

Light & Electronic Structure

CHEM 50 · Dr. Karmach

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

- Relate a light wave's wavelength, frequency, and energy using $c = \lambda\nu$ and $E = h\nu$
- Calculate the energy of a photon from its wavelength or frequency
- State how many electrons each sublevel (s, p, d, f) holds and why
- Write the ground-state electron configuration of an atom using the building-up order
- Predict periodic trends in atomic radius, ionization energy, electronegativity, metallic character, and electron affinity from an element's position, including the exceptions to the trends
- Explain the effect of ionization on size: a cation is smaller and an anion larger than its atom, and an isoelectronic series ranks by nuclear charge

Today's route

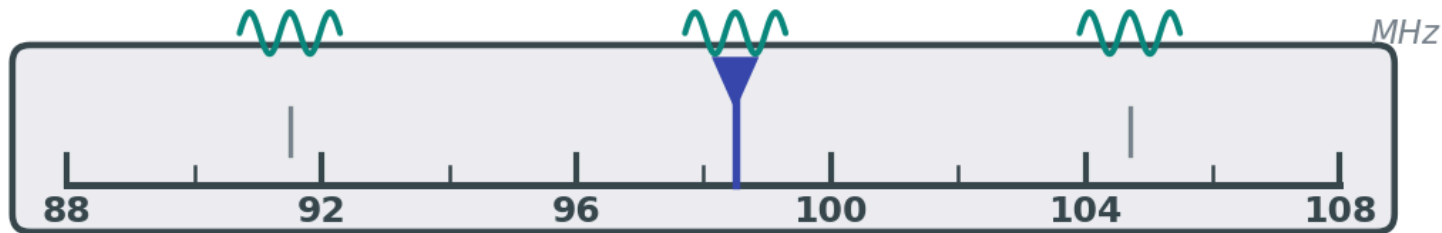
1. Wavelength & Frequency
2. Photon Energy
3. The Bohr Model & Line Spectra
4. Quantum Numbers
5. Orbital Shapes & Nodes
6. Sublevels & Orbitals
7. Electron Configurations
8. Orbital Diagrams & Valence Electrons
9. Electron Configurations of Ions
10. Exceptions: Chromium & Copper
11. Periodic Trends
12. Ion Size & Isoelectronic Series
13. Paramagnetism & Diamagnetism

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

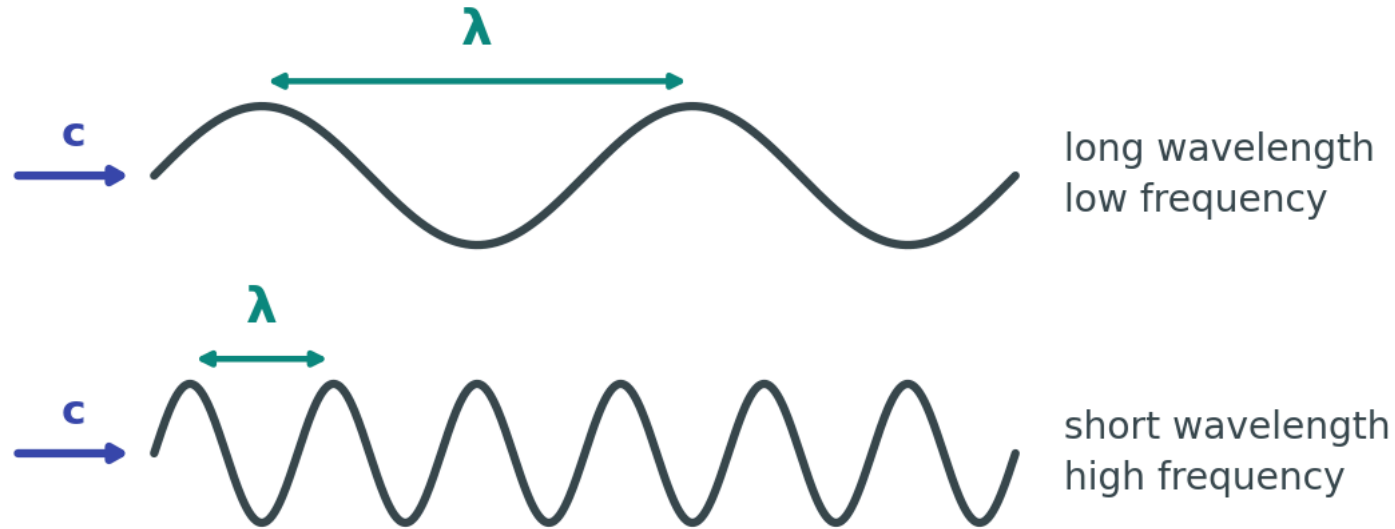


every station is the same kind of wave — only the number changes

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.

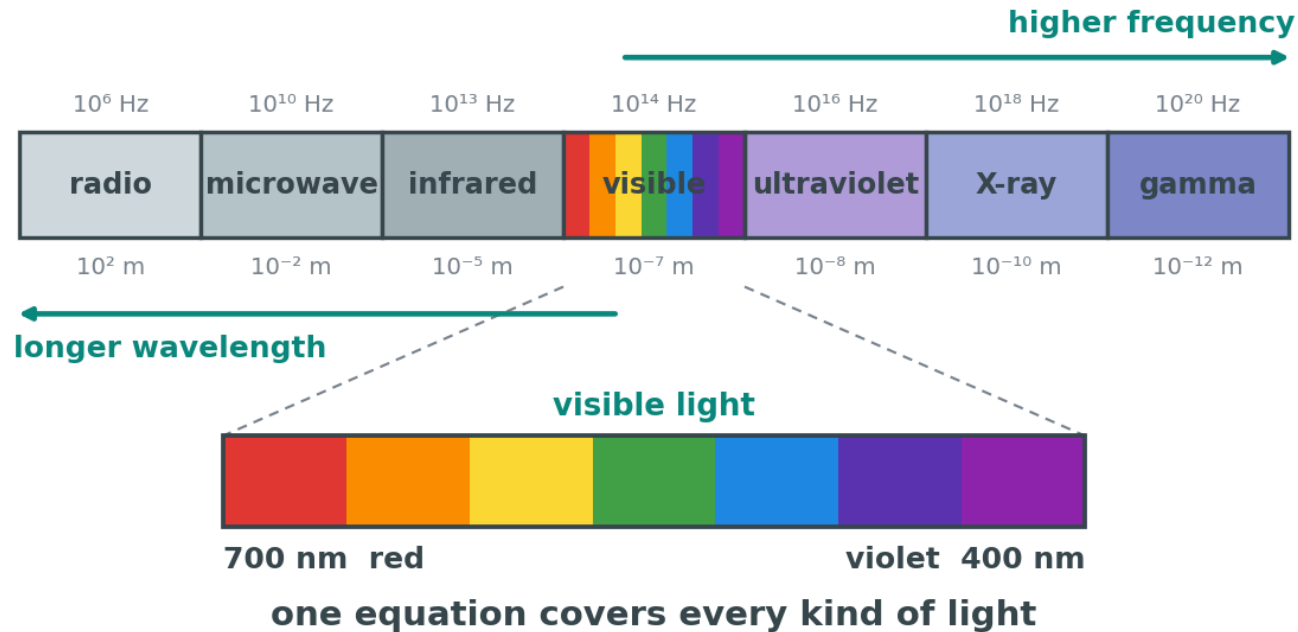
The equation: $c = \lambda \nu$

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

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

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

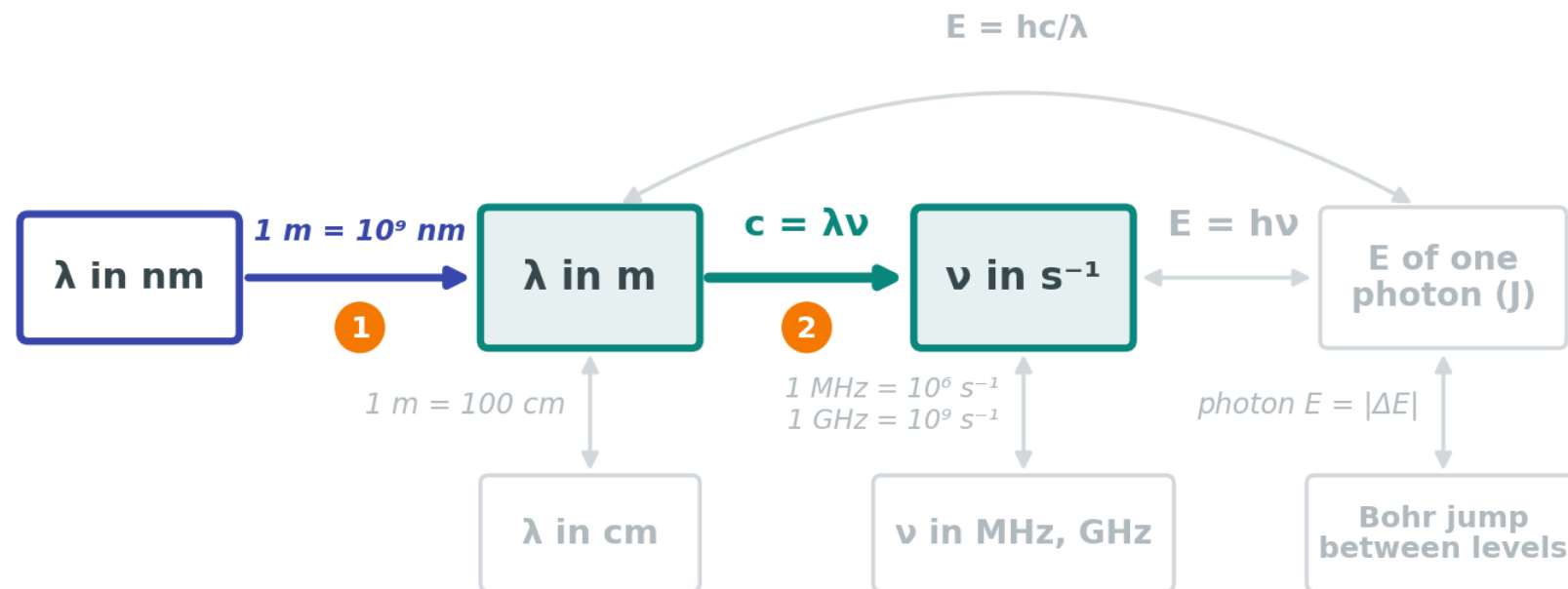
$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{6.00 \times 10^{-7} \text{ m}} = 5.00 \times 10^{14} \text{ s}^{-1}$$

🧠 **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

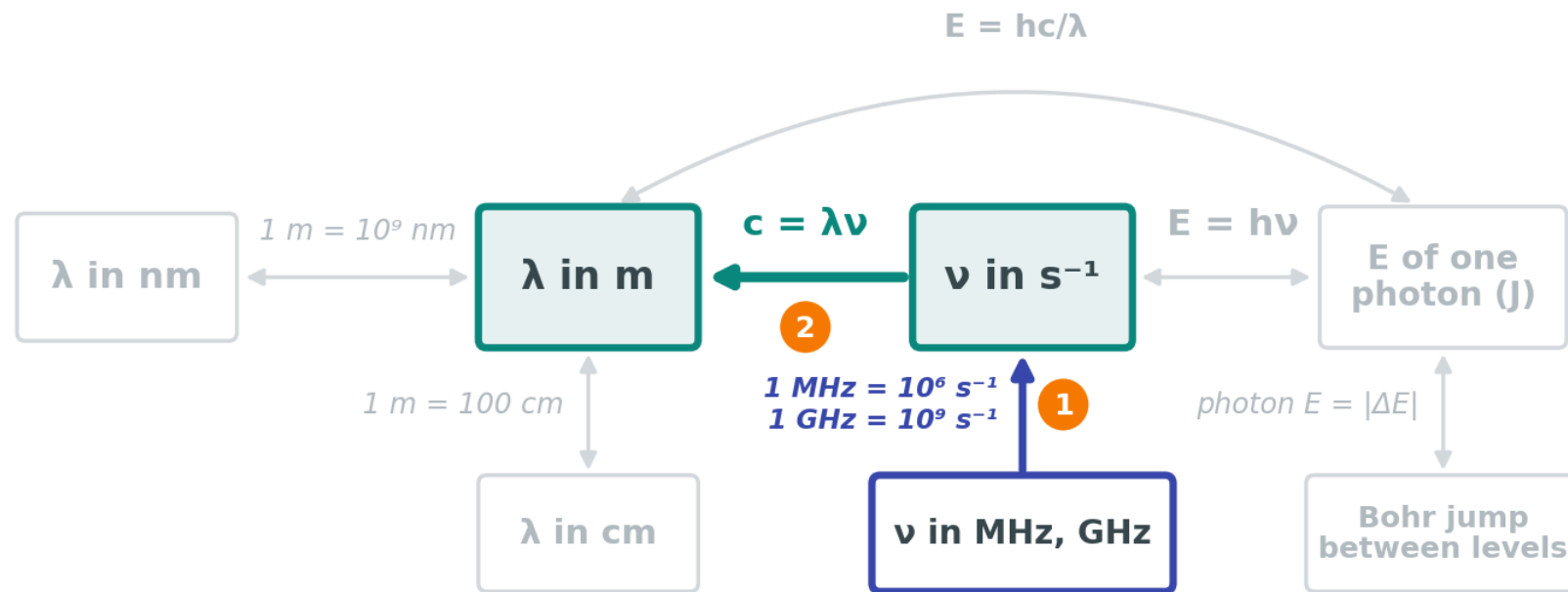
$$\lambda = \frac{3.00 \times 10^8 \text{ m} \cdot \cancel{\text{s}^{-1}}}{1.00 \times 10^8 \cancel{\text{s}^{-1}}} = 3.00 \text{ m}$$

 **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 · $c = 3.00 \times 10^8$ m/s · correct $\nu = 5.00 \times 10^{14}$ s⁻¹

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

Multiplying c by the wavelength. $3.00 \times 10^8 \times 6.00 \times 10^{-7} = 180$. Its units are m²/s, not a frequency. Rearrange $c = \lambda\nu$ to $\nu = c/\lambda$, with λ underneath.

