

Electronic Structure & Periodic Properties

Preparation for General Chemistry · Dr. Karmach

By the end of this unit, you can...

- Relate a light wave's wavelength, frequency, and photon energy with $c = \lambda\nu$ and $E = h\nu$, and rank colors of light by each
- Explain why excited atoms emit line spectra rather than continuous ones, and compare Bohr transitions by photon energy and wavelength
- Describe the shapes of s, p, and d orbitals and count the orbitals and electrons a sublevel or shell holds
- Write full and noble-gas electron configurations for atoms through the d block and for their common ions
- Draw orbital diagrams with the building-up order, the Pauli principle, and Hund's rule, and count valence electrons from the configuration or the group number
- Predict atomic radius, ionization energy, and ion size from position on the periodic table, including an isoelectronic series

Today's route

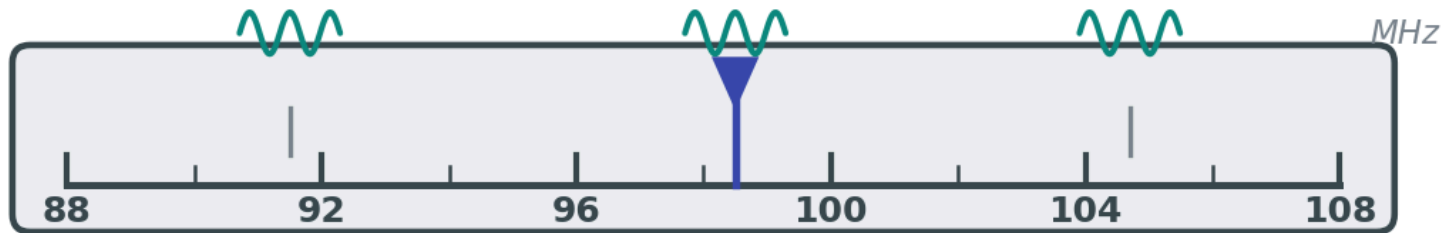
1. Wavelength & Frequency
2. Photon Energy
3. The Bohr Model & Line Spectra
4. Sublevels & Orbitals
5. Electron Configurations
6. Electron Configurations of Ions
7. Orbital Diagrams & Valence Electrons
8. Periodic Trends
9. Ion Size & Isoelectronic Series

1 · Wavelength & Frequency

Use $c = \lambda\nu$ to find a wavelength or a frequency for any kind of light, converting nanometers to meters so the units match the speed of light.

Every station, the same kind of wave

one dial, many stations

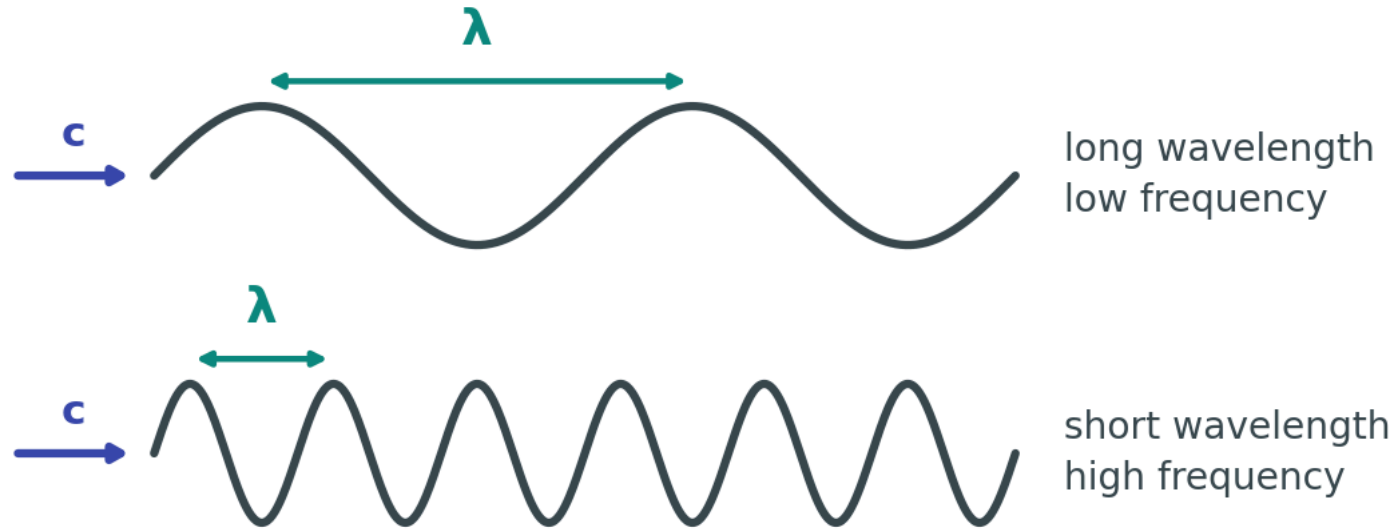


every station is the same kind of wave — only the number changes

Turn the tuner across the dial and the station changes. Every one sends the same kind of radio wave. Only the number is different.

One speed for all light

both travel at c — the shorter the wave, the more crests pass each second



Light is an electromagnetic wave. In a vacuum, every kind travels at one speed, $c = 3.00 \times 10^8$ m/s. Stretch the wave and fewer crests pass each second; squeeze it and more do.

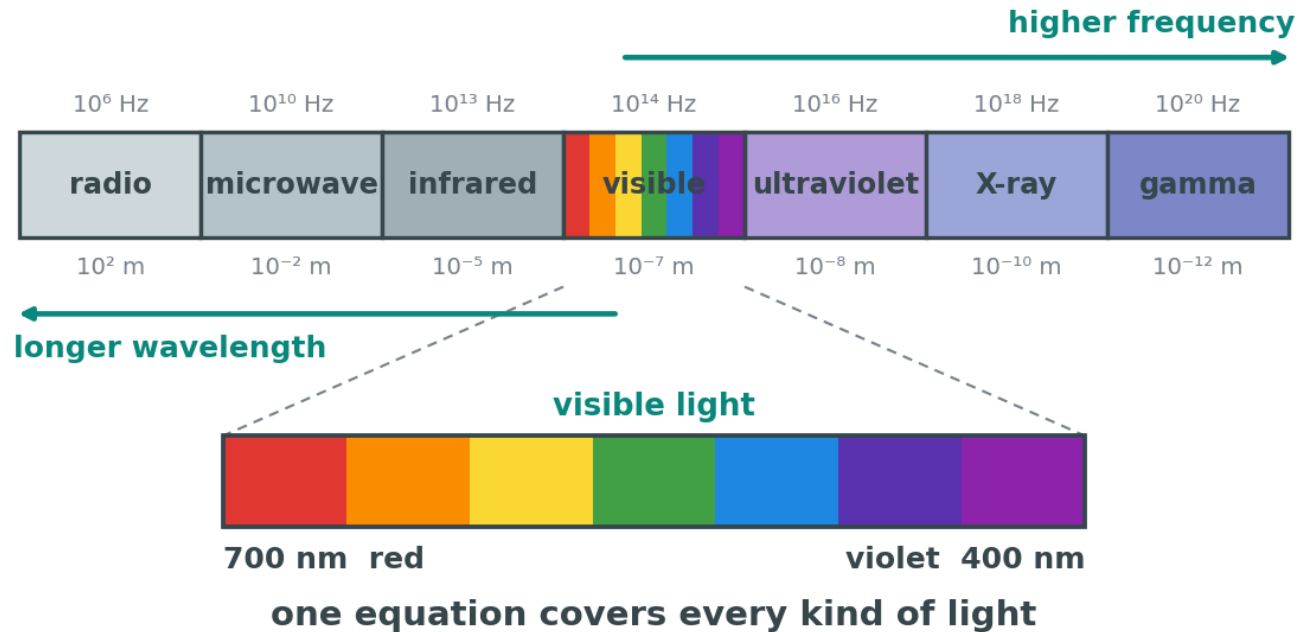
The equation: $c = \lambda \nu$

$$c = \lambda \times \nu$$

$$c = 3.00 \times 10^8 \text{ m/s} \cdot \lambda \text{ in meters (m)} \cdot \nu \text{ in s}^{-1}, \text{ also called hertz (Hz)}$$

Wavelength λ is the distance from one crest to the next, in meters. Frequency ν is how many crests pass each second, in s^{-1} . Their product is always the fixed speed c .

The electromagnetic spectrum



The same equation covers every kind of light, from radio waves meters long to gamma rays smaller than an atom. Visible light is a thin band: red near 700 nm, violet near 400 nm.

memory hook: ROY G BIV

red · orange · yellow · green · blue · indigo · violet: the visible colors, 700 nm down to 400 nm



The method

1. **Identify the unknown.** λ or ν ; the other two are given.
2. **Match units to c .** Wavelength in meters, frequency in s^{-1} .
3. **Rearrange for the unknown:** $\nu = c/\lambda$, or $\lambda = c/\nu$.
4. **Substitute and cancel units.**



Worked example 1: frequency of orange light

$$c = \lambda \nu$$

given: $\lambda = 600 \text{ nm (orange)}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

A lamp glows with **orange light** at **600 nm**. Find its frequency. ($c = 3.00 \times 10^8 \text{ m/s}$; $1 \text{ m} = 10^9 \text{ nm}$)

Identify the unknown, and match its units to c before dividing.

✓ Worked example 1: solution

$$c = \lambda \nu$$

given: $\lambda = 600 \text{ nm (orange)}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

Step 1 · Identify the unknown

Step 2 · Match units to c

Step 3 · Rearrange for the unknown

Step 4 · Substitute and cancel units

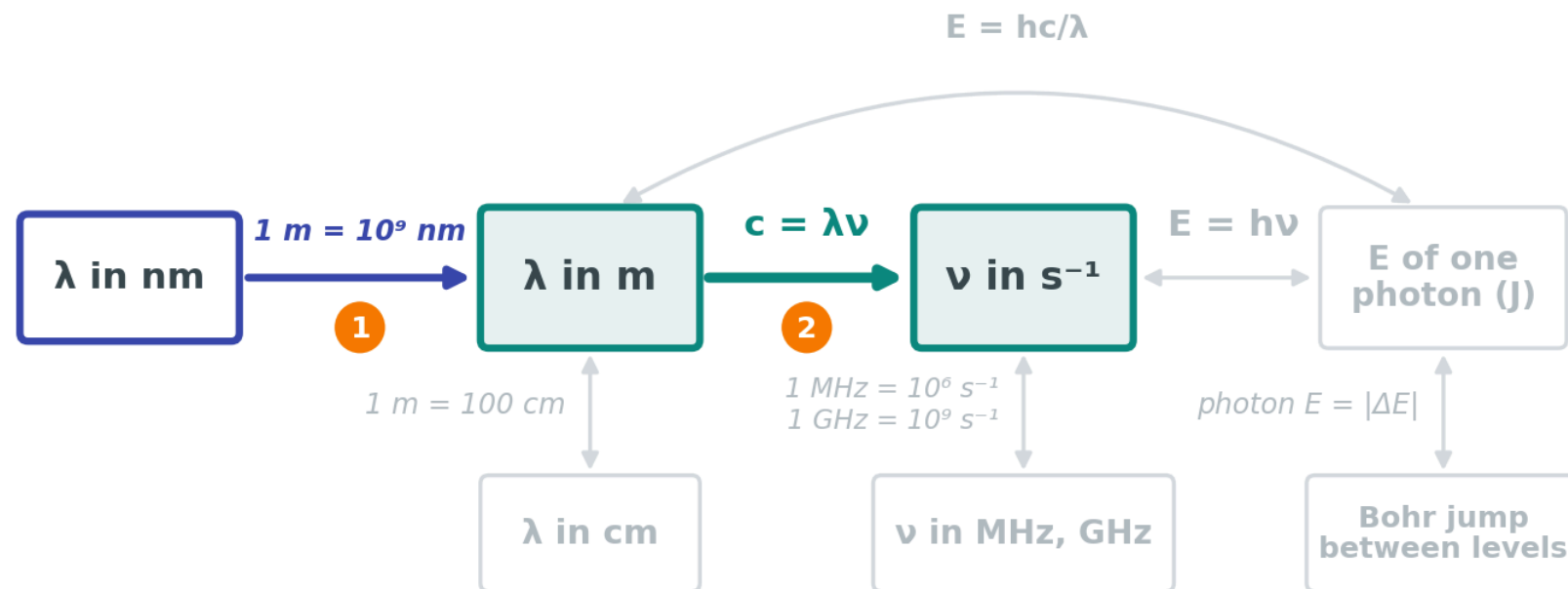
$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{6.00 \times 10^{-7} \text{ m}} = 5.00 \times 10^{14} \text{ s}^{-1}$$

🧠 **Sense check:** 600 nm is orange light, and $5.00 \times 10^{14} \text{ s}^{-1}$ lands in the middle of the visible range. ✓

✓ Worked example 1: the route on the map

$$c = \lambda \nu$$

given: $\lambda = 600 \text{ nm}$ · found: $\nu = 5.00 \times 10^{14} \text{ s}^{-1}$



🧠 **Sense check:** Arrow 1 puts λ in meters to match c . Arrow 2 is $c = \lambda \nu$, solved for ν . ✓



Worked example 2: wavelength of a radio station

$$c = \lambda \nu$$

given: $\nu = 100.0 \text{ MHz}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ

An FM station broadcasts at **100.0 MHz** ($1 \text{ MHz} = 10^6 \text{ s}^{-1}$). Find the wavelength of its radio wave. ($c = 3.00 \times 10^8 \text{ m/s}$)

This time the frequency is given and the wavelength is unknown.

✓ Worked example 2: solution

$$c = \lambda \nu$$

given: $\nu = 100.0 \text{ MHz}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ

Step 1 · Identify the unknown

Step 2 · Match units to c

Step 3 · Rearrange for the unknown

Step 4 · Substitute and cancel units

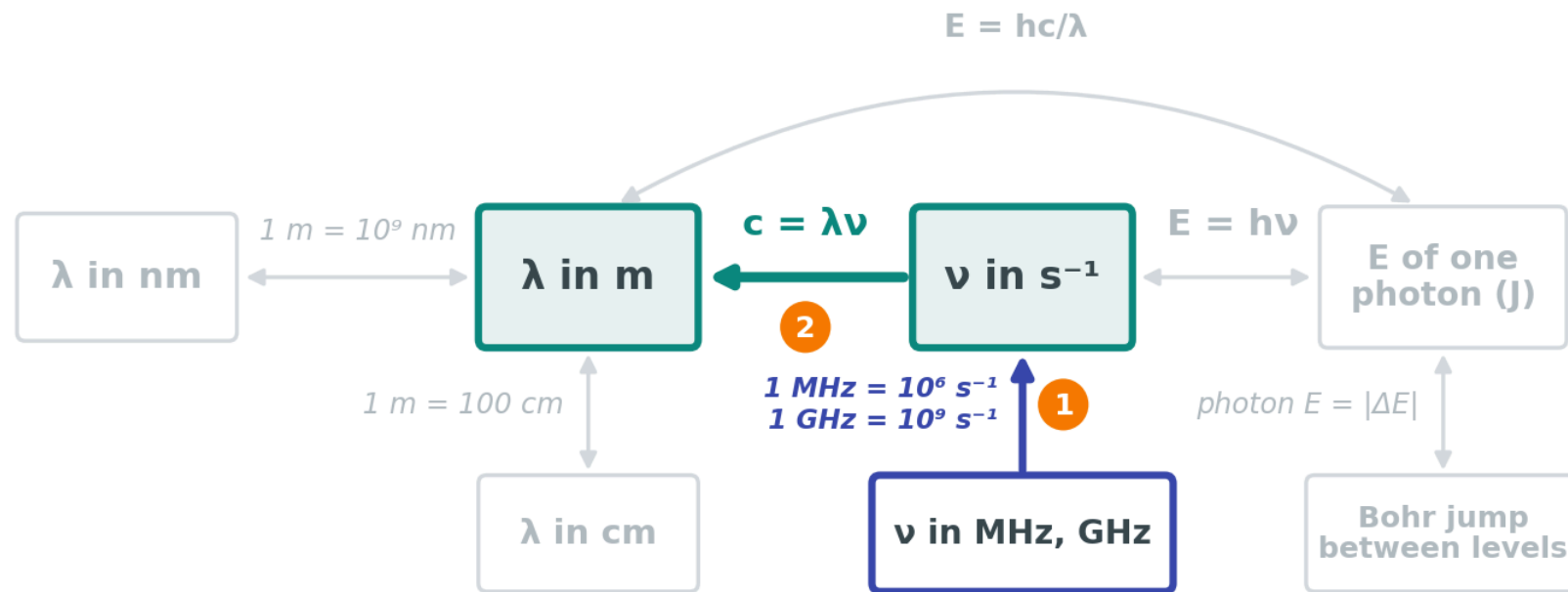
$$\lambda = \frac{3.00 \times 10^8 \text{ m} \cdot \cancel{\text{s}^{-1}}}{1.00 \times 10^8 \cancel{\text{s}^{-1}}} = 3.00 \text{ m}$$

 **Sense check:** A 3-meter wave, far longer than any visible light. Long waves and low frequencies travel together, which is why these are called radio waves. ✓

✓ Worked example 2: the route on the map

$$c = \lambda \nu$$

given: $\nu = 100.0 \text{ MHz}$ · found: $\lambda = 3.00 \text{ m}$



🧠 **Sense check:** The same equation, run the other way. Arrow 1 puts ν in s^{-1} ; arrow 2 is $c = \lambda \nu$, solved for λ . ✓

Your turn: frequency of green light

$$c = \lambda \nu$$

given: $\lambda = 500 \text{ nm (green)}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

A leaf reflects **green light** near **500 nm**. Convert it to meters, then divide c by it.

$$\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^8 \text{ m/s}}{\text{ } \text{m}} = \text{ } \text{s}^{-1}$$

$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{5.00 \times 10^{-7} \text{ m}} = 6.00 \times 10^{14} \text{ s}^{-1}$$

 **Sense check:** 500 nm is green, between red and violet, and $6.00 \times 10^{14} \text{ s}^{-1}$ falls right in the visible range. ✓

Where this goes wrong

$$\nu = c/\lambda$$

600 nm orange light $\cdot c = 3.00 \times 10^8 \text{ m/s}$ $\cdot$ correct $\nu = 5.00 \times 10^{14} \text{ s}^{-1}$

Leaving the wavelength in nanometers. $3.00 \times 10^8 \div 600 = 5.00 \times 10^5 \text{ s}^{-1}$, a factor of 10^9 too small. c is meters per second, so λ must be in meters: $600 \text{ nm} = 6.00 \times 10^{-7} \text{ m}$.

Multiplying c by the wavelength. $3.00 \times 10^8 \times 6.00 \times 10^{-7} = 180$. Its units are m^2/s , not a frequency. Rearrange $c = \lambda\nu$ to $\nu = c/\lambda$, with λ underneath.

Putting the speed of light on the bottom. $\lambda \div c = 6.00 \times 10^{-7} \div 3.00 \times 10^8 = 2.00 \times 10^{-15} \text{ s}$, a time in seconds, not a frequency. Frequency needs c on top.

Practice 1: comparing colors

$$c = \lambda \nu$$

visible colors: red · orange · yellow · green · blue · violet

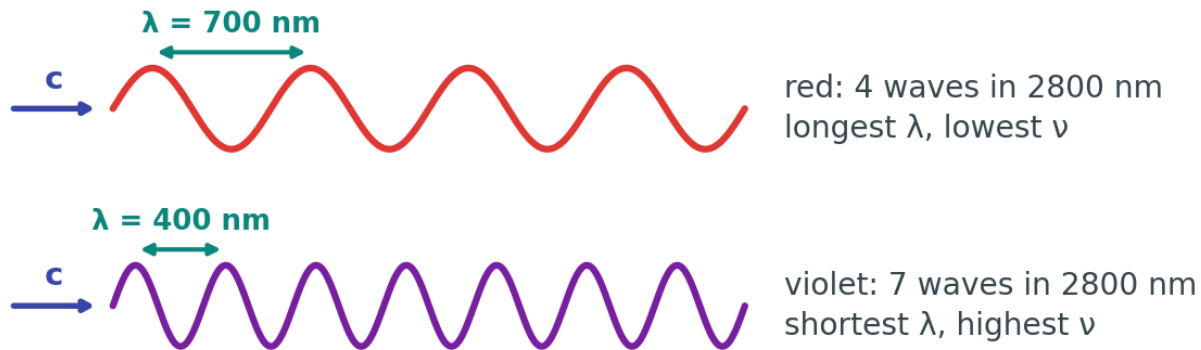
Which color has the **longest wavelength**, and which has the **highest frequency**?

- A. red, then violet
- B. red, then red
- C. violet, then red
- D. red, then the same for every color

✓ Practice 1 · answer: A

$$c = \lambda \nu$$

c is fixed · a longer λ means a lower ν · red 700 nm, violet 400 nm



B let λ and ν rise together; with c fixed, the longer wave has the lower frequency. C reversed the colors; ROY G BIV runs from red at 700 nm to violet at 400 nm. D read one speed as one frequency; each color has its own λ , so its own ν .

🧠 **Sense check:** Red: longest wave, lowest frequency. Violet: shortest wave, highest frequency. ✓



Practice 2

$$c = \lambda \nu$$

given: $\lambda = 400. \text{ nm (violet)}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

Violet light sits at **400. nm**. What is its frequency, in s^{-1} ?

- A. 7.50×10^5
- B. 7.50×10^{14}
- C. 1.20×10^2
- D. 1.33×10^{-15}

✓ Practice 2 · answer: B

$$v = c/\lambda$$

$$\text{given: } \lambda = 400. \text{ nm} = 4.00 \times 10^{-7} \text{ m} \cdot c = 3.00 \times 10^8 \text{ m/s}$$

$$v = \frac{3.00 \times 10^8 \text{ m/s}}{4.00 \times 10^{-7} \text{ m}} = 7.50 \times 10^{14} \text{ s}^{-1} \text{ (answer B)}$$

A left the wavelength in nanometers: $3.00 \times 10^8 \div 400. = 7.50 \times 10^5 \text{ s}^{-1}$, smaller by 10^9 . C multiplied c by λ : $3.00 \times 10^8 \times 4.00 \times 10^{-7} = 1.20 \times 10^2$, with units m^2/s . D put c on the bottom: $4.00 \times 10^{-7} \div 3.00 \times 10^8 = 1.33 \times 10^{-15} \text{ s}$, a time.

🧠 **Sense check:** 400. nm is violet, the short-wavelength edge, so its frequency is the highest in the visible range. ✓



Practice 3

$$c = \lambda \nu$$

given: $\nu = 315 \text{ MHz}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in m

A car's key fob unlocks the doors with a radio signal at **315 MHz**. What is the signal's wavelength, in meters?

- A. 9.52×10^5
- B. 1.05
- C. 0.952
- D. 952

✓ Practice 3 · answer: C

$$\lambda = c/v$$

given: $v = 315 \text{ MHz} = 3.15 \times 10^8 \text{ s}^{-1}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in m

$$\lambda = \frac{3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1}}{3.15 \times 10^8 \text{ s}^{-1}} = 0.952 \text{ m (answer C)}$$

A left the frequency in MHz: $3.00 \times 10^8 \div 315 = 9.52 \times 10^5 \text{ m}$, a wave nearly a thousand kilometers long. B put c on the bottom: $3.15 \times 10^8 \div 3.00 \times 10^8 = 1.05$, with units m^{-1} . D used 10^3 for mega: $3.00 \times 10^8 \div 3.15 \times 10^5 = 952 \text{ m}$. Mega means 10^6 .

🧠 **Sense check:** Just under a meter: a radio wave, shorter than an FM wave because 315 MHz is a higher frequency.

✓



Practice 4

$$c = \lambda \nu$$

given: $\nu = 5.00 \text{ GHz}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in cm

A Wi-Fi router transmits at **5.00 GHz**. How long is its wave, in centimeters?

- A. 6.00×10^9
- B. 16.7
- C. 0.0600
- D. 6.00

✓ Practice 4 · answer: D

$$\lambda = c/v$$

given: $v = 5.00 \text{ GHz} = 5.00 \times 10^9 \text{ s}^{-1}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in cm

$$\lambda = \frac{3.00 \times 10^8 \text{ m} \cdot \cancel{\text{s}^{-1}}}{5.00 \times 10^9 \cancel{\text{s}^{-1}}} = 0.0600 \cancel{\text{ m}} \times \frac{100 \text{ cm}}{1 \cancel{\text{ m}}} = \mathbf{6.00 \text{ cm (answer D)}}$$

A left the frequency in GHz: $3.00 \times 10^8 \div 5.00 = 6.00 \times 10^7 \text{ m}$, which is $6.00 \times 10^9 \text{ cm}$, a wave longer than the Earth. B put c on the bottom: $5.00 \times 10^9 \div 3.00 \times 10^8 = 16.7$, with units m^{-1} . C stopped halfway: $3.00 \times 10^8 \div 5.00 \times 10^9 = 0.0600$ is the wavelength in meters; the last hop, $\times (100 \text{ cm} / 1 \text{ m})$, was never done.

