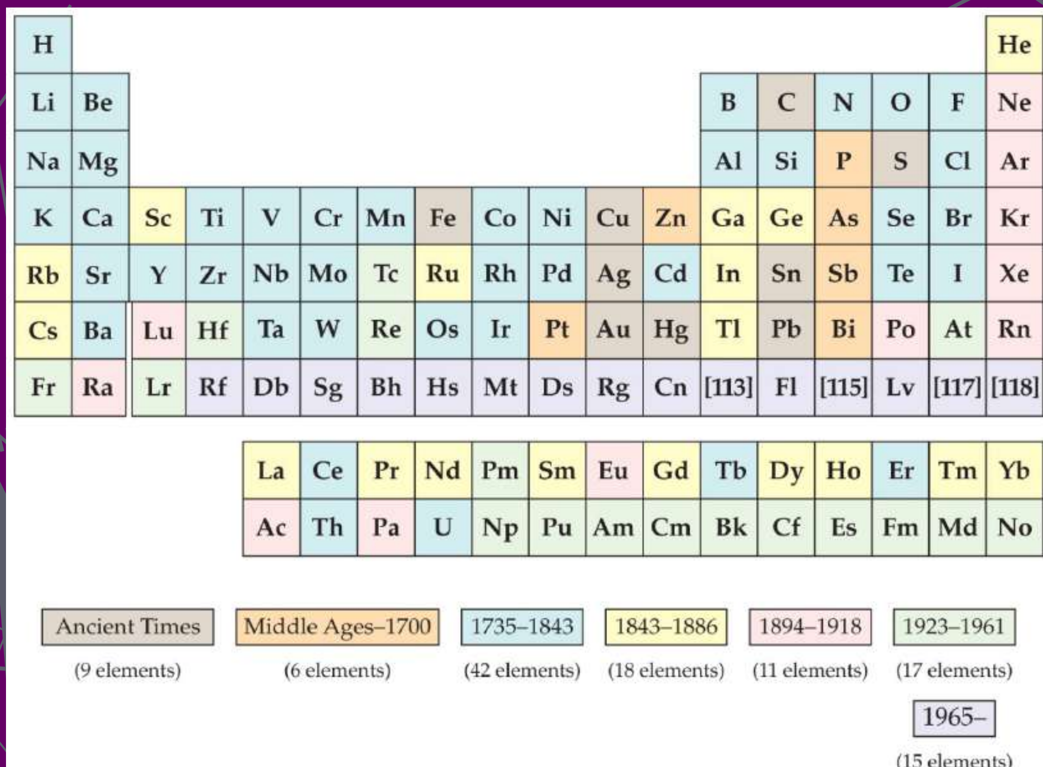
Chapter 7
Periodic Properties of the Elements

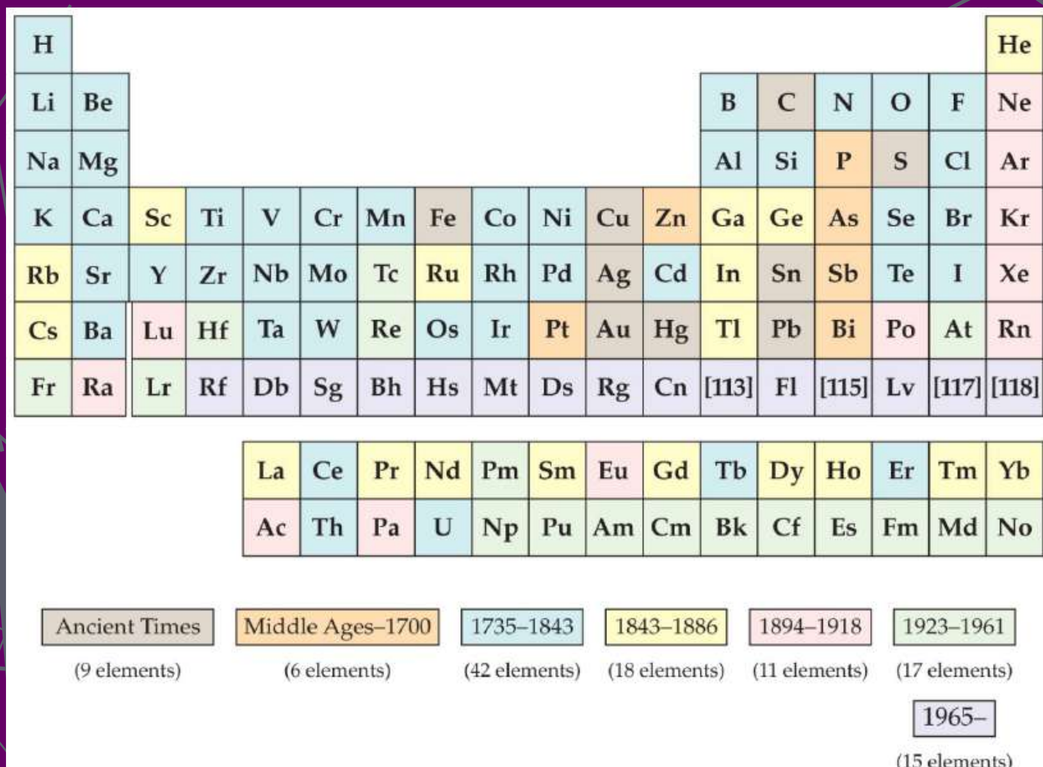
7.1 Development of the Periodic Table

Development of the Periodic Table



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

Development of the Periodic Table



- Copper, silver, and gold have all been known since ancient times, whereas most of the other metals have not. Why do you think this is the case?

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

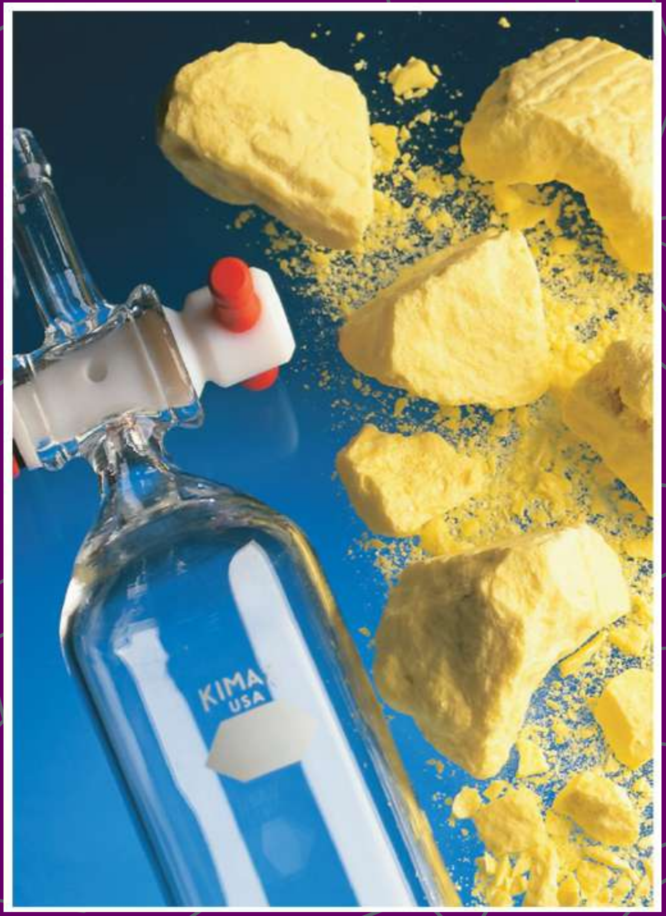
- Chemists mostly credit Mendeleev because he also used chemical properties to organize the table and predicted some missing elements and their expected properties, including germanium.

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.

Development of Periodic Table

- Elements in the same group generally have similar chemical properties.
- Physical properties are not necessarily similar, however.



7.1 Give It Some Thought

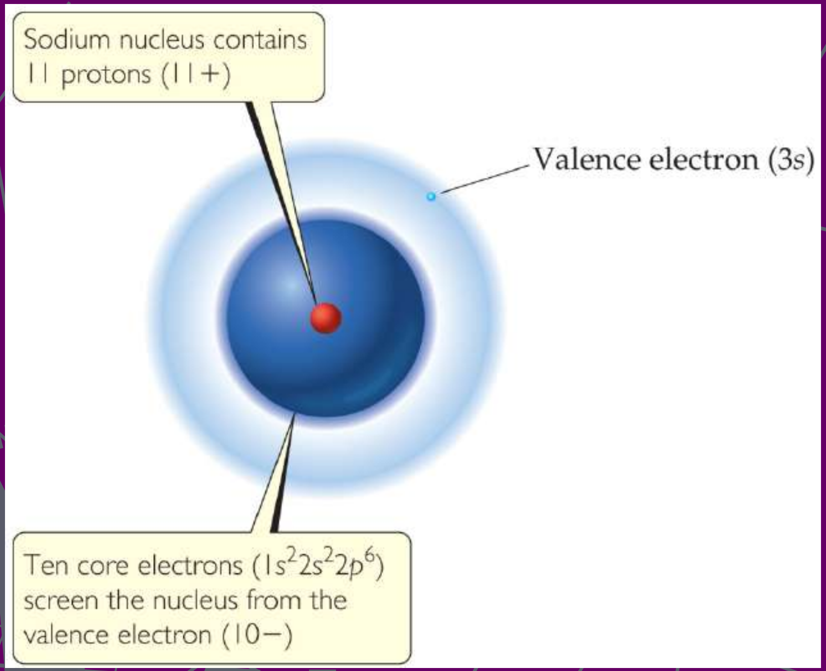
- Can you find an example other than Ar and K where the order of the elements would be different if the elements were arranged in order of increasing atomic weight?

7.2 Effective Nuclear Charge

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

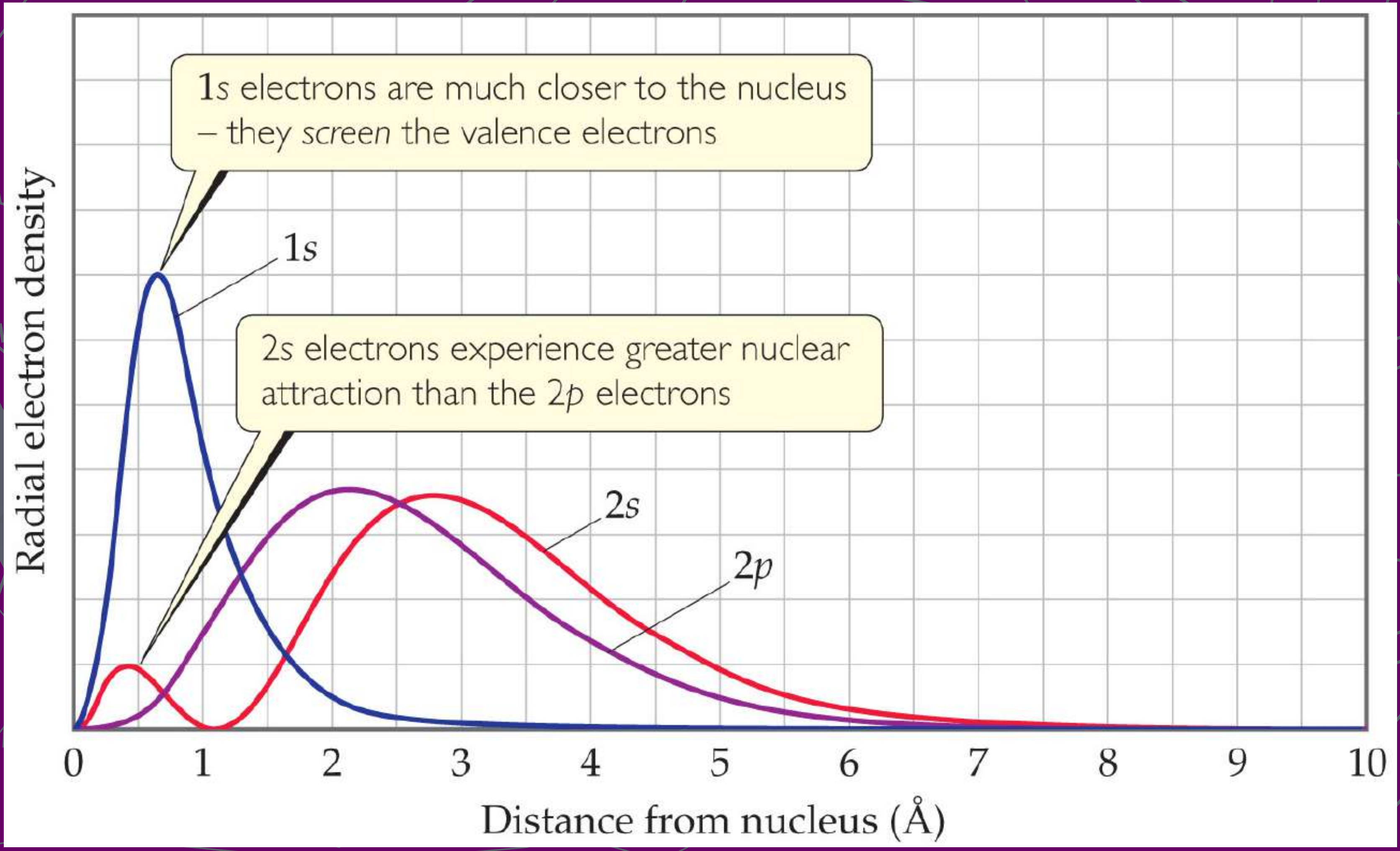
- The effective nuclear charge, Z_{eff}, is found this way:

