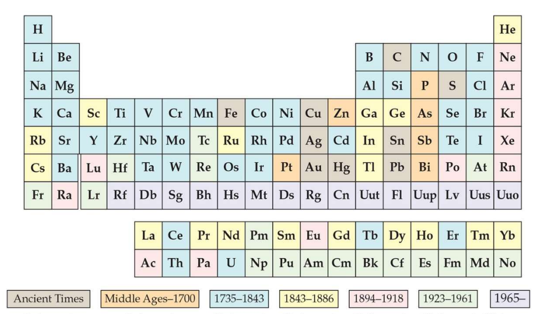
Chapter 7
Periodic Properties of the Elements

Learning Outcomes:

- ➢Explain the meaning of effective nuclear charge, Z_{eff}, and how Z_{eff} depends on nuclear charge and electron configuration.
- ➢Predict the trends in atomic radii, ionic radii, ionization energy, and electron affinity by using the periodic table.
- ➢Explain how the radius of an atom changes upon losing electrons to form a cation or gaining electrons to form an anion.
- ➢Write the electron configurations of ions.
- ➢Explain how the ionization energy changes as we remove successive electrons, and the jump in ionization energy that occurs when the ionization corresponds to removing a core electron.
- ➢Explain how irregularities in the periodic trends for electron affinity can be related to electron configuration.
- ➢Explain the differences in chemical and physical properties of metals and nonmetals, including the basicity of metal oxides and the acidity of nonmetal oxides.
- ➢Correlate atomic properties, such as ionization energy, with electron configuration, and explain how these relate to the chemical reactivity and physical properties of the alkali and alkaline earth metals (groups 1A and 2A).
- ➢Write balanced equations for the reactions of the group 1A and 2A metals with water, oxygen, hydrogen, and the halogens.
- ➢List and explain the unique characteristics of hydrogen.
- ➢Correlate the atomic properties (such as ionization energy, electron configuration, and electron affinity) of group 6A, 7A, and 8A elements with their chemical reactivity and physical properties.

Development of Periodic Table



- •Dmitri Mendeleev and Lothar Meyer (~1869) independently came to the same conclusion about how elements should be grouped in the periodic table.
- •Henry Moseley (1913) developed the concept of atomic numbers (the number of protons in the nucleus of an atom)

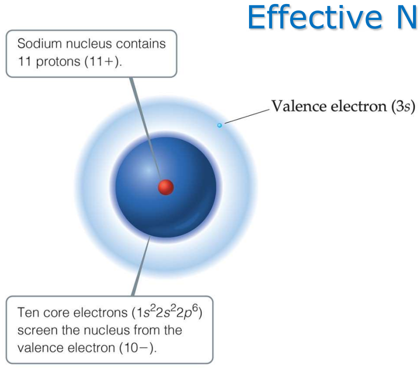
Predictions and the Periodic Table

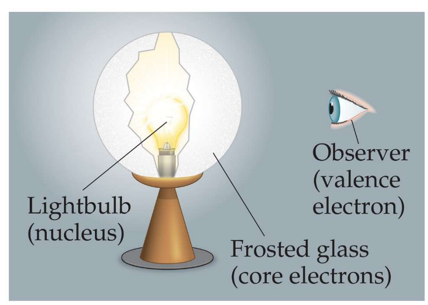
TABLE 7.1 Comparison of the Properties of Eka-Silicon Predicted by Mendeleev with the Observed Properties of Germanium

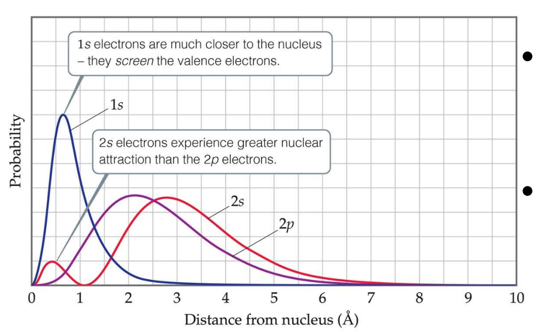
Property	Mendeleev's Predictions for Eka-Silicon (made in 1871)	Observed Properties of Germanium (discovered in 1886)
Atomic weight	72	72.59
Density (g/cm^3)	5.5	5.35
Specific heat (J/g-K)	0.305	0.309
Melting point (°C)	High	947
Color	Dark gray	Grayish white
Formula of oxide	XO_2	GeO_2
Density of oxide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

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.

