Cl Chlorine 35.45	18 Ar Argon 39.95
19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.88	23 V Vanadium 50.94	24 Cr Chromium 51.99	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.38	31 Ga Gallium 69.72	32 Ge Germanium 72.63	33 As Arsenic 74.92	34 Se Selenium 78.97	35 Br Bromine 79.90	36 Kr Krypton 83.80
37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.95	43 Tc Technetium 98.91	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.91	46 Pd Palladium 106.42	47 Ag Silver 107.87	48 Cd Cadmium 112.41	49 In Indium 114.82	50 Sn Tin 118.71	51 Sb Antimony 121.76	52 Te Tellurium 127.6	53 I Iodine 126.90	54 Xe Xenon 131.29
55 Cs Cesium 132.91	56 Ba Barium 137.33	57-71 Lanthanides	72 Hf Hafnium 178.49	73 Ta Tantalum 180.95	74 W Tungsten 183.85	75 Re Rhenium 186.21	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.97	80 Hg Mercury 200.59	81 Tl Thallium 204.38	82 Pb Lead 207.20	83 Bi Bismuth 208.98	84 Po Polonium [208.98]	85 At Astatine 209.98	86 Rn Radon 222.02
87 Fr Francium 223.02	88 Ra Radium 226.03	89-103 Actinides	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [277]	109 Mt Meitnerium [278]	110 Ds Darmstadtium [281]	111 Rg Roentgenium [282]	112 Cn Copernicium [285]	113 Nh Nihonium [286]	114 Fl Flerovium [289]	115 Mc Moscovium [289]	116 Lv Livermorium [293]	117 Ts Tennessine [294]	118 Og Oganesson [294]

General Chemistry

202-SN1-RE

with Olivia Bibollet-Bahena

Office: 5th floor

57 La Lanthanum 138.91	58 Ce Cerium 140.12	59 Pr Praseodymium 140.91	60 Nd Neodymium 144.24	61 Pm Promethium 144.91	62 Sm Samarium 150.36	63 Eu Europium 151.96	64 Gd Gadolinium 157.25	65 Tb Terbium 158.93	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93	68 Er Erbium 167.26	69 Tm Thulium 168.93	70 Yb Ytterbium 173.06	71 Lu Lutetium 174.97
89 Ac Actinium 227.03	90 Th Thorium 232.04	91 Pa Protactinium 231.04	92 U Uranium 238.03	93 Np Neptunium 237.05	94 Pu Plutonium 244.06	95 Am Americium 243.06	96 Cm Curium 247.07	97 Bk Berkelium 247.07	98 Cf Californium 251.08	99 Es Einsteinium [254]	100 Fm Fermium 257.10	101 Md Mendelevium 258.10	102 No Nobelium 259.10	103 Lr Lawrencium [262]

- Alkali Metal
- Alkaline Earth
- Transition Metal
- Basic Metal
- Metalloid
- Nonmetal
- Halogen
- Noble Gas
- Lanthanide
- Actinide

Aufbau Principle - Summary

Aufbau Principle (“building up”): A guide for determining the filling order of orbitals.

As protons are added one by one to the nucleus to build up the elements, electrons are similarly added to orbitals.

Hund's Rule

The lowest energy configuration for an atom is the one having the maximum number of unpaired electrons allowed by the **Pauli principle** in a particular set of degenerate (same energy) orbitals.

Electron Configurations of Multielectron Atoms

- **Electron Configuration:** A description of which orbitals are occupied by electrons
- **Degenerate Orbitals:** Orbitals that have the same energy level—for example, the three p orbitals in a given subshell
- **Ground-State Electron Configuration:** The lowest-energy configuration

Electronic Configuration and the Periodic Table

Corresponds to quantum number n

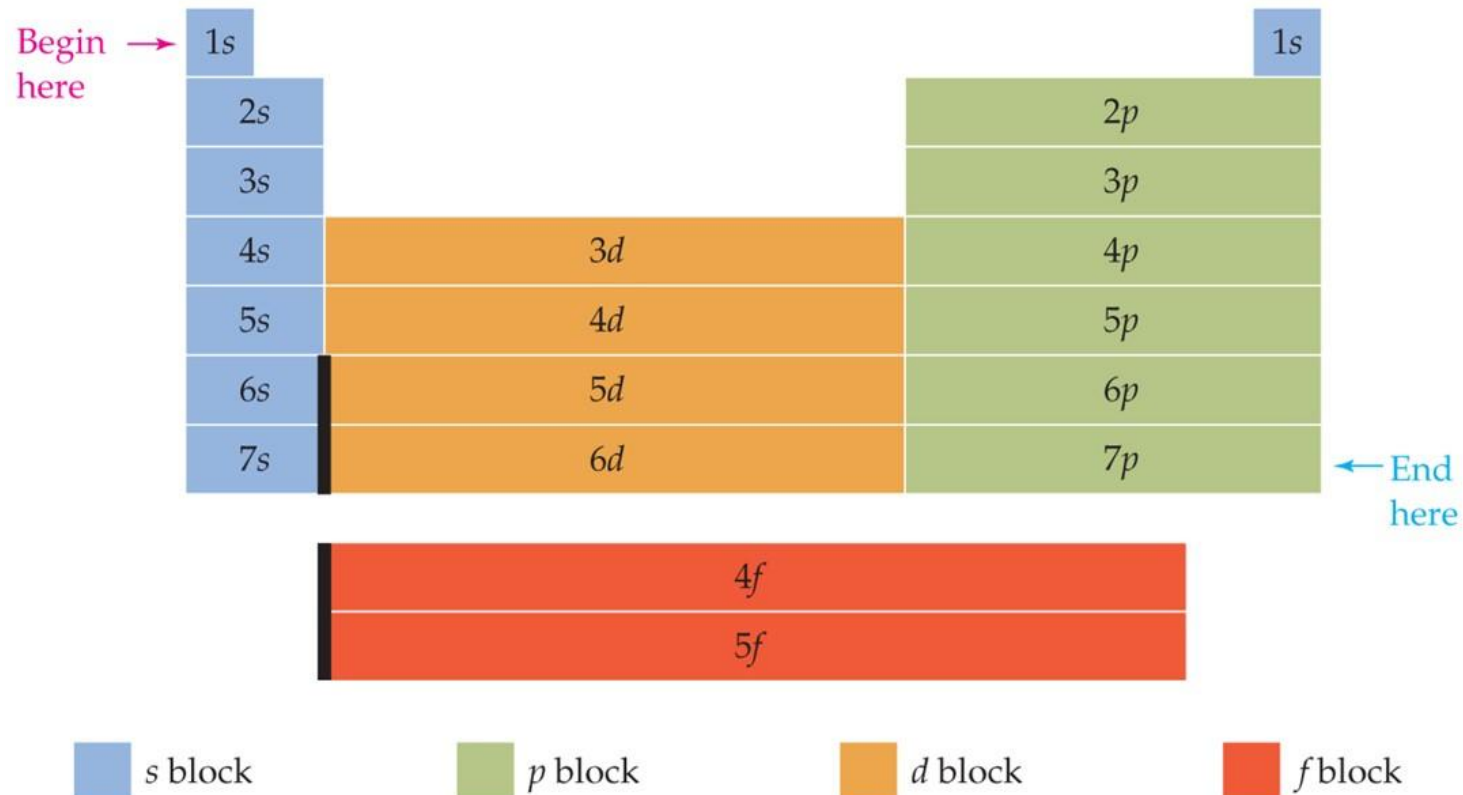


Group

Period	1A	2A							3A	4A	5A	6A	7A	8A
1	1s													1s
2	2s										2p			
3	3s										3p			
4	4s				3d					4p				
5	5s				4d					5p				
6	6s	La			5d					6p				
7	7s	Ac			6d									
								4f						
								5f						

Electron Configurations and the Periodic Table

The arrangement of the periodic table provides a method for remembering the order of orbital filling. Beginning at the top left and moving across successive rows, the order is $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p$ and so on.



Electronic Configuration and Orbital Diagram

- A notation that shows how many electrons an atom has in each of its occupied electron orbitals.

Oxygen: $1s^2 2s^2 2p^4$

Oxygen: $1s$ $2s$ $2p$



Electron Configurations of Multielectron Atoms

	Electron Configuration	Shorthand Configuration
P:	-----	-----
K:	-----	-----
Sc:	-----	-----

Anomalous Electron Configurations

	Expected Configuration	Actual Configuration
Cr:	$[\text{Ar}] 4s^2 3d^4$	$[\text{Ar}] 4s^1 3d^5$
Cu:	$[\text{Ar}] 4s^2 3d^9$	$[\text{Ar}] 4s^1 3d^{10}$

Exceptions to Hund's Rule

4s orbitals are only **half filled**.
Reason is beyond our scope.

K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
$4s^1$	$4s^2$	$4s^23d^1$	$4s^23d^2$	$4s^23d^3$	$4s^1 3d^5$	$4s^23d^5$	$4s^23d^6$	$4s^23d^7$	$4s^23d^8$	$4s^1 3d^{10}$	$4s^23d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$

Mo

Ag

Au

Rg

Chromium $4s^1 3d^5$

Molybdenum $5s^1 4d^5$

Copper $4s^1 3d^{10}$

Silver $5s^1 4d^{10}$

Gold $6s^1 5d^{10}$

Roentgenium $7s^1 6d^{10}$

Valence Electrons

- The electrons in the **outermost** principal quantum level of an atom.