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

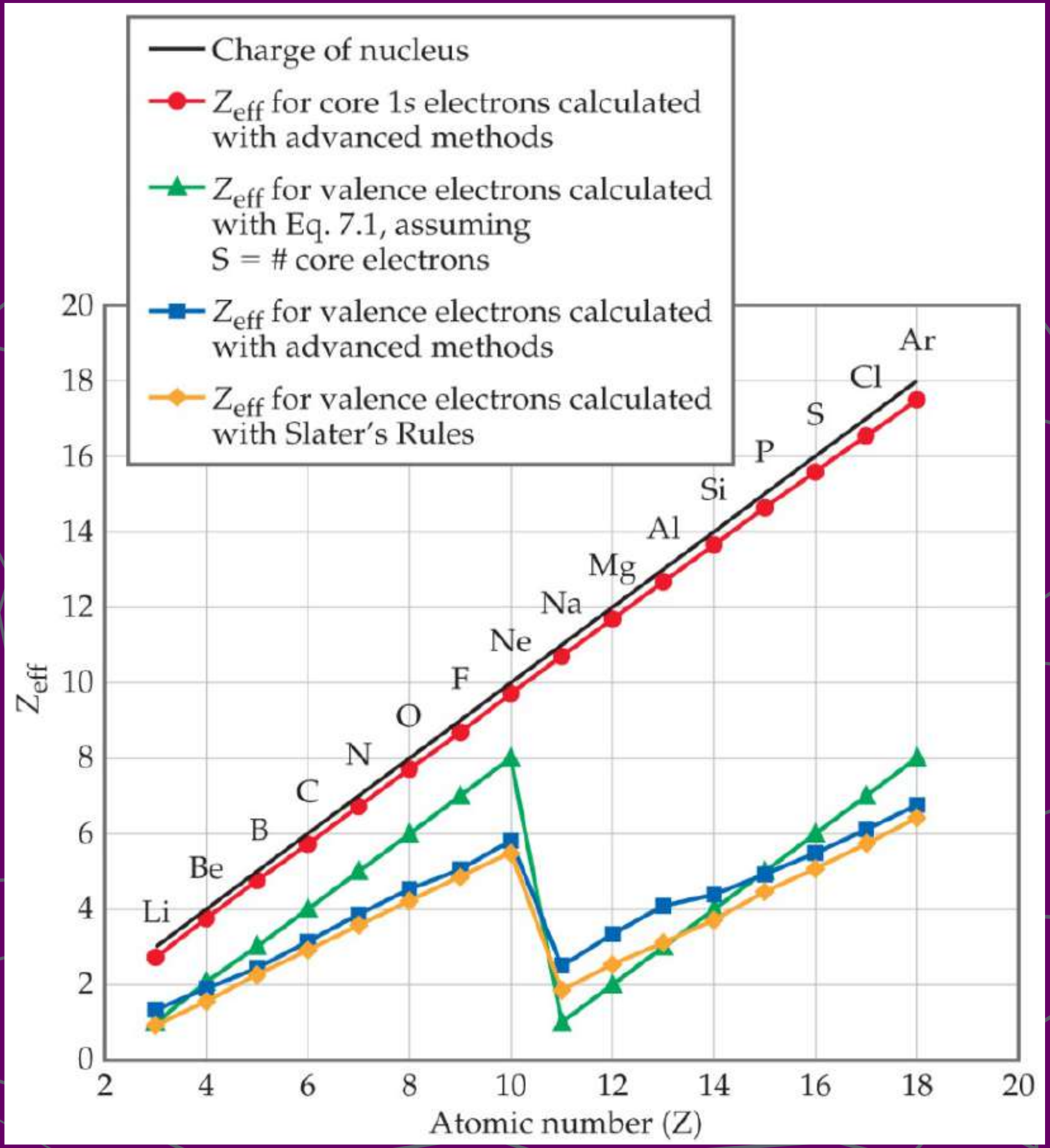
where Z is the atomic number and S is a screening constant, usually close to the number of inner electrons.

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

Based on this figure, is it possible for an electron in a 2*s* orbital to be closer to the nucleus than an electron in a 1*s* orbital?



Effective Nuclear Charge Increases across a Period



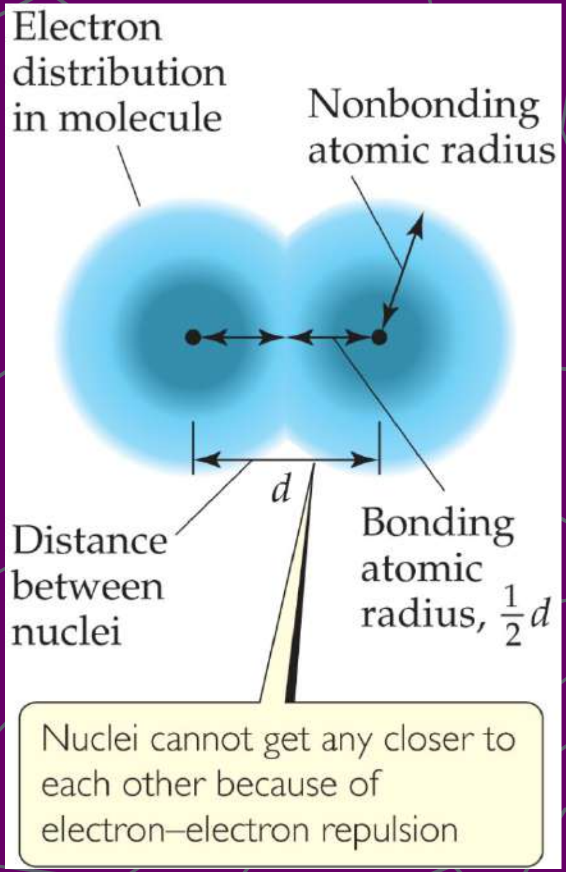
7.2 Give It Some Thought

- Which would you expect to experience a greater effective nuclear charge: a 2*p* electron of a Ne atom or a 3*s* electron of a Na atom?

7.3 Sizes of Atoms and Ions

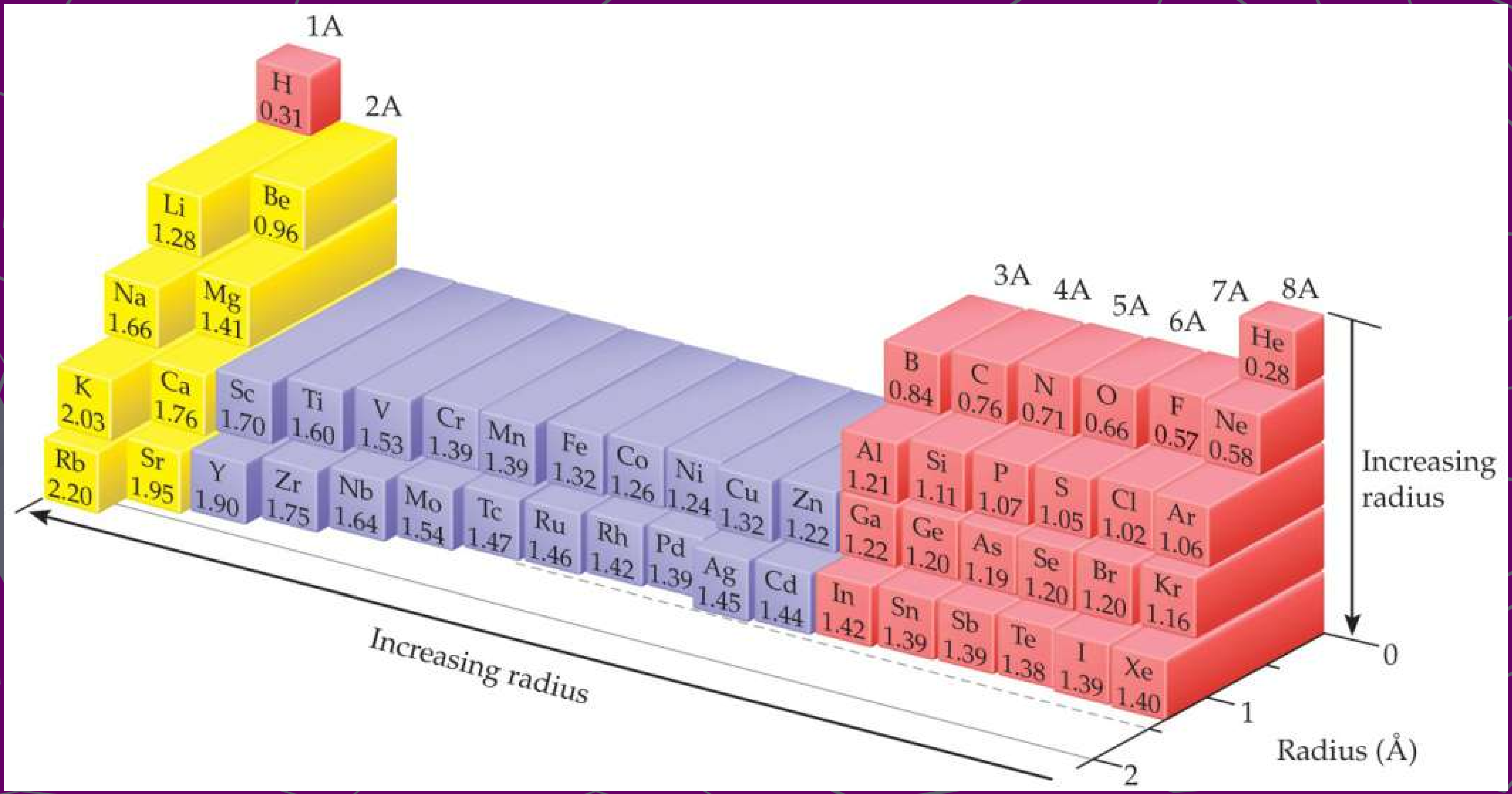
What Is the Size of an Atom?

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

- Which part of the periodic table (top/bottom, left/right) has the elements with the largest atoms?



Sample Exercise 7.1

- 1) Natural gas used in home heating and cooking is odorless. Because natural gas leaks pose the danger of explosion or suffocation, various smelly substances are added to detect leaks. One substance is methyl mercaptan, CH_3SH. Use Figure 7.6 to predict the lengths of the C-S, C-H, and S-H bonds in the molecule.
- 2) Use Figure 7.7 to predict which will be greater: the P-Br bond length in PBr_3 or the As-Cl bond length in $AsCl_3$.

Practice Exercise 1

Hypothetical elements X and Y form a molecule XY_2, in which both Y atoms are bonded to atom X (and not to one another). The X—X distance in the elemental form of X is 2.04 Å, and the Y—Y distance in elemental Y is 1.68 Å. What would you predict for the X—Y distance in the XY_2 molecule?

