Preview

- Lesson Starter
- Objectives
- Mendeleev and Chemical Periodicity
- Moseley and the Periodic Law
- The Modern Periodic Table

Lesson Starter

Share what you have learned previously about the periodic table.

End Of Slide

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Objectives

- **Explain** the roles of Mendeleev and Moseley in the development of the periodic table.
- **Describe** the modern periodic table.
- **Explain** how the periodic law can be used to predict the physical and chemical properties of elements.
- **Describe** how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.

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Mendeleev and Chemical Periodicity

- Mendeleev noticed that when the elements were arranged in order of increasing atomic mass, certain similarities in their chemical properties appeared at regular intervals.
- Repeating patterns are referred to as *periodic.*
- Mendeleev created a table in which elements with similar properties were grouped together—a periodic table of the elements.

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Mendeleev and Chemical Periodicity, *continued*

- After Mendeleev placed all the known elements in his periodic table, several empty spaces were left.
- In 1871 Mendeleev predicted the existence and properties of elements that would fill three of the spaces.
- By 1886, all three of these elements had been discovered.

End of Slide

Properties of Some Elements Predicted By Mendeleev

Predicted elements	Element and year discovered	Properties	Predicted properties	Observed properties
Ekaaluminum	gallium 1875	density of metal	6.0 g/mL	5.96 g/mL
		melting point	low	30°C
		oxide formula	Ea_2O_3	Ga_2O_3
Ekaboron	scandium 1877	density of metal	3.5 g/mL	3.86 g/mL
		oxide formula	Eb_2O_3	Sc_2O_3
		solubility of oxide	dissolves in acid	dissolves in acid
Ekasilicon	germanium 1886	melting point	high	900°C
		density of metal	5.5 g/mL	5.47 g/mL
		color of metal	dark gray	grayish white
		oxide formula	EsO_2	GeO_2
		density of oxide	4.7 g/mL	4.70 g/mL
		chloride formula	$EsCl_4$	$GeCl_4$

Moseley and the Periodic Law

- In 1911, the English scientist Henry Moseley discovered that the elements fit into patterns better when they were arranged according to atomic number, rather than atomic weight.

- **The Periodic Law** states that the physical and chemical properties of the elements are periodic functions of their atomic numbers.

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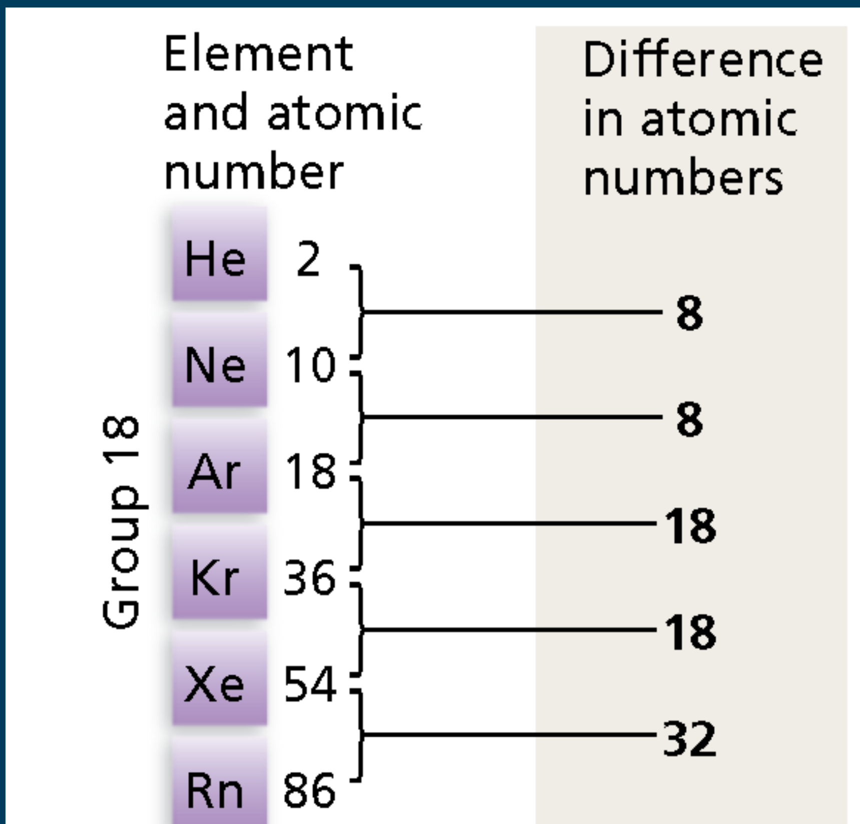
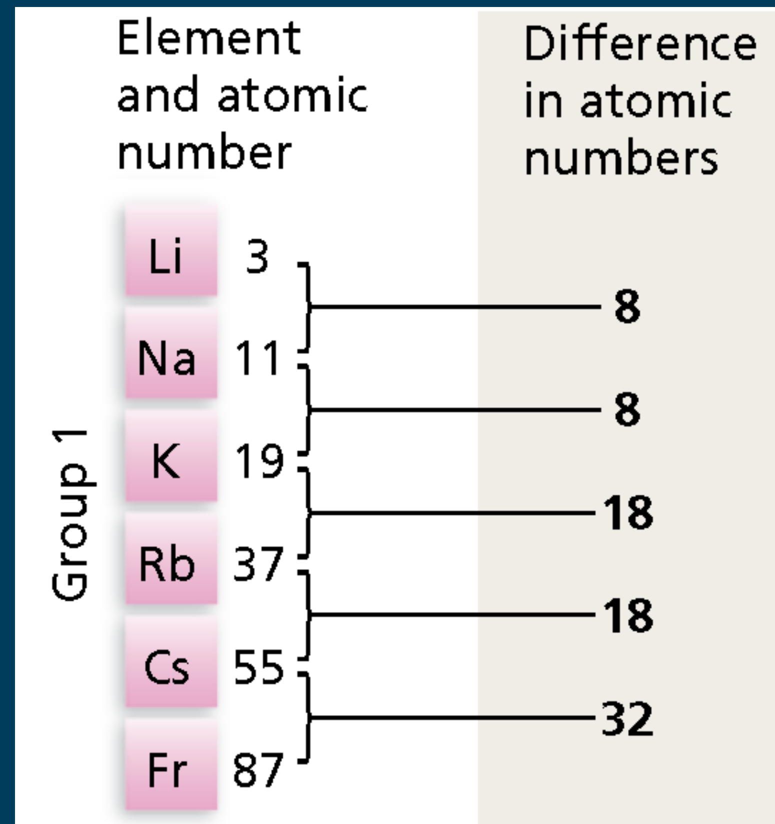
Periodicity of Atomic Numbers





