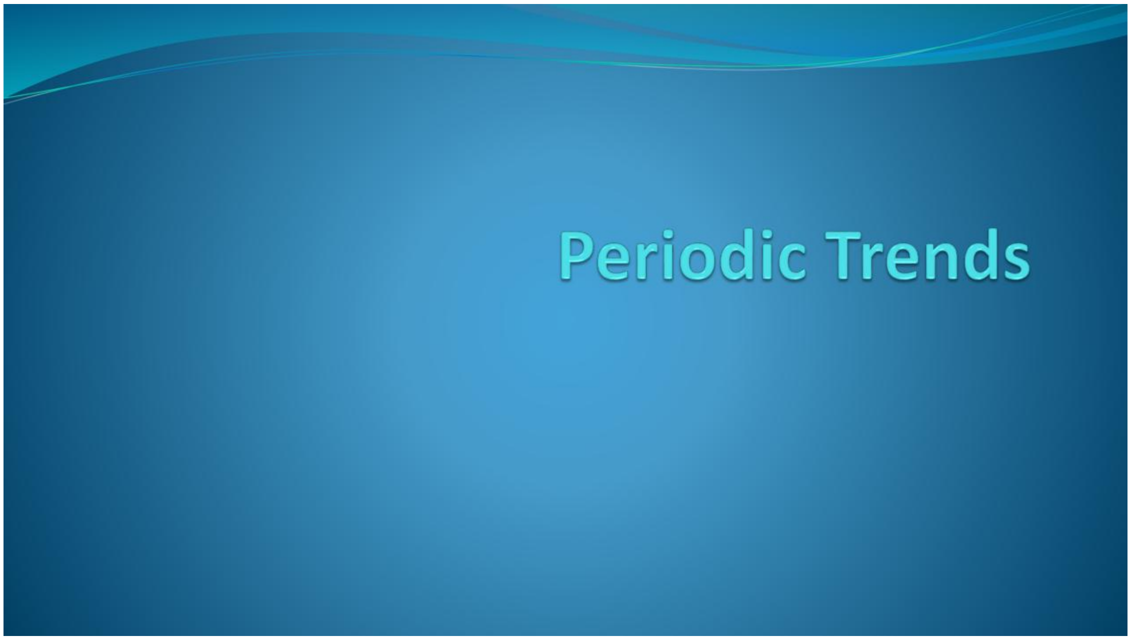
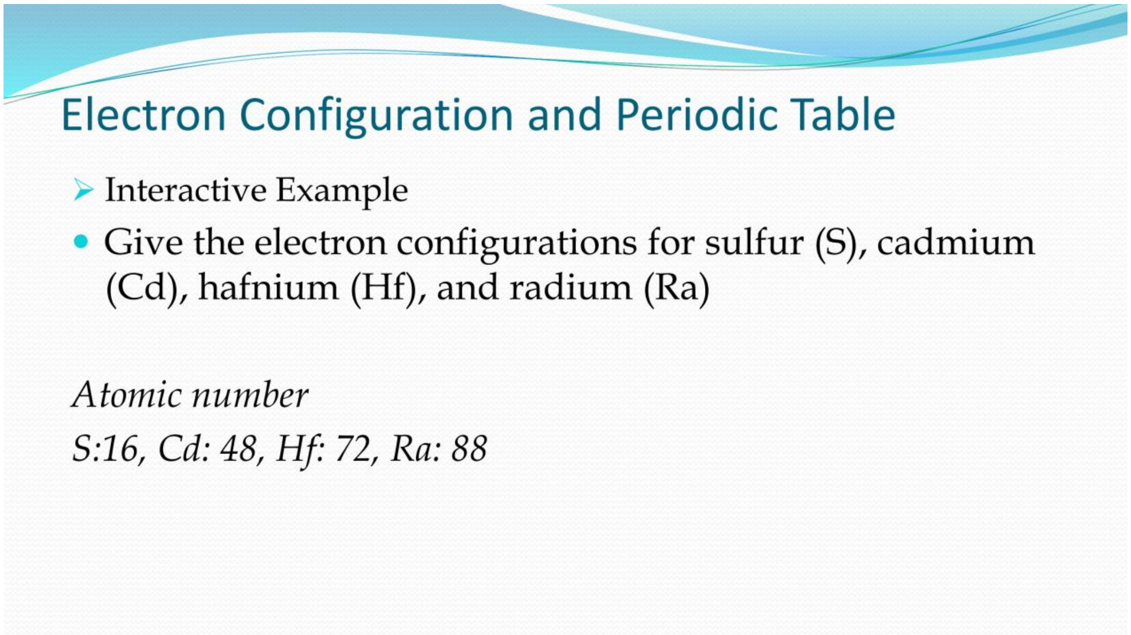


Periodic Trends

Contents

Periodic Trends — Page 3



Electron Configuration and Periodic Table

- Interactive Example
- Give the electron configurations for sulfur (S), cadmium (Cd), hafnium (Hf), and radium (Ra)

Atomic number

S:16, Cd: 48, Hf: 72, Ra: 88

Page 3 Annotation

Title: Electron Configuration and Periodic Table – Interactive Example

Clear Explanation:

Practice writing electron configurations for four elements:

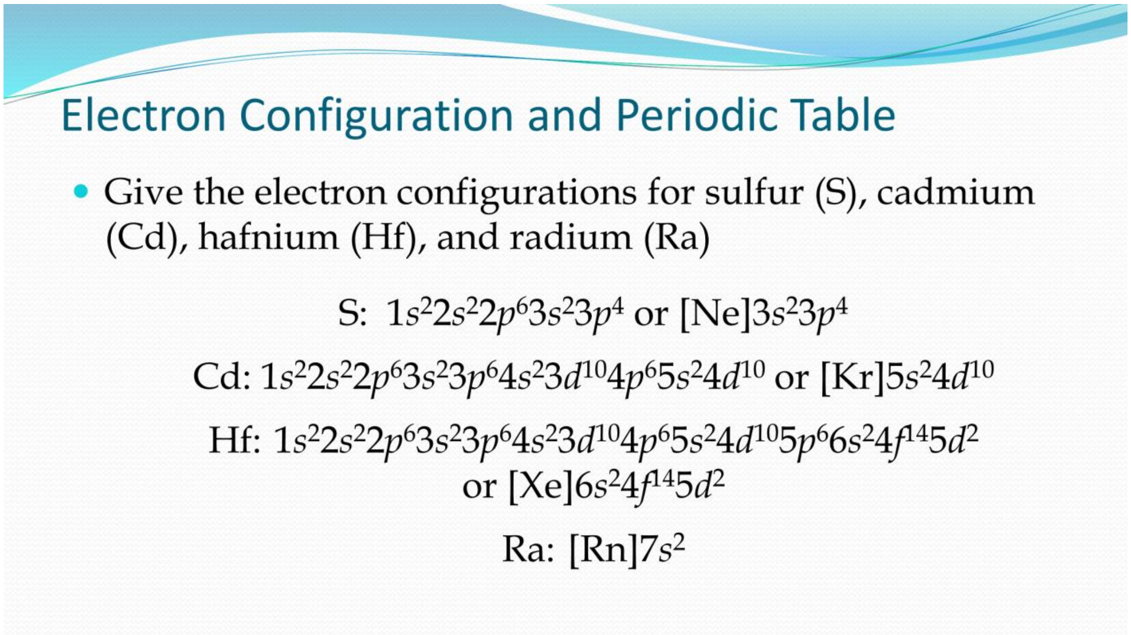
- Sulfur (S, atomic number 16)
- Cadmium (Cd, atomic number 48)
- Hafnium (Hf, atomic number 72)
- Radium (Ra, atomic number 88)

Key Points:

- Use the periodic table to determine the order of orbital filling
- Write both full configuration and noble-gas abbreviated form

Why It Matters:

This is essential practice for reading the periodic table correctly and understanding block locations (s, p, d, f).



Electron Configuration and Periodic Table

- Give the electron configurations for sulfur (S), cadmium (Cd), hafnium (Hf), and radium (Ra)

S: $1s^22s^22p^63s^23p^4$ or $[\mathrm{Ne}]3s^23p^4$

Cd: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}$ or $[\mathrm{Kr}]5s^24d^{10}$

Hf: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^2$
or $[\mathrm{Xe}]6s^24f^{14}5d^2$

Ra: $[\mathrm{Rn}]7s^2$

Page 4 Annotation

Title: Electron Configuration Answers

Clear Explanation:

Correct electron configurations:

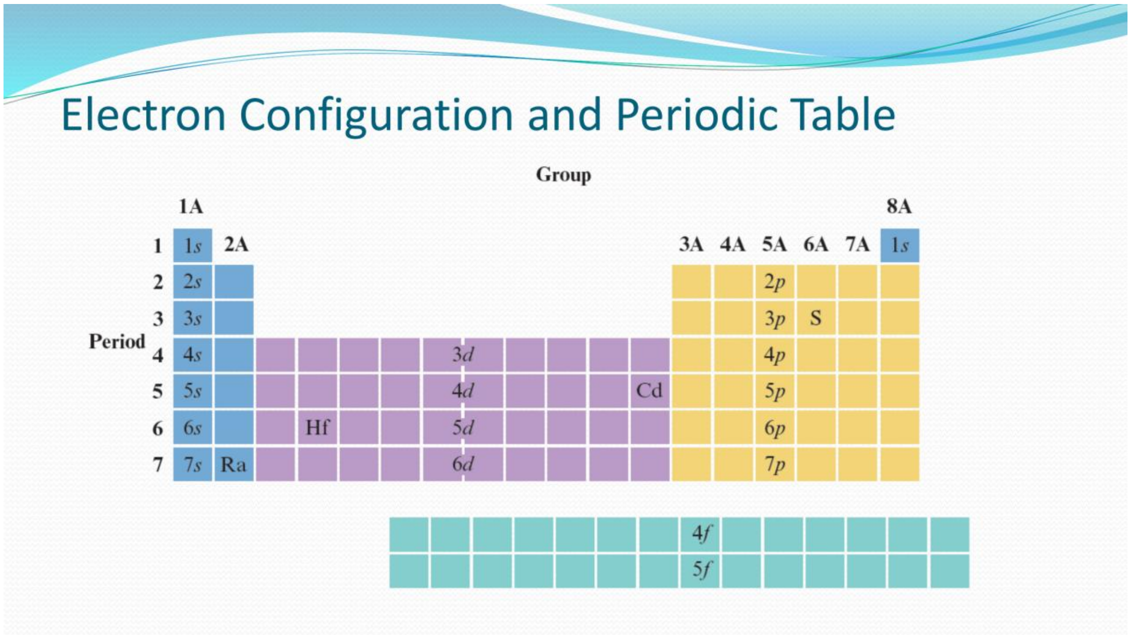
- S: 1s^2 2s^2 2p^6 3s^2 3p^4 or [Ne] 3s^2 3p^4
- Cd: 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} or [Kr] 5s^2 4d^{10}
- Hf: [Xe] 6s^2 4f^{14} 5d^2
- Ra: [Rn] 7s^2

Key Points:

- Noble-gas core notation saves time
- Shows d-block and f-block filling

Why It Matters:

These examples cover main-group, transition, and heavy elements.



Electron Configuration and Periodic Table

Page 5 Annotation

Title: Periodic Table with Blocks Highlighted

Clear Explanation:

The periodic table is divided into blocks based on which orbital is being filled:

- s-block (left 2 columns)
- p-block (right 6 columns)
- d-block (middle 10 columns)
- f-block (two rows at the bottom)

Elements in the same block have similar valence electron configurations.

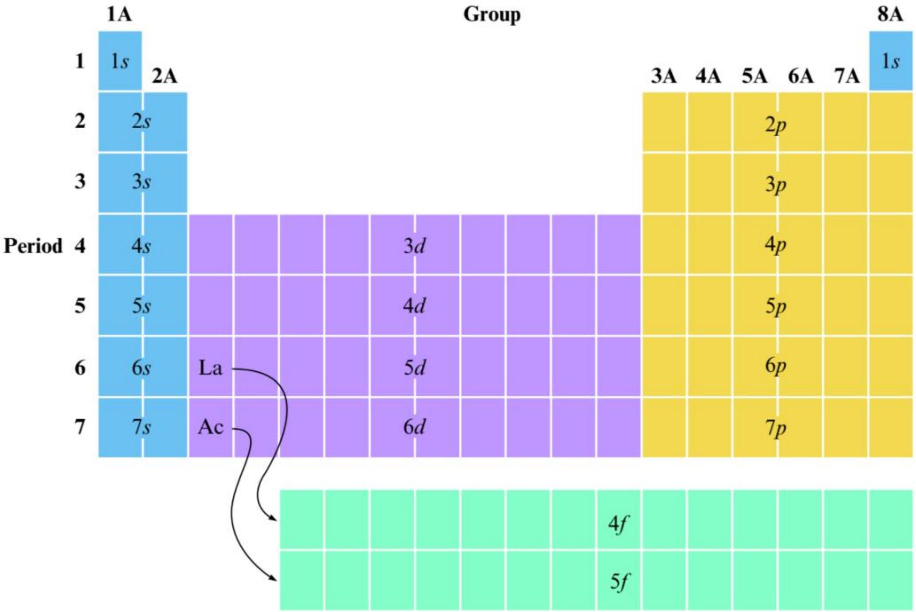
Key Points:

- s-block: groups 1A–2A
- p-block: groups 3A–8A
- d-block: transition metals
- f-block: lanthanoids and actinoids

Why It Matters:

The block tells you the valence orbital and helps predict chemical properties.

Periodic Trends — Page 6



Page 6 Annotation

Title: Periodic Table with Blocks Highlighted

Clear Explanation:

The periodic table is divided into blocks based on which orbital is being filled:

- s-block (left 2 columns)
- p-block (right 6 columns)
- d-block (middle 10 columns)
- f-block (two rows at the bottom)

Elements in the same block have similar valence electron configurations.

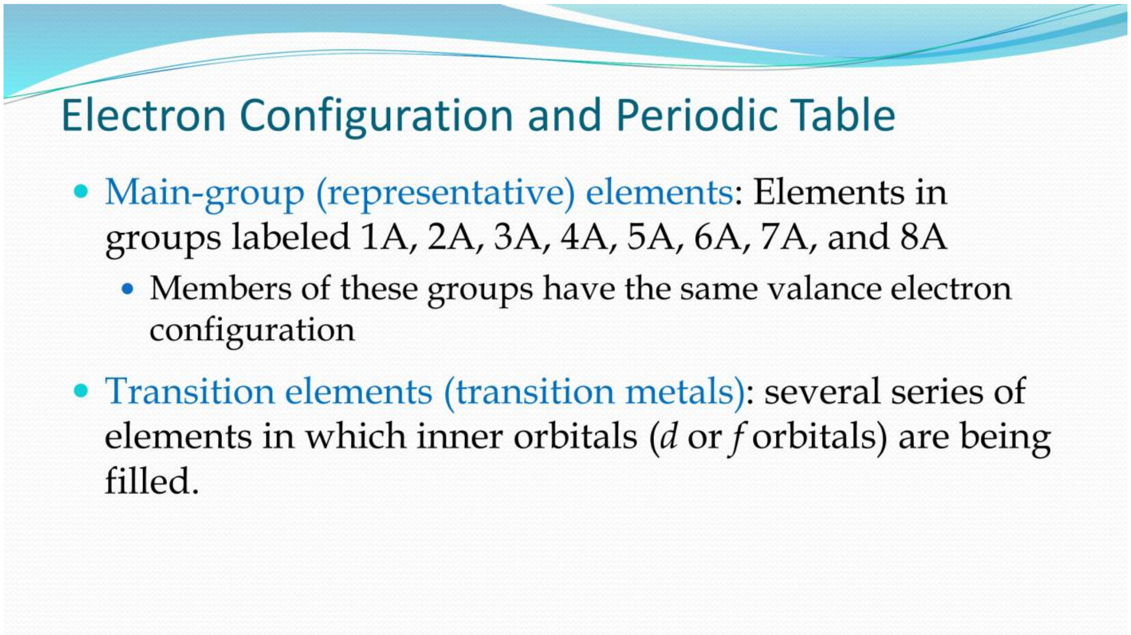
Key Points:

- s-block: groups 1A–2A
- p-block: groups 3A–8A
- d-block: transition metals
- f-block: lanthanoids and actinoids

Why It Matters:

The block tells you the valence orbital and helps predict chemical properties.

Periodic Trends — Page 7



Electron Configuration and Periodic Table

- Main-group (representative) elements: Elements in groups labeled 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A
  - Members of these groups have the same valance electron configuration
- Transition elements (transition metals): several series of elements in which inner orbitals (*d* or *f* orbitals) are being filled.

