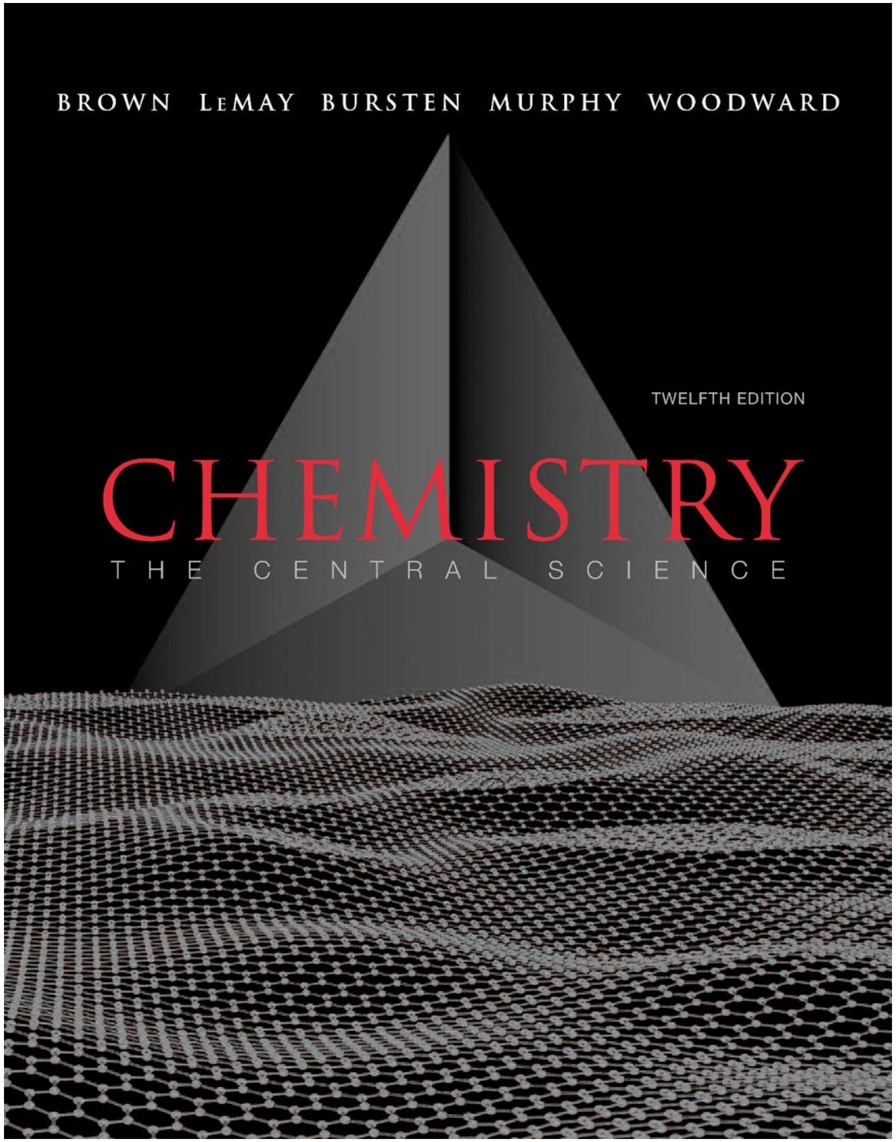

Lecture Presentation

Chapter 7

Periodic Properties of the Elements

John D. Bookstaver
St. Charles Community College
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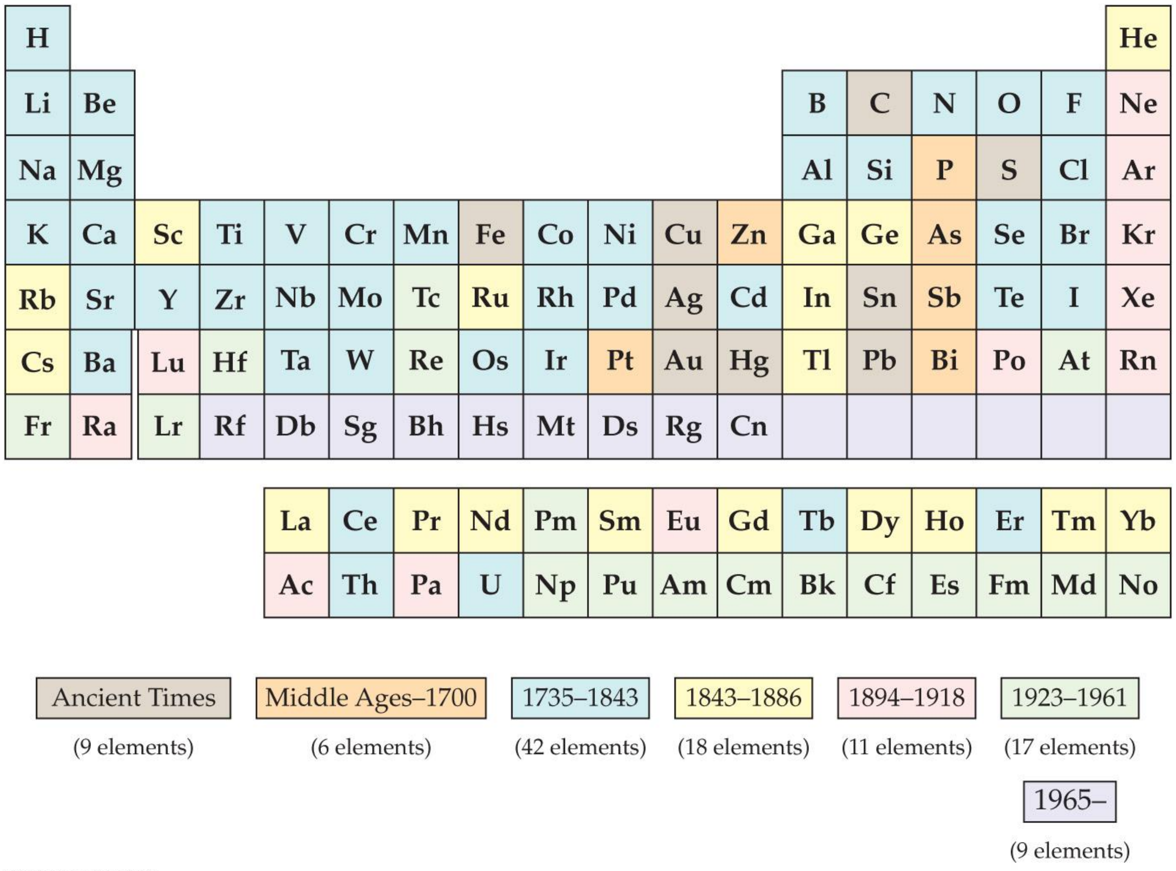
Development of Periodic Table



Dmitri Mendeleev and Lothar Meyer independently came to the same conclusion about how elements should be grouped.

Development of Periodic Table

TABLE 7.1 • Comparison of the Properties of Eka-Silicon Predicted by Mendeleev with the Observed Properties of Germanium