The Modern Periodic Table

- **The Periodic Table** is an arrangement of the elements in order of their atomic numbers so that elements with similar properties fall in the same column, or group.

Periodic Table Overview

Click below to watch the Visual Concept.

Visual Concept

Preview

- Lesson Starter
- Objectives
- Periods and Blocks of the Periodic Table

Lesson Starter

- Name as many properties shared by elements of the same group in the periodic table as possible.
- Describe what you already know about an element just by looking at its position in the periodic table.
- Identify any noticeable trends.

End Of Slide

Objectives

- **Explain** the relationship between electrons in sublevels and the length of each period of the periodic table.

- **Locate** and name the four blocks of the periodic table. Explain the reasons for these names.

Objectives, *continued*

- **Discuss** the relationship between group configurations and group numbers.

- **Describe** the locations in the periodic table and the general properties of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.

Periods and Blocks of the Periodic Table

- Elements are arranged vertically in the periodic table in groups that share similar chemical properties.
- Elements are also organized horizontally in rows, or *periods*.
- The length of each period is determined by the number of electrons that can occupy the sublevels being filled in that period.
- The periodic table is divided into four blocks, the *s, p, d,* and *f* blocks. The name of each block is determined by the electron sublevel being filled in that block.

Relating Period Length and Sublevels Filled

Click below to watch the Visual Concept.

