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Periodic Trends: Radius, Ionization Energy, Electronegativity & More

Study Guide — Periodic Table & Trends

High-school/pre-med-beginner practice on periodic trends with conceptual traps: atomic/ionic radius, effective nuclear charge, shielding, first ionization energy (including common exceptions), electron affinity, electronegativity, and metallic character. Emphasis on reasoning, not memorization.

40 items — Study Guide with Answers

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1 **Across a period from left to right (main-group elements), atomic radius generally decreases primarily because:**

- A Electrons are added to new shells, pushing the outer electrons farther out
- B **The number of protons increases while electrons are added to the same shell, increasing effective nuclear charge** ✓
- C The number of neutrons decreases, shrinking the nucleus
- D Atoms lose electron shells as you move right
- E Electrons become heavier across a period

▶ **Explanation:** Moving left→right, proton number rises but added electrons go into the same principal shell, so shielding doesn't increase much. Higher effective nuclear charge pulls electrons closer, shrinking atomic radius.

2 **Down a group, atomic radius generally increases mainly because:**

- A Effective nuclear charge increases strongly down the group
- B **Electron shells are added, and shielding increases** ✓
- C Atoms lose protons down the group
- D Valence electron number increases down the group
- E Atoms become less stable and fall apart

▶ **Explanation:** Each step down adds a new principal energy level (shell). Increased distance and shielding outweigh the increased nuclear charge, so radius increases.

3 **Which element has the largest atomic radius?**

- A Li

B Na

C **K** ✓

D Mg

E Cl

▶ **Explanation:** Atomic radius increases down a group and decreases across a period. K is below Na and Li in Group 1, so it has the largest radius among the choices.

4 **Which has the smallest atomic radius?**

A Na

B Mg

C Al

D Si

E **Cl** ✓

▶ **Explanation:** Across Period 3, atomic radius generally decreases left→right as effective nuclear charge increases. Chlorine is farthest right among these, so it's smallest.

5 **Which statement best defines effective nuclear charge (Z_{eff}) as experienced by a valence electron?**

A The total number of neutrons in the nucleus

B The actual charge of the nucleus (atomic number) with no corrections

C **The net positive pull from the nucleus after accounting for shielding by inner electrons** ✓

D The attraction between two atoms in a covalent bond

E The number of electrons in the outer shell

▶ **Explanation:** Valence electrons feel less than the full nuclear charge because inner electrons repel (shield) them. Z_{eff} summarizes this net attraction.

6 **Across a period, first ionization energy generally increases because:**

- A Atomic radius increases, so electrons are harder to remove
- B **Effective nuclear charge increases, so electrons are held more tightly** ✓
- C Shielding increases dramatically, making electrons harder to remove
- D Valence electrons move into a completely new shell each step
- E Protons disappear as you move right

▶ **Explanation:** Higher Z_{eff} across a period pulls valence electrons closer and holds them more strongly, so more energy is needed to remove one (higher first ionization energy).

7 **Down a group, first ionization energy generally decreases because:**

- A The nucleus becomes less positive down the group
- B **Electrons are added to higher shells farther from the nucleus and more shielded** ✓
- C Atoms become smaller down the group
- D Valence electron number decreases down the group
- E The octet rule stops applying

▶ **Explanation:** Outer electrons are farther from the nucleus and shielded by more inner electrons, reducing the net pull and making electron removal easier (lower ionization energy).

8 **Which element is expected to have the highest first ionization energy?**

A Na

B Mg

C Al

D Cl

E **Ne** ✓

▶ **Explanation:** Ionization energy generally increases toward the top-right. Noble gases have very high ionization energies due to filled shells; Ne is highest among these choices.

9 **A student expects ionization energy to increase smoothly across Period 3 from Na to Ar. But there are small "drops" (exceptions). Which pair best represents a well-known drop in first ionization energy?**

A Na → Mg

B **Mg → Al** ✓

C Al → Si

D Si → P

E Cl → Ar

▶ **Explanation:** From Mg to Al, the first electron removed from Al is in a higher-energy p orbital (3p) compared with Mg (3s), so it's slightly easier to remove, causing a dip.