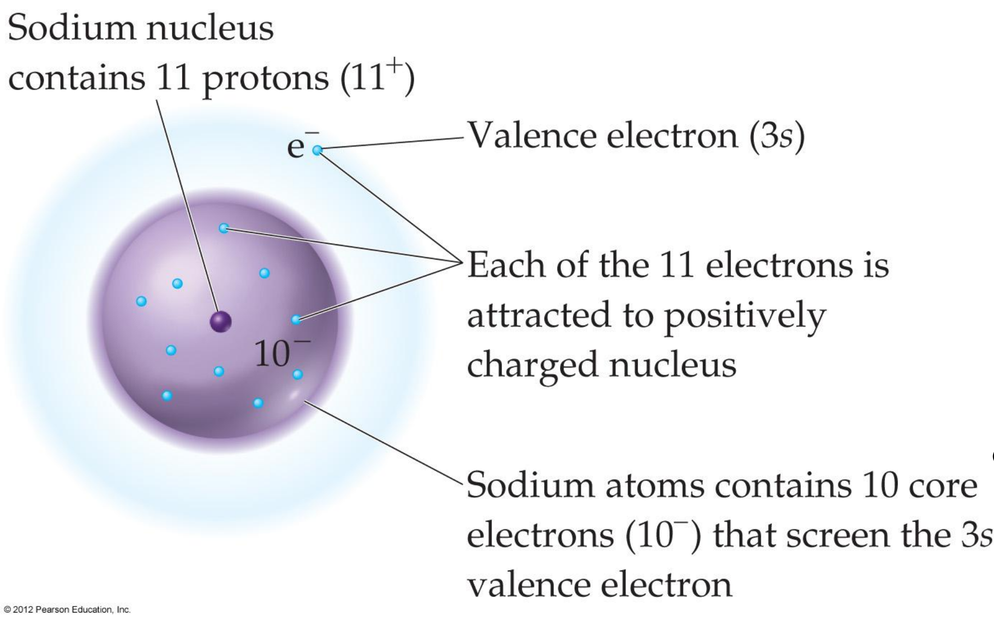
Property	Mendeleev's Predictions for Eka-Silicon (made in 1871)	Observed Properties of Germanium (discovered in 1886)
Atomic weight	72	72.59
Density (g/cm^3)	5.5	5.35
Specific heat (J/g-K)	0.305	0.309
Melting point (°C)	High	947
Color	Dark gray	Grayish white
Formula of oxide	XO_2	GeO_2
Density of oxide (g/cm^3)	4.7	4.70
Formula of chloride	XCl_4	$GeCl_4$
Boiling point of chloride (°C)	A little under 100	84

Mendeleev, for instance, predicted the discovery of germanium (which he called eka-silicon) as an element with an atomic weight between that of zinc and arsenic, but with chemical properties similar to those of silicon.

Periodic Trends

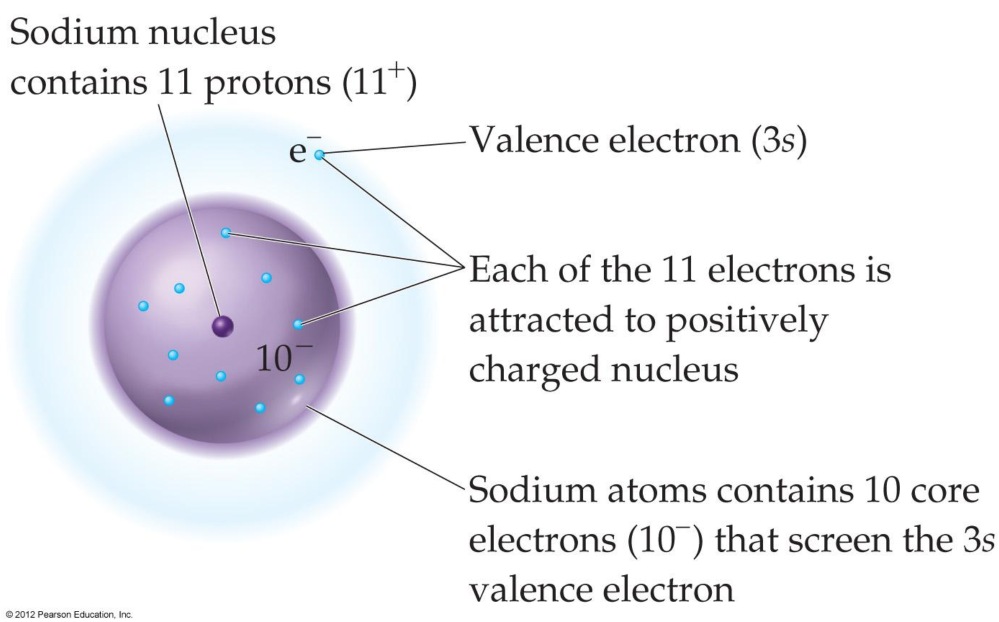
- In this chapter, we will rationalize observed trends in
  - Sizes of atoms and ions.
  - Ionization energy.
  - Electron affinity.

Effective Nuclear Charge



- In a many-electron atom, electrons are both attracted to the nucleus and repelled by other electrons.
- The nuclear charge that an electron experiences depends on both factors.

Effective Nuclear Charge



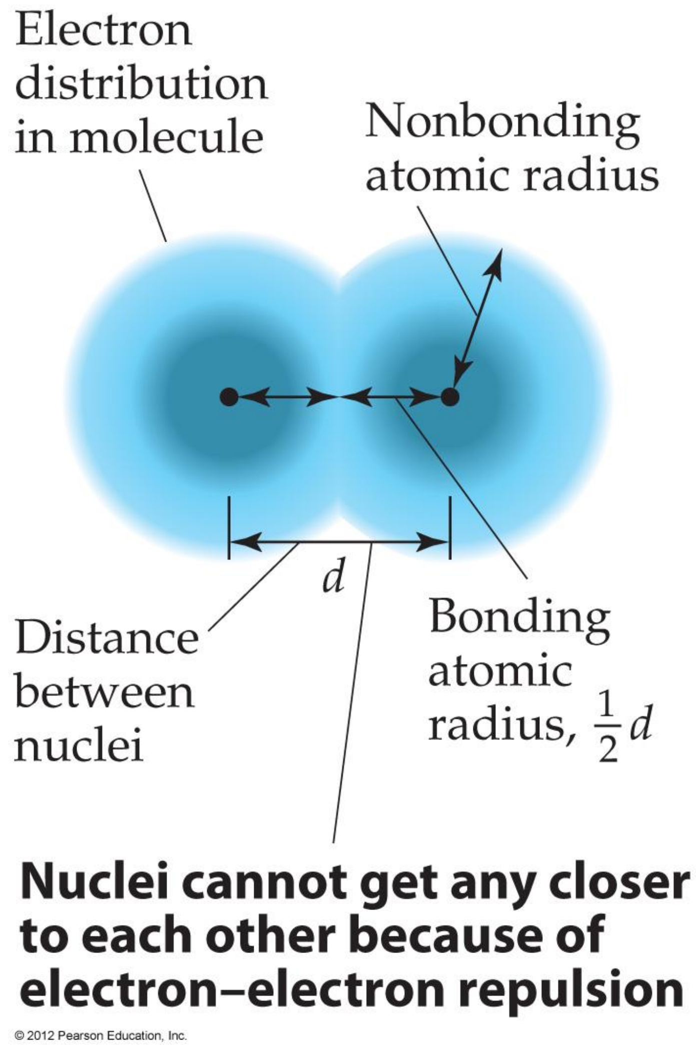
The effective nuclear charge, Z_{eff}, is found this way:

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

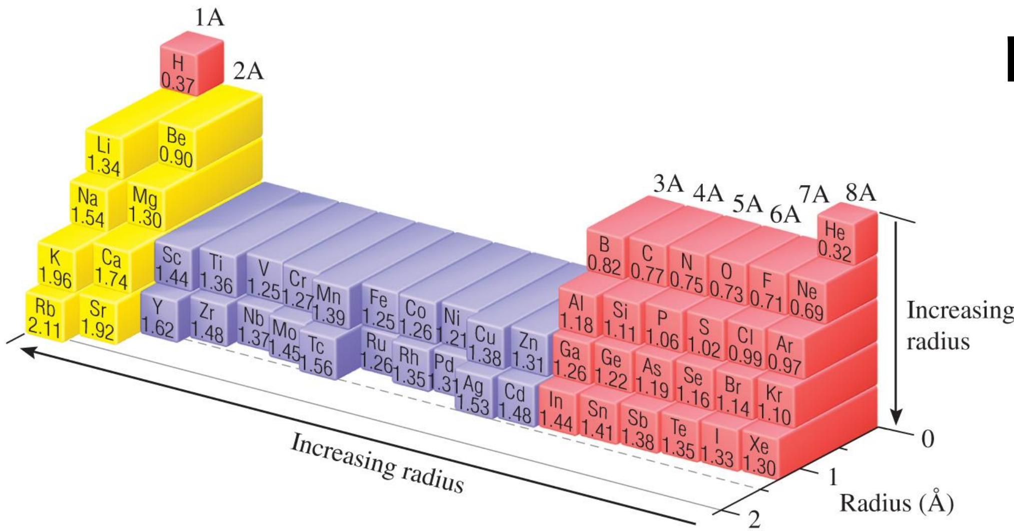
where Z is the atomic number and S is a screening constant, usually close to the number of inner electrons.