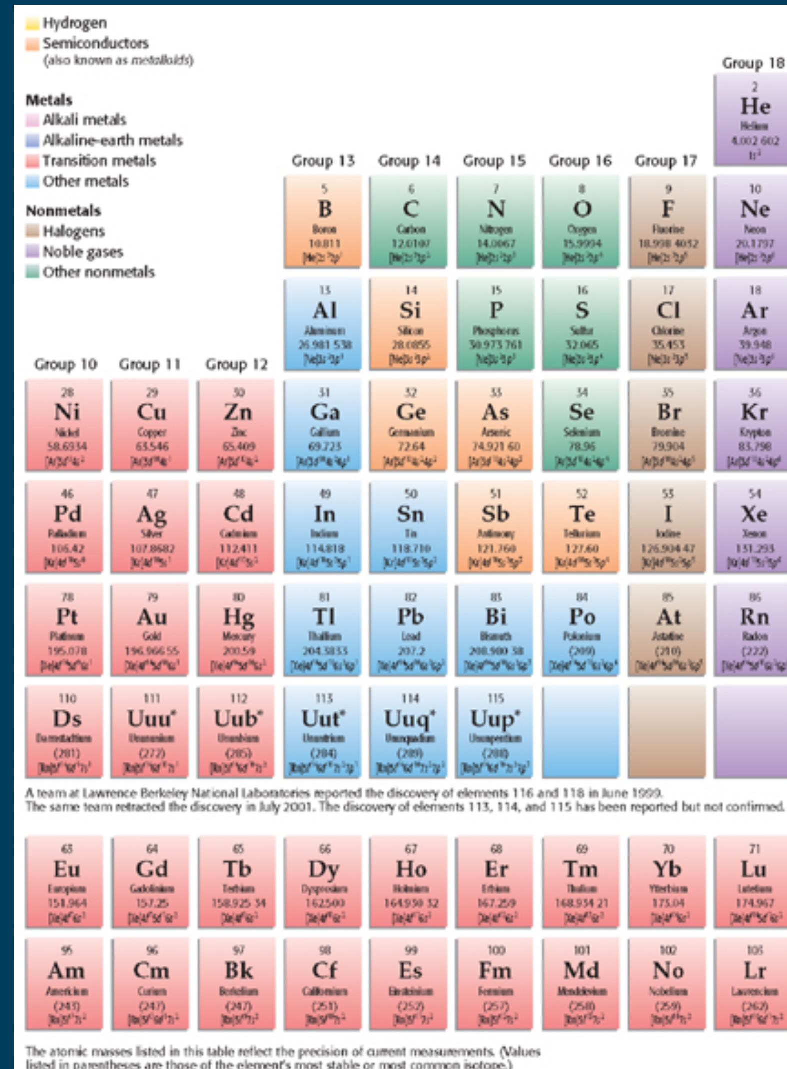
Visual Concept

Blocks of the Periodic Table Based on Sublevel

Click below to watch the Visual Concept.

Visual Concept

Periodic Table of the Elements



Periods and Blocks of the Periodic Table, *continued*

- The elements of Group 1 of the periodic table are known as the **alkali metals.**
  - lithium, sodium, potassium, rubidium, cesium, and francium
  - In their pure state, all of the alkali metals have a silvery appearance and are soft enough to cut with a knife.
- The elements of Group 2 of the periodic table are called the **alkaline-earth metals.**
  - beryllium, magnesium, calcium, strontium, barium, and radium
  - Group 2 metals are less reactive than the alkali metals, but are still too reactive to be found in nature in pure form.

Periods and Blocks of the Periodic Table, *continued*

- Hydrogen has an electron configuration of $1s^1$, but despite the ns^1 configuration, it does not share the same properties as the elements of Group 1.
    - Hydrogen is a unique element.
- Like the Group 2 elements, helium has an ns^2 group configuration. Yet it is part of Group 18.
    - Because its highest occupied energy level is filled by two electrons, helium possesses special chemical stability.

Relationship Between Periodicity and Electron Configurations

Period number	Number of elements in period	Sublevels in order of filling
1	2	1*s*
2	8	2*s* 2*p*
3	8	3*s* 3*p*
4	18	4*s* 3*d* 4*p*
5	18	5*s* 4*d* 5*p*
6	32	6*s* 4*f* 5*d* 6*p*
7	29 (to date)	7*s* 5*f* 6*d*, etc.

Periods and Blocks of the Periodic Table, *continued*

Sample Problem A

a. Without looking at the periodic table, identify the group, period, and block in which the element that has the electron configuration [Xe]6s^2 is located.

b. Without looking at the periodic table, write the electron configuration for the Group 1 element in the third period. Is this element likely to be more reactive or less reactive than the element described in (a)?

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Periods and Blocks of the Periodic Table, *continued*

Sample Problem A Solution

a. The element is in Group 2, as indicated by the group configuration of ns^2.

It is in the sixth period, as indicated by the highest principal quantum number in its configuration, 6.

The element is in the *s* block.

End Of Slide

Periods and Blocks of the Periodic Table, *continued*

Sample Problem A Solution, *continued*

b. In a third-period element, the highest occupied energy level is the third main energy level, $n = 3$. The 1*s*, 2*s*, and 2*p* sublevels are completely filled.