Ne: $1s^2 2s^2 2p^6$ (valence electrons = 8)

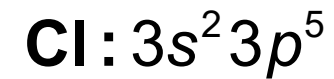
- The elements in the **same group** on the periodic table have the same valence electron configuration.
- The number of valence electrons is the same as the groups labeled 1A, 2A, 3A, 4A, 5A, 6A, 7A and 8A.

1A									8A
H $1s^1$									He $1s^2$
	2A			3A	4A	5A	6A	7A	
Li $2s^1$	Be $2s^2$			B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
Na $3s^1$	Mg $3s^2$			Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$

Electron Configurations and the Periodic Table

- **Valence Shell:** Outermost shell

Group	Valence-Shell Electron Configuration	
1A	ns^1	(1 total)
2A	ns^2	(2 total)
3A	ns^2np^1	(3 total)
4A	ns^2np^2	(4 total)
5A	ns^2np^3	(5 total)
6A	ns^2np^4	(6 total)
7A	ns^2np^5	(7 total)
8A	ns^2np^6	(8 total)



Exercise

In which of the following groups do all the elements have the same number of valence electrons?

- a) P, S, Cl
- b) Ag, Cd, Ar
- c) Na, Ca, Ba
- d) P, As, Se
- e) None of these

Ions: Electron Configuration

Stable Compounds

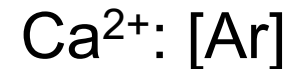
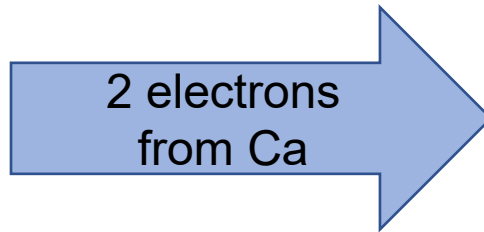
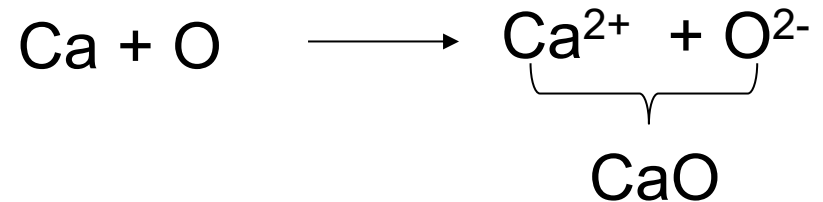
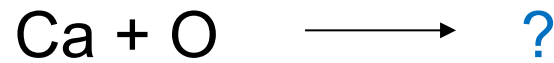
- Quantum Mechanics has enhanced our understanding of what constitutes a stable compound
- Atoms in stable compounds usually have a noble gas electron configuration.

Electron Configurations in Stable Compounds

- When *two nonmetals* react to form a covalent bond, they share electrons in a way that completes the valence electron configurations of both atoms.
- When a *nonmetal and a main-group metal* react to form a binary ionic compound,
 - the ions form so that the valence electron configuration of the *nonmetal* achieves the electron configuration of the next noble gas atom
 - and
 - the valence orbitals of the *metal* are emptied to obtain the previous noble gas configuration.

Electronic Configurations of Ions

- The formula of an ionic compound can be predicted by considering the valence electron configurations of the two atoms.
- Atoms want to achieve a noble gas configuration.



$\underbrace{\hspace{10em}}$
O²⁻ is the same as [Ne]

Topic 4c – Atomic Periodicity

The Predictive Power of the Quantum Mechanical Model

Elements in the **same column (group)** of the periodic table have similar chemical and physical properties because they have the same number of valence electrons in the same kinds of orbitals.

8A
2 He $1s^2$
10 Ne $2s^2 2p^6$
18 Ar $3s^2 3p^6$
36 Kr $4s^2 4p^6$
54 Xe $5s^2 5p^6$
86 Rn $6s^2 6p^6$
Noble gases

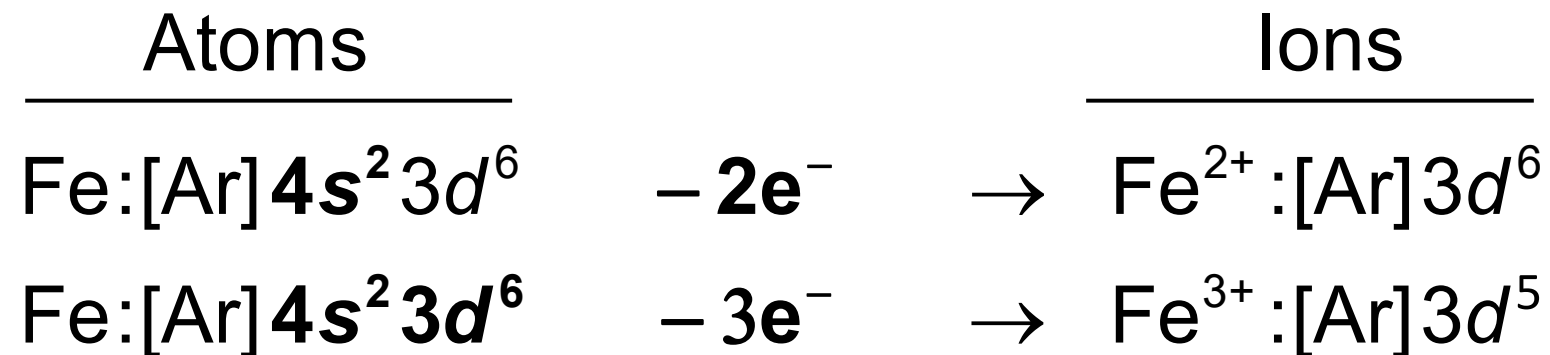
Electron Configuration and Ion Charge

Elements That Form Ions with Predictable Charges

	1A	2A											3A	4A	5A	6A	7A	8A
1	Li ⁺														N ³⁻	O ²⁻	F ⁻	
2	Na ⁺	Mg ²⁺											Al ³⁺			S ²⁻	Cl ⁻	
3	K ⁺	Ca ²⁺														Se ²⁻	Br ⁻	
4	Rb ⁺	Sr ²⁺														Te ²⁻	I ⁻	
5	Cs ⁺	Ba ²⁺																

© 2010 Pearson Education, Inc.

Electron Configurations of Ions



Keep in mind the general pattern of removing electrons.

Periodic Trends in Atomic Properties

Atomic Properties:

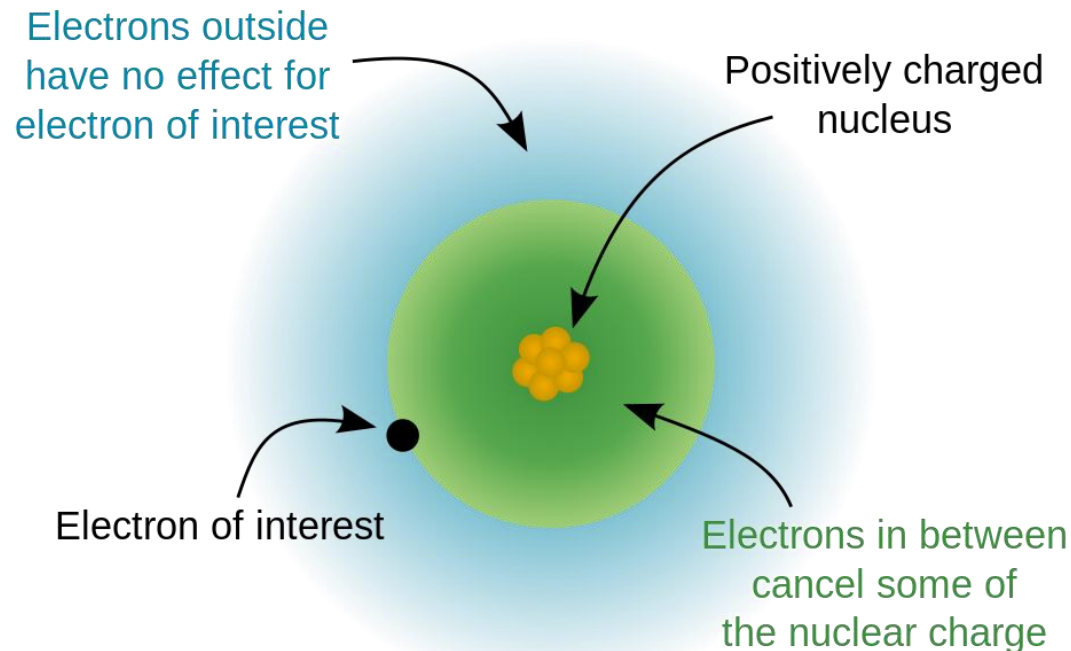
- 1) Ionization energy
- 2) Atomic radius
- 3) Ionic radius
- 4) Electronegativity
- 5) Electron affinity

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

Variations in atomic properties can be predicted using the periodic table.