Page 7 Annotation

Title: Main-group and Transition Elements

Clear Explanation:

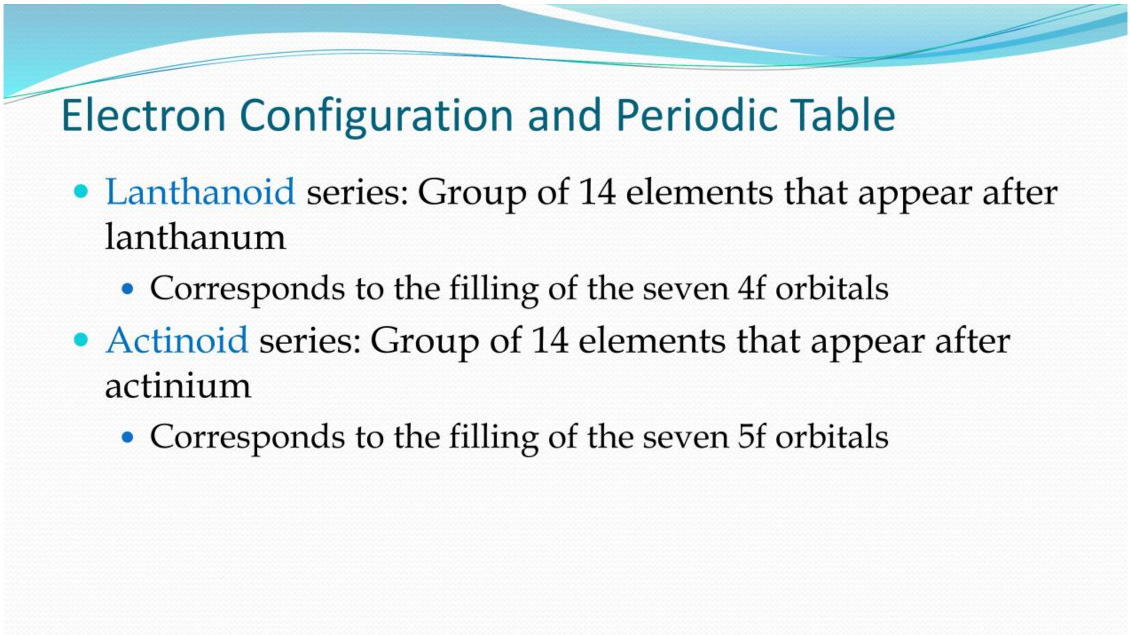
- **Main-group (representative) elements**: Groups 1A–8A. All elements in the same group have the same number of valence electrons and similar chemical properties.
- **Transition elements (transition metals)**: d-block. Inner d orbitals are being filled.

Key Points:

- Main-group: same valence configuration within a group
- Transition: d orbitals being filled

Why It Matters:

Main-group elements are more predictable; transition metals often have multiple oxidation states.



Electron Configuration and Periodic Table

- Lanthanoid series: Group of 14 elements that appear after lanthanum
  - Corresponds to the filling of the seven 4f orbitals
- Actinoid series: Group of 14 elements that appear after actinium
  - Corresponds to the filling of the seven 5f orbitals

Page 8 Annotation

Title: Lanthanoid and Actinoid Series

Clear Explanation:

- **Lanthanoid series**: 14 elements after lanthanum (La); 4f orbitals are filled.
- **Actinoid series**: 14 elements after actinium (Ac); 5f orbitals are filled.

These are the two rows at the bottom of the periodic table.

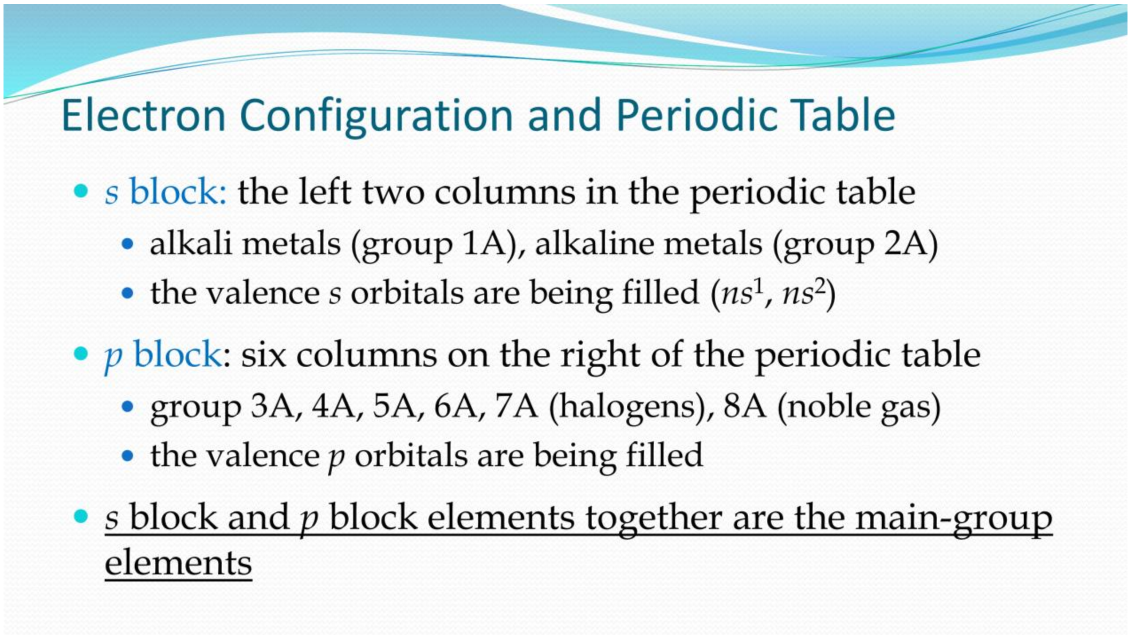
Key Points:

- Lanthanoids: 4f filling
- Actinoids: 5f filling

Why It Matters:

These elements have very similar properties because the f orbitals are deep inside the atom.

Periodic Trends — Page 9



Electron Configuration and Periodic Table

- *s* block: the left two columns in the periodic table
  - alkali metals (group 1A), alkaline metals (group 2A)
  - the valence *s* orbitals are being filled (ns^1, ns^2)
- *p* block: six columns on the right of the periodic table
  - group 3A, 4A, 5A, 6A, 7A (halogens), 8A (noble gas)
  - the valence *p* orbitals are being filled
- *s* block and *p* block elements together are the main-group elements

Page 9 Annotation

Title: s-block and p-block

Clear Explanation:

- **s-block**: Left two columns (groups 1A and 2A). Valence electrons are in s orbitals (ns^1 or ns^2).
- **p-block**: Right six columns (groups 3A–8A). Valence electrons are in p orbitals.

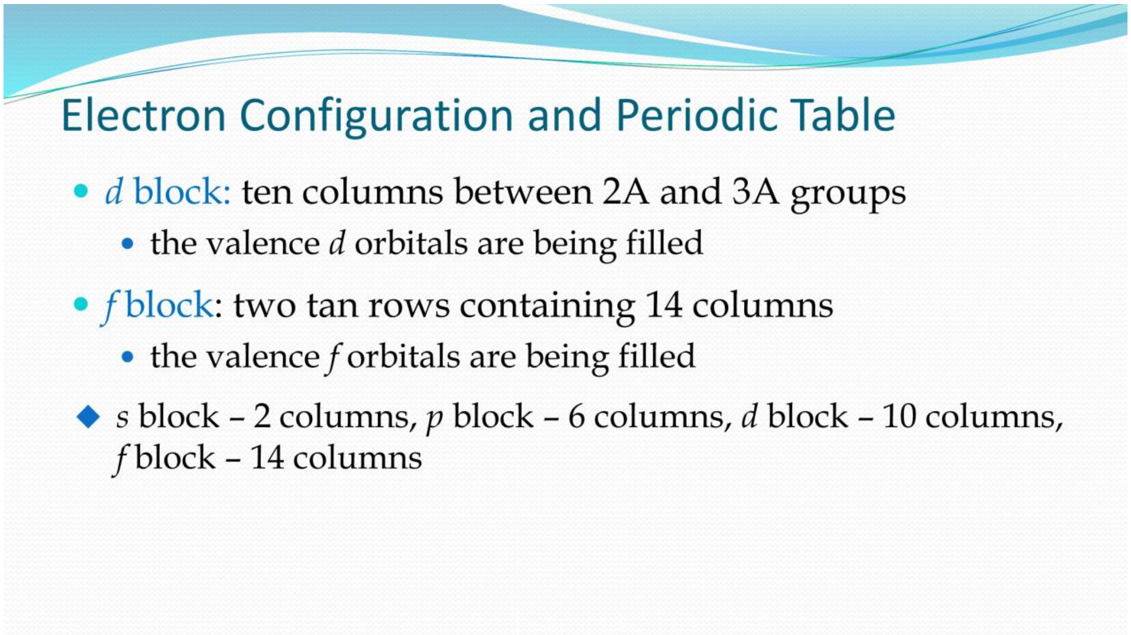
Together, s-block and p-block make up the **main-group elements**.

Key Points:

- s-block: alkali and alkaline earth metals
- p-block: includes halogens and noble gases

Why It Matters:

Most everyday elements and compounds come from the s- and p-blocks.



Electron Configuration and Periodic Table

- *d* block: ten columns between 2A and 3A groups
  - the valence *d* orbitals are being filled
- *f* block: two tan rows containing 14 columns
  - the valence *f* orbitals are being filled

◆ *s* block – 2 columns, *p* block – 6 columns, *d* block – 10 columns, *f* block – 14 columns

Page 10 Annotation

Title: d-block and f-block

Clear Explanation:

- **d-block**: 10 columns between groups 2A and 3A. Valence d orbitals are being filled.
- **f-block**: Two rows of 14 elements each (lanthanoids and actinoids). Valence f orbitals are being filled.

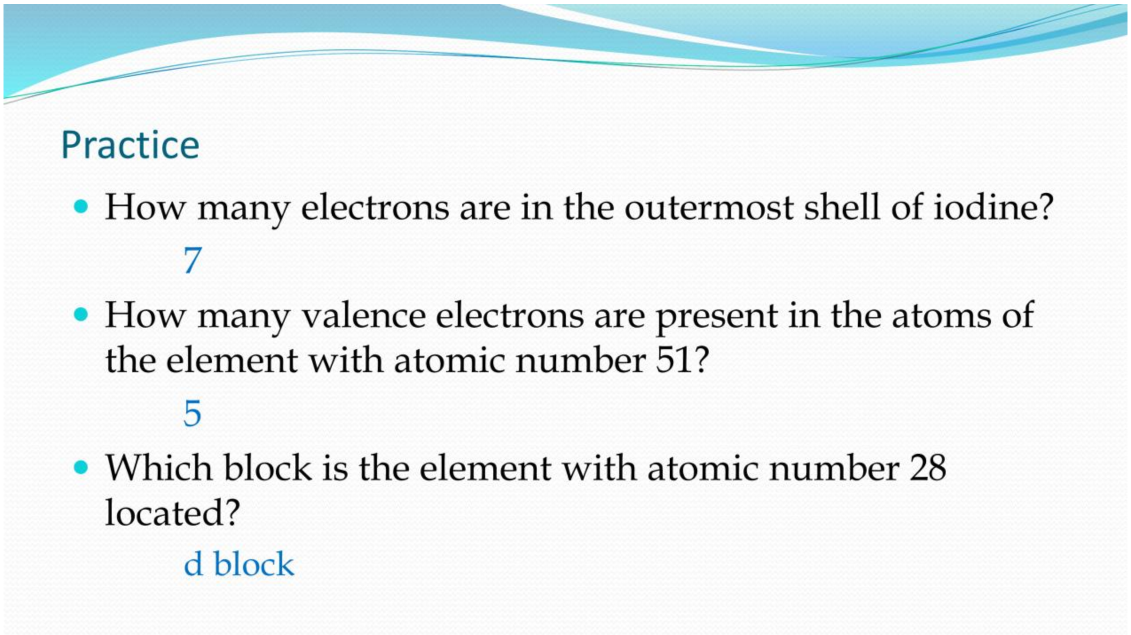
Key Points:

- d-block: transition metals
- f-block: inner transition metals

Why It Matters:

Understanding blocks helps predict electron configuration and reactivity.

Periodic Trends — Page 11



Practice

- How many electrons are in the outermost shell of iodine?

  7

- How many valence electrons are present in the atoms of the element with atomic number 51?

  5

- Which block is the element with atomic number 28 located?

  d block

Page 11 Annotation

Title: Practice Questions

Clear Explanation:

1. How many electrons are in the outermost shell of iodine? → **7**
2. How many valence electrons in element 51 (antimony)? → **5**
3. Which block is element 28 (nickel) in? → **d-block**

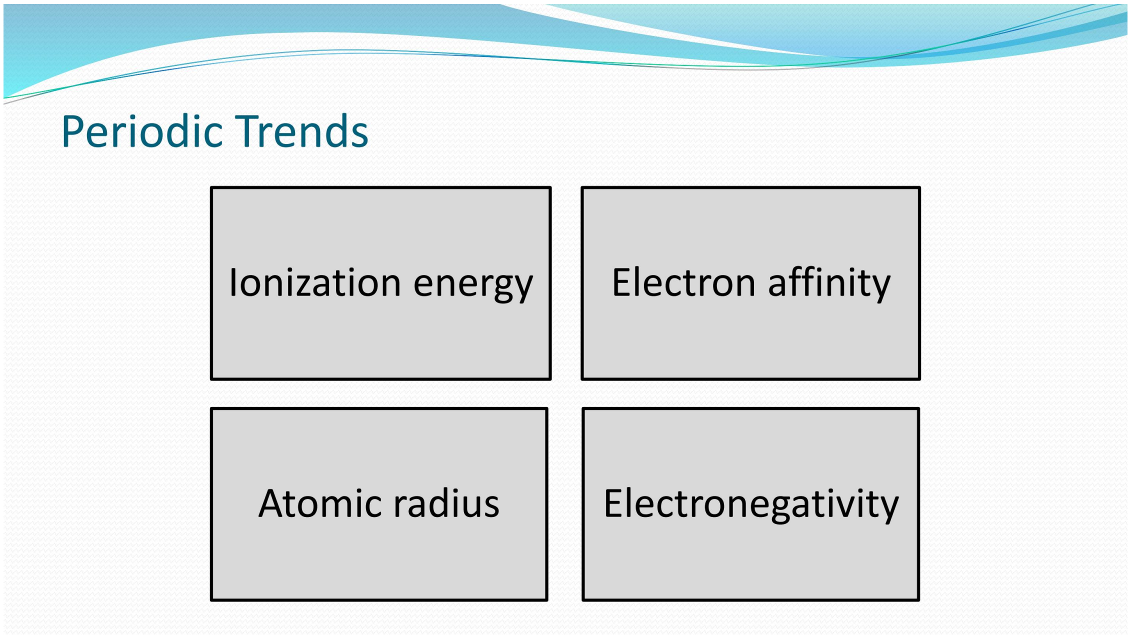
Key Points:

- Outermost shell = valence electrons
- Block determined by which orbital is filling last

Why It Matters:

Quick review to check understanding of electron configuration and periodic table blocks.