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



Practice 1

$$c = \lambda \nu$$

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 1 · 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. ✓



Worked example 3: frequency of ultraviolet light

$$c = \lambda \nu$$

given: $\lambda = 250 \text{ nm}$ (ultraviolet) · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

A sterilizing lamp emits **ultraviolet light** at **250 nm**. Find its frequency. ($c = 3.00 \times 10^8 \text{ m/s}$; $1 \text{ m} = 10^9 \text{ nm}$)

A common first attempt: divide c by 250 with no unit change. Test the result.

✓ Worked example 3: solution

$$c = \lambda \nu$$

given: $\lambda = 250 \text{ nm}$ (ultraviolet) · $c = 3.00 \times 10^8 \text{ m/s}$ · wanted: ν

A common first attempt

$$\nu = \frac{3.00 \times 10^8 \text{ m/s}}{250 \text{ m}} = 1.20 \times 10^6 \text{ s}^{-1} \times$$

Step 1 · Identify the unknown

Step 2 · Match units to c

c is meters per second, so λ must be in meters: $250 \text{ nm} \times (1 \text{ m} / 10^9 \text{ nm}) = 2.50 \times 10^{-7} \text{ m}$.

 **Sense check:** In meters the wavelength reads 2.50×10^{-7} , smaller than the bare 250 by 10^9 . That missing factor is the whole error. ✓

✓ Worked example 3: frequency

$$v = c/\lambda$$

$$\lambda = 2.50 \times 10^{-7} \text{ m (250 nm)} \cdot c = 3.00 \times 10^8 \text{ m/s} \cdot \text{wanted: } v$$

Step 3 • Rearrange for the unknown

Step 4 • Substitute and cancel units

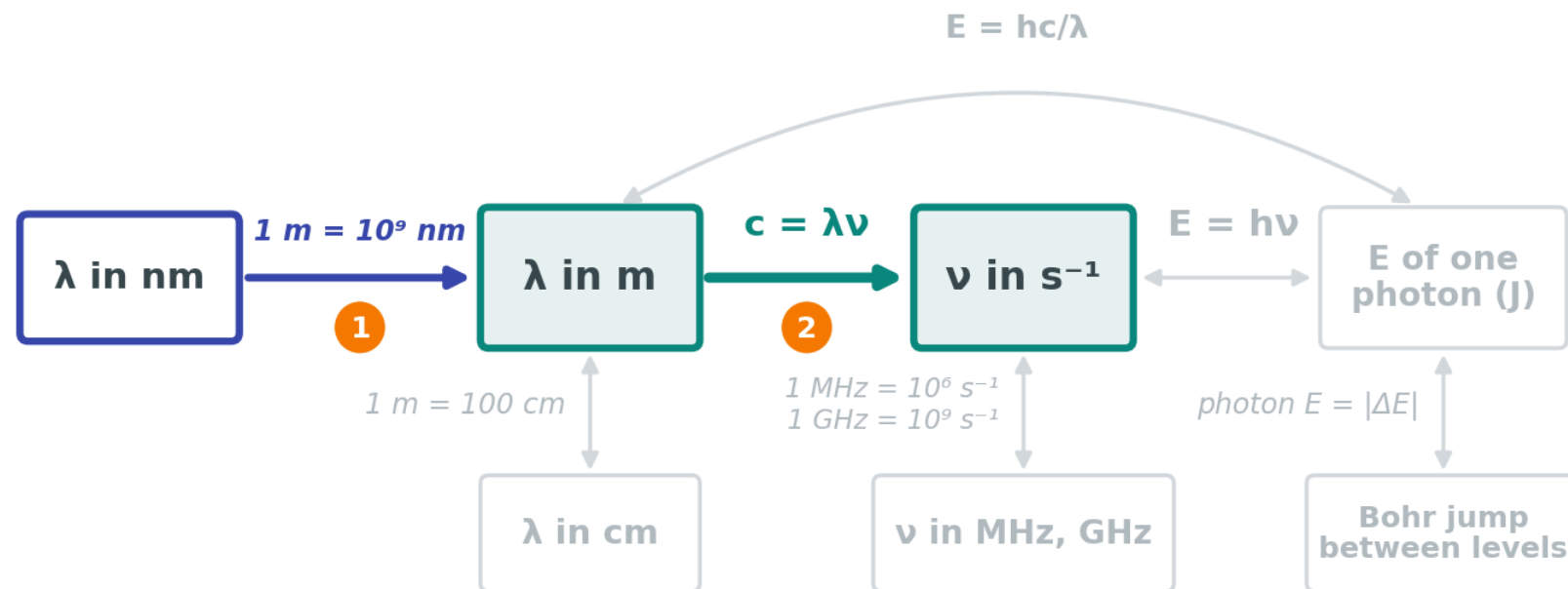
$$v = \frac{3.00 \times 10^8 \text{ m/s}}{2.50 \times 10^{-7} \text{ m}} = 1.20 \times 10^{15} \text{ s}^{-1}$$

 **Sense check:** Ultraviolet sits past violet, so its frequency runs higher than visible light. The unconverted route gave $1.20 \times 10^6 \text{ s}^{-1}$, a radio frequency for a UV lamp. ✗ vs ✓

✓ Worked example 3: the route on the map

$$c = \lambda \nu$$

given: $\lambda = 250 \text{ nm}$ · found: $\nu = 1.20 \times 10^{15} \text{ s}^{-1}$



🧠 **Sense check:** Arrow 1 is the hop the first attempt skipped: nanometers to meters. Arrow 2 is $c = \lambda \nu$, solved for ν . ✓

 Practice 2

$$c = \lambda \nu$$

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 2 · 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

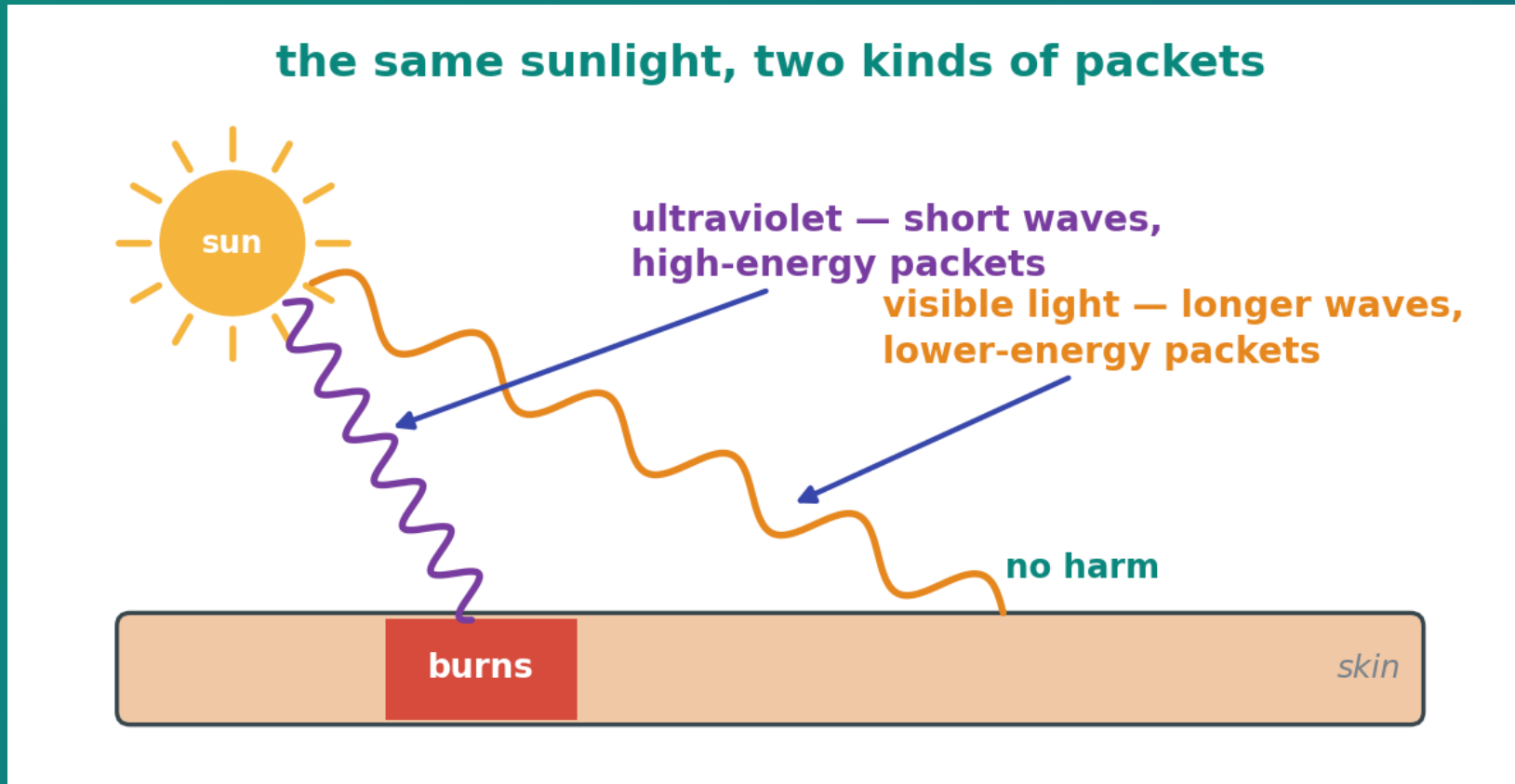
1. Blue light has a wavelength of 450 nm. Solve $c = \lambda\nu$ for its frequency in symbols. Which unit cancels, and which one survives?
2. An X-ray has a wavelength of 0.100 nm. Set up $c = \lambda\nu$ for its frequency. Is that frequency higher or lower than visible light's?

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

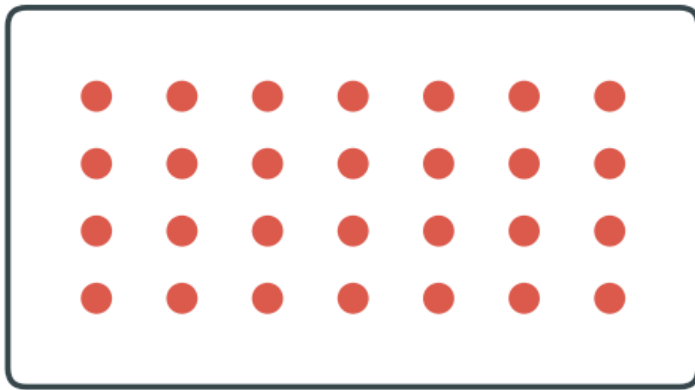


Ultraviolet light burns skin. The visible light beside it does not. The difference is not brightness. Each packet of ultraviolet carries far more energy than a packet of visible light.

Light comes in packets

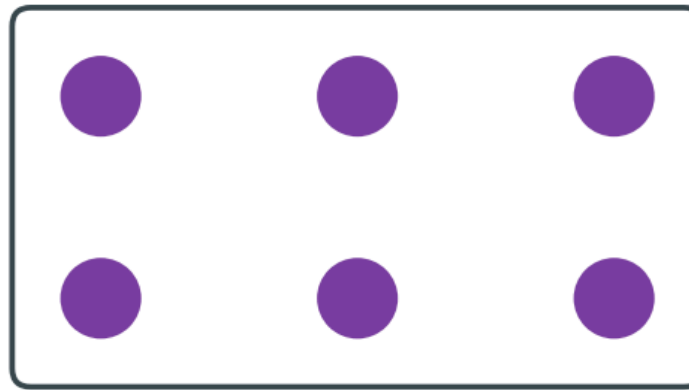
brightness is how MANY packets; frequency sets the energy of EACH

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**.