Periodic Trends

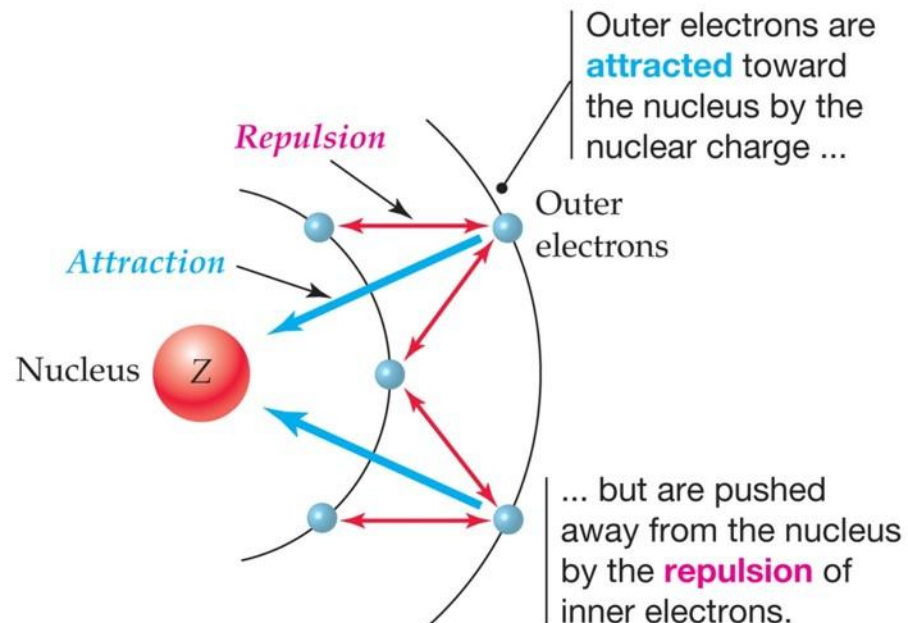
- The nuclear charge that an electron would experience if there were no intervening electrons is Z , the atomic number.
- The nuclear charge that an electron experiences is reduced by intervening electrons (by the shielding effect) to a value of Z_{eff} , called the **effective nuclear charge**.
- The key to understanding all periodic trends is to understand Z_{eff} !



Orbital Energy Levels in Multielectron Atoms

- **Effective Nuclear Charge (Z_{eff})**: The nuclear charge actually felt by an electron.

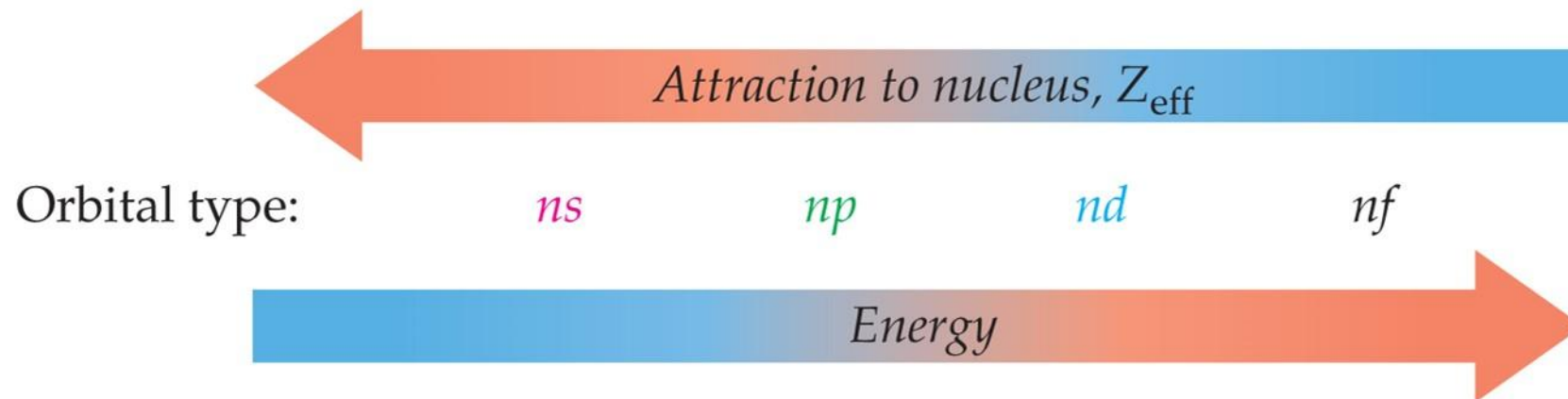
$$Z_{\text{eff}} = Z_{\text{actual}} - \text{Electron shielding}$$



Orbital Energy Levels in Multielectron Atoms

- **Effective Nuclear Charge (Z_{eff})**: The nuclear charge actually felt by an electron.

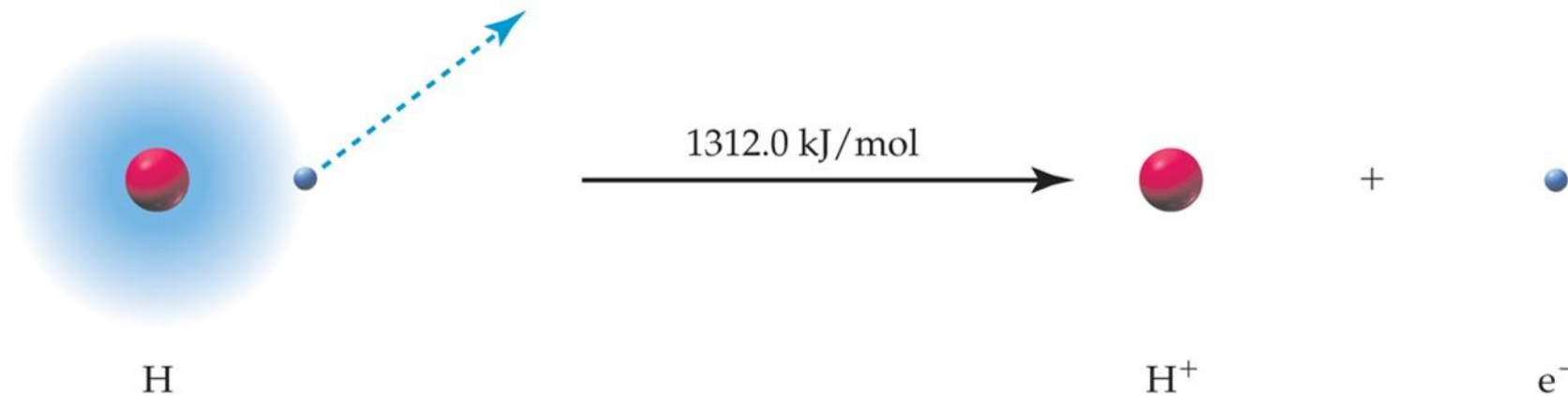
$$Z_{\text{eff}} = Z_{\text{actual}} - \text{Electron shielding}$$



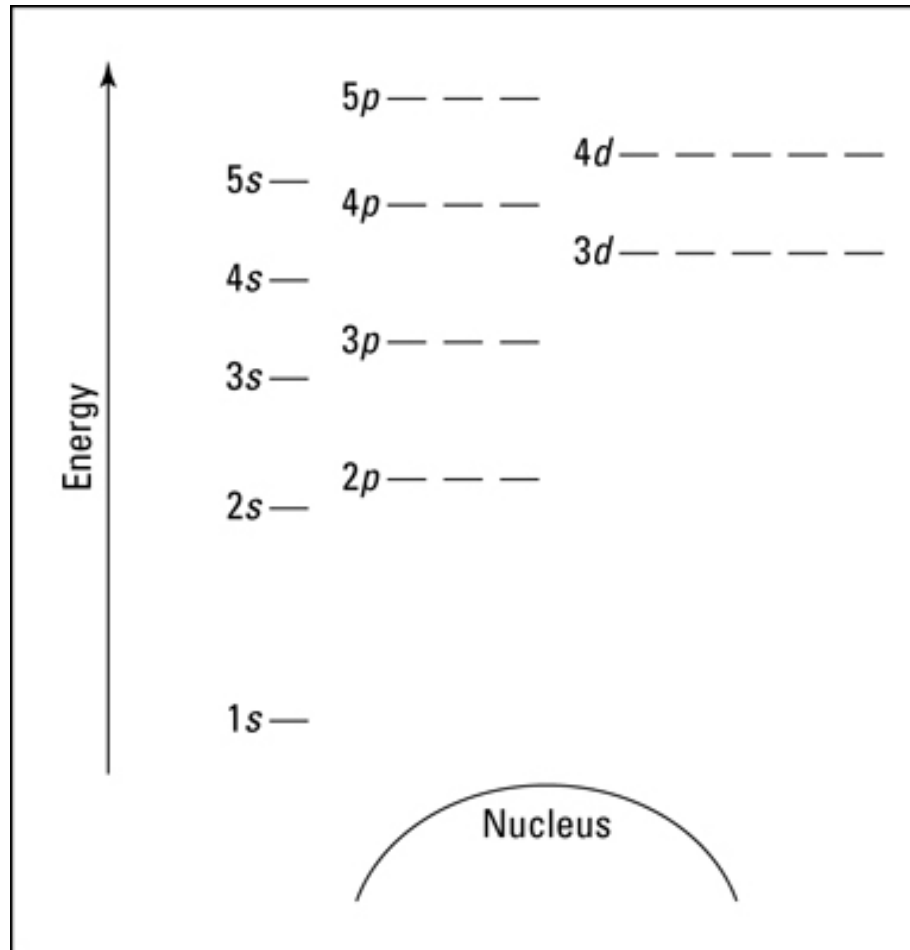
Ionization Energy

Ionization Energy

Ionization Energy (E_i): The amount of energy necessary to remove the highest-energy electron from an isolated neutral atom in the gaseous state.



Polyelectronic Atoms



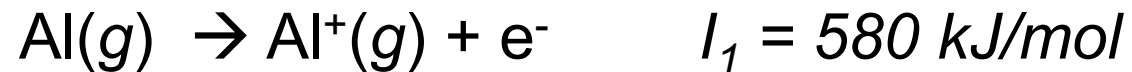
The 3p orbital is higher in energy than the 3s orbital because of the penetration effect.

Ionization Energy

Ionization energy: the energy required to remove an electron from a gaseous atom or ion.



For example, Al(g)



Why is $I_2 > I_1$?

Ionization Energy

Ionization energy: the energy required to remove an electron from a gaseous atom or ion.

