CHEM 103
CHEMISTRY I

CHAPTER 7
PERIODIC PROPERTIES OF ELEMENTS

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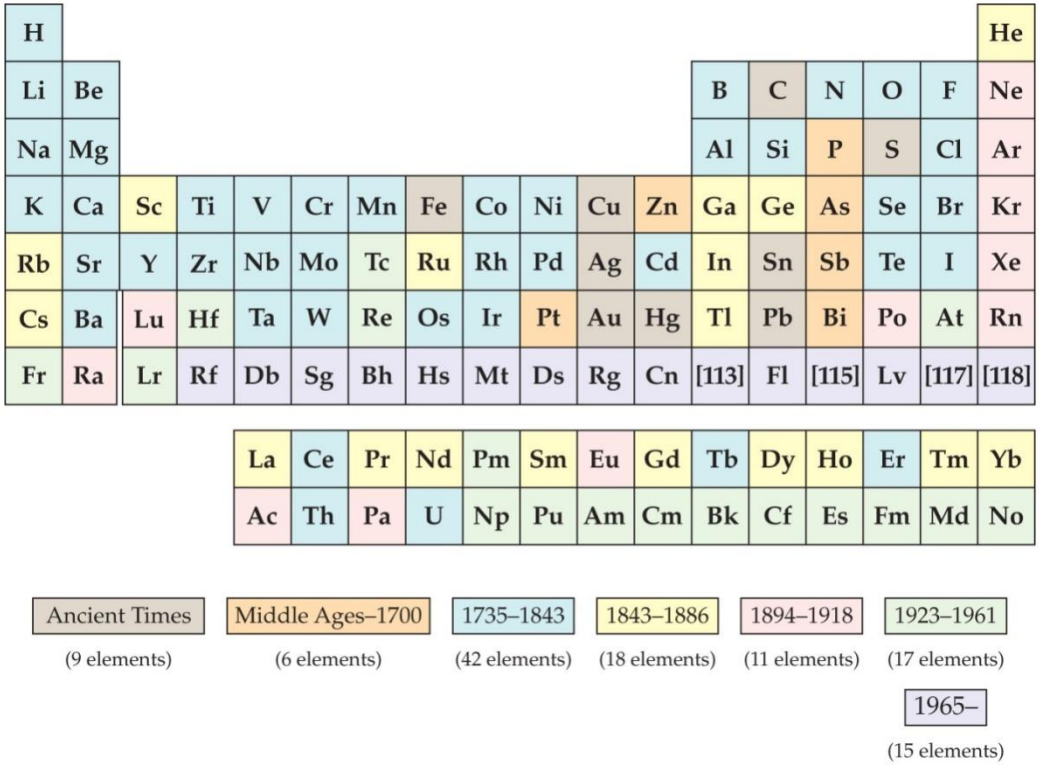
Development of the Periodic Table



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

Mendeleev and the 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

Chemists mostly credit Mendeleev because he also used chemical properties to organize the table and predicted some missing elements and their expected properties, including germanium.

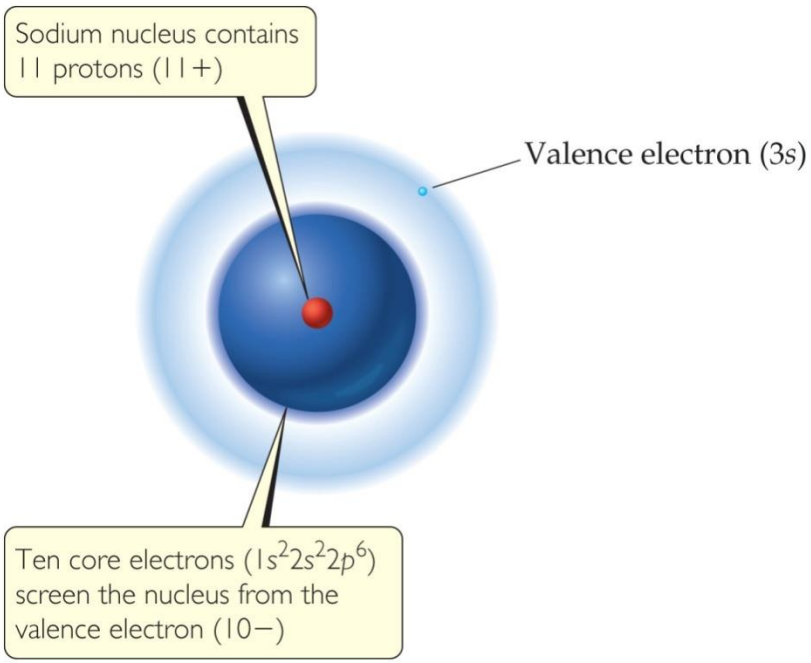
Atomic Number

- Mendeleev's table was based on atomic masses. It was the most fundamental property of elements known at the time.
- About 35 years later, the nuclear atom was discovered by Ernest Rutherford.
- Henry Moseley developed the concept of atomic number experimentally. The number of protons was considered the basis for the periodic property of elements.