(a) 0.84 Å

(b) 1.02 Å

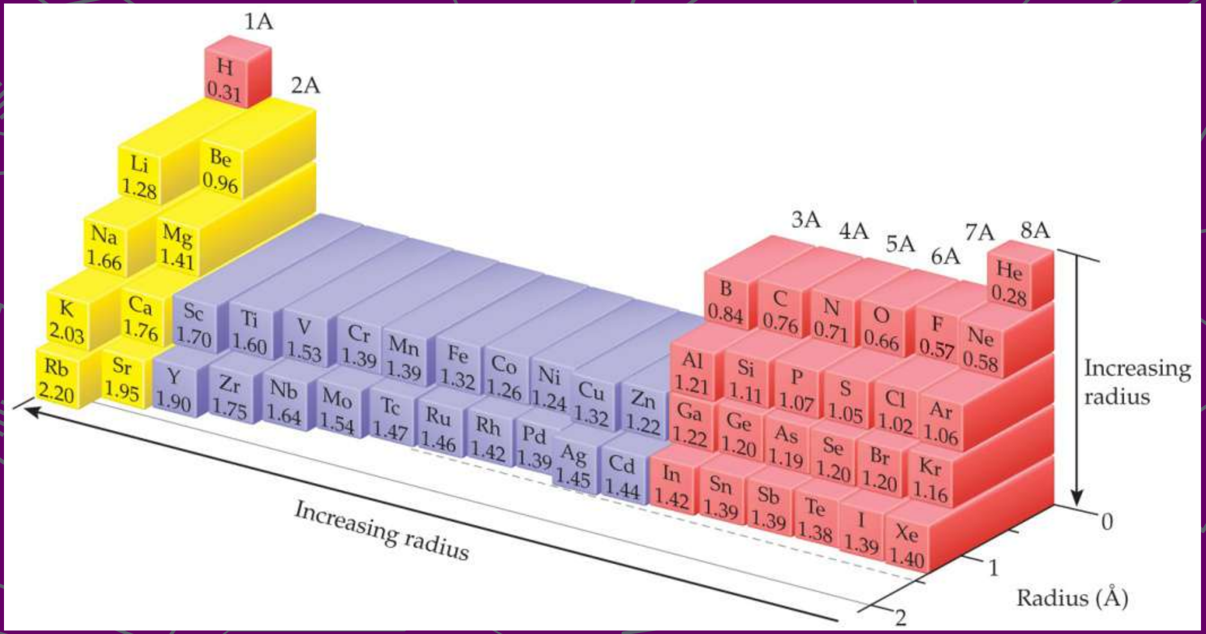
(c) 1.86 Å

(d) 2.70 Å

(e) 3.72 Å

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 ↑).



Sample Exercise 7.2

- Referring to the periodic table, arrange (as much as possible) the atoms B, C, Al, and Si in order of increasing size.
- Arrange Be, C, K, and Ca in order of increasing atomic radius.

Practice Exercise 1

By referring to the periodic table, but not to Figure 7.7, place the following atoms in order of increasing bonding atomic radius: N, O, P, Ge.

(a) N < O < P < Ge

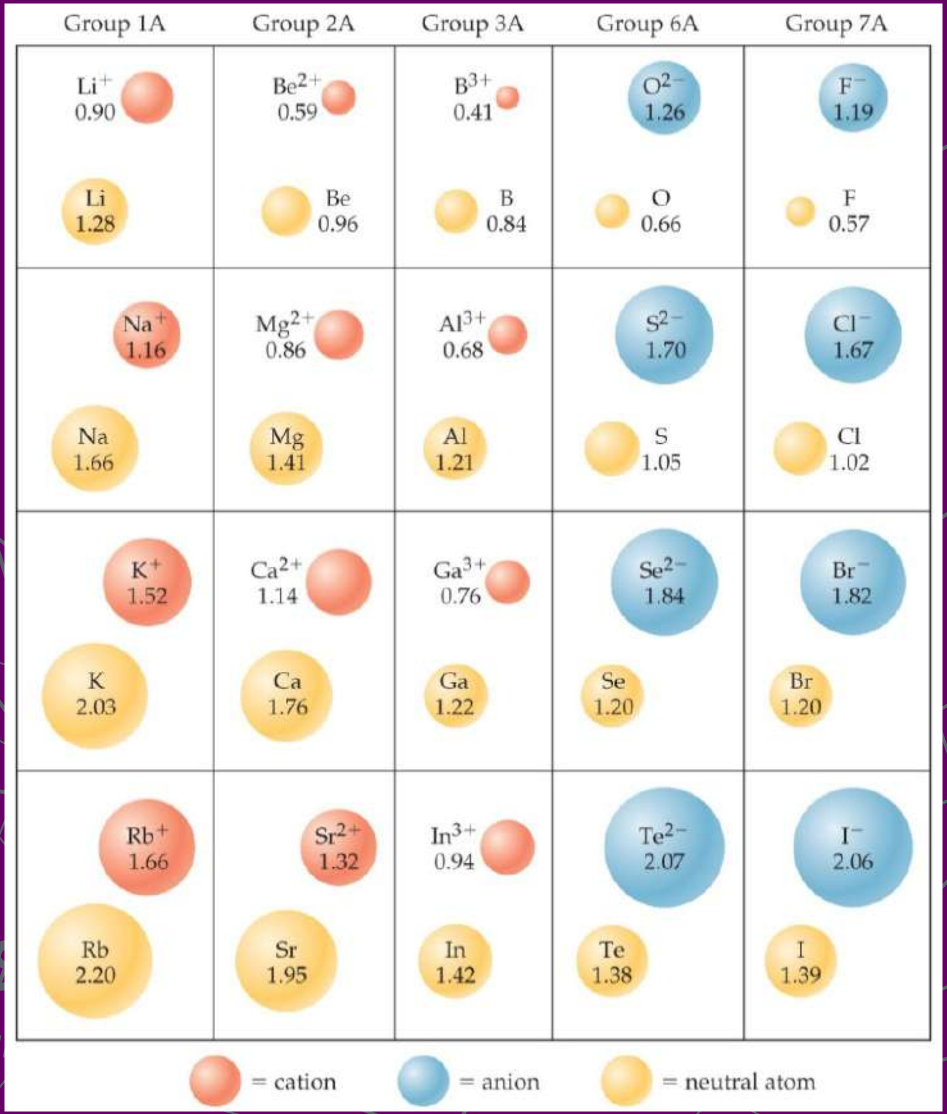
(b) P < N < O < Ge

(c) O < N < Ge < P

(d) O < N < P < Ge

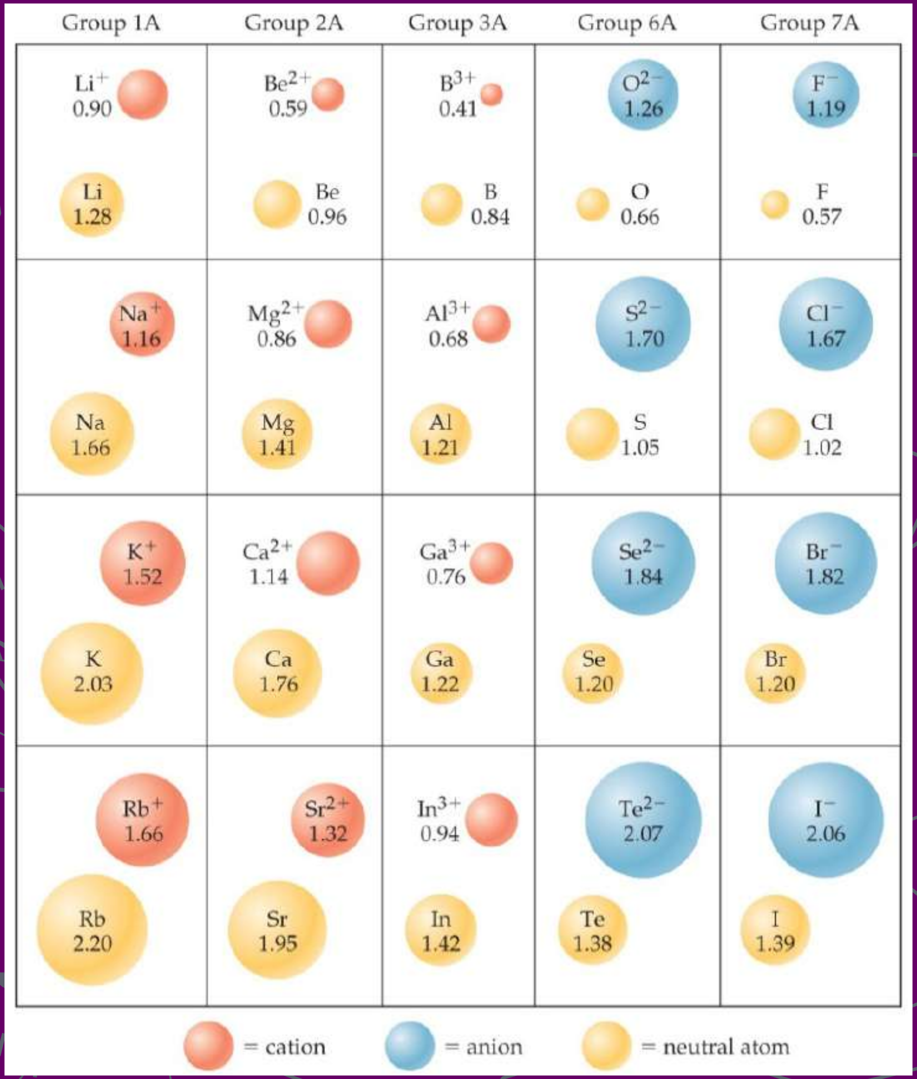
(e) N < P < Ge < O

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

Sample Exercise 7.3 Atomic and Ionic Radii

Arrange these atoms and ions in order of decreasing radius: Mg^{2+}, Ca^{2+}, and Ca.

Practice Exercise

Which of the following atoms and ions is largest: S^{2-}, S, O^{2-}?

Practice Exercise 1

Arrange the following atoms and ions in order of increasing ionic radius: F^-, S^{2-}, Cl^-, and Se^{2-}.