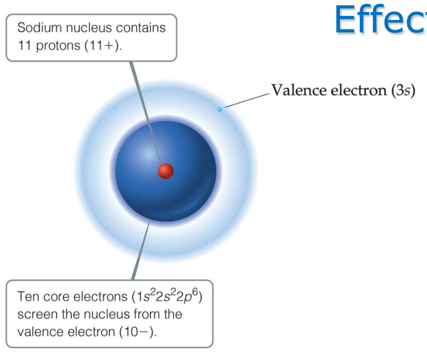
Effective Nuclear Charge







- 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.

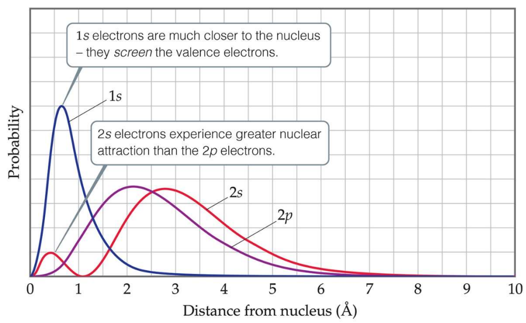


Effective Nuclear Charge

The effective nuclear charge, Z_{eff}, is:

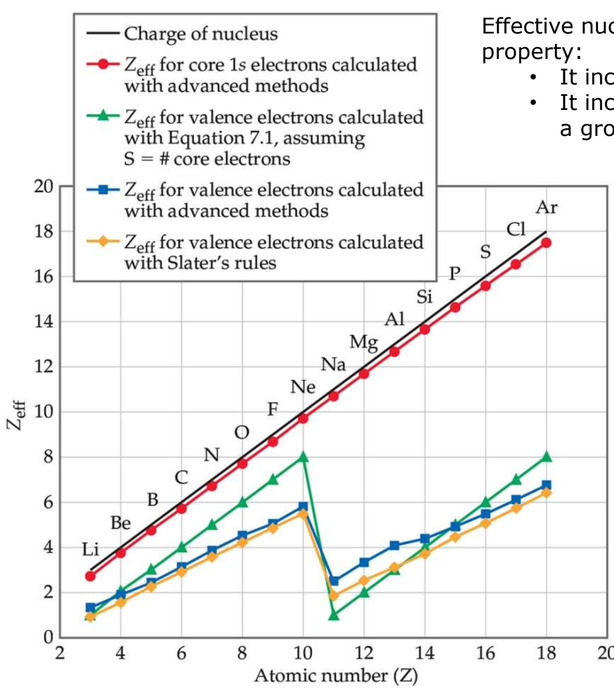
$$Z_{eff} = Z - S$$

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



The valence electrons do not screen as efficiently as core electrons.

Z_{eff} is a representation of the *average* electrical field experienced by a single electron.

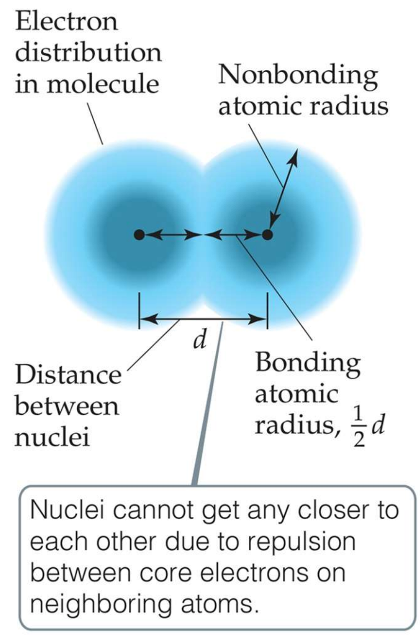


Effective nuclear charge is a periodic property:

- It increases across a period.
- It increases *slightly* down a group.