 **Sense check:** A few centimeters: shorter than a radio wave, far longer than visible light. That is the microwave band, where Wi-Fi lives. ✓



Check yourself

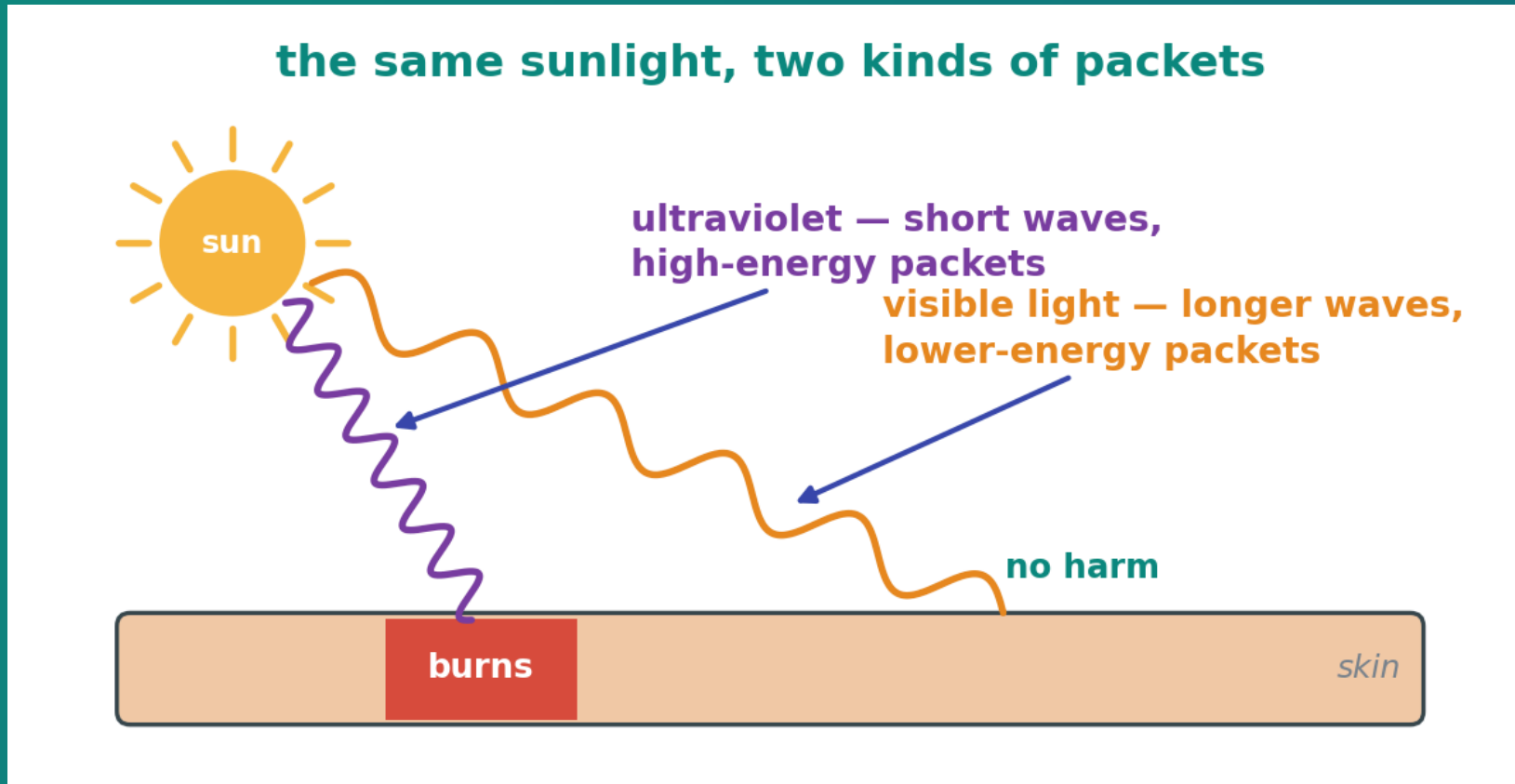
1. Blue light has a wavelength of 450 nm. Solve $c = \lambda\nu$ for its frequency in symbols. Which unit cancels, and which one survives?
2. An X-ray has a wavelength of 0.100 nm. Set up $c = \lambda\nu$ for its frequency. Is that frequency higher or lower than visible light's?

Frequency also fixes a photon's energy through $E = h \cdot \nu$, with h a constant. A higher frequency means more energy per photon, which is why ultraviolet light burns skin and radio waves pass through harmlessly.

2 • Photon Energy

Find the energy of one photon from its frequency with $E = h\nu$, or from its wavelength with $E = hc/\lambda$, remembering that a shorter wavelength and a higher frequency both mean more energy per photon.

Sunlight and sunscreen

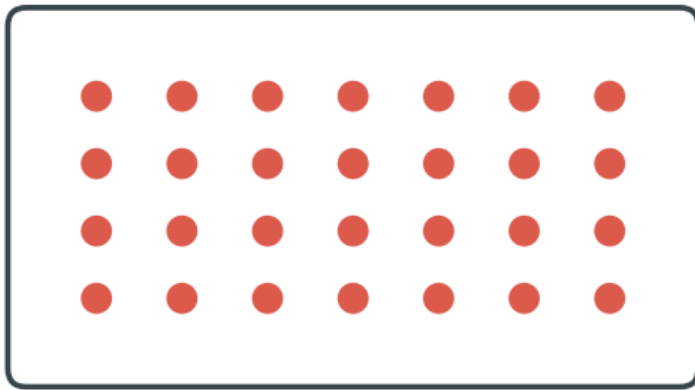


Ultraviolet light burns skin. The visible light beside it does not. The difference is not brightness. Each packet of ultraviolet carries far more energy than a packet of visible light.

Light comes in packets

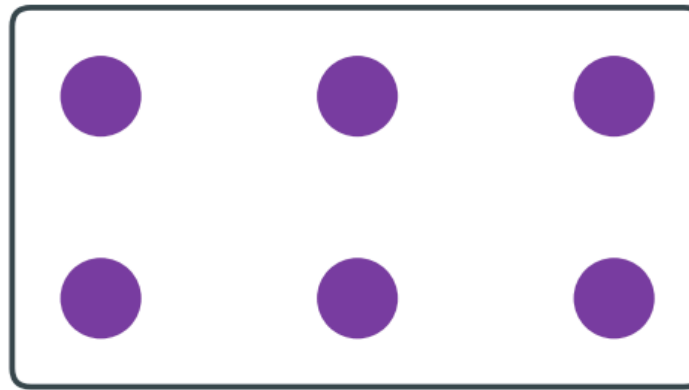
brightness is how MANY packets; frequency sets the energy of EACH

bright red light



many packets
each carries little energy

dim violet light



few packets
each carries much more energy

dot size = energy per packet

Light arrives in discrete packets called photons. Each photon carries a fixed energy. A brighter beam sends more photons, not more energetic ones. Frequency sets each photon's energy.

The photon energy formula

$s^{-1} \times s$ cancels — E comes out in joules

$$\boxed{E} = \boxed{h} \times \boxed{\nu} \rightarrow \text{put } \nu = c/\lambda$$

energy of one photon
J

Planck's constant
 6.626×10^{-34}
J·s

frequency of the light
 s^{-1}

with $\nu = c/\lambda$ from the wave relationship, $E = h \cdot c / \lambda$

One photon's energy equals Planck's constant times the light's frequency. Planck's constant, h , is fixed at 6.626×10^{-34} J·s. A higher frequency means a higher-energy photon.

Wave-particle duality

wave: $c = \lambda \nu$ · particle: $E = h\nu$

one ν sits in both equations: the same light, described both ways

Light is both at once: a wave with a wavelength and a frequency, and a stream of photons each carrying $E = h\nu$. This double identity is called **wave-particle duality**.

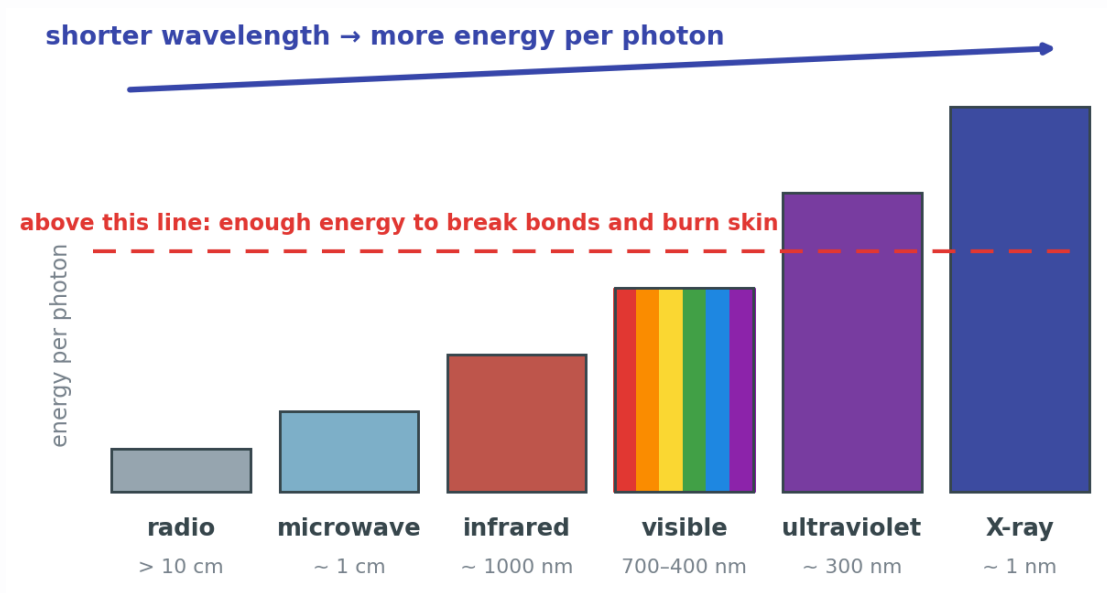
From wavelength: $E = hc/\lambda$

$$\nu = c/\lambda \rightarrow E = h \cdot \nu = h \cdot c / \lambda$$

a wavelength gives a frequency, and a frequency gives an energy · convert nm to meters first

A light wave's frequency and wavelength are tied by $\nu = c/\lambda$. Substitute that into $E = h\nu$ to get a photon's energy straight from its wavelength.

Shorter wavelength, more energy



Across the spectrum, wavelength shrinks and frequency climbs together, so the energy per photon climbs too. Radio photons are feeble. Ultraviolet photons carry enough energy to break bonds in skin.

memory hook: ROY G BIV climbs in energy: red lowest, violet highest
infra means below: infrared sits under red · ultra means beyond: ultraviolet lies past violet



The method

1. **Identify the given** and the unknown. Convert a wavelength to meters first.
2. **Choose the formula:** $E = h\nu$ from a frequency, $E = hc/\lambda$ from a wavelength.
3. **Substitute and cancel units.**
4. **Check the size:** shorter wavelength, more energy.



Worked example 1: energy from a frequency

$$E = h \cdot \nu$$

given: $\nu = 5.00 \times 10^{14} \text{ s}^{-1}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · wanted: E

A photon of orange light has a frequency of $5.00 \times 10^{14} \text{ s}^{-1}$. What is its energy? ($h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$)

Identify the given and the unknown, then choose the formula.

✓ Worked example 1: solution

$$E = h \cdot \nu$$

given: $\nu = 5.00 \times 10^{14} \text{ s}^{-1}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · wanted: E

Step 1 · Identify the given

Step 2 · Choose the formula

Step 3 · Substitute and cancel units

$$E = 6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 5.00 \times 10^{14} \text{ s}^{-1} = 3.31 \times 10^{-19} \text{ J}$$

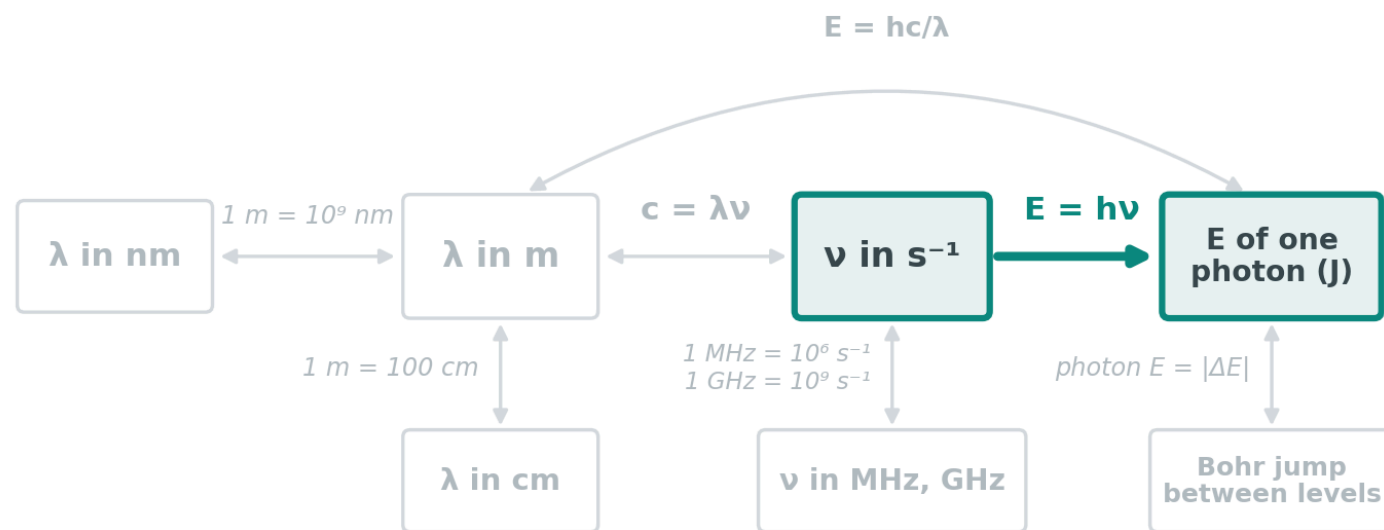
Step 4 · Check the size

 **Sense check:** One photon of visible light carries about 10^{-19} J . A single packet of light holds only a tiny amount of energy. ✓

✓ Worked example 1: the route on the map

$$E = h \cdot \nu$$

given: $\nu = 5.00 \times 10^{14} \text{ s}^{-1}$ · found: $E = 3.31 \times 10^{-19} \text{ J}$



🧠 **Sense check:** A frequency sits one arrow from energy. One formula, $E = h\nu$, and no unit hop. ✓

Practice: three lamps

$$E = h \cdot c / \lambda$$

one photon from each lamp · no numbers needed

Rank one photon from each lamp from lowest to highest energy: an ultraviolet sterilizing lamp, a green laser pointer, an infrared heat lamp.

- A. ultraviolet < green < infrared
- B. infrared < green < ultraviolet
- C. green < ultraviolet < infrared
- D. all three carry the same energy per photon

✓ Three lamps: answer B

$$E = h \cdot c / \lambda$$

infrared: longest λ · green: middle · ultraviolet: shortest λ

longest λ , lowest $E \rightarrow$ **infrared < green < ultraviolet (answer B)**

A ranked energy rising with wavelength, the $E = h\lambda$ slip; λ sits in the denominator of $E = hc/\lambda$. C took the heat lamp's warmth for photon energy: a heat lamp warms by sending many weak photons. D mixed up brightness and energy: wavelength fixes each photon's energy, and brightness counts photons.

 **Sense check:** Infrared sits below red, ultraviolet past violet: energy per photon climbs from infrared to ultraviolet.

✓



Practice: energy from a frequency

$$E = h \cdot \nu$$

given: $\nu = 6.52 \times 10^{14} \text{ s}^{-1}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · wanted: E

A blue LED emits light with a frequency of $6.52 \times 10^{14} \text{ s}^{-1}$. What is the energy, in joules, of one photon?

- A. 4.32×10^{-19}
- B. 1.02×10^{-48}
- C. 1.44×10^{-27}
- D. 3.05×10^{-40}

✓ Energy from a frequency: answer A

$$E = h \cdot \nu$$

given: $\nu = 6.52 \times 10^{14} \text{ s}^{-1}$ · wanted: E

$$E = 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \times 6.52 \times 10^{14} \text{ s}^{-1} = 4.32 \times 10^{-19} \text{ J (answer A)}$$

B divided instead of multiplying: $6.626 \times 10^{-34} / 6.52 \times 10^{14} = 1.02 \times 10^{-48}$, in $\text{J} \cdot \text{s}^2$, not joules.
C divided by c as well: $h\nu/c = 1.44 \times 10^{-27}$. D put the frequency where λ goes: $hc/\nu = 3.05 \times 10^{-40}$.

🧠 **Sense check:** Blue light has a shorter wavelength than orange, and $4.32 \times 10^{-19} \text{ J}$ lies above orange light's $3.31 \times 10^{-19} \text{ J}$. ✓



Worked example 2: energy from a wavelength

$$E = h \cdot c / \lambda$$

given: $\lambda = 400 \text{ nm}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: E

A photon of violet light has a wavelength of **400 nm**. What is its energy? ($1 \text{ m} = 10^9 \text{ nm}$)

Convert the wavelength to meters, then substitute.

✓ Worked example 2: solution

$$E = h \cdot c / \lambda$$

given: $\lambda = 400 \text{ nm}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: E

Step 1 · Identify the given

Step 2 · Choose the formula

Step 3 · Substitute and cancel units

$$E = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 3.00 \times 10^8 \text{ m/s}}{400 \times 10^{-9} \text{ m}} = 4.97 \times 10^{-19} \text{ J}$$

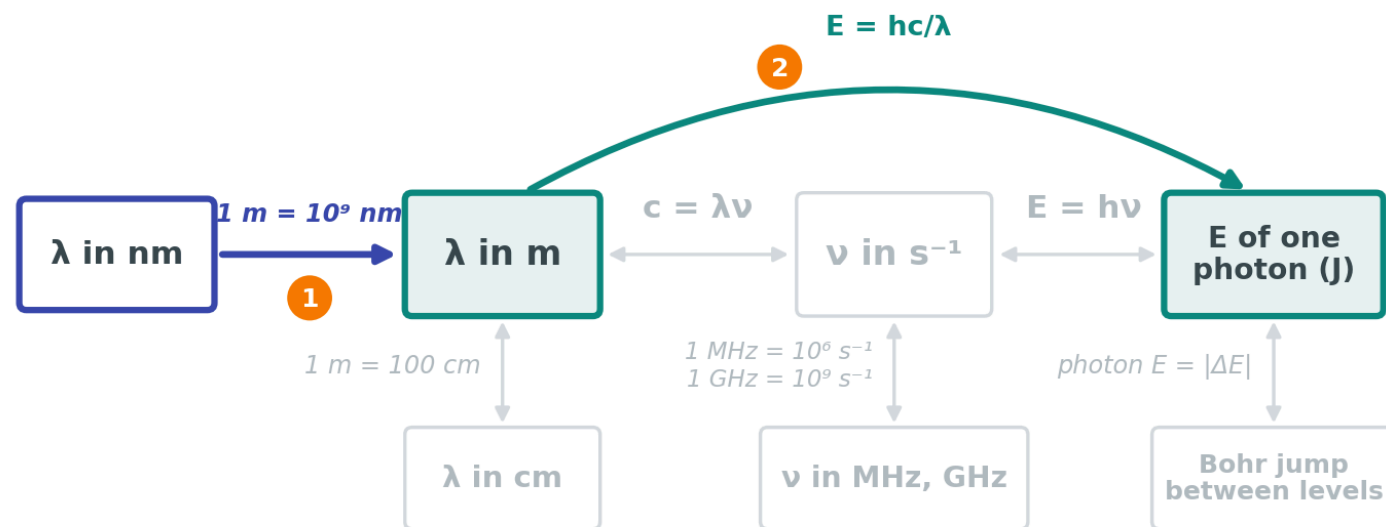
Step 4 · Check the size

 **Sense check:** Violet light (400 nm) lands near $5 \times 10^{-19} \text{ J}$ per photon. A shorter wavelength carries more energy, and 400 nm is near the short end of visible light. ✓

✓ Worked example 2: the route on the map

$$E = h \cdot c / \lambda$$

given: $\lambda = 400 \text{ nm}$ · found: $E = 4.97 \times 10^{-19} \text{ J}$



🧠 **Sense check:** Two moves: nanometers to meters, then $E = hc/\lambda$. The combined formula passes over the frequency box. ✓

Your turn: energy of a green photon

$$E = h \cdot c / \lambda$$

given: $\lambda = 500 \text{ nm}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: E

A photon of **green light** has a wavelength of **500 nm**. ($1 \text{ m} = 10^9 \text{ nm}$)

$$E = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 3.00 \times 10^8 \text{ m/s}}{\text{ } \text{m}} = \text{ } \text{J}$$

Convert 500 nm to meters, fill the denominator, then compute.

$$E = \frac{6.626 \times 10^{-34} \text{ J}\cdot\text{s} \times 3.00 \times 10^8 \text{ m/s}}{500 \times 10^{-9} \text{ m}} = 3.98 \times 10^{-19} \text{ J}$$

 **Sense check:** Green light (500 nm) sits mid-spectrum, and its photon energy lands mid-range too, near $4 \times 10^{-19} \text{ J}$.



Where this goes wrong

$$E = h \cdot c / \lambda$$

$$\lambda = 400 \text{ nm} \cdot h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s} \cdot c = 3.00 \times 10^8 \text{ m/s} \cdot \text{correct } E = 4.97 \times 10^{-19} \text{ J}$$

Stopping at the frequency. $\nu = c/\lambda = 7.50 \times 10^{14} \text{ s}^{-1}$ is only the frequency, one step short. Finish with $E = h\nu$: multiply by $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ so s^{-1} and s cancel into joules.

Leaving the wavelength in nanometers. Dividing by 400 instead of $400 \times 10^{-9} \text{ m}$ gives $4.97 \times 10^{-28} \text{ J}$, smaller by a factor of 10^9 . Convert first: $400 \text{ nm} \times (1 \text{ m} / 10^9 \text{ nm})$.

Pairing energy with wavelength. $E = h\lambda = 2.65 \times 10^{-40}$ has units of $\text{J}\cdot\text{s}\cdot\text{m}$, not joules. Energy pairs with frequency: $E = h\nu$, or $E = hc/\lambda$.



Practice: a neon sign

$$E = h \cdot c / \lambda$$

given: $\lambda = 640. \text{ nm}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: E

A neon sign glows with light of wavelength **640. nm**. What is the energy, in joules, of one photon?

- A. 3.11×10^{-28}
- B. 4.24×10^{-40}
- C. 3.11×10^{-19}
- D. 3.11×10^{-37}

✓ A neon sign: answer C

$$E = h \cdot c / \lambda$$

given: $\lambda = 640. \text{ nm} \times (1 \text{ m} / 10^9 \text{ nm}) = 6.40 \times 10^{-7} \text{ m}$ · wanted: E

$$E = \frac{6.626 \times 10^{-34} \text{ J} \cdot \text{s} \times 3.00 \times 10^8 \text{ m/s}}{6.40 \times 10^{-7} \text{ m}} = 3.11 \times 10^{-19} \text{ J (answer C)}$$

A left λ in nanometers: $hc/640. = 3.11 \times 10^{-28}$. B paired energy with wavelength: $h\lambda = 4.24 \times 10^{-40}$. D flipped the factor: $640. \text{ nm} \times (10^9 \text{ m} / 1 \text{ nm}) = 6.40 \times 10^{11} \text{ m}$, so $hc/\lambda = 3.11 \times 10^{-37}$.

 **Sense check:** Neon's red-orange glow sits just below orange light's $3.31 \times 10^{-19} \text{ J}$, as a longer wavelength should.

✓



Practice: wavelength from an energy

$$E = h \cdot c / \lambda$$

given: $E = 3.60 \times 10^{-19} \text{ J}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in nm

A photon carries $3.60 \times 10^{-19} \text{ J}$. What is its wavelength, in nanometers?

- A. 5.52×10^{-7}
- B. 5.52×10^{-16}
- C. 1.84×10^{-6}
- D. 552

✓ Wavelength from an energy: answer D

$$\lambda = h \cdot c / E$$

given: $E = 3.60 \times 10^{-19} \text{ J}$ · wanted: λ in nm

$$\lambda = \frac{6.626 \times 10^{-34} \text{ J} \cdot \text{s} \times 3.00 \times 10^8 \text{ m/s}}{3.60 \times 10^{-19} \text{ J}} = 5.52 \times 10^{-7} \text{ m} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = \mathbf{552 \text{ nm (answer D)}}$$

A skipped the last unit hop: 5.52×10^{-7} is the wavelength in meters, not nanometers. B flipped the conversion factor: $5.52 \times 10^{-7} \text{ m} \times (1 \text{ m} / 10^9 \text{ nm}) = 5.52 \times 10^{-16}$, and meters do not cancel. C dropped c : $\nu = E/h = 5.43 \times 10^{14} \text{ s}^{-1}$, then $\lambda = 1/\nu = 1.84 \times 10^{-15}$, which is 1.84×10^{-6} after the hop; $\lambda = c/\nu$ needs c on top.