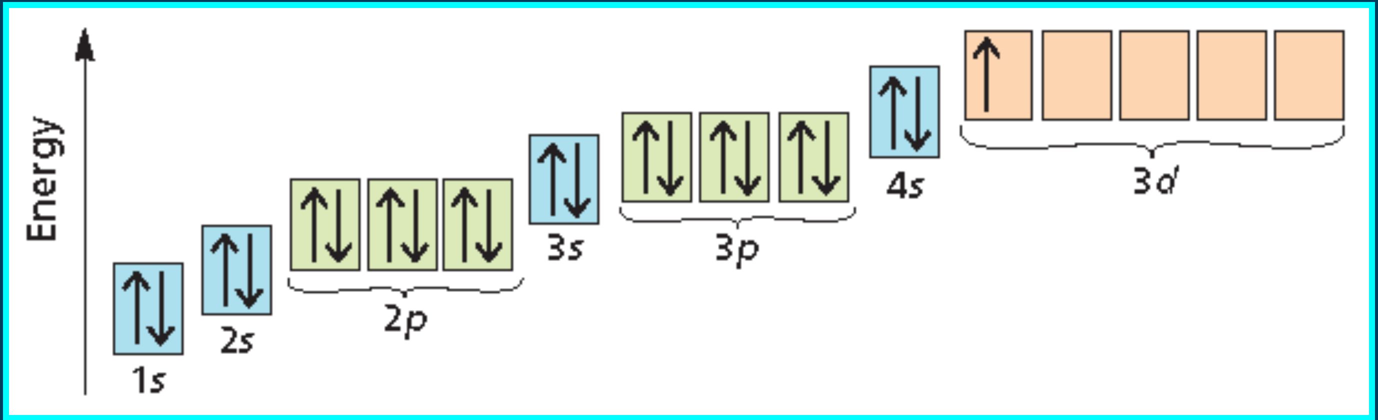
This element has the following configuration:

$$1s^2 2s^2 2p^6 3s^1 \text{ or } [\text{Ne}]3s^1$$

Because it is in Group 1, this element is likely to be more reactive than the element described in (a), which is in Group 2.

Periods and Blocks of the Periodic Table

- The *d* sublevel first appears when *n* = 3.
- The 3*d* sublevel is slightly higher in energy than the 4*s* sublevel, so these are filled in the order 4*s*3*d*.



- The *d*-block elements are metals with typical metallic properties and are often referred to as **transition elements.**

End Of Slide

Periods and Blocks of the Periodic Table, *continued*

Sample Problem B

An element has the electron configuration [Kr]$4d^5 5s^1$. Without looking at the periodic table, identify the period, block, and group in which this element is located. Then, consult the periodic table to identify this element and the others in its group.

Periods and Blocks of the Periodic Table, *continued*

Sample Problem B Solution

- The number of the highest occupied energy level is 5, so the element is in the fifth period.
- There are five electrons in the *d* sublevel, which means that it is incompletely filled. The *d* sublevel can hold 10 electrons. Therefore, the element is in the *d* block.
- For d-block elements, the number of electrons in the *ns* sublevel (1) plus the number of electrons in the $(n - 1)d$ sublevel (5) equals the group number, 6.
- This Group 6 element is molybdenum. The others in Group 6 are chromium, tungsten, and seaborgium.

End Of Slide

Periods and Blocks of the Periodic Table, *continued*

- The *p*-block elements consist of all the elements of Groups 13–18 except helium.
- The *p*-block elements together with the *s*-block elements are called the **main-group elements.**
- The properties of elements of the *p* block vary greatly.
  - At its right-hand end, the *p* block includes all of the *nonmetals* except hydrogen and helium.
  - All six of the *metalloids* are also in the *p* block.
  - At the left-hand side and bottom of the block, there are eight *p*-block metals.

Periods and Blocks of the Periodic Table, *continued*

- The elements of Group 17 are known as the **halogens.**
  - fluorine, chlorine, bromine, iodine, and astatine
  - The halogens are the most reactive nonmetals.
  - They react vigorously with most metals to form examples of the type of compound known as salts.
- The metalloids, or semiconducting elements, are located between nonmetals and metals in the *p* block.
- The metals of the *p* block are generally harder and denser than the *s*-block alkaline-earth metals, but softer and less dense than the *d*-block metals.

Periods and Blocks of the Periodic Table, *continued*

Sample Problem C

Without looking at the periodic table, write the outer electron configuration for the Group 14 element in the second period. Then, name the element, and identify it as a metal, nonmetal, or metalloid.

Periods and Blocks of the Periodic Table, *continued*

Sample Problem C Solution

- The group number is higher than 12, so the element is in the *p* block.
- The total number of electrons in the highest occupied *s* and *p* sublevels is therefore equal to the group number minus 10 (14 − 10 = 4).
- Two electrons are in the *s* sublevel, so two electrons must also be present in the 2*p* sublevel.
- The outer electron configuration is $2s^2 2p^2$.
- The element is carbon, C, which is a nonmetal.

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Periods and Blocks of the Periodic Table, *continued*

- In the periodic table, the *f*-block elements are wedged between Groups 3 and 4 in the sixth and seventh periods.
  - Their position reflects the fact that they involve the filling of the 4*f* sublevel.
- The first row of the *f* block, the *lanthanides,* are shiny metals similar in reactivity to the Group 2 alkaline metals.
- The second row of the *f* block, the *actinides,* are between actinium and rutherfordium. The actinides are all radioactive.