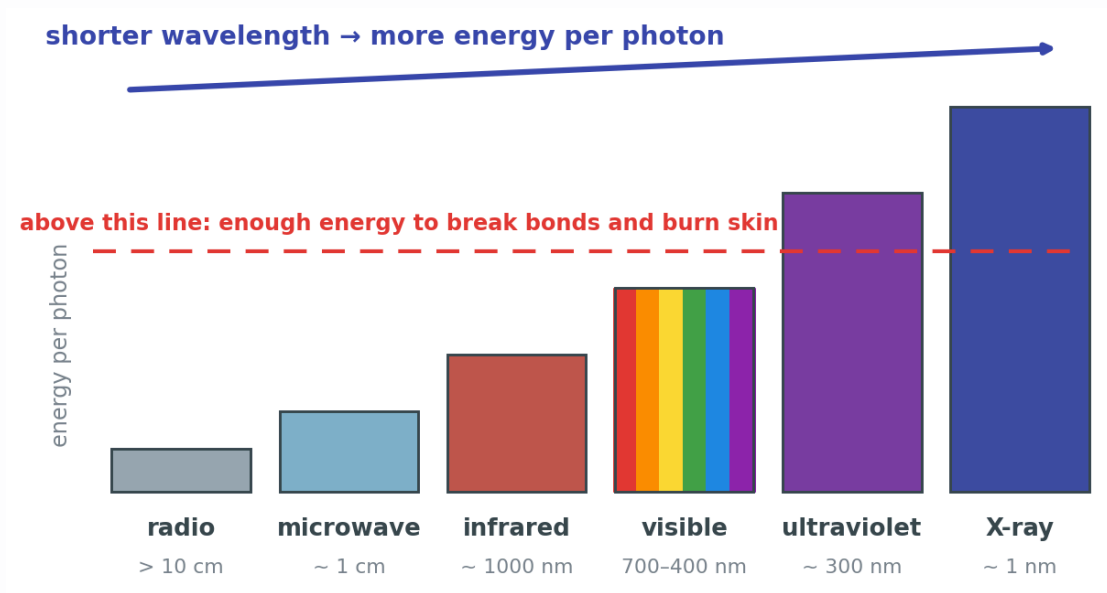
From wavelength: $E = hc/\lambda$

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

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

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

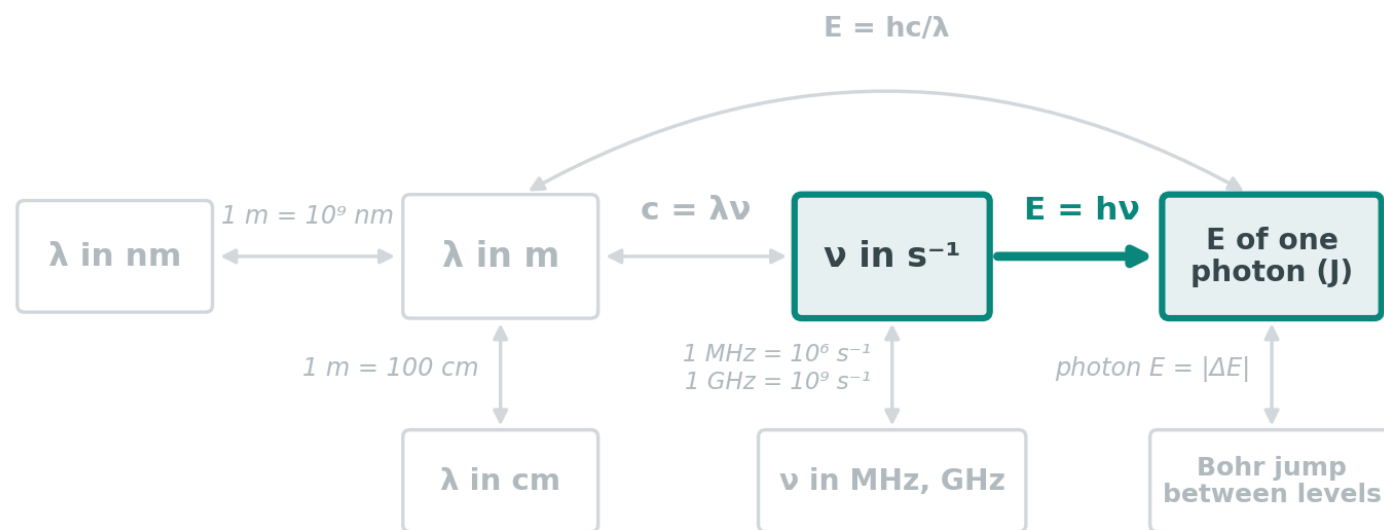
Step 4 · Check the size

🧠 **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. ✓



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

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

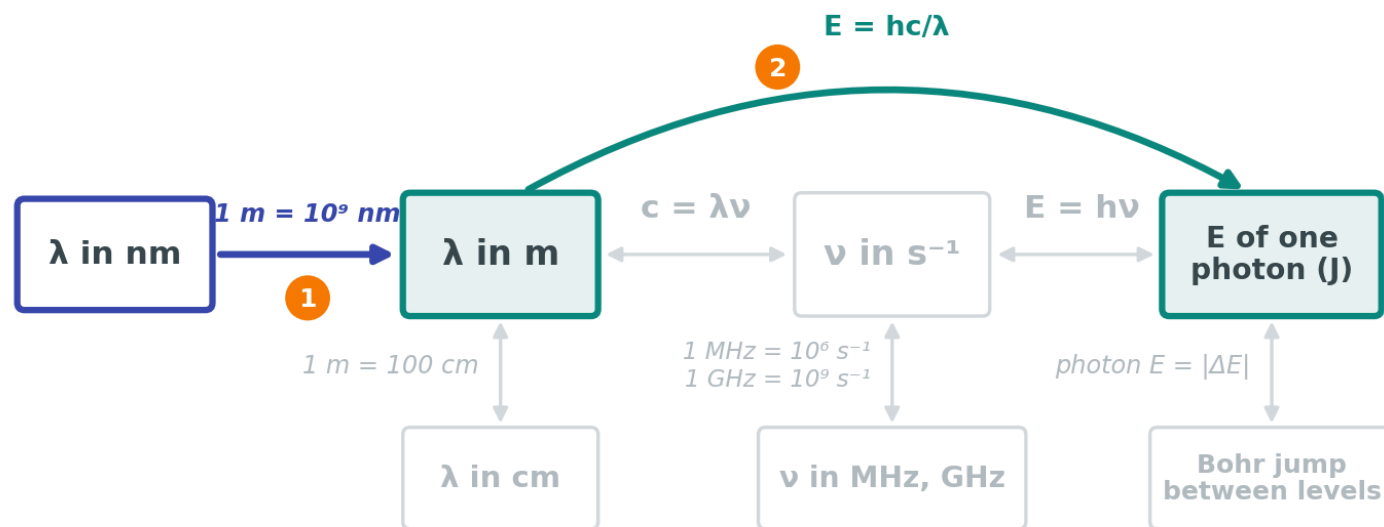
Step 4 · Check the size

 **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 1

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

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

✓ Practice 1 · 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.





Worked example 3: comparing two photons

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

$$\text{red } \lambda = 650 \text{ nm} \cdot \text{ultraviolet } \lambda = 325 \text{ nm} \cdot h = 6.626 \times 10^{-34} \text{ J}\cdot\text{s} \cdot c = 3.00 \times 10^8 \text{ m/s}$$

A red photon has a wavelength of **650 nm**; an ultraviolet photon has a wavelength of **325 nm**. Which carries more energy, and by how much?

A common first attempt: the 650 nm photon has the larger number, so call it the more energetic one. Test it.

✓ Worked example 3: solution

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

red $\lambda = 650 \text{ nm}$ · ultraviolet $\lambda = 325 \text{ nm}$ · wanted: which E is larger

A common first attempt

Step 1 · Identify the given

Step 2 · Choose the formula

Both photons use $E = hc/\lambda$, with the same h and c . Only λ differs, and 325 nm is exactly half of 650 nm.

 **Sense check:** λ in the denominator means the shorter 325 nm wavelength gives the larger energy, not the longer 650 nm one. ✓

✓ Worked example 3: which carries more

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

$$\text{red } \lambda = 650 \times 10^{-9} \text{ m} \cdot \text{ultraviolet } \lambda = 325 \times 10^{-9} \text{ m}$$

Step 3 • Substitute and cancel units

$$E_{\text{red}} = \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{650 \times 10^{-9}} = 3.06 \times 10^{-19} \text{ J}$$

$$E_{\text{UV}} = \frac{6.626 \times 10^{-34} \times 3.00 \times 10^8}{325 \times 10^{-9}} = 6.12 \times 10^{-19} \text{ J}$$

🧠 **Sense check:** Same numerator, smaller wavelength: the larger energy goes to the ultraviolet photon. ✓

✓ Worked example 3: which carries more

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

$$\text{red } \lambda = 650 \times 10^{-9} \text{ m} \cdot \text{ultraviolet } \lambda = 325 \times 10^{-9} \text{ m}$$

Step 4 • Check the size

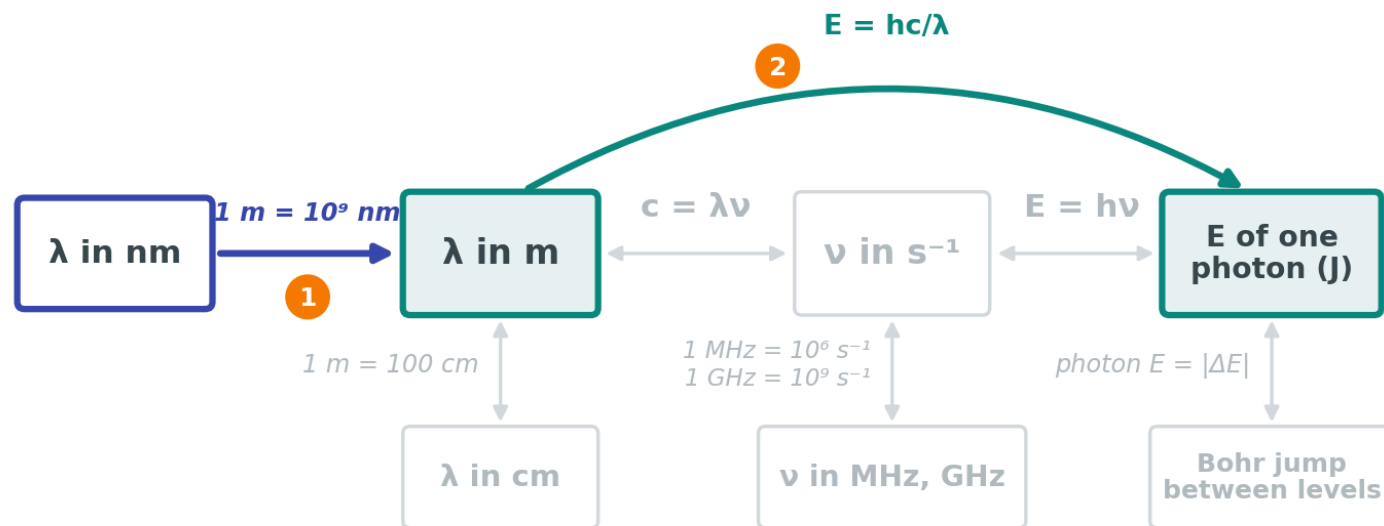
 **Sense check:** Half the wavelength, double the energy: $650 / 325 = 2$, and $6.12 \times 10^{-19} \text{ J}$ is twice $3.06 \times 10^{-19} \text{ J}$. The ultraviolet photon carries the larger energy, and that is why ultraviolet burns skin and the red light beside it does not.

✓

✓ Worked example 3: the route on the map

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

given: $\lambda = 650 \text{ nm}$ and 325 nm · found: $E = 3.06 \times 10^{-19} \text{ J}$ and $6.12 \times 10^{-19} \text{ J}$



🧠 Sense check: Each photon takes the same two moves: nanometers to meters, then $E = hc/\lambda$. Only the value of λ differs. ✓



Practice 2

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

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}

✓ Practice 2 · 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.

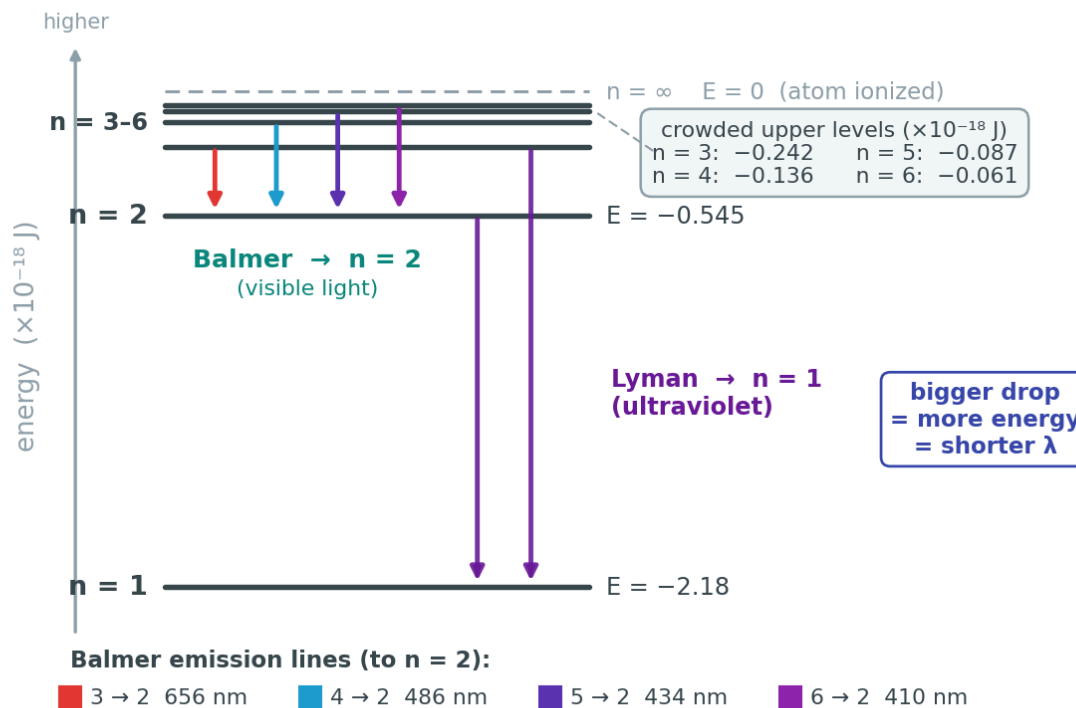
A **line spectrum** holds only a few wavelengths, bright lines in the dark: light from excited, low-pressure gas atoms.

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



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



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 $-2.18 \times 10^{-18} \text{ J}$

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

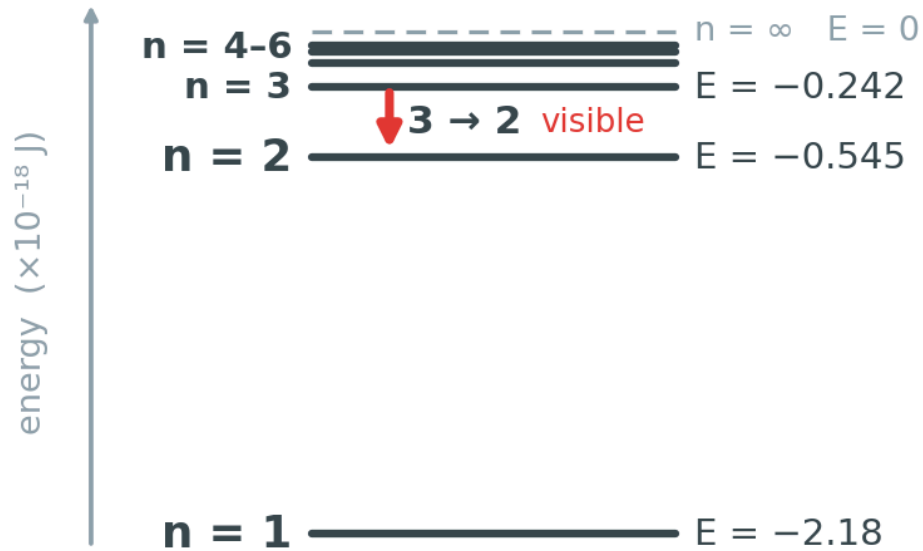
negative $\rightarrow$ the atom loses energy • the photon carries $|\Delta E| = 3.03 \times 10^{-19} \text{ J}$

🧠 **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). ✓

 Worked example 2: $n = 4 \rightarrow n = 1$

hydrogen: $n_i = 4 \rightarrow n_f = 1$

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

Same equation, a much bigger drop: all the way down to the ground level. Predict where this photon lands.