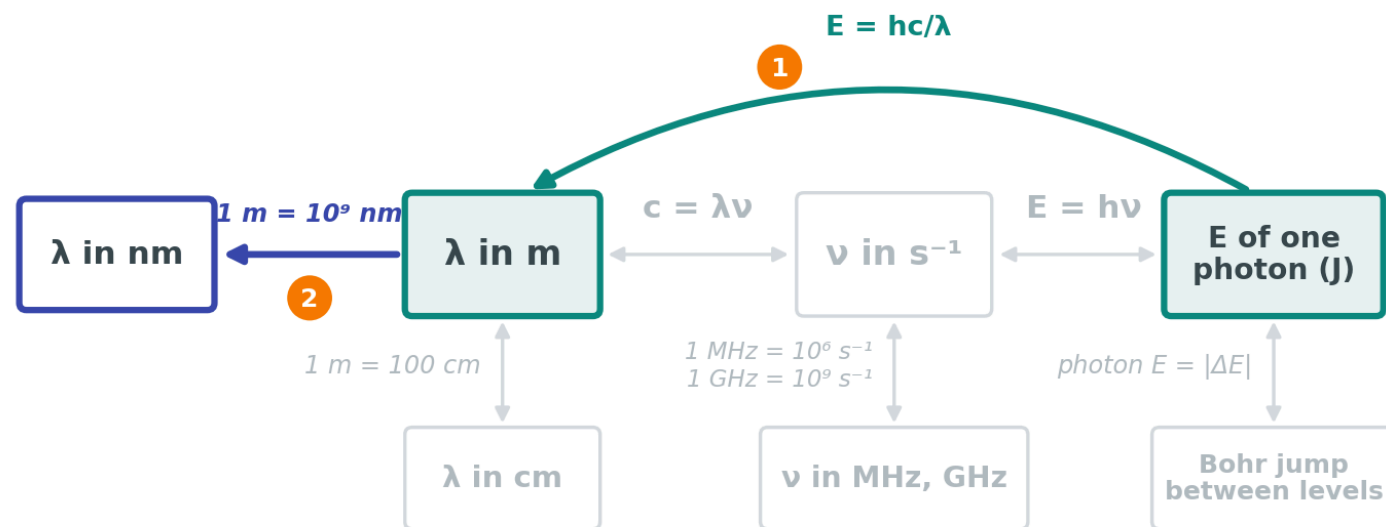
 **Sense check:** 552 nm sits in the green part of the visible band, and $3.60 \times 10^{-19} \text{ J}$ is a typical visible-photon energy.



✓ Wavelength from an energy: the route on the map

$$\lambda = h \cdot c / E$$

given: $E = 3.60 \times 10^{-19} \text{ J}$ · found: $\lambda = 552 \text{ nm}$



 **Sense check:** The wavelength-to-energy route, run in reverse: $\lambda = hc/E$ gives meters, then 10^9 nm per meter gives nanometers. ✓



Practice: the energy gap

$$E = h \cdot c / \lambda$$

photon A: $\lambda = 630. \text{ nm}$ · photon B: $\lambda = 210. \text{ nm}$ · $h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$

Photon A has a wavelength of **630. nm**; photon B has a wavelength of **210. nm**. How much more energy, in J, does photon B carry than photon A?

- A. 9.47×10^{-19}
- B. 6.31×10^{-19}
- C. 4.73×10^{-19}
- D. 6.31×10^{-28}
- E. -6.31×10^{-19}

✓ The energy gap: answer B

$$\Delta E = E_B - E_A = hc/\lambda_B - hc/\lambda_A$$

photon A: $\lambda = 6.30 \times 10^{-7} \text{ m}$ · photon B: $\lambda = 2.10 \times 10^{-7} \text{ m}$ · wanted: how much more B carries

$$\begin{aligned} \Delta E &= \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{2.10 \times 10^{-7}} - \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{6.30 \times 10^{-7}} = 9.47 \times 10^{-19} \text{ J} - 3.16 \times 10^{-19} \text{ J} \\ &= \mathbf{6.31 \times 10^{-19} \text{ J (answer B)}} \end{aligned}$$

A stopped at $E_B = 9.47 \times 10^{-19} \text{ J}$ and never subtracted E_A . C subtracted the wavelengths first: $630. - 210. = 420. \text{ nm}$, then $hc/(4.20 \times 10^{-7} \text{ m}) = 4.73 \times 10^{-19} \text{ J}$; subtract energies, not wavelengths. D left both λ in nm: $hc/210. - hc/630. = 6.31 \times 10^{-28} \text{ J}$. E subtracted backwards: $E_A - E_B = -6.31 \times 10^{-19} \text{ J}$.

🧠 **Sense check:** 210. nm is one-third of 630. nm, so $E_B = 3E_A$ and the gap is $2 \times 3.155 \times 10^{-19} \text{ J} = 6.31 \times 10^{-19} \text{ J}$. ✓



Check yourself

1. A photon's frequency doubles. What happens to its energy, and which formula tells you in one line?
2. Two photons: one at 350 nm, one at 700 nm. Without a calculator, which carries more energy, and roughly how many times more?

When an electron drops from a higher energy level to a lower one inside an atom, it emits a photon whose energy is exactly the gap between the levels. $E = hc/\lambda$ then turns that gap into the color of light you see.

3 · The Bohr Model & Line Spectra

Explain why an excited hydrogen atom emits a line spectrum rather than a continuous one, and calculate the energy of the photon released when an electron falls between principal energy levels.

Two kinds of glow

Continuous spectrum

every wavelength present — a hot glowing solid



Hydrogen line spectrum

only a few wavelengths — excited H atoms (emission)



A lightbulb pours out every color: a smooth rainbow. A tube of hydrogen gas glows too, but a prism splits it into just a few lines. Why skip the rest?

Line vs. continuous spectra

A **continuous spectrum** holds every wavelength, no gaps: light from a hot solid.

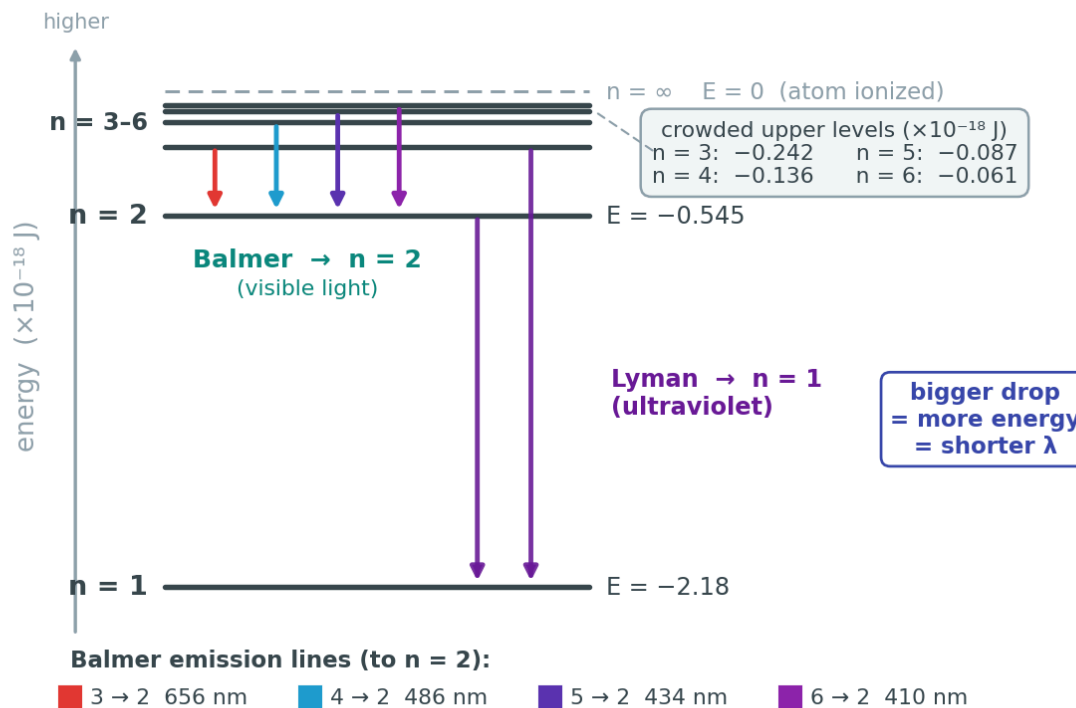
A **line spectrum** holds only a few wavelengths, bright lines in the dark: light from excited, low-pressure gas atoms.

continuous → every wavelength · line → only a few
the gaps are the clue: an atom can release only certain photon energies

Bohr's idea

Bohr's electron sits on a ladder of fixed levels, never between rungs. **Absorbing** energy lifts it to a higher level; falling back **emits** that gap as a **photon**. Only certain gaps exist, so only certain lines appear.

Emission: the electron falls to a lower level and a photon leaves



Reading a jump off the ladder

Every jump is final level minus initial level. Read both off the ladder ($\times 10^{-18}$ J): $n = 1$ at -2.18 , $n = 3$ at -0.242 .

$$3 \rightarrow 1: \Delta E = E_{\text{final}} - E_{\text{initial}} = -2.18 - (-0.242) = -1.94$$

negative: the atom loses energy, so a photon of 1.94×10^{-18} J leaves

Run the same jump upward: $1 \rightarrow 3$.

$$1 \rightarrow 3: \Delta E = -0.242 - (-2.18) = +1.94$$

positive: the atom must absorb 1.94×10^{-18} J · no photon leaves

The energy of a level, and of a jump

Each hydrogen level has a fixed energy:

$$E_n = -2.18 \times 10^{-18} \text{ J} \times (1/n^2)$$

$n = 1$ is deepest ($-2.18 \times 10^{-18} \text{ J}$); higher levels crowd toward 0

A jump changes the energy by the difference between the levels:

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \times (1/n_f^2 - 1/n_i^2)$$

emission: $n_f < n_i \rightarrow \Delta E$ negative; the photon carries $|\Delta E|$ · absorption: $n_f > n_i \rightarrow \Delta E$ positive

Bigger drops release more energy and shorter wavelengths.

Which series lands where

The final level fixes the region of the spectrum.

to $n = 1 \rightarrow$ ultraviolet (Lyman)

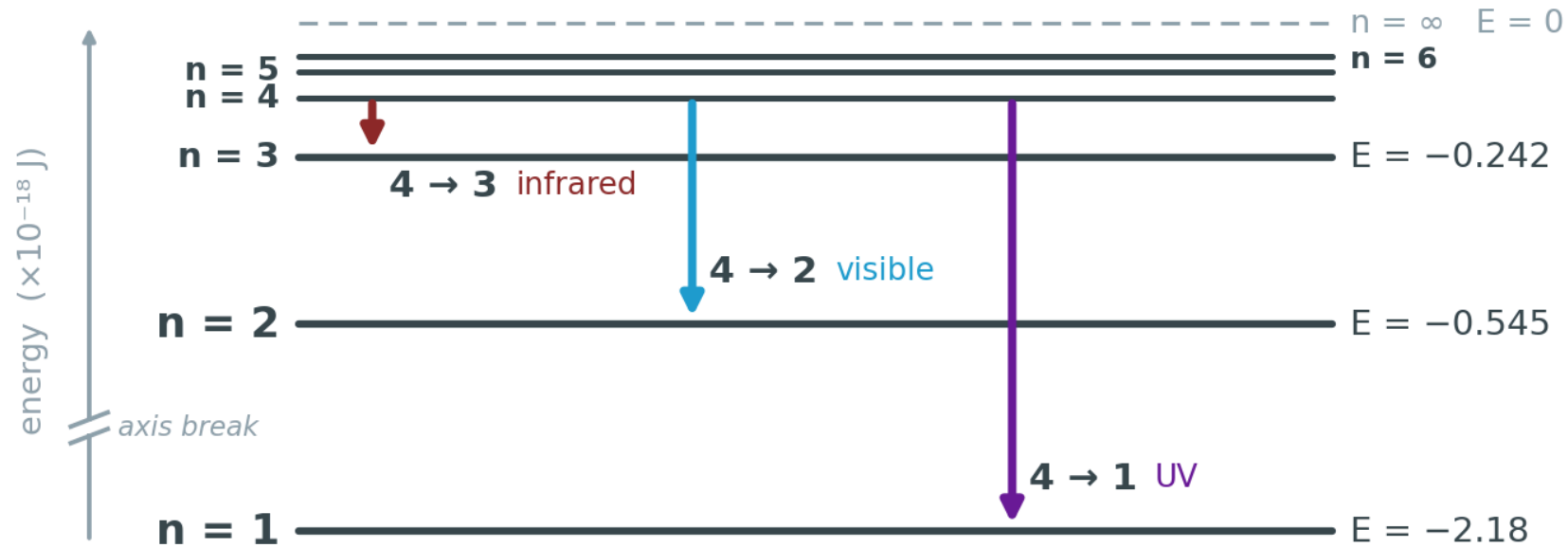
to $n = 2 \rightarrow$ visible (Balmer) · to $n = 3 \rightarrow$ infrared (Paschen)

Jumps ending on $n = 2$ give the visible lines your eye can see; those ending on $n = 1$ are higher-energy ultraviolet.

memory hook: Lyman, Balmer, Paschen land on 1, 2, 3

say the three names in that order and the landing levels count themselves off

Comparing drops on the ladder



Three electrons fall from $n = 4$. Rank their photons from most energy to least.

$4 \rightarrow 1 > 4 \rightarrow 2 > 4 \rightarrow 3$ in photon energy

bigger drop $\rightarrow$ more energy $\rightarrow$ shorter λ · the smallest drop, $4 \rightarrow 3$, gives the longest wavelength



The method

1. **Identify n_i and n_f .** Emission means n_f is the lower level.
2. **Compute $1/n_f^2$ and $1/n_i^2$.** Square, then invert.
3. **Subtract:** $(1/n_f^2 - 1/n_i^2)$.
4. **Multiply by -2.18×10^{-18} J.** ΔE is negative; the photon carries its magnitude.

 Worked example 1: $n = 3 \rightarrow n = 2$

hydrogen: $n_i = 3 \rightarrow n_f = 2$

given: the two levels · wanted: ΔE and the emitted photon's energy

This is the first line of the Balmer series. Find the energy released, then place it in the spectrum.

✓ Worked example 1: solution

$$n_i = 3, \quad n_f = 2$$

emission: the electron drops to the lower level

Step 1 · Identify n_i and n_f

Step 2 · Compute $1/n_f^2$ and $1/n_i^2$

Step 3 · Subtract

$$1/2^2 - 1/3^2 = 0.2500 - 0.1111 = 0.1389$$

Step 4 · Multiply by $-2.18 \times 10^{-18} \text{ J}$

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \times 0.1389 = -3.03 \times 10^{-19} \text{ J}$$

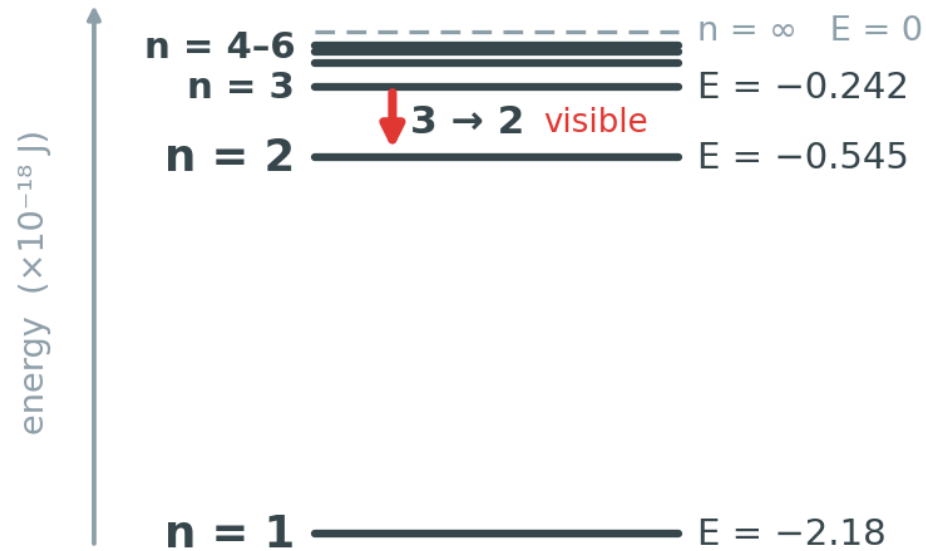
negative $\rightarrow$ the atom loses energy · the photon carries $|\Delta E| = 3.03 \times 10^{-19} \text{ J}$

 **Sense check:** $3.03 \times 10^{-19} \text{ J}$ is a visible photon ($\lambda \approx 656 \text{ nm}$, red): hydrogen's red Balmer line. A small gap gives low energy and a long wavelength.

✓ Worked example 1: the jump on the ladder

$$n_i = 3 \rightarrow n_f = 2$$

found: $\Delta E = -3.03 \times 10^{-19} \text{ J}$ · the photon carries $3.03 \times 10^{-19} \text{ J}$



🧠 **Sense check:** One level down to $n = 2$: a small drop, so a low-energy visible photon (656 nm, red). ✓

Your turn: $n = 4 \rightarrow n = 2$

Fill each blank, then confirm the photon lands in the visible range.

$$1/2^2 - 1/4^2 = 0.2500 - \underline{\hspace{1cm}} = \underline{\hspace{1cm}}$$

quantity	value
$1/n_f^2 = 1/2^2$	0.2500
$1/n_i^2 = 1/4^2$	<u> </u>
difference	<u> </u>
ΔE	<u> </u>

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \times (0.2500 - 0.0625) = -2.18 \times 10^{-18} \times 0.1875 = -4.09 \times 10^{-19} \text{ J}$$

photon = $4.09 \times 10^{-19} \text{ J} \approx 486 \text{ nm}$, blue-green: the second Balmer line

⚠ Where this goes wrong

Forgetting to square n. Using $(1/n_f - 1/n_i)$ instead of $(1/n_f^2 - 1/n_i^2)$ changes the answer. For $3 \rightarrow 2$ that gives 0.1667, not 0.1389: a different, wrong energy. Square each n first.

Subtracting in the wrong order. Writing $(1/n_i^2 - 1/n_f^2)$ flips the sign of ΔE . Keep final minus initial: $(1/n_f^2 - 1/n_i^2)$.

Reporting ΔE as the photon's energy. ΔE for emission is negative because the atom loses energy. The photon it emits carries the positive magnitude, $|\Delta E|$. A photon never has negative energy.

Expecting a continuous rainbow. A lone excited atom can release only fixed gaps, so it emits lines, not a smear. The rainbow comes from hot solids, not from single atoms.

Practice 1: longest wavelength

Four hydrogen electrons each fall to $n = 2$. Which one emits light of the **longest wavelength**?

- A. $n = 5 \rightarrow n = 2$
- B. $n = 6 \rightarrow n = 2$
- C. $n = 4 \rightarrow n = 2$
- D. $n = 3 \rightarrow n = 2$

✓ Practice 1 · answer: D

Longest wavelength means least energy, which means the smallest drop.

D: $3 \rightarrow 2$ is the smallest drop to $n = 2 \rightarrow \lambda \approx 656 \text{ nm}$, red

$4 \rightarrow 2 \approx 486 \text{ nm}$ · $5 \rightarrow 2 \approx 434 \text{ nm}$ · $6 \rightarrow 2 \approx 410 \text{ nm}$: each a bigger drop, each a shorter λ

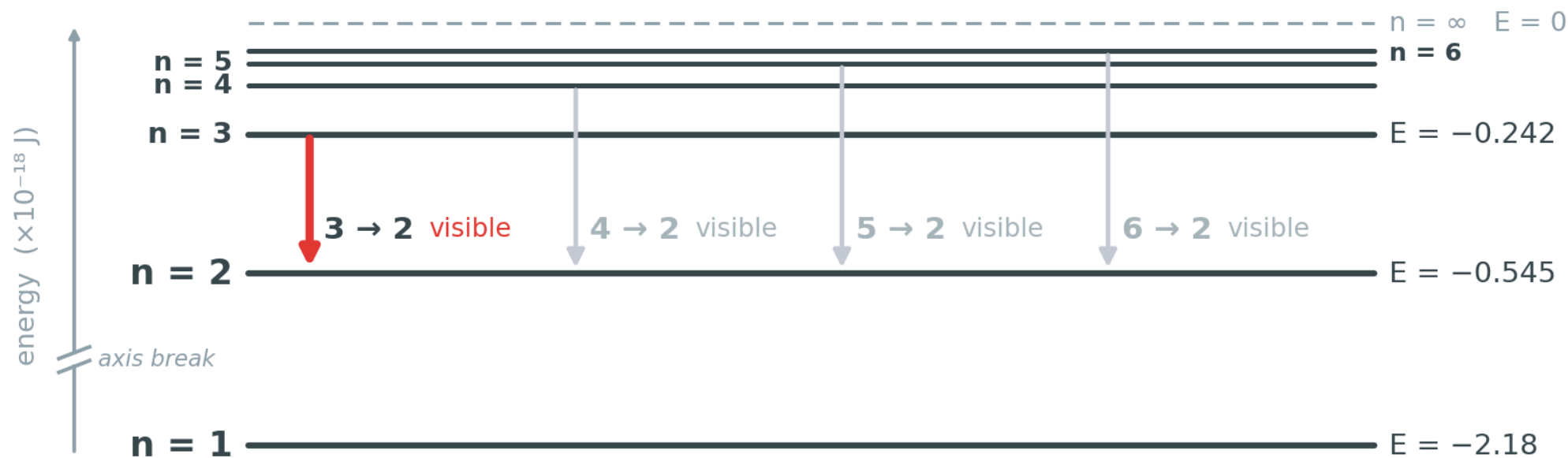
B flipped the relation: $6 \rightarrow 2$ is the biggest drop in the set, so it has the most energy and the shortest wavelength. A and C are also bigger drops than $3 \rightarrow 2$.

 **Sense check:** All four land on $n = 2$, so all four are visible. The smallest gap gives the longest wavelength. ✓

✓ Practice 1: the jumps on the ladder

four drops to $n = 2$

smallest gap: $3 \rightarrow 2$ (answer D) · biggest gap: $6 \rightarrow 2$



🧠 **Sense check:** Same landing level, different starts: the shortest arrow carries the least energy. ✓

Practice 2: which is visible?

Which transition emits a photon of **visible** light?

A. $n = 2 \rightarrow n = 1$

B. $n = 3 \rightarrow n = 2$

C. $n = 4 \rightarrow n = 3$

D. $n = 5 \rightarrow n = 4$

✓ Practice 2 · answer: B

Only jumps that land on $n = 2$ (the Balmer series) fall in the visible range.

$n = 3 \rightarrow n = 2 \Rightarrow$ visible (656 nm, red)
to $n = 2 \rightarrow$ visible · to $n = 1 \rightarrow$ UV · to $n = 3 \rightarrow$ infrared

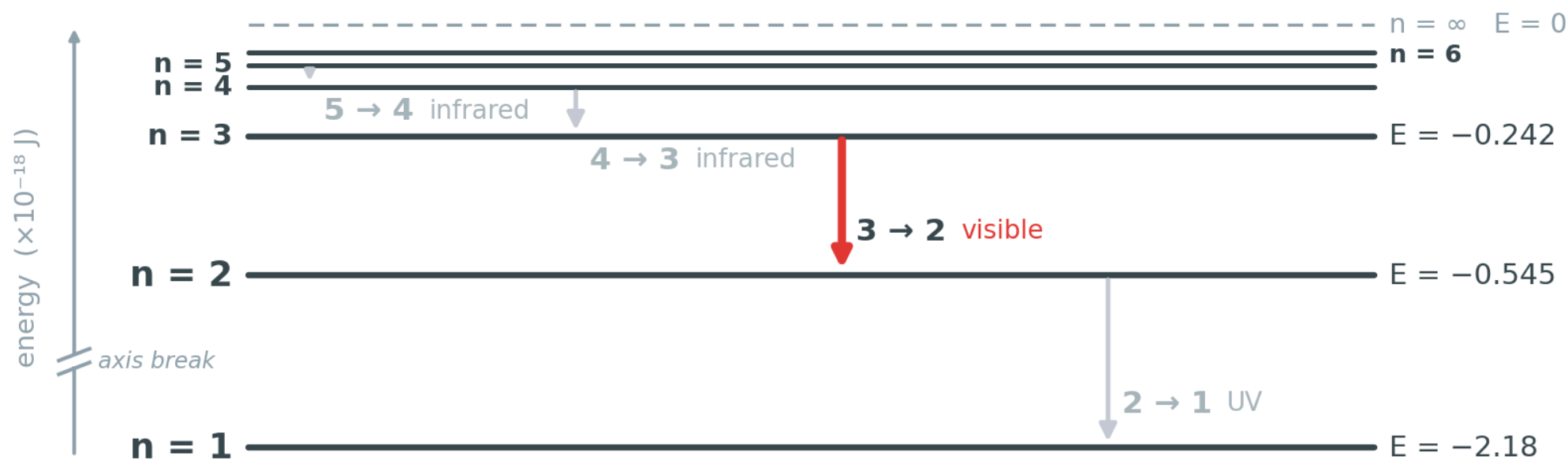
A ($2 \rightarrow 1$) is a Lyman jump: ultraviolet. C ($4 \rightarrow 3$) and D ($5 \rightarrow 4$) land on $n = 3$ or higher: infrared. Only B lands on $n = 2$.

