

Topic 1.7: Periodic Trends

Worksheet · AP Chemistry

UNIT 1 · ATOMIC STRUCTURE AND PROPERTIES

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How to Use This Worksheet

Work through the sections in order. Section A checks the ideas, Section B builds ranking fluency, and Section C mixes everything the way an exam would.

One rule governs this whole topic: never explain a trend by restating it. Every justification must pass through Coulomb's law: what happens to nuclear charge, shielding, shell and distance, and therefore to the attraction.

Do not look at the answer key until you have attempted every question. The key gives full working, so compare your method, not just your final answer.

Section A: Recall and Concepts

A1. Define effective nuclear charge, and explain why an electron in the same shell shields another electron only weakly while a core electron shields strongly.

A2. A sodium atom and a chlorine atom are in the same period. State which has the larger radius and explain the difference using nuclear charge and shielding.

A3. Explain, in terms of shells and electron repulsion, why Na^+ is much smaller than a sodium atom while Cl^- is much larger than a chlorine atom.

A4. State the general trends in first ionization energy across a period and down a group, and name the two Period 2 elements that break the smooth pattern across the period.

A5. The successive ionization energies of an element are 496, 4,562, 6,910 and 9,543 kJ/mol. State what the pattern tells you about the element's position in the periodic table, and why the large jump occurs where it does.

Section B: Core Practice

B1. Rank magnesium, silicon and sulfur in order of increasing atomic radius, and justify the order.

B2. S^{2-} , Cl^- , K^+ and Ca^{2+} all contain 18 electrons. Rank them in order of decreasing radius and justify the ranking.

B3. An element has successive ionization energies of 590, 1,145, 4,912 and 6,474 kJ/mol. State its group in the periodic table and explain how the values show it.

B4. Compare the effective nuclear charge felt by the valence electrons of sodium and of chlorine, and state one trend this difference explains besides atomic radius.

B5. State which is larger, Fe^{2+} or Fe^{3+} , and explain your answer in terms of electron count and attraction.

B6. Rank carbon, nitrogen, oxygen and fluorine in order of increasing electronegativity and explain the pattern.

Section C: Exam Style

C1. A student writes: "Atomic radius increases across a period because each atom has more electrons than the one before it." Identify the error and write the correct explanation, referring to shells, shielding and effective nuclear charge.

C2. The successive ionization energies of element X are 801, 2,427, 3,660, 25,026 and 32,827 kJ/mol. (a) State the group of element X. (b) Explain why the fourth value is so much larger than the third. (c) Predict, with a reason, whether the first ionization energy of the element below X in the group is larger or smaller than 801 kJ/mol.

C3. Magnesium oxide contains Mg^{2+} and O^{2-} ions; sodium fluoride contains Na^{+} and F^{-} ions. The ions are of roughly similar size in the two compounds. (a) Using Coulomb's law, compare the force of attraction between the ions in the two compounds. (b) State one observable property of the compounds that this difference in attraction helps to explain, and why.

Answer Key

A1. Effective nuclear charge is the net positive charge an electron experiences after shielding by the other electrons is taken into account. Core electrons sit between the nucleus and a valence electron and block the nuclear charge almost completely. Electrons in the same shell are at roughly the same distance from the nucleus as each other, spend little time between a given electron and the nucleus, and so shield only weakly.

A2. Sodium has the larger radius. Both atoms have their valence electrons in the $n = 3$ shell with the same core beneath, so shielding is similar. Chlorine has seven more protons, so its valence electrons feel a much larger effective nuclear charge and are pulled in much closer to the nucleus.

A3. Na^+ forms when sodium loses its only $n = 3$ electron: the whole outer shell disappears, and the remaining 10 electrons (in shells $n = 1$ and 2) are drawn in by the same 11 protons, so the ion is far smaller. Cl^- forms when chlorine gains an electron into the same $n = 3$ shell: the nuclear charge is unchanged at 17 protons, but 18 electrons repel each other more strongly and the cloud expands, so the ion is larger.