Periodic Trends — Page 12



Periodic Trends

Ionization energy

Electron affinity

Atomic radius

Electronegativity

Page 12 Annotation

Title: The Four Major Periodic Trends

Clear Explanation:

This unit focuses on four important trends:

- Ionization energy
- Electron affinity
- Atomic radius
- Electronegativity

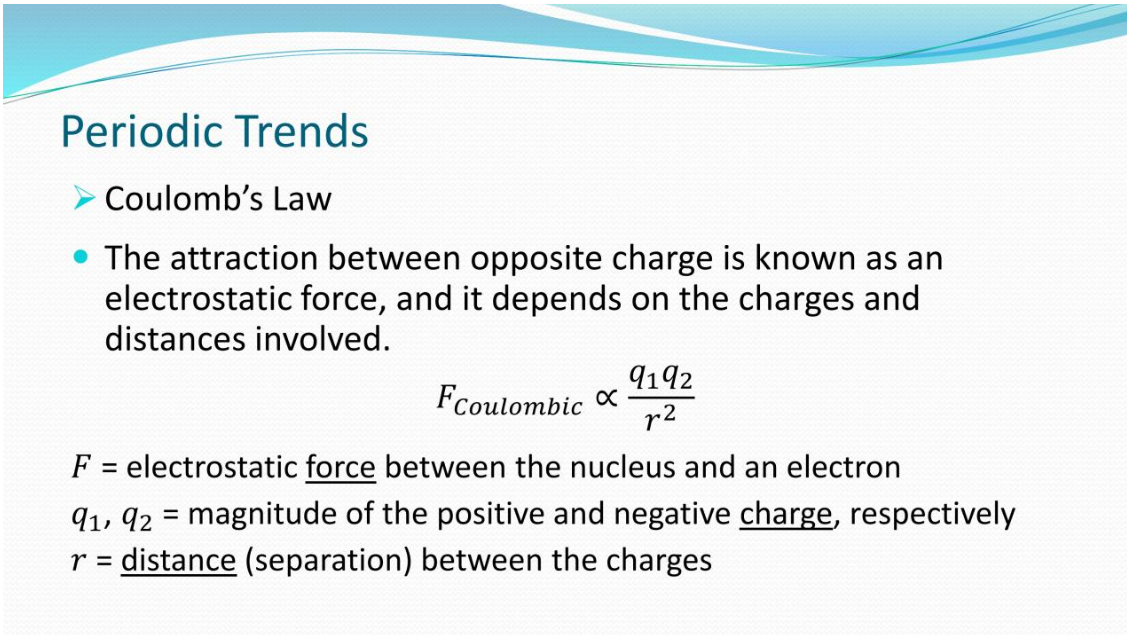
Key Points:

- These trends depend on nuclear charge, distance from nucleus, and shielding.

Why It Matters:

These trends explain why elements behave differently and help predict chemical reactions.

Periodic Trends — Page 13



Periodic Trends

- Coulomb's Law
- The attraction between opposite charge is known as an electrostatic force, and it depends on the charges and distances involved.

$$F_{Coulombic} \propto \frac{q_1 q_2}{r^2}$$

F = electrostatic force between the nucleus and an electron

q_1, q_2 = magnitude of the positive and negative charge, respectively

r = distance (separation) between the charges

Page 13 Annotation

Title: Coulomb’s Law

Clear Explanation:

Coulomb’s Law describes the electrostatic force between charged particles:

F \propto \frac{q_1 q_2}{r^2}

In atoms, the nucleus (positive) attracts electrons (negative). Stronger force = electrons held more tightly.

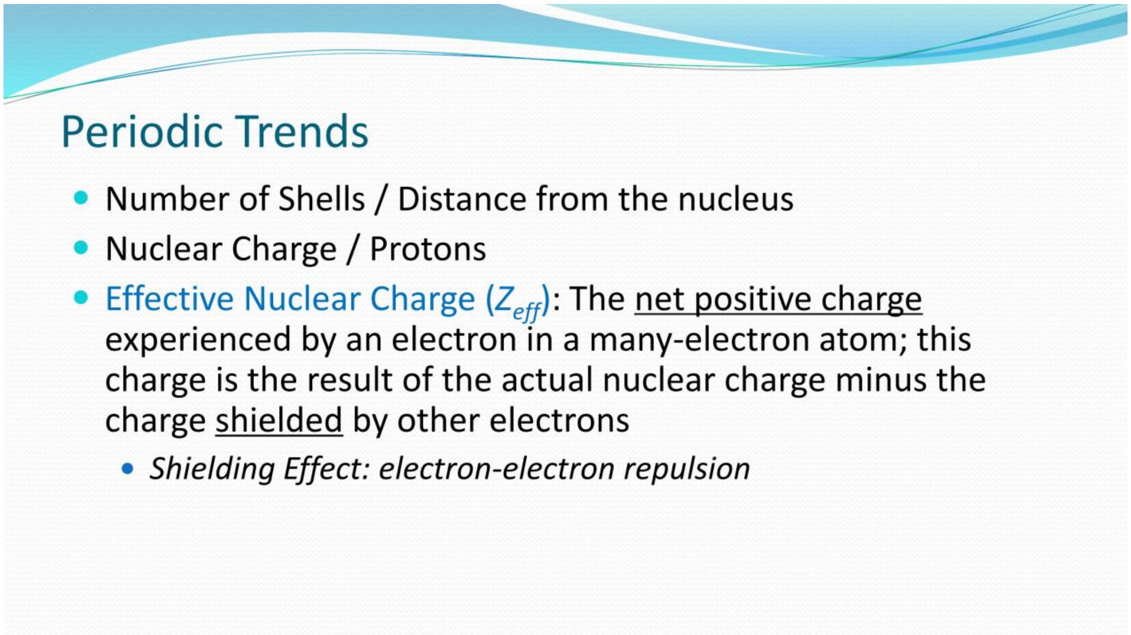
Key Points:

- Larger charge difference → stronger attraction
- Smaller distance (r) → stronger attraction

Why It Matters:

This law is the fundamental reason behind all periodic trends.

Periodic Trends — Page 14



Periodic Trends

- Number of Shells / Distance from the nucleus
- Nuclear Charge / Protons
- Effective Nuclear Charge (Z_{eff}): The net positive charge experienced by an electron in a many-electron atom; this charge is the result of the actual nuclear charge minus the charge shielded by other electrons
  - *Shielding Effect: electron-electron repulsion*

Page 14 Annotation

Title: Factors Affecting Periodic Trends

Clear Explanation:

Three main factors control trends:

4. Number of electron shells / distance from nucleus
5. Nuclear charge (number of protons)
6. **Effective nuclear charge (Z_eff)**: net positive charge felt by valence electrons (actual nuclear charge minus shielding)

Key Points:

- More shells → larger radius, weaker attraction
- More protons → stronger attraction
- Shielding Effect: inner electrons block outer electrons from full nuclear charge

Why It Matters:

These factors explain why trends go up/down/left/right on the periodic table.

Periodic Trends

- **An analogy for effective nuclear charge.**

We envision the nucleus as a light bulb, the core electrons as a frosted glass lampshade, and a valence electron as an observer. The amount of light seen by the observer depends on the intensity of the light bulb and the screening by the frosted glass lampshades.

Page 15 Annotation

Title: Analogy for Effective Nuclear Charge (Z_eff)

Clear Explanation:

Imagine the nucleus as a bright lightbulb, core electrons as a frosted glass lampshade, and a valence electron as an observer. The amount of “light” (nuclear attraction) the valence electron feels depends on:

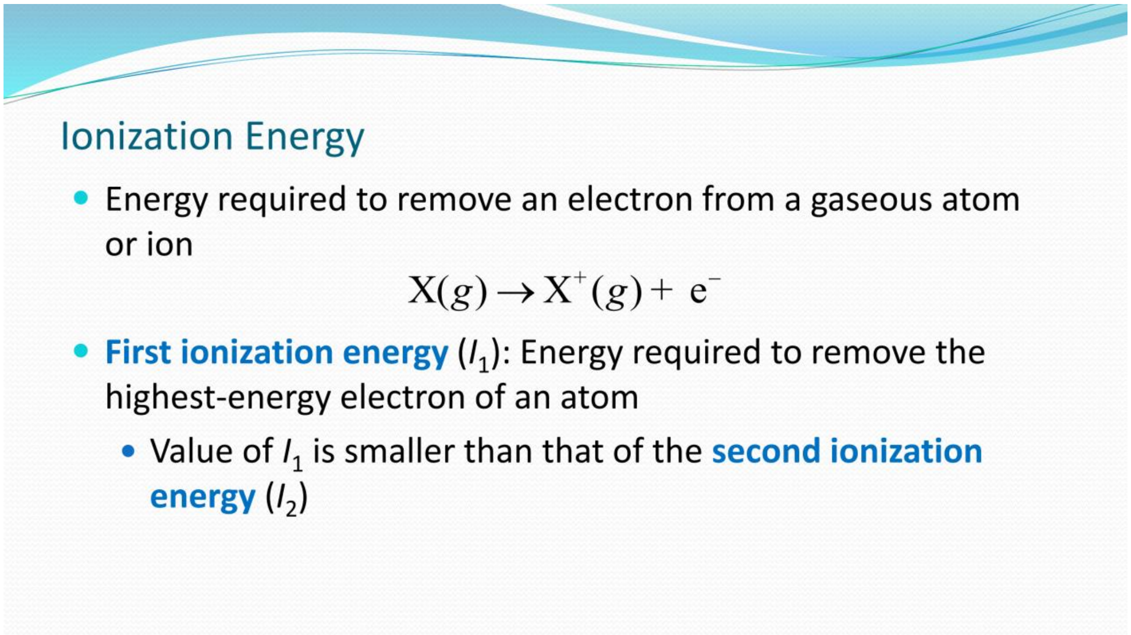
- How bright the bulb is (nuclear charge)
- How much the lampshade blocks it (shielding)

Key Points:

- More protons = brighter bulb
- More core electrons = thicker lampshade (more shielding)

Why It Matters:

This analogy makes the abstract concept of Z_eff easy to understand and remember.



Ionization Energy

- Energy required to remove an electron from a gaseous atom or ion

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

- **First ionization energy** (I_1): Energy required to remove the highest-energy electron of an atom
  - Value of I_1 is smaller than that of the **second ionization energy** (I_2)

Page 16 Annotation

Title: Ionization Energy

Clear Explanation:

Ionization energy is the energy required to remove an electron from a gaseous atom or ion.

The equation is: X(g) \rightarrow X^+(g) + e^- .

First ionization energy (I_1) is the energy needed to remove the outermost (highest-energy) electron from a neutral atom.

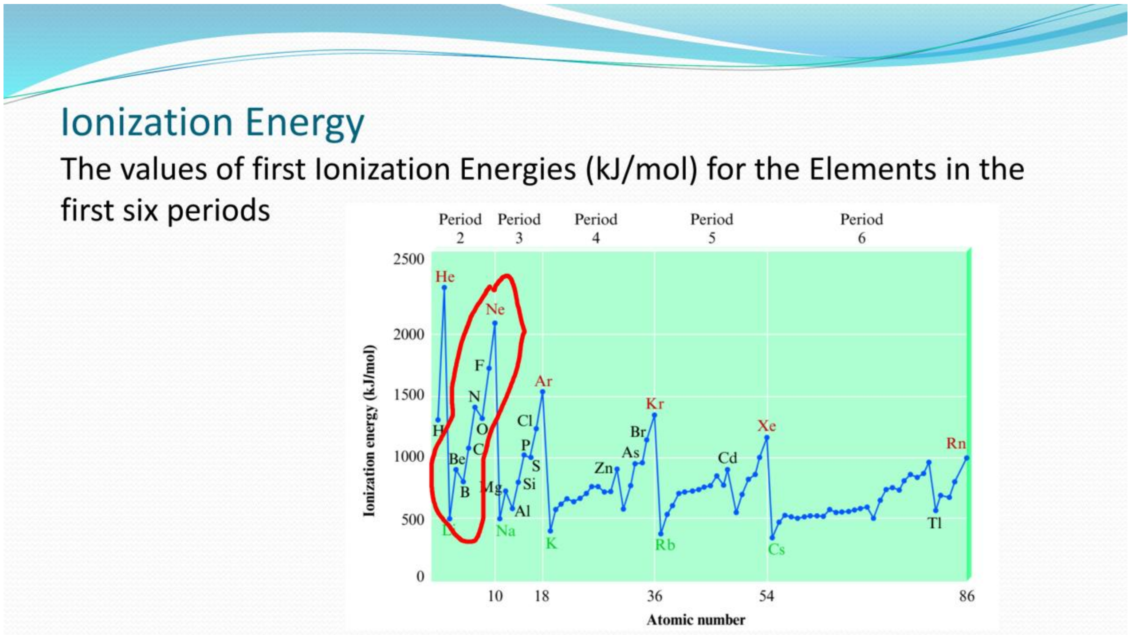
Key Points:

- Always measured for gaseous atoms
- I_1 is always smaller than the second ionization energy (I_2)
- Removing the first electron is relatively easier than removing a second one from a positively charged ion

Why It Matters:

Ionization energy tells us how strongly an atom holds onto its electrons. It is one of the most important periodic trends.

Periodic Trends — Page 17



Ionization Energy

The values of first Ionization Energies (kJ/mol) for the Elements in the first six periods

Page 17 Annotation

Title: Ionization Energy – Graph of First Ionization Energies

Clear Explanation:

This graph shows the first ionization energies (in kJ/mol) for elements in the first six periods.

You can clearly see the overall pattern: ionization energy generally increases across a period (left to right) and decreases down a group (top to bottom). Noble gases (He, Ne, Ar, etc.) have the highest values.

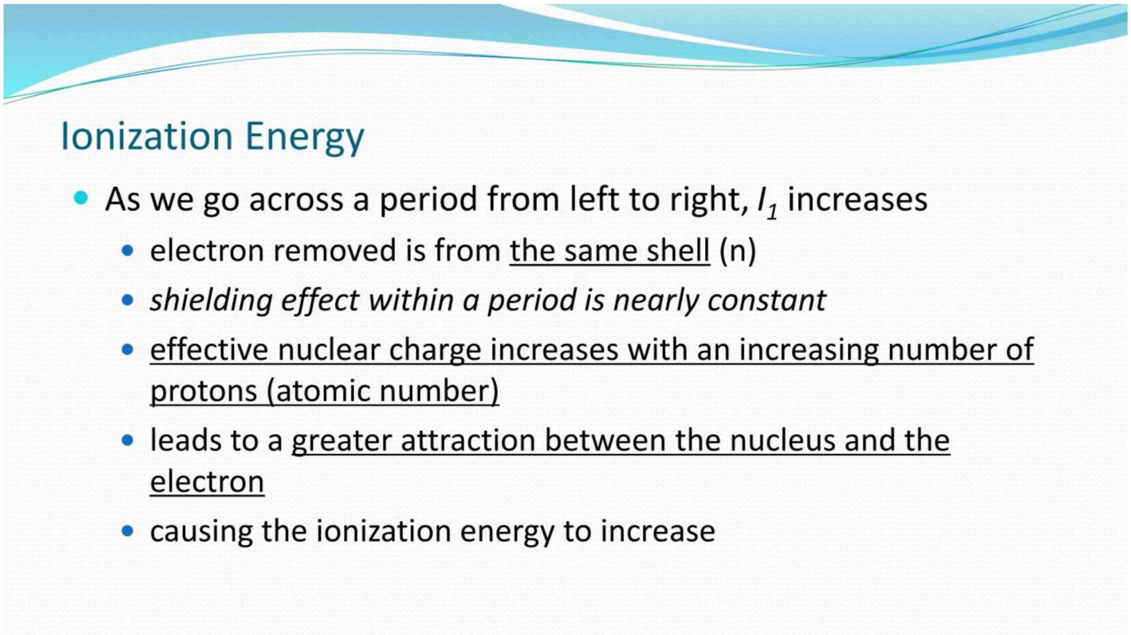
Key Points:

- Peaks at noble gases
- Drops sharply after each noble gas (start of new period)
- Small dips within periods (e.g., Be → B, N → O)

Why It Matters:

The graph visually demonstrates the periodic trends we will explain in the next slides.

Periodic Trends — Page 18



Ionization Energy

- As we go across a period from left to right, I_1 increases
  - electron removed is from the same shell (n)
  - *shielding effect within a period is nearly constant*
  - effective nuclear charge increases with an increasing number of protons (atomic number)
  - leads to a greater attraction between the nucleus and the electron
  - causing the ionization energy to increase

Page 18 Annotation

Title: Ionization Energy – Trend Across a Period

Clear Explanation:

As we go across a period from left to right, the first ionization energy (I_1) generally **increases**.