🧠 **Sense check:** Where a jump ends sets its region: $n = 2$ is visible, $n = 1$ is ultraviolet, $n = 3$ and up are infrared.

✓ Practice 2: the jumps on the ladder

four drops, four landing levels

lands on $n = 2$: $3 \rightarrow 2$ (answer B) · $n = 1$: UV · $n = 3$ or $n = 4$: infrared



🧠 **Sense check:** The landing level, not the size of the start, picks the region. ✓

Practice 3: shortest wavelength

Each transition emits a photon. Which emitted photon has the **shortest wavelength**?

A. $n = 2 \rightarrow n = 1$

B. $n = 3 \rightarrow n = 2$

C. $n = 4 \rightarrow n = 2$

D. $n = 5 \rightarrow n = 2$

✓ Practice 3 · answer: A

Shortest wavelength means highest energy, which means the biggest $(1/n_f^2 - 1/n_i^2)$.

$$\text{A: } 1/1^2 - 1/2^2 = 0.7500 \rightarrow \Delta E = -1.64 \times 10^{-18} \text{ J}$$

a UV photon: larger than every Balmer jump below it

B, C, and D all land on $n = 2$ (visible): $3 \rightarrow 2 = 3.03 \times 10^{-19} \text{ J}$, $4 \rightarrow 2 = 4.09 \times 10^{-19} \text{ J}$, $5 \rightarrow 2 = 4.58 \times 10^{-19} \text{ J}$: each smaller than A, so each a longer wavelength.

🧠 **Sense check:** A drops all the way to $n = 1$, the largest gap in the set, so it carries the most energy and the shortest wavelength.

✓ Practice 3: the jumps on the ladder

three drops to $n = 2$, one to $n = 1$

biggest gap: $2 \rightarrow 1$ (answer A), even though its n values are the smallest



🧠 **Sense check:** The gap between $n = 2$ and $n = 1$ is bigger than every gap above it. ✓

Practice 4: ΔE of a Balmer line

hydrogen: $n_i = 6 \rightarrow n_f = 2$
given: the two levels · wanted: ΔE in J

A hydrogen electron falls from $n = 6$ to $n = 2$. What is ΔE for this jump, in joules?

- A. -7.27×10^{-19}
- B. -4.84×10^{-19}
- C. -5.45×10^{-19}
- D. -6.06×10^{-19}
- E. 4.84×10^{-19}

✓ Practice 4 · answer: B

hydrogen: $n_i = 6 \rightarrow n_f = 2$

$$1/n_f^2 = 1/2^2 = 0.2500 \cdot 1/n_i^2 = 1/6^2 = 0.0278$$

$$\begin{aligned}\Delta E &= -2.18 \times 10^{-18} \text{ J} \times (0.2500 - 0.0278) = -2.18 \times 10^{-18} \text{ J} \times 0.2222 \\ &= -4.84 \times 10^{-19} \text{ J (answer B)}\end{aligned}$$

$$\text{check: } E_{\text{final}} - E_{\text{initial}} = -5.45 \times 10^{-19} \text{ J} - (-6.06 \times 10^{-20} \text{ J}) = -4.84 \times 10^{-19} \text{ J} \checkmark$$

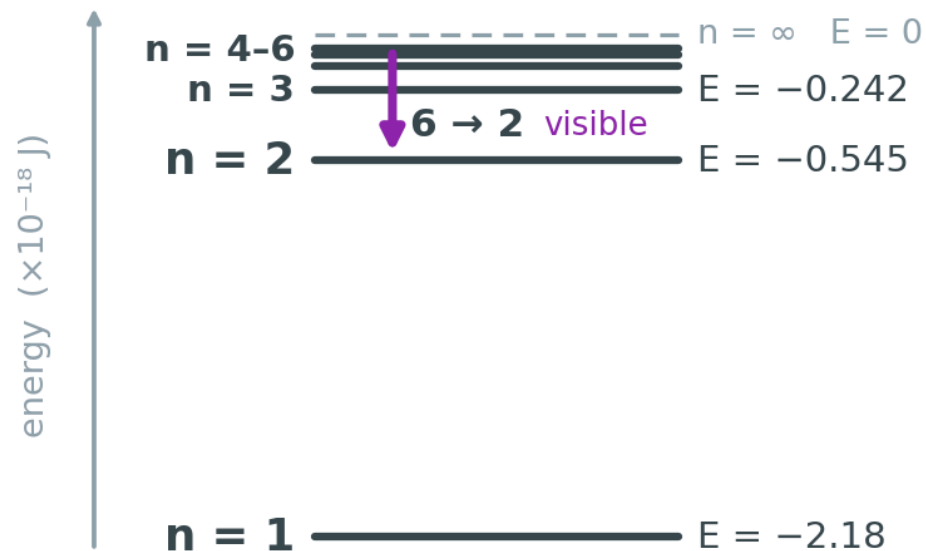
A did not square n : $-2.18 \times 10^{-18} \text{ J} \times (1/2 - 1/6) = -7.27 \times 10^{-19}$. C stopped at E_{final} : $-5.45 \times 10^{-19} \text{ J}$ is the energy of $n = 2$, not the jump. D dropped the double negative: $-5.45 \times 10^{-19} - 6.06 \times 10^{-20} = -6.06 \times 10^{-19}$. E reported the photon's energy, or subtracted initial – final: ΔE for emission is negative.

 **Sense check:** Negative: the atom loses energy. The photon carries $4.84 \times 10^{-19} \text{ J}$, the violet Balmer line (410 nm). ✓

✓ Practice 4: the jump on the ladder

$$n_i = 6 \rightarrow n_f = 2$$

found: $\Delta E = -4.84 \times 10^{-19} \text{ J}$ · the photon carries $4.84 \times 10^{-19} \text{ J}$



🧠 **Sense check:** A drop to $n = 2$ from high on the ladder: still visible, at the violet end. ✓



Practice 5: wavelength of a Lyman line

hydrogen: $n_i = 7 \rightarrow n_f = 1$

given: the two levels · $2.18 \times 10^{-18} \text{ J}$ · $h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$ · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: λ in nm

A hydrogen electron falls from $n = 7$ to $n = 1$. What wavelength of light is emitted, in nanometers?

- A. 2.14×10^{-18}
- B. 91.2
- C. 9.31×10^{-8}
- D. 93.1
- E. 9.31×10^{-17}

✓ Practice 5 · answer: D

hydrogen: $n_i = 7 \rightarrow n_f = 1$
given: the two levels · wanted: λ in nm

$$|\Delta E| = 2.18 \times 10^{-18} \text{ J} \times (1.0000 - 0.0204) = 2.18 \times 10^{-18} \text{ J} \times 0.9796 = 2.136 \times 10^{-18} \text{ J}$$

$$\lambda = \frac{6.626 \times 10^{-34} \text{ J} \cdot \text{s} \times 3.00 \times 10^8 \text{ m/s}}{2.136 \times 10^{-18} \text{ J}} = 9.31 \times 10^{-8} \text{ m} \times \frac{10^9 \text{ nm}}{1 \text{ m}} = \mathbf{93.1 \text{ nm (answer D)}}$$

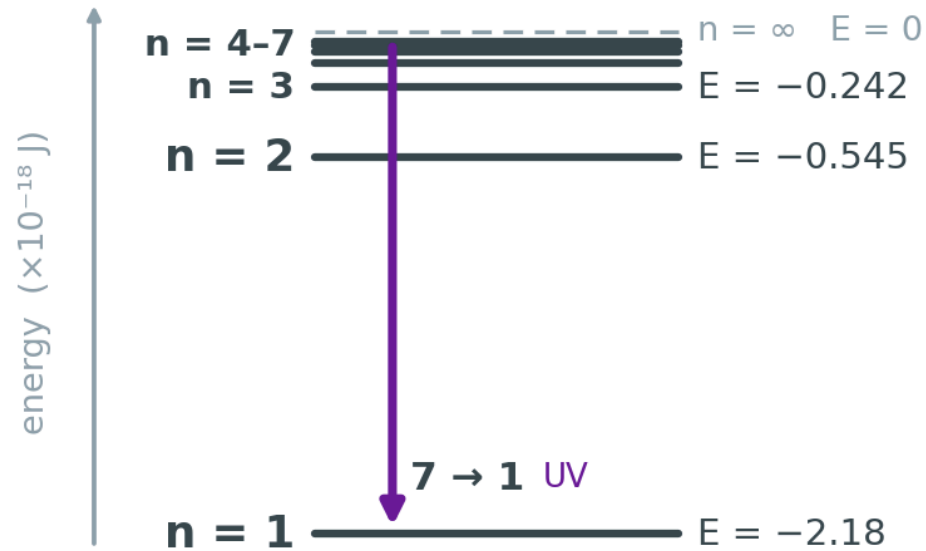
A stopped at the photon's energy, one step short of λ . B used the $n = 1$ level's energy alone: $2.18 \times 10^{-18} \text{ J}$ drops the $1/7^2$ term and gives 91.2 nm. C skipped the m to nm hop: 9.31×10^{-8} is meters. E flipped the factor: $9.31 \times 10^{-8} \text{ m} \times (1 \text{ m} / 10^9 \text{ nm}) = 9.31 \times 10^{-17}$, and meters do not cancel.

🧠 **Sense check:** Lands on $n = 1$: ultraviolet, shorter than any Balmer line. ✓

✓ Practice 5: the jump on the ladder

$$n_i = 7 \rightarrow n_f = 1$$

$$\text{found: } |\Delta E| = 2.14 \times 10^{-18} \text{ J} \cdot \lambda = 93.1 \text{ nm}$$

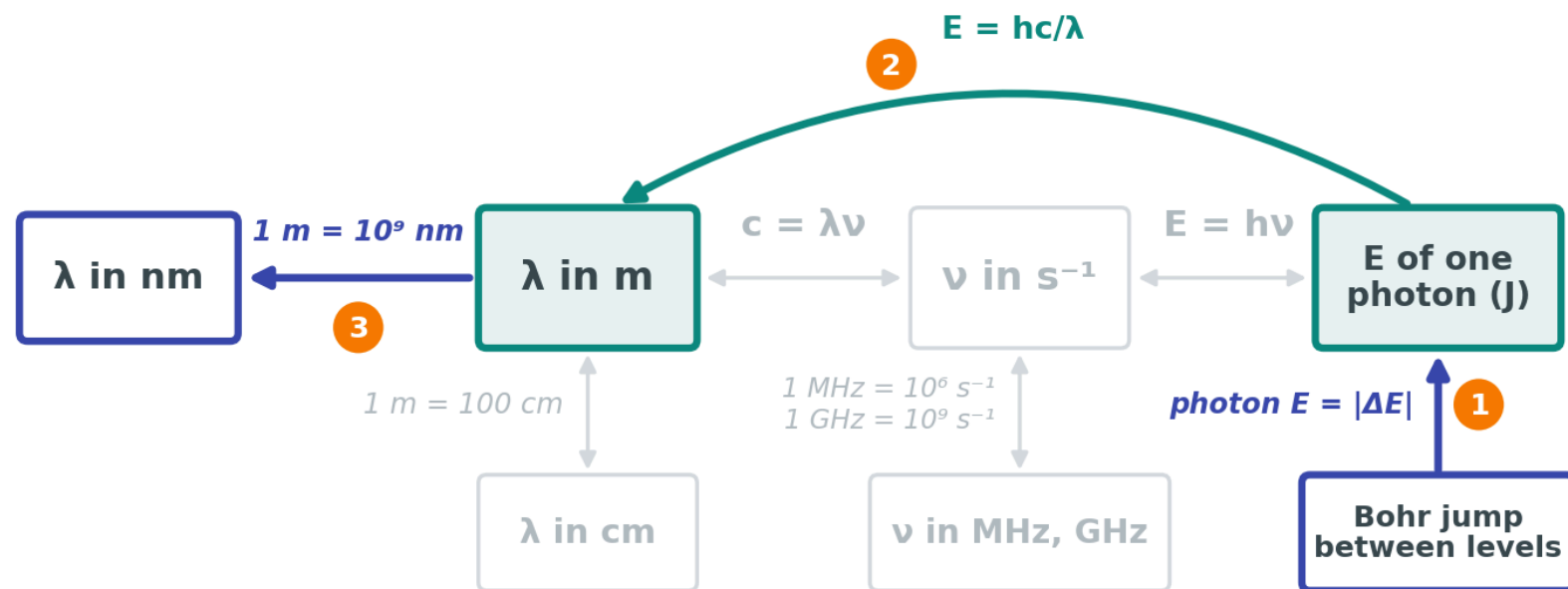


🧠 **Sense check:** From high on the ladder all the way down to $n = 1$: a large gap, so an ultraviolet photon. ✓

✓ Practice 5: the route on the light map

$$n_i = 7 \rightarrow n_f = 1$$

jump $\rightarrow |\Delta E|$ is the photon's $E \rightarrow \lambda = hc/E$, in m $\rightarrow$ nm



🧠 **Sense check:** Three arrows: jump to energy, energy to meters, meters to nanometers. ✓



Check yourself

1. An electron in hydrogen falls from $n = 5$ to $n = 2$. Compute ΔE , and state whether the photon is visible or ultraviolet.
2. In one sentence, explain why hydrogen produces a line spectrum instead of a continuous one.

Every element has its own set of energy-level gaps, so every element has its own line-spectrum fingerprint. The levels are real; the next question is how many electrons each one holds. Counting that capacity, level by level, builds an atom's full electron arrangement.

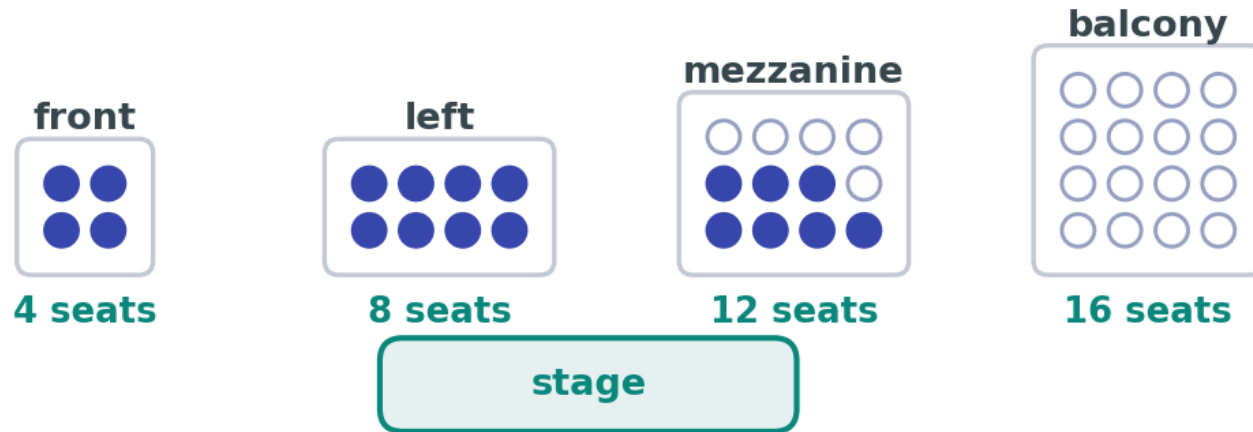
4 · Sublevels & Orbitals

Count the orbitals and the maximum electrons in any sublevel or shell, using that every orbital holds two electrons.

Filling a concert hall

Each section holds a fixed number of seats

$4 + 8 + 12 + 16 = 40$ — the hall's capacity is a sum you can count



A concert hall fills one section at a time. Each section holds a fixed number of seats. Add the seats section by section, and the hall's capacity is known exactly.

One orbital, at most two electrons

An electron shell is divided into sublevels. Each sublevel is built from orbitals, and every orbital holds at most two electrons. That limit is the Pauli exclusion principle. It fixes every capacity that follows.

1 orbital → 2 electrons maximum

the two electrons pair with opposite spins · no orbital holds a third

Four sublevels, four orbital counts

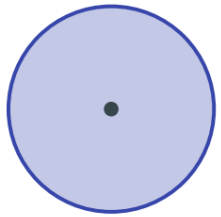
sublevel	Every orbital holds 2 electrons <i>orbitals</i>	<i>max electrons</i>
s		$1 \times 2 = \mathbf{2} \text{ e}^-$
p		$3 \times 2 = \mathbf{6} \text{ e}^-$
d		$5 \times 2 = \mathbf{10} \text{ e}^-$
f		$7 \times 2 = \mathbf{14} \text{ e}^-$

The sublevels are named s, p, d, and f, built from 1, 3, 5, or 7 orbitals apiece.

memory hook: the odd numbers, doubled
orbitals $1 \cdot 3 \cdot 5 \cdot 7 \rightarrow$ electrons $2 \cdot 6 \cdot 10 \cdot 14$

Three orbital shapes

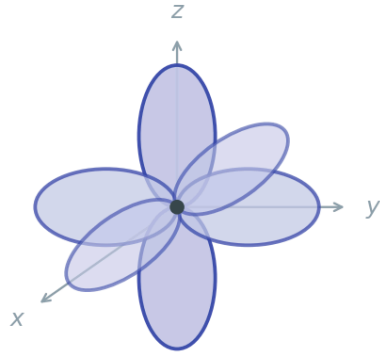
s orbital



sphere

1 per level

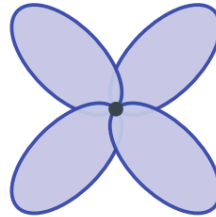
p orbitals



dumbbells along x, y, and z

3 per level, from n = 2

d orbital



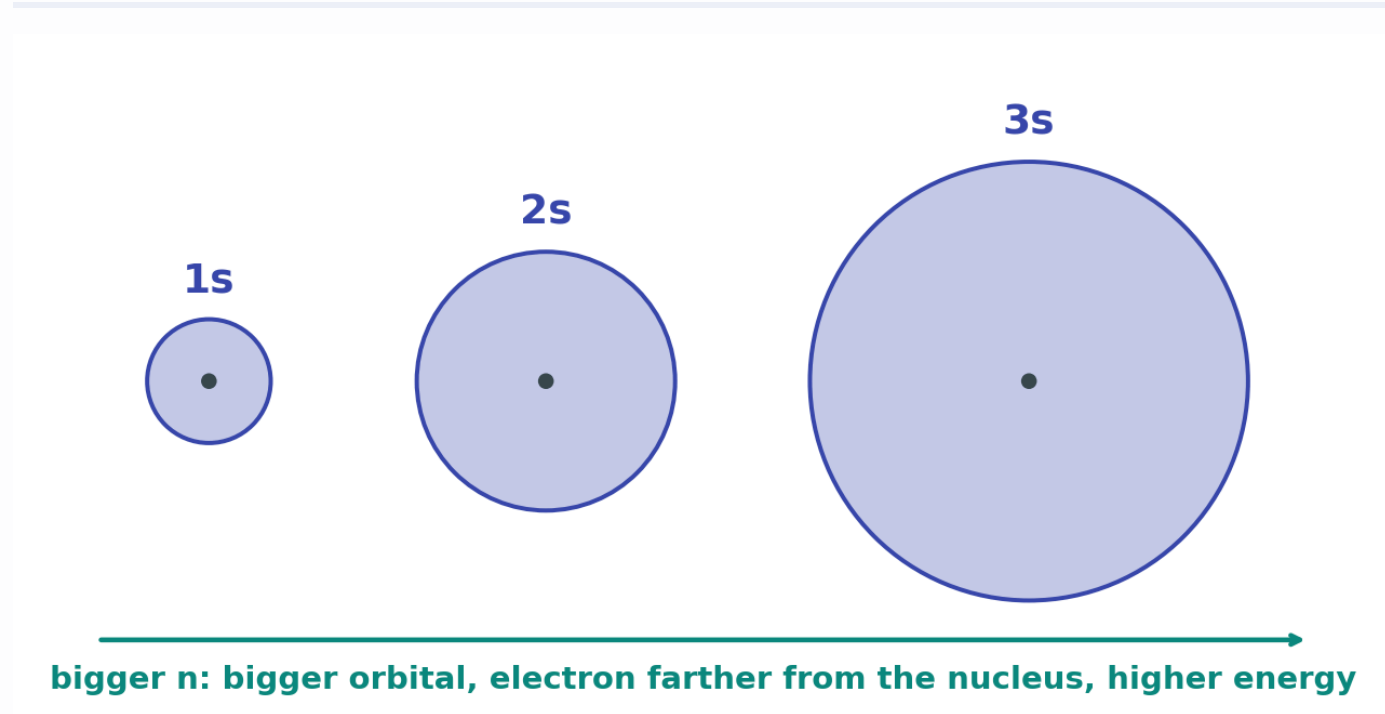
cloverleaf

5 per level, from n = 3

An orbital is the region where its electrons are most likely found. An s orbital is a sphere, the three p orbitals are dumbbells along x, y, and z, and most d orbitals are cloverleaves.

s = sphere · p = dumbbell · d = mostly cloverleaf
the letter sets the shape · every 2p, 3p, or 4p orbital is a dumbbell

Bigger n, bigger orbital



1s, 2s, and 3s are all spheres. The shell number n sets the size: the higher the shell, the larger the orbital and the farther its electrons sit from the nucleus.

$$1s < 2s < 3s$$

same shape, larger sphere · the letter sets the shape, n sets the size

A shell holds $2n^2$ electrons

shell n	sublevels	orbitals (n^2)	max electrons ($2n^2$)
1	1s	1	2
2	2s 2p	4	8
3	3s 3p 3d	9	18
4	4s 4p 4d 4f	16	32

n sublevels · n^2 orbitals · $2n^2$ electrons

shell 3 adds up: $2 + 6 + 10 = 18 = 2 \times 3^2$

Shell number n contains exactly n sublevels. Their orbitals total n^2 , and since each orbital holds two electrons, the shell's capacity is $2n^2$. Add the sublevels or use the formula: the count matches.



The method

1. **Name the sublevel.** The type sets its orbitals: s = 1, p = 3, d = 5, f = 7.
2. **Double the orbitals.** Each orbital holds 2 electrons.
3. **For a whole shell, add its sublevels,** or use $2n^2$.

Worked example 1: the 3p sublevel

3p sublevel

given: a p-type sublevel · wanted: its maximum electrons

Filled completely, how many electrons does the 3p sublevel hold? Name the type, then count.

✓ Worked example 1: solution

3p sublevel

given: a p-type sublevel · wanted: its maximum electrons

Step 1 · Name the sublevel

Step 2 · Double the orbitals

$3 \text{ orbitals} \times 2 = 6 \text{ electrons}$

each orbital holds 2 · a full 3p sublevel holds 6

 **Sense check:** Three orbitals, two electrons apiece. Six is the most the 3p sublevel can hold, the same as any p sublevel.

Worked example 2: the 3d sublevel

3d sublevel

given: a d-type sublevel · wanted: its maximum electrons

A d sublevel has 5 orbitals. A common first attempt: 5 orbitals, so 5 electrons. Name the type, then count carefully.

✓ Worked example 2: solution

3d sublevel

given: a d-type sublevel · wanted: its maximum electrons

A common first attempt

5 orbitals → 5 electrons?

that counts orbitals, not electrons: each orbital still holds 2 ×

Step 1 · Name the sublevel

Step 2 · Double the orbitals

5 orbitals × 2 = 10 electrons

five orbitals, two electrons each · a full 3d sublevel holds 10

 **Sense check:** Ten, not five. The orbital count and the electron count differ by exactly the factor of two that every orbital carries.