10 **Another classic dip in first ionization energy in Period 3 occurs between:**

A Na → Mg

B Mg → Al

C **P → S** ✓

D S → Cl

E Cl → Ar

▶ **Explanation:** From P to S, an electron pairing occurs in a p orbital in sulfur, increasing electron–electron repulsion slightly and making it easier to remove one electron than expected—producing a small dip.

11 **Which statement best describes electronegativity?**

A The tendency of an atom to lose electrons to form a cation

B **The tendency of an atom to attract bonding (shared) electrons** ✓

C The energy needed to remove an electron from a gaseous atom

D The number of protons in an atom

E The mass number of an atom

▶ **Explanation:** Electronegativity measures how strongly an atom attracts shared electrons in a chemical bond. It's different from ionization energy (removing electrons).

12 **Electronegativity generally:**

A Decreases left to right and increases down a group

B **Increases left to right and decreases down a group** ✓

C Increases left to right and increases down a group

D Decreases left to right and decreases down a group

E Has no relationship to periodic position

▶ **Explanation:** Moving right, Z_{eff} increases and atoms attract bonding electrons more strongly. Moving down, atoms are larger and bonding electrons are farther from the nucleus, lowering electronegativity.

13 **Which element is the most electronegative?**

A O

B **F** ✓

C Cl

D Na

E Ne

▶ **Explanation:** Fluorine is the most electronegative element: it is small and has a high effective nuclear charge on its valence shell, strongly attracting shared electrons.

14 **Which bond is expected to be the most polar (largest electronegativity difference)?**

A H–H

B C–H

C **H–F** ✓

D Cl–Cl

E N–N

▶ **Explanation:** H–F has a large electronegativity difference (F is very electronegative). Bonds between identical atoms (H–H, Cl–Cl, N–N) are nonpolar.

15 **Which element most readily forms a 1+ ion (loses one electron easily)?**

- A **Na** ✓
- B Mg
- C Al
- D Cl
- E Ar

▶ **Explanation:** Alkali metals (Group 1) have one valence electron and low ionization energy, so they most readily lose it to form 1+ ions.

16 **Which element most readily forms a 1− ion (gains one electron easily)?**

- A Na
- B Mg
- C **Cl** ✓
- D Ar
- E Al

▶ **Explanation:** Halogens (Group 17) have 7 valence electrons and strongly tend to gain 1 electron to complete an octet, forming 1− ions like Cl−.

17 **Which species is larger in radius?**

- A **Na** ✓
- B Na^+

C They are equal because they have the same nucleus

D Cannot be compared

E Na^{+} is larger because it is charged

▶ **Explanation:** Cations are smaller than their neutral atoms because they have fewer electrons (often losing an outer shell) and the remaining electrons are pulled closer by the same number of protons.

18 Which species is larger in radius?

A Cl

B **Cl—** ✓

C They are equal because they have the same nucleus

D Cl is larger because atoms are always larger than ions

E Cl— is smaller because negative charge pulls electrons inward

▶ **Explanation:** Anions are larger than their neutral atoms because adding electrons increases repulsion among electrons and reduces the effective pull per electron, expanding the electron cloud.

19 O2—, F—, Ne, Na^{+}, and Mg^{2+} are isoelectronic (10 electrons). Which is the smallest?

A O2—

B F—

C Ne

D Na^{+}

E **Mg^{2+}** ✓

▶ **Explanation:** In an isoelectronic series, more protons means stronger attraction on the same electron cloud, so radius decreases as nuclear charge increases. Mg^{2+} has the most protons.

20 **Which has the highest electronegativity?**

A Li

B C

C O

D **F** ✓

E Na

▶ **Explanation:** Fluorine is the most electronegative element, reflecting its strong attraction for shared electrons due to its small size and high effective nuclear charge.

21 **Which element is the most metallic (most likely to lose electrons) among these?**

A **Na** ✓

B Si

C P

D Cl

E Ar

▶ **Explanation:** Metallic character increases toward the left and down the periodic table. In Period 3, Na is farthest left and most readily loses an electron (very metallic).