Reasons:

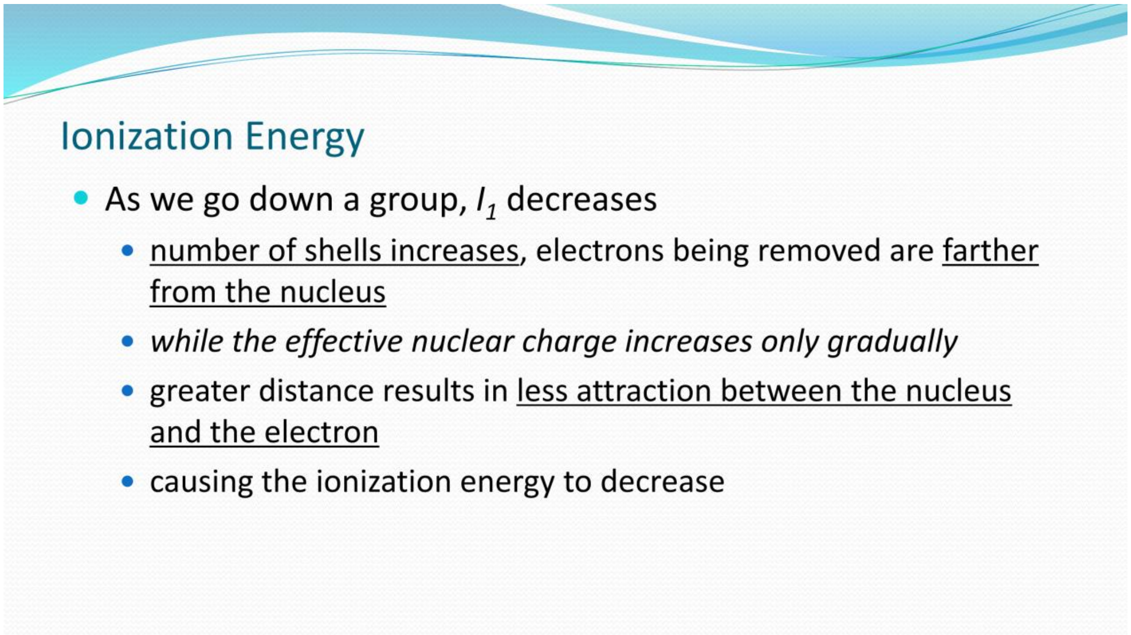
- All valence electrons are in the same shell (same n)
- Shielding effect is nearly constant
- Number of protons (nuclear charge) increases
- Effective nuclear charge (Z_eff) increases → stronger attraction between nucleus and valence electrons

Key Points:

- Same shell → electrons are at similar distance
- More protons → stronger pull
- Harder to remove an electron

Why It Matters:

This explains why metals (left side) lose electrons easily while nonmetals (right side) hold them tightly.



Ionization Energy

- As we go down a group, I_1 decreases
  - number of shells increases, electrons being removed are farther from the nucleus
  - *while the effective nuclear charge increases only gradually*
  - greater distance results in less attraction between the nucleus and the electron
  - causing the ionization energy to decrease

Page 19 Annotation

Title: Ionization Energy – Trend Down a Group

Clear Explanation:

As we go down a group, the first ionization energy (I_1) generally **decreases**.

Reasons:

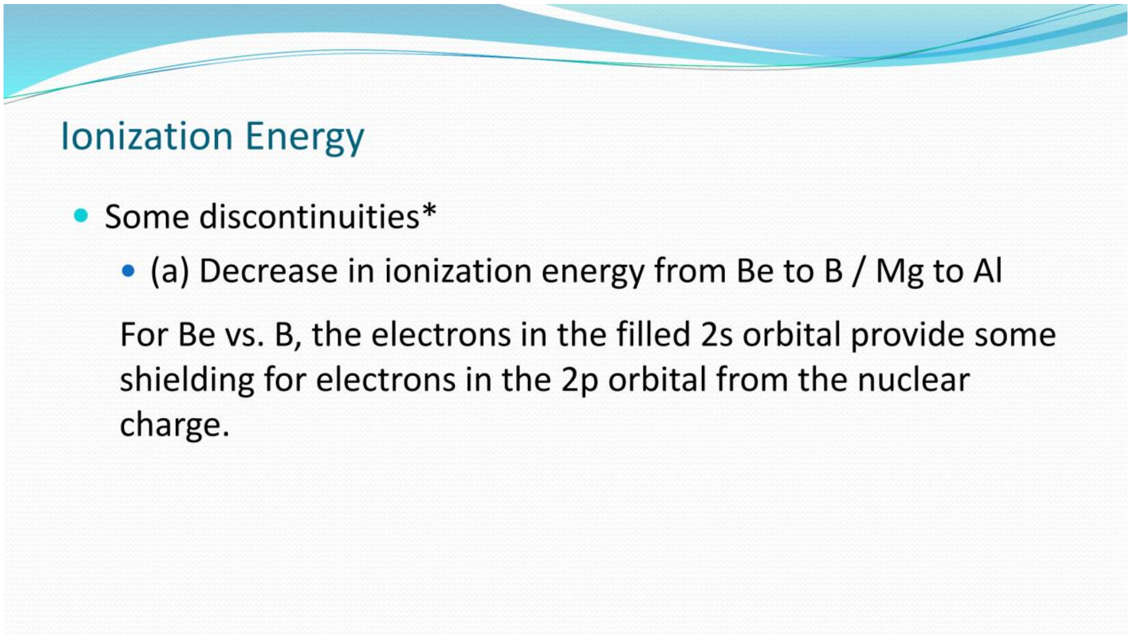
- Number of electron shells increases
- Valence electrons are farther from the nucleus
- Effective nuclear charge increases only slightly
- Greater distance weakens the attraction between nucleus and valence electrons

Key Points:

- More shells → larger atomic radius
- Outer electrons are shielded more effectively
- Easier to remove an electron

Why It Matters:

This is why alkali metals (Group 1A) become more reactive as you go down the group.



Ionization Energy

- Some discontinuities*
  - (a) Decrease in ionization energy from Be to B / Mg to Al

  For Be vs. B, the electrons in the filled 2s orbital provide some shielding for electrons in the 2p orbital from the nuclear charge.

Page 20 Annotation

Title: Ionization Energy – Some Discontinuities (Part a)

Clear Explanation:

There are small drops in ionization energy that do not follow the general trend:

- From Be to B (and Mg to Al)

Reason:

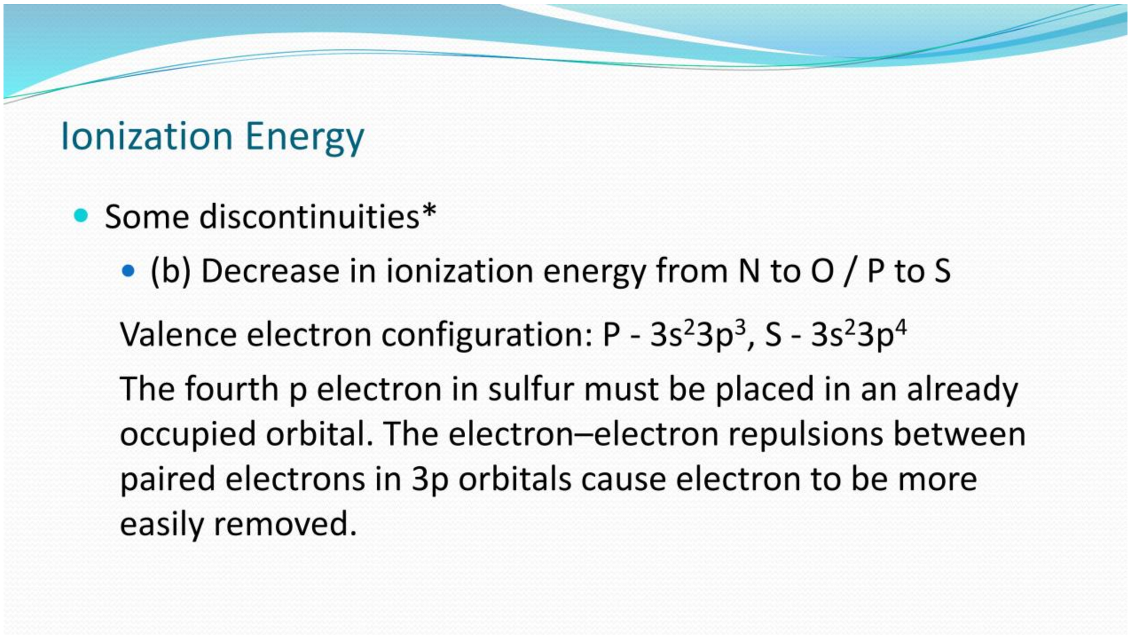
The electron removed from B (or Al) is in a higher-energy p orbital, while Be (or Mg) has a filled 2s (or 3s) orbital. The filled s orbital provides extra shielding for the p electron, making it easier to remove.

Key Points:

- Be: $1s^2 2s^2$ (filled s orbital)
- B: $1s^2 2s^2 2p^1$ (p electron is shielded)

Why It Matters:

These exceptions help us understand orbital energy levels and shielding effects more deeply.



Ionization Energy

- Some discontinuities*
  - (b) Decrease in ionization energy from N to O / P to S

Valence electron configuration: P - $3s^23p^3$, S - $3s^23p^4$

The fourth p electron in sulfur must be placed in an already occupied orbital. The electron–electron repulsions between paired electrons in 3p orbitals cause electron to be more easily removed.

Page 21 Annotation

Title: Ionization Energy – Some Discontinuities (Part b)

Clear Explanation:

Another small drop occurs from N to O (and P to S).

Reason:

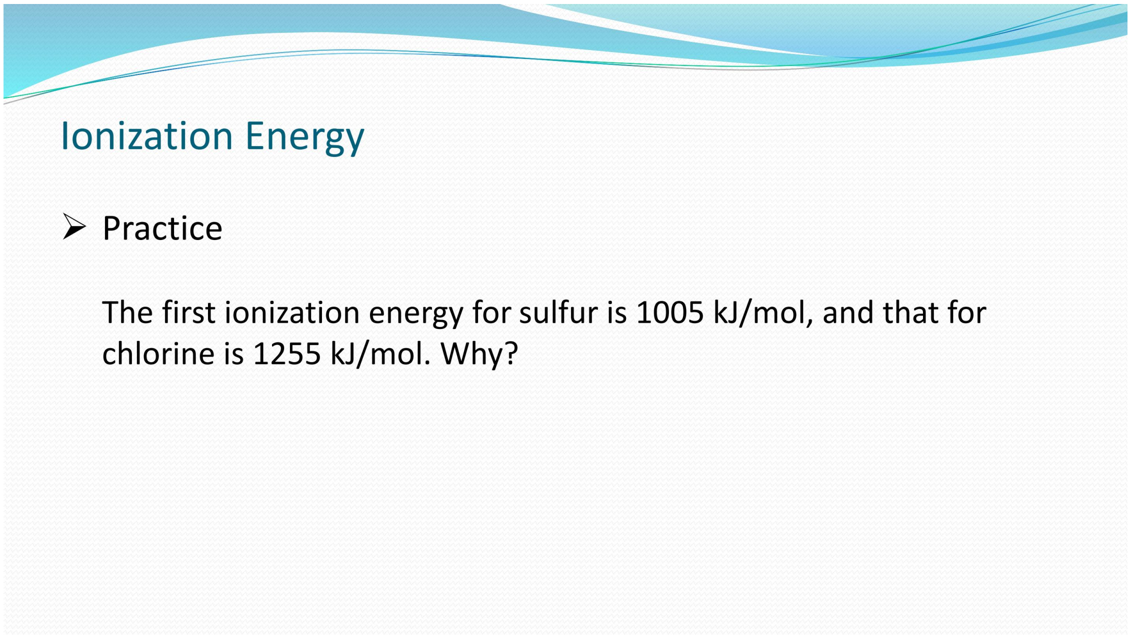
N has a half-filled p subshell ($2p^3$), which is unusually stable. O has one paired electron in the 2p orbital ($2p^4$). The paired electrons repel each other, so it is easier to remove one electron from O.

Key Points:

- N: $2p^3$ (half-filled, stable)
- O: $2p^4$ (paired electrons → repulsion)

Why It Matters:

Half-filled and fully-filled subshells have extra stability, which affects ionization energy.



Ionization Energy

- Practice

The first ionization energy for sulfur is 1005 kJ/mol, and that for chlorine is 1255 kJ/mol. Why?

Page 22 Annotation

Title: Practice – Sulfur vs Chlorine

Clear Explanation:

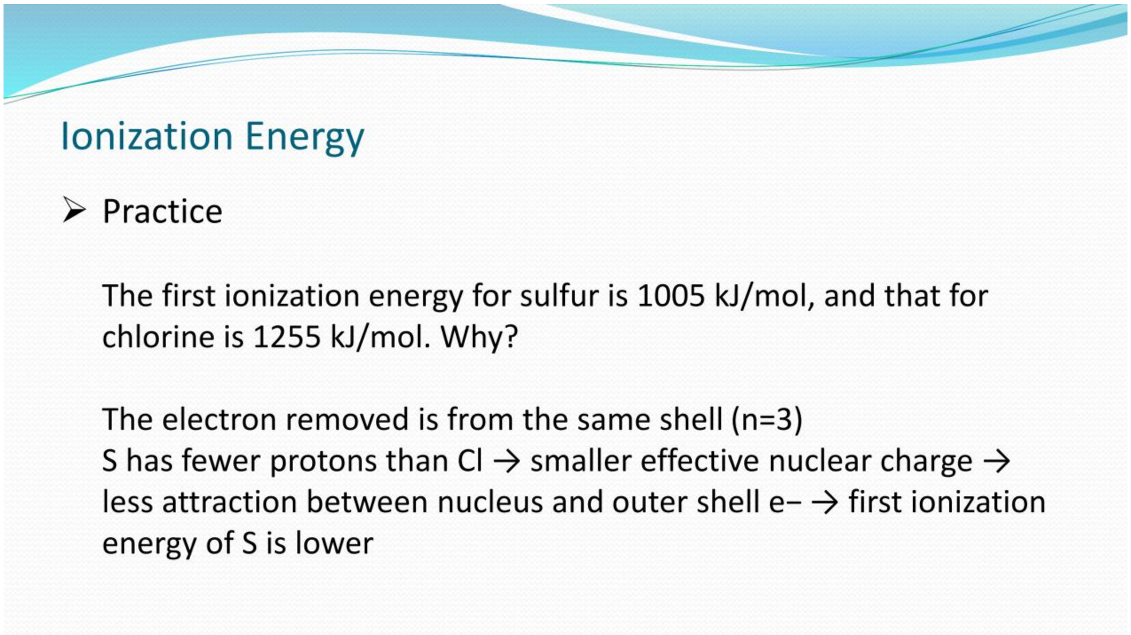
The first ionization energy for sulfur is 1005 kJ/mol, and for chlorine it is 1255 kJ/mol. Why is sulfur’s value lower?

Key Points:

- Both elements are in the same period (n = 3)
- Chlorine has more protons → higher effective nuclear charge
- Stronger attraction in Cl makes it harder to remove an electron

Why It Matters:

This practice question tests your understanding of the across-a-period trend.



Ionization Energy

- Practice

The first ionization energy for sulfur is 1005 kJ/mol, and that for chlorine is 1255 kJ/mol. Why?

The electron removed is from the same shell (n=3)
S has fewer protons than Cl → smaller effective nuclear charge → less attraction between nucleus and outer shell e− → first ionization energy of S is lower

Page 23 Annotation

Title: Practice Answer – Sulfur vs Chlorine

Clear Explanation:

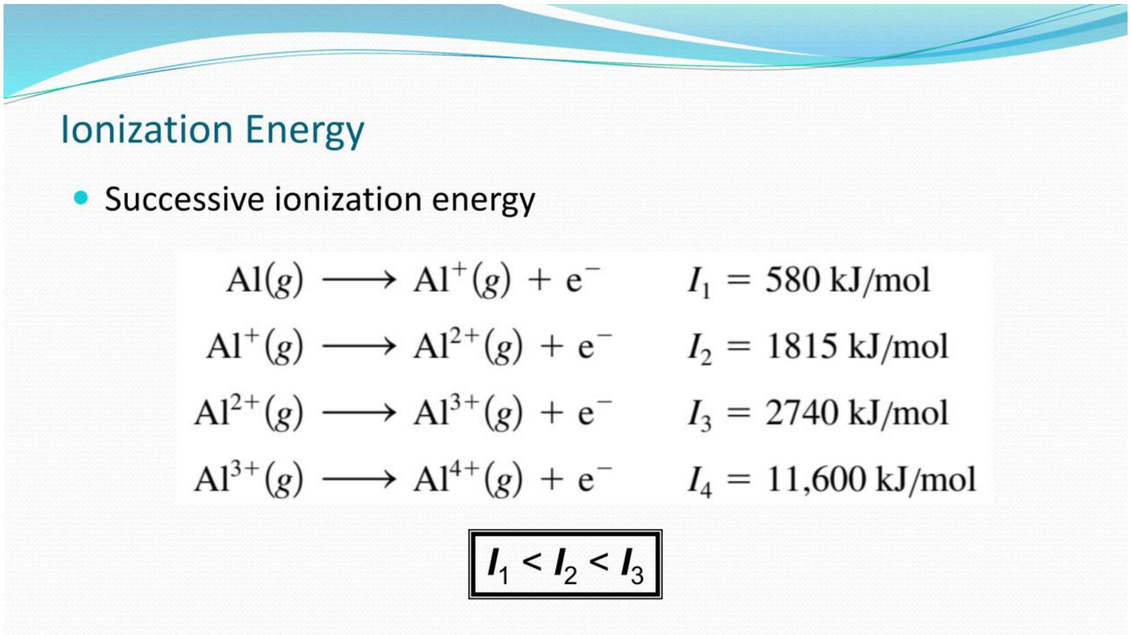
Answer:
The electron is removed from the same shell (n=3) in both atoms.
Sulfur has fewer protons than chlorine, so it has a smaller effective nuclear charge.
Weaker attraction between the nucleus and the outer electrons → lower ionization energy for sulfur.