Periodicity

- Periodicity is the repetitive pattern of a property for elements based on atomic number.
- The following properties are discussed in this chapter:
  - Sizes of atoms and ions
  - Ionization energy
  - Electron affinity
  - Some group chemical property trends
- First, we will discuss a fundamental property that leads to may of the trends, effective nuclear charge.

Effective Nuclear Charge



- Many properties depend on attractions between valence electrons and the nucleus.
- Electrons are both attracted to the nucleus and repelled by other electrons.
- The forces an electron experiences depend on both factors.

Effective Nuclear Charge

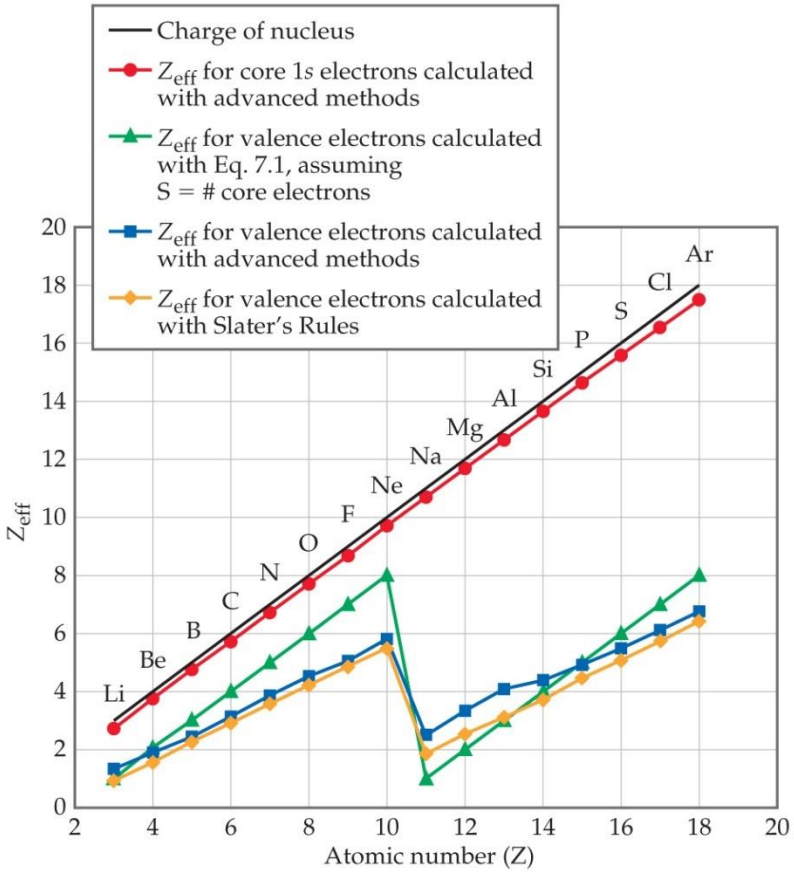
- The effective nuclear charge, Z_{eff}, is found this way:

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

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

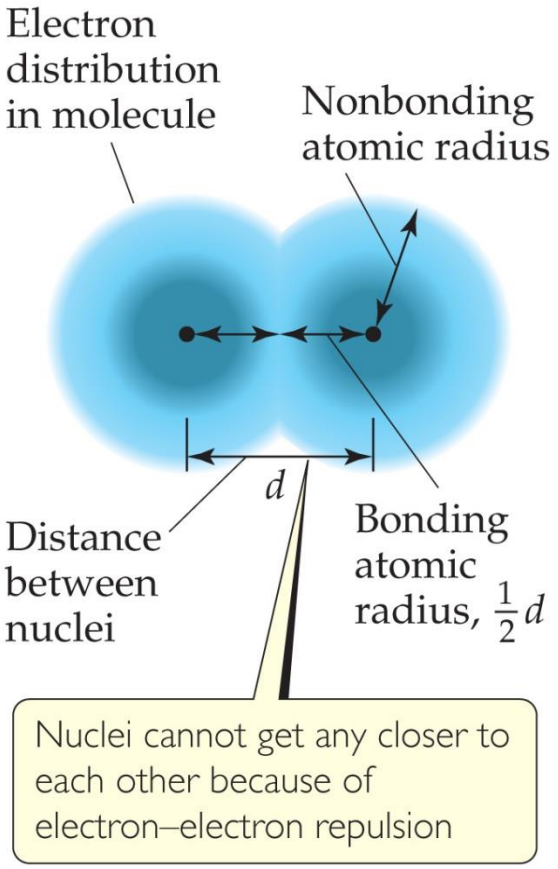
- Effective nuclear charge is a periodic property:
  - It increases across a period.