What Is the Size of an Atom?

The **bonding atomic radius** is defined as one-half of the distance between covalently bonded nuclei.



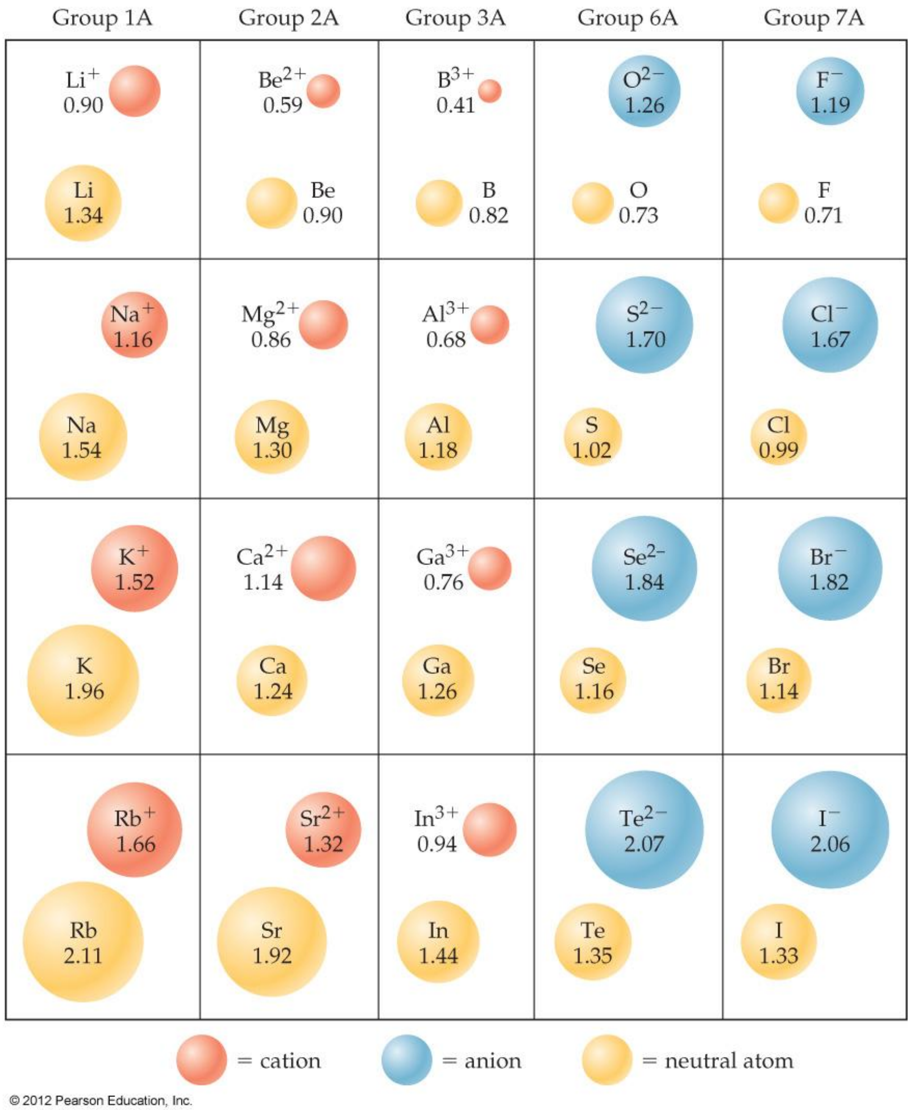
Sizes of Atoms



The bonding atomic radius tends to

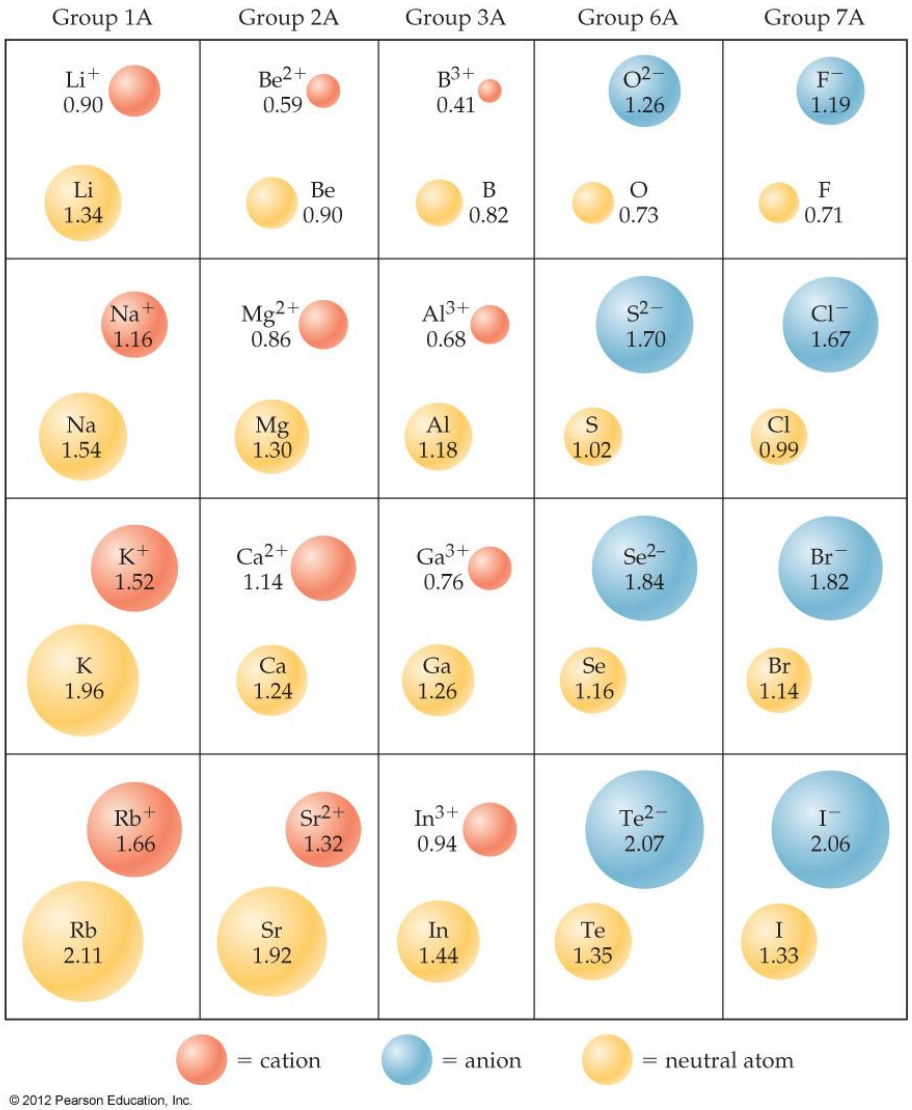
- Decrease from left to right across a row
  (due to increasing Z_{eff}).
- Increase from top to bottom of a column
  (due to the increasing value of n).