Key Points:

- Same shell, different nuclear charge
- Z_eff is the key factor

Why It Matters:

Reinforces the main reason for the left-to-right increase in ionization energy.



Ionization Energy

- Successive ionization energy

$$\mathrm{Al}(g) \longrightarrow \mathrm{Al}^{+}(g) + e^{-} \qquad I_1 = 580\ \mathrm{kJ/mol}$$

$$\mathrm{Al}^{+}(g) \longrightarrow \mathrm{Al}^{2+}(g) + e^{-} \qquad I_2 = 1815\ \mathrm{kJ/mol}$$

$$\mathrm{Al}^{2+}(g) \longrightarrow \mathrm{Al}^{3+}(g) + e^{-} \qquad I_3 = 2740\ \mathrm{kJ/mol}$$

$$\mathrm{Al}^{3+}(g) \longrightarrow \mathrm{Al}^{4+}(g) + e^{-} \qquad I_4 = 11{,}600\ \mathrm{kJ/mol}$$

$$\boxed{I_1 < I_2 < I_3}$$

Page 24 Annotation

Title: Successive Ionization Energies

Clear Explanation:

Successive ionization energies are the energies required to remove the 2nd, 3rd, 4th… electrons.

Example for aluminum:

I_1 = 580 kJ/mol, I_2 = 1815 kJ/mol, I_3 = 2740 kJ/mol, I_4 = 11,600 kJ/mol

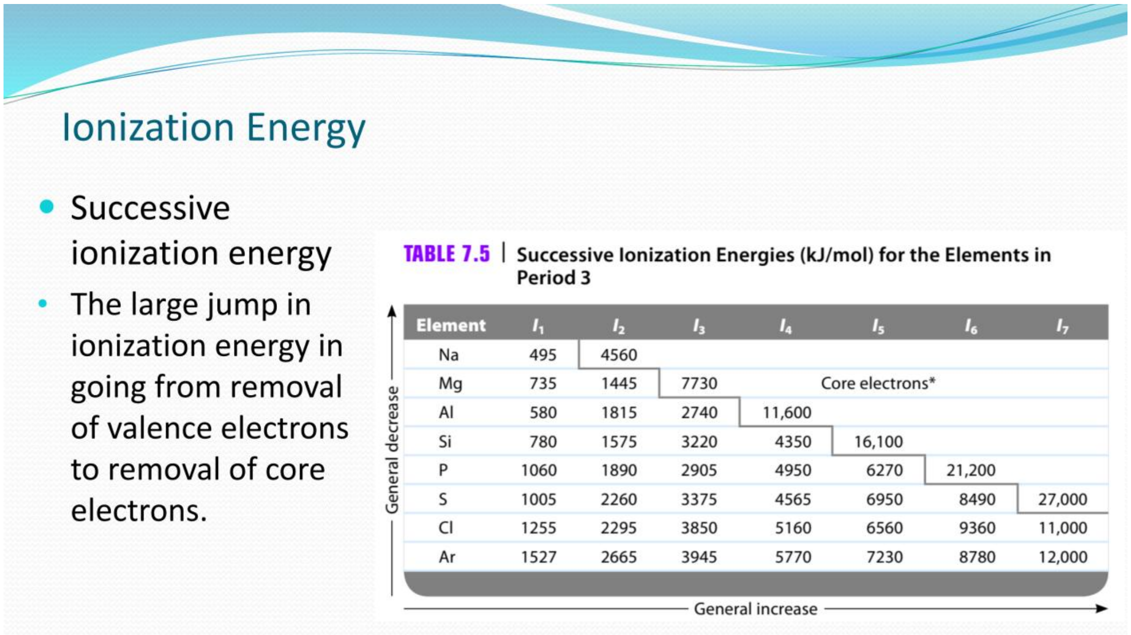
Notice that $I_1 < I_2 < I_3$, and there is a huge jump when a core electron is removed.

Key Points:

- Each successive ionization energy is larger
- Very large jump when removing core (inner-shell) electrons

Why It Matters:

The big jump helps identify how many valence electrons an element has.



Ionization Energy

- Successive ionization energy
- The large jump in ionization energy in going from removal of valence electrons to removal of core electrons.

TABLE 7.5 | Successive Ionization Energies (kJ/mol) for the Elements in Period 3

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	495	4560					
Mg	735	1445	7730		Core electrons*		
Al	580	1815	2740	11,600			
Si	780	1575	3220	4350	16,100		
P	1060	1890	2905	4950	6270	21,200	
S	1005	2260	3375	4565	6950	8490	27,000
Cl	1255	2295	3850	5160	6560	9360	11,000
Ar	1527	2665	3945	5770	7230	8780	12,000

Page 25 Annotation

Title: Successive Ionization Energies – Table for Period 3

Clear Explanation:

This table shows the first seven ionization energies for elements in Period 3.

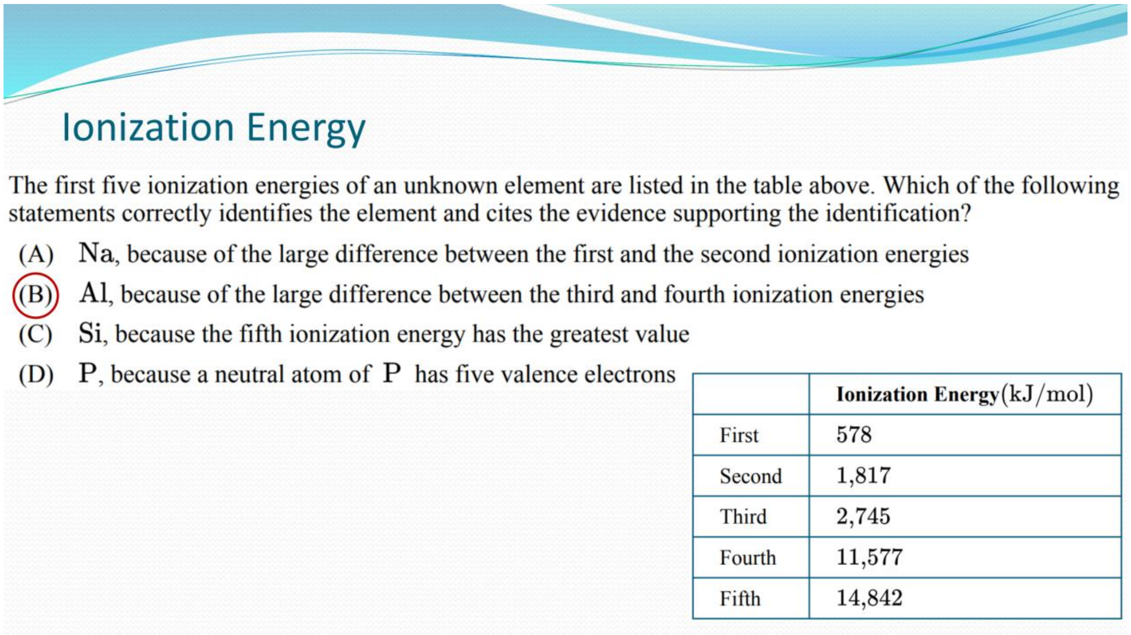
Notice the very large increase when moving from valence electrons to core electrons (e.g., from I_3 to I_4 for Al).

Key Points:

- Gradual increase within valence electrons
- Huge jump when core electrons are removed
- The number of valence electrons = number of small I values before the big jump

Why It Matters:

This table is classic evidence for electron shell structure and is often used in exam questions.



Ionization Energy

The first five ionization energies of an unknown element are listed in the table above. Which of the following statements correctly identifies the element and cites the evidence supporting the identification?

(A) Na, because of the large difference between the first and the second ionization energies

[✓ (B)] Al, because of the large difference between the third and fourth ionization energies

(C) Si, because the fifth ionization energy has the greatest value

(D) P, because a neutral atom of P has five valence electrons

	Ionization Energy (kJ/mol)
First	578
Second	1,817
Third	2,745
Fourth	11,577
Fifth	14,842

Page 26 Annotation

Title: Ionization Energy – Practice Question

Clear Explanation:

The first five ionization energies of an unknown element are given in a table.

Which statement correctly identifies the element?

Options:

(A) Na

(B) Al

(C) Si

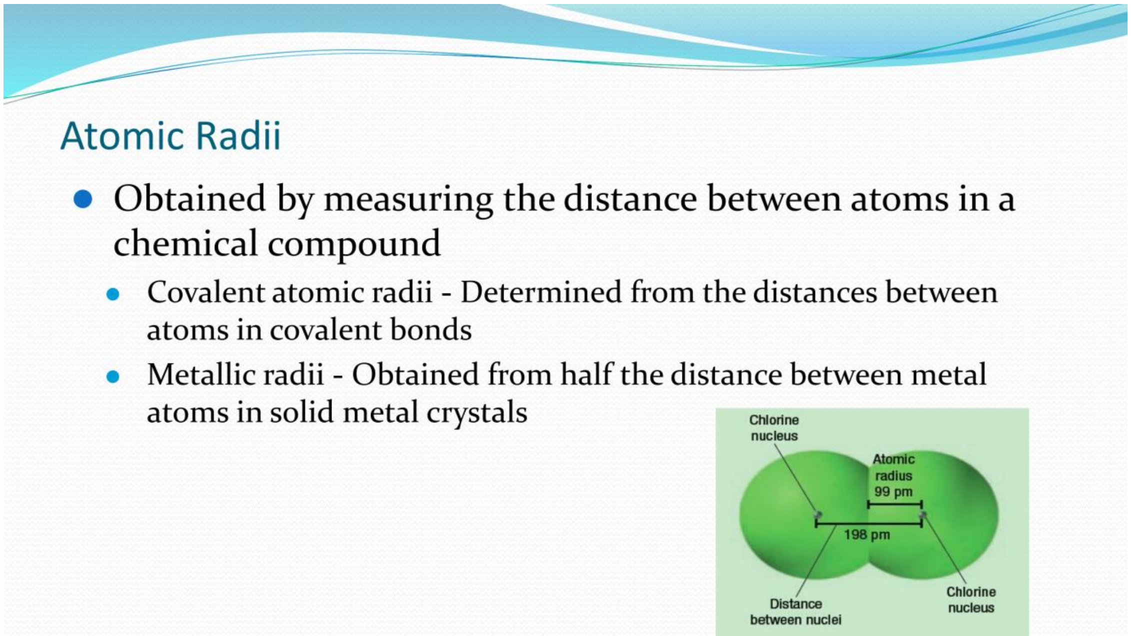
(D) P

Key Points:

- Look for the biggest jump in ionization energy values.
- The huge jump happens when a core electron (inner shell) is removed.
- For this element, the big jump is between the 3rd and 4th ionization energies → it has 3 valence electrons.

Why It Matters:

This is a classic AP-style question that tests your understanding of successive ionization energies and valence electrons.



Atomic Radii

- Obtained by measuring the distance between atoms in a chemical compound
  - Covalent atomic radii - Determined from the distances between atoms in covalent bonds
  - Metallic radii - Obtained from half the distance between metal atoms in solid metal crystals

Page 27 Annotation

Title: Atomic Radii – Definition and Measurement

Clear Explanation:

Atomic radius is the size of an atom, usually measured as half the distance between the nuclei of two identical atoms bonded together. There are two common ways to measure it:

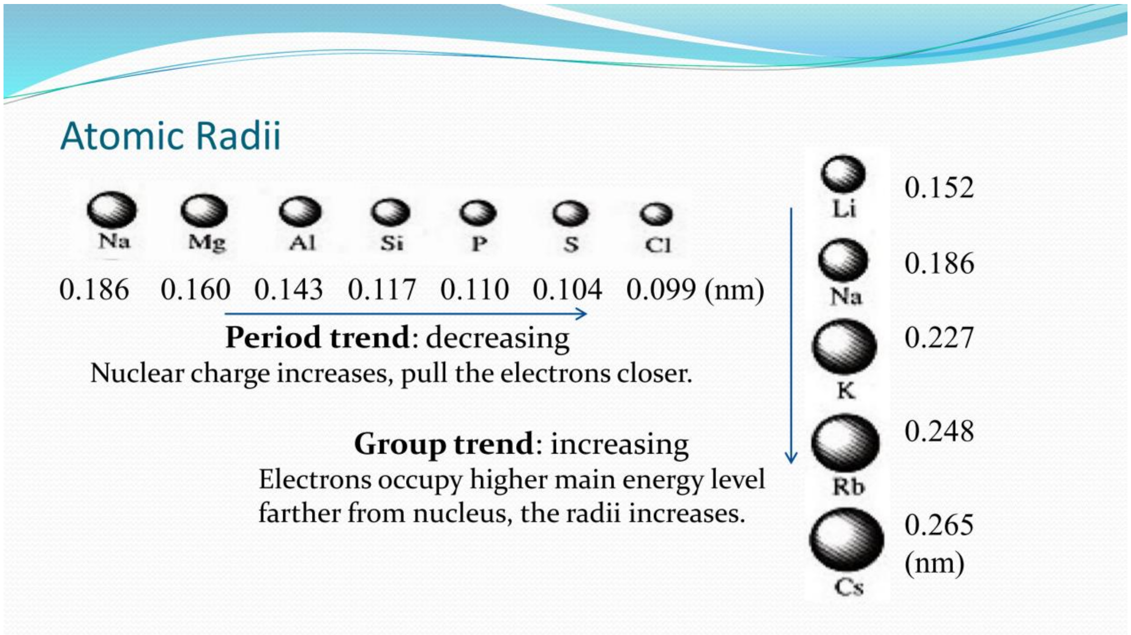
- **Covalent atomic radius**: measured from covalent bonds (nonmetals)
- **Metallic radius**: measured from the distance between metal atoms in a solid crystal (half the distance)

Key Points:

- Measured in picometers (pm) or nanometers (nm)
- Different measurement methods for metals vs nonmetals

Why It Matters:

Atomic radius is one of the four major periodic trends and directly affects many other properties (ionization energy, reactivity, etc.).



Atomic Radii

Na	Mg	Al	Si	P	S	Cl
0.186	0.160	0.143	0.117	0.110	0.104	0.099 (nm)

Period trend: decreasing
Nuclear charge increases, pull the electrons closer.

Group trend: increasing
Electrons occupy higher main energy level
farther from nucleus, the radii increases.

Li	0.152
Na	0.186
K	0.227
Rb	0.248
Cs	0.265 (nm)

Page 28 Annotation

Title: Atomic Radii – Period and Group Trends (with graph)

Clear Explanation:

The graph shows atomic radii for elements across periods and down groups.

Period trend: Atomic radius **decreases** from left to right.

Group trend: Atomic radius **increases** from top to bottom.

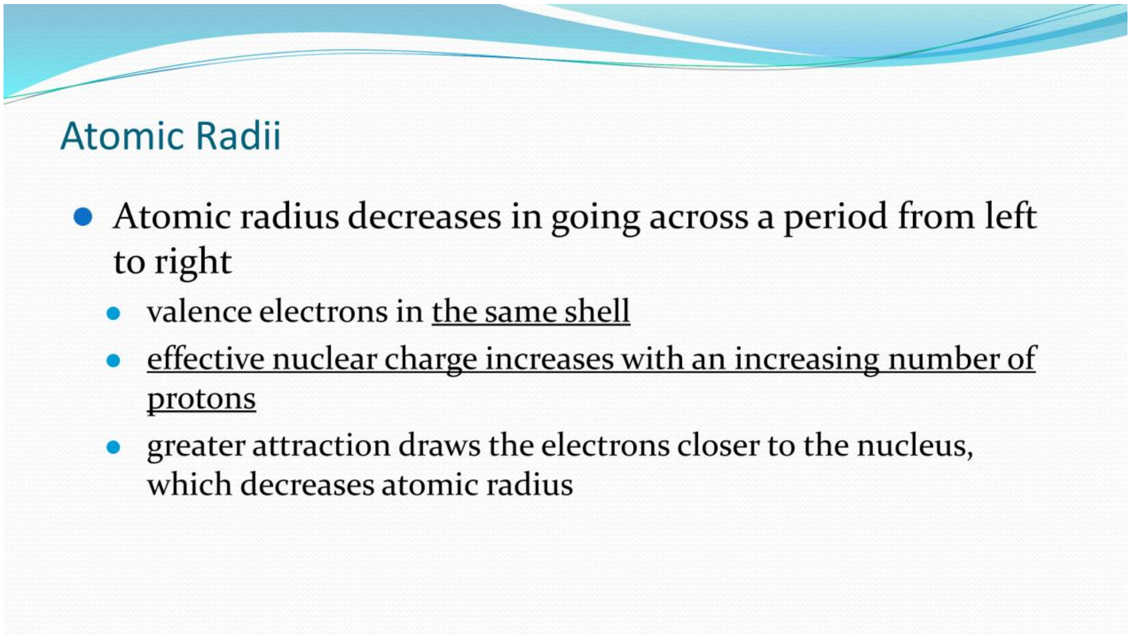
Key Points:

- Across a period: nuclear charge increases while electrons are in the same shell → stronger pull
- Down a group: new shells are added → valence electrons are farther from nucleus

Why It Matters:

This visual clearly shows the two main directions of the atomic radius trend.