Your turn: fill the $n = 3$ shell

the $n = 3$ shell: 3s, 3p, 3d

three sublevels: double each orbital count, then add

sublevel	orbitals	electrons
3s	1	_____
3p	3	_____
3d	5	_____
whole shell	9	_____

Double each orbital count, then total the shell.

$$3s \rightarrow 2 \cdot 3p \rightarrow 6 \cdot 3d \rightarrow 10 \cdot \text{shell} \rightarrow 2 + 6 + 10 = 18$$

$18 = 2 \times 3^2$: the sublevels add to the shell's $2n^2$ capacity

⚠ Where this goes wrong

Reporting orbitals as electrons. A d sublevel has 5 orbitals, so "5 electrons" looks right. Each orbital holds 2, so the count is $5 \times 2 = 10$ electrons.

Giving the whole shell's capacity for one sublevel. Asked for the 3d electrons, answering 18 reports the entire $n = 3$ shell. One sublevel is not the shell: 3d holds 10.

Using $2n$ for a shell instead of $2n^2$. The $n = 4$ shell is not $2 \times 4 = 8$. Square n first: $2 \times 4^2 = 32$ electrons.

Miscounting d or f orbitals. A d sublevel has 5 orbitals, an f has 7, not the reverse. Calling d seven orbitals gives $7 \times 2 = 14$, too many.



Practice 1

the 5f sublevel

an f-type sublevel in the $n = 5$ shell · wanted: orbitals, not electrons

How many orbitals make up the 5f sublevel?

- A. 7
- B. 14
- C. 25
- D. 50

✓ Practice 1 · answer: A

5f → f type → 7 orbitals (answer A)

the 5 names the shell · an f sublevel always has 7 orbitals

B doubled the orbitals: $7 \times 2 = 14$ is the electron count, but the question asks for orbitals. C gave $n^2 = 5^2 = 25$, the orbitals in the whole $n = 5$ shell. D gave $2n^2 = 50$, the electron capacity of the whole $n = 5$ shell.

 **Sense check:** Every f sublevel has 7 orbitals, whatever its shell. Match the number to the question: orbitals stop at step 1 of the method.



Practice 2

$n = 1$ full · $n = 2$ full · $3s$ full · $3p$ holds 3 · nothing beyond
given: where the filling stops · wanted: the atom's total electrons

An atom's electrons completely fill the $n = 1$ and $n = 2$ shells and the $3s$ sublevel. The $3p$ sublevel holds 3 more, and nothing lies beyond it. How many electrons does the atom have?

- A. 9
- B. 12
- C. 13
- D. 15

✓ Practice 2 · answer: D

$n = 1: 2 \times 1^2 = 2 \cdot n = 2: 2 \times 2^2 = 8 \cdot 3s: 2 \cdot 3p: 3 \rightarrow 2 + 8 + 2 + 3 = 15$ (answer D)
two full shells by $2n^2$ · then the partly filled third shell, sublevel by sublevel

A counted orbitals, not electrons: $1 + 4 + 1 + 3 = 9$ boxes. B stopped before the 3p: $2 + 8 + 2 = 12$ leaves out the three 3p electrons. C skipped the 3s: $2 + 8 + 3 = 13$ drops the full 3s pair the stem names.

🧠 **Sense check:** Fifteen electrons with the filling stopped at $3p^3$: that atom is phosphorus, $Z = 15$. ✓



Check yourself

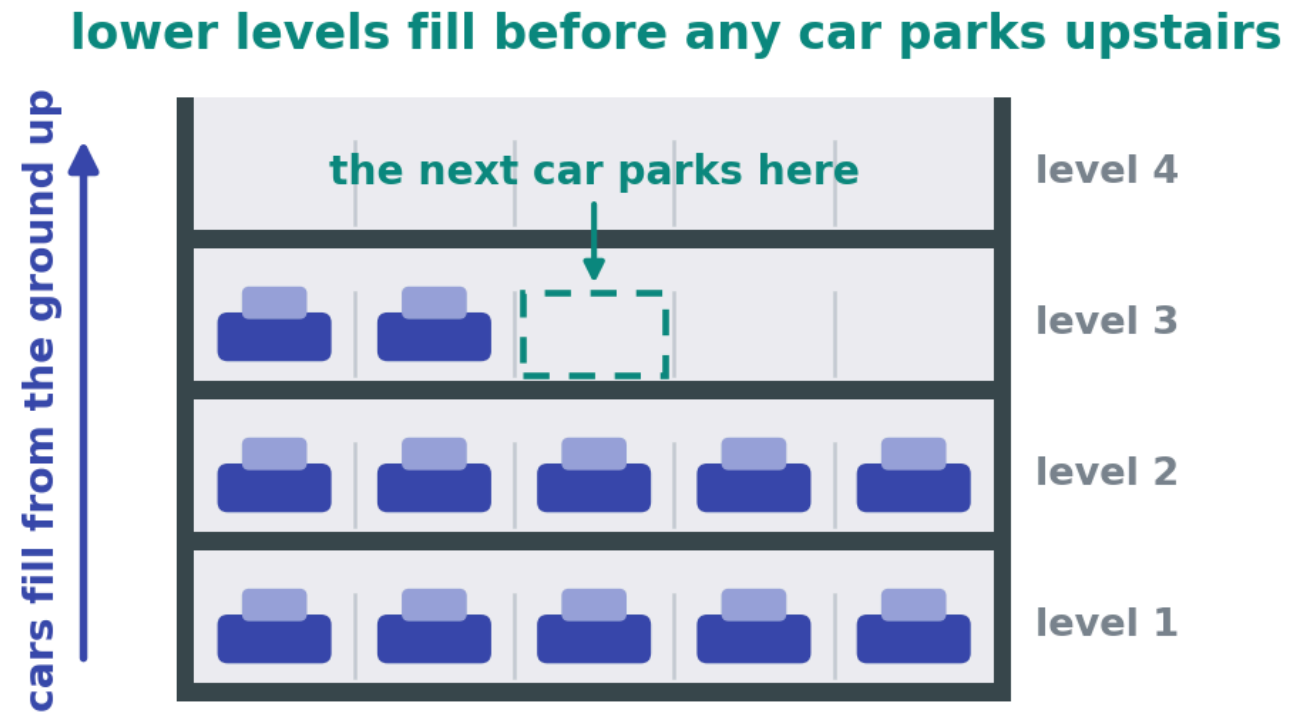
1. A full 4d sublevel: how many orbitals, and how many electrons? Name the type, then double.
2. Which shell first includes an f sublevel, and what is that shell's total capacity, $2n^2$?
3. Name the shape of a 2p orbital and of a 3s orbital. Which is larger, a 1s or a 3s orbital?

Every sublevel now has a known size. Filling them in order of increasing energy (1s, then 2s, then 2p) builds an atom's **electron configuration**, the ground-state arrangement of all its electrons.

5 • Electron Configurations

Write the ground-state electron configuration of any atom through the d block, using the building-up order and noble-gas core notation.

Filling from the ground up



A parking structure fills from the ground up. Every space on a lower level is taken before a single car parks on the level above.

Lowest levels fill first

Electrons fill sublevels from the lowest energy up, each one completely before the next. Listing every sublevel with its electron count gives the electron configuration. The superscripts must add up to the atom's total.

neon: $1s^2 2s^2 2p^6$

$2 + 2 + 6 = 10$ electrons · neon's atomic number is 10: the count matches ✓

Reading a configuration

shell number: $n = 2$

electrons in that sublevel: 6

2p⁶

sublevel: p

neon: $1s^2 2s^2 2p^6$

one term per sublevel, lowest energy first

superscripts: $2 + 2 + 6 = 10$ electrons

Each term names one sublevel. The number is the shell, n . The letter is the sublevel. The superscript counts the electrons in that sublevel. The superscripts add up to the atom's electrons.

The building-up order

Sublevels do not fill in simple numerical order. They fill by increasing energy, which the diagonal arrows trace: 1s, 2s, 2p, 3s, 3p, then 4s before 3d. A few elements are exceptions to this order.

follow each arrow, top to bottom

n	s	p	d
1	1s		
2	2s	2p	
3	3s	3p	3d
4	4s	4p	4d
5	5s	5p	5d

building-up order

1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p

4s fills before 3d

Reading the order from the table

The periodic table follows this same order. Reading left to right across a period gives the sublevels in turn: s block, then d block, then p block.

reading left to right across a period follows the building-up order

H															He
Li	Be														Ne
Na															
K	Ca	Sc													Kr

s block **d block (n - 1)** **p block**



The method

1. **Count the electrons:** a neutral atom's atomic number.
2. **Fill in the building-up order:** 1s, 2s, 2p, 3s, 3p, 4s, 3d.
3. **Fill sublevels to capacity:** s 2, p 6, d 10.
4. **Check the superscripts sum to the count.**



Guided example: hydrogen to carbon

Step 1 · Count the electrons

H 1 · He 2 · Li 3 · Be 4 · B 5 · C 6

atomic numbers · each element has one more electron than the one before

Build the six configurations in order. Each new electron enters the lowest sublevel with room.

✓ Guided example: solution

H 1 · He 2 · Li 3 · Be 4 · B 5 · C 6
atomic numbers · the electrons to place

Step 2 · Fill in the building-up order

H: $1s^1$ · He: $1s^2$
 $1s$ holds 2: full at helium

Step 3 · Fill sublevels to capacity

Li: $1s^2 2s^1$ · Be: $1s^2 2s^2$
 $2s$ holds 2: full at beryllium

 **Sense check:** Helium fills $1s$ and beryllium fills $2s$. The next electron needs a new sublevel: $2p$.

✓ Guided example: solution, continued

H 1 · He 2 · Li 3 · Be 4 · B 5 · C 6
atomic numbers · the electrons to place

Step 3 · Fill sublevels to capacity

B: $1s^2 2s^2 2p^1$ · C: $1s^2 2s^2 2p^2$
2p holds 6: carbon has placed 2 of them

Step 4 · Check the superscripts sum to the count

B: $2 + 2 + 1 = 5$ ✓ · C: $2 + 2 + 2 = 6$ ✓
each sum matches the atomic number

 **Sense check:** Each element adds one electron to the lowest sublevel with room: either the last superscript goes up by one, or a new sublevel opens with ¹ (Li $2s^1$, B $2p^1$).



Worked example 1: oxygen

Step 1 • Count the electrons

O: atomic number 8
a neutral atom has 8 electrons to place

Oxygen sits in the p block of period 2. Build its ground-state configuration from the lowest sublevel up.

✓ Worked example 1: solution

O: atomic number 8
8 electrons to place

Step 2 · Fill in the building-up order

Step 3 · Fill sublevels to capacity

$1s^2 2s^2 2p^4$
 $2 + 2 + 4 = 8$ electrons placed

Step 4 · Check the superscripts sum to the count

The superscripts total 8, matching oxygen's 8 electrons.

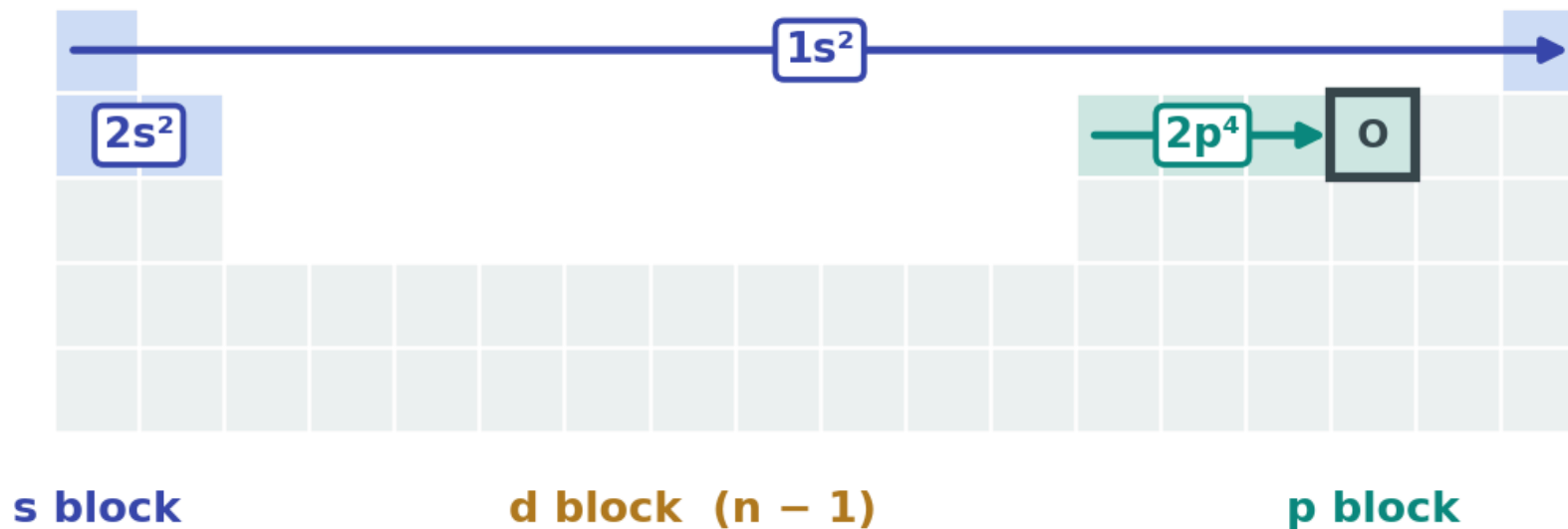
 **Sense check:** The final sublevel need not be full: 2p holds 6 but carries only 4 here. Configurations often end on a partly filled sublevel.

✓ Worked example 1: the route on the table



given: atomic number 8 · found: $2 + 2 + 4 = 8$ electrons

the route to oxygen: $1s^2 2s^2 2p^4$ · 8 electrons



🧠 **Sense check:** Period 1 gives $1s^2$, period 2 gives $2s^2$, then 2p up to oxygen, the fourth p-block cell: $2p^4$. ✓



Worked example 2: iron

Step 1 • Count the electrons

Fe: atomic number 26

26 electrons to place

Iron is a d-block metal in period 4. After $3p^6$, 18 electrons are placed and 8 remain. A common first attempt sends all 8 straight into 3d. Build the full configuration.

✓ Worked example 2: solution

Fe: atomic number 26

26 electrons to place

A common first attempt



$2 + 2 + 6 + 2 + 6 + 8 = 26$ ✓ count · ✗ order: 4s is lower in energy than 3d

Step 2 · Fill in the building-up order

Step 3 · Fill sublevels to capacity

Step 4 · Check the superscripts sum to the count



$2 + 2 + 6 + 2 + 6 + 2 + 6 = 26$ ✓ · matches iron's 26 electrons

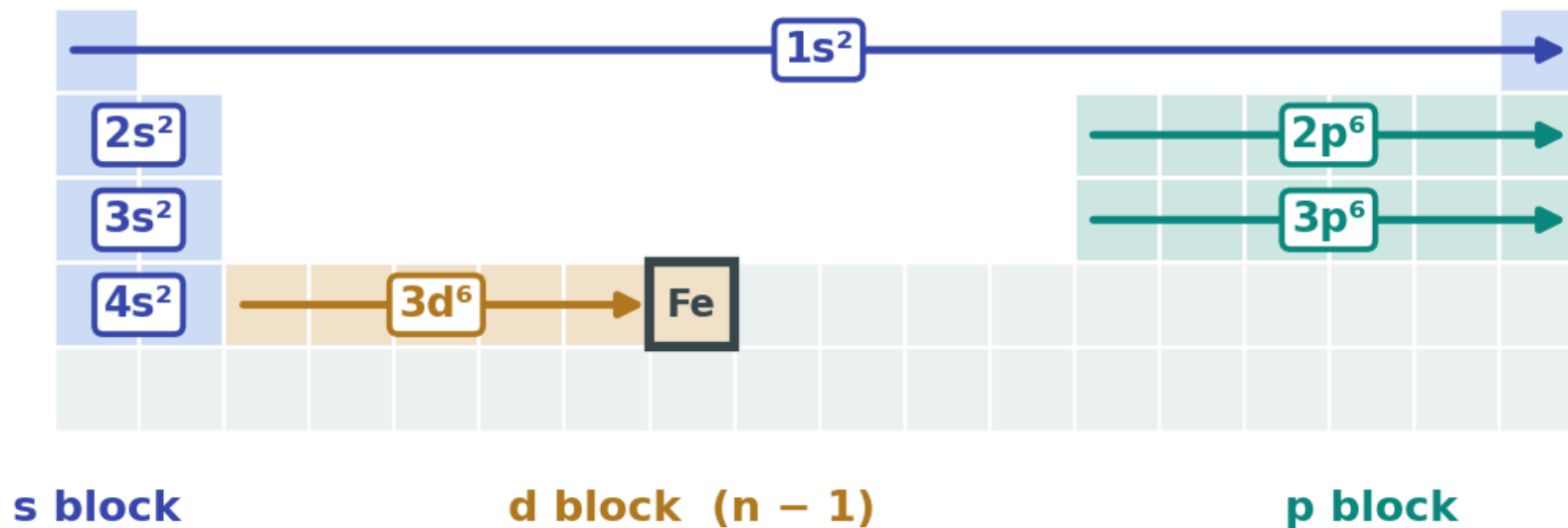
 **Sense check:** Both forms sum to 26, so the count alone cannot flag the error. The fixed order does: 4s fills before 3d.

✓ Worked example 2: the route on the table



given: atomic number 26 · found: $2 + 2 + 6 + 2 + 6 + 2 + 6 = 26$ electrons

the route to iron: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ · 26 electrons



🧠 **Sense check:** Period 4 opens with $4s^2$ (K, Ca) before any d. Iron is the sixth d-block cell: $3d^6$. ✓

Your turn: nickel

Nickel has atomic number 28. Fill the blanks, then confirm the superscripts sum to 28.



sublevel	electrons
through $3p^6$	18
4s	<u> </u>
3d	<u> </u>



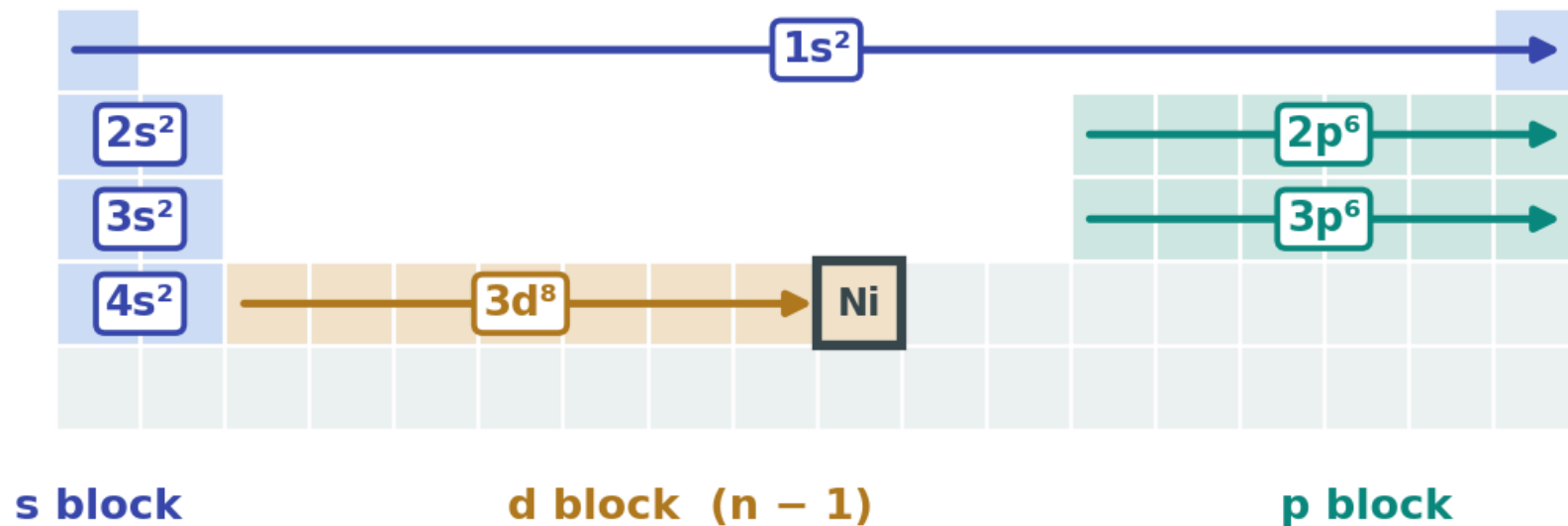
$$2 + 2 + 6 + 2 + 6 + 2 + 8 = 28 \checkmark \cdot 4s \text{ takes } 2, \text{ then } 3d \text{ takes } 28 - 18 - 2 = 8$$

✓ Nickel: the route on the table



given: atomic number 28 · found: $2 + 2 + 6 + 2 + 6 + 2 + 8 = 28$ electrons

the route to nickel: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ · 28 electrons



🧠 **Sense check:** The same route as iron, two cells farther. Nickel is the eighth d-block cell: $3d^8$. ✓

⚠ Where this goes wrong

Filling 3d before 4s. For iron, 26 electrons, writing $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8$ sums to 26, but 4s is lower in energy than 3d. The ground state is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$. The count passes; only the order catches this.

Superscripts that miss the count. Stopping calcium at $1s^2 2s^2 2p^6 3s^2 3p^6$ gives $2 + 2 + 6 + 2 + 6 = 18$. That is argon, not calcium. Calcium, 20 electrons, needs $4s^2$: two more.

Overfilling a sublevel. Writing $2p^8$ to reach the count faster claims p holds 8. Each s holds 2, p holds 6, d holds 10. Move to the next sublevel when the current one is full.

Reading only the valence electrons. A configuration ending $3s^2 3p^2$ has 4 valence electrons, but the atom is not element 4. Every superscript counts toward the atomic number, not just the last sublevel.



Practice 1



a ground-state configuration

This configuration describes the ground state of which element?

- A. Nitrogen
- B. Oxygen
- C. Phosphorus
- D. Boron

✓ Practice 1 · answer: A



$2 + 2 + 3 = 7$ electrons: answer A, nitrogen (Z 7)

The superscripts add to 7, so the atomic number is 7: nitrogen. B, oxygen, has one electron too many: $2 + 2 + 4 = 8$. C, phosphorus, shares the p^3 ending but sits a period lower, at 15: $2 + 2 + 6 + 2 + 3 = 15$. D, boron, matches only the 5 valence electrons, $2 + 3 = 5$, not all 7.

 **Sense check:** The superscripts count every electron, not just the outer ones. Their sum, 7, is the atomic number.



Practice 2

$$K \cdot Z = 19$$

a neutral atom in its ground state

Which lists potassium's 19 electrons correctly, sublevel by sublevel?

- A. K: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$
- B. K: $1s^2 2s^2 2p^6 3s^2 3p^7$
- C. K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
- D. K: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

✓ Practice 2 · answer: D



$$2 + 2 + 6 + 2 + 6 + 1 = 19 \checkmark \cdot \text{answer D, potassium's 19 electrons}$$

After $3p^6$, 18 electrons are placed and 1 remains; it enters 4s. A put it in 3d: $2 + 2 + 6 + 2 + 6 + 1 = 19$ counts right, but 4s is lower in energy than 3d. B wrote $3p^7$: 19 again, but a p sublevel holds only 6. C filled 4s to capacity: $2 + 2 + 6 + 2 + 6 + 2 = 20$, calcium.

 **Sense check:** Two of the wrong answers also sum to 19. The count, the order and each sublevel's capacity must all check.

Noble-gas core notation

A configuration that begins with a noble gas can be abbreviated by that gas's symbol in brackets. [Ne] stands for $1s^22s^22p^6$; [Ar] for $1s^22s^22p^63s^23p^6$. Write the core, then the sublevels beyond it.



[Ne] = 10 electrons · + $3s^1$ = 11 total, sodium's count ✓



Worked example 3: titanium

Step 1 • Count the electrons

Ti: atomic number 22

22 electrons to place

Titanium is a d-block metal in period 4. Write its full ground-state configuration, then abbreviate it with a noble-gas core.

✓ Worked example 3: solution

Ti: atomic number 22

22 electrons to place

Step 2 · Fill in the building-up order

Step 3 · Fill sublevels to capacity



$$2 + 2 + 6 + 2 + 6 + 2 + 2 = 22 \text{ electrons placed}$$

Collapse to the noble-gas core

Step 4 · Check the superscripts sum to the count

$1s^2 2s^2 2p^6 3s^2 3p^6$ is argon, so it becomes [Ar]. The electrons still total 22.



$$[\text{Ar}] = 18 \cdot 18 + 2 + 2 = 22 \checkmark \cdot \text{titanium's 22 electrons}$$

🧠 **Sense check:** [Ar] hides 18 electrons but keeps them: $18 + 2 + 2 = 22$, titanium's atomic number.



Practice 3

Br: atomic number 35
35 electrons to place

Which is the noble-gas core configuration of bromine?