A4. Across a period, first ionization energy generally increases, because radius shrinks and effective nuclear charge rises, so the outer electron is held more tightly. Down a group it decreases, because the outer electron is in a higher shell, farther away and more shielded. In Period 2 the two exceptions are boron (lower than beryllium) and oxygen (lower than nitrogen).

A5. The huge jump between the first and second values (496 to 4,562, about nine times larger) shows the second electron is far harder to remove than the first: the first electron is a valence electron, while the second must come from a complete inner shell, much closer to the nucleus and feeling almost the full nuclear charge. The element therefore has exactly one valence electron and sits in Group 1. (The element is sodium.)

B1. Sulfur < silicon < magnesium. All three have valence electrons in the $n = 3$ shell with the same core, so the ranking follows effective nuclear charge: sulfur ($Z = 16$) pulls hardest and is smallest, magnesium ($Z = 12$) pulls weakest and is largest.

B2. $\text{S}^{2-} > \text{Cl}^- > \text{K}^+ > \text{Ca}^{2+}$. All four species have exactly 18 electrons in the same shells, so distance and shielding are equal; only the proton count differs: 16, 17, 19 and 20. More protons pulling on the same electron cloud means a smaller radius, so Ca^{2+} , with the most protons, is the smallest and S^{2-} , with the fewest, is the largest.

B3. Group 2. The first two values rise gently (590 to 1,145), then the third jumps to 4,912, over four times the second. The cliff after the second electron means two electrons are easy valence electrons and the third comes from an inner shell, so the atom has two valence electrons.

B4. Chlorine's valence electrons feel the larger effective nuclear charge: both elements have a neon-like core of 10 electrons shielding them, but chlorine's nucleus carries 17 protons against

sodium's 11. Besides radius, this also explains why chlorine's first ionization energy is far higher than sodium's (and why its electronegativity is much greater): its outer electrons are gripped much more strongly.

B5. Fe^{2+} is larger. Fe^{3+} has one fewer electron but the same nuclear charge of 26 protons, so the remaining electrons feel a stronger net pull per electron, with less electron-electron repulsion, and are drawn in closer. Removing more electrons from the same element always shrinks the ion.

B6. Carbon < nitrogen < oxygen < fluorine. All four are Period 2 elements with valence electrons in the $n = 2$ shell, so distance and shielding are similar while nuclear charge rises from 6 to 9. The bond pair is pulled progressively more strongly toward the nucleus, so electronegativity rises steadily across the period, ending at fluorine, the most electronegative element of all.

C1. The error is the claim that electron count decides size. Across a period the added electrons all enter the same outer shell, and electrons in the same shell shield each other only weakly. Meanwhile each step adds a full proton to the nucleus, so the effective nuclear charge felt by the outer shell rises steadily. By Coulomb's law the stronger attraction pulls the whole cloud inward, and the radius decreases across the period even though the atoms contain more electrons.

C2. (a) Group 13. The values rise gradually for the first three electrons (801, 2,427, 3,660) and then the fourth is nearly seven times the third, so the atom has three valence electrons. (Element X is boron.) (b) The first three electrons removed are the valence electrons ($2s^2 2p^1$ pattern, $n = 2$ shell). The fourth electron must come from the full $1s$ shell: it is far closer to the nucleus, almost unshielded, and feels nearly the full nuclear charge, so removing it takes an order of magnitude more energy. (c) Smaller. The element below (aluminium) holds its valence electrons in the $n = 3$ shell, farther from the nucleus and shielded by an extra complete shell, so despite its larger nuclear charge the outer electron is held less tightly and is easier to remove.

C3. (a) Coulomb's law: the force is proportional to the product of the charges divided by the square of the distance. The distances are similar, but the charge product in MgO is $(+2) \times (-2) = 4$ in magnitude, against $(+1) \times (-1) = 1$ in NaF . The attraction in magnesium oxide is therefore roughly four times as strong. (b) Melting point (or lattice enthalpy, or hardness). Melting requires pulling the ions apart against their attraction; the roughly fourfold stronger attraction in MgO means far more energy is needed, so MgO melts at a far higher temperature (about $2,850^\circ\text{C}$ against about 990°C for NaF).

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