End of Slide

Periods and Blocks of the Periodic Table, *continued*

Sample Problem D

Name the block and group in which each of the following elements is located in the periodic table. Then, use the periodic table to name each element. Identify each element as a metal, nonmetal, or metalloid. Finally, describe whether each element has high reactivity or low reactivity.

a. $[Xe]4f^{14}5d^{9}6s^{1}$

b. $[Ne]3s^{2}3p^{5}$

c. $[Ne]3s^{2}3p^{6}$

d. $[Xe]4f^{6}6s^{2}$

Periods and Blocks of the Periodic Table, *continued*

Sample Problem D Solution

a. The 4*f* sublevel is filled with 14 electrons. The 5*d* sublevel is partially filled with nine electrons. Therefore, this element is in the *d* block.

The element is the transition metal platinum, Pt, which is in Group 10 and has a low reactivity.

b. The incompletely filled *p* sublevel shows that this element is in the *p* block.

A total of seven electrons are in the *ns* and *np* sublevels, so this element is in Group 17, the halogens.

The element is chlorine, Cl, and is highly reactive.

Periods and Blocks of the Periodic Table, *continued*

Sample Problem D Solution, *continued*

c. This element has a noble-gas configuration and thus is in Group 18 in the *p* block.

The element is argon, Ar, which is an unreactive nonmetal and a noble gas.

d. The incomplete 4*f* sublevel shows that the element is in the *f* block and is a lanthanide.

Group numbers are not assigned to the *f* block.

The element is samarium, Sm. All of the lanthanides are reactive metals.

End of Slide

Preview

Lesson Starter

- Define *trend.*
- Describe some trends you can observe, such as in fashion, behavior, color, design, and foods.
- How are trends used to classify?

Objectives

- **Define** *atomic* and *ionic radii, ionization energy, electron affinity,* and *electronegativity.*
- **Compare** the periodic trends of atomic radii, ionization energy, and electronegativity, and state the reasons for these variations.
- **Define** *valence electrons,* and state how many are present in atoms of each main-group element.
- **Compare** the atomic radii, ionization energies, and electronegativities of the *d*-block elements with those of the main-group elements.

Atomic Radii

- The boundaries of an atom are fuzzy, and an atom's radius can vary under different conditions.
- To compare different atomic radii, they must be measured under specified conditions.
- **Atomic radius** may be defined as one-half the distance between the nuclei of identical atoms that are bonded together.

End Of Slide

Atomic Radius

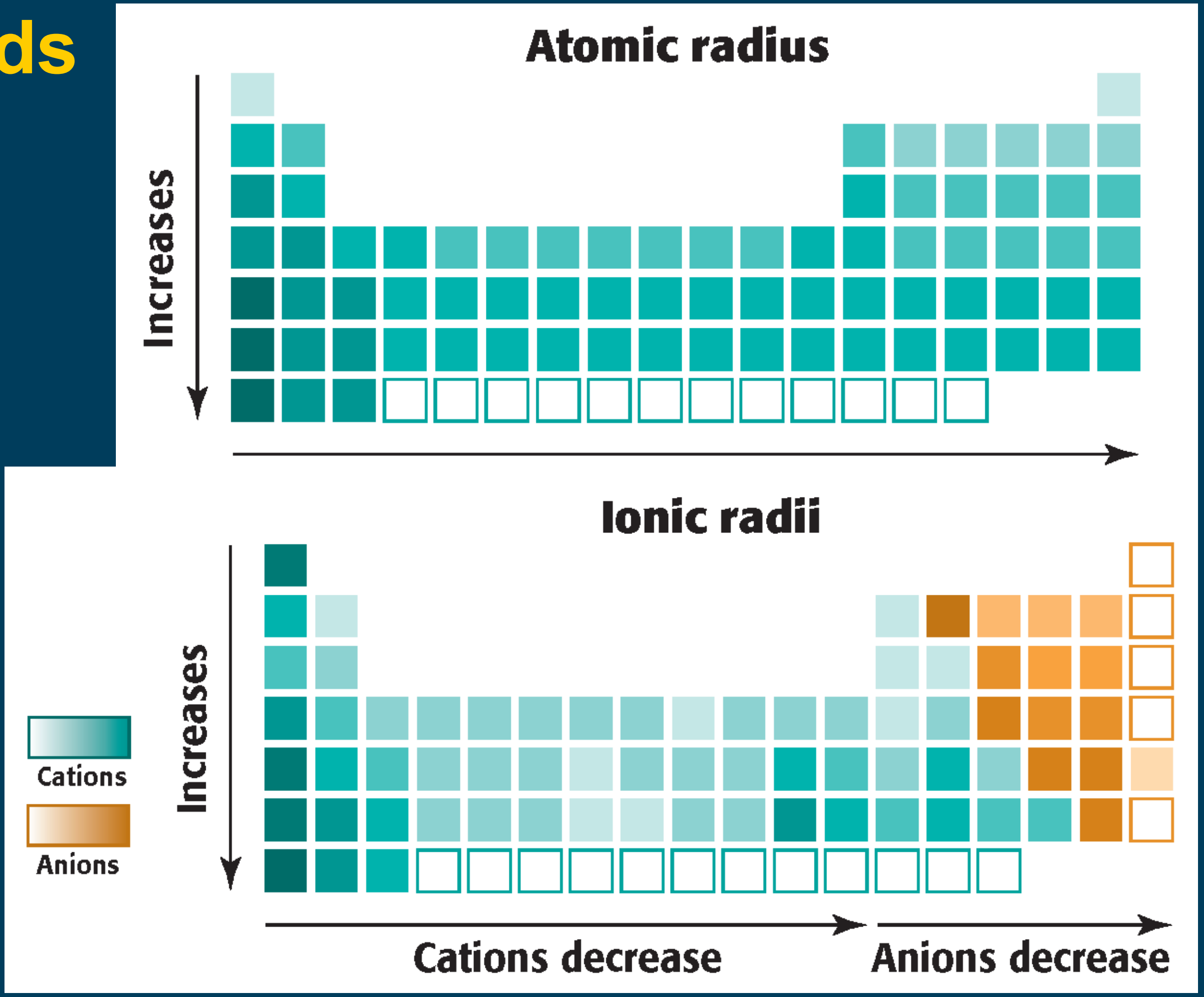
Click below to watch the Visual Concept.