Periodic Trends — Page 29



Atomic Radii

- Atomic radius decreases in going across a period from left to right
  - valence electrons in the same shell
  - effective nuclear charge increases with an increasing number of protons
  - greater attraction draws the electrons closer to the nucleus, which decreases atomic radius

Page 29 Annotation

Title: Atomic Radii – Trend Across a Period

Clear Explanation:

Atomic radius **decreases** when going from left to right across a period.

Reasons:

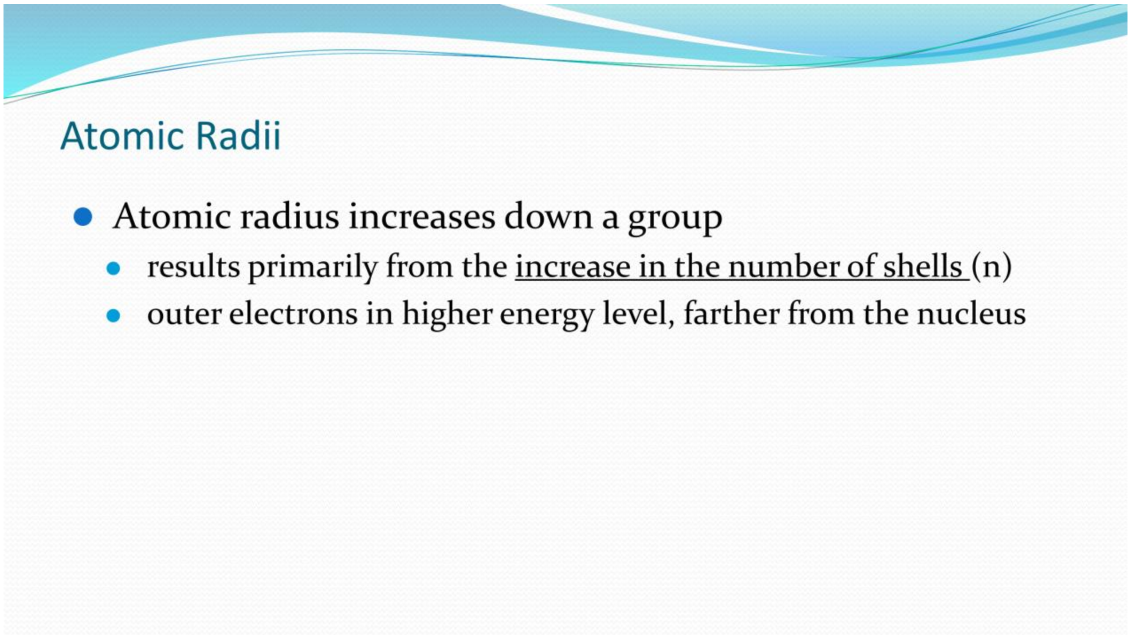
- Valence electrons are in the **same shell** (same distance from nucleus)
- Number of protons increases → effective nuclear charge (Z_eff) increases
- Stronger attraction pulls the electrons closer to the nucleus

Key Points:

- Same shell
- Increasing Z_eff
- Electrons are drawn closer

Why It Matters:

This explains why atoms on the right side of the periodic table are smaller.



Atomic Radii

- Atomic radius increases down a group
  - results primarily from the increase in the number of shells (n)
  - outer electrons in higher energy level, farther from the nucleus

Page 30 Annotation

Title: Atomic Radii – Trend Down a Group

Clear Explanation:

Atomic radius **increases** when going down a group.

Reasons:

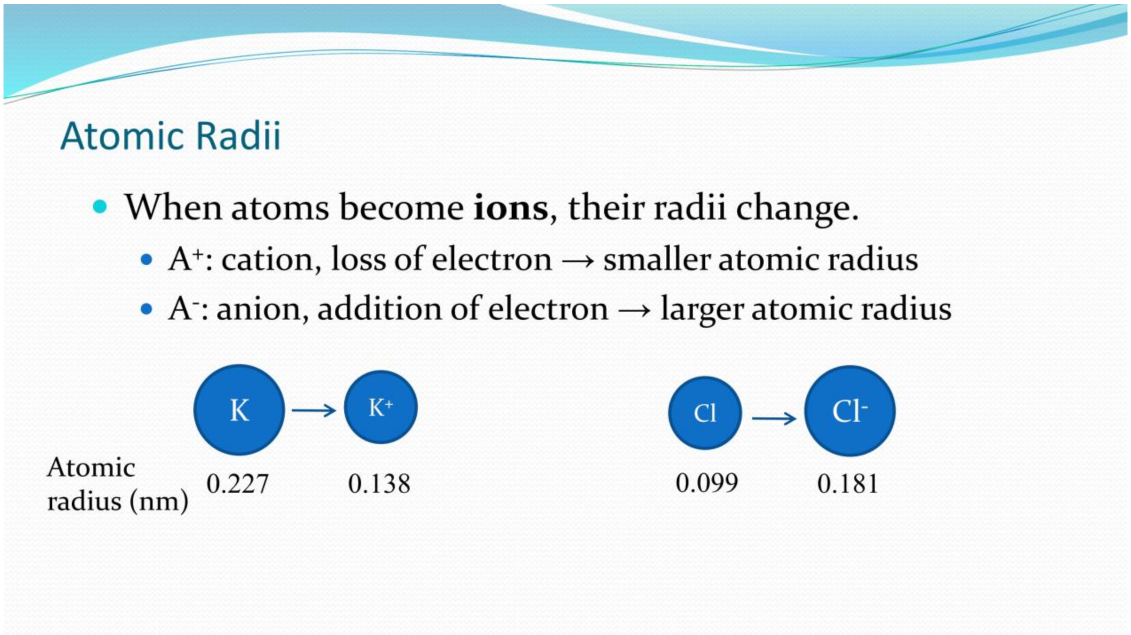
- Each new period adds a new electron shell (higher principal quantum number n)
- Valence electrons are in a higher energy level, farther from the nucleus
- Shielding effect increases

Key Points:

- More electron shells
- Valence electrons farther from nucleus
- Weaker attraction

Why It Matters:

This is why atoms get larger as you go down any group.



Atomic Radii

- When atoms become **ions**, their radii change.
  - A^+: cation, loss of electron → smaller atomic radius
  - A^-: anion, addition of electron → larger atomic radius

Page 31 Annotation

Title: Atomic Radii of Ions

Clear Explanation:

When atoms form ions, their radii change:

- **Cation (A^+)**: loses electrons → **smaller** radius
- **Anion (A^-)**: gains electrons → **larger** radius

Example data:

K (0.227 nm) → K^+ (0.138 nm)
Cl (0.099 nm) → Cl^- (0.181 nm)

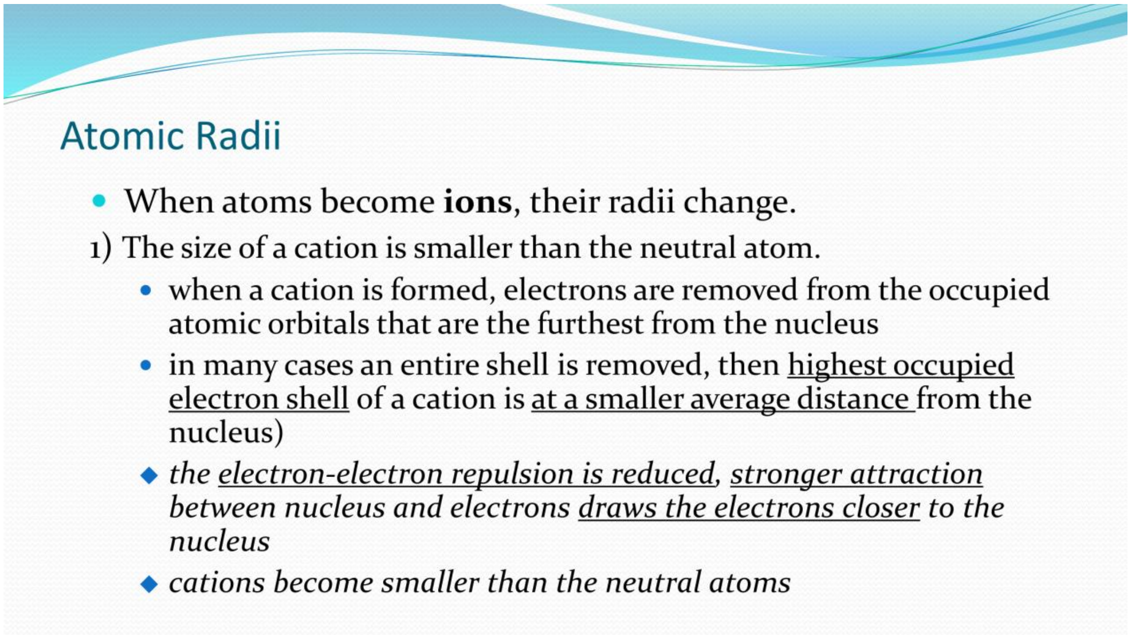
Key Points:

- Cations are smaller than neutral atoms
- Anions are larger than neutral atoms

Why It Matters:

Ionic radius is important for understanding ionic compounds, lattice energy, and solubility.

Periodic Trends — Page 32



Atomic Radii

- When atoms become **ions**, their radii change.

1) The size of a cation is smaller than the neutral atom.
   - when a cation is formed, electrons are removed from the occupied atomic orbitals that are the furthest from the nucleus
   - in many cases an entire shell is removed, then highest occupied electron shell of a cation is at a smaller average distance from the nucleus)
   - *the electron-electron repulsion is reduced, stronger attraction between nucleus and electrons draws the electrons closer to the nucleus*
   - *cations become smaller than the neutral atoms*

Page 32 Annotation

Title: Why Cations Are Smaller

Clear Explanation:

When a cation forms:

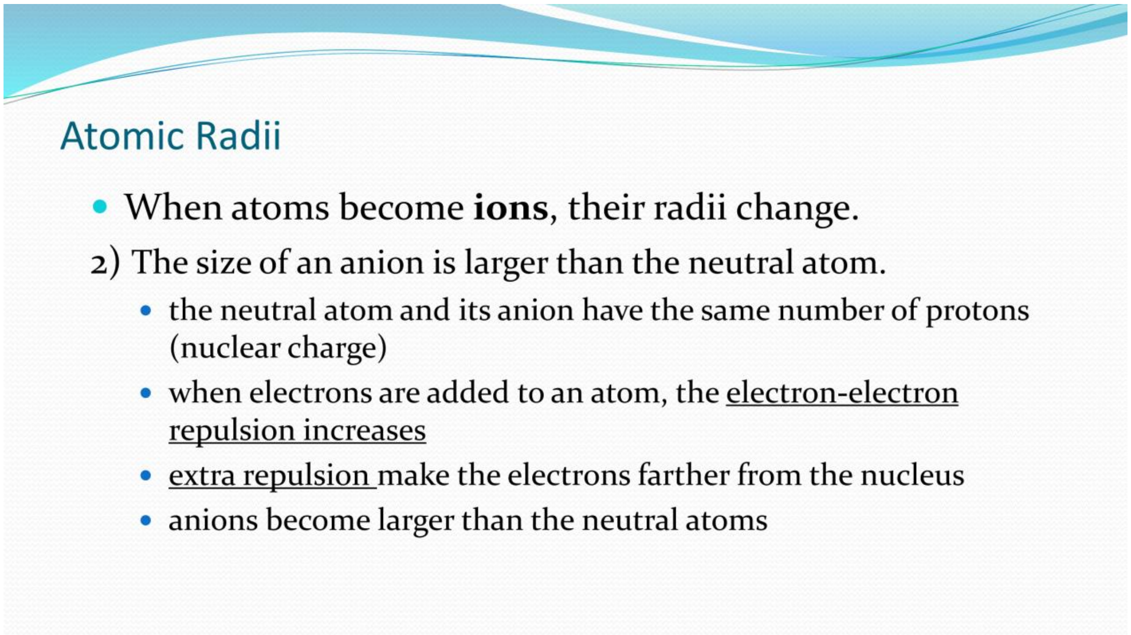
- Electrons are removed from the outermost shell
- In many cases an entire shell is lost
- Fewer electrons → less electron-electron repulsion
- The remaining electrons are pulled closer by the nucleus

Key Points:

- Loss of outer electrons
- Reduced repulsion
- Stronger nuclear attraction

Why It Matters:

This explains the dramatic size decrease when metals form positive ions.



Atomic Radii

- When atoms become **ions**, their radii change.

2) The size of an anion is larger than the neutral atom.
  - the neutral atom and its anion have the same number of protons (nuclear charge)
  - when electrons are added to an atom, the electron-electron repulsion increases
  - extra repulsion make the electrons farther from the nucleus
  - anions become larger than the neutral atoms

Page 33 Annotation

Title: Why Anions Are Larger

Clear Explanation:

When an anion forms:

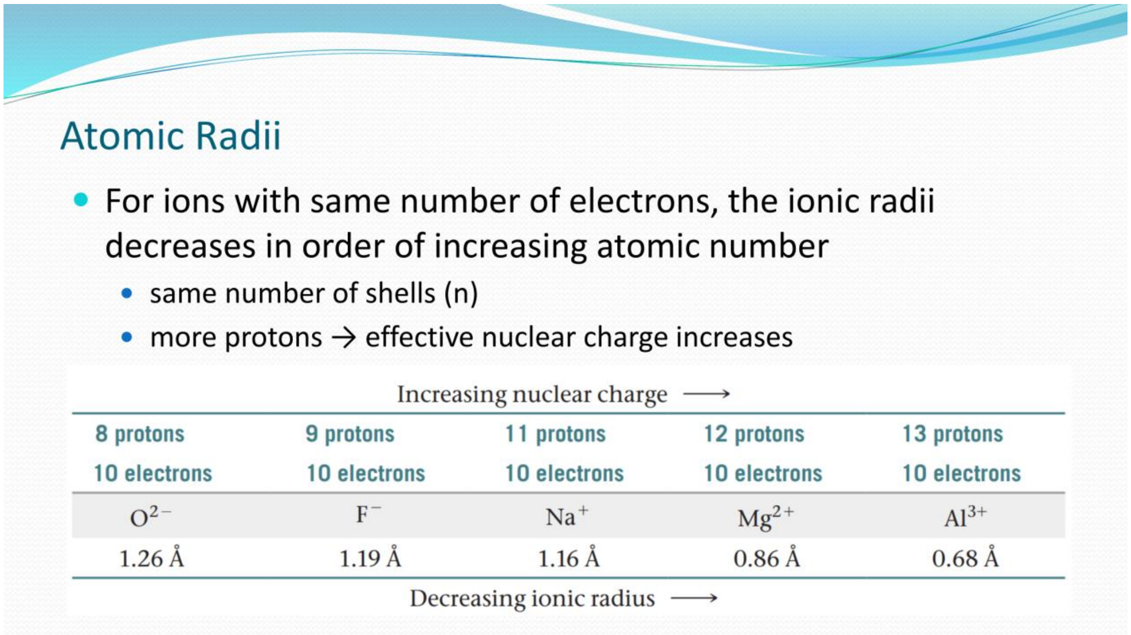
- Extra electrons are added to the same outer shell
- Same nuclear charge (same number of protons)
- Increased electron-electron repulsion
- Electrons spread out farther from the nucleus

Key Points:

- Same number of protons
- More electrons → greater repulsion
- Larger size

Why It Matters:

Nonmetals become significantly larger when they gain electrons to form anions.