(a) $F^- < S^{2-} < Cl^- < Se^{2-}$

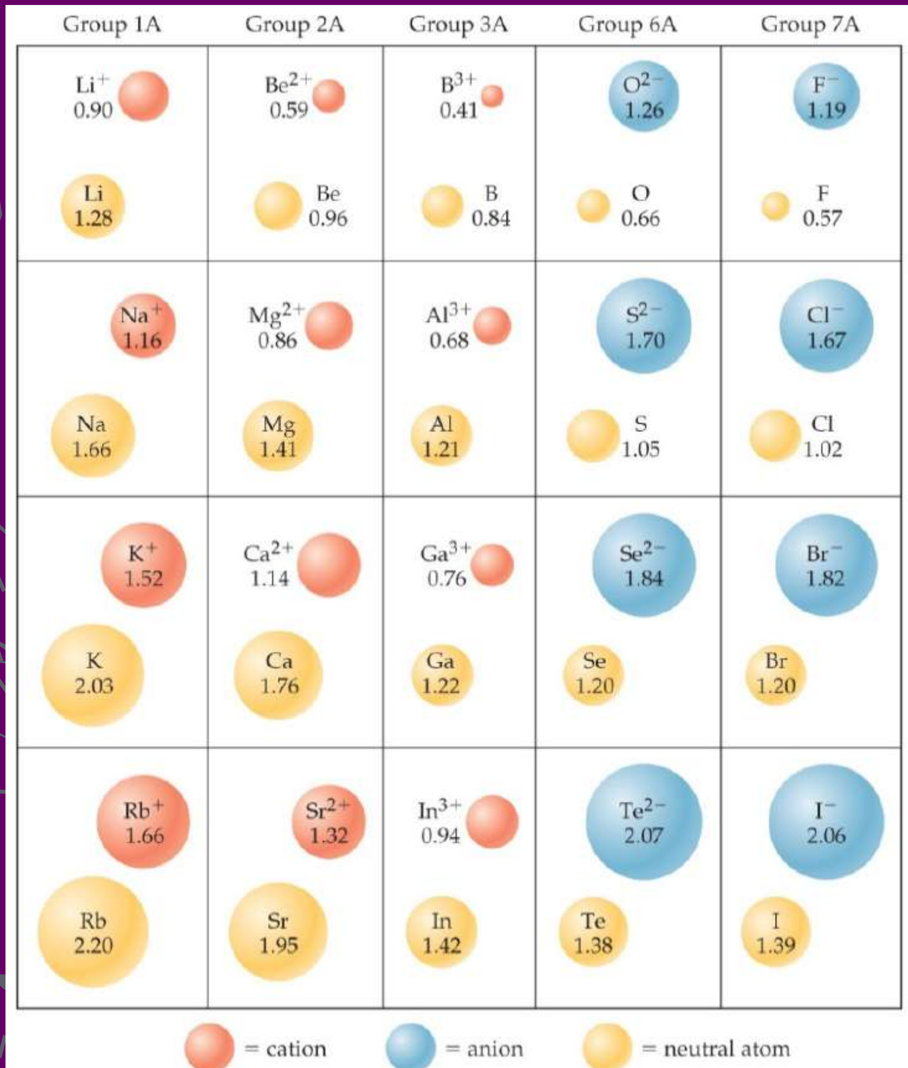
(b) $F^- < Cl^- < S^{2-} < Se^{2-}$

(c) $F^- < S^{2-} < Se^{2-} < Cl^-$

(d) $Cl^- < F^- < Se^{2-} < S^{2-}$

(e) $S^{2-} < F^- < Se^{2-} < Cl^-$

Sizes of Ions



- How do cations of the same charge change in radius as you move down a column in the periodic table?

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

Arrange the ions K^+, Cl^-, Ca^{2+}, and S^{2-} in order of decreasing size.

In the isoelectronic series Ca^{2+}, Cs^+, Y^{3+}, which ion is largest?

Practice Exercise 1

Arrange the following atoms and ions in order of increasing ionic radius: Br^-, Rb^+, Se^{2-}, Sr^{2+}, Te^{2-}.

(a) $Sr^{2+} < Rb^+ < Br^- < Se^{2-} < Te^{2-}$

(b) $Br^- < Sr^{2+} < Se^{2-} < Te^{2-} < Rb^+$

(c) $Rb^+ < Sr^{2+} < Se^{2-} < Te^{2-} < Br^-$

(d) $Rb^+ < Br^- < Sr^{2+} < Se^{2-} < Te^{2-}$

(e) $Sr^{2+} < Rb^+ < Br^- < Te^{2-} < Se^{2-}$

7.4 Ionization Energy

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

7.4 Give It Some Thought

- Light can be used to ionize atoms and ions as shown in these two equations:

$Na(g) \rightarrow Na^{+}(g) + e^{-}$

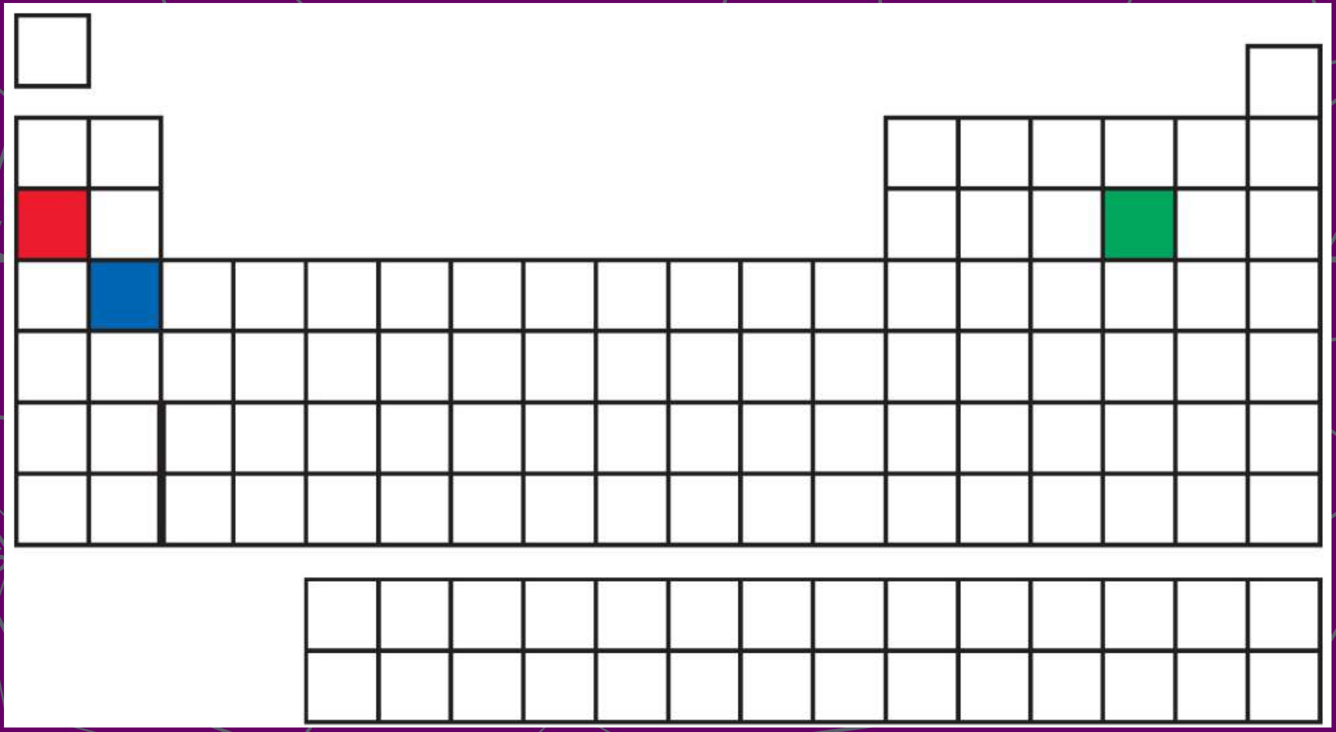
$Na^{+}(g) \rightarrow Na^{2+}(g) + e^{-}$

Which would require shorter wavelength radiation?

- Which would you expect to be greater: I_1 for a boron atom or I_2 for a carbon atom?

Sample Exercise 7.5 Trends in Ionization Energy

Three elements are indicated in the periodic table. Based on their locations, predict the one with the largest second ionization energy.



Practice Exercise 1

The third ionization energy of bromine is the energy required for which of the following processes?

(a) $Br(g) \rightarrow Br^{+}(g) + e^{-}$

(b) $Br^{+}(g) \rightarrow Br^{2+}(g) + e^{-}$

(c) $Br(g) \rightarrow Br^{2+}(g) + 2e^{-}$

(d) $Br(g) \rightarrow Br^{3+}(g) + 3e^{-}$

(e) $Br^{2+}(g) \rightarrow Br^{3+}(g) + e^{-}$

Practice Exercise 2

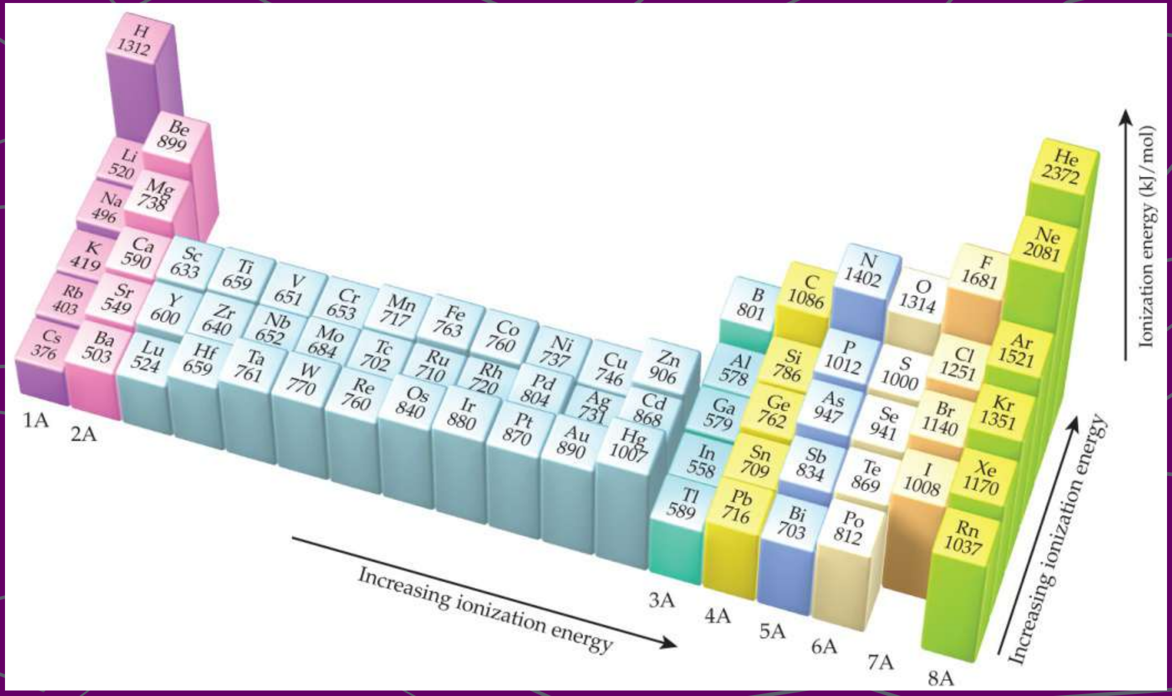
Which has the greater third ionization energy, Ca or S?

Periodic Trends in First Ionization Energy (I_1)

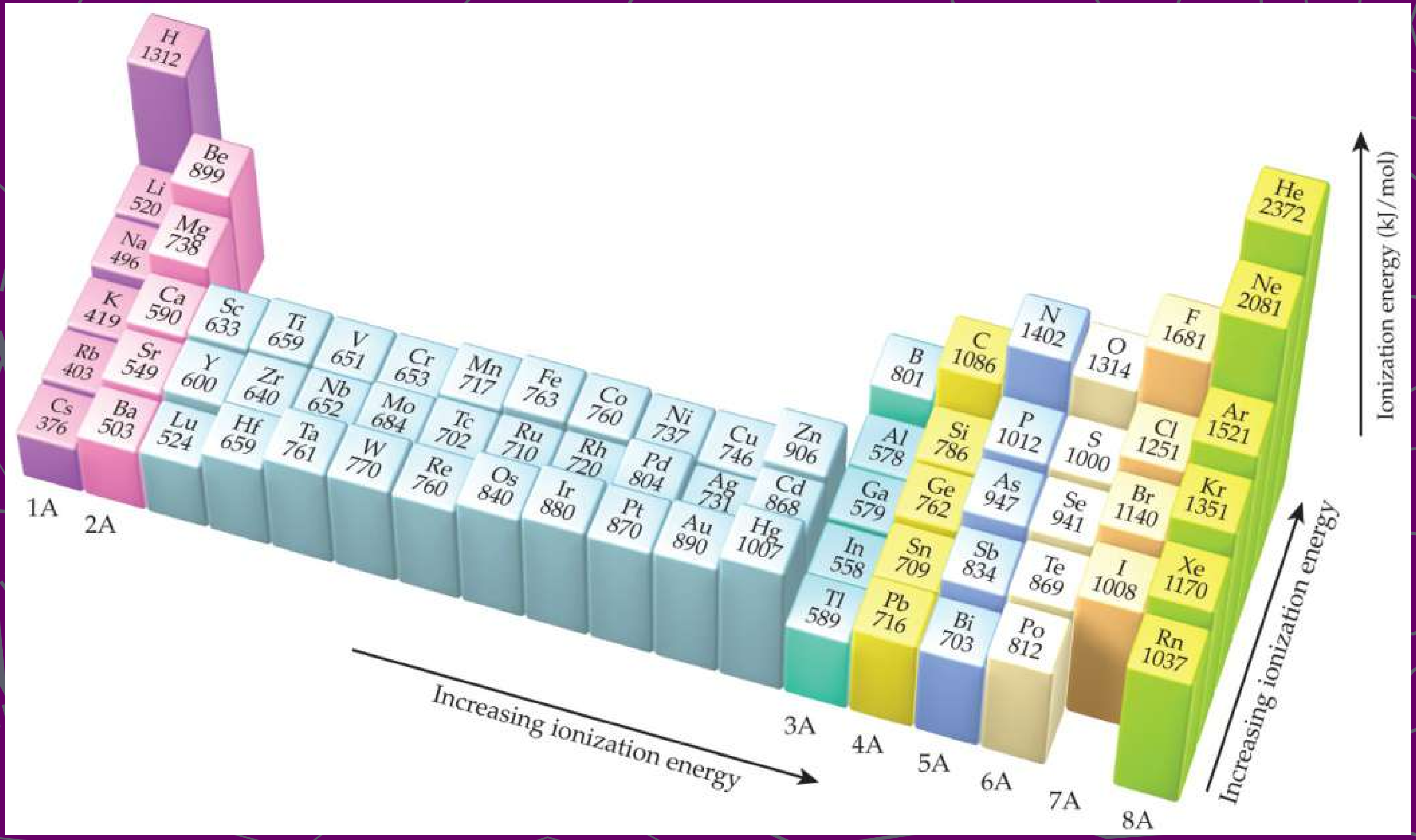
1) I_1 generally increases across a period.
2) I_1 generally decreases down a group.
3) The *s*- and *p*-block elements show a larger range of values for I_1. (The *d*-block generally increases slowly across the period; the *f*-block elements show only small variations.)