Sizes of Ions



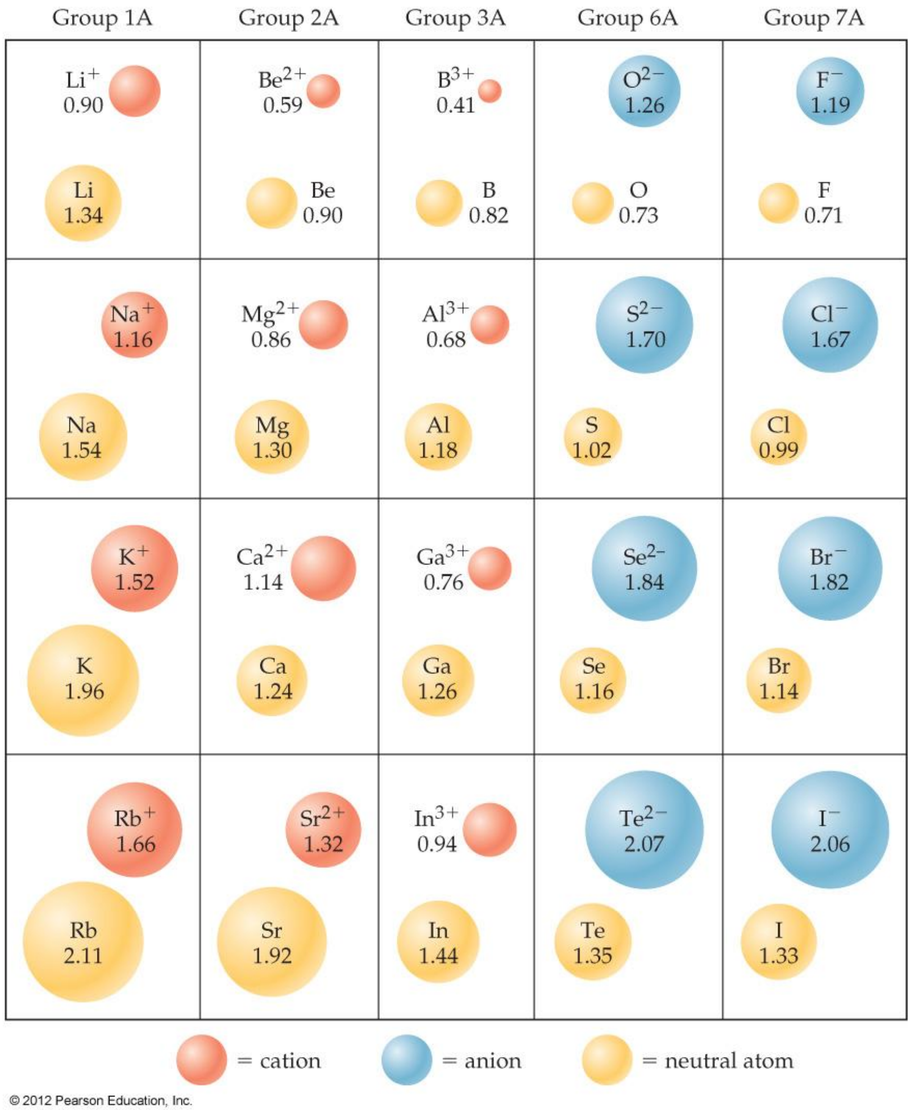
- Ionic size depends upon
  – The nuclear charge.
  – The number of electrons.
  – The orbitals in which electrons reside.

Sizes of Ions



- Cations are smaller than their parent atoms:
  - The outermost electron is removed and repulsions between electrons are reduced.

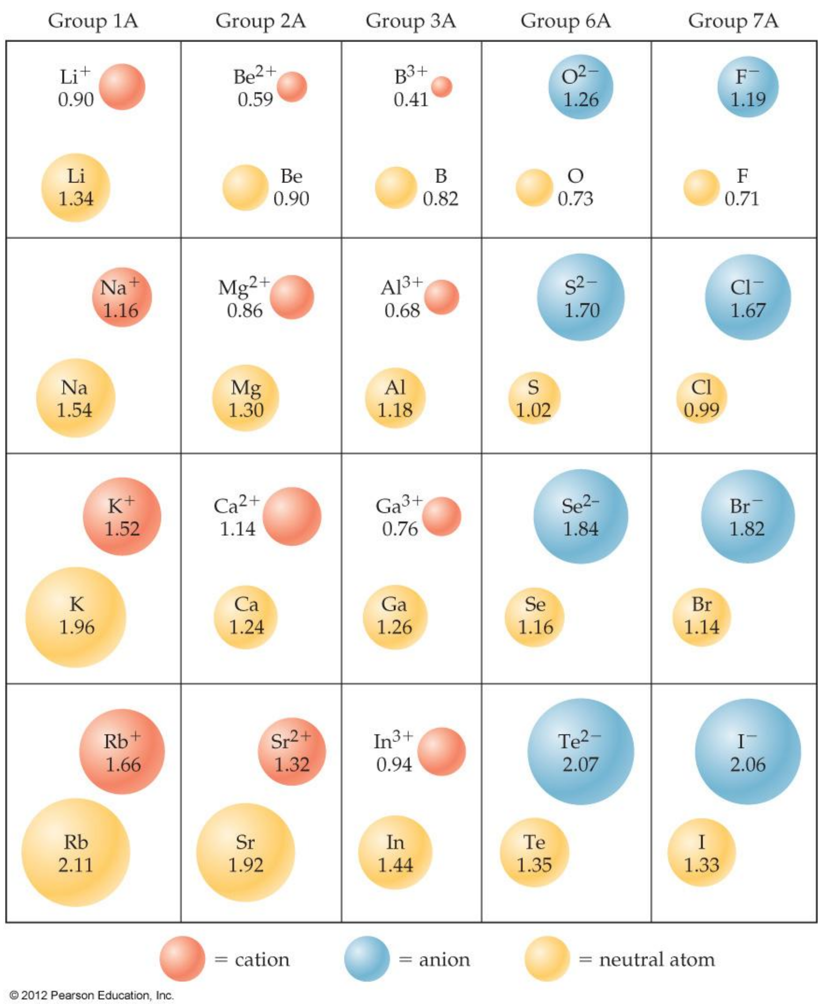
Sizes of Ions



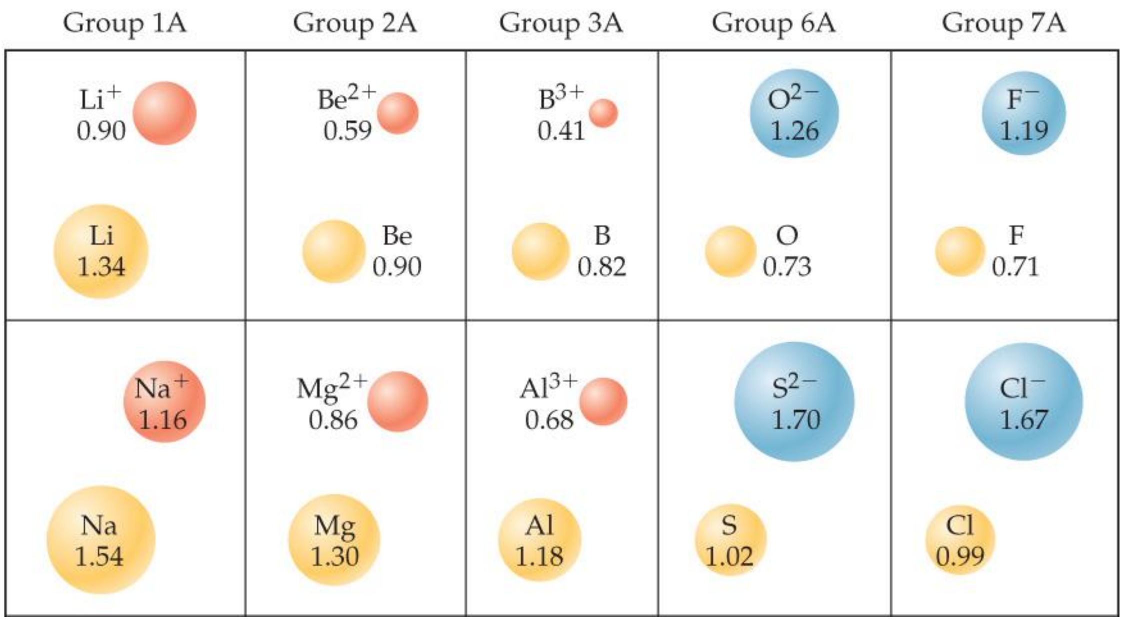
- Anions are larger than their parent atoms”
  - Electrons are added and repulsions between electrons are increased.

Sizes of Ions

- Ions increase in size as you go down a column:
  - This increase in size is due to the increasing value of *n*.



Sizes of Ions



- In an **isoelectronic series**, ions have the same number of electrons.
- Ionic size decreases with an increasing nuclear charge.

Ionization Energy

- The **ionization energy** is the amount of energy required to remove an electron from the ground state of a gaseous atom or ion.
  - The first ionization energy is that energy required to remove the first electron.
  - The second ionization energy is that energy required to remove the second electron, etc.

