

Periodic Trends

Why do atomic radius, ionization energy and electronegativity change across the periodic table?

●● Intermediate

Atomic Structure & Periodicity

Last reviewed 3 October 2026

Before you start: it helps to understand [Inside the Atom: Protons, Neutrons and Electrons](#) and [Electron Configuration](#).

What is it?

Periodic trends are the regular, predictable changes in the properties of elements as you move **across a period** (left to right) or **down a group** (top to bottom) of the periodic table. The four most important are:

- **Atomic radius:** the size of an atom.
- **First ionization energy:** the energy needed to remove one electron from each atom in a mole of gaseous atoms.
- **Electronegativity:** how strongly an atom attracts the shared electrons in a chemical bond.
- **Ionic radius:** the size of an ion.

KEY IDEA

Two factors explain almost every trend: the **number of shells** (more shells means outer electrons are farther from the nucleus) and the **effective nuclear charge** (the net positive charge an outer electron actually feels).

Why does it matter?

- **It predicts reactivity.** Group 1 metals get more reactive down the group because their outer electron is easier to remove.
- **It predicts bonding.** The difference in electronegativity between two atoms tells you whether a bond is nonpolar, polar or ionic.
- **It organises facts.** Instead of memorising the properties of 118 elements, you learn a few patterns and the reasons behind them.

How does it work?

1. Effective nuclear charge

Inner electrons **shield** outer electrons from the full nuclear charge. A simple estimate:

$$Z_{\text{eff}} \approx Z - S$$

where Z is the number of protons and S is the number of inner (core) electrons. For sodium, $Z_{\text{eff}} \approx 11 - 10 = 1$; for chlorine, in the same period, $Z_{\text{eff}} \approx 17 - 10 = 7$.

- **Across a period:** protons are added but the core stays the same, so Z_{eff} **increases**.
- **Down a group:** a new shell is added each period, so outer electrons are **farther away**.

2. Atomic radius

- **Decreases across a period:** a higher Z_{eff} pulls the same shell closer.
- **Increases down a group:** each period adds a shell.

3. First ionization energy

- **Increases across a period:** outer electrons are held more tightly.
- **Decreases down a group:** the outer electron is farther away and more shielded. Group 1: Li 520.2, Na 495.8, K 418.8, Rb 403.0, Cs 375.7 kJ/mol.

4. Electronegativity

- **Increases across a period** and **decreases down a group**, for the same reasons.
- Fluorine is the most electronegative element (3.98 on the Pauling scale); caesium and francium are the least.

5. Ionic radius

- **Cations are smaller** than their atoms: often a whole shell is lost, and the remaining electrons feel more nuclear pull.
- **Anions are larger** than their atoms: the extra electrons repel each other, and the same nuclear charge holds more electrons.
- In an **isoelectronic series** (same number of electrons), more protons means a smaller ion: $\text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+}$ (all have 10 electrons).

∞ THINK OF IT LIKE THIS

Think of the nucleus as a magnet and the outer electrons as paper clips. Across a period the magnet gets stronger while the clips stay at the same distance, so they are held more tightly. Down a group the clips sit farther away, behind more layers of other clips, so they are held more loosely.

Q MORE PRECISELY

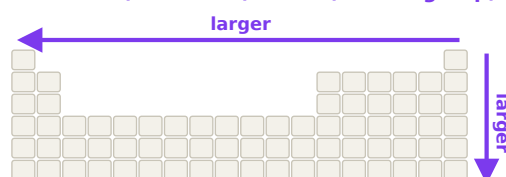
Ionization energy does not rise perfectly smoothly across a period. In period 2 there are two dips. **Boron** (800.6 kJ/mol) is lower than beryllium (899.5) because its outer electron is in a 2p orbital, slightly higher in energy than 2s. **Oxygen** (1313.9) is lower than nitrogen (1402.3) because oxygen's fourth 2p electron is paired, and repulsion between the paired electrons makes one of them easier to remove.