Effective Nuclear Charge Increases across a Period



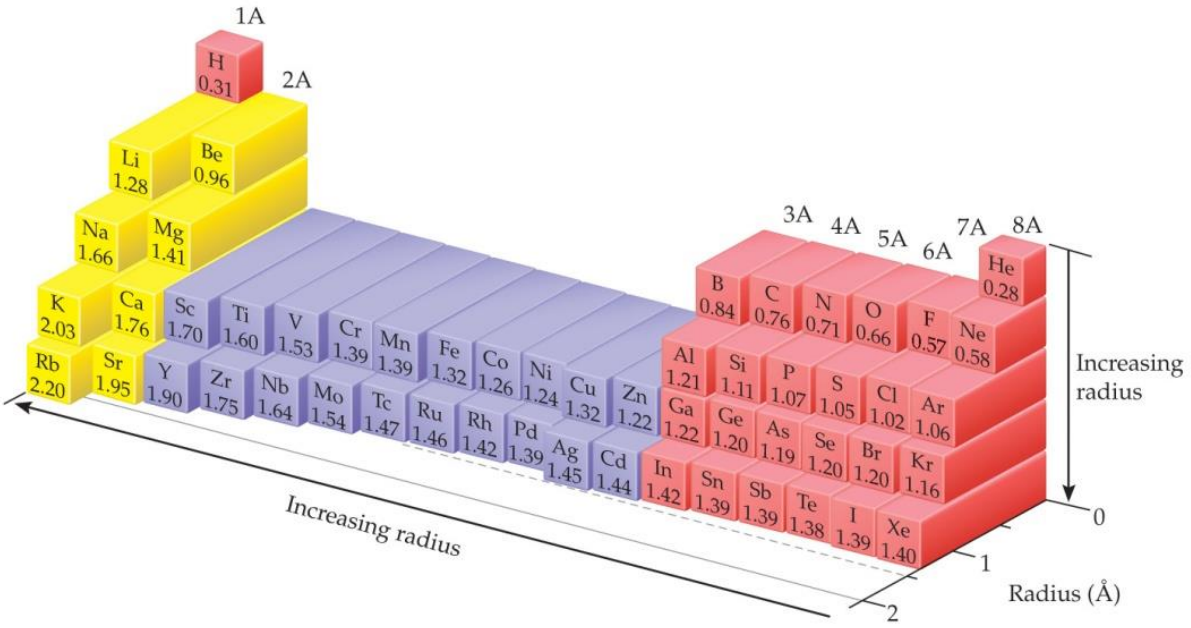
What Is the Size of an Atom?

The **nonbonding atomic radius** or **van der Waals radius** is half of the shortest distance separating two nuclei during a collision of atoms.

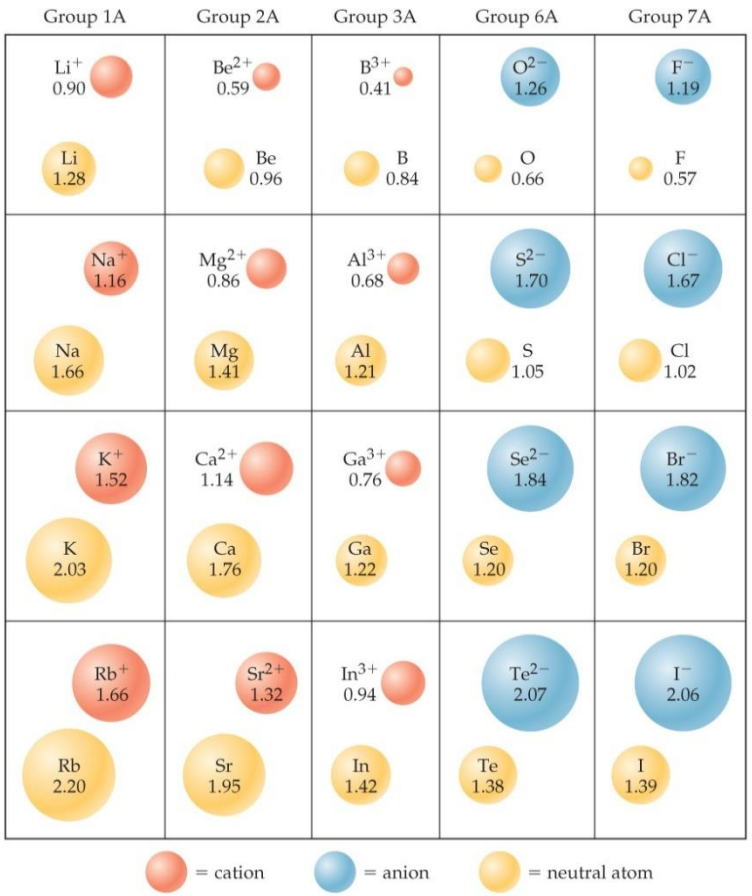


Sizes of Atoms

- The bonding atomic radius is half the internuclear distance when atoms are bonded.
- The bonding atomic radius tends to
  - — decrease from left to right across a period (Z_{eff} ↑).
  - — increase from top to bottom of a group (n ↑).