22 **Which element is the best oxidizing agent (most likely to gain electrons) among these?**

- A Na
- B Mg
- C **Cl** ✓
- D Ar
- E Al

▶ **Explanation:** Strong oxidizers tend to gain electrons easily. Halogens (like Cl) have high electronegativity and strong tendency to gain one electron to complete an octet.

23 **Which statement best describes electron shielding?**

- A Valence electrons block protons from attracting neutrons
- B **Inner electrons repel outer electrons, reducing the nucleus's pull on valence electrons** ✓
- C Neutrons repel electrons, pushing them outward
- D Electrons become heavier in larger atoms
- E Shielding occurs only in ions, not atoms

▶ **Explanation:** Inner (core) electrons repel valence electrons and reduce how strongly valence electrons feel the nuclear attraction. This helps explain trends down groups.

24 **Which element would you expect to have the lowest first ionization energy?**

- A Li
- B Na

C **K** ✓

D Mg

E Cl

▶ **Explanation:** Group 1 elements have low ionization energies, and ionization energy decreases down the group. K is below Na and Li, so it's lowest here.

25 Which comparison is correct for atomic radius?

A Na is smaller than Cl because Na is more metallic

B **Na is larger than Cl because atomic radius decreases across a period** ✓

C Na and Cl are the same size because they are in the same period

D Cl is larger than Na because Cl has more protons

E Cl is larger than Na because electronegativity increases

▶ **Explanation:** Across a period, increasing effective nuclear charge pulls electrons closer, so atoms get smaller. Na is far left in Period 3 and is larger than Cl on the right.

26 Which set of trends is generally correct as you move left → right across a period (main group)?

A Atomic radius increases; electronegativity decreases

B **Atomic radius decreases; ionization energy increases; electronegativity increases** ✓

C Atomic radius decreases; ionization energy decreases; electronegativity decreases

D Atomic radius increases; ionization energy increases; electronegativity decreases

E No consistent trends exist

▶ **Explanation:** Across a period, Z_{eff} increases, pulling electrons closer (smaller radius) and making them harder to remove (higher ionization energy). Atoms also attract bonding electrons more strongly (higher electronegativity).

27 **Which set of trends is generally correct as you move down a group?**

- A Atomic radius decreases; ionization energy increases; electronegativity increases
- B **Atomic radius increases; ionization energy decreases; electronegativity decreases** ✓
- C Atomic radius increases; ionization energy increases; electronegativity increases
- D Atomic radius decreases; ionization energy decreases; electronegativity increases
- E Trends only apply to noble gases

▶ **Explanation:** Down a group, shells and shielding increase, so atoms get larger, lose electrons more easily (lower ionization energy), and attract bonding electrons less strongly (lower electronegativity).

28 **Which is more likely to form a negative ion based on periodic trends?**

- A Na
- B Mg
- C Al
- D **Cl** ✓
- E Ar

▶ **Explanation:** Nonmetals on the right (especially halogens) have high electronegativity/electron affinity and tend to gain electrons to complete an octet. Metals on the left tend to lose electrons.

29 **Which statement best explains why noble gases have very low tendency to gain electrons (low electron affinity, conceptually)?**

A They have empty valence shells

B They already have a full valence shell, so adding an electron would require a new, higher-energy shell ✓

C They are metals and metals cannot gain electrons

D They have no protons

E Their nuclei are negatively charged

▶ **Explanation:** Noble gases have stable filled valence shells. Adding an electron typically places it into a new shell (higher energy) and is not favorable, so their electron affinity is low.

30 **A student says: "Higher electronegativity means higher ionization energy." Which response is most accurate?**

A Always true with no exceptions

B Usually correlated across a period because both reflect stronger attraction to electrons, but they are different properties ✓

C Always false because electronegativity and ionization energy are opposites

D Only true for metals, not nonmetals

E Only true for noble gases

▶ **Explanation:** Both trends often increase across a period due to rising effective nuclear charge, so they correlate. But electronegativity refers to attracting shared electrons in a bond, while ionization energy refers to removing an electron from a gaseous atom—different concepts with different details/exceptions.