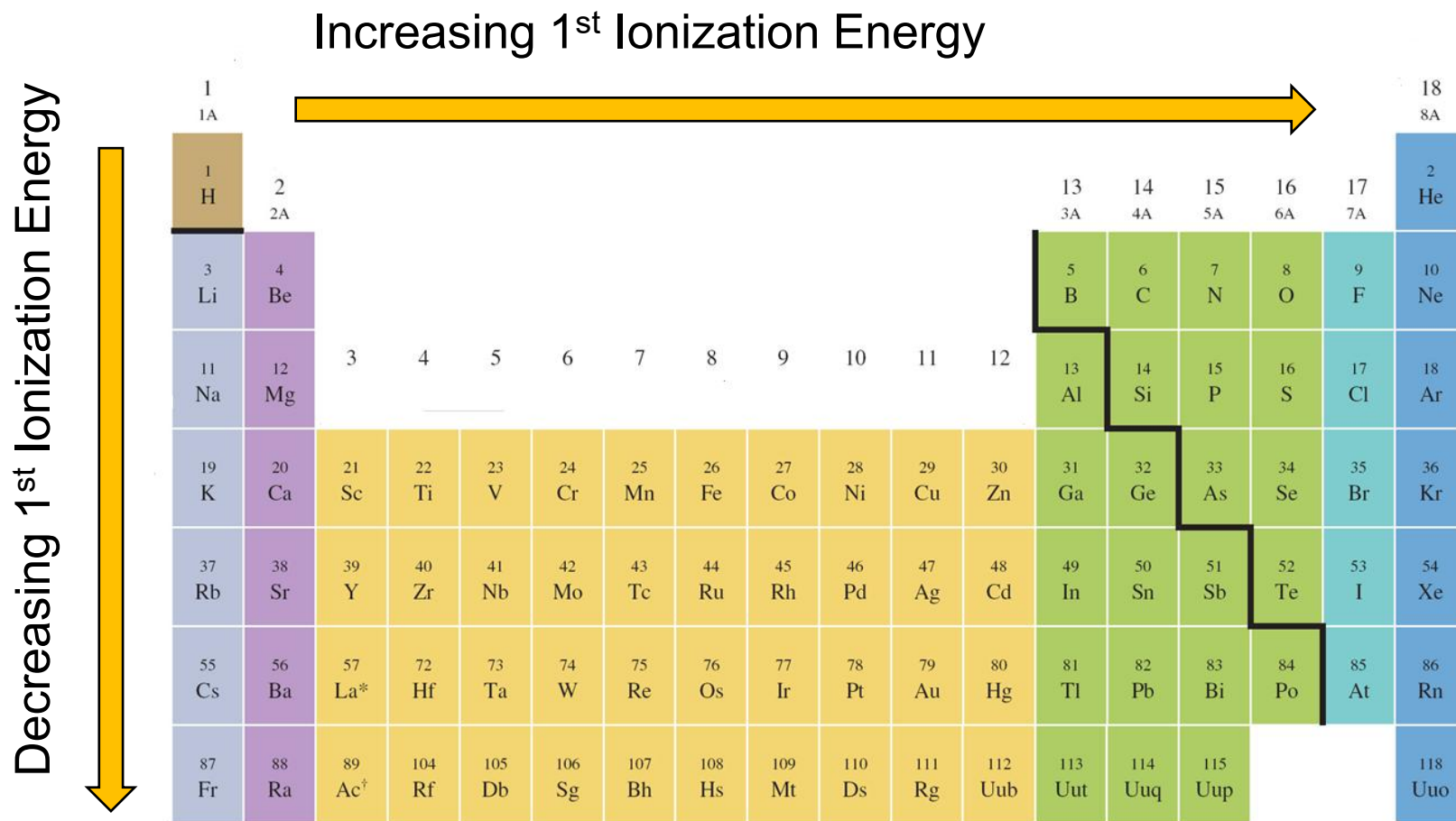


For example, Al(*g*)

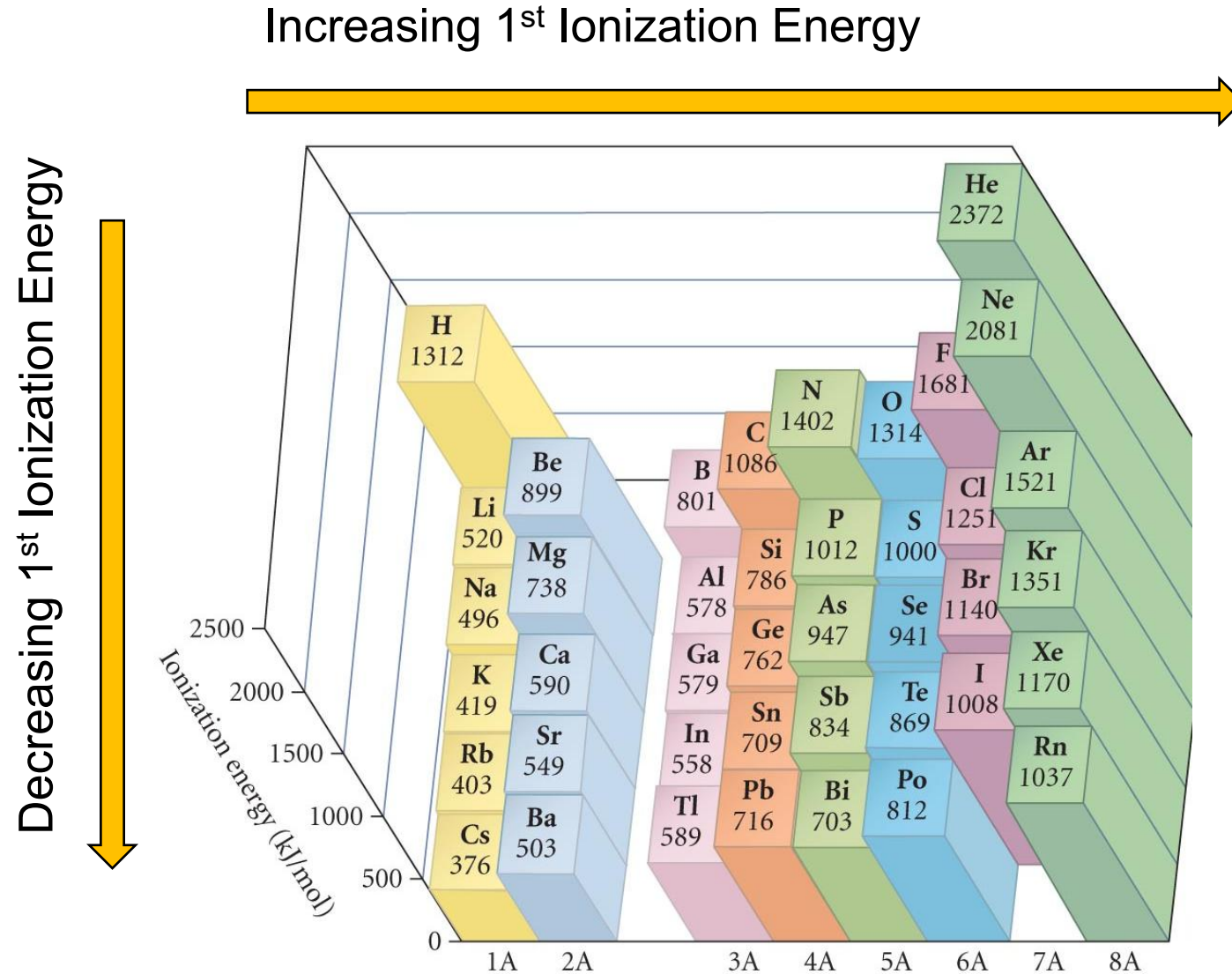


Ionization of core electrons causes a big jump in ionization energy!