✓ Worked example 2: solution

$$n_i = 4, \quad n_f = 1$$

a Lyman jump: down to the ground level $n = 1$

Step 1 • Identify n_i and n_f

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

Step 3 • Subtract

$$1/1^2 - 1/4^2 = 1.0000 - 0.0625 = 0.9375$$

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

$$\Delta E = -2.18 \times 10^{-18} \text{ J} \times 0.9375 = -2.04 \times 10^{-18} \text{ J}$$

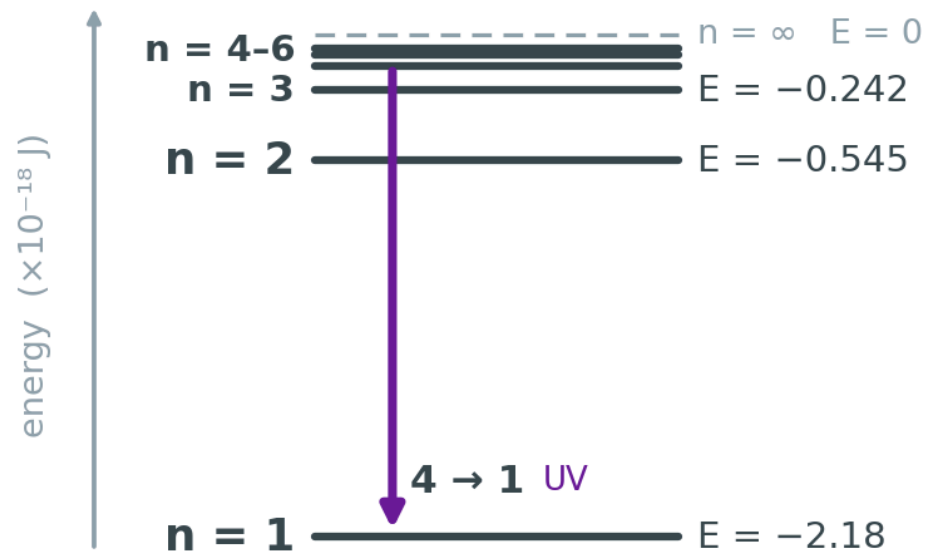
photon = $2.04 \times 10^{-18} \text{ J}$: far more than any Balmer jump

 **Sense check:** $2.04 \times 10^{-18} \text{ J}$ lands in the ultraviolet ($\lambda \approx 97 \text{ nm}$), the Lyman series. The bigger the drop, the more energy and the shorter the wavelength.

✓ Worked example 2: the jump on the ladder

$$n_i = 4 \rightarrow n_f = 1$$

found: $\Delta E = -2.04 \times 10^{-18} \text{ J}$ · the photon carries $2.04 \times 10^{-18} \text{ J}$



🧠 **Sense check:** All the way down to $n = 1$: a far bigger drop than $3 \rightarrow 2$, so an ultraviolet photon (97 nm). ✓

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: 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 1 · 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 2: 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. 106
- C. 9.31×10^{-8}
- D. 93.1
- E. 9.31×10^{-17}

✓ Practice 2 · 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 did not square n : $2.18 \times 10^{-18} \text{ J} \times (1/1 - 1/7) = 1.87 \times 10^{-18} \text{ J}$ gives 106 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 3: 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 3 · answer: B

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

$n = 3 \rightarrow n = 2 \Rightarrow \text{visible (656 nm, red)}$
to $n = 2 \rightarrow \text{visible} \cdot$ to $n = 1 \rightarrow \text{UV} \cdot$ to $n = 3 \rightarrow \text{infrared}$

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.



Check yourself

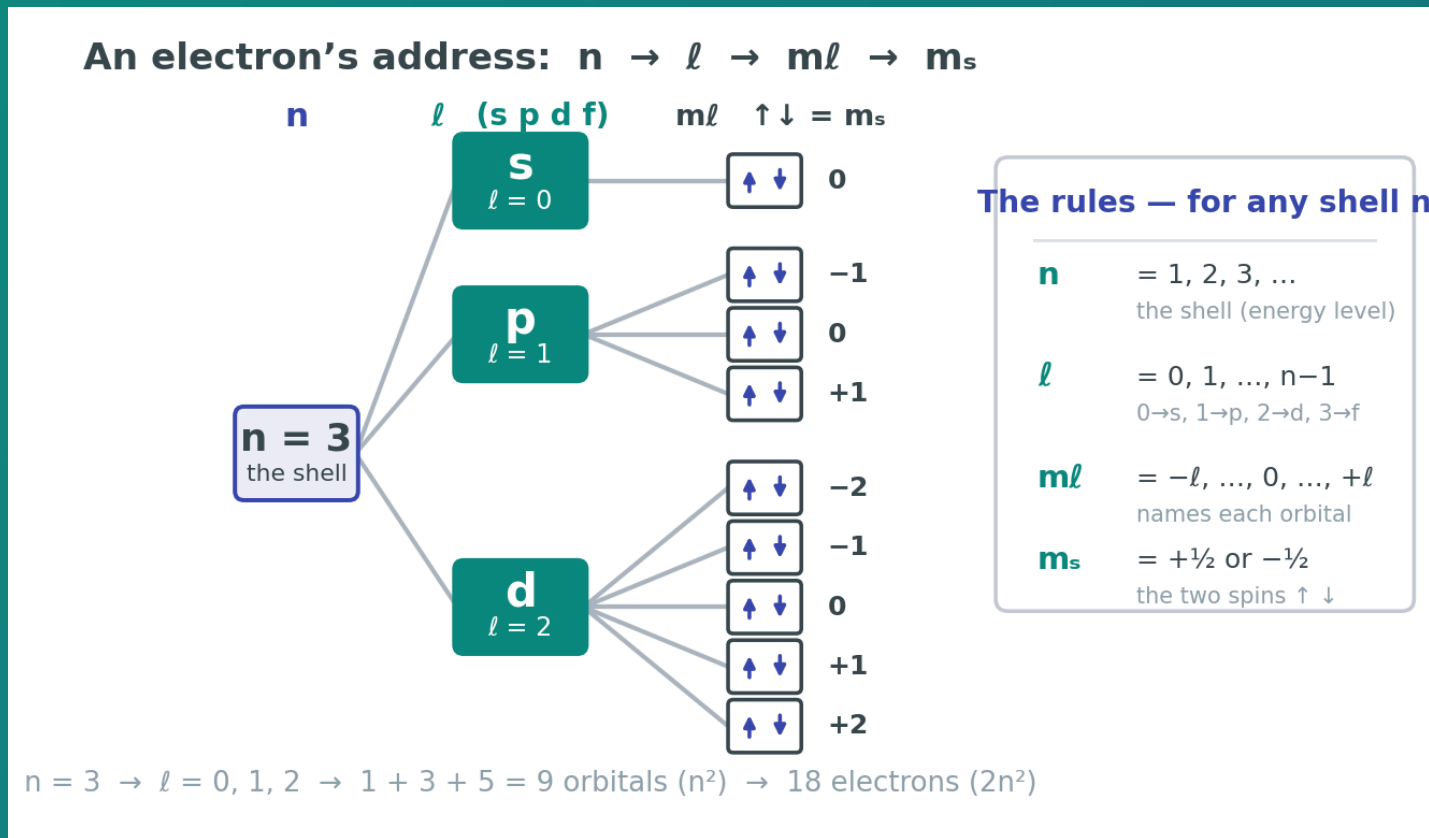
1. An electron in hydrogen falls from $n = 5$ to $n = 2$. Compute ΔE , and state whether the photon is visible or ultraviolet.
2. In one sentence, explain why hydrogen produces a line spectrum instead of a continuous one.

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 • Quantum Numbers

Assign a valid set of quantum numbers (n , ℓ , m_ℓ , m_s) to any electron, and explain what those numbers and the Heisenberg uncertainty principle say about where an electron is.

Every electron has an address



A mailing address narrows a location step by step. An electron's address does the same: state = n , city = ℓ , street = $m\ell$, house = m_s .

What a label like 3d holds

$$3d \rightarrow n = 3 \cdot \ell = 2$$

the digit is the energy level n · the letter is the subshell: $s = 0, p = 1, d = 2, f = 3$

ahead: $(2\ell + 1)$ orbitals $\times$ 2 electrons is why s will hold 2 · p 6 · d 10 · f 14

A label such as 3d carries two quantum numbers: the digit is n , the letter is ℓ . The configurations you will write next are strings of these labels, and four rules set their sizes.

Four numbers, four questions

Each quantum number answers one question, and each narrows the electron's location further than the last, exactly as an address line does.

$$n \cdot \ell \cdot m_\ell \cdot m_s$$

shell (state) → subshell (city) → orbital (street) → spin (house): one electron's full address

$$n = 1, 2, 3, \dots \cdot \ell = 0 \dots n-1 \text{ (s, p, d, f)} \cdot m_\ell = -\ell \dots +\ell \cdot m_s = \pm 1/2$$

the allowed values, in address order

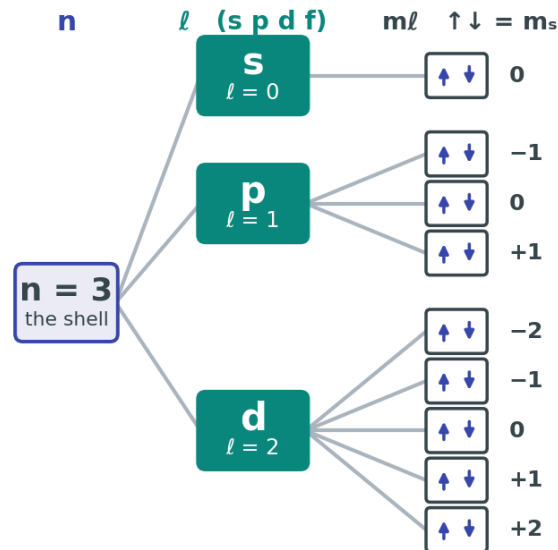
The rules chain together

Each number's range depends on the one before it. Read the tree left to right: the shell sets which subshells exist, and each subshell sets how many orbitals it holds.

$$\ell \leq n-1 \quad \cdot \quad |m_\ell| \leq \ell$$

ℓ can never reach n · m_ℓ can never exceed ℓ

An electron's address: $n \rightarrow \ell \rightarrow m_\ell \rightarrow m_s$



The rules — for any shell n

- n = 1, 2, 3, ...
the shell (energy level)
- ℓ = 0, 1, ..., $n-1$
0→s, 1→p, 2→d, 3→f
- m_ℓ = $-\ell, \dots, 0, \dots, +\ell$
names each orbital
- m_s = $+\frac{1}{2}$ or $-\frac{1}{2}$
the two spins $\uparrow \downarrow$

$n = 3 \rightarrow \ell = 0, 1, 2 \rightarrow 1 + 3 + 5 = 9$ orbitals (n^2) $\rightarrow 18$ electrons ($2n^2$)

Why an address, not a path

$$\Delta x \cdot \Delta p \geq \hbar/2$$

Heisenberg's uncertainty principle: position (Δx) and momentum (Δp) cannot both be pinned down at once

Pin down where an electron is and you lose where it is going. So no fixed orbits: an orbital holds the electron about 90% of the time, and the numbers name that region.

What each number tells you

n · shell → size and energy

the state: larger n , larger orbital, higher energy

ℓ · subshell → shape

the city: s sphere · p dumbbell · d cloverleaf

m_ℓ · orbital → orientation

the street: the three p orbitals point along x, y, z

m_s · spin → direction

the house: ↑ or ↓

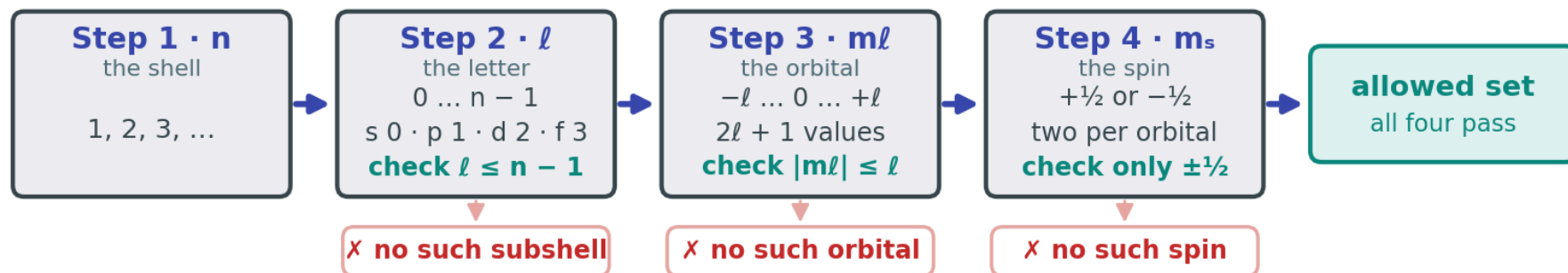
Two electrons sharing an orbital match on n , ℓ , and m_ℓ , so they must differ in m_s . No two electrons in an atom carry the same four numbers: the Pauli exclusion principle.