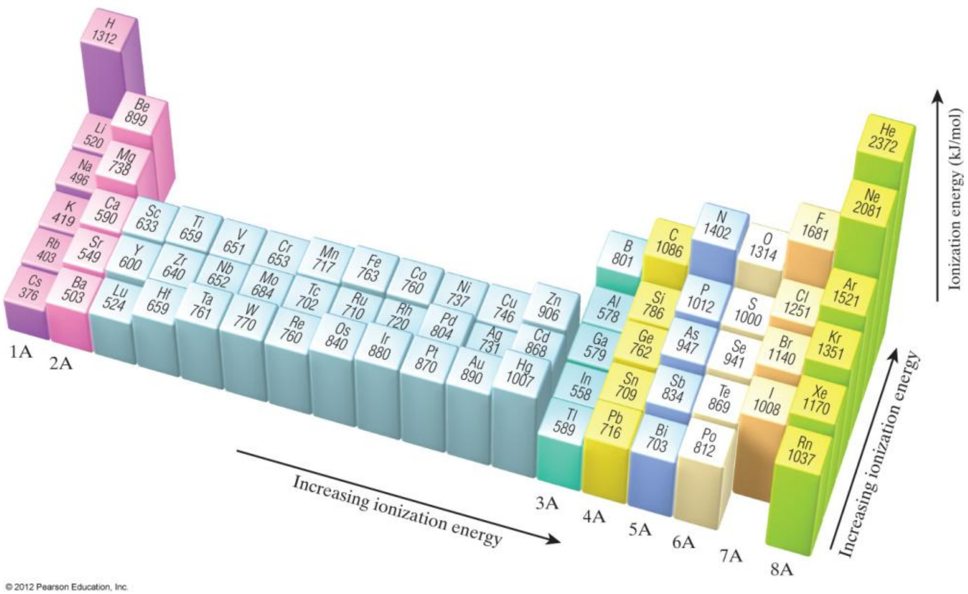
Ionization Energy

- It requires more energy to remove each successive electron.
- When all valence electrons have been removed, the ionization energy takes a quantum leap.

TABLE 7.2 • Successive Values of Ionization Energies, *I*, for the Elements Sodium through Argon (kJ/mol)

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4562			(inner-shell electrons)		
Mg	738	1451	7733				
Al	578	1817	2745	11,577			
Si	786	1577	3232	4356	16,091		
P	1012	1907	2914	4964	6274	21,267	
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018
Ar	1521	2666	3931	5771	7238	8781	11,995

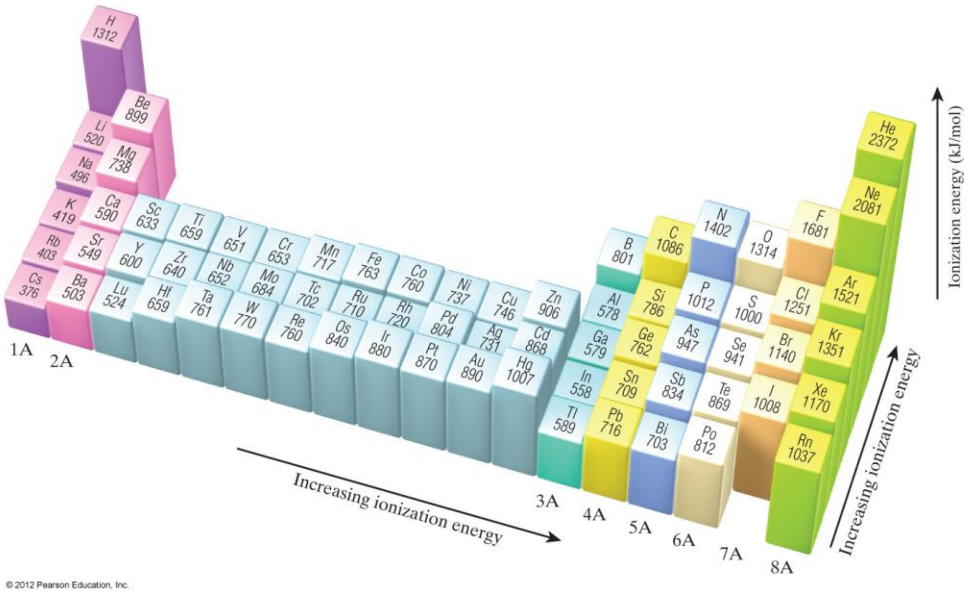
Trends in First Ionization Energies



- As one goes down a column, less energy is required to remove the first electron.
  - For atoms in the same group, Z_{eff} is essentially the same, but the valence electrons are farther from the nucleus.

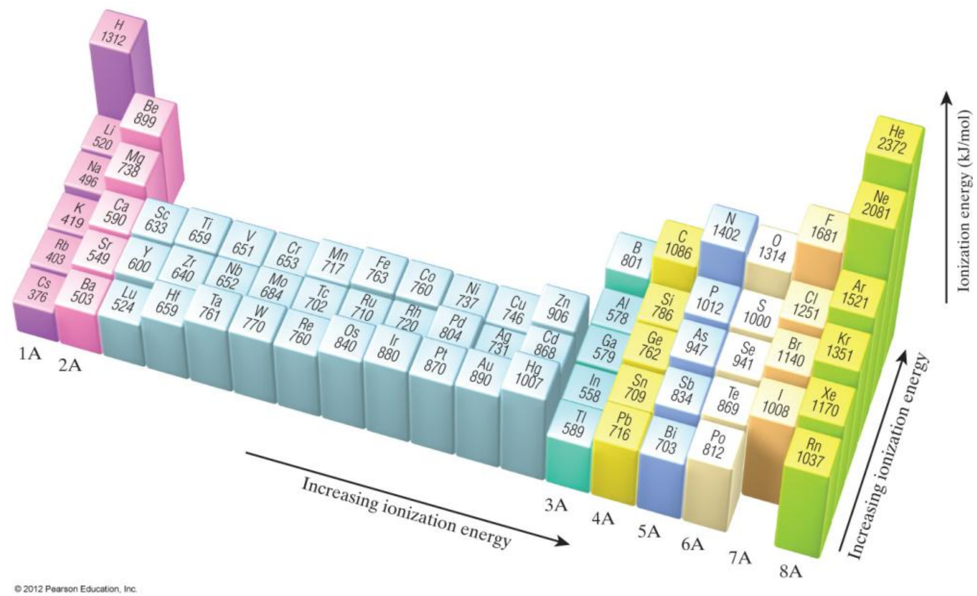
Trends in First Ionization Energies

- Generally, as one goes across a row, it gets harder to remove an electron.
  - As you go from left to right, Z_{eff} increases.



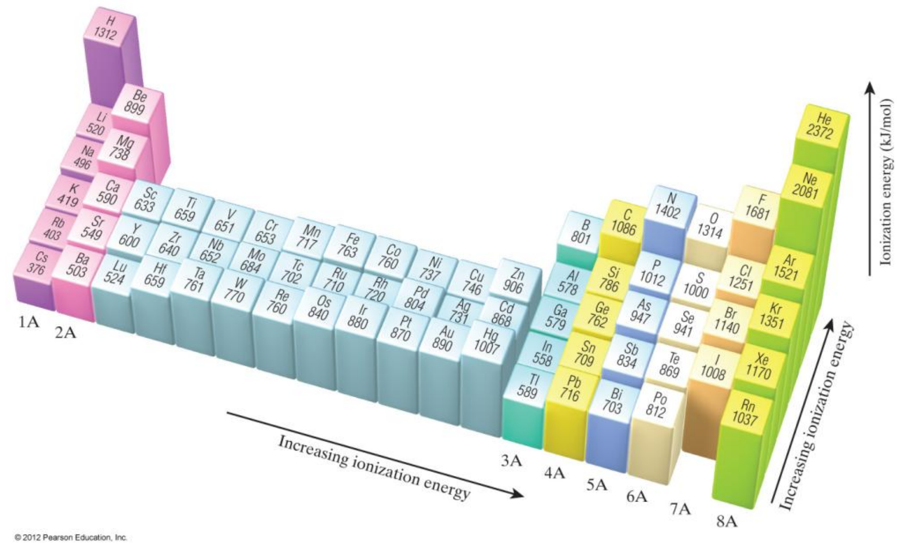
Trends in First Ionization Energies

However, there are two apparent discontinuities in this trend.