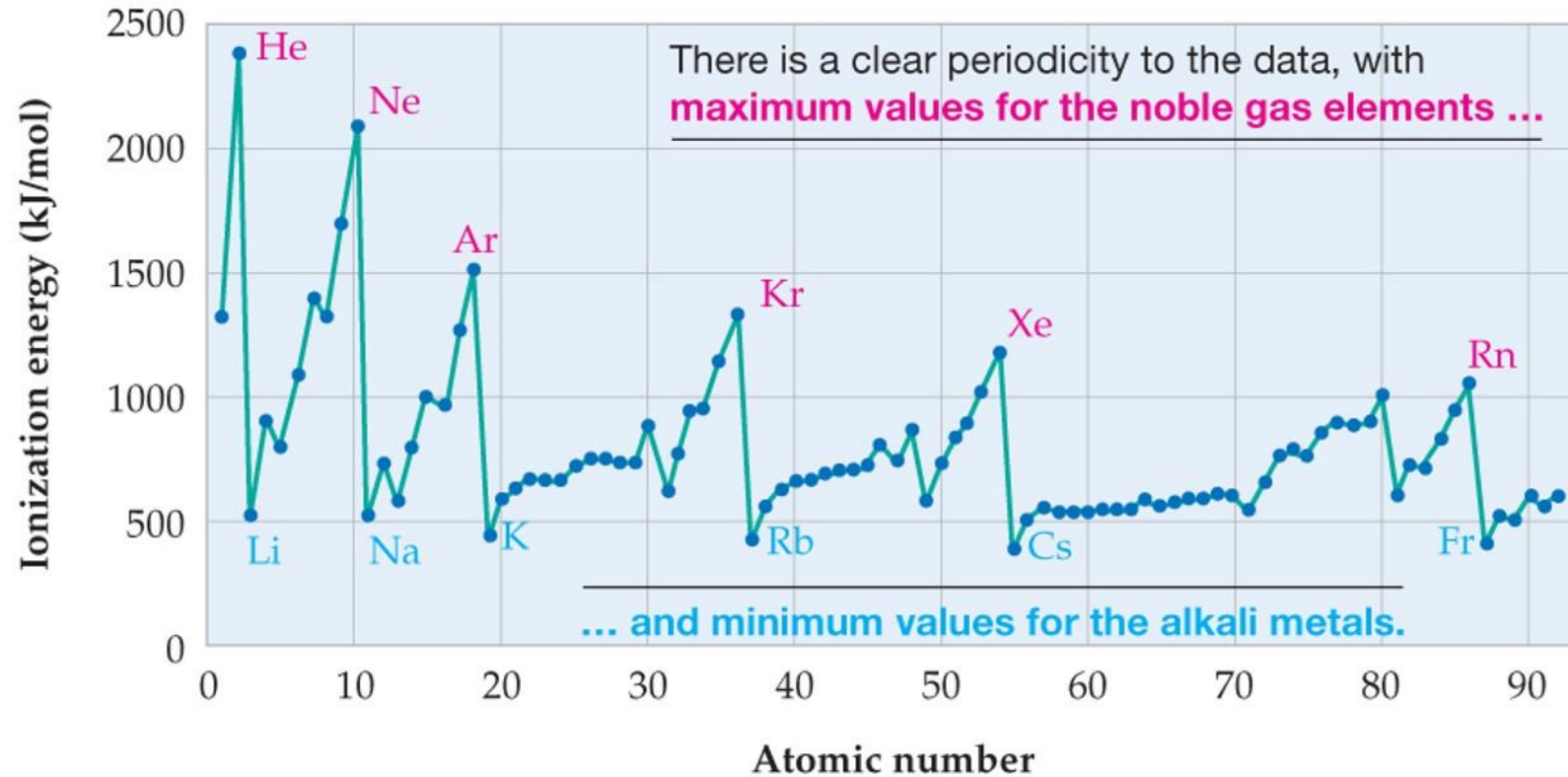
Ionization Energy and the Periodic Table



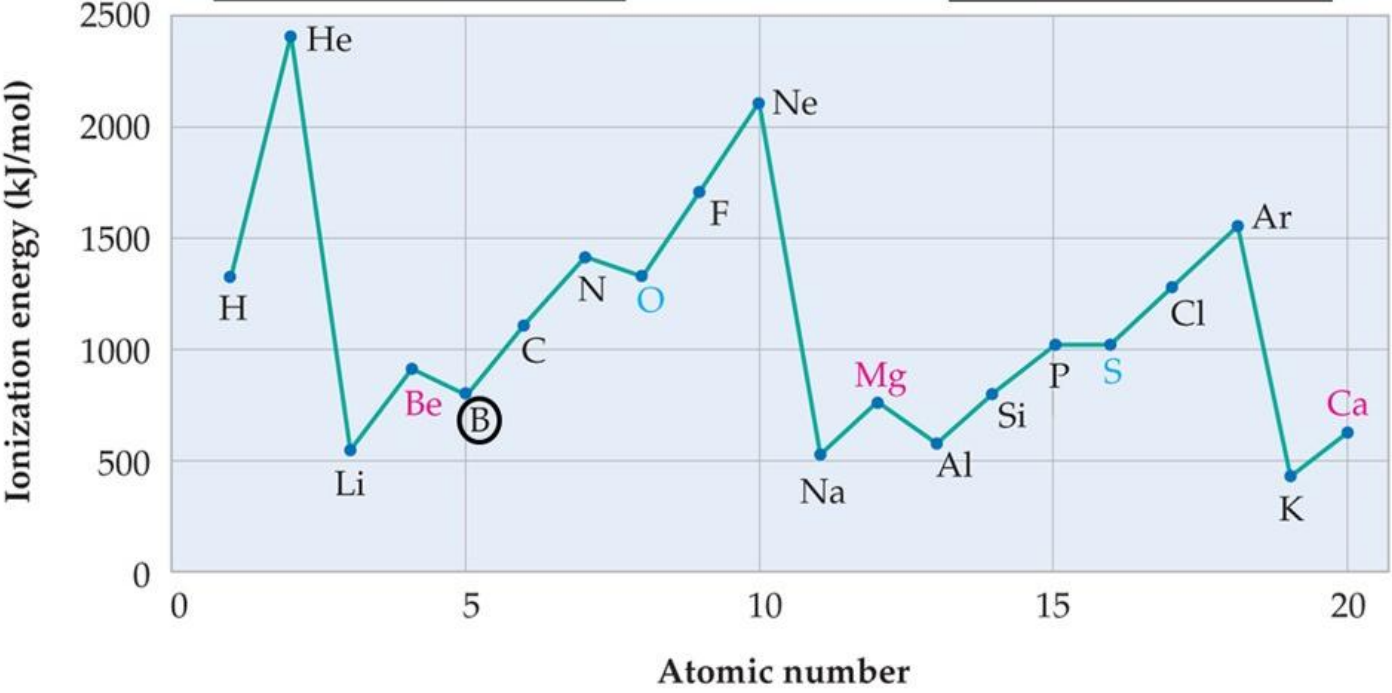
Ionization Energy and the Periodic Table



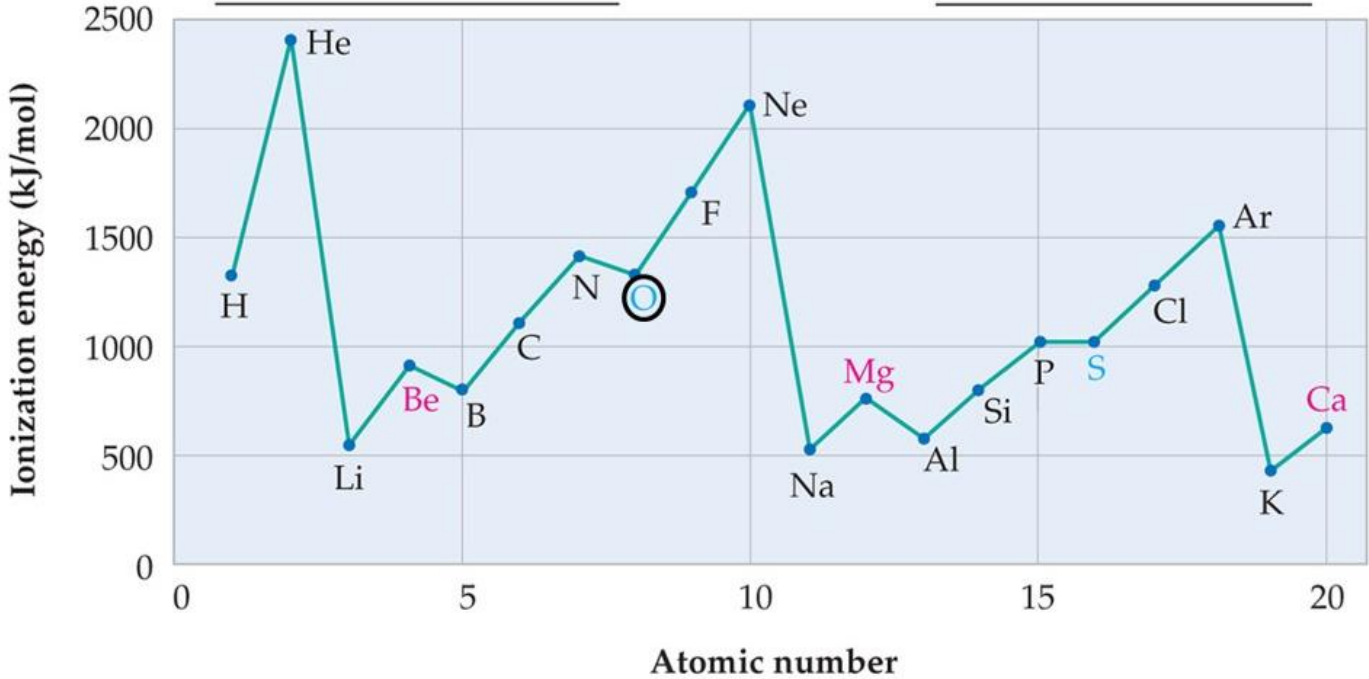
Ionization Energy and the Periodic Table



Ionization Energy and the Periodic Table



Ionization Energy and the Periodic Table



Ionization Energy - Summary

First ionization energy **increases** going from left to right along a period because of the increasing effective nuclear charge. The electrons are more strongly bound going from left to right.

First ionization energy **decreases** going down a group because electrons are farther from the nucleus. As n increases, the size of the orbital increases, and the electron is easier to remove.

Exercise

Which atom would require **more** energy to remove an electron? Why?

Na

Cl

Which atom would require **more** energy to remove an electron? Why?

Li

Cs

Exercise

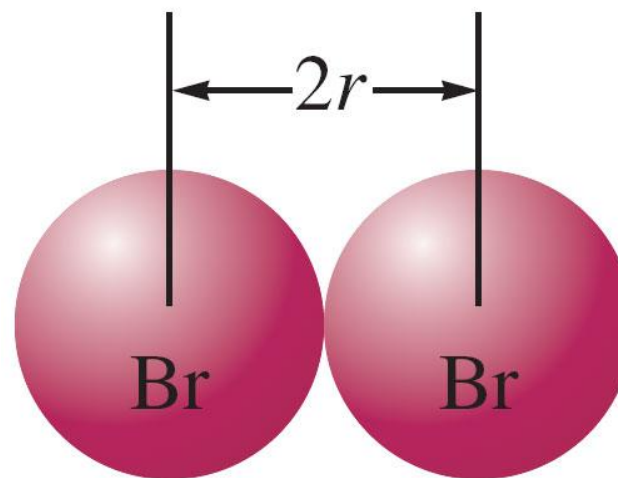
Which has the larger **second** ionization energy?
Why?