Sizes of Atoms

- The **nonbonding atomic radius,** or **van der Waals radius,** is half of the shortest distance separating two nuclei during a collision of atoms.
- The **bonding atomic radius** is defined as one-half of the distance between covalently bonded nuclei.



Periodic Trends in Atomic Radii

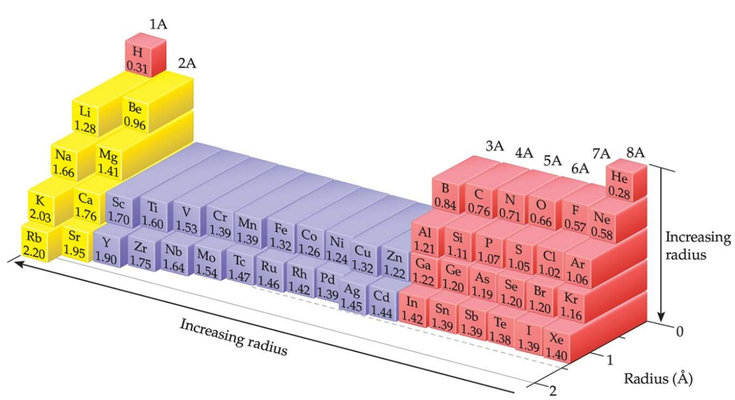
Bonding atomic radius tends to…

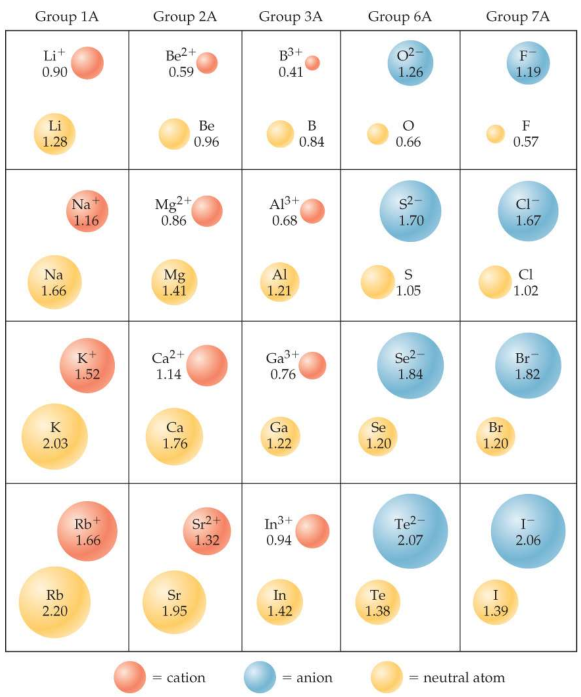
…decrease from left to right across a row

due to increasing Z_{eff}.

…increase from top to bottom of a column

due to increasing value of n





Sizes of Ions

Ionic size depends upon:

- Nuclear charge, number of electrons, orbitals in which electrons reside.
- Cations are smaller than their parent atoms.
  - The outermost electron is removed and repulsions are reduced.
- Anions are larger than their parent atoms.
  - Electrons are added and repulsions are increased.

Sizes of Ions

- In an **isoelectronic series**, ions have the same number of electrons.
- Ionic size decreases with an increasing nuclear charge.
- Increasing nuclear charge with decreasing ionic radius as atomic number increases.

Increasing nuclear charge ⟶

8 protons 10 electrons	9 protons 10 electrons	11 protons 10 electrons	12 protons 10 electrons	13 protons 10 electrons
O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}
1.26 Å	1.19 Å	1.16 Å	0.86 Å	0.68 Å

Decreasing ionic radius ⟶

Ionization Energy

Amount of energy required to remove an electron from the ground state of a gaseous atom or ion.