The method

1. n = the digit in the label (3 in 3d).
2. ℓ = the letter: s=0, p=1, d=2, f=3. Check $\ell \leq n-1$.
3. $m\ell$ = $-\ell$ to $+\ell$. Check $|m\ell| \leq \ell$.
4. m_s = $+\frac{1}{2}$ or $-\frac{1}{2}$.

Run the checks in order. One $\times$ forbids the whole set.





Worked example 1: a 3d electron

an electron in a 3d orbital

wanted: one valid set of quantum numbers (n , ℓ , m_ℓ , m_s)

Give a complete set of four quantum numbers that could describe this electron. Read the label 3d, then apply each rule in turn.

✓ Worked example 1: solution

Step 1 · n from the shell number

Step 2 · ℓ from the letter

Step 3 · m_ℓ from $-\ell$ to $+\ell$

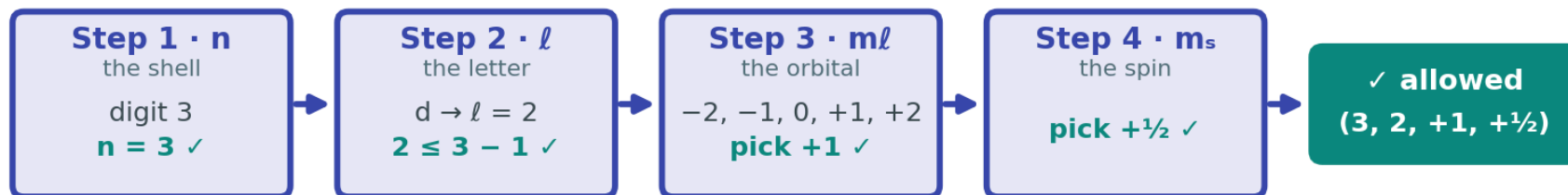
Step 4 · $m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

$$(n, \ell, m_\ell, m_s) = (3, 2, +1, +\frac{1}{2})$$

a valid set · n and ℓ are fixed by "3d"; m_ℓ and m_s pick the seat

 **Sense check:** The 3d subshell holds $(2 \times 2 + 1) \times 2 = 10$ electrons (5 orbitals, 2 spins each), so ten different valid sets share $n = 3$, $\ell = 2$. The label fixes n and ℓ ; m_ℓ and m_s choose among the openings.

an electron in a 3d orbital





Guided example: checking a set

$$n = 3, \ell = 1, m_\ell = -2, m_s = +\frac{1}{2}$$

given: one set (n, ℓ, m_ℓ, m_s) · decide: allowed or forbidden

Run the four method steps in order on this set. Name the first rule it breaks, if any.

✓ Guided example: solution

$n = 3, \ell = 1, m_\ell = -2, m_s = +\frac{1}{2}$
given: one set · decide: allowed or forbidden

Step 1 · n from the shell number

Step 2 · ℓ from the letter

Step 3 · m_ℓ from $-\ell$ to $+\ell$

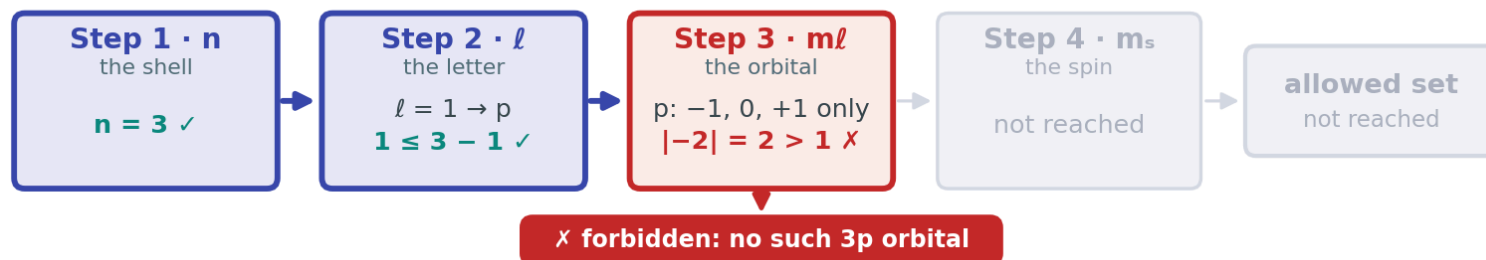
Step 4 · $m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

$$|m_\ell| = |-2| = 2 > \ell = 1 \times$$

forbidden: a 3p subshell has three orbitals, and none carries $m_\ell = -2$

🧠 **Sense check:** To keep 3p, m_ℓ must be $-1, 0$, or $+1$. To keep $m_\ell = -2$, ℓ must be at least 2: a 3d electron.

checking $n = 3, \ell = 1, m_\ell = -2, m_s = +\frac{1}{2}$



Your turn: a 4p electron

An electron sits in a 4p orbital. Fill each quantum number, checking it against its rule.

quantum number	value
n	_____
ℓ	_____
$m\ell$ (choose one)	_____
m_s	_____

$$n = 4, \ell = 1, m\ell = -1 \text{ (or 0 or +1), } m_s = +\frac{1}{2} \text{ (or } -\frac{1}{2})$$

$$4 \rightarrow n = 4 \cdot p \rightarrow \ell = 1 (\leq 3 \checkmark) \cdot m\ell \text{ from } -1 \text{ to } +1 \cdot \text{spin } \pm\frac{1}{2}$$

⚠ Where this goes wrong

Letting ℓ equal n . For $n = 3$, writing $\ell = 3$ is forbidden: ℓ stops at $n-1 = 2$. The subshells in $n = 3$ are only s, p, d ($\ell = 0, 1, 2$).

Pushing m_ℓ past ℓ . A p subshell has $\ell = 1$, so m_ℓ is $-1, 0$, or $+1$. Writing $m_\ell = +2$ for a p electron breaks $|m_\ell| \leq \ell$; $+2$ first appears in a d subshell.

Giving m_s any value but $\pm\frac{1}{2}$. Spin is only $+\frac{1}{2}$ or $-\frac{1}{2}$. A "0" or "+1" for m_s is never allowed: an electron has exactly two spin states.

Reading m_ℓ 's range from n . m_ℓ runs from $-\ell$ to $+\ell$, not $-n$ to $+n$. It is ℓ , the shape number, that sets how many orbitals a subshell has ($2\ell+1$).

Practice 1

Which set of quantum numbers is **not** allowed for an electron in an atom?

A. $n = 3, \ell = 2, m_\ell = -2, m_s = +\frac{1}{2}$

B. $n = 2, \ell = 1, m_\ell = 0, m_s = -\frac{1}{2}$

C. $n = 1, \ell = 1, m_\ell = 0, m_s = +\frac{1}{2}$

D. $n = 4, \ell = 0, m_\ell = 0, m_s = -\frac{1}{2}$

✓ Practice 1 · answer: C

$n = 1, \ell = 1 \rightarrow n - 1 = 0$, but $\ell = 1 \times$
 answer C: the $n = 1$ shell holds only $\ell = 0$ (the 1s subshell)

🧠 **Sense check:** Scan the set in order: ℓ against n first, then m_ℓ against ℓ , then m_s . C breaks at the very first check.

choice C: $n = 1, \ell = 1, m_\ell = 0, m_s = +\frac{1}{2}$





Practice 2

	statement
1	An electron in a 5f orbital may have $m_\ell = -3$
2	An electron with $n = 5$ and $\ell = 1$ sits in a 5s orbital
3	The $n = 4$ shell holds s, p, d, and f subshells

Three statements about electrons in atoms. Which statements are correct?

- A. Statements 1 and 3 only
- B. Statement 3 only
- C. Statements 1, 2 and 3
- D. Statement 1 only

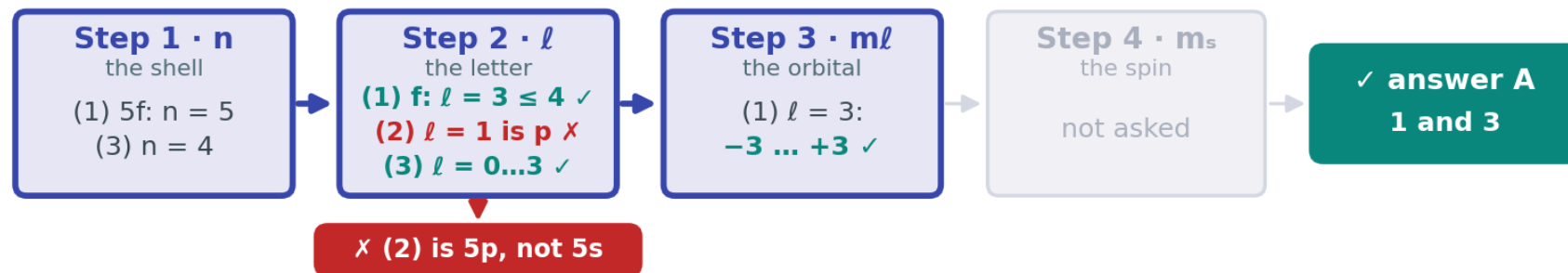
✓ Practice 2 · answer: A

correct: 1 and 3 · false: 2 (answer A)

1: f is $\ell = 3$, so m_ℓ runs -3 to $+3$ ✓ · 2: $\ell = 1$ is p, so the orbital is 5p ✗ · 3: $\ell \leq n-1 = 3$, so s, p, d, f ✓

🧠 **Sense check:** Check each statement against its own rule: m_ℓ against ℓ , the letter against ℓ , ℓ against n . Two survive.

three statements, three checks





Practice 3

Four of these sets describe real electrons. Which set can describe no electron in any atom?

A. $n = 4, \ell = 2, m_\ell = +2, m_s = +\frac{1}{2}$

B. $n = 7, \ell = 0, m_\ell = 0, m_s = -\frac{1}{2}$

C. $n = 4, \ell = 3, m_\ell = 0, m_s = +\frac{1}{2}$

D. $n = 5, \ell = 2, m_\ell = -1, m_s = +\frac{1}{2}$

E. $n = 3, \ell = 2, m_\ell = -3, m_s = +\frac{1}{2}$

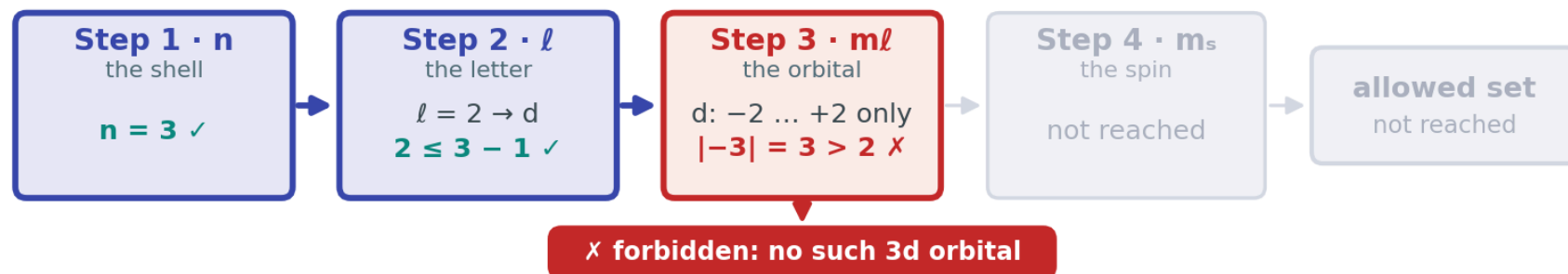
✓ Practice 3 · answer: E

$n = 3, \ell = 2 \rightarrow m_\ell = -2 \dots +2$, but $m_\ell = -3 \times$ (answer E)

$\ell = 2 \leq n-1 = 2 \checkmark \cdot |m_\ell| = 3 > \ell = 2 \times \cdot$ a 3d subshell has five orbitals, none with $m_\ell = -3$

🧠 **Sense check:** A through D each look unusual and still pass every check. Only E breaks a rule: ℓ sets m_ℓ 's range, never n .

choice E: $n = 3, \ell = 2, m_\ell = -3, m_s = +\frac{1}{2}$





Worked example 2: is this set allowed?

$$n = 2, \ell = 2, m_\ell = 0, m_s = +\frac{1}{2}$$

decide: valid or forbidden, and name the rule

A student writes this set for an electron. Check each number against its rule before accepting it.