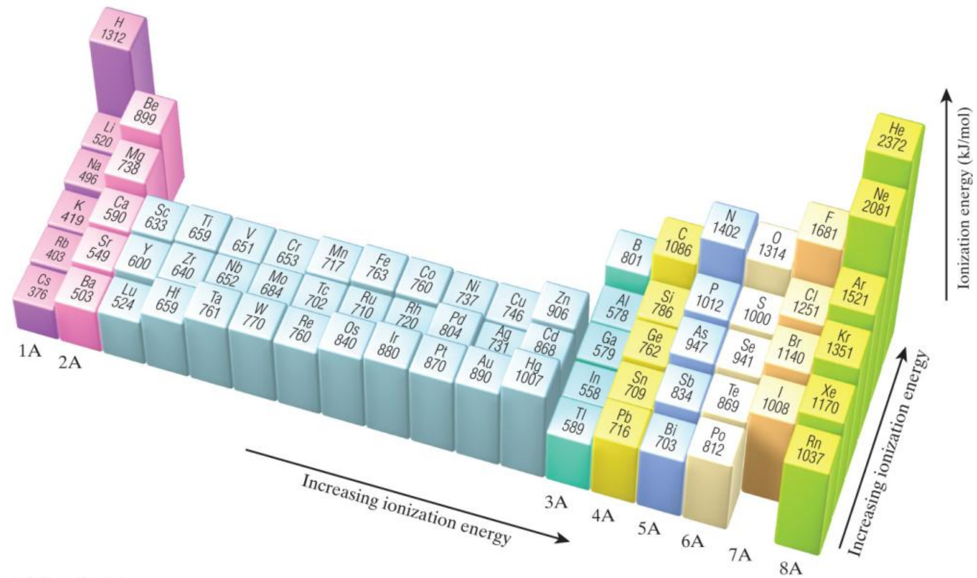
Trends in First Ionization Energies

- The first occurs between Groups IIA and IIIA.
- In this case the electron is removed from a *p* orbital rather than an *s* orbital.
  - The electron removed is farther from the nucleus.
  - There is also a small amount of repulsion by the *s* electrons.



Trends in First Ionization Energies

- The second discontinuity occurs between Groups VA and VIA.
  – The electron removed comes from a doubly occupied orbital.
  – Repulsion from the other electron in the orbital aids in its removal.



Electron Affinity

Electron affinity is the energy change accompanying the addition of an electron to a gaseous atom:

$$\mathrm{Cl} + e^{-} \longrightarrow \mathrm{Cl}^{-}$$

Trends in Electron Affinity

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

In general, electron affinity becomes more exothermic as you go from left to right across a row.

Trends in Electron Affinity

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

There are again, however, two discontinuities in this trend.

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Trends in Electron Affinity

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

The first occurs between Groups IA and IIA.

- The added electron must go in a *p* orbital, not an *s* orbital.
- The electron is farther from the nucleus and feels repulsion from the *s* electrons.

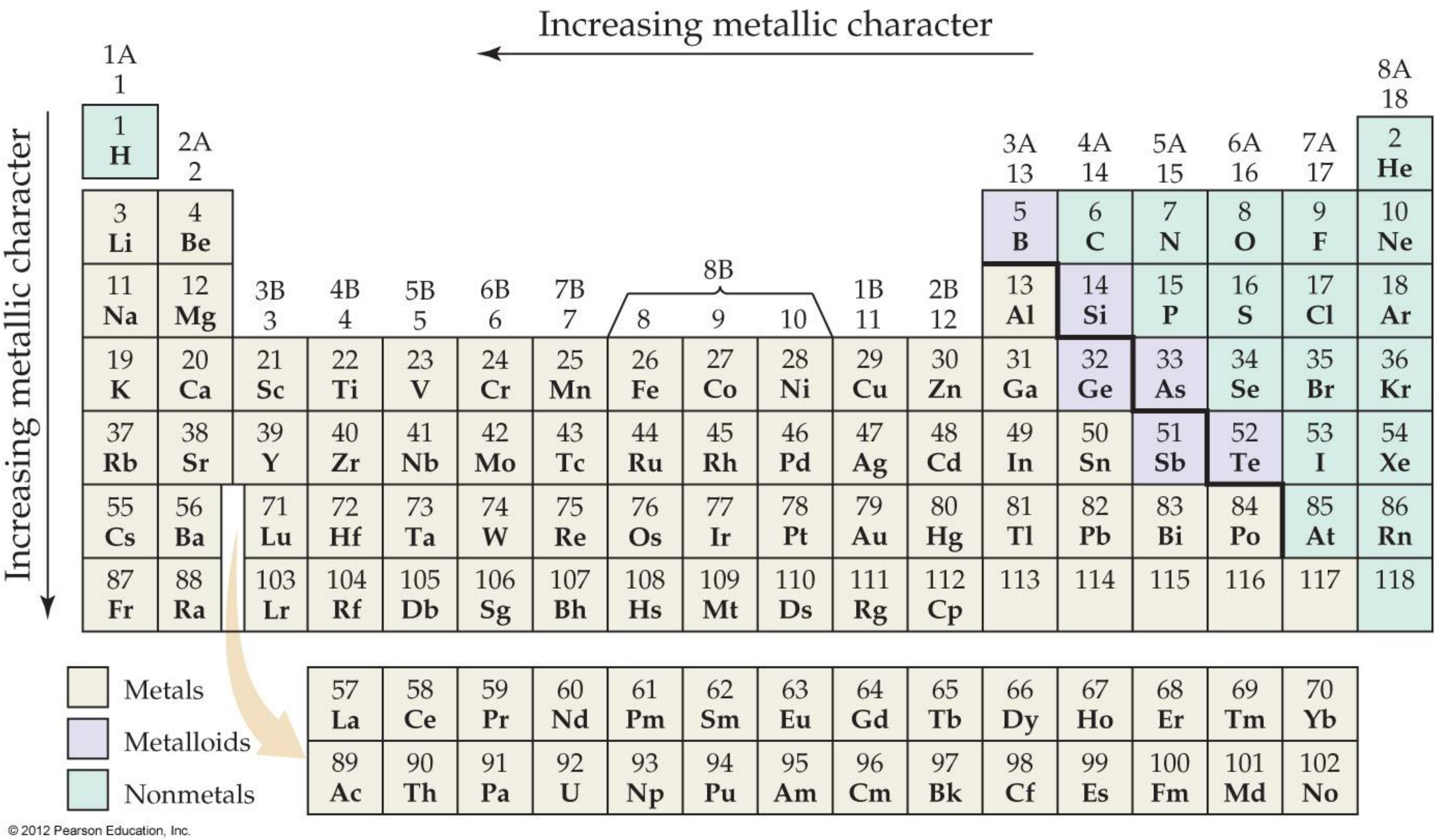
Trends in Electron Affinity

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

The second discontinuity occurs between Groups IVA and VA.

- Group VA has no empty orbitals.
- The extra electron must go into an already occupied orbital, creating repulsion.

Properties of Metal, Nonmetals,



Metals versus Nonmetals

TABLE 7.3 • Characteristic Properties of Metals and Nonmetals

Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal oxides are ionic solids that are basic	Most nonmetal oxides are molecular substances that form acidic solutions
Tend to form cations in aqueous solution	Tend to form anions or oxyanions in aqueous solution

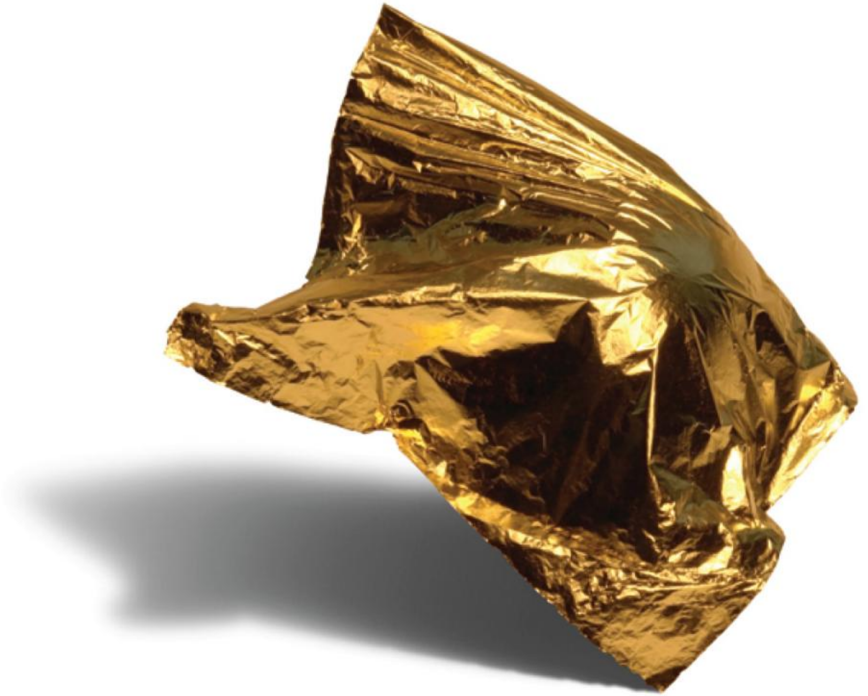
Differences between metals and nonmetals tend to revolve around these properties.