Lithium or Beryllium

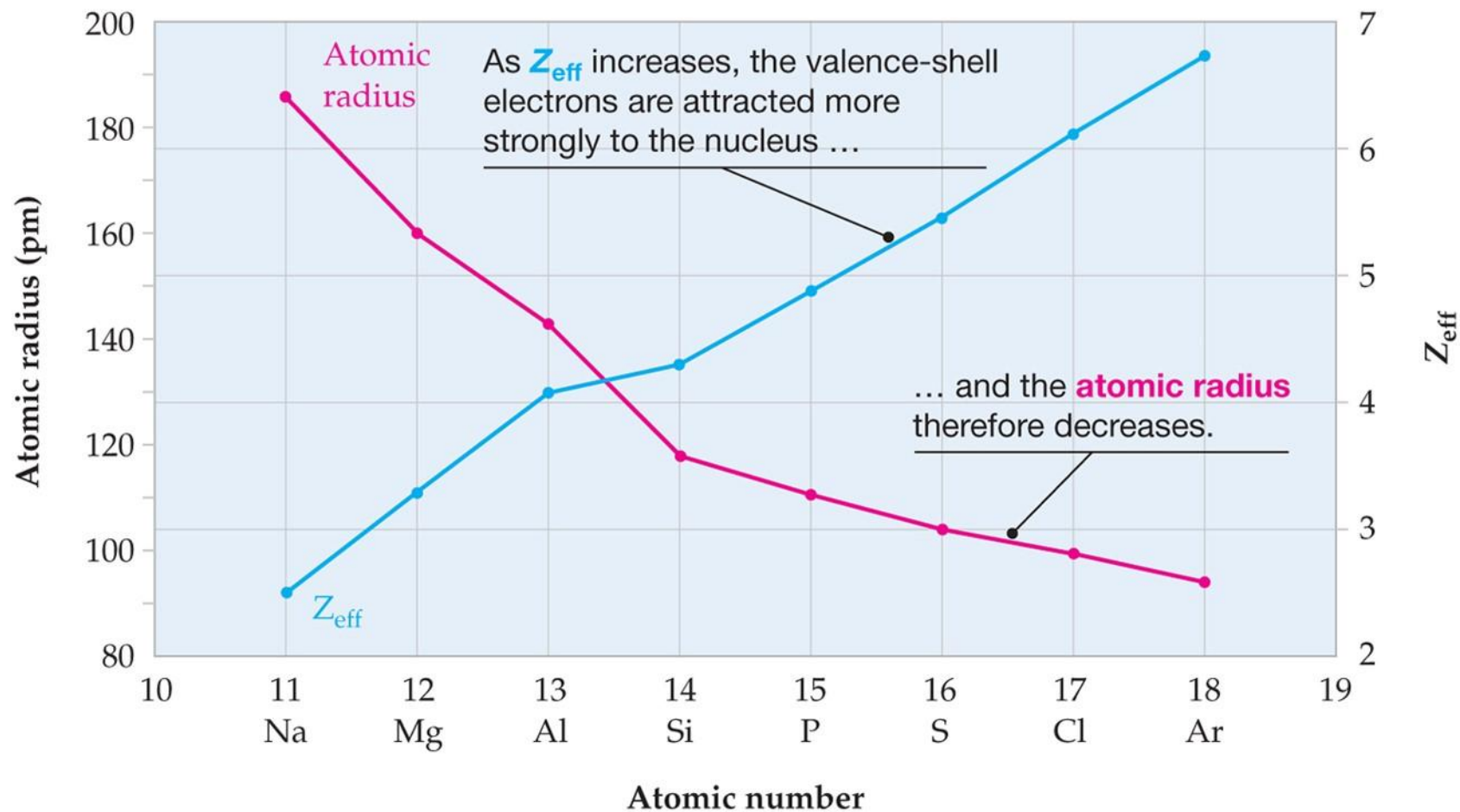
Atomic Radius

Atomic Radius

- The atomic radius in chemical compounds is obtained by measuring the distances between atoms.
- It is half the distance between the nuclei in a molecule consisting of identical atoms.



Electron Configurations and Periodic Properties: Atomic Radii



Atomic Radius and the Periodic Table

- The atomic radius **decreases** going from left to right across a period because of the increasing effective nuclear charge. The valence electrons are held more tightly, decreasing the size of the atom.
- Atomic radius **increases** down a group because of the increases in the orbital sizes.

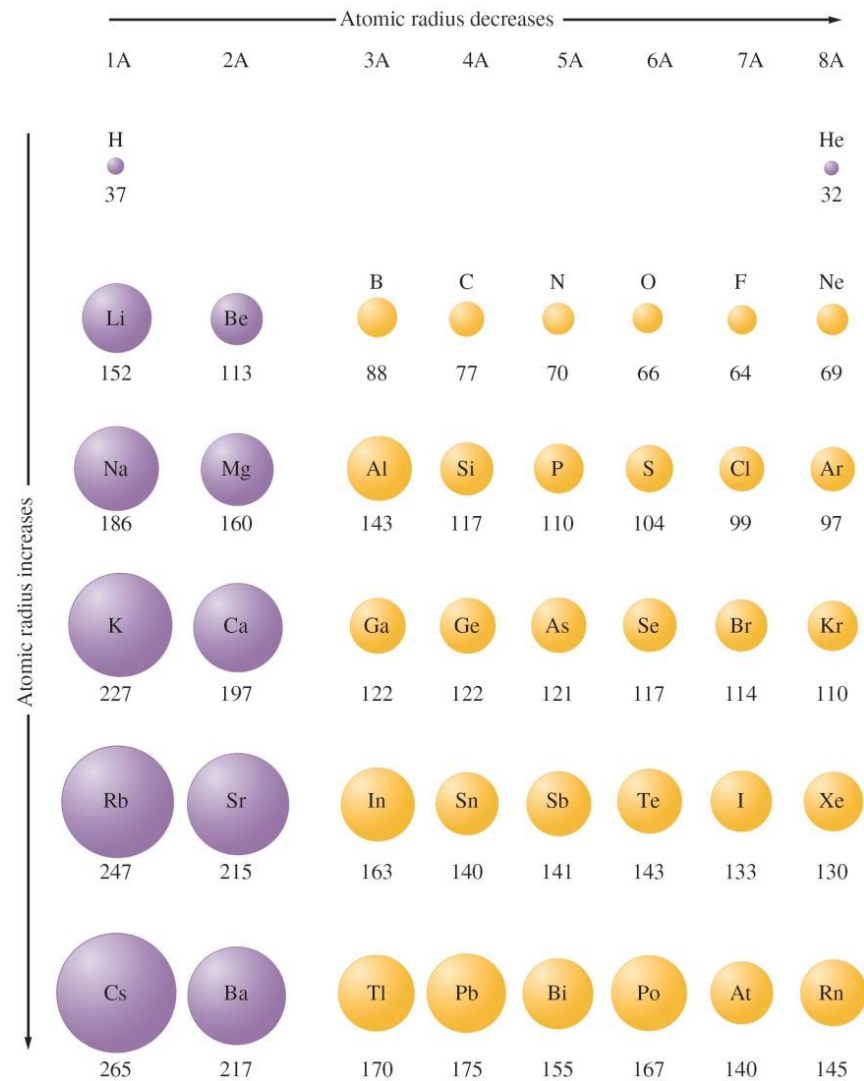
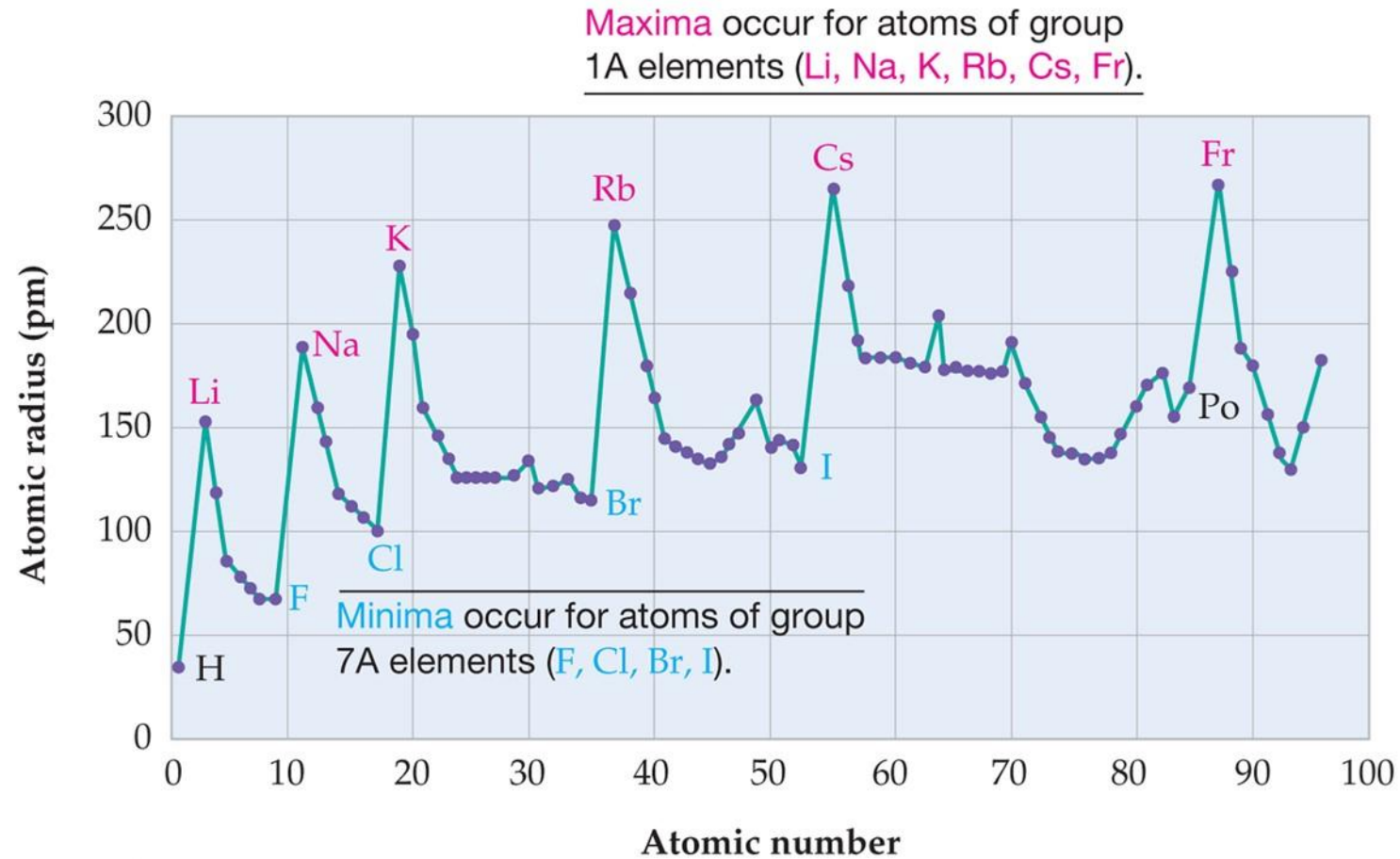