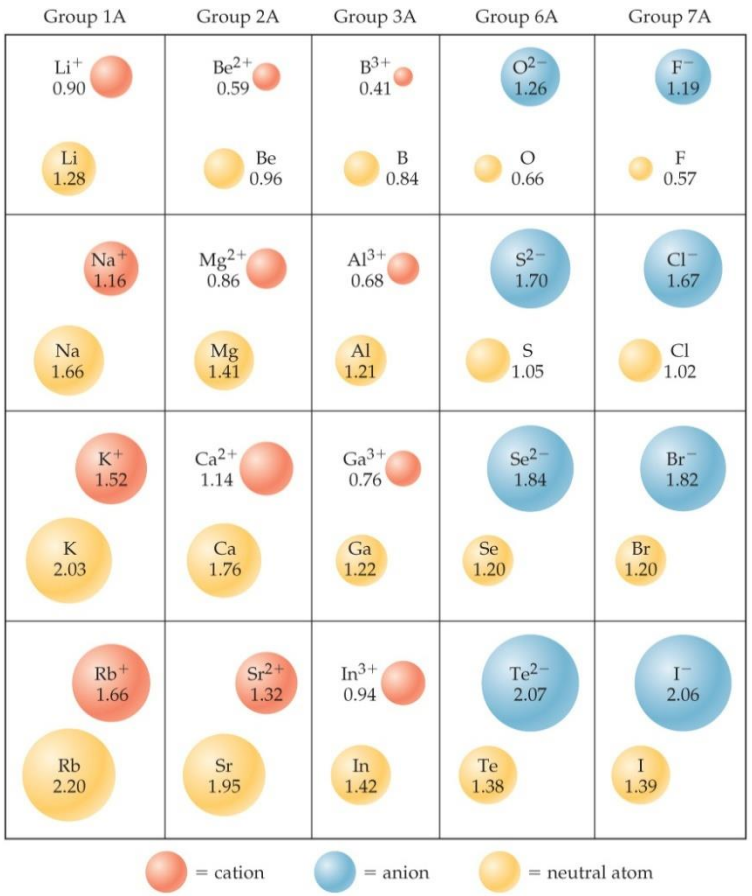


Sizes of Ions



- Determined by interatomic distances in ionic compounds
- Ionic size depends on
  - 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.
- Anions are larger than their parent atoms:
  - Electrons are added and repulsions between electrons are increased.

Size of Ions—Isoelectronic Series

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

➢ **An Isoelectronic Series (10 electrons)**

- Note increasing nuclear charge with decreasing ionic radius as atomic number increases

O^{2-}	F^{-}	Na^{+}	Mg^{2+}	Al^{3+}
1.26 Å	**1.19 Å**	**1.16 Å**	**0.86 Å**	**0.68 Å**

Ionization Energy (*I*)

- The **ionization energy** is the minimum 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.
- Note: the higher the ionization energy, the more difficult it is to remove an electron!

Ionization Energy

- It requires more energy to remove each successive electron.
- When all valence electrons have been removed, it takes a great deal more energy to remove the next electron.