- A. $[\text{Kr}]4s^23d^{10}4p^5$
- B. $[\text{Ar}]4s^24p^5$
- C. $[\text{Ar}]4s^23d^{10}4p^5$
- D. $[\text{Ar}]4s^23d^{10}4p^6$

✓ Practice 3 · answer: C



$$[\text{Ar}] = 18 \cdot 18 + 2 + 10 + 5 = 35 \checkmark \cdot \text{answer C, bromine's 35 electrons}$$

Argon is the noble gas before bromine. After [Ar], 4s fills, then 3d to 10, then 4p. A used krypton, the noble gas after bromine: $36 + 2 + 10 + 5 = 53$. B skipped 3d: $18 + 2 + 5 = 25$. D filled 4p to capacity: $18 + 2 + 10 + 6 = 36$, krypton itself.

🧠 **Sense check:** Period 4 crosses the whole d block before its p block: $3d^{10}$ sits between $4s^2$ and $4p$.

Practice 4

Which element is paired with a wrong ground-state configuration?

- A. Sc: $[\text{Ar}]4s^23d^1$
- B. Ge: $[\text{Ar}]4s^23d^84p^4$
- C. Cl: $[\text{Ne}]3s^23p^5$
- D. F: $1s^22s^22p^5$

✓ Practice 4 · answer: B

Ge: $[\text{Ar}]4s^23d^84p^4$ ✗

$18 + 2 + 8 + 4 = 32$ ✓ count · ✗ order: 3d fills to 10 before 4p · Ge is $[\text{Ar}]4s^23d^{10}4p^2$, $18 + 2 + 10 + 2 = 32$ (answer B)

The count passes, so only the order flags B: after $4s^2$, 3d takes all 10 before any electron enters 4p. A is correct: Sc, $18 + 2 + 1 = 21$, with 4s filled before 3d. C is correct: Cl, $10 + 2 + 5 = 17$, and [Ne] is the noble gas before Cl. D is correct: F, $2 + 2 + 5 = 9$; a last sublevel may be partly filled.

🧠 **Sense check:** A matching count cannot clear a configuration. Check the order too: 4s, then 3d to 10, then 4p.

Check yourself

1. Write the ground-state configuration of sulfur (Z 16), and confirm the superscripts sum to 16.
2. In period 4, which fills first, 4s or 3d? Give the noble-gas-core configuration of calcium (Z 20).

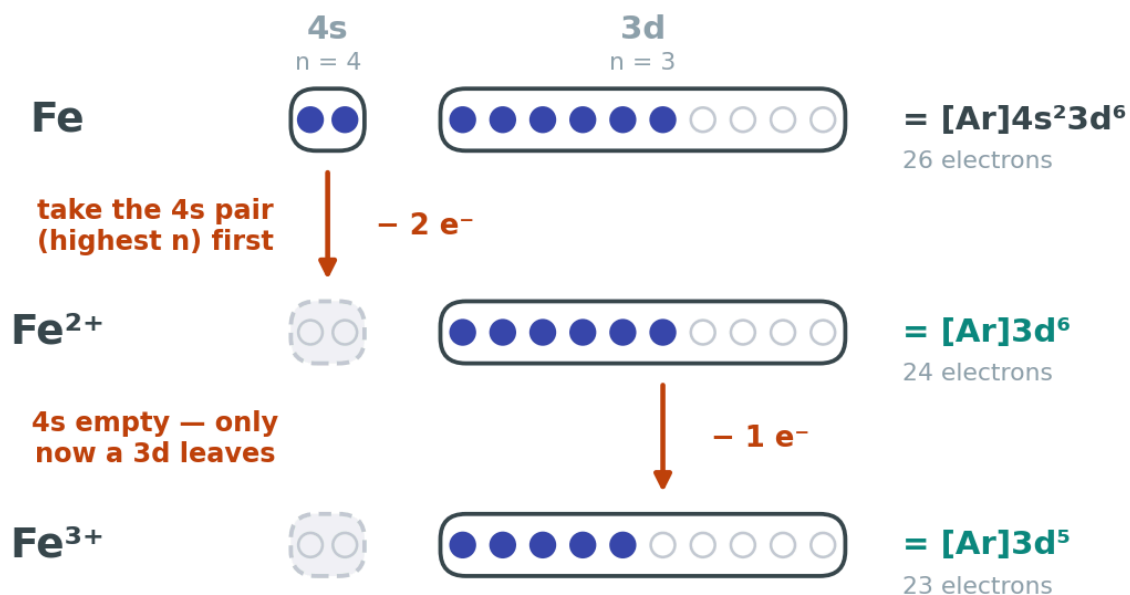
Atoms rarely keep every electron. An ion's configuration starts from the neutral atom's: electrons leave from the highest n first, and gained ones continue the filling order.

6 · Electron Configurations of Ions

Write the ground-state configuration of any ion by removing electrons from the highest n first.

Emptying from the top

Ionizing iron: the highest-n electrons leave first



4s fills before 3d — and empties before 3d too: always remove the highest n first.

When cars leave a parking structure, the top floor empties first. Atoms lose electrons the same way: the highest-energy shell goes first, not the last one filled.

Counting an ion's electrons

An ion's electron count no longer equals its atomic number. The protons never change. A positive ion lost electrons; a negative ion gained them.

lose one electron per + · gain one electron per –

K^+ : $19 - 1 = 18$ electrons · N^{3-} : $7 + 3 = 10$ electrons · protons unchanged

A cation is a neutral atom, minus electrons

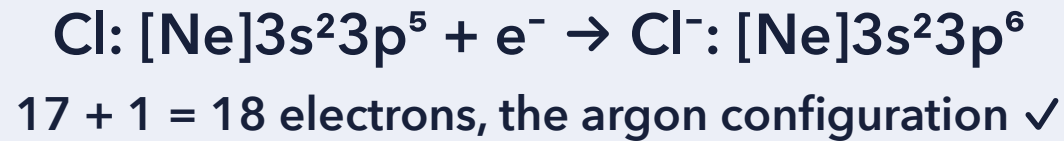
Start from the neutral atom and remove one electron per unit of positive charge, taking the highest principal level n first. For main-group atoms those are also the last electrons added.



remove the single $n = 3$ electron (the $3s$) · $11 - 1 = 10$ electrons, matching the $1+$ charge

An anion adds electrons in the filling order

An anion gains electrons, and they continue the normal building-up order. Chlorine gains one electron to finish its 3p sublevel and reach argon's configuration.





The method

1. **Write the neutral atom** in noble-gas form.
2. **Count the charge:** one electron lost per plus, gained per minus.
3. **Remove from the highest n first.** Anions add in filling order.
4. **Check:** protons – electrons = charge.

Guided example: three ions

Step 1 • Write the neutral atom



$2 + 2 + 4 = 8 \cdot 2 + 2 + 5 = 9 \cdot 2 + 2 + 6 + 2 = 12$ electrons, each the atomic number

Write the configurations of O^{2-} , F^- and Mg^{2+} , then compare them.

Count the charge first: the two anions gain electrons, and magnesium loses them.

✓ Guided example: solution

O: $1s^2 2s^2 2p^4$ · F: $1s^2 2s^2 2p^5$ · Mg: $1s^2 2s^2 2p^6 3s^2$
8, 9 and 12 electrons in the neutral atoms

Step 2 · Count the charge

O^{2-} : $8 + 2 = 10$ · F^- : $9 + 1 = 10$ · Mg^{2+} : $12 - 2 = 10$
10 electrons in each ion

Step 3 · Remove from the highest n first

O^{2-} , F^- and Mg^{2+} : $1s^2 2s^2 2p^6 = [Ne]$
 $2 + 2 + 6 = 10$ electrons each ✓

Step 4 · Check

🧠 **Sense check:** Protons – electrons: $8 - 10 = -2$, $9 - 10 = -1$, $12 - 10 = +2$ ✓ All three match neon.



Practice 1



strontium ions give fireworks their red color

Which is the ground-state electron configuration of Sr^{2+} ?

- A. $[\text{Kr}]5s^24d^2$
- B. $[\text{Kr}]5s^1$
- C. $[\text{Kr}]5s^2$
- D. $[\text{Kr}]$

✓ Practice 1 · answer: D

Sr: $[\text{Kr}]5s^2 \rightarrow \text{Sr}^{2+}$: $[\text{Kr}]$ (answer D)

$36 + 2 = 38$ electrons · lose 2 · $38 - 2 = 36$, krypton's count ✓

A added two electrons for the 2+ charge: $38 + 2 = 40$, continuing into 4d. B removed only one: $38 - 1 = 37$, which is Sr^+ . C removed nothing: $[\text{Kr}]5s^2$ is neutral strontium, 38 electrons.

🧠 **Sense check:** A 2+ ion lost two electrons. Strontium's 5s pair leaves, and krypton's configuration remains. ✓



Practice 2



the sulfide ion in zinc sulfide, a glow-in-the-dark pigment

Which configuration does the sulfide ion have in its ground state?

- A. $[\text{Ne}]3s^23p^4$
- B. $[\text{Ne}]3s^23p^6$
- C. $[\text{Ne}]3s^23p^2$
- D. $[\text{Ne}]3s^23p^44s^2$

✓ Practice 2 · answer: B

S: $[\text{Ne}]3s^23p^4 \rightarrow \text{S}^{2-}: [\text{Ne}]3s^23p^6$ (answer B)

$10 + 2 + 4 = 16$ electrons · gain 2 · $16 + 2 = 18 = 10 + 2 + 6$, argon's count ✓

A gained nothing: $[\text{Ne}]3s^23p^4$ is neutral sulfur, 16 electrons. C removed two electrons for the 2− charge: $16 - 2 = 14$. D put the gained pair in 4s: $10 + 2 + 4 + 2 = 18$ counts right, but 3p still had room.

🧠 **Sense check:** A 2− ion gained two electrons. They finish 3p, so sulfide matches argon. ✓

The transition-metal twist: lose ns before $(n-1)d$

The building-up order fills $4s$ before $3d$. Once $3d$ is occupied it sinks below $4s$, leaving $4s$ as the outermost shell: the highest n , and the first to leave.

remove the highest n first: $n = 4$ (the $4s$) beats $n = 3$ (the $3d$)

$4s$ fills first, and $4s$ empties first: strip ns before touching $(n-1)d$

Worked example 1: iron's cations

Fe: atomic number 26 $\rightarrow$ $[\text{Ar}]4s^23d^6$

18 (argon core) + 2 + 6 = 26 electrons ✓

Iron gives up electrons to form Fe^{2+} and Fe^{3+} . A common first attempt pulls them out of 3d, since 3d was filled last. Build both ions correctly.

✓ Worked example 1: solution

Fe: atomic number 26 $\rightarrow$ [Ar]4s²3d⁶

18 (argon core) + 2 + 6 = 26 electrons ✓

Step 1 • Write the neutral atom

Step 2 • Count the charge

Step 3 • Remove from the highest n first

Fe²⁺: [Ar]3d⁶

4s emptied, 3d untouched • 18 + 6 = 24 electrons • 26 – 24 = 2 $\rightarrow$ 2+ ✓

Step 3 • Remove from the highest n first

Fe³⁺: [Ar]3d⁵

18 + 5 = 23 electrons • 26 – 23 = 3 $\rightarrow$ 3+ ✓

Step 4 • Check

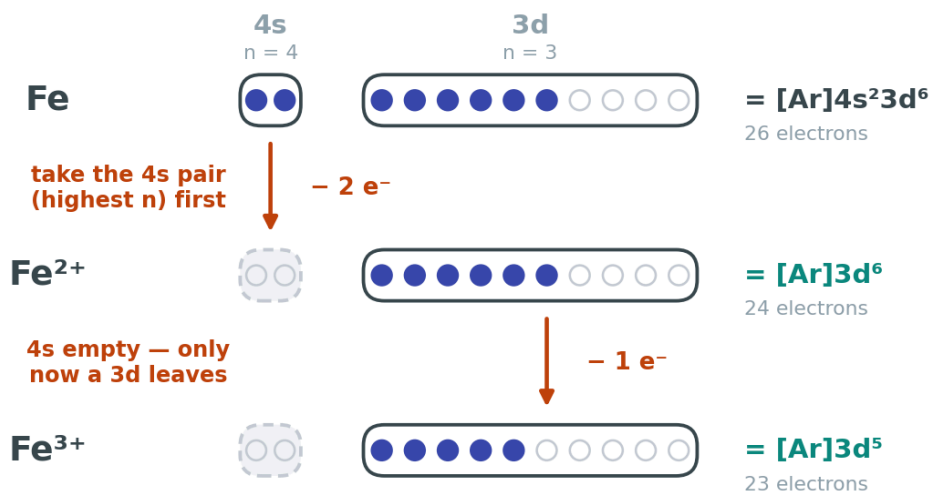
🧠 **Sense check:** Iron keeps its 26 protons in both ions. Highest n leaves first, so Fe²⁺ is [Ar]3d⁶, never [Ar]4s²3d⁴.

✓ Worked example 1: sublevel by sublevel



found: $\text{Fe}^{2+} = [\text{Ar}]3d^6$ (24 electrons) · $\text{Fe}^{3+} = [\text{Ar}]3d^5$ (23 electrons)

Ionizing iron: the highest-n electrons leave first



4s fills before 3d — and empties before 3d too: always remove the highest n first.

 **Sense check:** The 4s sublevel empties first. Only then does a 3d electron leave. ✓

🧩 Your turn: manganese

Mn: atomic number 25 → [Ar]4s²3d⁵

18 + 2 + 5 = 25 electrons ✓

step	question	answer
1 · write the neutral atom	noble-gas form of Mn	[Ar]4s ² 3d ⁵
2 · count the charge	electrons lost for Mn ²⁺	_____
3 · remove from the highest n first	which sublevel empties?	_____
4 · check	resulting configuration	_____

Complete the steps for Mn²⁺.

Mn²⁺: [Ar]3d⁵

the 4s pair leaves · 18 + 5 = 23 electrons · 25 – 23 = 2 → 2+ ✓



Practice 3



the ion that turns nickel salt solutions green

Which configuration belongs to Ni^{2+} ?

- A. $[\text{Ar}]3\text{d}^8$
- B. $[\text{Ar}]4\text{s}^23\text{d}^6$
- C. $[\text{Ar}]4\text{s}^23\text{d}^{10}$
- D. $[\text{Ar}]4\text{s}^23\text{d}^8$

✓ Practice 3 · answer: A

Ni: $[\text{Ar}]4s^23d^8 \rightarrow \text{Ni}^{2+}: [\text{Ar}]3d^8$ (answer A)

$18 + 2 + 8 = 28$ electrons · the 4s pair leaves · $18 + 8 = 26$ · $28 - 26 = 2 \rightarrow 2+$ ✓

B took both electrons from 3d, the last filled: $18 + 2 + 6 = 26$ counts right, but 4s ($n = 4$) empties first. C added two electrons for the 2+ charge: $28 + 2 = 30$. D removed nothing: $[\text{Ar}]4s^23d^8$ is neutral nickel, 28 electrons.

🧠 **Sense check:** Like iron, nickel empties 4s before any 3d electron leaves. ✓

⚠ Where this goes wrong

Removing 3d before 4s. [Ar]4s²3d⁴ treats "last filled" as "first removed." Highest n leaves first: 4s before 3d. Fe²⁺ is [Ar]3d⁶.

Forgetting to start neutral. You can't remove from a shell never filled: write the full neutral configuration first, then subtract.

Miscounting the charge. A 3+ ion lost three electrons, not protons. Fe is always 26 p⁺; 26 – 23 e⁻ = 3+ ✓.

Expecting a noble-gas ion. Main-group ions reach a noble gas; transition-metal cations do not. Fe³⁺ is [Ar]3d⁵, five electrons past argon.



Practice 4



the cathode of a spent alkaline battery

Which is the ground-state electron configuration of the manganese ion in Mn_2O_3 ?

- A. $[\text{Ar}]3\text{d}^5$
- B. $[\text{Ar}]4\text{s}^23\text{d}^2$
- C. $[\text{Ar}]3\text{d}^4$
- D. $[\text{Ar}]3\text{d}^1$
- E. $[\text{Ar}]4\text{s}^23\text{d}^5$

✓ Practice 4 · answer: C

Mn_2O_3 : 3 oxides $\times 2- = 6-$, shared by 2 Mn: $6 \div 2 = 3+$ each
the 2 in Mn_2 counts manganese atoms, not charge

Mn: $[\text{Ar}]4s^23d^5 \rightarrow \text{Mn}^{3+}$: $[\text{Ar}]3d^4$ (answer C)
the 4s pair first, then one 3d · $18 + 4 = 22$ electrons · $25 - 22 = 3 \rightarrow 3+ \checkmark$

A read the 2 in Mn_2 as the charge: $[\text{Ar}]3d^5$ is Mn^{2+} , $18 + 5 = 23$ electrons. B took all three from 3d: $18 + 2 + 2 = 22$ counts right, but 4s ($n = 4$) empties first. D put the whole $6-$ on one Mn: Mn^{6+} , $25 - 6 = 19$ electrons, $[\text{Ar}]3d^1$. E removed nothing: that is neutral manganese, 25 electrons.

 **Sense check:** The formula sets the charge; the configuration still empties 4s before 3d. ✓

Check yourself

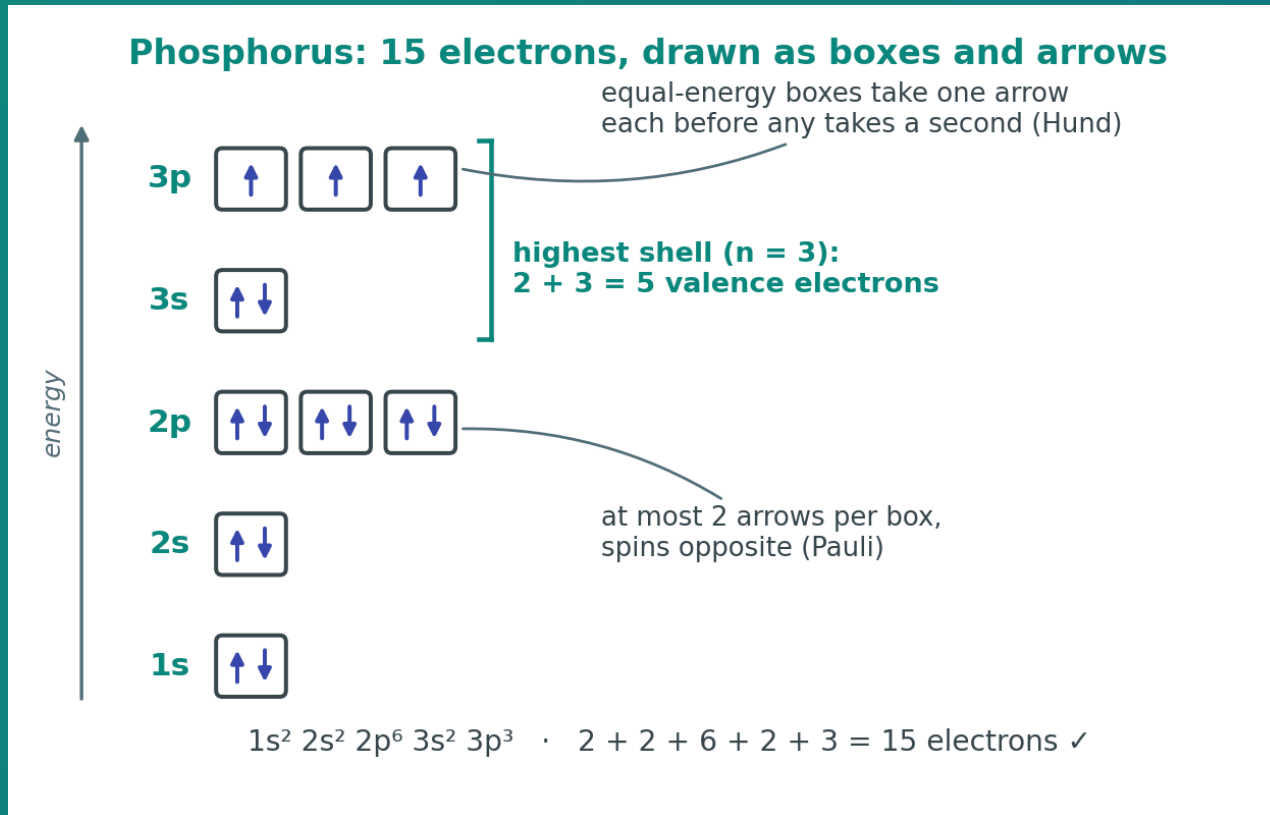
1. Cobalt is $[\text{Ar}]4s^23d^7$. Write the configuration of Co^{2+} and name which electrons left first.
2. Write the configuration of Br^- in noble-gas form, and name the noble gas it matches.

Each sublevel here is still one total. Next, every orbital gets its own box, and the arrows show which electrons pair up and which sit alone.

7 · Orbital Diagrams & Valence Electrons

Draw the orbital diagram of any element through calcium and of every d-block atom that follows the filling order, read a diagram back into its configuration, and count valence electrons from the configuration or from the group number.

Where the electrons sit



A phosphorus atom carries 15 electrons. Drawn as boxes and arrows, they land the same way in every atom, and the outermost few do all the chemistry.

One box is one orbital

An orbital diagram draws each orbital as a box and each electron as an arrow. An orbital holds at most two electrons, and a pair's spins must be opposite: the Pauli exclusion principle.

1 box = 1 orbital → at most $\uparrow\downarrow$

two arrows per box, never parallel · a third has no spin left to take

Period 2, one electron at a time

		1s	2s	2p	configuration	
H	Z = 1	↑			1s ¹	
He	Z = 2	↑↓			1s ²	
Li	Z = 3	↑↓	↑		1s ² 2s ¹	1s is full: 2s opens
Be	Z = 4	↑↓	↑↓		1s ² 2s ²	
B	Z = 5	↑↓	↑↓	↑	1s ² 2s ² 2p ¹	2s is full: 2p opens

teal arrow: the electron each element adds

Each next element adds one electron to the lowest open box. 1s holds two, so lithium's third electron opens 2s. Boron's fifth opens 2p.

Equal-energy boxes fill singly first

The three p boxes of a sublevel share one energy. Paired electrons sit close and repel, so electrons spread out, one per box with parallel spins, before any box takes a second: Hund's rule.

2p³: ↑ ↑ ↑ ✓ · ↑↓ ↑ (one box empty) ✕

three singles, spins parallel · pairing starts only when every box holds one

memory hook: the bus-seat rule

riders take an empty seat before sitting beside a stranger · electrons take an empty box before pairing

Carbon to neon: singles, then pairs

		1s	2s	2p	configuration	
C	Z = 6				$1s^2 2s^2 2p^2$	empty box, not a pair (Hund)
N	Z = 7				$1s^2 2s^2 2p^3$	one in each 2p box
O	Z = 8				$1s^2 2s^2 2p^4$	pairing begins
F	Z = 9				$1s^2 2s^2 2p^5$	
Ne	Z = 10				$1s^2 2s^2 2p^6$	n = 2 is full

teal arrow: the electron each element adds

Carbon's second 2p electron takes an empty box, not a partner. Nitrogen puts one in each. From oxygen on, every box already holds one, so each new electron pairs.

Core and valence electrons

The electrons in the highest occupied shell are the valence electrons; everything beneath is the core. Valence electrons sit outermost, so they meet other atoms first and do the bonding.



core $2 + 2 + 6 = 10$ · valence $2 + 3 = 5$, the whole shell with $n = 3$

The group number counts them

Same column, same number of outer electrons. Groups 1 and 2 state the valence count directly. Groups 13 to 18 count past the ten transition-metal columns: subtract 10.

N: group 15 $\rightarrow$ 15 – 10 = 5 valence electrons

Ca: group 2 $\rightarrow$ 2 valence electrons · the configurations agree: $2s^2 2p^3$ and $4s^2$



The method

1. **Count the electrons:** the atomic number.
2. **Fill the boxes lowest-energy first:** 1s, 2s, 2p, 3s, 3p, 4s, 3d.
3. **Within a sublevel, singly first:** one per box, then pair.
4. **Mark the highest shell:** its electrons are the valence electrons.



Worked example 1: nitrogen

Step 1 • Count the electrons

N: atomic number 7 • group 15
7 electrons to place

Draw nitrogen's orbital diagram, then count its valence electrons two ways: from the diagram and from the group number.

✓ Worked example 1: solution

Step 2 · Fill the boxes lowest-energy first

Step 3 · Within a sublevel, singly first

1s: $\uparrow\downarrow$ · 2s: $\uparrow\downarrow$ · 2p: $\uparrow \uparrow \uparrow$

$2 + 2 + 3 = 7$ · 3 unpaired · pairing would begin only with a fourth

Step 4 · Mark the highest shell

The highest occupied shell is $n = 2$, holding $2s^2$ and $2p^3$. The 1s pair is the core.

valence: $2 + 3 = 5$ · group check: $15 - 10 = 5$ ✓

both counts land on 5 valence electrons

 **Sense check:** Nitrogen holds 5 valence electrons, 3 of them unpaired; only the diagram shows which sit alone.

Your turn: silicon

Si: atomic number 14 · group 14

14 electrons to place · highest occupied shell: $n = 3$

step	answer
configuration	$1s^2 2s^2 2p^6 3s^2 3p$ _____
the 3p boxes	_____
valence electrons	_____
unpaired electrons	_____

Fill the four blanks by the method, then check the valence count against the group number.