Visual Concept

Atomic Radii, *continued*

- *Atoms tend to be smaller the farther to the right they are found across a period.*
- The trend to smaller atoms across a period is caused by the increasing positive charge of the nucleus, which attracts electrons toward the nucleus.
- *Atoms tend to be larger the farther down in a group they are found.*
- The trend to larger atoms down a group is caused by the increasing size of the electron cloud around an atom as the number electron sublevels increases.

Periodic Trends of Radii



Atomic Radii, *continued*

Sample Problem E

Of the elements magnesium, Mg, chlorine, Cl, sodium, Na, and phosphorus, P, which has the largest atomic radius? Explain your answer in terms of trends of the periodic table.

Atomic Radii, *continued*

Sample Problem E Solution

- Sodium has the largest atomic radius
- All of the elements are in the third period. Of the four, sodium has the lowest atomic number and is the first element in the period. Atomic radii decrease across a period.

End Of Slide

Ionization Energy

- An **ion** is an atom or group of bonded atoms that has a positive or negative charge.
  - Sodium (Na), for example, easily loses an electron to form Na^+.
- Any process that results in the formation of an ion is referred to as **ionization.**
- The energy required to remove one electron from a neutral atom of an element is the **ionization energy, *IE*** (or first ionization energy, IE_1).

End Of Slide

Ion

Click below to watch the Visual Concept.

Visual Concept

Ionization Energy, *continued*

- *In general, ionization energies of the main-group elements increase across each period.*
  - This increase is caused by increasing nuclear charge.
  - A higher charge more strongly attracts electrons in the same energy level.
- *Among the main-group elements, ionization energies generally decrease down the groups.*
  - Electrons removed from atoms of each succeeding element in a group are in higher energy levels, farther from the nucleus.
  - The electrons are removed more easily.