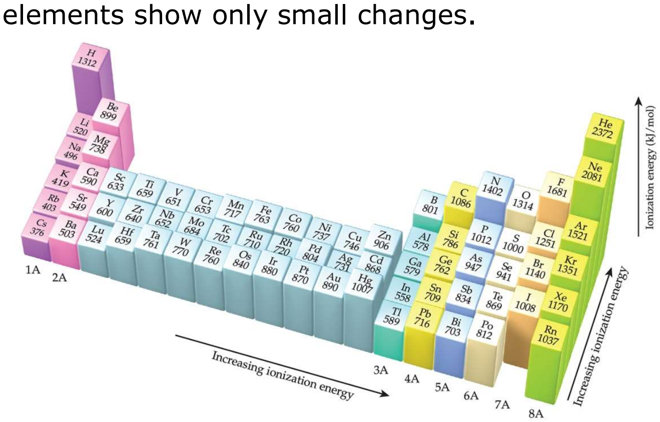
- First ionization energy is that energy required to remove first electron.
- Second ionization energy is that energy required to remove second electron, etc.

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					
Mg	738	1451	7733		(inner-shell electrons)		
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

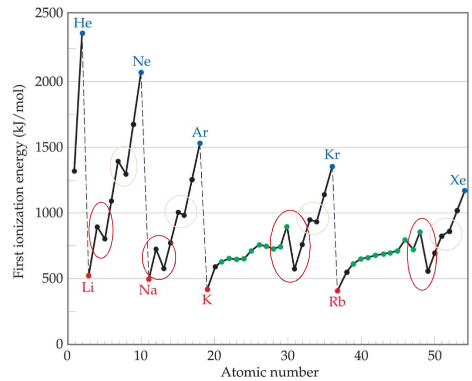
Trends in First Ionization Energies

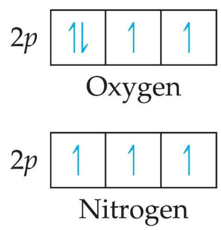
- I_1 generally decreases down a group. For atoms in the same group, Z_{eff} is essentially the same, but the valence electrons are farther from the nucleus.
- I_1 generally increases across a period.
- The *s*- and *p*-block elements show a larger range of values for I_1.
- The *d*-block generally increases slightly across the period; the *f*-block elements show only small changes.



Irregularities in First Ionization Energies

- Discontinuity occurs between Groups IIA and IIIA.
- Electron removed from *p*-orbital rather than *s*-orbital
  - Electron farther from nucleus
  - Small amount of repulsion by *s* electrons.
- The second occurs between Groups VA and VIA.
  - Electron removed comes from doubly occupied orbital.
  - Repulsion from other electron in orbital helps in its removal.

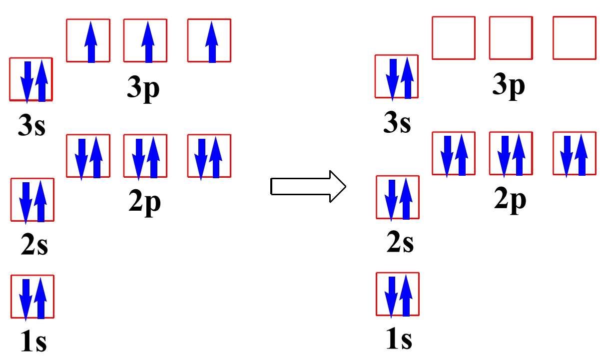




Electronic Configurations of Ions

To form cations from elements remove e⁻'s from the subshell with the highest n.

P([Ne]$3s^2 3p^3$) → P^{3+}([Ne]$3s^2$) + 3e⁻

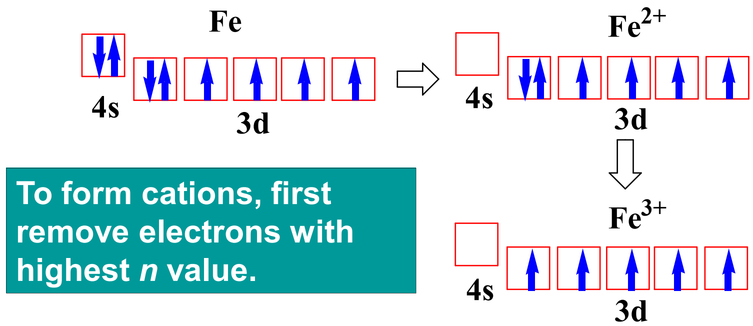


P P^{3+}

Electronic Configurations of Ions

For transition metals, remove *n*s electrons and then (*n* - 1)d electrons.