Atomic Radii

- For ions with same number of electrons, the ionic radii decreases in order of increasing atomic number
  - same number of shells (n)
  - more protons → effective nuclear charge 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 ⟶

Page 34 Annotation

Title: Ionic Radii for Isoelectronic Ions

Clear Explanation:

For ions with the **same number of electrons** (isoelectronic), the ionic radius **decreases** as atomic number increases.

Example (all have 10 electrons):

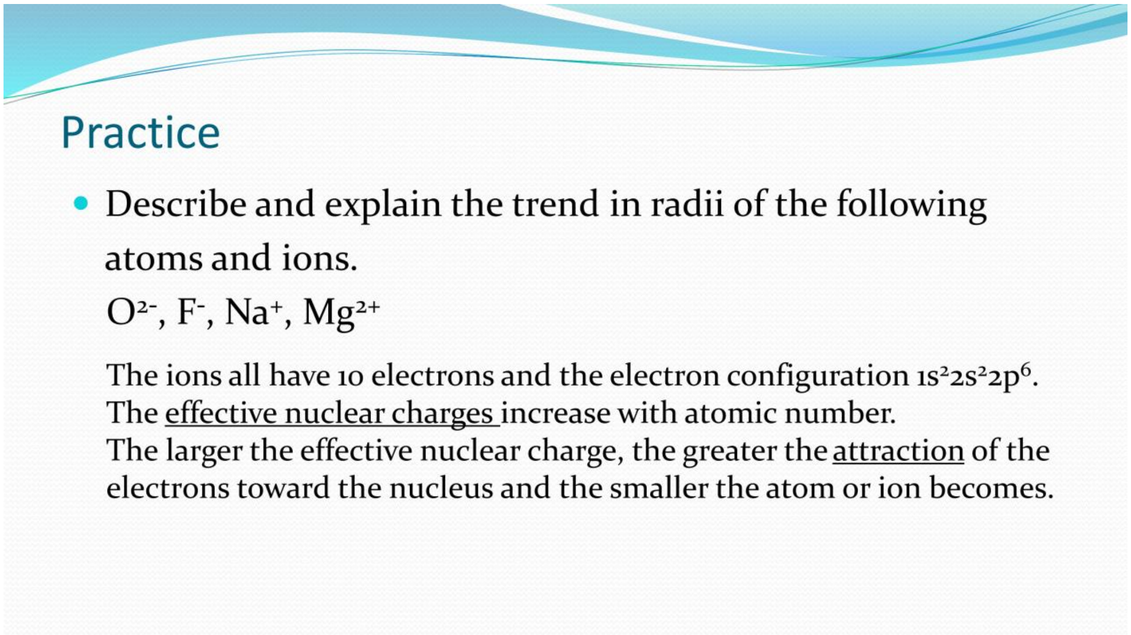
$O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$

Key Points:

- Same electron configuration ($1s^2 2s^2 2p^6$)
- More protons → higher effective nuclear charge
- Stronger attraction → smaller radius

Why It Matters:

This is a very common exam question type.



Practice

- Describe and explain the trend in radii of the following atoms and ions.
  O^{2-}, F^{-}, Na^{+}, Mg^{2+}

The ions all have 10 electrons and the electron configuration $1s^2 2s^2 2p^6$.
The effective nuclear charges increase with atomic number.
The larger the effective nuclear charge, the greater the attraction of the electrons toward the nucleus and the smaller the atom or ion becomes.

Page 35 Annotation

Title: Practice – Isoelectronic Ions

Clear Explanation:

Describe and explain the trend in radii for the isoelectronic ions: O^{2-} , F^{-} , Na^{+} , Mg^{2+} .

All have 10 electrons (same configuration $1s^2 2s^2 2p^6$).

The effective nuclear charge increases with atomic number.

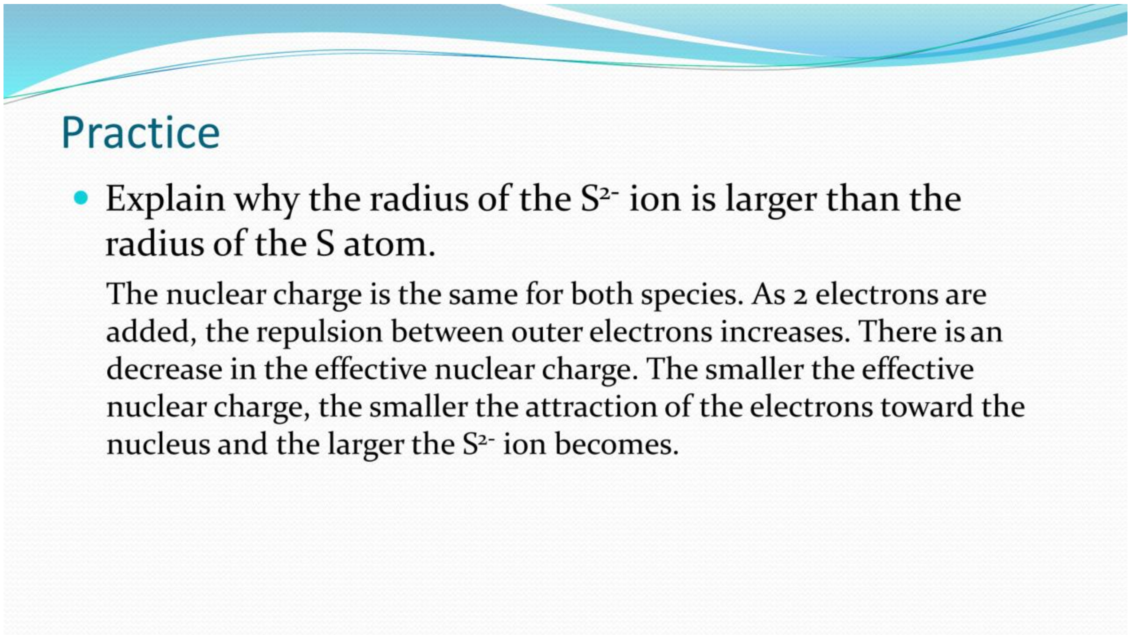
The larger the effective nuclear charge, the stronger the attraction, and the smaller the ion becomes.

Key Points:

- Same number of electrons
- Increasing nuclear charge
- Decreasing ionic radius

Why It Matters:

This practice reinforces the concept of effective nuclear charge for isoelectronic species.



Practice

- Explain why the radius of the S^{2-} ion is larger than the radius of the S atom.

  The nuclear charge is the same for both species. As 2 electrons are added, the repulsion between outer electrons increases. There is an decrease in the effective nuclear charge. The smaller the effective nuclear charge, the smaller the attraction of the electrons toward the nucleus and the larger the S^{2-} ion becomes.

Page 36 Annotation

Title: Electron Affinity

Clear Explanation:

Electron affinity is the energy released (or absorbed) when a neutral gaseous atom gains an electron to form a negative ion.

The general equation is:

$X(g) + e^- \rightarrow X^-(g)$

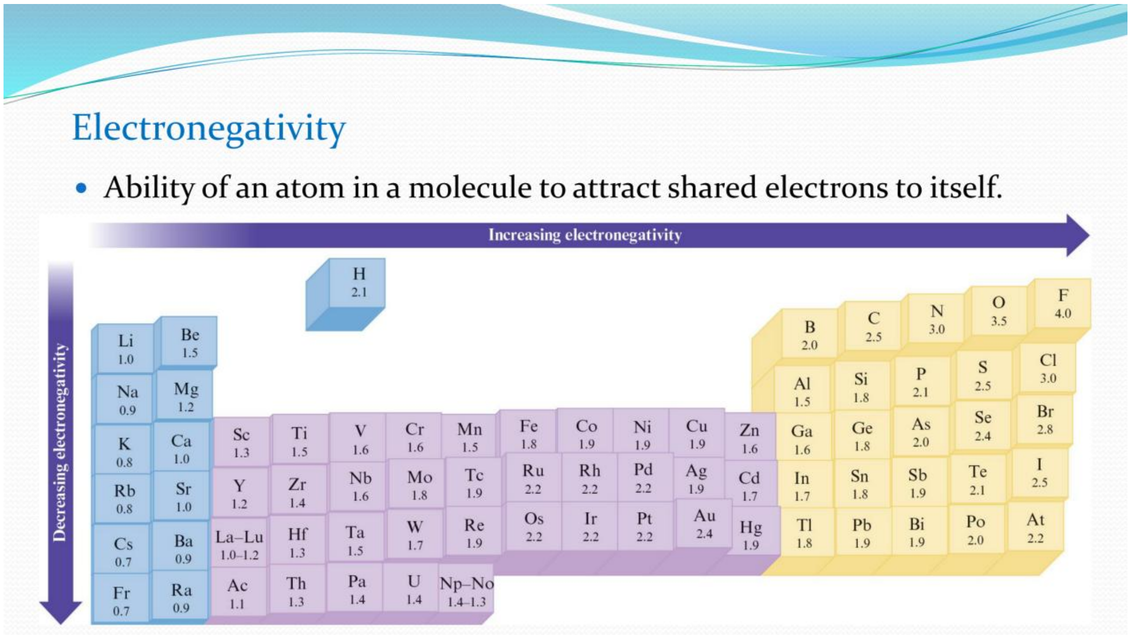
A **more negative** value means the atom releases more energy when it gains an electron (the process is more favorable).

Key Points:

- Measured in kJ/mol
- Halogens (Group 7A) have the highest (most negative) electron affinities
- Noble gases have positive (unfavorable) values

Why It Matters:

Electron affinity helps explain why nonmetals (especially halogens) readily form negative ions and are very reactive.



Electronegativity

- Ability of an atom in a molecule to attract shared electrons to itself.

Increasing electronegativity

Decreasing electronegativity

H 2.1																
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La–Lu 1.0–1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np–No 1.4–1.3										

Page 37 Annotation

Title: Electron Affinity – Periodic Trends

Clear Explanation:

Across a period (left to right): Electron affinity generally becomes **more negative** (stronger attraction for electrons).

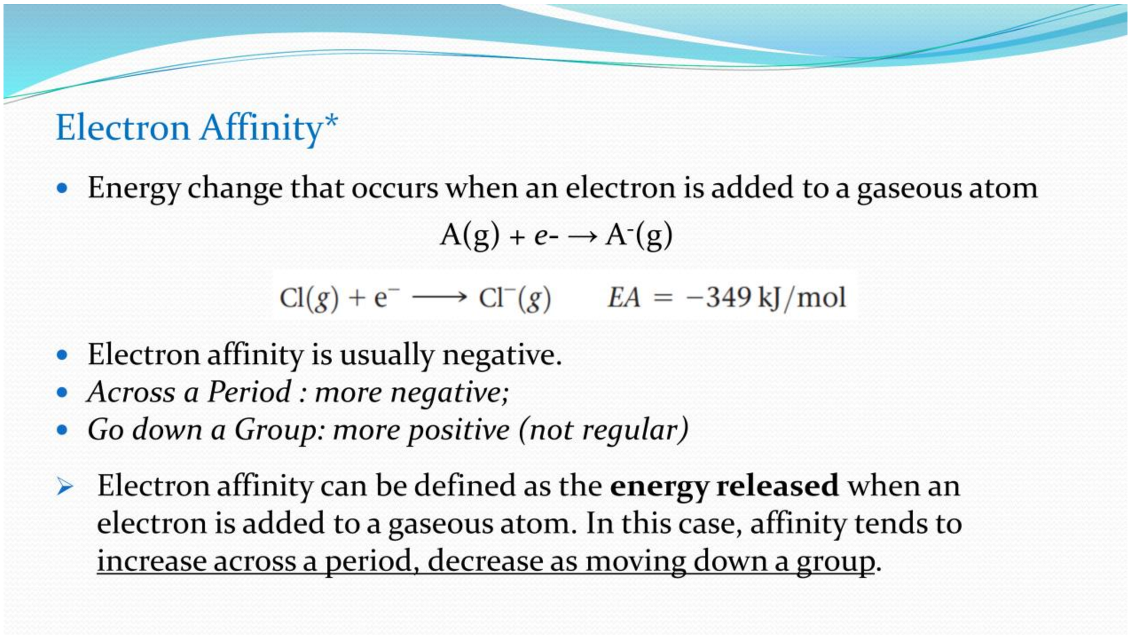
Down a group (top to bottom): Electron affinity generally becomes **less negative** (weaker attraction).

Key Points:

- Same reasons as ionization energy: increasing Z_eff across a period
- More shells down a group → weaker attraction
- Exceptions exist (e.g., F is slightly less than Cl due to small size and repulsion)

Why It Matters:

This trend explains why nonmetals on the right side of the periodic table gain electrons easily.



Electron Affinity*

- Energy change that occurs when an electron is added to a gaseous atom

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

$$Cl(g) + e^{-} \longrightarrow Cl^{-}(g) \qquad EA = -349\ \text{kJ/mol}$$

- Electron affinity is usually negative.
- *Across a Period : more negative;*
- *Go down a Group: more positive (not regular)*

➢ Electron affinity can be defined as the **energy released** when an electron is added to a gaseous atom. In this case, affinity tends to increase across a period, decrease as moving down a group.

Page 38 Annotation

Title: Electronegativity

Clear Explanation:

Electronegativity is a measure of an atom’s ability to attract shared electrons toward itself in a chemical bond.
It is a relative scale (no units). The most widely used is the **Pauling scale**, where fluorine (F) has the highest value of 4.0.

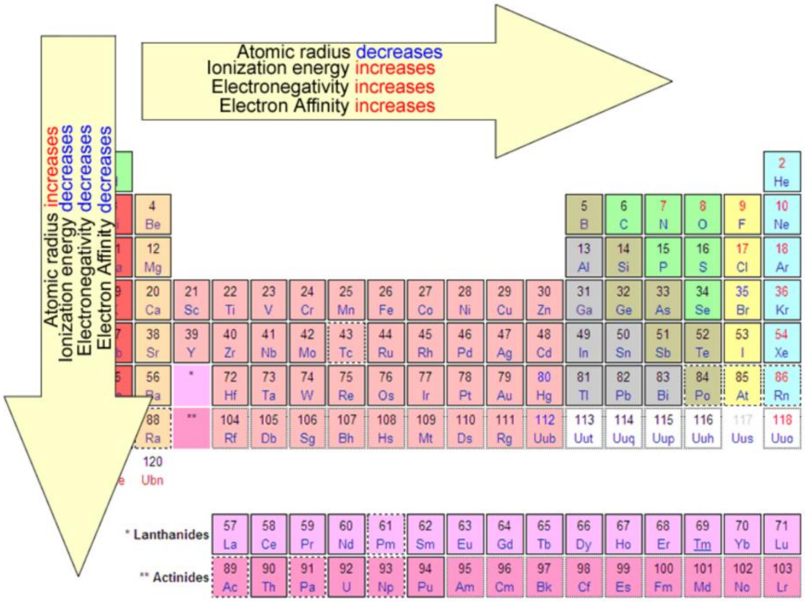
Key Points:

- No direct measurement — it is a calculated value
- Fluorine is the most electronegative element
- Cesium (Cs) is the least electronegative

Why It Matters:

Electronegativity difference between two atoms determines bond type (nonpolar covalent, polar covalent, or ionic).

Periodic Trends — Page 39



Page 39 Annotation

Title: Electronegativity – Periodic Trends

Clear Explanation:

Across a period: Electronegativity **increases** (left to right).

Down a group: Electronegativity **decreases** (top to bottom).

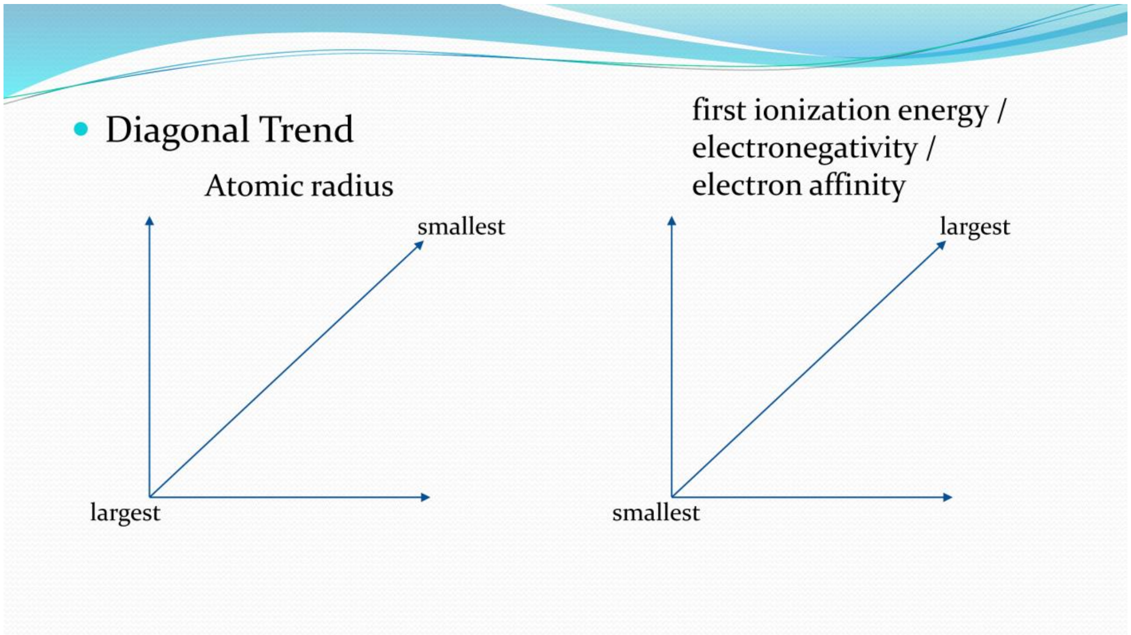
These trends are very similar to ionization energy and electron affinity because they all depend on effective nuclear charge and atomic size.