Figure 5.1



Exercise

Arrange the elements **oxygen**, **fluorine**, and **sulfur** according to increasing:

- Ionization energy
- Atomic size













Ionic Radius

Ionic Radius and the Periodic Table

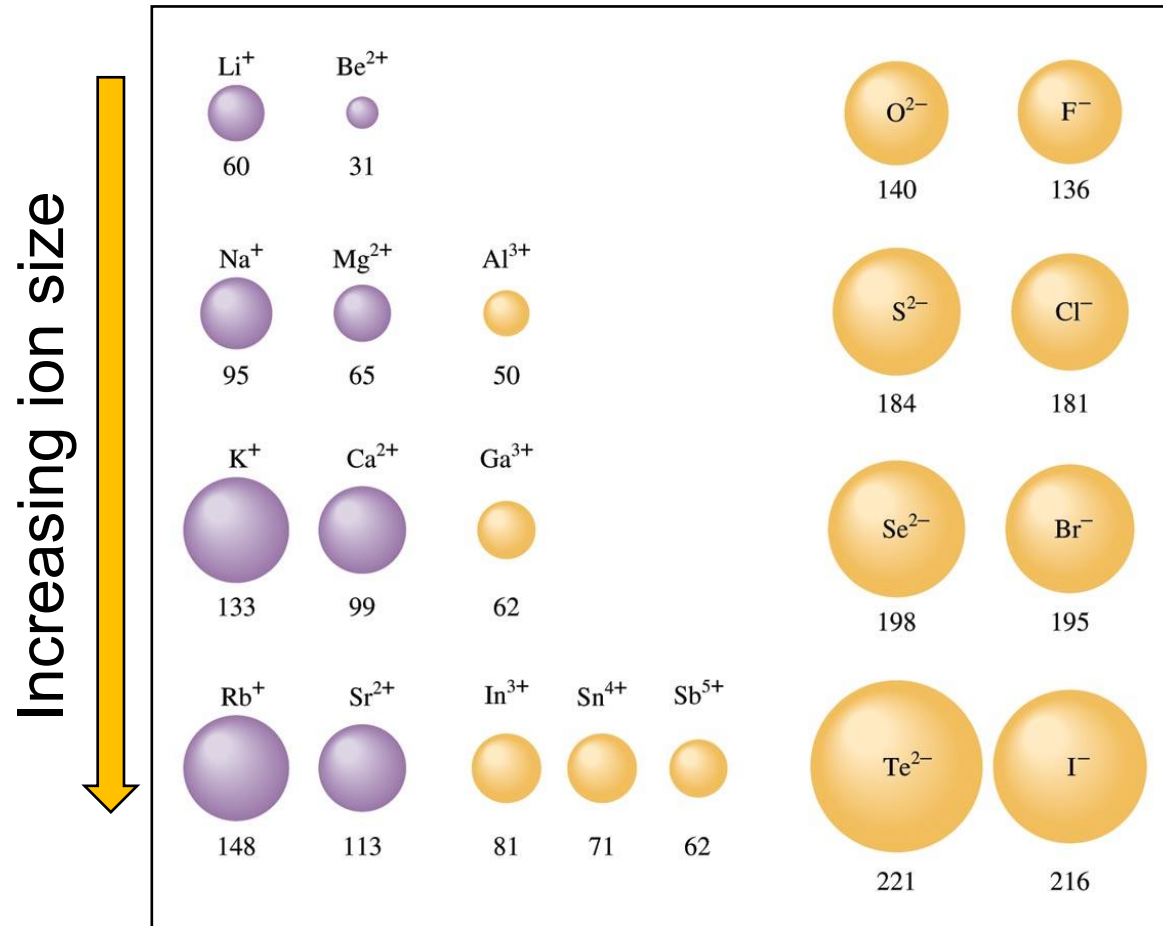
- When an electron is removed from an atom to form a cation, the atom shrinks; the radius of a cation is always smaller than that of the parent atom.
- The addition of electrons to a neutral atom produces an anion significantly larger than its parent atom.

Ionic Radii

Period 3

Atom	Na  186 pm	Mg  160 pm	Al  143 pm	P  110 pm	S  103 pm	Cl  99 pm
Ion	Na ⁺  102 pm	Mg ²⁺  72 pm	Al ³⁺  53 pm	P ³⁻  212 pm	S ²⁻  184 pm	Cl ⁻  181 pm

Ionic Radius and the Periodic Table

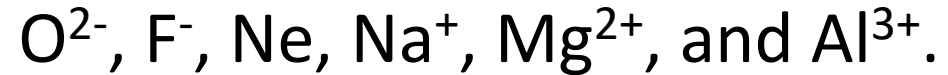


- The size of ions **increases** down a group.
- Change going across a period is complicated by the transition from metal (which lose electrons) to non-metal (which gain electrons).

Isoelectronic Series

- A series of ions/atoms containing the same number of electrons.

For example:



These ions/atom all have 10 electrons.

Exercise

Which one is the largest ion? S^{2-} Te^{2-} O^{2-} Se^{2-}

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

Exercise

Which one is the largest ion?

Se^{2-} Sr^{2+} Br^- Rb^+

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

Exercise

For the following isoelectronic series,



- What is the **electron configuration** for each species?
- Determine the **number of electrons** for each species.
- Determine the **number of protons** for each species.
- Rank the species according to **increasing radius**.
- Rank the species according to **increasing ionization energy**.

Electronegativity & Electron Affinity

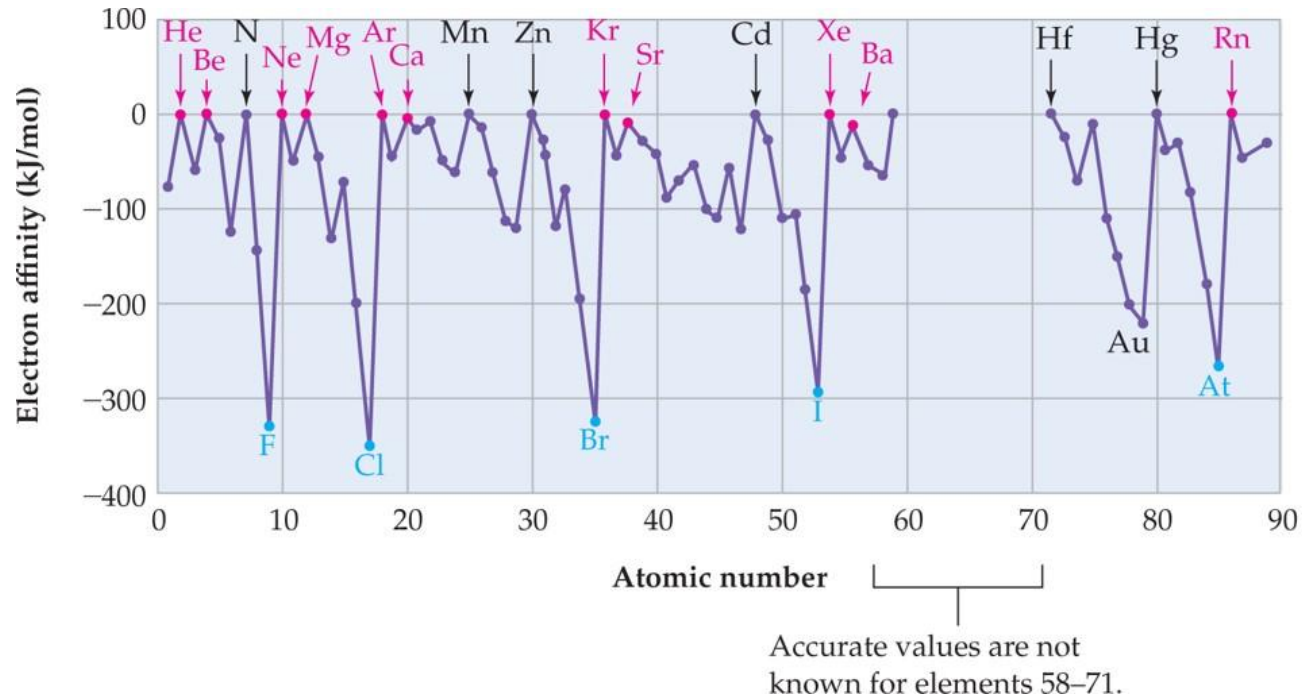
Electron Affinity

Electron affinity is the energy change associated with the addition of an electron to an atom in the gas phase.