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.

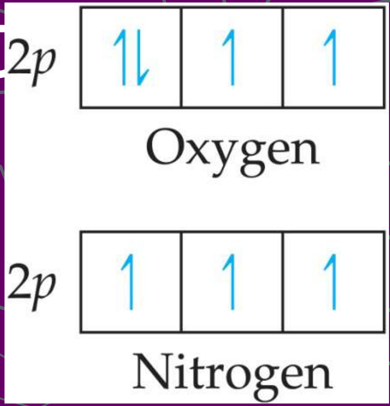


The value for astatine, At, is missing in this figure. To the nearest 100 kJ/mol, what estimate would you make for the first ionization energy of At?

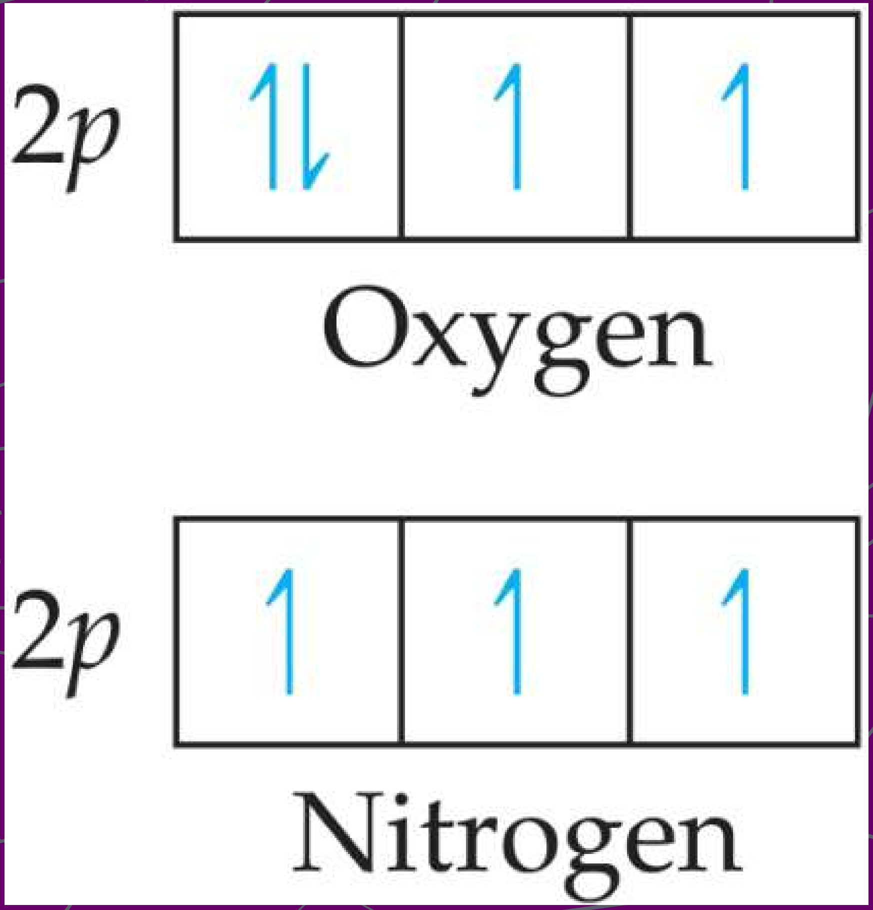


Irregularities in the General Trend

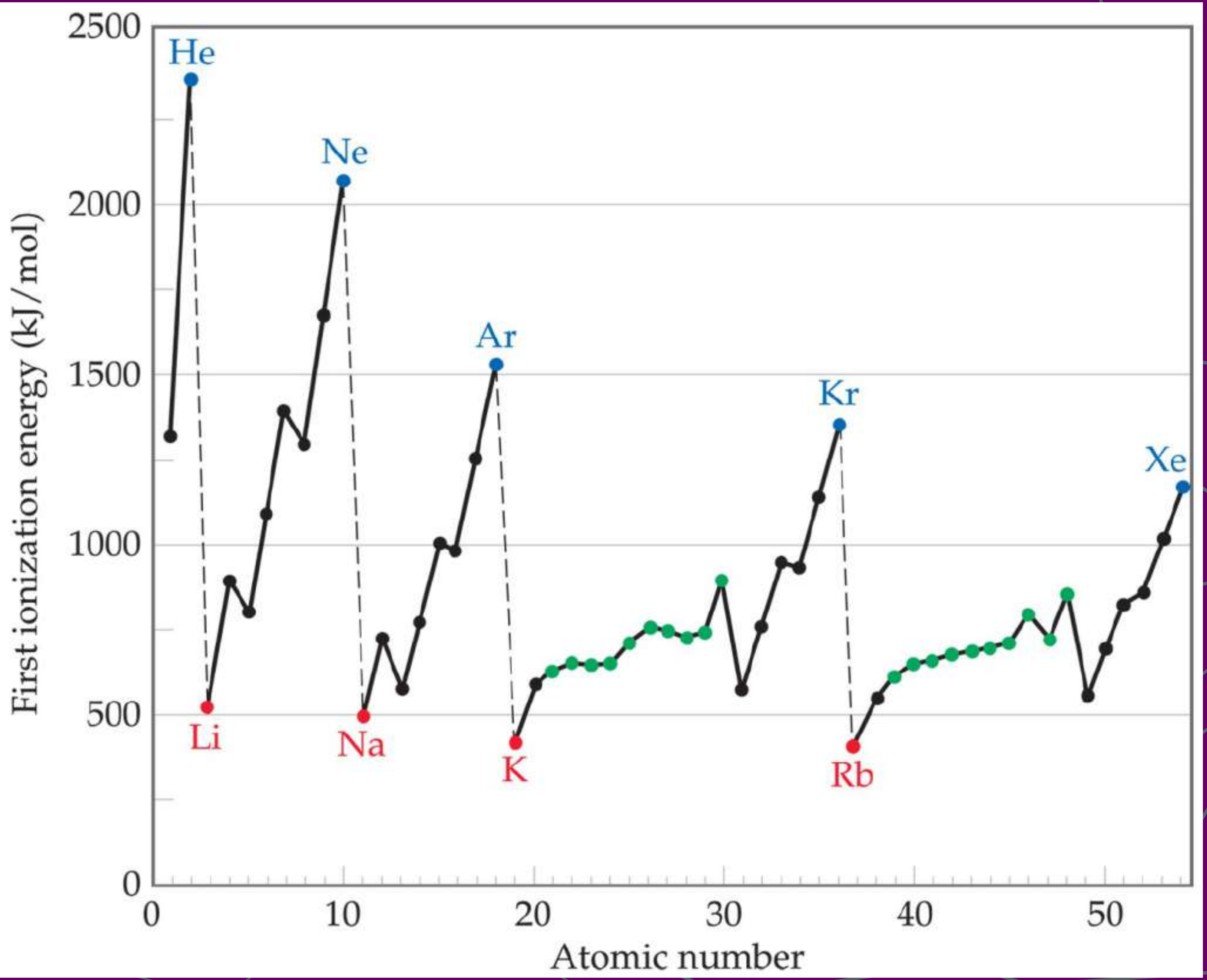
- The trend is **not** followed when the added valence electron in the next element
  - enters a new sublevel (higher energy sublevel);
  - is the first electron to pair in one orbital of the sublevel (elect sions lower energy).



Why is it easier to remove a 2*p* electron from an oxygen atom than from a nitrogen atom?



Trends in First Ionization Energies



Sample Exercise 7.6 Periodic Trends in Ionization Energy

Referring to a periodic table, arrange the following atoms in order of increasing first ionization energy: Ne, Na, P, Ar, K.

Practice Exercise

Which has the lowest first ionization energy, B, Al, C, or Si? Which has the highest first ionization energy?

Practice Exercise 1

Consider the statements about first ionization energies:

(i) Because the effective nuclear charge for Mg is greater than that for Be, the first ionization energy of Mg is greater than that of Be.

(ii) The first ionization energy of O is less than that of N because in O we must pair electrons in the 2*p* orbitals.

(iii) The first ionization energy of Ar is less than that of Ne because a 3*p* electron in Ar is farther from the nucleus than a 2*p* electron in Ne.

Which of the statements (i), (ii), and (iii) is or are true?

(a) Only one of the statements is true.

(b) Statements (i) and (ii) are true.

(c) Statements (i) and (iii) are true.

(d) Statements (ii) and (iii) are true.

(e) All three statements are true.

Electron Configurations of Ions

- Cations: The electrons are lost from the *highest* energy level (*n* value).
  - Li^+ is $1s^2$ (losing a $2s$ electron).
  - Fe^{2+} is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ (losing two $4s$ 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 $2p$).

Would Cr^{3+} and V^{2+} have the same or different electron configurations?

Sample Exercise 7.7 Electron Configurations of Ions

Write the electron configuration for:

(a) Ca^{2+}

(b) Co^{3+}

(c) S^{2-}.

Practice Exercise 1

The ground electron configuration of a Tc atom is $[Kr]5s^2 4d^5$.

What is the electron configuration of a Tc^{3+} ion?

(a) $[Kr]4d^4$

(b) $[Kr]5s^2 4d^2$

(c) $[Kr]5s^1 4d^3$

(d) $[Kr]5s^2 4d^8$

(e) $[Kr]4d^{10}$

Practice Exercise 2

Write the electron configurations for:

(a) Ga^{3+}

(b) Cr^{3+}

(c) Br^{-}.

7.5 Electron Affinity

Electron Affinity

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

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

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

Why are the electron affinities of the Group 4A elements more negative than those of the Group 5A elements?

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

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!