✓ Worked example 2: solution

Check ℓ against n

$\ell = 2$ with $n = 2 \rightarrow n-1 = 1$, but $2 > 1 \times$
 forbidden: ℓ has reached n , which the rule never allows

The fix

 **Sense check:** $m_\ell = 0$ and $m_s = +\frac{1}{2}$ are both fine on their own. One broken rule (ℓ too large for its shell) is enough to forbid the whole set.

checking $n = 2, \ell = 2, m_\ell = 0, m_s = +\frac{1}{2}$



Check yourself

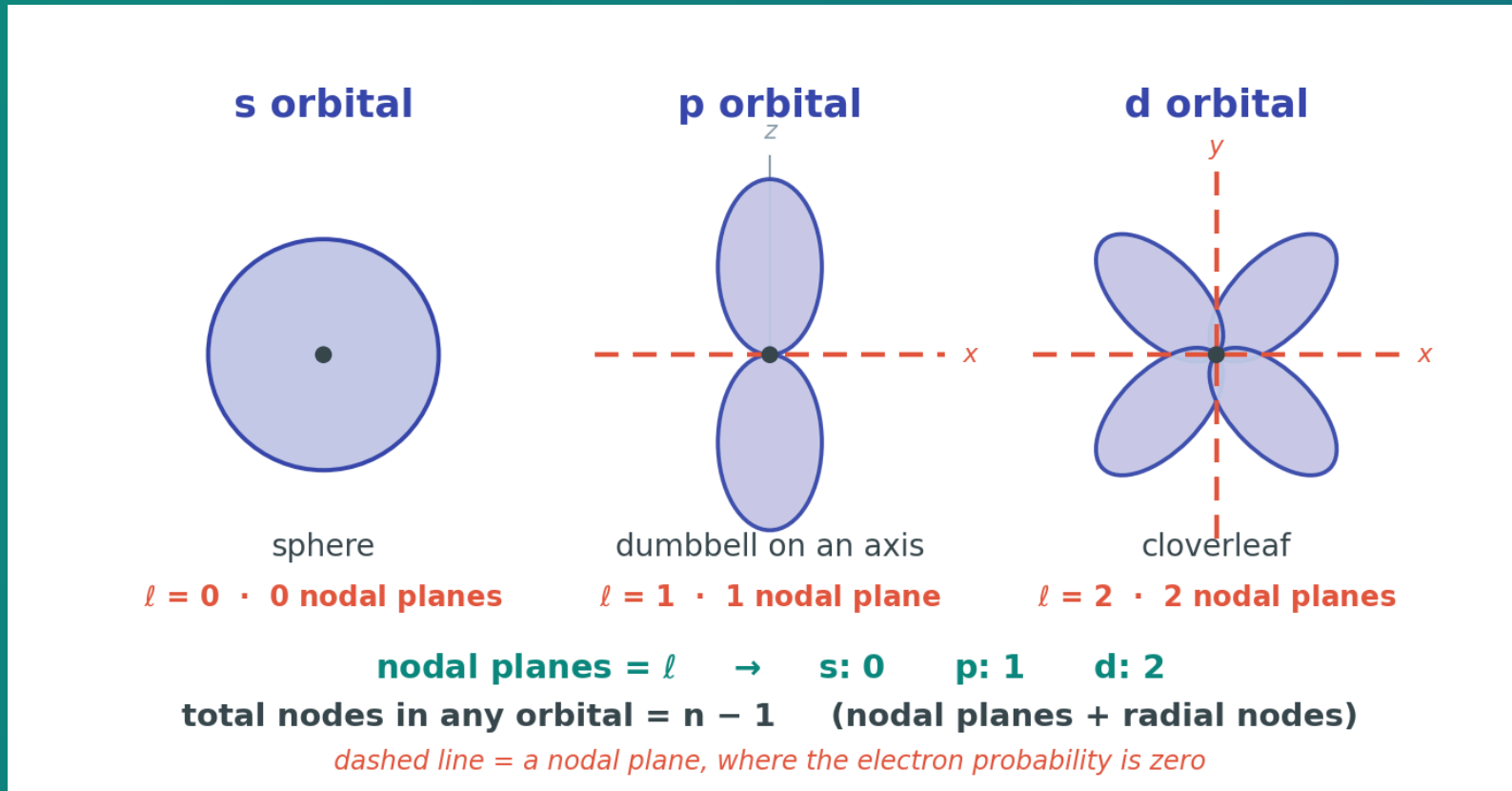
1. Give one valid set of quantum numbers for an electron in a 2p orbital, and state how many valid sets a full 2p subshell allows.
2. Why is the set $n = 3$, $\ell = 0$, $m_\ell = +1$, $m_s = +\frac{1}{2}$ forbidden? Name the rule it breaks.

The four quantum numbers label every electron uniquely, and no two in an atom share all four. Filling those addresses from the lowest energy upward (1s, then 2s, then 2p) builds the atom's **electron configuration**.

5 • Orbital Shapes & Nodes

Describe the shapes of s, p, and d orbitals and count their nodes, using nodal planes = ℓ and total nodes = $n - 1$.

The shape of the cloud



An electron is not a dot on a path. The best we can do is map where it likely sits: a region with a shape, set by the sublevel.

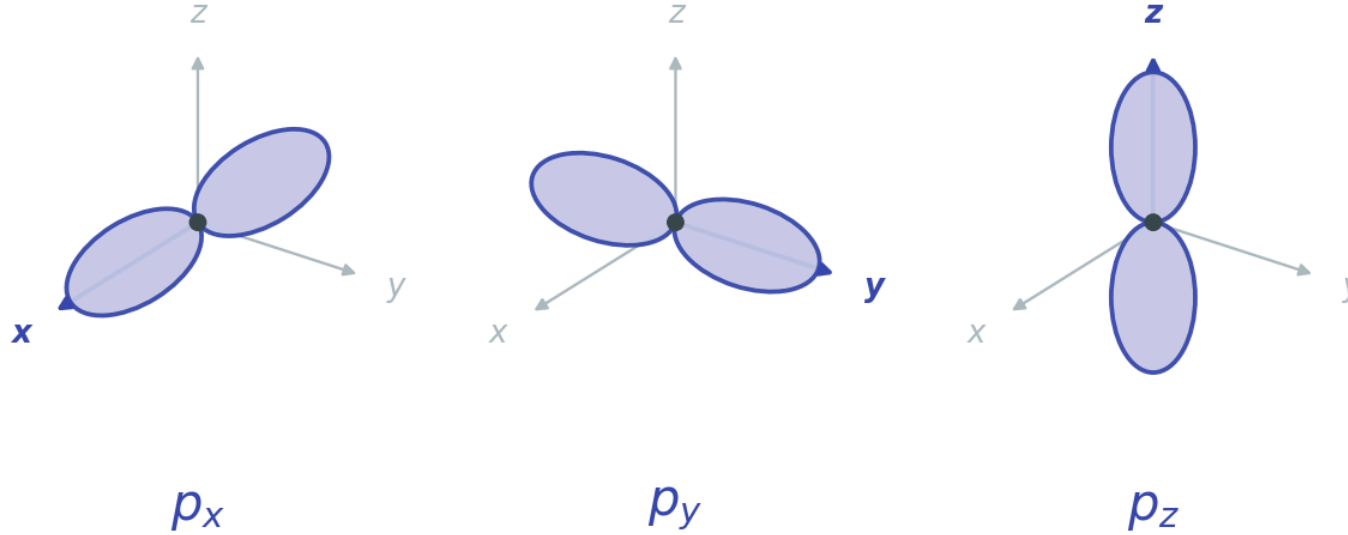
What n and ℓ already describe

$$3p \rightarrow n = 3 \cdot \ell = 1 \cdot m_\ell = -1, 0, +1$$

ℓ sets the shape · $2\ell + 1 = 3$ orbitals of that shape · n sets the size and the node count

Quantum numbers give each electron an address. The same numbers describe its orbital. The letter fixes the shape, m_ℓ counts the orientations of that shape, and n sets how far out the cloud reaches.

An orbital is a region of probability



Three p orbitals — one identical dumbbell each, pointing along x, y, and z.

An orbital is the region where an electron most likely is. The sublevel letter sets the shape, and the three p orbitals differ only in the axis they point along.

orbital = the electron's 90% region

s → sphere · p → dumbbell on an axis · d → cloverleaf

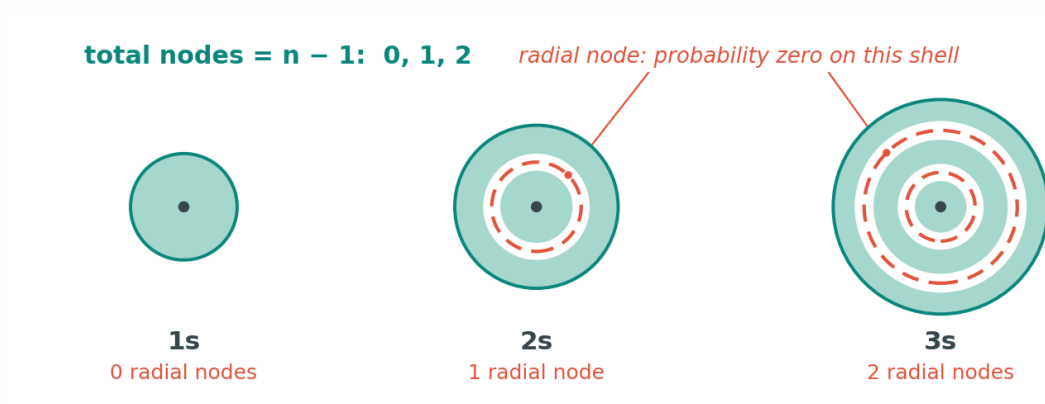
Nodes: where probability is zero

A node is a surface where the electron's probability drops to zero. The flat nodal planes through the nucleus are counted by ℓ : s has none, p has one, d has two.

nodal planes = ℓ

$s \rightarrow 0 \cdot p \rightarrow 1 \cdot d \rightarrow 2 \cdot$ flat surfaces cutting through the nucleus

Two kinds of node, one total



Total nodes equal $n - 1$. Whatever ℓ leaves behind shows up as radial nodes: spherical shells where the probability dips to zero. More nodes mean a higher-energy orbital.

$$\text{total nodes} = n - 1 \quad \cdot \quad \text{radial nodes} = n - \ell - 1$$

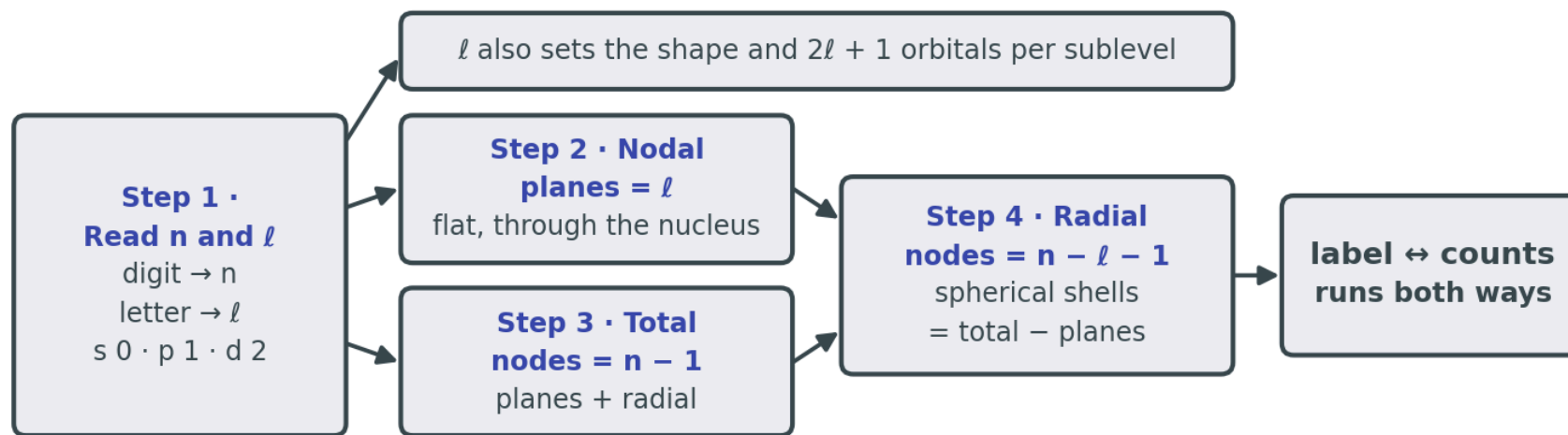
planes (ℓ) + spheres ($n - \ell - 1$) = $n - 1$ · a 1s orbital has $1 - 1 = 0$ nodes



The method

1. **Read n and ℓ** from the label: $s=0$, $p=1$, $d=2$.
2. **Nodal planes = ℓ** , flat, through the nucleus.
3. **Total nodes = $n - 1$** .
4. **Radial nodes = $n - \ell - 1$** , spherical shells.

Two numbers set every count. The route runs both ways.





Worked example 1: the 3p orbital

3p orbital

given: the label · wanted: nodal planes and total nodes

How many nodal planes does the 3p orbital have, and how many nodes in all? Read the label first.

✓ Worked example 1: solution

3p orbital: $n = 3 \cdot p \rightarrow \ell = 1$

Step 1 · Read n and ℓ

Step 2 · Nodal planes = ℓ

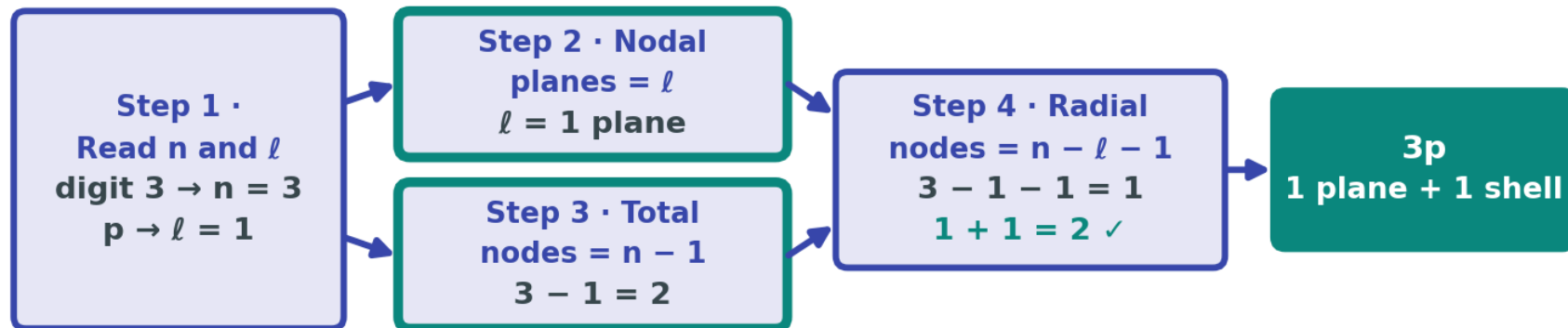
Step 3 · Total nodes = $n - 1$

Step 4 · Radial nodes = $n - \ell - 1$

total = $3 - 1 = 2$ · radial = $3 - 1 - 1 = 1$

1 plane + 1 spherical shell = 2 nodes ✓

🧠 **Sense check:** A 2p orbital has the same plane but only $2 - 1 = 1$ node. The 3p extra is a radial shell. ✓





Worked example 2: the 3s orbital

3s orbital

given: the label · wanted: nodal planes and total nodes

A sphere has no flat surface through its center, so a common first attempt calls the 3s orbital node-free. Count both kinds before agreeing.