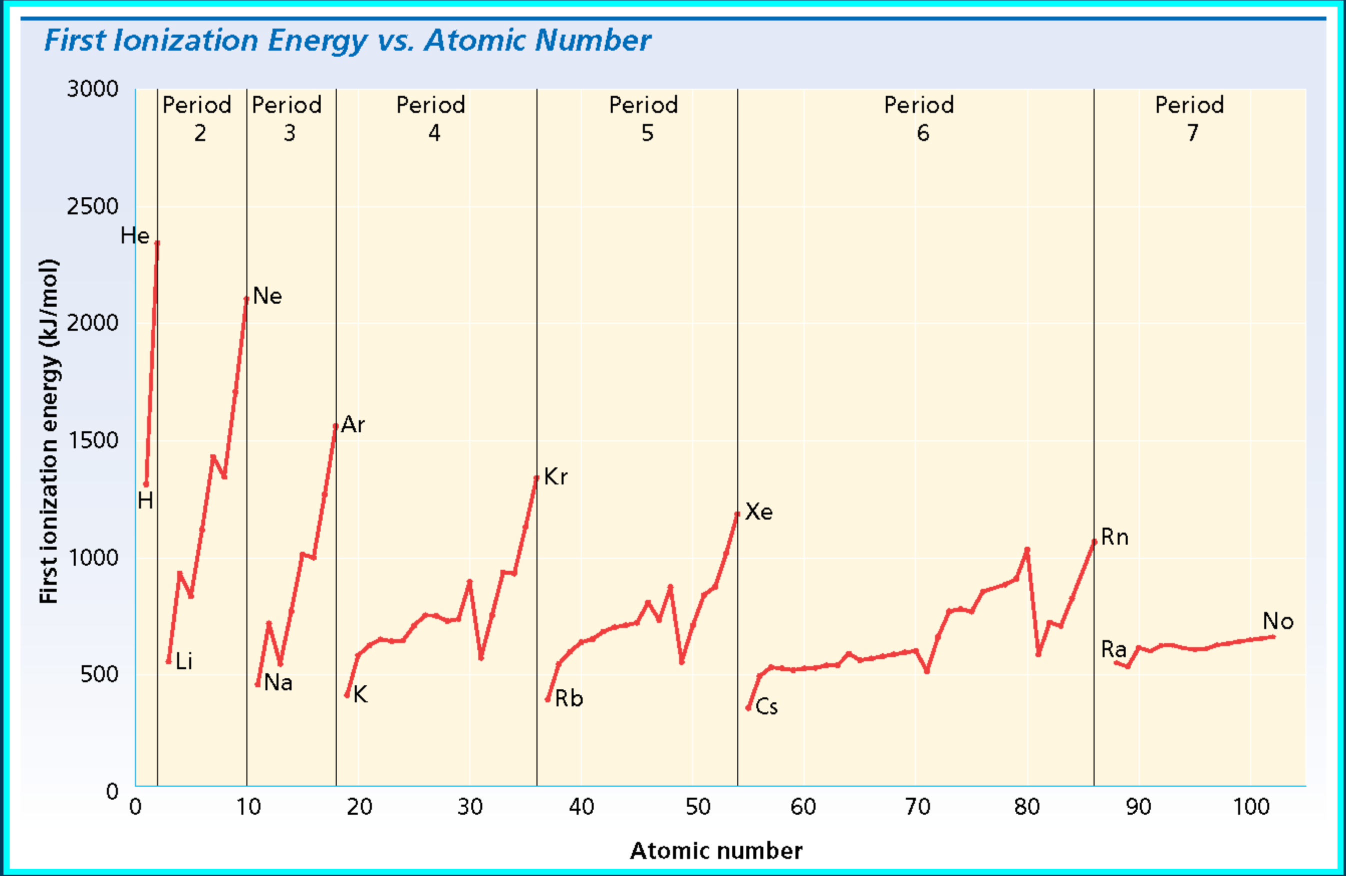
Ionization

Click below to watch the Visual Concept.

Visual Concept

Ionization Energy, *continued*

Periodic trends in ionization energy are shown in the graph below.



Ionization Energy, *continued*

Sample Problem F

Consider two main-group elements, A and B. Element A has a first ionization energy of 419 kJ/mol. Element B has a first ionization energy of 1000 kJ/mol. Decide if each element is more likely to be in the *s* block or *p* block. Which element is more likely to form a positive ion?

End Of Slide

Ionization Energy, *continued*

Sample Problem F Solution

- Element A has a very low ionization energy, which means that atoms of A lose electrons easily.
- Element A is most likely to be an *s*-block metal because ionization energies increase across the periods.
- Element B has a very high ionization energy which means that atoms of B have difficulty losing electrons.
- Element B would most likely lie at the end of a period in the *p* block.
- Element A is more likely to form a positive ion because it has a much lower ionization energy than element B does.

Electron Affinity

- The energy change that occurs when an electron is acquired by a neutral atom is called the atom's **electron affinity.**

- *Electron affinity generally increases across periods.*
  - Increasing nuclear charge along the same sublevel attracts electrons more strongly

- *Electron affinity generally decreases down groups.*
  - The larger an atom's electron cloud is, the farther away its outer electrons are from its nucleus.

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Electron Affinity

Click below to watch the Visual Concept.