$Fe([Ar]4s^2 3d^6) \rightarrow Fe^{2+}([Ar]3d^6) + 2e^-$



$Fe^{2+}([Ar]3d^6) \rightarrow Fe^{3+}([Ar]3d^5) + e^-$

Electron Affinity

Energy change accompanying addition of electron to gaseous atom:

$$Cl(g) + e^- \longrightarrow Cl^-(g)$$

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.

Three notable exceptions include the following:

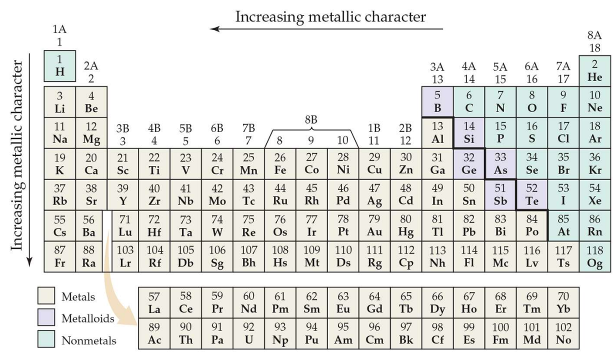
1) Group 2A: *s* sublevel is full
2) Group 5A: *p* sublevel is half-full
3) Group 8A: *p* sublevel is full

Halogens have the most negative electron affinity.

Metal, Nonmetals, and Metalloids

TABLE 7.3 Characteristic Properties of Metals and Nonmetals

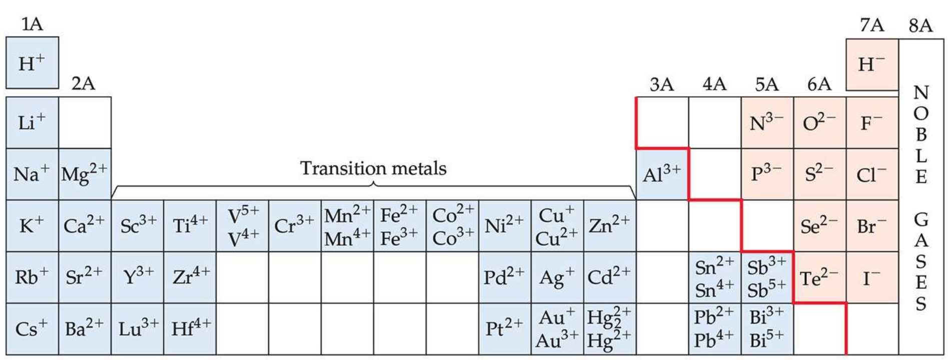
Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, and some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal 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



Metals versus Nonmetals

Metals tend to form cations.

Nonmetals tend to form anions.



Metals

Nickel 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 $Ni(NO_3)_2$.

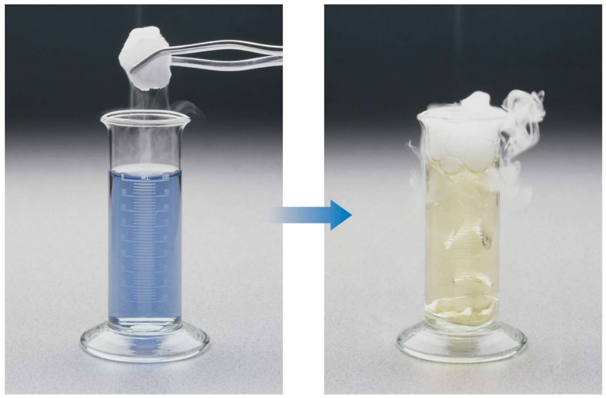
- Compounds formed between metals and nonmetals tend to be ionic.
- Metal oxides tend to be basic.

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2\ (aq)$$

Nonmetals

- Dull, brittle substances that are poor conductors of heat and electricity.
- Large negative electron affinity, tend to gain electrons in reactions with metals to acquire noble gas configuration.
- Substances containing only nonmetals are molecular compounds.