$1s^2 2s^2 2p^6 3s^2 3p^2 \cdot 3p: \uparrow \uparrow$ (one box empty)

$2 + 2 + 6 + 2 + 2 = 14 \checkmark \cdot \text{valence } 2 + 2 = 4 = 14 - 10 \cdot 2 \text{ unpaired}$

⚠ Where this goes wrong

A third arrow in one box. Writing lithium as $1s^3$ crams three electrons into one orbital. Only two spin directions exist, so a third electron has no partner slot: it must move up. Lithium is $1s^2 2s^1$.

Pairing before every box has one. Drawing $3p^4$ as $\uparrow\downarrow \uparrow\downarrow$ (one box empty) shows no unpaired electrons. Singles come first: $\uparrow\downarrow \uparrow \uparrow$, leaving $6 - 4 = 2$ unpaired. The count changes with the drawing.

Counting every electron as valence. Chlorine holds 17 electrons, but $2 + 2 + 6 = 10$ of them are core. Valence electrons live only in the highest shell: $3s^2 3p^5$ gives $2 + 5 = 7$.

Reading a p-block group number straight. Sulfur sits in group 16 yet holds 6 valence electrons, not 16. Groups 13 to 18 count past the ten transition-metal columns: $16 - 10 = 6$.



Practice 1

Cl: atomic number 17 · group 17



How many valence electrons does a chlorine atom hold?

- A. 5
- B. 17
- C. 7
- D. 10

✓ Practice 1 · answer: C

Cl: $3s^23p^5 \rightarrow 2 + 5 = 7$ valence electrons (answer C)

group check: $17 - 10 = 7$ · the two counts agree

B counts the whole atom: 17 is every electron, core included. A reads only the last superscript: $3p^5$ gives 5, dropping the 3s pair in the same shell. D counts the core, $2 + 2 + 6 = 10$, the electrons that stay out of the chemistry.

 **Sense check:** Valence means the full highest shell, 3s and 3p together. Seven, one short of argon's eight, is why chlorine takes up one electron so readily.



Practice 2

Pb: group 14 · period 6

How many valence electrons does lead have?

- A. 14
- B. 4
- C. 2
- D. 6

✓ Practice 2 · answer: B

Pb: group 14 $\rightarrow 14 - 10 = 4$ valence electrons (answer B)

highest shell $n = 6$: $6s^2 6p^2 \rightarrow 2 + 2 = 4$ · the two counts agree

A read group 14 straight, skipping the 10 that groups 13 to 18 count past; counting the filled $5d^{10}$ as valence lands on the same $10 + 2 + 2 = 14$. C read only the last superscript, $6p^2$, and dropped the $6s$ pair. D read the period, 6: it numbers the highest shell, not its electrons.

 **Sense check:** Lead sits below carbon and silicon, so it shares their 4 valence electrons. A heavy atom still bonds with only its outer few. ✓



Practice 3

O: atomic number 8 · group 16

How many unpaired electrons does an oxygen atom have?

- A. 2
- B. 4
- C. 0
- D. 6

✓ Practice 3 · answer: A

1s: $\uparrow\downarrow$ · 2s: $\uparrow\downarrow$ · 2p: $\uparrow\downarrow$ $\uparrow$ $\uparrow$

$2 + 2 + 4 = 8$ ✓ · three 2p singles, then the fourth pairs · $4 - 2 = 2$ unpaired (answer A)

B read the superscript of $2p^4$ as four lone arrows, but three boxes hold only three singles.
C paired first, $\uparrow\downarrow$ $\uparrow\downarrow$ with one box empty, which breaks Hund's rule. D counted every valence electron as single, $2 + 4 = 6$, but the 2s pair is matched.

🧠 **Sense check:** Four electrons in three boxes: one pair and two singles. Oxygen has 2 unpaired electrons. ✓



Practice 4

C: atomic number 6 · a student draws the valence shell as 2s: $\uparrow\downarrow$ · 2p: $\uparrow\downarrow$ (two boxes empty)

$1s^2 2s^2 2p^2 \cdot 2 + 2 + 2 = 6$ electrons placed · the drawing shows 0 unpaired

Which rule does that drawing break, and how many unpaired electrons does a carbon atom really have?

- A. Pauli exclusion principle; 2 unpaired electrons
- B. No rule is broken; 0 unpaired electrons
- C. Hund's rule; 4 unpaired electrons
- D. Hund's rule; 2 unpaired electrons

✓ Practice 4 · answer: D

$2p^2$: $\uparrow \uparrow$ (one box empty) ✓ · $\uparrow\downarrow$ (two boxes empty) ✗

Hund's rule: equal-energy boxes take one electron each before any pairs · 2 unpaired (answer D)

The drawing pairs the two 2p electrons while two empty 2p boxes wait. That breaks Hund's rule. Spread them out, one per box with parallel spins, and both sit alone.

A named the wrong rule: Pauli limits a box to two opposite spins, and $\uparrow\downarrow$ obeys it. B accepted the drawing and its count of 0. C counted every valence electron as single, $2 + 2 = 4$, but the 2s pair is matched.

 **Sense check:** A wrong drawing changes the unpaired count. Redraw by the rules first, then count the lone arrows: two. ✓



Worked example 3: vanadium

Step 1 • Count the electrons

V: atomic number 23 • period 4, d block • [Ar] holds 18

$$23 - 18 = 5 \text{ electrons to place after argon}$$

Draw vanadium's orbital diagram past [Ar], then count its unpaired electrons.

✓ Worked example 3: solution

Step 2 • Fill the boxes lowest-energy first



$$18 + 2 + 3 = 23 \checkmark \cdot 4s \text{ fills before } 3d$$

Step 3 • Within a sublevel, singly first



$$2 + 3 = 5 \text{ after } [\text{Ar}] \checkmark \cdot \text{the } 4s \text{ pair adds } 0 \text{ unpaired} \cdot 3 \text{ unpaired, all in } 3d$$

 **Sense check:** Hund's rule runs the same on five d boxes as on three p boxes. Every [Ar] electron and the 4s pair sit paired, so vanadium's 3 unpaired electrons all sit in 3d.



Practice 5

Mn: $Z = 25$ · group 7

How many of manganese's electrons sit unpaired?

- A. 7
- B. 1
- C. 3
- D. 5

✓ Practice 5 · answer: D



$18 + 2 + 5 = 25$ ✓ · five electrons, five 3d boxes: one each · 5 unpaired (answer D)

A counted all $25 - 18 = 7$ electrons past [Ar] as single, the group number too, but the 4s pair is matched. B paired before every box held one: $\uparrow\downarrow \uparrow\downarrow \uparrow$ leaves $5 - 2 \times 2 = 1$. C put all 7 into 3d before 4s: $[\text{Ar}]3d^7$ pairs 2 of them, leaving $5 - 2 = 3$.

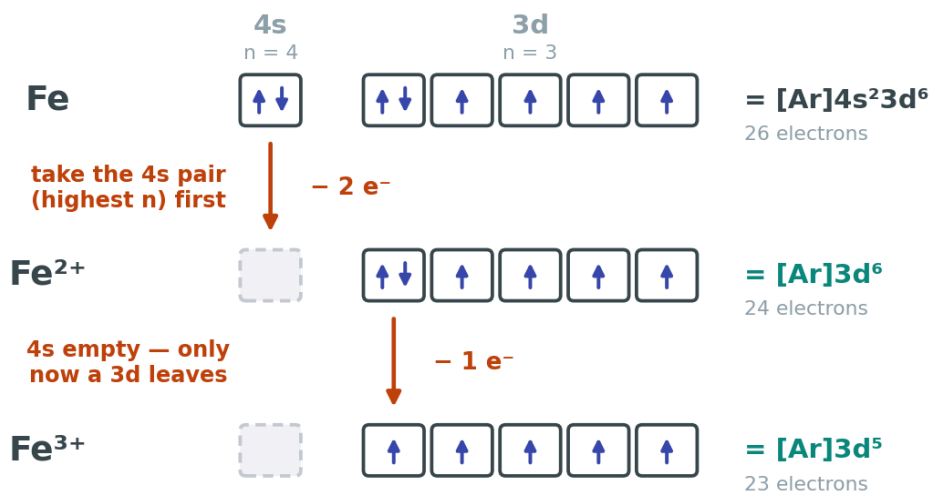
🧠 **Sense check:** 4s takes its pair first, then five 3d boxes take one electron each. The 3d sublevel is half full: 5 unpaired. ✓

✓ Iron's ions, box by box



$18 + 2 + 6 = 26 \rightarrow 26 - 2 = 24 \rightarrow 26 - 3 = 23$ electrons · the 4s pair leaves first

Ionizing iron: the highest-n electrons leave first



4s fills before 3d — and empties before 3d too: always remove the highest n first.

 **Sense check:** Hund's rule holds in ions too. Fe²⁺: one 3d pair, 5 – 1 = 4 unpaired. Fe³⁺: 5 unpaired. ✓

Check yourself

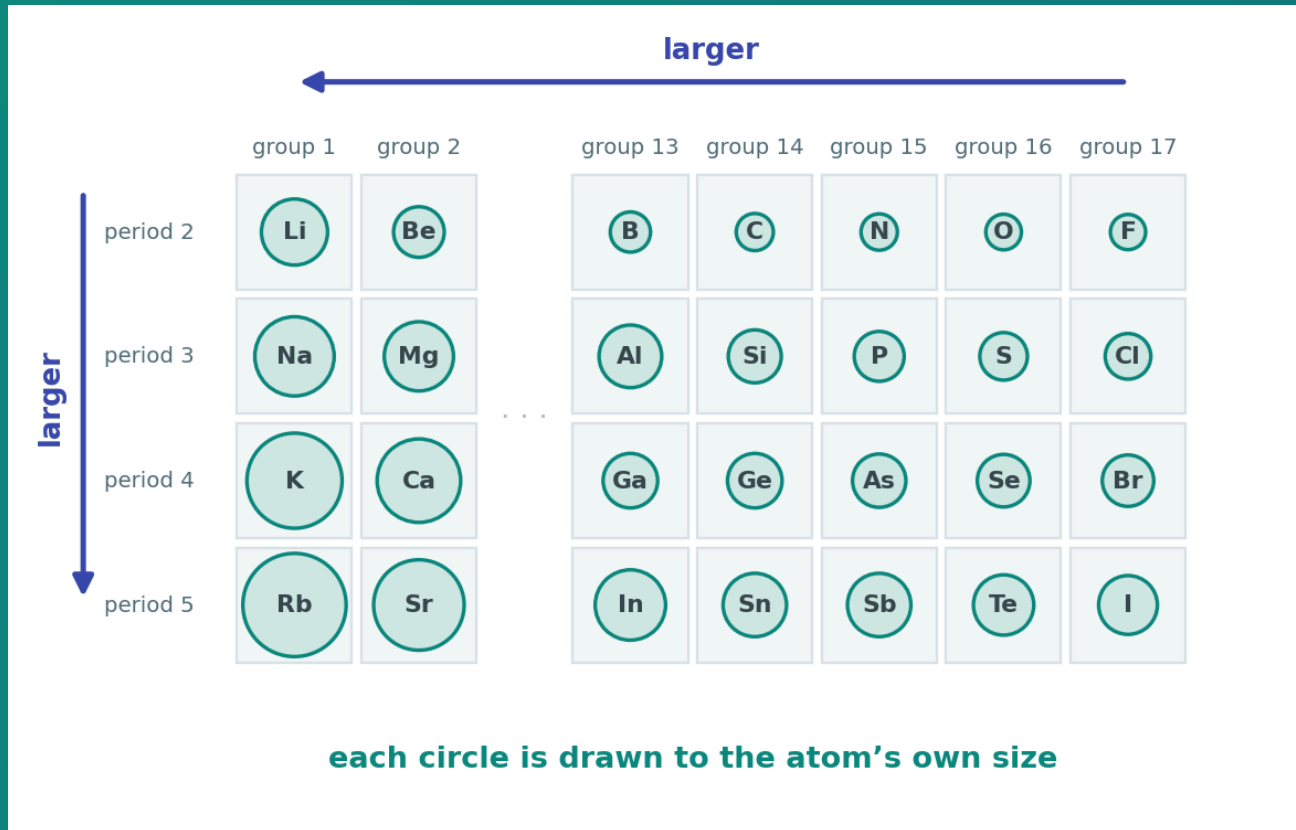
1. Draw the orbital diagram of magnesium, atomic number 12, and count its valence and unpaired electrons.
2. Aluminum sits in group 13. Give its valence count from the group number, then confirm it from the configuration ending $3s^23p^1$.

A configuration compresses the diagram; the diagram shows what the superscripts hide: which electrons sit alone. The valence count is the number to keep. The valence electrons sit outermost, so they are what every neighbor meets first: what an atom holds tight, gives up, or shares is decided by exactly these few.

8 • Periodic Trends

Predict which of two elements has the larger atomic radius, the higher ionization energy, the greater electronegativity, or the more metallic character from where each sits on the periodic table.

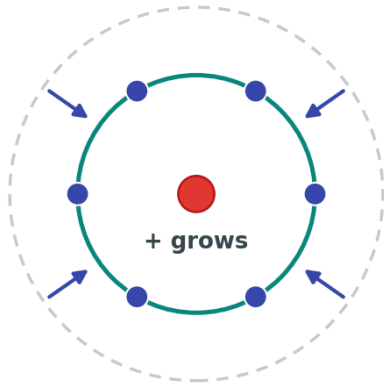
An atom's place predicts its size



Atoms are not all one size. Low and to the left, an atom is large; high and to the right, it is small. Position tells you which.

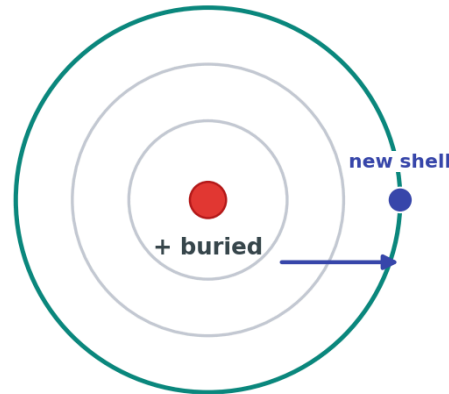
Two forces set every trend

Across a period



more protons pull the same shell in → smaller

Down a group



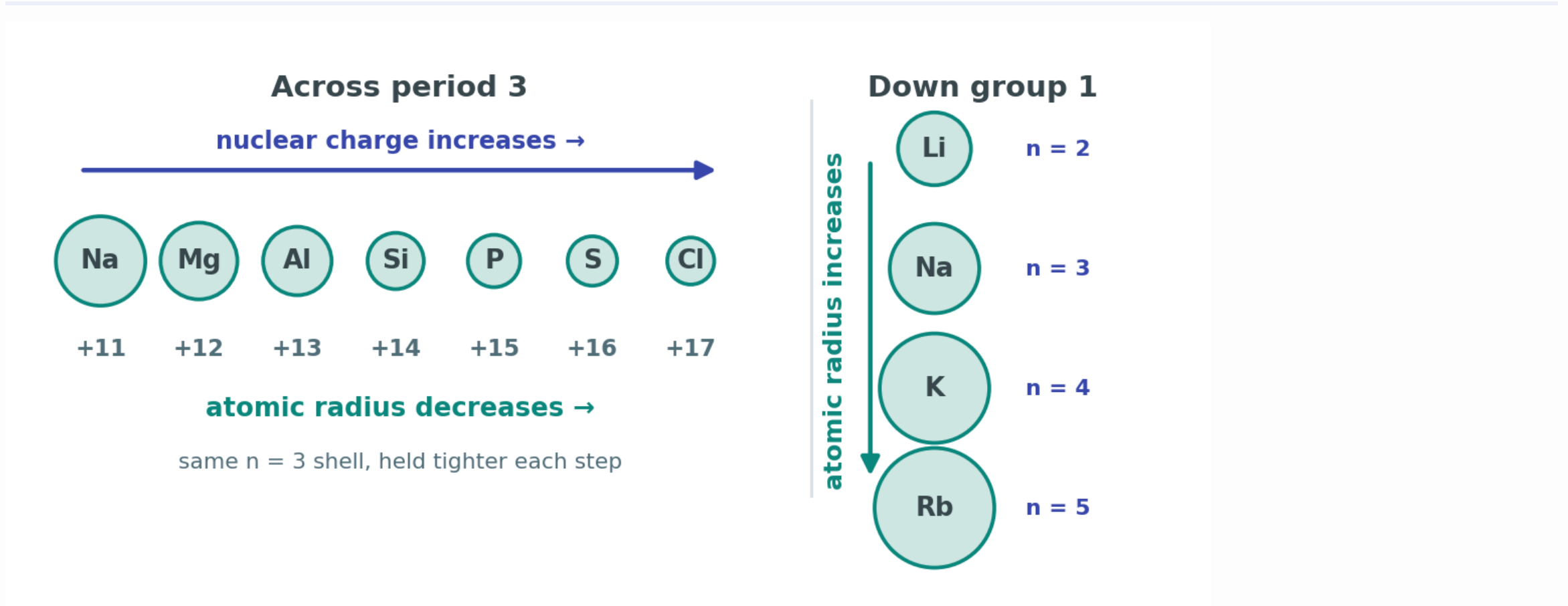
each new shell seats electrons farther out → larger

Two pulls decide every trend. Across a period, each added proton raises Z_{eff} , drawing the same shell inward. Down a group, each new shell adds shielding and distance.

effective nuclear charge (Z_{eff}) · shielding

Z_{eff} : the net pull the outer electrons actually feel · shielding: the inner shells that block part of the nucleus's pull

Atomic radius



Atomic radius is how far the outer electrons sit from the nucleus. It shrinks left to right as the nuclear charge climbs, and grows down a group as each new shell is added.

Ionization energy

Ionization energy is the energy to pull one electron off a gaseous atom. A tighter grip costs more. It rises across a period and falls down a group, mirroring atomic radius.

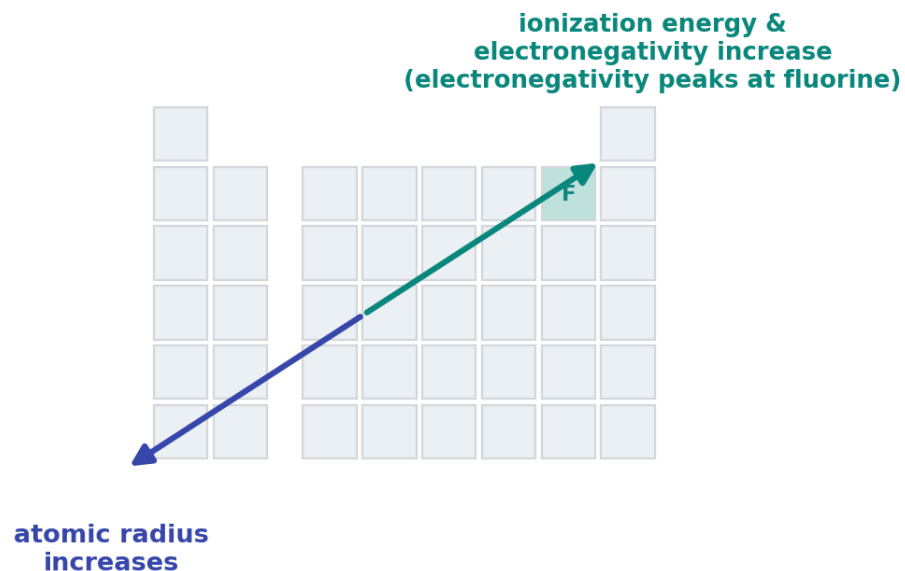
across period 2: Li 520 → C 1086 → F 1681 kJ/mol

nuclear charge climbs, the same shell grips harder: ionization energy increases

down group 1: Li 520 → Na 496 → K 419 kJ/mol

each new shell sits farther out, easier to strip: ionization energy decreases

Electronegativity



Electronegativity is how strongly a bonded atom pulls shared electrons toward itself. It increases across a period and up a group, peaking at fluorine.

memory hook: everything climbs toward fluorine
except atomic radius and metallic character, which run away from it

Metallic character

Metallic character is how readily an atom gives up electrons. It runs opposite to ionization energy: strongest at the lower left of the table, weakest at the upper right.

across period 3: Na (metal) → Si (metalloid) → Cl (nonmetal)
electrons held tighter each step: metallic character decreases across a period

down group 14: C (nonmetal) → Si, Ge (metalloids) → Sn, Pb (metals)
outer electrons sit farther out, given up more easily: metallic character increases down



The method

1. **Place the two elements.** Same period, or same group?
2. **Name the trend** for that direction: radius, ionization energy, electronegativity, or metallic character.
3. **Apply it.** State which element wins, and give the reason from nuclear charge or shells.



Worked example 1: sodium and sulfur

sodium (Na) and sulfur (S): same period 3

Na: +11 nucleus · S: +16 nucleus · both fill through the $n = 3$ shell

Sodium sits at the left of period 3, sulfur well to its right.

Which atom has the larger atomic radius?

✓ Worked example 1: solution

sodium (Na) and sulfur (S): same period 3

Na: +11 nucleus · S: +16 nucleus · both fill through the $n = 3$ shell

Step 1 · Place the two elements

Step 2 · Name the trend

Step 3 · Apply it

Na vs S → sodium is larger (≈ 186 pm vs ≈ 104 pm)

same $n = 3$ shell · Na's +11 nucleus grips it more loosely than S's +16

 **Sense check:** Sodium's shell feels the weaker pull, so it holds the same electrons farther out. Across a period, the growing nuclear charge wins.

✓ Worked example 1: the path on the table

sodium (Na) and sulfur (S): same period 3
found: sodium is larger (≈ 186 pm vs ≈ 104 pm)

atomic radius (pm)

period	1	2	group	13	14	15	16	17	18
2	Li	Be		B	C	N	O	F	Ne
3	Na 186	Mg		Al	Si	P	S 104	Cl	Ar
4	K	Ca		Ga	Ge	As	Se	Br	Kr

1 Na → S across period 3
radius decreases

Na > S

🧠 **Sense check:** One leg along one row. A single trend decides the pair: radius shrinks to the right. ✓



Worked example 2: lithium and potassium

lithium (Li) and potassium (K): same group 1

Li: outer electron in $n = 2$ · K: outer electron in $n = 4$

Lithium and potassium are both alkali metals, potassium two rows below.

A common first attempt: potassium holds far more protons, so its outer electron should be the hardest to remove. Which atom has the higher ionization energy?

✓ Worked example 2: solution

lithium (Li) and potassium (K): same group 1

Li: outer electron in $n = 2$ · K: outer electron in $n = 4$

A common first attempt

Step 1 · Place the two elements

Step 2 · Name the trend

Step 3 · Apply it

Li vs K → lithium is higher (520 vs 419 kJ/mol)

Li's $n = 2$ electron sits close and poorly shielded, costing the most to remove

🧠 **Sense check:** Potassium is the bigger atom, yet its electron leaves more easily. Down a group, distance beats a larger nuclear charge.

✓ Worked example 2: the path on the table

lithium (Li) and potassium (K): same group 1
found: lithium is higher (520 vs 419 kJ/mol)

ionization energy (kJ/mol)

period	1	2	group	13	14	15	16	17	18
2	Li 520	Be		B	C	N	O	F	Ne
3	Na	Mg		Al	Si	P	S	Cl	Ar
4	K 419	Ca		Ga	Ge	As	Se	Br	Kr

A blue arrow labeled '1' points from the Li cell to the K cell, indicating the path from period 2 to period 4 in group 1.

1 Li → K down group 1
ionization energy falls

Li > K

🧠 **Sense check:** One leg down one column. Each row down adds a shell, so ionization energy falls. ✓

Your turn: aluminum and sulfur

aluminum (Al) and sulfur (S): same period 3

Al: +13 nucleus · S: +16 nucleus · both fill through the $n = 3$ shell

step	question	answer
1 · place them	same period, or same group?	same _____
2 · name the trend	which way does radius run?	radius _____ across a period
3 · apply it	which atom is larger?	_____

Complete the three steps.

Al vs S → aluminum is larger

same $n = 3$ shell · radius decreases left to right · Al's +13 grips it less tightly than S's +16

⚠ Where this goes wrong

Reading the trend backwards. Radius does not grow across a period. Left to right the nuclear charge climbs while the shell stays the same, so the atoms shrink. Sodium (+11) is larger than chlorine (+17), not smaller.