Metals versus Nonmetals

- Metals tend to form cations.
- Nonmetals tend to form anions.

1A	2A	Transition metals										3A	4A	5A	6A	7A	8A
H^+																H^-	NOBLE GASES
Li^+														N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}	Sc^{3+}	Ti^{4+}	V^{5+} V^{4+}	Cr^{3+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}								Pd^{2+}	Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}	Sb^{3+} Sb^{5+}	Te^{2-}	I^-	
Cs^+	Ba^{2+}								Pt^{2+}	Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}	Bi^{3+} Bi^{5+}			

Metals



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Metals tend to be lustrous, malleable, ductile, and good conductors of heat and electricity.

Metals

- Compounds formed between metals and nonmetals tend to be ionic.
- Metal oxides tend to be basic.

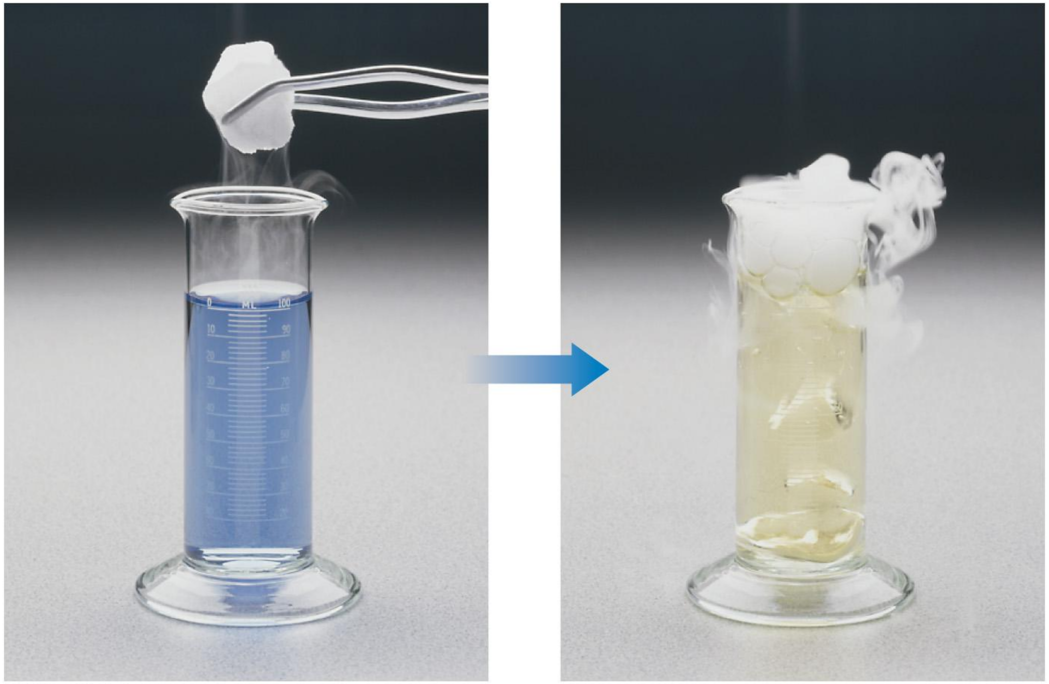
Nickle oxide (NiO), nitric acid (HNO_3), and water

NiO is insoluble in water but reacts with HNO_3 to give a green solution of the salt $Ni(NO_3)_2$

Nonmetals

- Nonmetals are dull, brittle substances that are poor conductors of heat and electricity.
- They tend to gain electrons in reactions with metals to acquire a noble-gas configuration.

Nonmetals



- Substances containing only nonmetals are molecular compounds.
- Most nonmetal oxides are acidic.

Metalloids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- Metalloids have some characteristics of metals and some of nonmetals.
- For instance, silicon looks shiny, but is brittle and a fairly poor conductor.

Group Trends

Alkali Metals

- **Alkali metals** are soft, metallic solids.
- The name comes from the Arabic word for ashes.

Alkali Metals

- They are found only in compounds in nature, not in their elemental forms.
- They have low densities and melting points.
- They also have low ionization energies.

TABLE 7.4 • Some Properties of the Alkali Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Lithium	$[He]2s^1$	181	0.53	1.34	520
Sodium	$[Ne]3s^1$	98	0.97	1.54	496
Potassium	$[Ar]4s^1$	63	0.86	1.96	419
Rubidium	$[Kr]5s^1$	39	1.53	2.11	403
Cesium	$[Xe]6s^1$	28	1.88	2.25	376

Alkali Metals

Li Na K

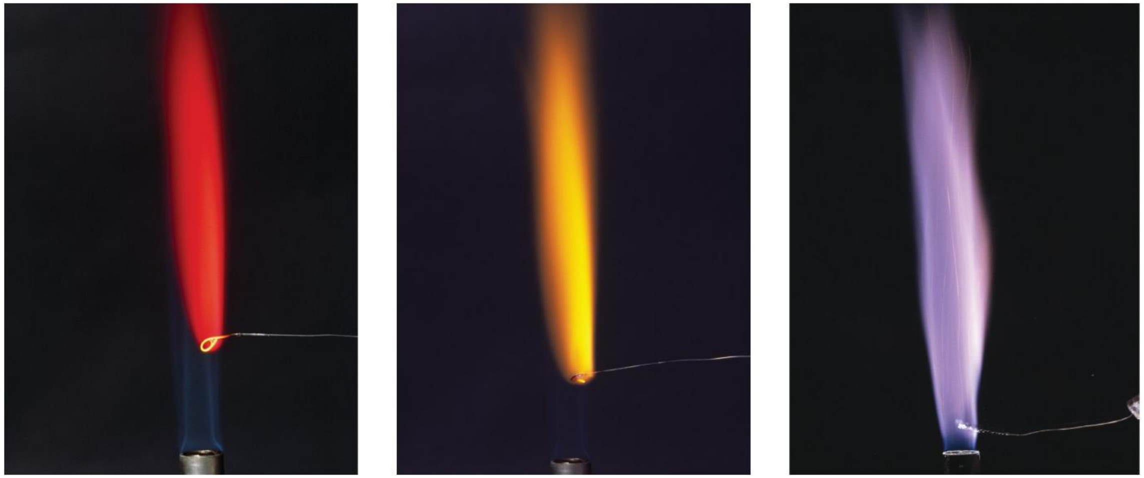
Their reactions with water are famously exothermic.

Alkali Metals

- Alkali metals (except Li) react with oxygen to form peroxides.
- K, Rb, and Cs also form superoxides:

$$K + O_2 \longrightarrow KO_2$$

- They produce bright colors when placed in a flame.