✓ Worked example 2: solution

3s orbital: $n = 3 \cdot s \rightarrow \ell = 0$

A common first attempt

Step 1 · Read n and ℓ

Step 2 · Nodal planes = ℓ

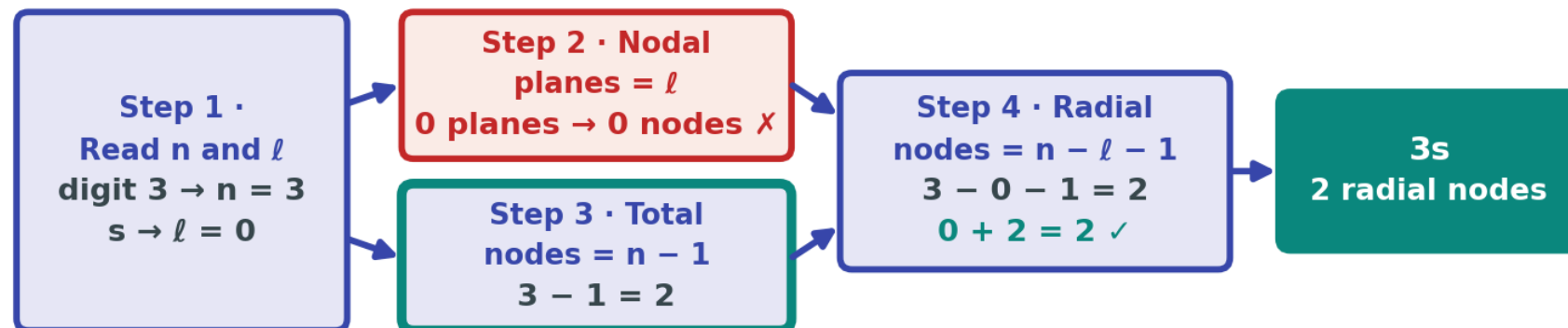
Step 3 · Total nodes = $n - 1$

Step 4 · Radial nodes = $n - \ell - 1$

total = $3 - 1 = 2$ · radial = $3 - 0 - 1 = 2$

0 planes + 2 spherical shells = 2 nodes ✓

🧠 **Sense check:** No planes is not no nodes. The 3s sphere hides two spherical shells. ✓



Your turn: the 4d orbital

4d orbital

read the label, then run the three counts

step	count
ℓ from the letter d	_____
nodal planes = ℓ	_____
total nodes = $n - 1$	_____
radial nodes = $n - \ell - 1$	_____

$$4d: \ell = 2 \cdot \text{planes} = 2 \cdot \text{total} = 4 - 1 = 3 \cdot \text{radial} = 4 - 2 - 1 = 1$$

$$2 \text{ planes} + 1 \text{ spherical shell} = 3 \text{ nodes } \checkmark$$

⚠ Where this goes wrong

Confusing nodal planes with total nodes. A 3p orbital has $\ell = 1$, so 1 nodal plane, but its total is $n - 1 = 2$. The extra node is a radial (spherical) node.

Giving a d orbital 0 nodal planes. A cloverleaf needs 2 planes to separate its four lobes: nodal planes $= \ell = 2$. For 3d, radial nodes $= 3 - 2 - 1 = 0$, so both of its nodes are planes.

Reading the 2p total as 2. A 2p orbital has $n = 2$, so $n - 1 = 1$ node: its single nodal plane. Radial nodes $= 2 - 1 - 1 = 0$; there are none.

Thinking p_x , p_y , and p_z have different shapes. They are one identical dumbbell, differing only in the axis they point along.

Practice 1

How many **radial nodes** does one 6d orbital have?

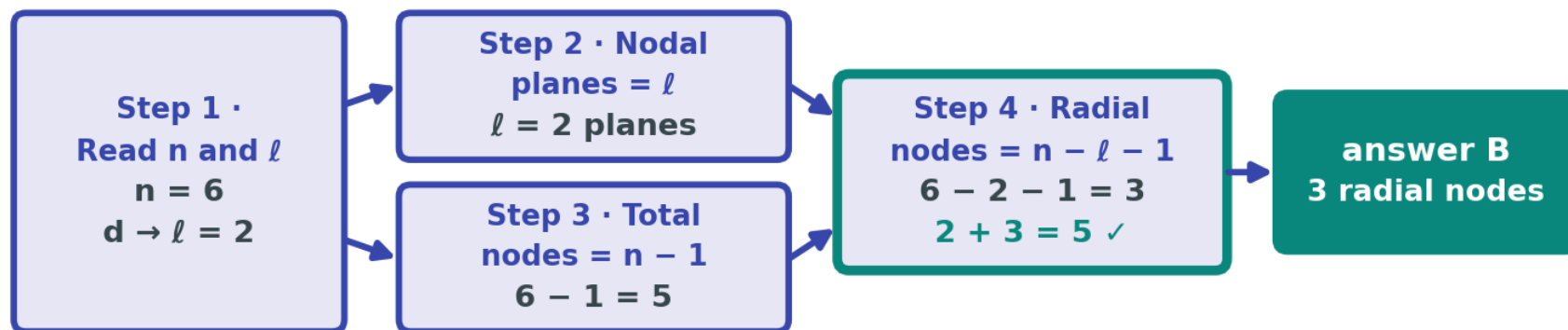
- A. 5
- B. 3
- C. 4
- D. 6

✓ Practice 1 · answer: B

6d: radial nodes = $n - \ell - 1 = 6 - 2 - 1 = 3$ (answer B)

$n = 6 \cdot d \rightarrow \ell = 2 \cdot 2 \text{ planes} + 3 \text{ radial shells} = 5 = 6 - 1 \checkmark$

🧠 **Sense check:** The two kinds add back to the total: 2 planes + 3 spherical shells = 5 = $n - 1$. ✓





Guided example: an orbital from its node counts

2 nodal planes · 2 radial nodes

given: both kinds of node · wanted: the orbital's label and its total node count

The label is the unknown here. Run the same four method steps, solving each one for n or ℓ .

✓ Guided example: solution

given: 2 nodal planes · 2 radial nodes · wanted: the label and the total

Step 1 · Read n and ℓ

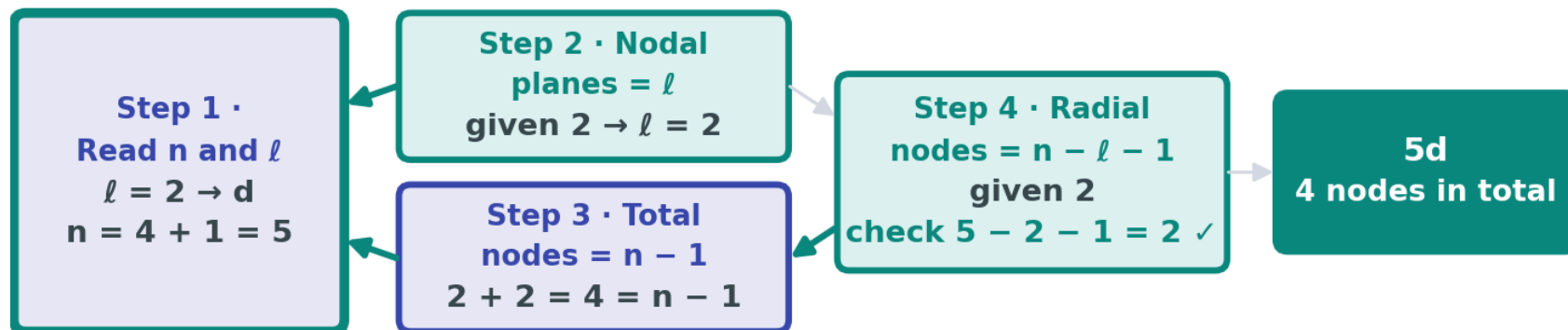
Step 2 · Nodal planes = ℓ

Step 3 · Total nodes = $n - 1$

$\ell = 2 \rightarrow d$ · total = $2 + 2 = 4 = n - 1 \rightarrow n = 4 + 1 = 5$
the orbital is 5d, with 4 nodes in total

Step 4 · Radial nodes = $n - \ell - 1$

🧠 **Sense check:** Check the label forward: $5 - 2 - 1 = 2$ radial nodes, the given count. ✓





Practice 2

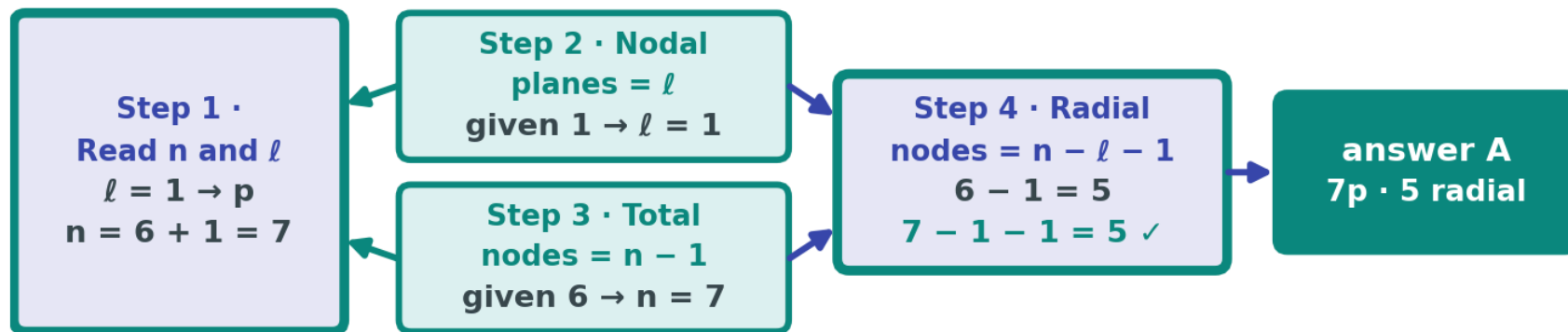
An orbital shows **1 nodal plane** and **6 nodes** in all. Which orbital is it, and how many radial nodes does it carry?

- A. orbital: 7p · radial nodes: 5
- B. orbital: 6p · radial nodes: 4
- C. orbital: 5p · radial nodes: 3
- D. orbital: 7p · radial nodes: 6

✓ Practice 2 · answer: A

planes = $\ell = 1 \rightarrow p \cdot n - 1 = 6 \rightarrow n = 6 + 1 = 7 \cdot \text{radial} = 7 - 1 - 1 = 5$ (answer A)
 7p: 1 nodal plane + 5 radial shells = 6 nodes ✓

🧠 **Sense check:** Planes name the letter, the total names the shell, and the radial count is what is left: $1 + 5 = 6$. ✓





Practice 3

Which description of one **7d** orbital gets its shape, the number of orbitals in its sublevel, and its total node count all right?

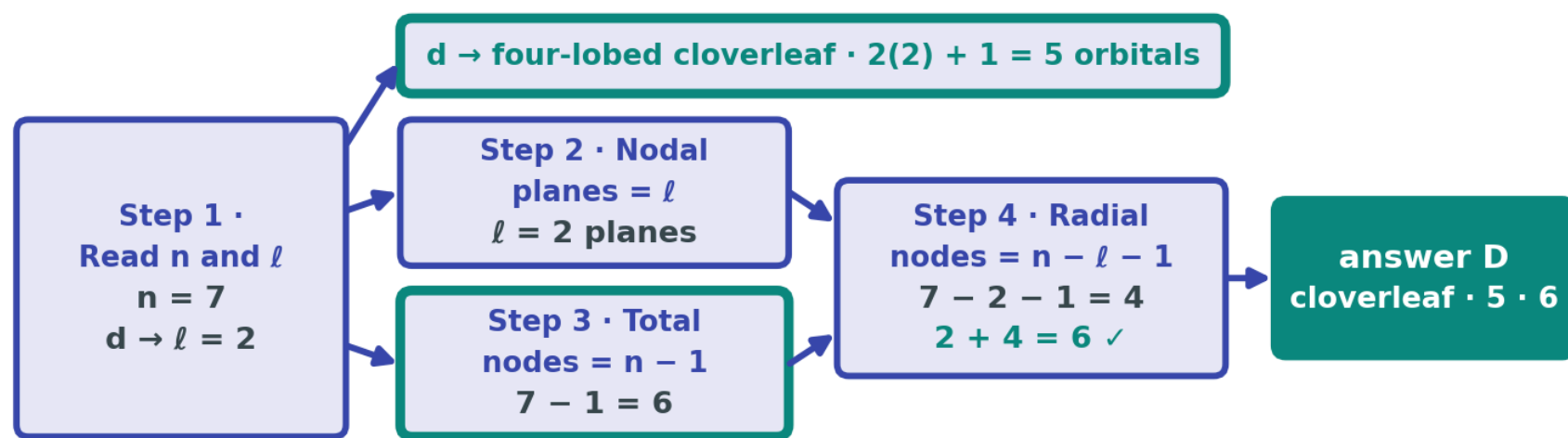
- A. two-lobed dumbbell · sublevel of 10 orbitals · 6 nodes in total
- B. four-lobed cloverleaf · sublevel of 10 orbitals · 7 nodes in total
- C. two-lobed dumbbell · sublevel of 5 orbitals · 7 nodes in total
- D. four-lobed cloverleaf · sublevel of 5 orbitals · 6 nodes in total

✓ Practice 3 · answer: D

7d: $\ell = 2 \rightarrow \text{cloverleaf} \cdot 2(2) + 1 = 5 \text{ orbitals} \cdot 7 - 1 = 6 \text{ nodes (answer D)}$

$$n = 7 \cdot 2 \text{ nodal planes} + 7 - 2 - 1 = 4 \text{ radial shells} = 6 \text{ nodes } \checkmark$$

🧠 **Sense check:** The letter sets the shape and the orbital count. The digit sets the total. Each reading comes from its own number. ✓



 Practice 4

one 4p orbital

$$n = 4 \cdot p \rightarrow \ell = 1$$

How many **nodes in total** does one 4p orbital have?