Nonmetals



- Most nonmetal oxides are acidic.
- Nonmetal oxides react with bases to form salts and water.

Metalloids

13	14	15	16
5 B	6 C	7 N	8 O
13 Al	14 Si	15 P	16 S
31 Ga	32 Ge	33 As	34 Se
49 In	50 Sn	51 Sb	52 Te

- B, Si, Ge, As, Sb, Te
- Have some characteristics of metals, some of nonmetals.
- For instance, silicon looks shiny (metallic luster), but is brittle and fairly poor conductor (semiconductor).

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
  - Hydrogen: nonmetal

Alkali Metals

- Soft, metallic solids.
- Name comes from Arabic word for ashes.
- Found only as compounds in nature.
- Have low densities and melting points.
- 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

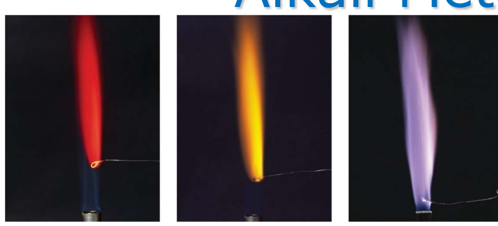
Alkali Metals

Li Na K

Reactions with water are exothermic

React with oxygen, hydrogen, and halogens.

Alkali Metals



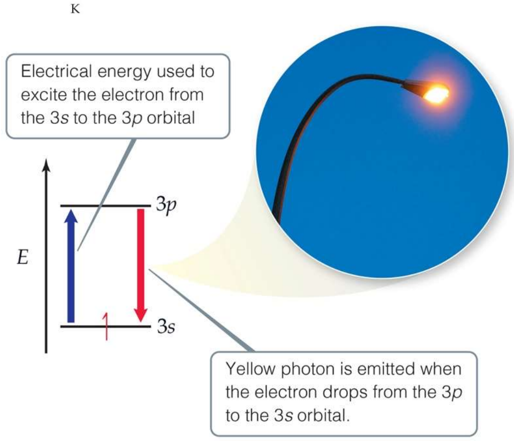
Li Na K

Alkali metals (except Li) react with oxygen to form peroxides.

- K, Rb, and Cs also form superoxides:

$$K + O_2 \longrightarrow KO_2$$

Produce bright colors when placed in flame.



Alkaline Earth Metals

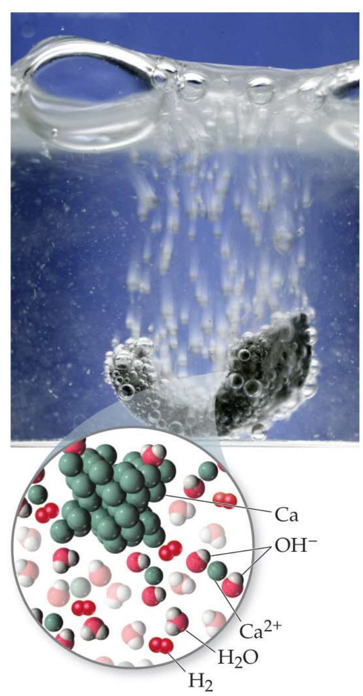
TABLE 7.5 Some Properties of the Alkaline Earth Metals

Element	Electron Configuration	Melting Point (°C)	Density (g/cm^3)	Atomic Radius (Å)	I_1 (kJ/mol)
Beryllium	$[He]2s^2$	1287	1.85	0.96	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

- Have higher densities and melting points than alkali metals.
- Have low ionization energies, but not as low as alkali metals.

Alkaline Earth Metals

- Be does not react with water, Mg reacts only with steam, but others react readily with water.
- Reactivity tends to increase as go down group.



Hydrogen

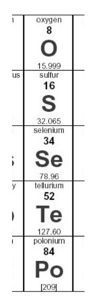
- $1s^1$ a metallic electron configuration like the other ns^1 elements
- We do think of acid compounds, like HCl, as having H^+, however they are really covalent in nature.
- When reacting with metals, hydride anions (H^-) form.
- Forms both ionic (H^- with metals) and molecular compounds (H^+ with nonmetals).