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

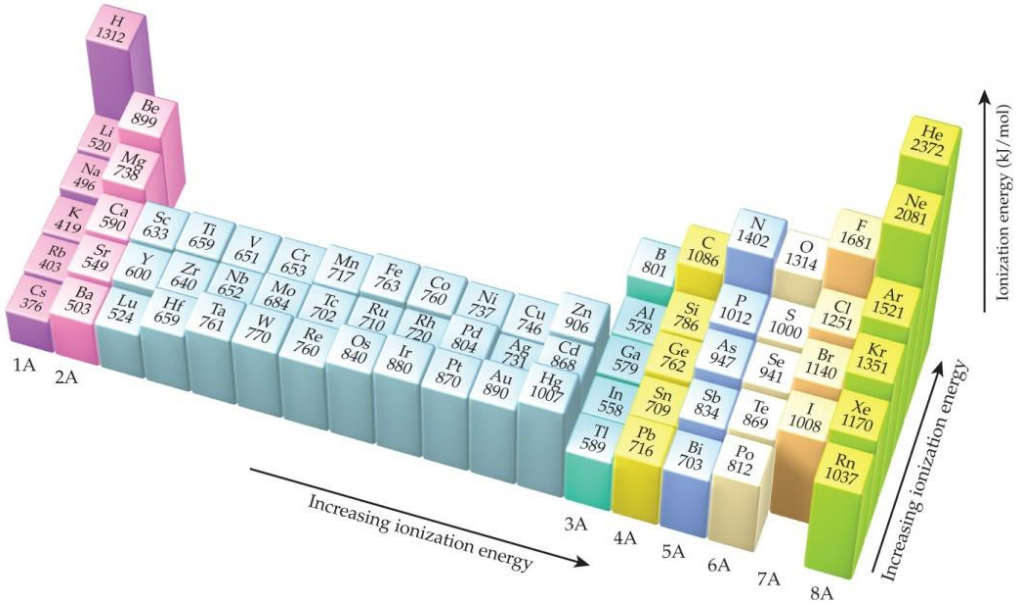
Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4562	(inner-shell electrons)				
Mg	738	1451	7733				
Al	578	1817	2745	11,577			
Si	786	1577	3232	4356	16,091		
P	1012	1907	2914	4964	6274	21,267	
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018
Ar	1521	2666	3931	5771	7238	8781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

Periodic Trends in First Ionization Energy (I_1)

1) I_1 generally increases across a period.
2) I_1 generally decreases down a group.
3) The *s*- and *p*-block elements show a larger range of values for I_1. (The *d*-block generally increases slowly across the period; the *f*-block elements show only small variations.)

Factors that Influence Ionization Energy

- Smaller atoms have higher *I* values.
- *I* values depend on effective nuclear charge and average distance of the electron from the nucleus.

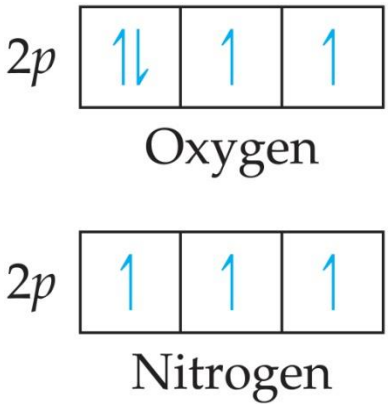


Irregularities in the General Trend

- The trend is **not** followed when the added valence electron in the next element

➢ enters a new sublevel (higher energy sublevel);

➢ is the first electron to pair in one orbital of the sublevel (electron repulsions lower energy).



Electron Configurations of Ions

- Cations: The electrons are lost from the *highest* energy level (*n* value).
  - Li^+ is $1s^2$ (losing a 2*s* electron).
  - Fe^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ (losing two 4*s* electrons).
- Anions: The electron configurations are filled to ns^2np^6; e.g., F^- is $1s^2 2s^2 2p^6$ (gaining one electron in 2*p*).

Electron Affinity

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

$$Cl + e^- \longrightarrow Cl^-$$

- It is typically *exothermic*, so, for most elements, it is negative!