The more negative the ΔE value, the more favorable it is to add an electron to an atom.

Electron Affinity



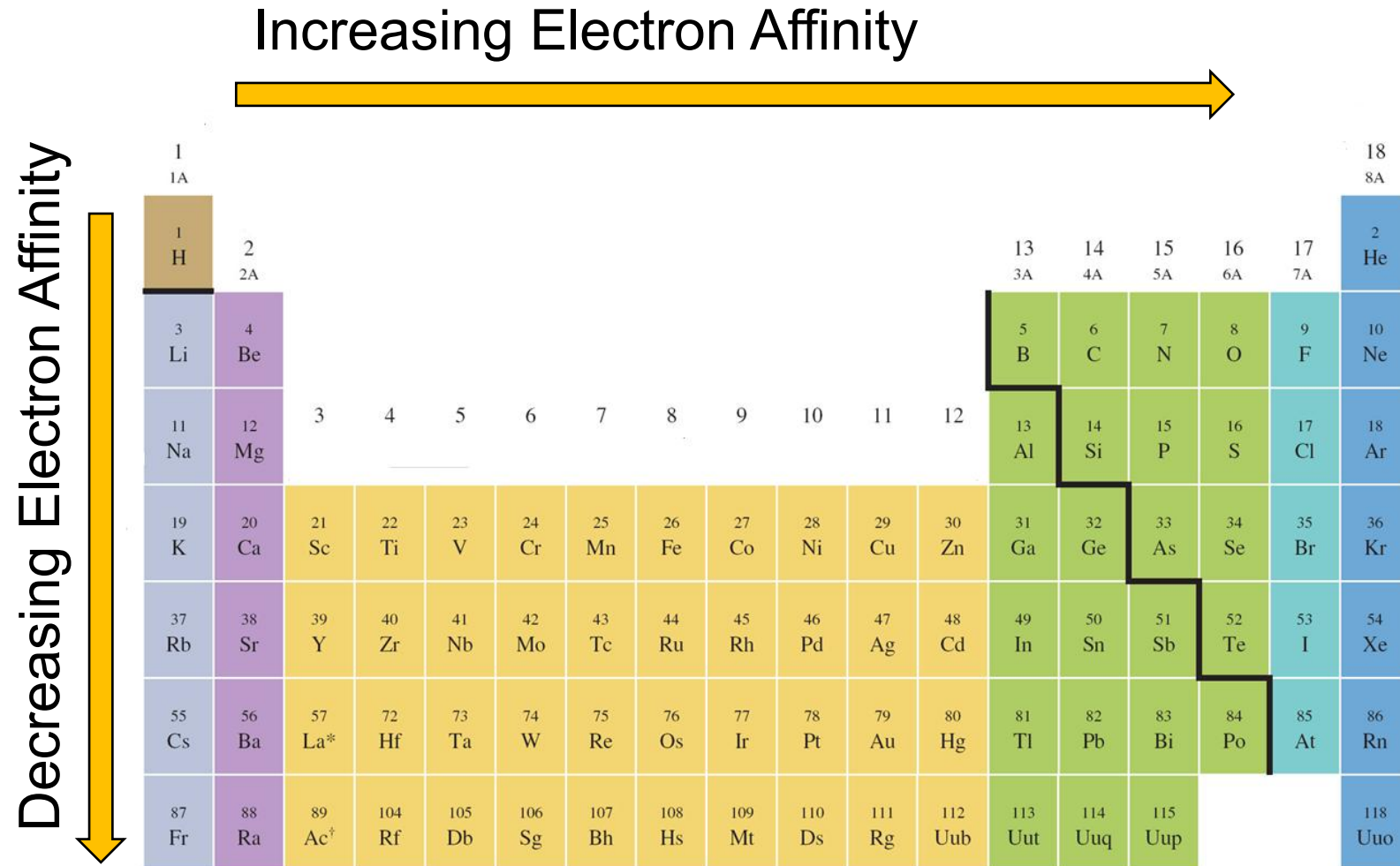
A negative value for E_{ea} , such as those for the **group 7A elements (halogens)**, means that energy is released when an electron adds to an atom.

A value of zero, such as those for the **group 2A elements (alkaline earths) and group 8A elements (noble gases)**, means that energy is absorbed but the exact amount can't be measured.

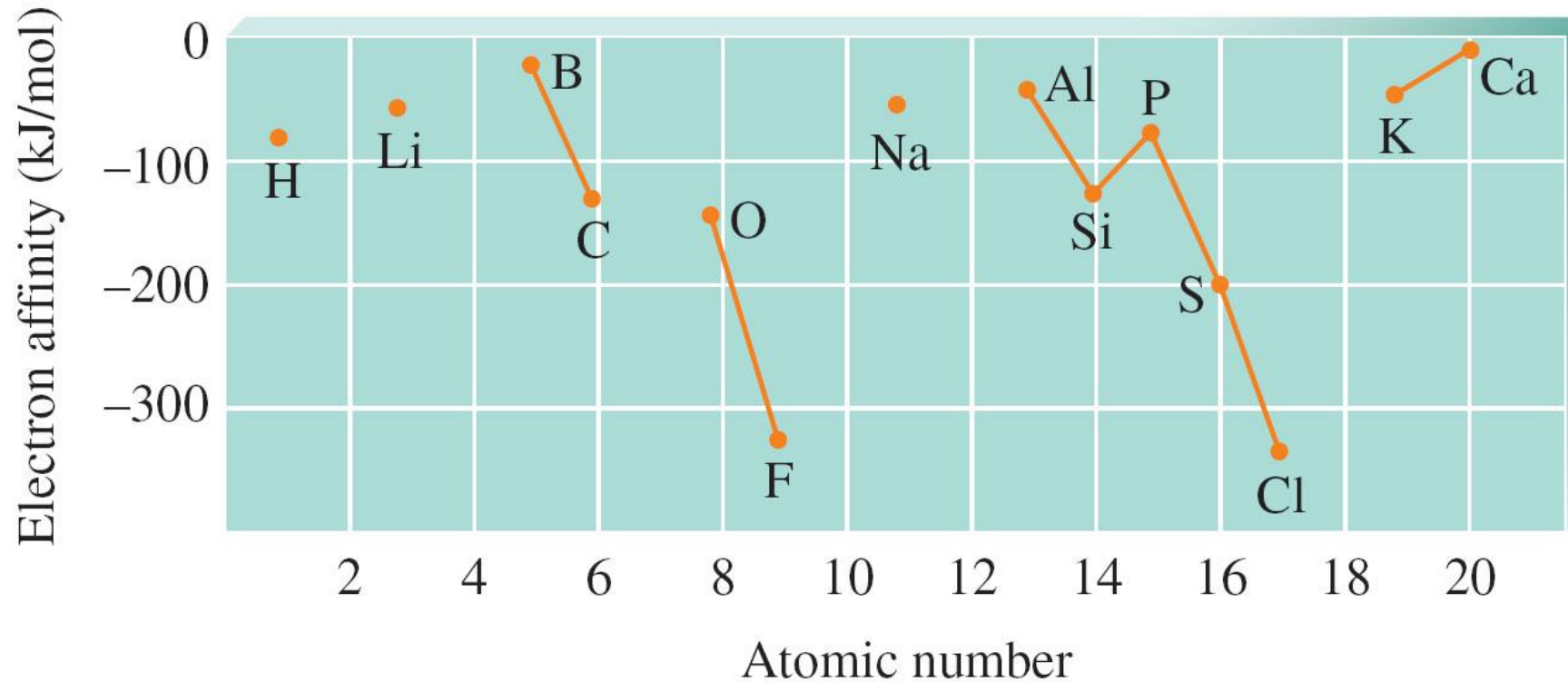
Electron Affinities (kJ/mol)

1A		2A	3A	4A	5A	6A	7A	8A
H -73								He >0
Li -60	Be >0		B -27	C -122	N >0	O -141	F -328	Ne >0
Na -53	Mg >0		Al -43	Si -134	P -72	S -200	Cl -349	Ar >0
K -48	Ca -2		Ga -30	Ge -119	As -78	Se -195	Br -325	Kr >0
Rb -47	Sr -5		In -30	Sn -107	Sb -103	Te -190	I -295	Xe >0

Electron Affinity and the Periodic Table



Electron Affinity



Shown are the electron affinity values for atoms among the first 20 elements that form stable, isolated X^- ions. The lines shown connect adjacent elements.

Electron Affinity

Electron affinity *tends* to **increase** going from left to right along a period because of the increasing effective nuclear charge.

Electron affinity *tends* to **decrease** going down a group because of the increasing orbital size. There is less attraction of an outer electron to the nucleus.

The Periodic Table Allows Us to Predict Many Properties

- Electron configurations
- Trends for:
 - Atomic size, ion radius, ionization energy, electronegativity, electron affinity
- Formula prediction for ionic compounds
- Covalent bond polarity ranking

Exercise

Given the orbital filling diagrams for the valence electrons of elements, rank them from lowest to highest ionization energy.