Visual Concept

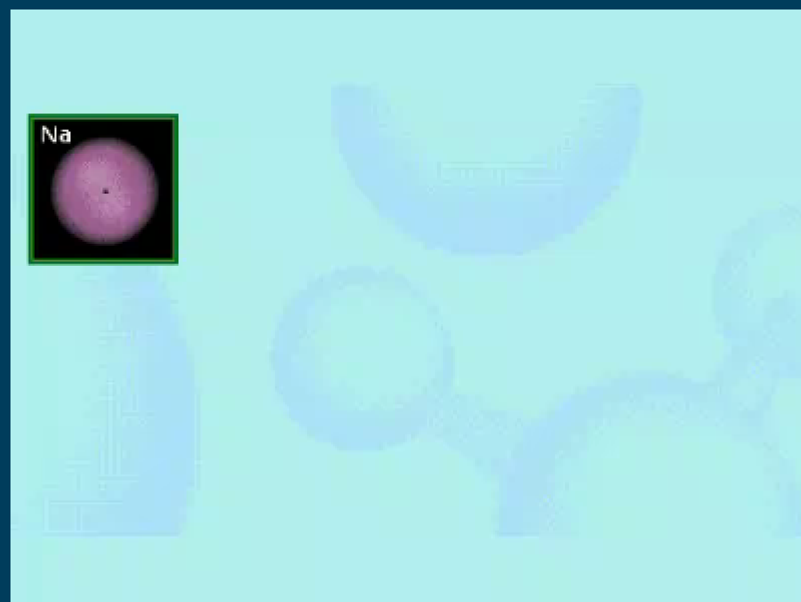
Ionic Radii

- A positive ion is known as a **cation.**
- The formation of a cation by the loss of one or more electrons always leads to a decrease in atomic radius.
  - The electron cloud becomes smaller.
  - The remaining electrons are drawn closer to the nucleus by its unbalanced positive charge.
- A negative ion is known as an **anion.**
- The formation of an anion by the addition of one or more electrons always leads to an increase in atomic radius.

End Of Slide

Comparing Cations and Anions

Click below to watch the Visual Concept.



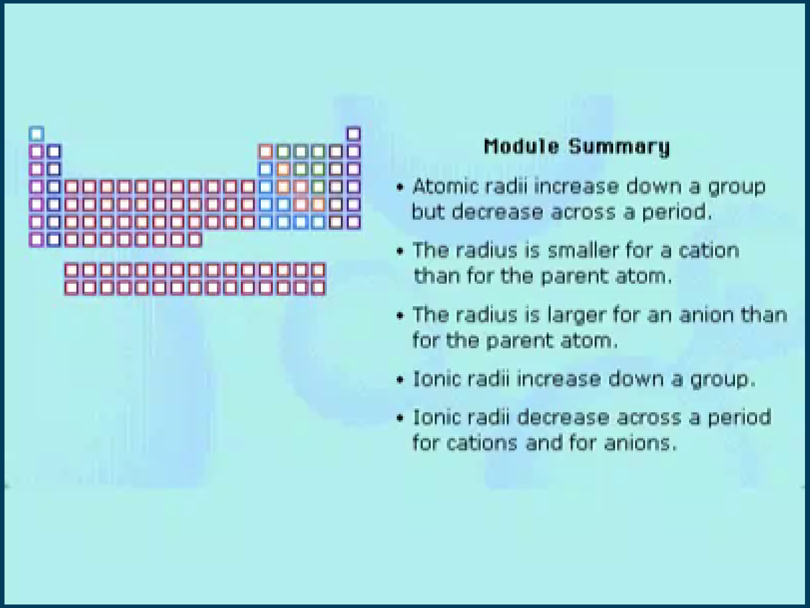
Ionic Radii, *continued*

- Cationic and anionic radii decrease across a period.
  - The electron cloud shrinks due to the increasing nuclear charge acting on the electrons in the same main energy level.
- The outer electrons in both cations and anions are in higher energy levels as one reads down a group.
  - There is a gradual increase of ionic radii down a group.

End Of Slide

Ionic Radius

Click below to watch the Visual Concept.



Valence Electrons

- Chemical compounds form because electrons are lost, gained, or shared between atoms.
- The electrons that interact in this manner are those in the highest energy levels.
- The electrons available to be lost, gained, or shared in the formation of chemical compounds are referred to as **valence electrons.**
  - Valence electrons are often located in incompletely filled main-energy levels.
  - example: the electron lost from the 3*s* sublevel of Na to form Na^+ is a valence electron.

Valence Electrons

Click below to watch the Visual Concept.

Visual Concept

Electronegativity

- Valence electrons hold atoms together in chemical compounds.
- In many compounds, the negative charge of the valence electrons is concentrated closer to one atom than to another.
- **Electronegativity** is a measure of the ability of an atom in a chemical compound to attract electrons from another atom in the compound.
- *Electronegativities tend to increase across periods, and decrease or remain about the same down a group.*

End Of Slide

Electronegativity

Click below to watch the Visual Concept.

Visual Concept

Electronegativity, *continued*

Sample Problem G

Of the elements gallium, Ga, bromine, Br, and calcium, Ca, which has the highest electronegativity? Explain your answer in terms of periodic trends.

Electronegativity, *continued*

Sample Problem G Solution

- All of these elements are in the fourth period.
- Bromine has the highest atomic number and is farthest to the right in the period.
- Bromine should have the highest electronegativity because electronegativity increases across the periods.

End of Chapter 5 Show