Key Points:

- Highest values in the upper right (nonmetals)
- Lowest values in the lower left (metals)
- Noble gases are usually not assigned electronegativity values

Why It Matters:

This is the most useful trend for predicting bond polarity and chemical behavior.



- Diagonal Trend

Page 40 Annotation

Title: Summary of All Four Periodic Trends

Clear Explanation:

Quick overview of the four major trends:

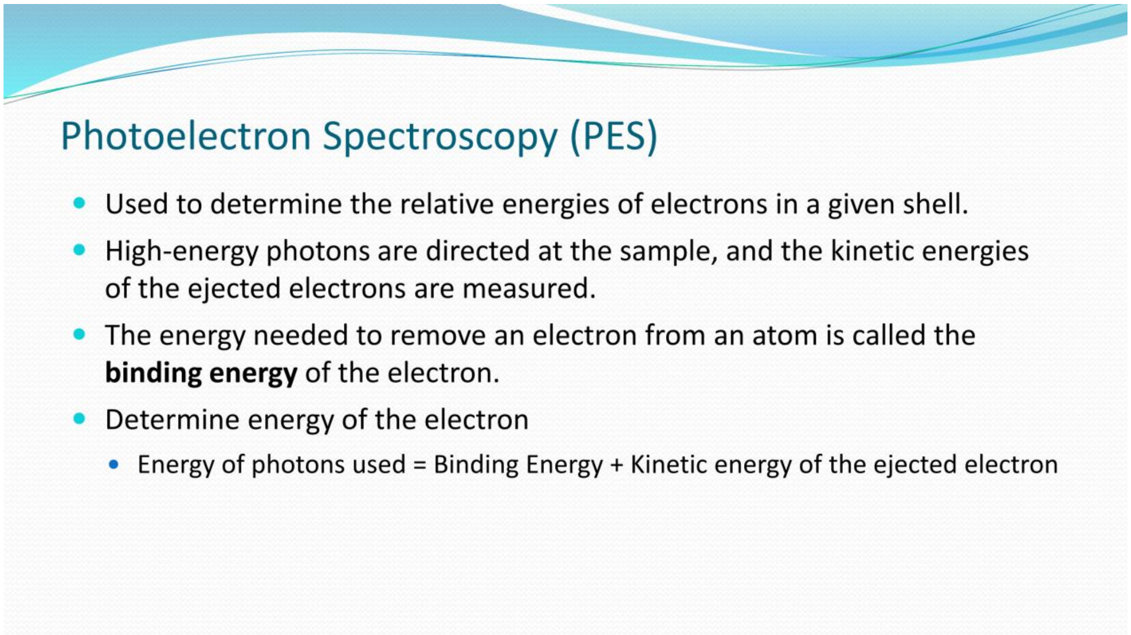
Trend	Across a Period (L→R)	Down a Group (Top→Bottom)
Atomic Radius	Decreases	Increases
Ionization Energy	Increases	Decreases
Electron Affinity	Increases (more negative)	Decreases
Electronegativity	Increases	Decreases

Key Points:

- All trends are explained by Z_eff and atomic size
- Opposite behavior for radius vs the other three

Why It Matters:

This table is extremely useful for quick review and exam questions.



Photoelectron Spectroscopy (PES)

- Used to determine the relative energies of electrons in a given shell.
- High-energy photons are directed at the sample, and the kinetic energies of the ejected electrons are measured.
- The energy needed to remove an electron from an atom is called the **binding energy** of the electron.
- Determine energy of the electron
  - Energy of photons used = Binding Energy + Kinetic energy of the ejected electron

Page 41 Annotation

Title: Practice – Predicting Trends

Clear Explanation:

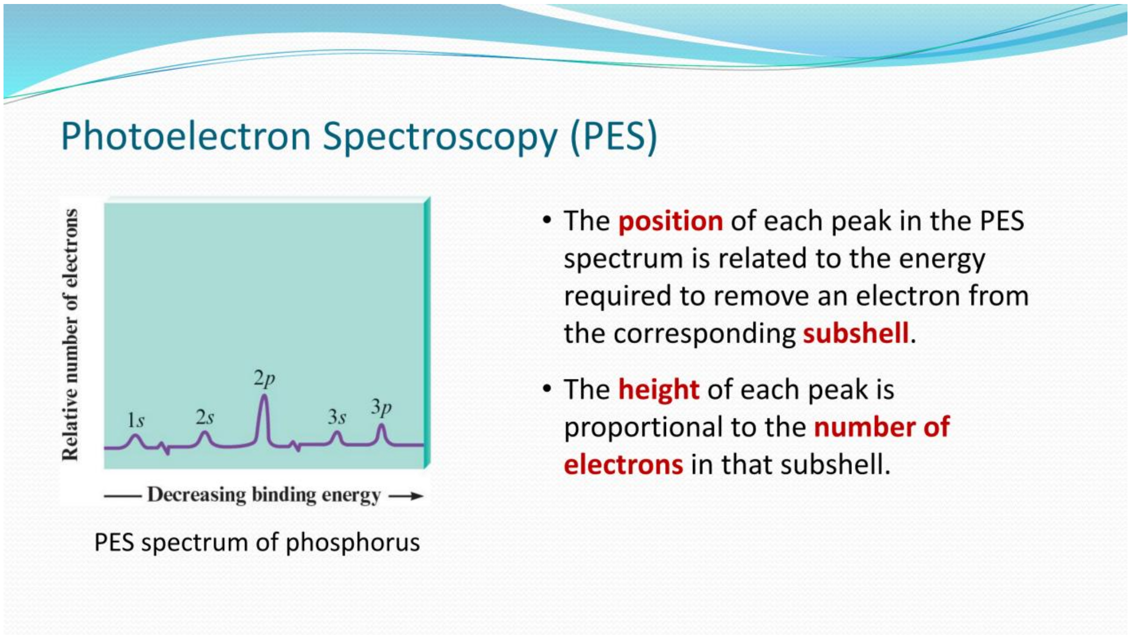
Which atom has the larger atomic radius: Na or Mg?
Which has the higher ionization energy: Cl or Br?
Which is more electronegative: O or S?

Key Points:

- Na > Mg (same period, Mg has higher Z_eff)
- Cl > Br (same group, Cl is higher up)
- O > S (same group, O is higher up)

Why It Matters:

Practice applying the trends to real elements.



Photoelectron Spectroscopy (PES)

- The **position** of each peak in the PES spectrum is related to the energy required to remove an electron from the corresponding **subshell**.
- The **height** of each peak is proportional to the **number of electrons** in that subshell.

PES spectrum of phosphorus

Page 42 Annotation

Title: Practice Answers & Explanation

Clear Explanation:

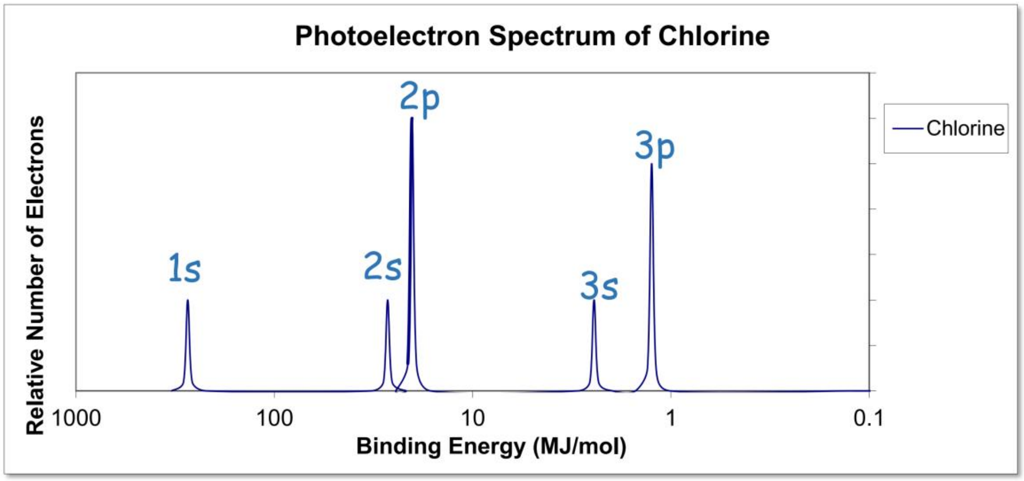
- Na has larger radius than Mg (across period, radius decreases)
- Cl has higher ionization energy than Br (down group, ionization energy decreases)
- O is more electronegative than S (down group, electronegativity decreases)

Key Points:

- Use the trends table from Page 40
- Always consider period and group position

Why It Matters:

These questions test whether you truly understand the direction of each trend.



3p peak essentially represents the first ionization energy, as an outermost electron is being removed.

Page 43 Annotation

Title: Unit Summary / Review

Clear Explanation:

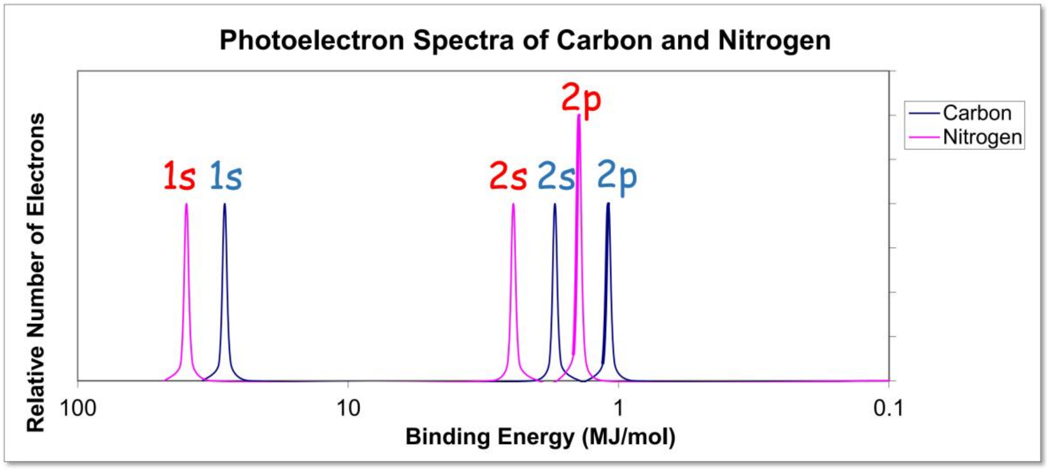
This final section reviews all key concepts: electron configuration blocks, the four periodic trends, effective nuclear charge, Coulomb’s Law, and exceptions.

Key Points:

- Electron configuration determines block and valence electrons
- All trends are driven by Z_eff and distance from nucleus
- Memorize the direction of each trend and the exceptions

Why It Matters:

You now have a complete understanding of why elements behave the way they do on the periodic table — this is foundational for all future chemistry topics.



Removal of electrons from nitrogen requires more energy, due to a greater nuclear charge.

Page 44 Annotation

Title: Unit Summary / Review

Clear Explanation:

This final section reviews all key concepts: electron configuration blocks, the four periodic trends, effective nuclear charge, Coulomb's Law, and exceptions.

Key Points:

- Electron configuration determines block and valence electrons
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- Memorize the direction of each trend and the exceptions

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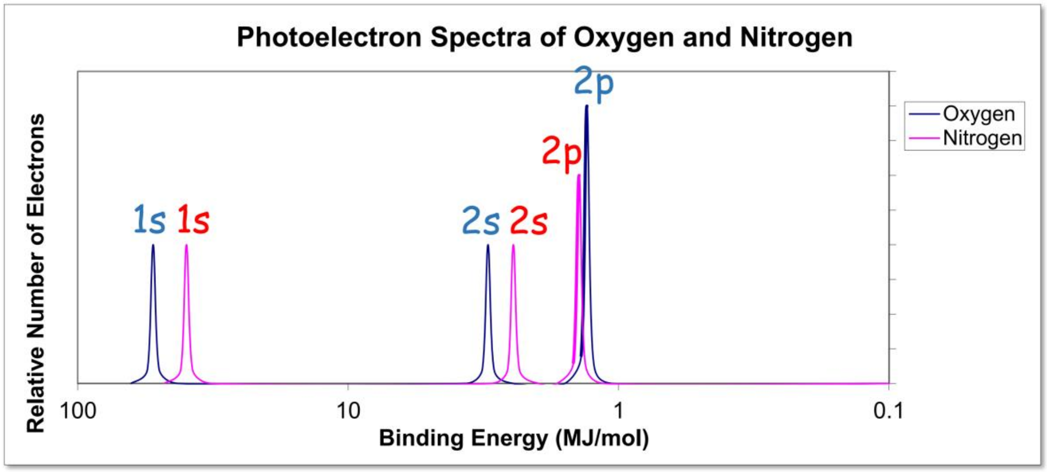
Periodic Trends — Page 45



2p electron is easier to remove in oxygen due to electron-electron repulsion is paired 2p orbital.

Page 45 Annotation

Title: Unit Summary / Review

Clear Explanation:

This final section reviews all key concepts: electron configuration blocks, the four periodic trends, effective nuclear charge, Coulomb's Law, and exceptions.

Key Points:

- Electron configuration determines block and valence electrons
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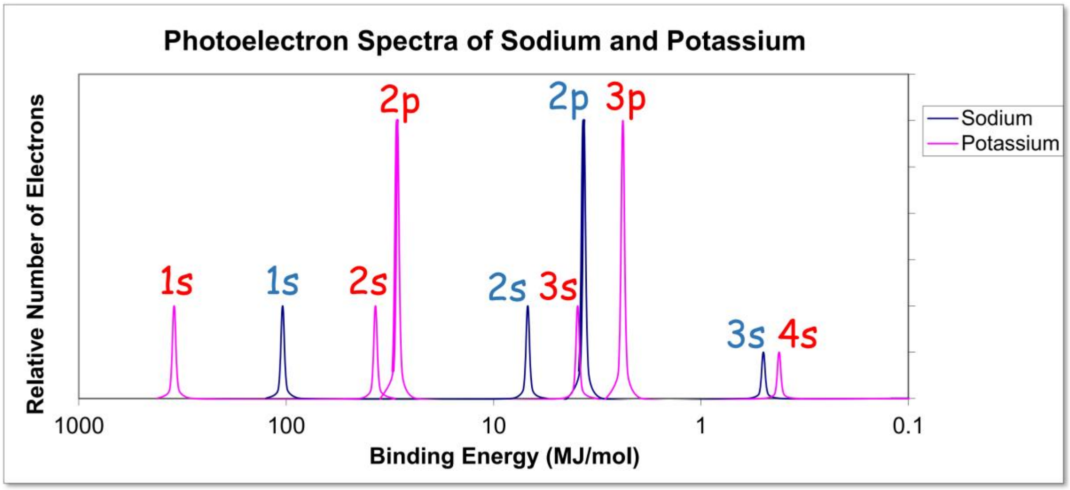

Electrons are harder to remove from potassium, due to the greater nuclear charge. The first ionization energy is smaller, because the electron is farther from the nucleus, and more shielded.

Page 46 Annotation

Title: Unit Summary / Review

Clear Explanation:

This final section reviews all key concepts: electron configuration blocks, the four periodic trends, effective nuclear charge, Coulomb’s Law, and exceptions.

Key Points:

- Electron configuration determines block and valence electrons
- All trends are driven by Z_eff and distance from nucleus
- Memorize the direction of each trend and the exceptions

Why It Matters:

You now have a complete understanding of why elements behave the way they do on the periodic table — this is foundational for all future chemistry topics.