Mixing up across and down. Across a period, a new proton pulls the same shell tighter. Down a group, a whole new shell is added. Same period → weigh nuclear charge; same group → count the shells.

Letting more protons mean a bigger atom. Extra protons pull electrons in, never push them out. Down a group an atom grows in spite of its larger charge, because each row opens a new, higher shell.

Treating radius and ionization energy as one trend. They run opposite. The bigger atom holds its outer electron more loosely, so a large radius comes with a low ionization energy.

Practice 1: oxygen or fluorine

oxygen (O) and fluorine (F)
wanted: the larger atomic radius

Which atom has the larger atomic radius, and why?

- A. Fluorine: its extra proton and electron take up more room.
- B. Fluorine: every periodic trend climbs toward fluorine.
- C. Oxygen: same $n = 2$ shell, and its $+8$ nucleus pulls it in less than fluorine's $+9$.
- D. Oxygen: its outer electrons sit in a higher shell than fluorine's.

✓ Practice 1 · answer: C

O vs F across period 2 → oxygen is larger (answer C)

same $n = 2$ shell · O's +8 nucleus pulls it in less than F's +9

A let more protons mean a bigger atom: fluorine's extra proton pulls the same shell in tighter. B carried the fluorine memory hook over to radius, the one trend that runs away from fluorine. D had the right atom for the wrong reason: both use the $n = 2$ shell, so the difference is nuclear charge, not shell height.

 **Sense check:** Across a period, the same shell feels a stronger pull at each step. Oxygen, one step left of fluorine, is the larger atom. ✓



Practice 2: magnesium or calcium

magnesium (Mg) and calcium (Ca)

wanted: the higher ionization energy

Which atom has the higher ionization energy, and why?

- A. Magnesium: its outer electrons sit in $n = 3$, closer to the nucleus than calcium's.
- B. Calcium: its +20 nucleus grips the outer electrons harder than magnesium's +12.
- C. Calcium: it is the larger atom, and a larger atom holds its electrons more tightly.
- D. Magnesium: its +12 nucleus holds fewer electrons, so each one is gripped harder.

✓ Practice 2 · answer: A

Mg vs Ca down group 2 → magnesium is higher (answer A)

738 vs 590 kJ/mol · Mg's outer electrons in $n = 3$, Ca's in $n = 4$

B let raw charge win: calcium's +20 nucleus sits under an extra shell, so its outer electrons are farther out and better shielded. C tied size to grip the wrong way: the larger atom holds its outer electrons more loosely. D had the right atom for the wrong reason: each neutral atom holds as many electrons as its nucleus has protons, so the count is no advantage. The difference is the shell: $n = 3$ against $n = 4$.

 **Sense check:** Down group 2, each new shell puts the outer electrons farther out. Magnesium's cost more to remove. ✓

Practice 3: boron or oxygen

boron (B) and oxygen (O): same period 2

B: +5 nucleus · O: +8 nucleus · both bond through the $n = 2$ shell

Boron and oxygen lie in the same period. Which atom is more electronegative, and why?

- A. Boron: electronegativity falls off across a period, so the element on the left attracts shared electrons more strongly.
- B. Boron: it is the larger atom, and a larger atom pulls a bonding pair in harder.
- C. Oxygen: electronegativity increases across a period, and oxygen sits farther right, closer to fluorine.
- D. Oxygen: it has more occupied shells than boron, so it reaches shared electrons better.

✓ Practice 3 · answer: C

B vs O across period 2 → oxygen (answer C)

B 2.04 · O 3.44 (Pauling) · both bond through $n = 2$, O's +8 nucleus outpulls B's +5

A ran the trend backwards: electronegativity increases, not decreases, across a period, so the right-hand atom wins. B confused the size trend: boron is the larger atom, but a larger atom pulls a shared pair less, not more. D reached for extra shells that are not there: both atoms bond through the $n = 2$ shell, and oxygen wins by nuclear charge, not shell count.

 **Sense check:** Across a period, electronegativity climbs toward fluorine. Oxygen, one step from it, outpulls boron.

✓

Worked example 3: ranking three atoms

calcium (Ca), magnesium (Mg), chlorine (Cl)

Mg & Ca: group 2 · Mg & Cl: period 3 · magnesium is the shared corner

No single row or column holds all three. Rank them by atomic radius, largest first.

✓ Worked example 3: two comparisons

calcium (Ca), magnesium (Mg), chlorine (Cl)

Mg & Ca: group 2 · Mg & Cl: period 3 · magnesium is the shared corner

Step 1 · Place the two elements

Step 2 · Name the trend

Ca vs Mg (group 2): radius increases downward → Ca is larger

≈ 197 pm vs ≈ 160 pm · calcium adds a shell ($n = 4$ vs $n = 3$)

 **Sense check:** Down group 2, calcium sits one row below magnesium, one shell farther out, so it is the larger of the two.

✓ Worked example 3: the ranking

calcium (Ca), magnesium (Mg), chlorine (Cl)
from Step 2: calcium is larger than magnesium

Step 3 • Apply it

Mg vs Cl (period 3): radius decreases to the right → Mg is larger
 $\approx 160 \text{ pm}$ vs $\approx 99 \text{ pm}$ · same $n = 3$ shell, Cl's +17 pulls harder

largest to smallest: calcium, magnesium, chlorine
 $\approx 197 \text{ pm}$, 160 pm , 99 pm : calcium the largest, chlorine the smallest

🧠 **Sense check:** Two clean comparisons through the shared corner rank all three, with no diagonal guesswork.

✓ Worked example 3: the path through the corner

calcium (Ca), magnesium (Mg), chlorine (Cl)

found: calcium, magnesium, chlorine ($\approx 197, 160, 99$ pm)

atomic radius (pm)

period	1	2	group	13	14	15	16	17	18
2	Li	Be		B	C	N	O	F	Ne
3	Na	Mg 160		Al	Si	P	S	Cl 99	Ar
4	K	Ca 197		Ga	Ge	As	Se	Br	Kr

□ **corner: Mg**

group 2 with Ca
period 3 with Cl

- 1 Ca → Mg up group 2
radius decreases
- 2 Mg → Cl across period 3
radius decreases

Ca > Mg > Cl

🧠 **Sense check:** Two legs through the corner, magnesium. Each leg is one single-trend comparison. ✓

Practice 4: sulfur or fluorine

sulfur (S) and fluorine (F)
wanted: the higher ionization energy

Which atom needs more energy to remove one electron, and why?

- A. Sulfur: its +16 nucleus outpulls fluorine's +9.
- B. Fluorine: it lies up and to the right of sulfur, and both moves raise ionization energy.
- C. Fluorine: it has fewer electrons, so each is easier to hold.
- D. Sulfur: ionization energy increases down a group, and sulfur sits one row lower.

✓ Practice 4 · answer: B

S vs F through the corner Cl → fluorine is higher (answer B)
S 1000 < Cl 1251 across period 3 · Cl 1251 < F 1681 up group 17 (kJ/mol)

A let raw charge win: sulfur's +16 nucleus sits under one more shell than fluorine's +9, so it holds its outer electrons more loosely. C had the right atom for the wrong reason: the electron count is not the cause. Fluorine's outer electrons sit in $n = 2$, closer in, and across a period the pull climbs. D ran the group trend backwards: moving down a group lowers the ionization energy.

 **Sense check:** Up and to the right, both trends point the same way. The other corner, oxygen, gives the same verdict. ✓

✓ Practice 4: the path through the corner

sulfur (S) and fluorine (F)

found: fluorine is higher (1681 vs 1000 kJ/mol)

ionization energy (kJ/mol)

period	1	2	group	13	14	15	16	17	18
2	Li	Be		B	C	N	O	F 1681	Ne
3	Na	Mg		Al	Si	P	S 1000	Cl 1251	Ar
4	K	Ca		Ga	Ge	As	Se	Br	Kr

Diagram illustrating the path from S to F through the corner element Cl. Arrows indicate the path: S (1000 kJ/mol) → Cl (1251 kJ/mol) → F (1681 kJ/mol). The path is labeled with 1 and 2.

□ **corner: Cl**
period 3 with S
group 17 with F

- 1 S → Cl across period 3
ionization energy rises
- 2 Cl → F up group 17
ionization energy rises

F > Cl > S

🧠 **Sense check:** Chlorine is the corner: one leg across period 3, one leg up group 17. Both legs raise the ionization energy. ✓



Practice 5: three atoms from their configurations

atom X: $[\text{Ar}]4s^1$ · atom Y: $[\text{Ne}]3s^1$ · atom Z: $[\text{Ne}]3s^23p^3$

given: three ground-state configurations · wanted: atomic radius, largest first

Which ranking by atomic radius runs largest to smallest?

- A. $Z > Y > X$
- B. $Y > X > Z$
- C. $X > Y > Z$
- D. $X > Z > Y$

✓ Practice 5 · answer: C

$X = \text{K (period 4, group 1)} \cdot Y = \text{Na (period 3, group 1)} \cdot Z = \text{P (period 3, group 15)}$

K vs Na, group 1: K adds a shell, larger · Na vs P, period 3: P's +15 grips the same shell tighter, smaller

largest to smallest: $\text{K} > \text{Na} > \text{P}$ (answer C)

$\approx 227 \text{ pm}, 186 \text{ pm}, 110 \text{ pm}$ · sodium is the corner both comparisons pass through

A reversed both trends: down group 1 the new $n = 4$ shell makes K the larger of K and Na, and across period 3 P's +15 draws the same shell in, so P is the smallest; $Z > Y > X$ is the correct list read backwards. B ran the group trend backwards: K sits a row below Na and opens a new shell, so K is the larger. D let Z's stronger nucleus push electrons out: a stronger pull draws the same shell inward, so P is the smallest.

 **Sense check:** Decode each configuration to its square on the table, then compare through the shared corner: down adds a shell, right tightens the grip. ✓



Check yourself

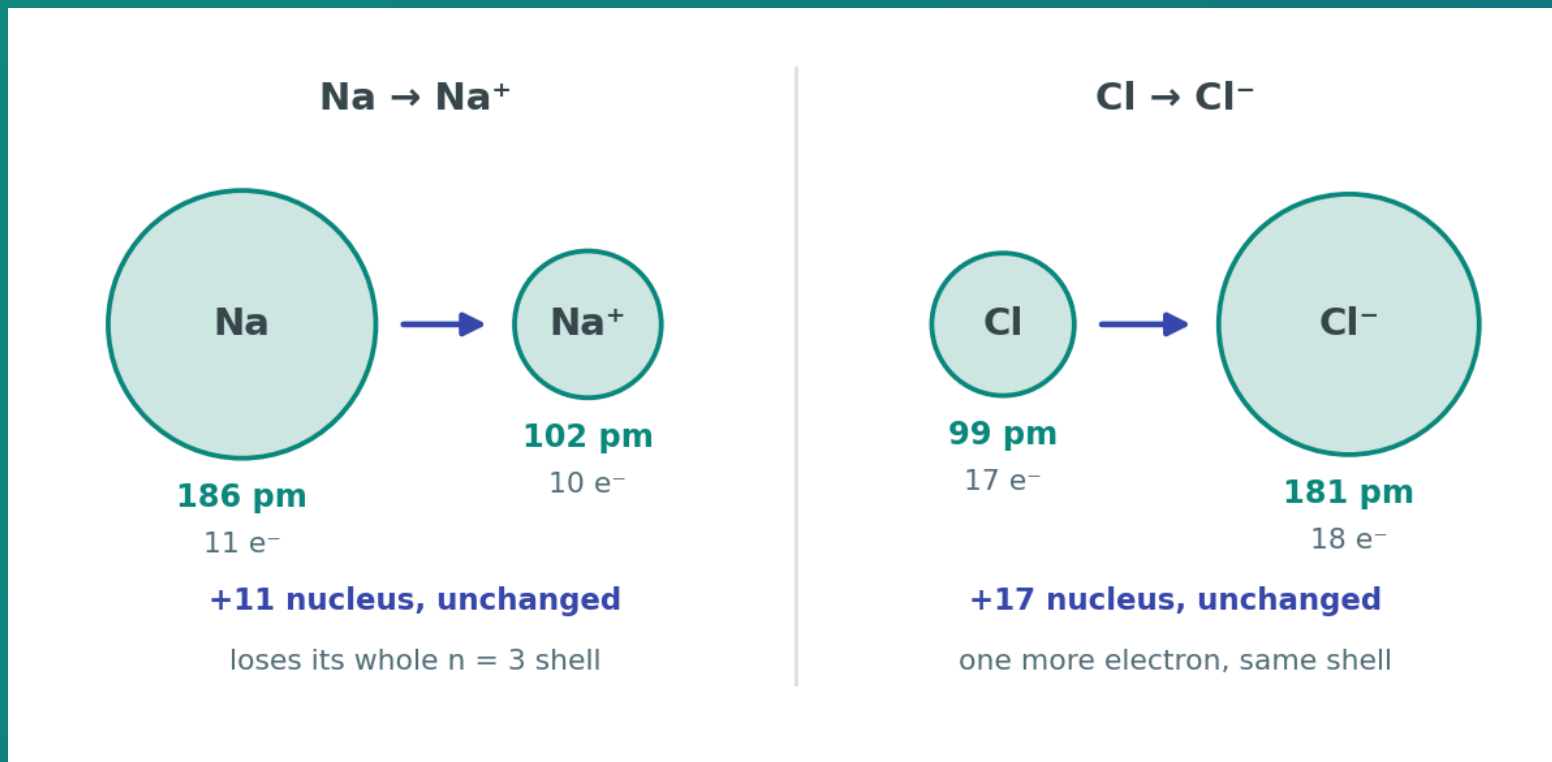
1. Chlorine and iodine sit in the same group. Which holds its electrons more tightly, and does that make it the larger or the smaller atom?
2. Across period 3, phosphorus lies left of chlorine. Rank their atomic radius and their ionization energy. Do the two rankings point the same way?

Electronegativity, the pull an atom keeps on shared electrons, is the trend that carries into bonding. When two bonded atoms differ in it, the shared pair sits closer to one, and the bond turns polar.

9 • Ion Size & Isoelectronic Series

Predict how ionization changes size, a cation smaller and an anion larger than its atom, and rank an isoelectronic series by nuclear charge.

Ionization changes an atom's size



Sodium metal and chlorine gas react to make table salt. In the crystal, each sodium is smaller than its atom, each chlorine larger. Losing or gaining electrons resized both.

What sets the size of an ion

Ionization never touches the nucleus: protons stay fixed while electrons are lost or gained. Size follows grip per electron. The same pull holds fewer electrons tighter, more electrons looser.

size follows pull per electron: protons fixed, electrons change
lose electrons → each one held tighter · gain electrons → the same pull spread thinner

A cation is smaller than its atom

Sodium's one $n = 3$ electron leaves, so the entire outer shell disappears. Ten electrons remain under the same $+11$ pull, each held tighter. The ion is close to half the atom's radius.

Na (186 pm) $\rightarrow$ Na⁺ (102 pm)

11 e⁻ $\rightarrow$ 10 e⁻ · +11 nucleus unchanged · outermost shell now $n = 2$

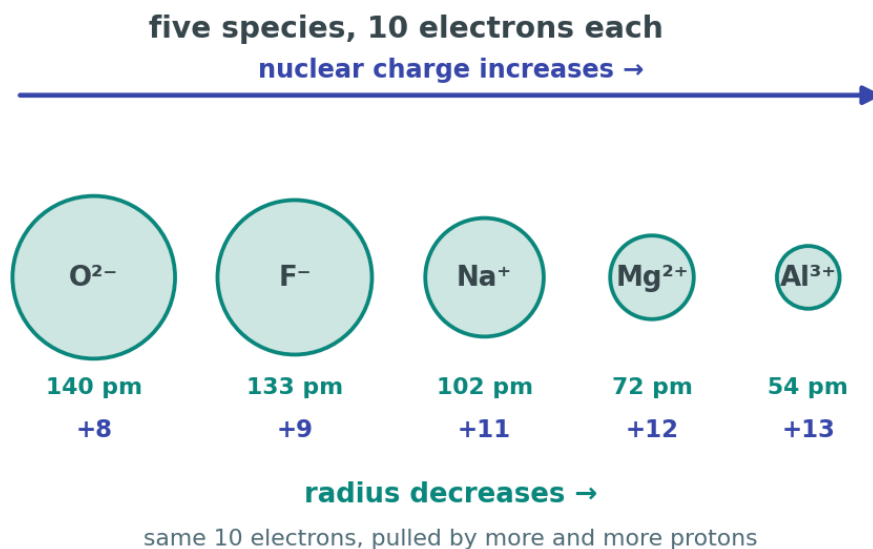
An anion is larger than its atom

Chlorine gains one electron into its $n = 3$ shell. The nucleus still pulls with $+17$, now spread over 18 electrons that repel each other more. The cloud swells.



$17 \text{ e}^- \rightarrow 18 \text{ e}^- \cdot +17 \text{ nucleus unchanged} \cdot \text{more repulsion in the same shell}$

Isoelectronic species: nuclear charge decides



O^{2-} , F^{-} , Na^{+} , Mg^{2+} , and Al^{3+} each hold 10 electrons: they are isoelectronic. Only the nuclear charge differs, so the strongest pull makes the smallest species.

memory hook: the size order is the proton count read backwards
+8 O^{2-} largest ... +13 Al^{3+} smallest: more protons, tighter grip on the same 10 electrons



The method

1. **Count the electrons.** Protons do not change; electrons are lost or gained.
2. **Compare pull to electrons.** Same element: fewer electrons means smaller. Same electron count: more protons means smaller.
3. **State the order,** largest first, with the reason.



Worked example 1: magnesium and its ion

magnesium (Mg) and its ion Mg^{2+}
 Mg : 12 p^+ , 12 e^- · Mg^{2+} : 12 p^+ , $12 - 2 = 10 \text{ e}^-$

Magnesium forms the $2+$ cation in compounds such as MgO .

Which is larger, Mg or Mg^{2+} ?

✓ Worked example 1: solution

magnesium (Mg) and its ion Mg^{2+}

Mg: 12 p^+ , 12 e^- · Mg^{2+} : 12 p^+ , $12 - 2 = 10 \text{ e}^-$

Step 1 · Count the electrons

Step 2 · Compare pull to electrons

Step 3 · State the order

$\text{Mg} (160 \text{ pm}) > \text{Mg}^{2+} (72 \text{ pm})$

the atom is larger · losing the $n = 3$ shell cut the radius by more than half

 **Sense check:** A 2+ cation keeps every proton and loses its whole outer shell. Cations are always smaller than their parent atoms.



Worked example 2: ranking an isoelectronic series



electrons: $8 + 2 \cdot 9 + 1 \cdot 11 - 1 \cdot 12 - 2 \cdot 13 - 3 = 10$ each

All five species hold the same 10 electrons.

Rank them by radius, largest first.

✓ Worked example 2: solution

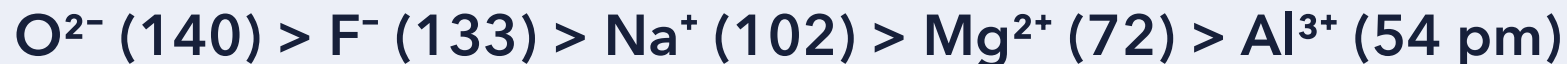


electrons: $8 + 2 \cdot 9 + 1 \cdot 11 - 1 \cdot 12 - 2 \cdot 13 - 3 = 10$ each

Step 1 · Count the electrons

Step 2 · Compare pull to electrons

Step 3 · State the order



radius runs opposite to nuclear charge: +8 largest, +13 smallest

 **Sense check:** The ranking is the proton count read backwards. Al^{3+} grips 10 electrons with 13 protons and ends up well under half of O^{2-} 's size.

Your turn: sulfur and sulfide

sulfur (S) and its ion S^{2-}

S: 16 p^+ , 16 e^- · S^{2-} : 16 p^+ , 16 + 2 = 18 e^-

step	question	answer
1 · count the electrons	how many electrons in S^{2-} ?	_____
2 · compare pull to electrons	same +16 pull, more electrons: tighter or looser?	_____
3 · state the order	which is larger?	_____

Complete the three steps.

S^{2-} (184 pm) > S (104 pm)

18 e^- share the +16 pull · more repulsion, less grip per electron: the anion is larger

⚠ Where this goes wrong

Changing the proton count. Ionization moves electrons only. Na^+ still holds 11 protons; the 1+ charge is the mismatch of 11 p^+ against 10 e^- , not a changed nucleus.

Expecting a small shrink. Na^+ is not slightly smaller than Na. At 102 pm against 186 pm it is close to half the size, because the whole $n = 3$ shell is gone.

Shrinking the anion. A gained electron adds repulsion under the same pull, so Cl^- (181 pm) is far larger than Cl (99 pm), never smaller.

Calling isoelectronic ions equal in size. Same electron count is not same size. O^{2-} and Al^{3+} both hold 10 electrons, yet span 140 pm to 54 pm. Nuclear charge decides.



Practice 1



electrons: $16 + 2 \cdot 17 + 1 \cdot 19 - 1 \cdot 20 - 2 = 18$ each · nuclei: +16, +17, +19, +20

All four species hold 18 electrons. Which order runs largest to smallest?

- A. $\text{S}^{2-} > \text{Cl}^- > \text{K}^+ > \text{Ca}^{2+}$
- B. $\text{Ca}^{2+} > \text{K}^+ > \text{Cl}^- > \text{S}^{2-}$
- C. $\text{K}^+ > \text{Ca}^{2+} > \text{S}^{2-} > \text{Cl}^-$
- D. They are equal: species with identical electron counts have identical radii.

✓ Practice 1 · answer: A

S^{2-} (184) > Cl^{-} (181) > K^{+} (138) > Ca^{2+} (100 pm) (answer A)

18 e^{-} each · the +16 nucleus grips loosest, the +20 nucleus tightest

B ran the pull backwards: more protons draw the same 18 electrons inward, so +20 gives the smallest ion, not the largest. C ranked the parent atoms (K 227 > Ca 197 > S 104 > Cl 99 pm): the atoms' sizes are irrelevant once all four species hold the same 18 electrons. D repeated the equal-size trap: isoelectronic means equal electron count, and the differing nuclear charges still set different radii.



Sense check: One rule ranks any isoelectronic set: radius runs opposite to nuclear charge. ✓

 Practice 2

electrons: $35 + 1 \cdot 37 - 1 \cdot 38 - 2 = 36$ each $\cdot \text{I}^-: 53 + 1 = 54$ $\cdot$ nuclei: $+35, +37, +38, +53$

Which of the four ions has the largest radius?

- A. Sr^{2+}
- B. Rb^+
- C. Br^-
- D. I^-

✓ Practice 2 · answer: D

$\text{I}^- (220) > \text{Br}^- (196) > \text{Rb}^+ (152) > \text{Sr}^{2+} (118 \text{ pm})$ (answer D)

among the 36-electron trio, +35 grips loosest: $\text{Br}^- \cdot \text{I}^-$ holds 54 electrons through the $n = 5$ shell, one shell beyond

A read the highest charge as the largest ion: Sr^{2+} has the most protons of the 36-electron trio, +38, so it grips its 36 electrons tightest and is the smallest (118 pm). B carried over the parent atom: rubidium is a large atom, but Rb^+ lost its whole $n = 5$ shell. C ranked the isoelectronic trio and stopped: Br^- beats Rb^+ and Sr^{2+} , but I^- sits a row lower with an extra shell, and an anion only adds to that size.

 **Sense check:** Two rules in order: same electron count, fewer protons wins; then a species with an extra shell is larger still. ✓



Check yourself

1. Rubidium forms Rb^+ . Which is larger, and what happened to the outermost shell?
2. N^{3-} , O^{2-} , and F^- each hold 10 electrons. Rank them largest to smallest and name the property that decides.

Ionic radii carry into lattice energy. The distance between ion centers is the sum of the two ionic radii, and smaller ions sit closer, making a tighter, more stable ionic solid.

Can you...?

- ☐ relate a light wave's wavelength, frequency, and photon energy with $c = \lambda\nu$ and $E = h\nu$, and rank colors of light by each?
- ☐ explain why excited atoms emit line spectra rather than continuous ones, and compare Bohr transitions by photon energy and wavelength?
- ☐ describe the shapes of s, p, and d orbitals and count the orbitals and electrons a sublevel or shell holds?
- ☐ write full and noble-gas electron configurations for atoms through the d block and for their common ions?
- ☐ draw orbital diagrams with the building-up order, the Pauli principle, and Hund's rule, and count valence electrons from the configuration or the group number?
- ☐ predict atomic radius, ionization energy, and ion size from position on the periodic table, including an isoelectronic series?

If any box stays empty, the practice site has a drill for it. 🧪