General Trend in Electron Affinity

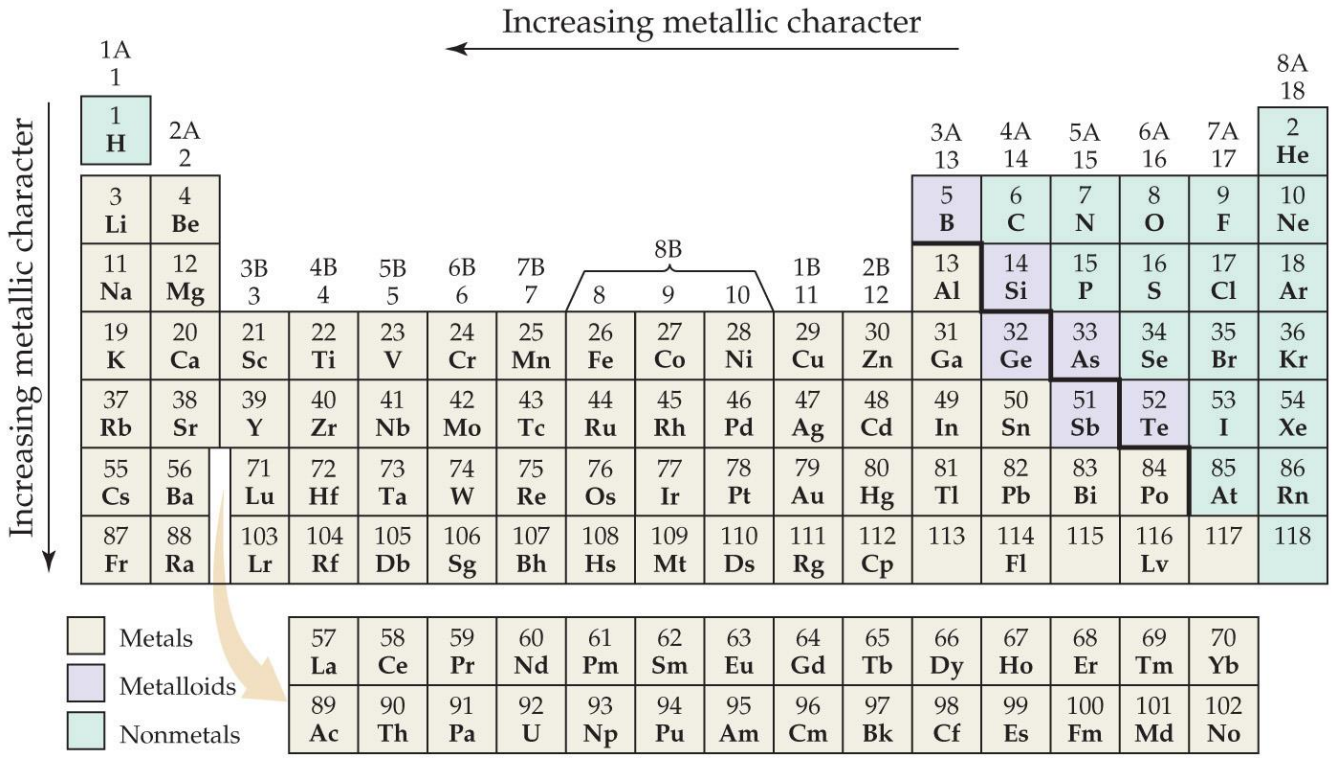
- Not much change in a group.
- Across a period, it generally increases. *Three* notable exceptions include the following:

1) Group 2A: *s* sublevel is full!
2) Group 5A: *p* sublevel is half-full!
3) Group 8A: *p* sublevel is full!

Note: the electron affinity for many of these elements is *positive* (X^- is unstable).

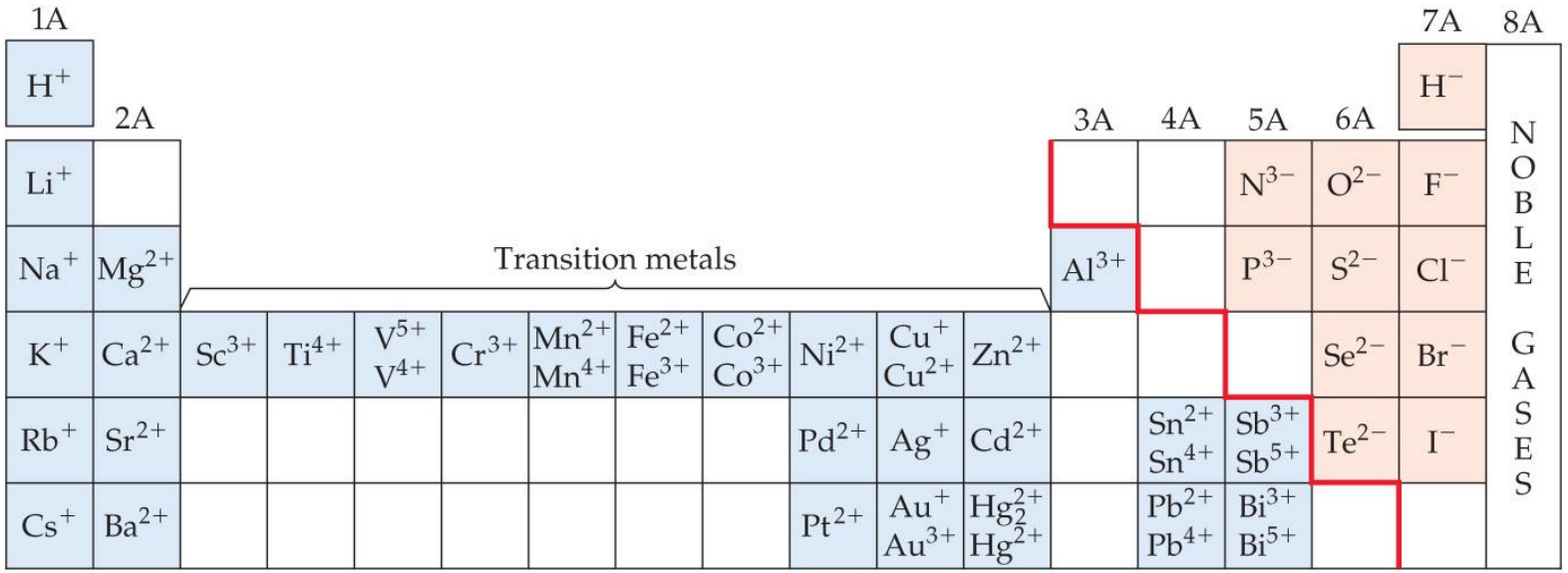
1A							8A
H −73	2A	3A	4A	5A	6A	7A	He > 0
Li −60	Be > 0	B −27	C −122	N > 0	O −141	F −328	Ne > 0
Na −53	Mg > 0	Al −43	Si −134	P −72	S −200	Cl −349	Ar > 0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr > 0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe > 0