Li Na K

Alkaline Earth Metals

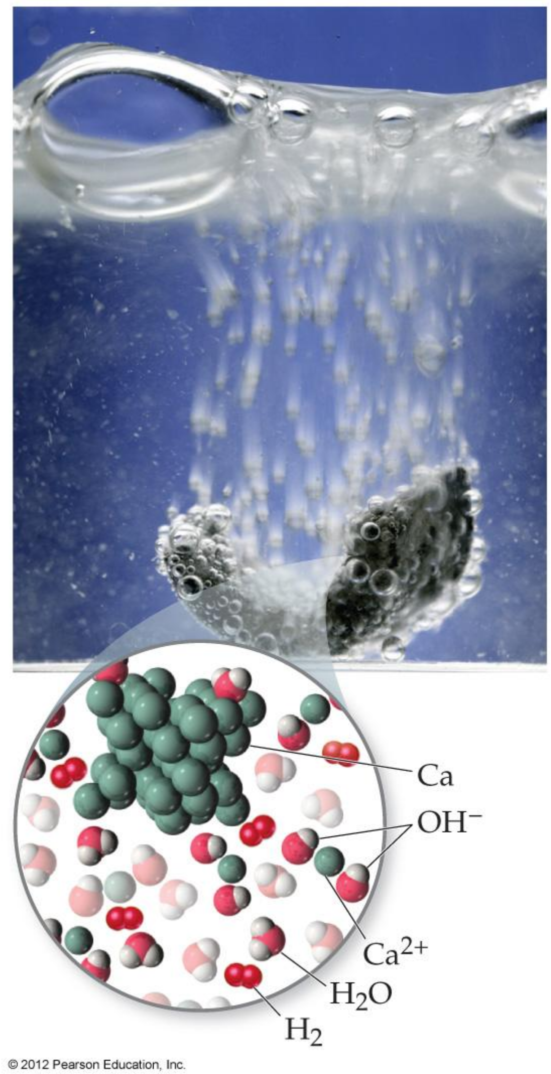
TABLE 7.5 • Some Properties of the Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Beryllium	$[He]2s^2$	1287	1.85	0.90	899
Magnesium	$[Ne]3s^2$	650	1.74	1.30	738
Calcium	$[Ar]4s^2$	842	1.55	1.74	590
Strontium	$[Kr]5s^2$	777	2.63	1.92	549
Barium	$[Xe]6s^2$	727	3.51	1.98	503

- Alkaline earth metals have higher densities and melting points than alkali metals.
- Their ionization energies are low, but not as low as those of alkali metals.

Alkaline Earth Metals

- Beryllium does not react with water, and magnesium reacts only with steam, but the other alkaline earth metals react readily with water.
- Reactivity tends to increase as you go down the group.



Group 6A

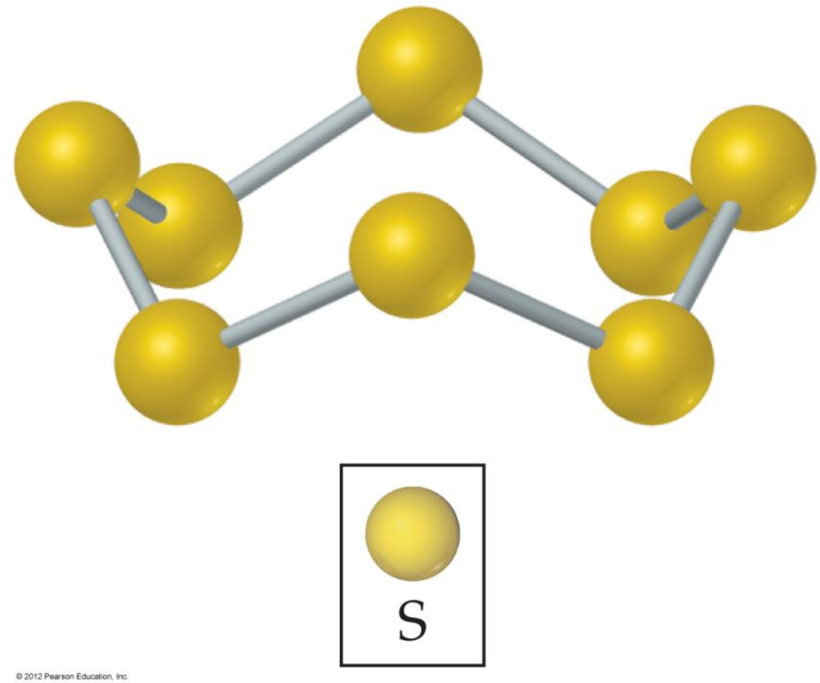
TABLE 7.6 • Some Properties of the Group 6A Elements

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Oxygen	$[He]2s^22p^4$	−218	1.43 g/L	0.73	1314
Sulfur	$[Ne]3s^23p^4$	115	1.96 g/cm^3	1.02	1000
Selenium	$[Ar]3d^{10}4s^24p^4$	221	4.82 g/cm^3	1.16	941
Tellurium	$[Kr]4d^{10}5s^25p^4$	450	6.24 g/cm^3	1.35	869
Polonium	$[Xe]4f^{14}5d^{10}6s^26p^4$	254	9.20 g/cm^3	—	812

- Oxygen, sulfur, and selenium are nonmetals.
- Tellurium is a metalloid.
- The radioactive polonium is a metal.

Sulfur

- Sulfur is a weaker oxidizer than oxygen.
- The most stable allotrope is S_8, a ringed molecule.



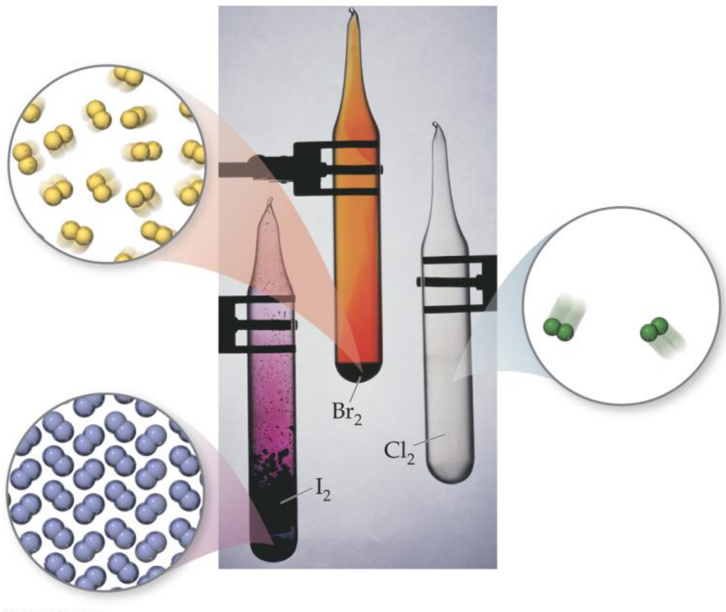
Group VIIA: Halogens

TABLE 7.7 • Some Properties of the Halogens

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Fluorine	$[He]2s^22p^5$	−220	1.69 g/L	0.71	1681
Chlorine	$[Ne]3s^23p^5$	−102	3.12 g/L	0.99	1251
Bromine	$[Ar]3d^{10}4s^24p^5$	−7.3	3.12 g/cm^3	1.14	1140
Iodine	$[Kr]4d^{10}5s^25p^5$	114	4.94 g/cm^3	1.33	1008

- The **halogens** are prototypical nonmetals.
- The name comes from the Greek words *halos* and *gennao*: "salt formers."

Group VIIA: Halogens



- They have large, negative electron affinities.
  - Therefore, they tend to oxidize other elements easily.
- They react directly with metals to form metal halides.
- Chlorine is added to water supplies to serve as a disinfectant.

Group VIIIA: Noble Gases

TABLE 7.8 • Some Properties of the Noble Gases

Element	Electron Configuration	Boiling Point (K)	Density (g/L)	Atomic Radius* (Å)	I_1 (kJ/mol)
Helium	$1s^2$	4.2	0.18	0.32	2372
Neon	$[He]2s^22p^6$	27.1	0.90	0.69	2081
Argon	$[Ne]3s^23p^6$	87.3	1.78	0.97	1521
Krypton	$[Ar]3d^{10}4s^24p^6$	120	3.75	1.10	1351
Xenon	$[Kr]4d^{10}5s^25p^6$	165	5.90	1.30	1170
Radon	$[Xe]4f^{14}5d^{10}6s^26p^6$	211	9.73	1.45	1037

*Only the heaviest of the noble-gas elements form chemical compounds. Thus, the atomic radii for the lighter noble-gas elements are estimated values.

- The **noble gases** have astronomical ionization energies.
- Their electron affinities are positive.
  - Therefore, they are relatively unreactive.
- They are found as monatomic gases.