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
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Rb −47	**Sr** −5	**In** −30	**Sn** −107	**Sb** −103	**Te** −190	**I** −295	**Xe** > 0

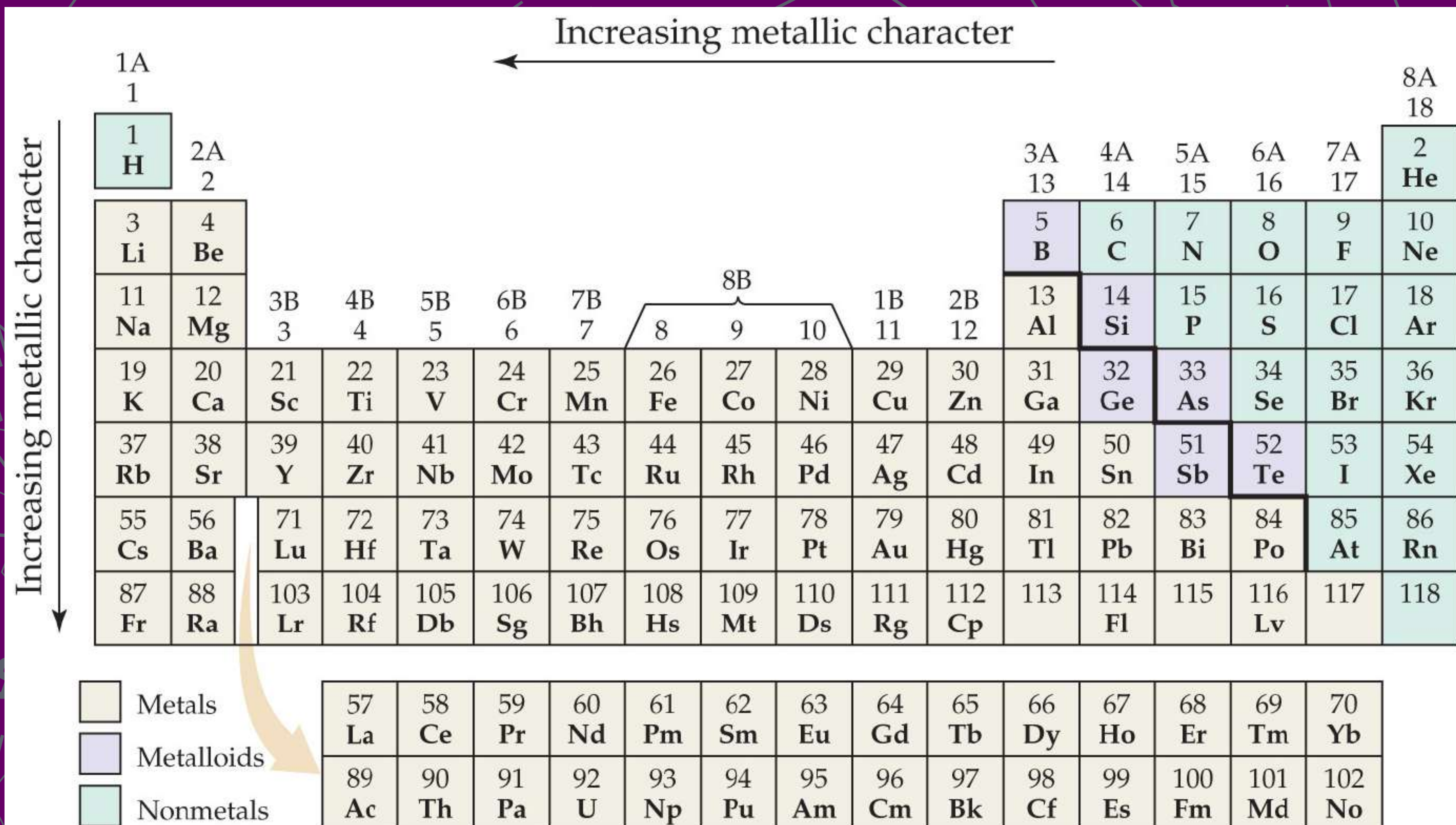
Note: the electron affinity for many of these elements is *positive* (X^- is unstable).

Section 7.5 Give It Some Thought

- What is the relationship between the value for the first ionization energy of a $Cl^{-}(g)$ ion and the electron affinity of $Cl(g)$?

7.6 Metals, Nonmetals, and Metalloids

How do the periodic trends in metallic character compare to those for ionization energy?



7.6 Give It Some Thought

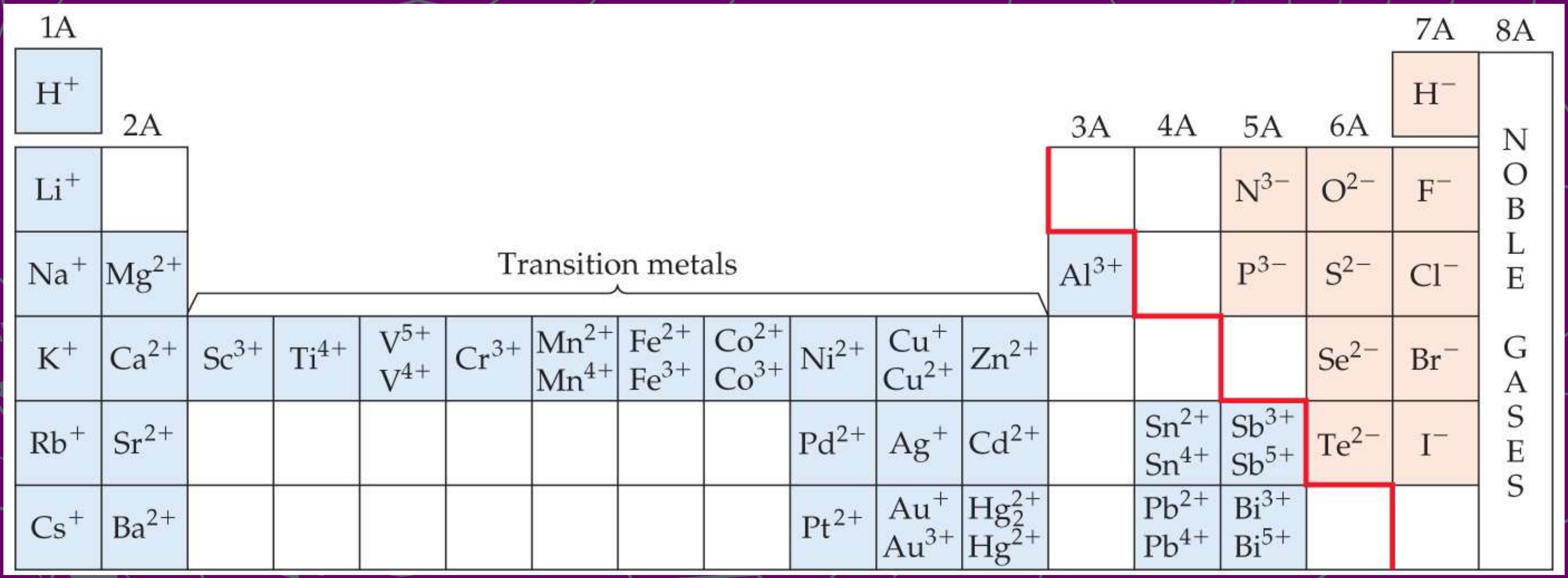
- Arsenic forms binary compounds with Cl and Mg. Will it be in the same oxidation sate in these two compounds?

Metals Differ from 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+}			

The red stepped line divides metals from nonmetals. How are common oxidation states divided by this line?



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 $Ni(NO_3)_2$

Would you expect NiO to dissolve in an aqueous solution of $NaNO_3$?

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 $Ni(NO_3)_2$

*****Reaction Trends*****

- Metal oxide + water → metal hydroxide
  - Ex: $Na_2O(s) + H_2O(l) \rightarrow 2\ NaOH(aq)$
- Metal oxide + acid → salt + water
  - Ex: $NiO(s) + 2\ HNO_3(aq) \rightarrow Ni(NO_3)_2(aq) + H_2O(l)$

Sample Exercise 7.8 Metal Oxides

(a) Would you expect scandium oxide to be a solid, liquid, or gas at room temperature? **(b)** Write the balanced chemical equation for the reaction of scandium oxide with nitric acid.

Practice Exercise

Write the balanced chemical equation for the reaction between copper(II) oxide and sulfuric acid.

Practice Exercise 1

Suppose that a metal oxide of formula M_2O_3 were soluble in water. What would be the major product or products of dissolving the substance in water?

(a) $MH_3(aq) + O_2(g)$

(b) $M(s) + H2(g) + O_2(g)$

(c) $M^{3+}(aq) + H_2O_2(aq)$

(d) $M(OH)_2(aq)$

(e) $M(OH)_3(aq)$

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.

*****Reaction Trends*****

- Nonmetal oxide + water → acid
  - Ex: $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$
- Nonmetal oxide + base → salt + water
  - $CO_2(g) + 2\ NaOH(aq) \rightarrow Na_2CO_3(aq) + H_2O(l)$

A compound ACl_3 (A is an element) has a melting point of -112°C. Would you expect it to be molecular or ionic? If you were told that element A was either scandium or phosphorus, which do you think is more likley?

Sample Exercise 7.9 Nonmetal Oxides

Write the balanced chemical equations for the reactions of solid selenium dioxide with **(a)** water, **(b)** aqueous sodium hydroxide.

Practice Exercise 1

Consider the following oxides: SO_2, Y_2O_3, MgO, Cl_2O, N_2O_5. How many are expected to form acidic solutions in water?

(a) 1

(b) 2

(c) 3

(d) 4

(e) 5

Practice Exercise 2

Write a balanced chemical equation for the reaction of solid tetraphosphorus hexoxide with water.

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).

7.7 Trends for Group 1A and Group 2A Metals

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.

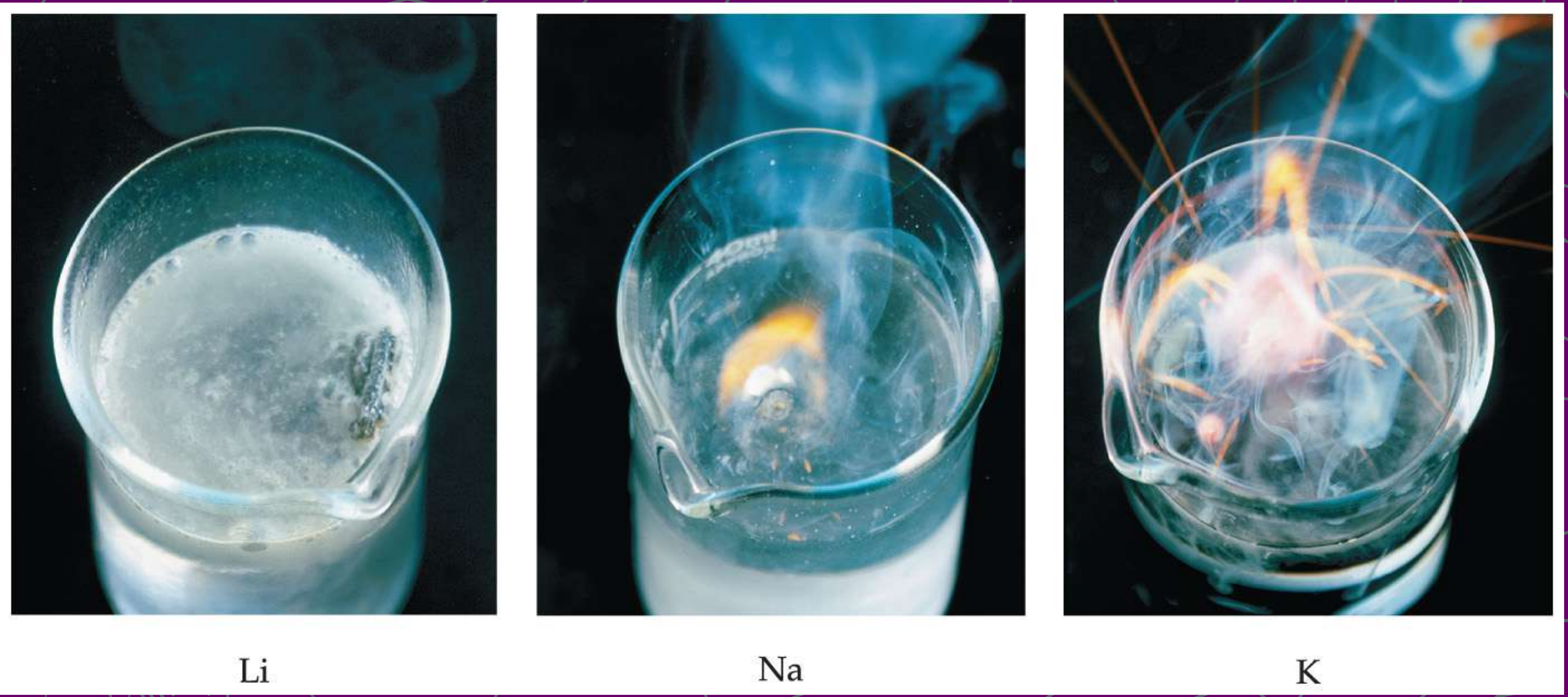
Alkali Metal Properties

- They have low densities and melting points.
- They also have low ionization energies.