Metal, Nonmetals, and Metalloids



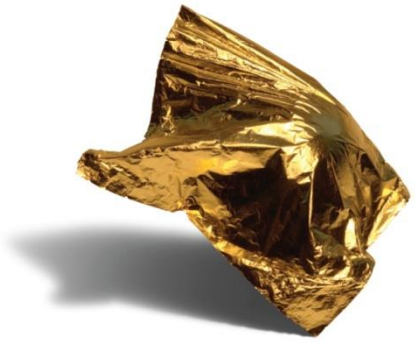
Metals Differ from Nonmetals

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



Metals

- Most of the elements in nature are metals.
- Properties of metals:
  - Shiny luster
  - Conduct heat and electricity
  - Malleable and ductile
  - Solids at room temperature (except mercury)
  - Low ionization energies/form cations easily



Metal Chemistry

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

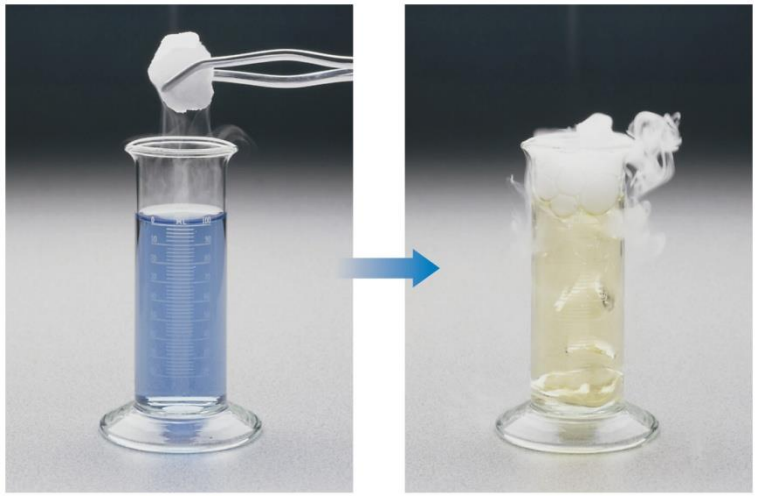
Nickle oxide (NiO), nitric acid (HNO_3), and water

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 found on the right hand side of the periodic table.
- Properties of nonmetals include the following:
- Solid, liquid, or gas (depends on element)
- Solids are dull, brittle, poor conductors
- Large negative electronegativity/form anions readily

Nonmetal Chemistry



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

Recap of a Comparison of the Properties of Metals and 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, and 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

Metalloids

- Metalloids have some characteristics of metals and some of nonmetals.
- Several metalloids are electrical semiconductors (computer chips).

Group Trends

- Elements in a group have similar properties.
- Trends also exist within groups.
- Groups Compared:

➢ Group 1A: The Alkali Metals

➢ Group 2A: The Alkaline Earth Metals

➢ Group 6A: The Oxygen Group

➢ Group 7A: The Halogens

➢ Group 8A: The Noble Gases

Alkali Metals

- **Alkali metals** are soft, metallic solids.
- They are found only in compounds in nature, not in their elemental forms.
- Typical metallic properties (luster, conductivity) are seen in them.

Alkali Metal Properties

- 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	1.28	520
Sodium	$[Ne]3s^1$	98	0.97	1.66	496
Potassium	$[Ar]4s^1$	63	0.86	2.03	419
Rubidium	$[Kr]5s^1$	39	1.53	2.20	403
Cesium	$[Xe]6s^1$	28	1.88	2.44	376

Alkali Metal Chemistry

Li Na K

Their reactions with water are famously exothermic.