Visualise it

Periodic trends at a glance

ATOMIC RADIUS

increases ← (to the left) and ↓ (down a group)



IONIZATION ENERGY and ELECTRONEGATIVITY

increase → (to the right) and ↑ (up a group)

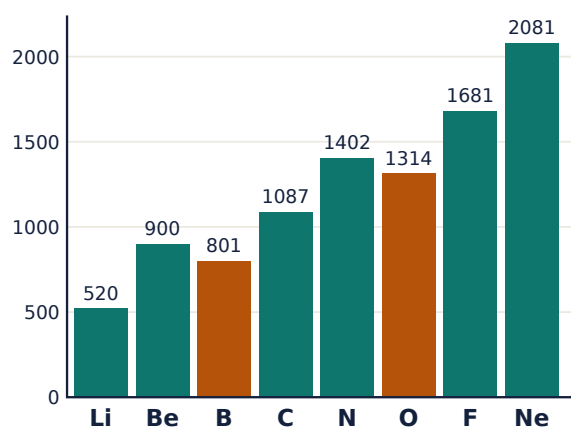


Main-group trends; noble gases have no usual electronegativity

Atomic radius grows towards the bottom left; ionization energy and electronegativity grow towards the top right.

First ionization energy, period 2

kJ/mol. The general trend rises across the period.



■ Dips at B and O (explained in the lesson)

Values rounded from NIST spectroscopic data.

First ionization energies across period 2. The general rise has two dips, at B and O.

Worked example

WORKED EXAMPLE: Ranking atomic radius

Question: Arrange Na, Mg, K and Cl in order of increasing atomic radius.

- Step 1** Na, Mg and Cl are in period 3, where radius decreases from left to right: $\text{Cl} < \text{Mg} < \text{Na}$.
- Step 2** K is directly below Na in group 1, with one more shell: $\text{K} > \text{Na}$.
- Step 3** Order: **$\text{Cl} < \text{Mg} < \text{Na} < \text{K}$**

WORKED EXAMPLE: Explaining an ionization-energy trend

Question: Why is the first ionization energy of sodium (495.8 kJ/mol) lower than that of lithium (520.2 kJ/mol)?

- Step 1** Na is $[\text{Ne}] 3s^1$ and Li is $[\text{He}] 2s^1$. Both have one outer s electron with $Z_{\text{eff}} \approx 1$.
- Step 2** Sodium's outer electron is in the **third** shell, farther from the nucleus and shielded by more inner electrons.
- Step 3** It is held less strongly, so **less energy** is needed to remove it.

Common mistake

COMMON MISTAKE: Thinking more protons make a bigger atom across a period

Across a period atoms get **smaller**, even though they have more protons and electrons. The added electrons go into the same shell, while the extra protons pull that shell in more tightly.

COMMON MISTAKE: Confusing ionization energy with electronegativity

Ionization energy is about **removing** an electron from an isolated gaseous atom. Electronegativity is about attracting **shared** electrons in a bond. They follow similar trends but measure different things.

COMMON MISTAKE: Assuming an ion is the same size as its atom

Na^+ is much smaller than Na (it has lost its whole third shell), and Cl^- is larger than Cl.

NOTATION NOTE

- Ionization energies are usually given in kJ/mol (sometimes in eV per atom).
- Electronegativity has no units; the Pauling scale is the most common.
- $Z_{\text{eff}} \approx Z - S$ is a simple estimate. More detailed methods (such as Slater's rules) give values that are not whole numbers.

Remember this

REMEMBER THIS

- Across a period: Z_{eff} increases, radius decreases, ionization energy and electronegativity increase.
- Down a group: more shells, radius increases, ionization energy and electronegativity decrease.
- Cations are smaller than their atoms; anions are larger.
- Period 2 ionization-energy dips: B (2p electron) and O (paired 2p electron).

References

1. Brown, T. L.; LeMay, H. E., Jr.; Bursten, B. E.; Murphy, C. J.; Woodward, P. M.; Stoltzfus, M. W. Chemistry: The Central Science, 15th ed.; Pearson, 2022.

Practise this topic with flashcards and a quiz at chemistryclarity.com/chemistry/periodic-trends/