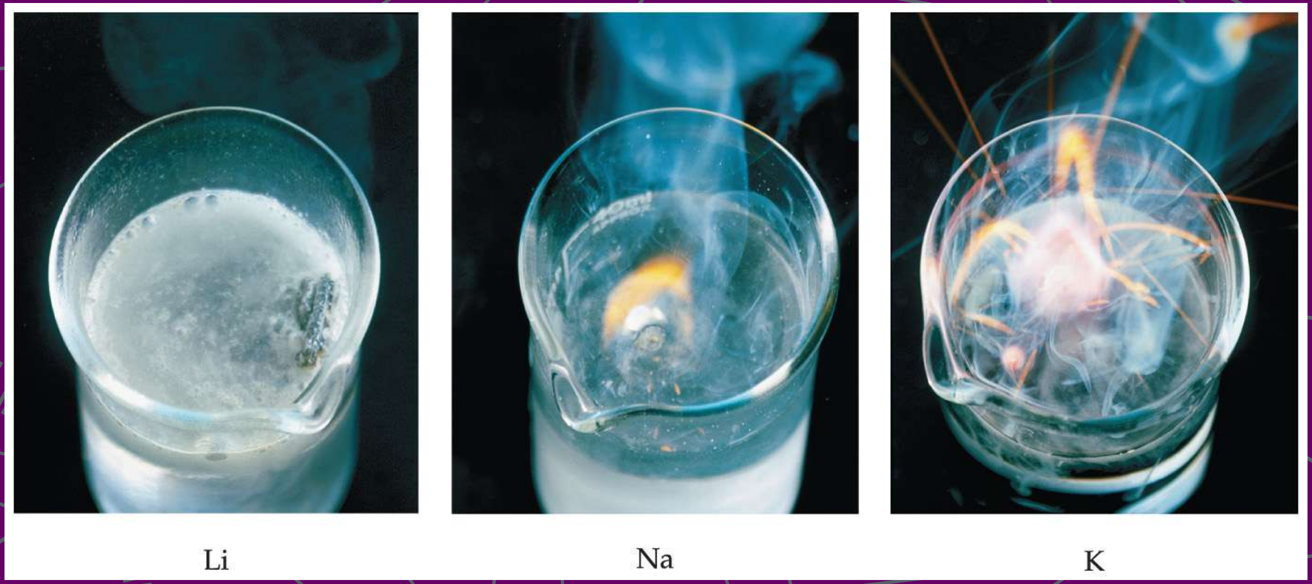
Table 7.4 Some Properties of the Alkali Metals

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

Would you expect rubidium metal to be more or less reactive with water than potassium metal?



Alkali Metal Chemistry



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:

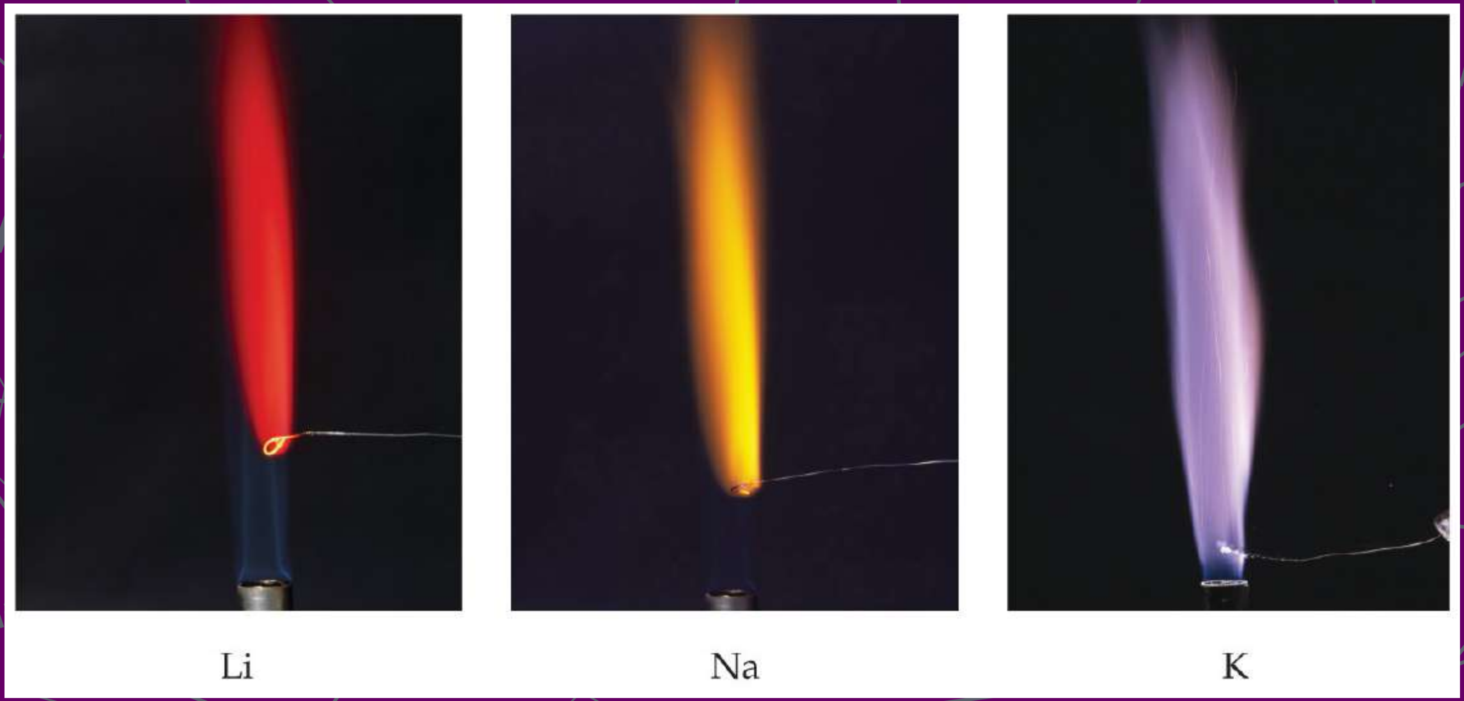
$$M + O_2 \longrightarrow MO_2$$

Alkali Metal Reaction Trend

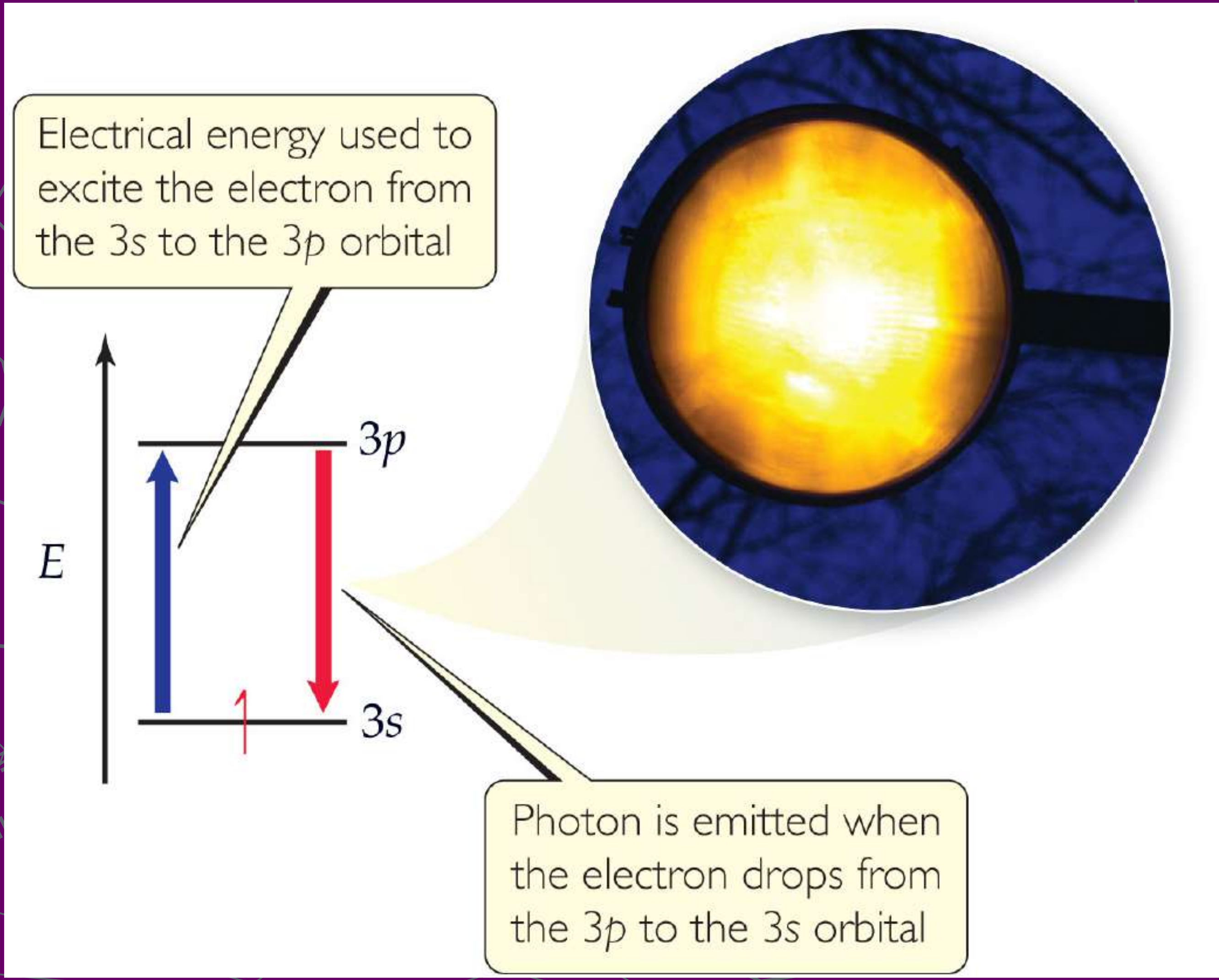
- Alkali metals react with water producing hydrogen gas and a solution of the alkali metal hydroxide.

Flame Tests

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



► If we had potassium vapor lamps, what color would they be?



7.7 Give It Some Thought

- Cesium tends to be the most reactive of the stable alkali metals (francium, Fr, is radioactive and has not been extensively studied). What *atomic* property of Cs is the most responsible for its high reactivity?

Sample Exercise 7.10

- 1) Write a balanced equation that predicts the reaction of cesium metal with:
  - A) $Cl_{2(g)}$
  - B) $H_2O_{(l)}$
  - C) $H_{2(g)}$

Practice Exercise 1

Consider the following three statements:

(i) Based on their positions in the periodic table, the expected product is the ionic oxide M_2O.

(ii) Some of the alkali metals produce metal peroxides or metal superoxides when they react with oxygen.

(iii) When dissolved in water, an alkali metal oxide produces a basic solution.

Which of the statements (i), (ii), and (iii) is or are true?

(a) Only one of the statements is true.

(b) Statements (i) and (ii) are true.

(c) Statements (i) and (iii) are true.

(d) Statements (ii) and (iii) are true.

(e) All three statements are true.

Practice Exercise 2

Write a balanced equation for the expected reaction between potassium metal and elemental sulfur, S(*s*).