Differences in Alkali Metal Chemistry

- Lithium reacts with oxygen to make an oxide:

$$4\ Li + O_2 \longrightarrow 2\ Li_2O$$

- Sodium reacts with oxygen to form a peroxide:

$$2\ Na + O_2 \longrightarrow Na_2O_2$$

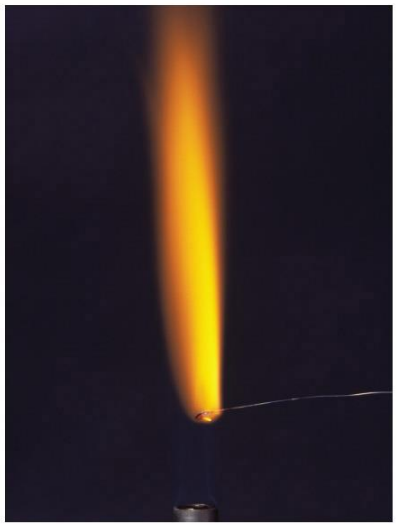
- K, Rb, and Cs also form superoxides:

$$M + O_2 \longrightarrow MO_2$$

Flame Tests

- Qualitative tests for alkali metals include their characteristic colors in flames.

Li



Na

K

Alkaline Earth Metals—Compare to Alkali Metals

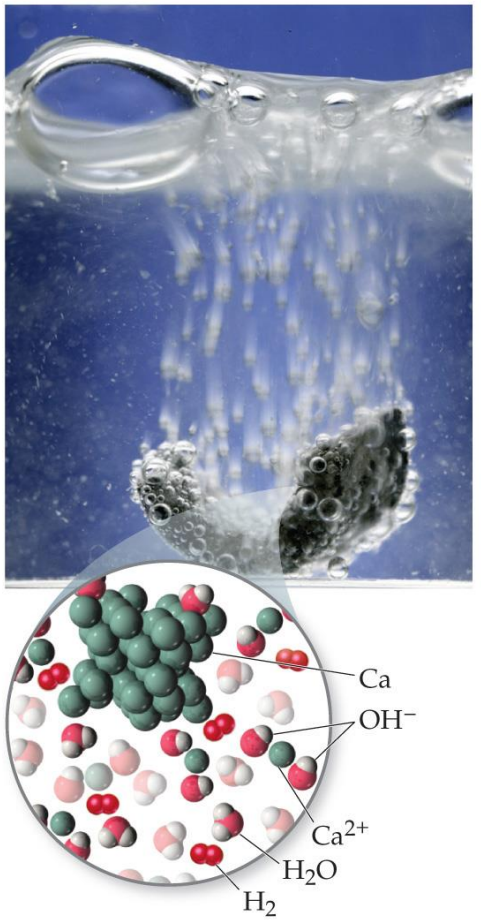
Table 7.5 Some Properties of the Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Beryllium	$[He]2s^2$	1287	1.85	0.96	899
Magnesium	$[Ne]3s^2$	650	1.74	1.41	738
Calcium	$[Ar]4s^2$	842	1.55	1.76	590
Strontium	$[Kr]5s^2$	777	2.63	1.95	549
Barium	$[Xe]6s^2$	727	3.51	2.15	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—Increasing in Metallic Character down the Group

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	$[He]2s^2 2p^4$	−218	1.43 g/L	0.66	1314
Sulfur	$[Ne]3s^2 3p^4$	115	1.96 g/cm^3	1.05	1000
Selenium	$[Ar]3d^{10} 4s^2 4p^4$	221	4.82 g/cm^3	1.20	941
Tellurium	$[Kr]4d^{10} 5s^2 5p^4$	450	6.24 g/cm^3	1.38	869
Polonium	$[Xe]4f^{14} 5d^{10} 6s^2 6p^4$	254	9.20 g/cm^3	1.40	812

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

Group 7A—Halogens

Table 7.7 Some Properties of the Halogens

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Fluorine	$[He]2s^2 2p^5$	−220	1.69 g/L	0.57	1681
Chlorine	$[Ne]3s^2 3p^5$	−102	3.12 g/L	1.02	1251
Bromine	$[Ar]4s^2 3d^{10} 4p^5$	−7.3	3.12 g/cm^3	1.20	1140
Iodine	$[Kr]5s^2 4d^{10} 5p^5$	114	4.94 g/cm^3	1.39	1008

- The **halogens** are typical nonmetals.
- They have highly negative electron affinities, so they exist as anions in nature.
- They react directly with metals to form metal halides.

Group 8A—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	0.28	2372
Neon	$[He]2s^22p^6$	27.1	0.90	0.58	2081
Argon	$[Ne]3s^23p^6$	87.3	1.78	1.06	1521
Krypton	$[Ar]4s^23d^{10}4p^6$	120	3.75	1.16	1351
Xenon	$[Kr]5s^24d^{10}5p^6$	165	5.90	1.40	1170
Radon	$[Xe]6s^24f^{14}5d^{10}6p^6$	211	9.73	1.50	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 very large ionization energies.
- Their electron affinities are positive (can't form stable anions).
- Therefore, they are relatively unreactive.
- They are found as monatomic gases.

Periodic Table of the Elements