31 **Which species is expected to have the largest radius?**

A Mg^{2+}

B Na^{+}

C Ne

D F—

E **O2—** ✓

▶ **Explanation:** These are isoelectronic (10 electrons). In an isoelectronic series, fewer protons means weaker pull on the same electron cloud, giving a larger radius. O2— has the fewest protons (8), so it's largest.

32 Which element would you expect to have the strongest tendency to form a 2+ ion (lose two electrons) based on periodic position?

A Na

B **Mg** ✓

C Al

D Cl

E Ar

▶ **Explanation:** Group 2 metals (like Mg) have two valence electrons and readily lose both to reach a noble-gas configuration, forming 2+ ions.

33 Which element would generally have the greatest tendency to attract electrons in a bond among Period 3 elements?

A Na

B Mg

C Si

- D S
- E **Cl ✓**

▶ **Explanation:** Electronegativity increases across a period from left to right (excluding noble gases in many simple tables). Chlorine is farthest right among these, so it is most electronegative.

34 Which pair best illustrates that ionization energy is NOT perfectly smooth because of subshell effects?

- A Li → Be
- B **Be → B ✓**
- C C → N
- D F → Ne
- E Ne → Na

▶ **Explanation:** Be has a filled 2s subshell, while B begins filling the higher-energy 2p subshell. The first electron removed from B is a 2p electron (higher energy, easier to remove), causing a dip from Be to B.

35 Which pair best illustrates an ionization energy dip due to electron pairing in the same subshell?

- A **N → O ✓**
- B O → F
- C F → Ne
- D Li → Be
- E Be → B

▶ **Explanation:** Nitrogen has a half-filled 2p subshell (three unpaired electrons), which is relatively stable. Oxygen has one paired electron in 2p, adding repulsion and making it slightly easier to remove an electron, so ionization energy dips from N to O.

36 Why do alkali metals tend to have low first ionization energies?

A They have nearly full valence shells and strongly attract electrons

B They have one valence electron that is relatively far from the nucleus and well-shielded ✓

C They have no inner electrons, so there is no shielding

D They have the highest electronegativity in each period

E They are noble gases

▶ **Explanation:** Group 1 elements have a single outer electron. It experiences significant shielding and is relatively easy to remove, giving low first ionization energies and high reactivity.

37 Which statement best explains why fluorine is very reactive as a nonmetal?

A It easily loses electrons to form F^+

B It has low electronegativity so it doesn't attract electrons strongly

C It has high electronegativity and strongly attracts one electron to complete its valence shell ✓

D It has a full valence shell already

E It is a metal with many delocalized electrons

▶ **Explanation:** Fluorine is a halogen with 7 valence electrons. Its small size and high effective nuclear charge make it strongly attract an extra electron (or shared electrons), making it highly reactive.

38 **Which species would have the strongest attraction between its nucleus and its outer electrons (conceptually)?**

- A K
- B Na
- C Li
- D F
- E **He** ✓

▶ **Explanation:** Helium is very small and its valence electrons are in the first shell, close to the nucleus with little shielding, so the nucleus holds them very tightly. This aligns with helium's high ionization energy.

39 **Which comparison is most accurate for ionic radii in the same period?**

- A Cations are generally larger than their neutral atoms
- B Anions are generally smaller than their neutral atoms
- C **Cations are generally smaller than their neutral atoms; anions are generally larger** ✓
- D Ionic size depends only on neutrons
- E Ionic size is always the same as atomic size

▶ **Explanation:** Losing electrons (cation) decreases electron–electron repulsion and can remove an entire shell, shrinking the ion. Gaining electrons (anion) increases repulsion and expands the electron cloud.

40 **Why does electronegativity usually decrease down a group?**

- A Because the nucleus becomes less positive

B Because valence electrons move closer to the nucleus

C **Because the atom becomes larger and shielding increases, so the nucleus attracts bonding electrons less strongly** ✓

D Because the number of valence electrons changes from 1 to 8

E Because the octet rule stops applying down groups

▶ **Explanation:** Down a group, bonding electrons are farther from the nucleus and more shielded by inner electrons. The net attraction for shared electrons decreases, lowering electronegativity.