Alkaline Earth Metals—Compare to Alkali Metals

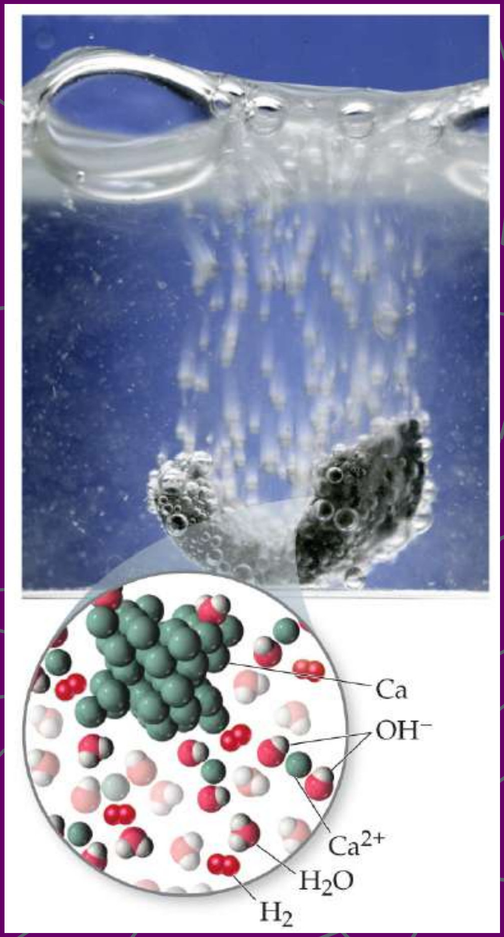
Table 7.5 Some Properties of the Alkaline Earth Metals

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.

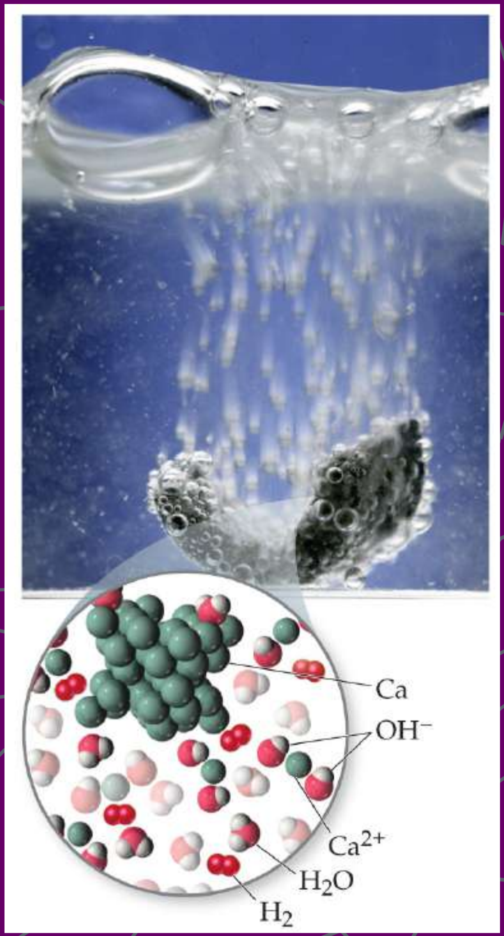


Alkaline Earth Metal Reaction Trends

- Beryllium or magnesium + H_2O yield metal oxide + hydrogen gas
- Calcium and elements below + H_2O yield metal hydroxide + hydrogen gas

Alkaline Earth Metals

- What is the cause of the bubbles that are formed? How could you test your answer?



7.7 Give It Some Thought

- Calcium carbonate, $CaCO_3$, is often used as a dietary calcium supplement for bone health. Although it is insoluble in water, it can be taken orally to allow for the delivery of $Ca^{2+}(aq)$ ions to the musculoskeletal system. Why is this the case? (*Hint:* Recall the reactions of metal carbonates from section 4.3).

7.8 Trends for Selected Nonmetals

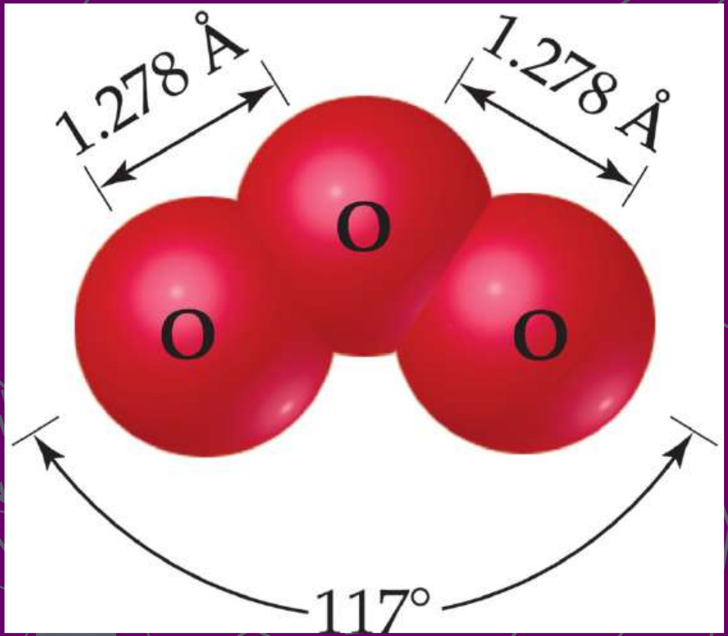
Group 6A—Increasing in Metallic Character down the Group

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.

Oxygen

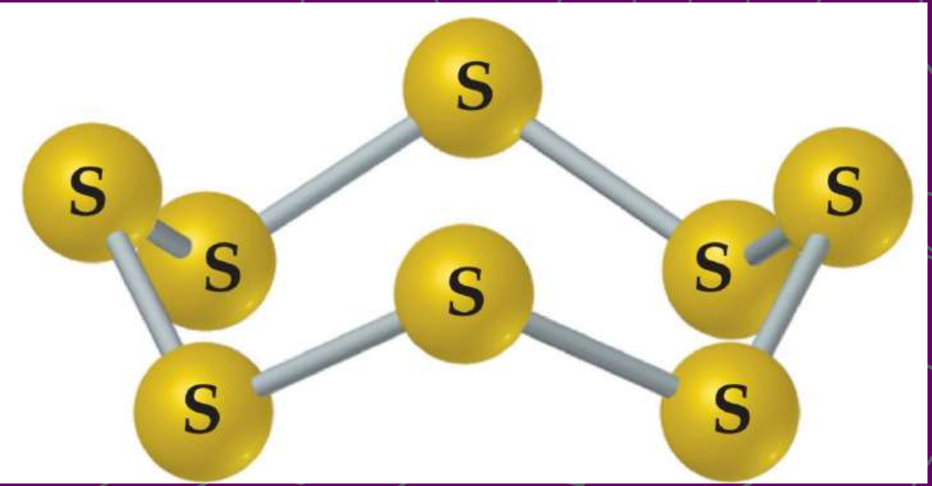


- There are two allotropes of oxygen:
  - O_2
  - O_3, ozone
- There can be three anions:
  - O^{2-}, oxide
  - O_2^{2-}, peroxide
  - O_2^{1-}, superoxide
- It tends to take electrons from other elements (oxidation).

- Why is it okay to store water in a bottle with a normal, nonventing cap?
- Hydrogen peroxide is so light sensitive and so is stored in brown bottles because its O-O bond is relatively weak. If we assume that the brown bottle absorbs all visible wavelengths of light, how might you estimate the energy of the O-O bond in hydrogen peroxide?

Sulfur

- Sulfur is a weaker oxidizer than oxygen.
- The most stable allotrope is S_8, a ringed molecule.
- Suppose it were possible to flatten the S_8 ring. What shape would you expect the flattened ring to have?



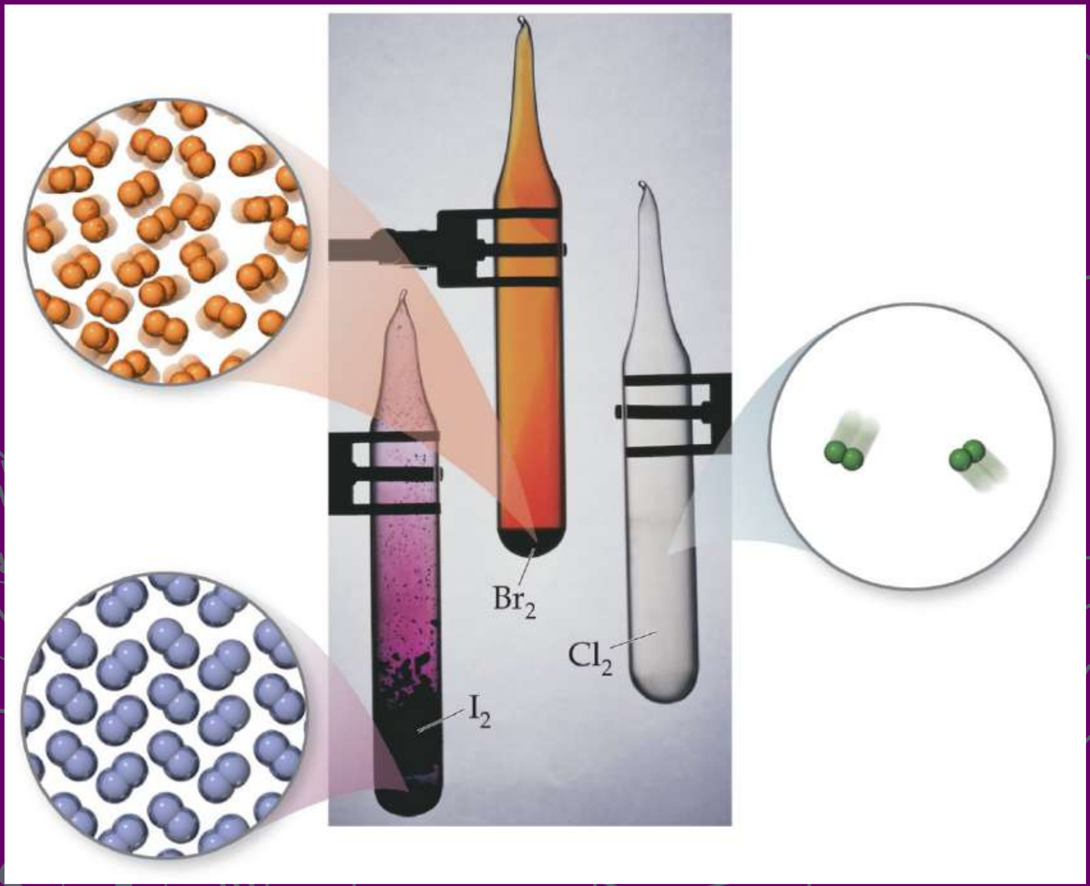
Group 7A—Halogens

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/cm³	1.20	1140
Iodine	$[Kr]5s^2 4d^{10} 5p^5$	114	4.94 g/cm³	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.

Halogens



- The halogens do not exist as X_8 molecules like sulfur and selenium do. Why do you think this is the case?
- Why are more molecules of I_2 seen in the molecular view relative to the number of Cl_2 molecules?

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	$[He]2s^22p^5$	−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

- Estimate the atomic radius and first ionization energy of an astatine atom.

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.

Group 8A—Noble Gases

- 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.

Noble Gases

- Xe forms three compounds:
  - XeF_2
  - XeF_4 (at right)
  - XeF_6
- Kr forms only one stable compound:
  - KrF_2
- The unstable HArF was synthesized in 2000.

Integrative Exercise

- Bismuth is the heaviest member of Group 5A and is used in Pepto Bismol.
  - A) The covalent atomic radii of thallium and lead are 1.48 Å and 1.47 Å respectively. Using this and Figure 7.6, predict the radius of bismuth.
  - B) What accounts for the general increase in atomic radius going down group 5A?
  - C) Bismuth is also used in low-melting alloys such as sprinkler systems. The element itself is a brittle, white crystalline solid. How do these characteristics fit with the fact that bismuth is in the same group as N and P?
  - D) Bi_2O_3 is a basic oxide. Write a balanced equation for its reaction with nitric acid. If 6.77 g of Bi_2O_3 is dissolved in dilute acidic solution to make up to 500 mL of solution, what is the molarity of the solution of Bi^{3+} ion?
  - E) ^{209}Bi is the heaviest stable isotope of any element. How many protons and neutrons are present in the nucleus?
  - F) The density of Bi at 25ºC is 9.808 g/cm^3. How many Bi atoms are present in a cube of the element that is 5.00 cm on each edge? How many moles of the element are present?