Group 6A - chalcogens

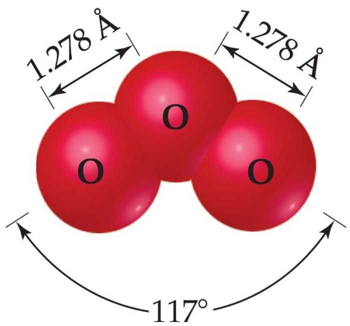
TABLE 7.6 Some Properties of the Group 6A Elements

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Oxygen	$[He]2s^22p^4$	−218	1.43 g/L	0.66	1314
Sulfur	$[Ne]3s^23p^4$	115	1.96 g/cm³	1.05	1000
Selenium	$[Ar]3d^{10}4s^24p^4$	221	4.82 g/cm³	1.20	941
Tellurium	$[Kr]4d^{10}5s^25p^4$	450	6.24 g/cm³	1.38	869
Polonium	$[Xe]4f^{14}5d^{10}6s^26p^4$	254	9.20 g/cm³	1.40	812

- Oxygen, sulfur, and selenium are nonmetals.
- Tellurium is a metalloid.
- Radioactive polonium is a metal, forms a cation.

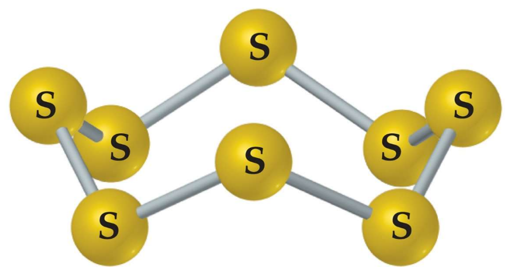


Oxygen and Sulfur



- Two allotropes:
  - O_2, dioxygen
  - O_3, ozone
- Three anions:
  - O^{2-}, oxide
  - O_2^{2-}, peroxide
  - O_2^{1-}, superoxide
- Tends to take electrons from other elements (oxidation)

- S is weaker oxidizing agent than oxygen.
- Most stable allotrope is S_8, a ringed molecule shaped like a crown.



Group VIIA: Halogens

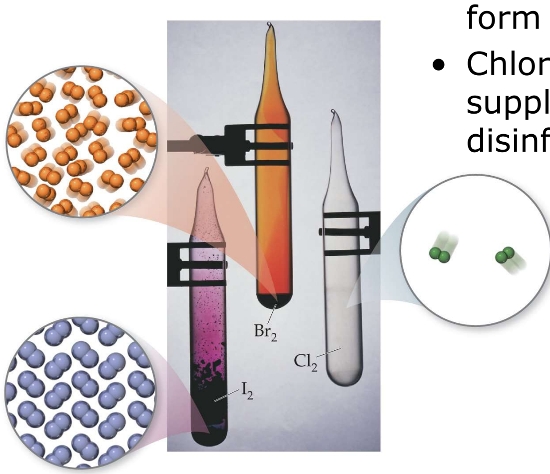
TABLE 7.7 Some Properties of the Halogens

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Fluorine	$[He]2s^22p^5$	−220	1.69 g/L	0.57	1681
Chlorine	$[Ne]3s^23p^5$	−102	3.12 g/L	1.02	1251
Bromine	$[Ar]4s^23d^{10}4p^5$	−7.3	3.12 g/cm^3	1.20	1140
Iodine	$[Kr]5s^24d^{10}5p^5$	114	4.94 g/cm^3	1.39	1008

- Prototypical nonmetals
- Name comes from the Greek *halos* and *gennao*: “salt formers”

Group VIIA: Halogens

- Large, negative electron affinities
  - Therefore, tend to oxidize other elements easily
- React directly with metals to form metal halides
- Chlorine added to water supplies to serve as 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	0.28	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.

- Monatomic gases
- Large ionization energies
- Positive electron affinities
- Relatively unreactive
  - Only fluorine (F) can remove electrons to form compounds
  - XeF_2, XeF_4 XeF_6, KrF_2 are known

XeF_4