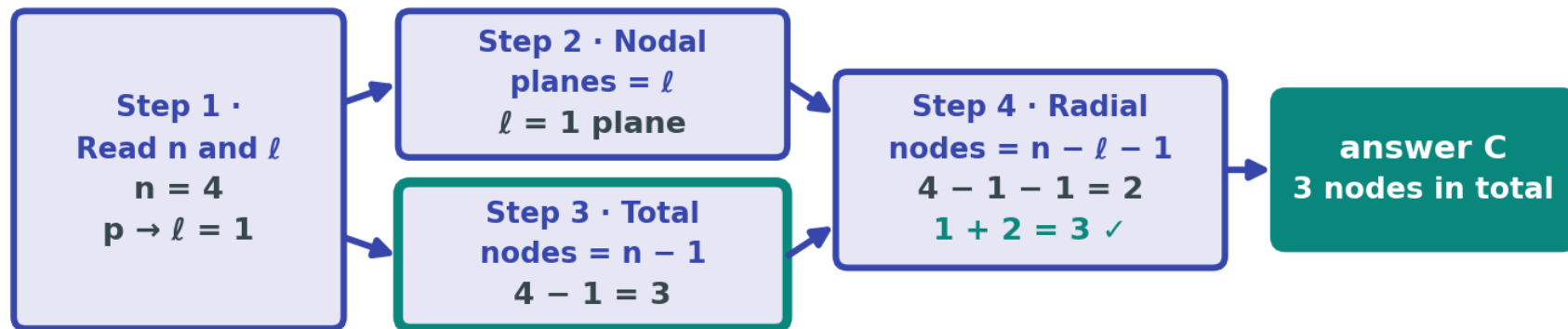
- A. 4
- B. 1
- C. 3
- D. 2

✓ Practice 4 · answer: C

4p: total nodes = $n - 1 = 4 - 1 = 3$ (answer C)

1 nodal plane (ℓ) + 2 radial shells ($n - \ell - 1$) = 3 ✓

🧠 **Sense check:** The two kinds always sum to $n - 1$: for 4p, one flat plane plus two spherical shells. ✓





Check yourself

1. A 5s orbital: how many nodal planes, and how many nodes in all?
2. Which has more nodal planes, 3d or 4p? Which has more total nodes?

Shapes come one to a sublevel letter, but not one to a shell: each s sublevel holds a single orbital, each p holds three, each d five. Counting those orbitals, and the two electrons each one holds, fixes how many electrons every sublevel and shell can carry.

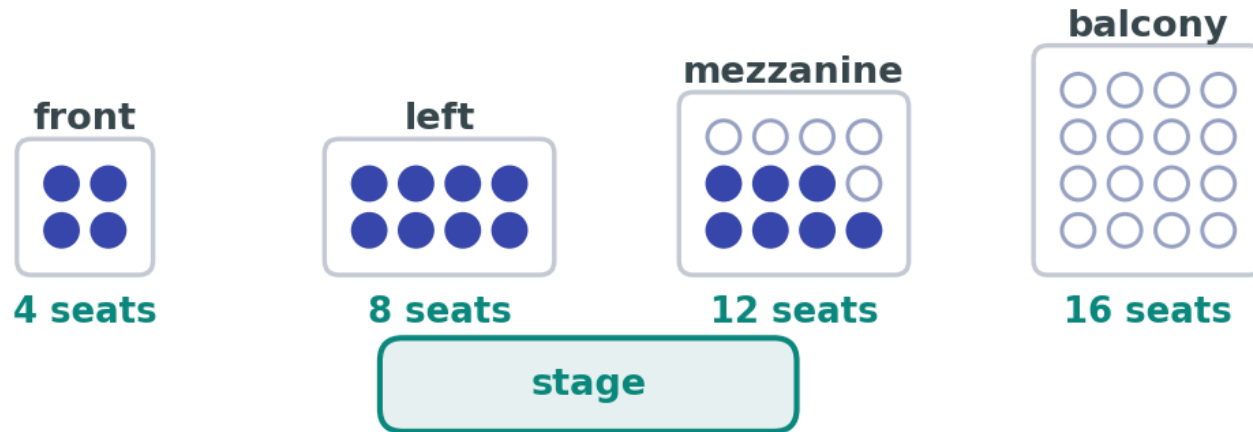
6 · Sublevels & Orbitals

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 $2 \cdot 6 \cdot 10 \cdot 14$

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.

Take-home: orbitals are not electrons

5 orbitals $\neq$ 5 electrons

the d sublevel: 5 orbitals $\times$ 2 = 10 electrons

Orbitals and electrons are not the same count. A sublevel's orbital count answers "how many orbitals"; doubling it answers "how many electrons". Match the number to the question.

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.

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

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

⚠ 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

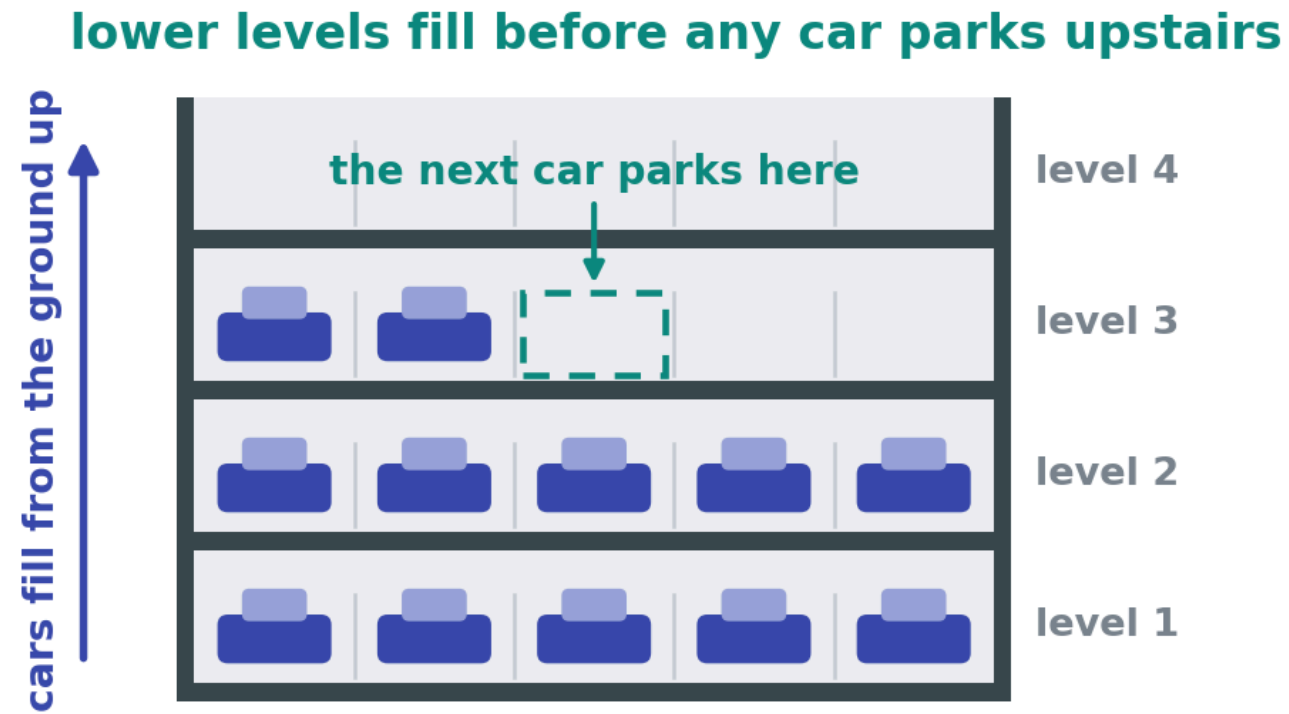
1. A full 4d sublevel: how many orbitals, and how many electrons? Name the type, then double.
2. Which shell first includes an f sublevel, and what is that shell's total capacity, $2n^2$?

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.

7 • 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 ✓

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

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.**



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.

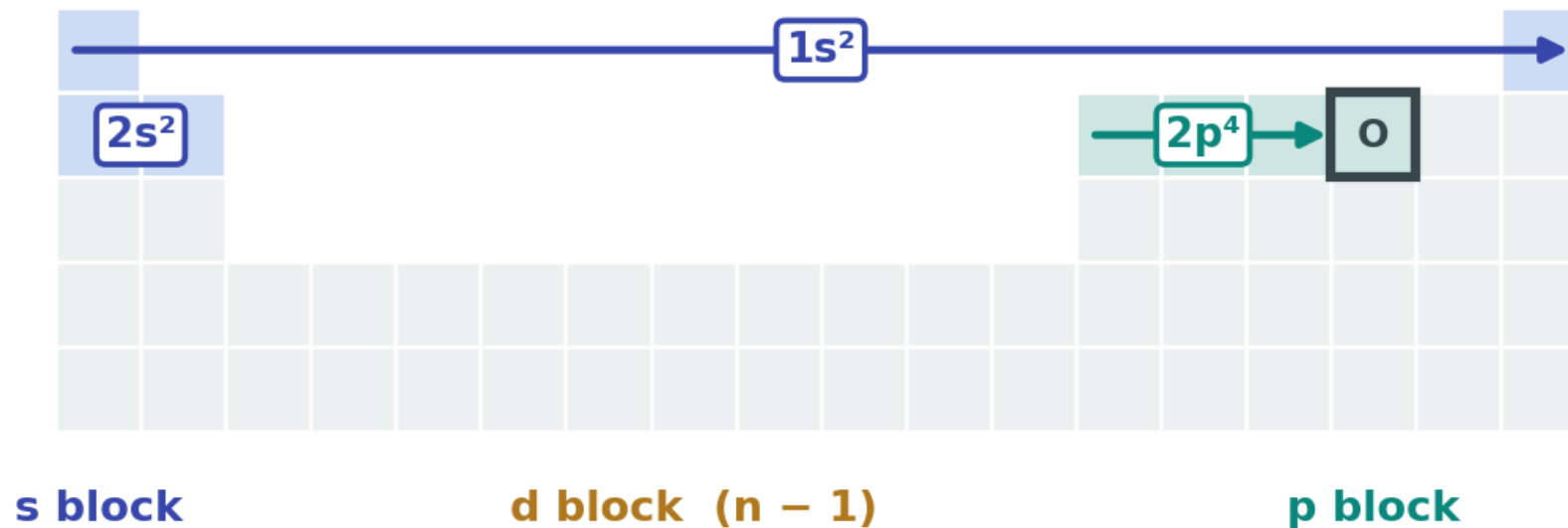
 **Sense check:** The final sublevel need not be full: 2p holds 6 but carries only 4 here. Configurations often end on a partly filled sublevel.

✓ 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

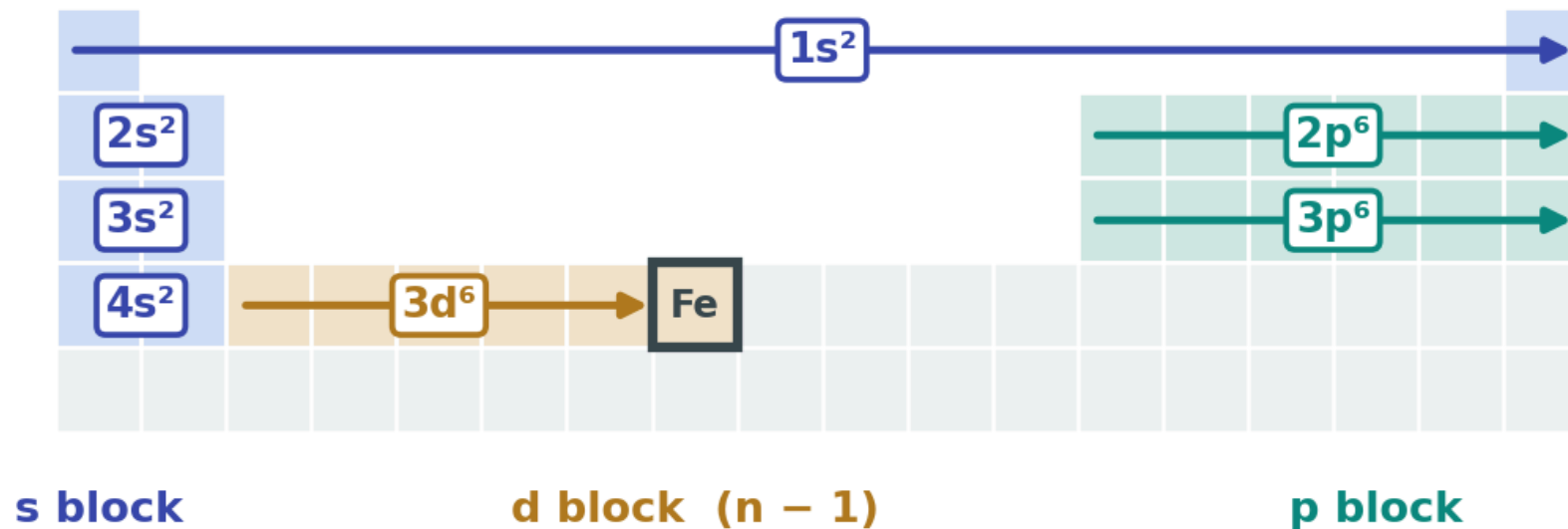
 **Sense check:** Both forms sum to 26, so the count alone cannot flag the error. The fixed order does: 4s fills before 3d.

✓ 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>



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

Where this goes wrong

Filling 3d before 4s. For potassium, 19 electrons, writing $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1$ sums to 19, but 4s is lower in energy than 3d. The ground state is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. 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.

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 $1s^2 2s^2 2p^6$; [Ar] for $1s^2 2s^2 2p^6 3s^2 3p^6$. 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.



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

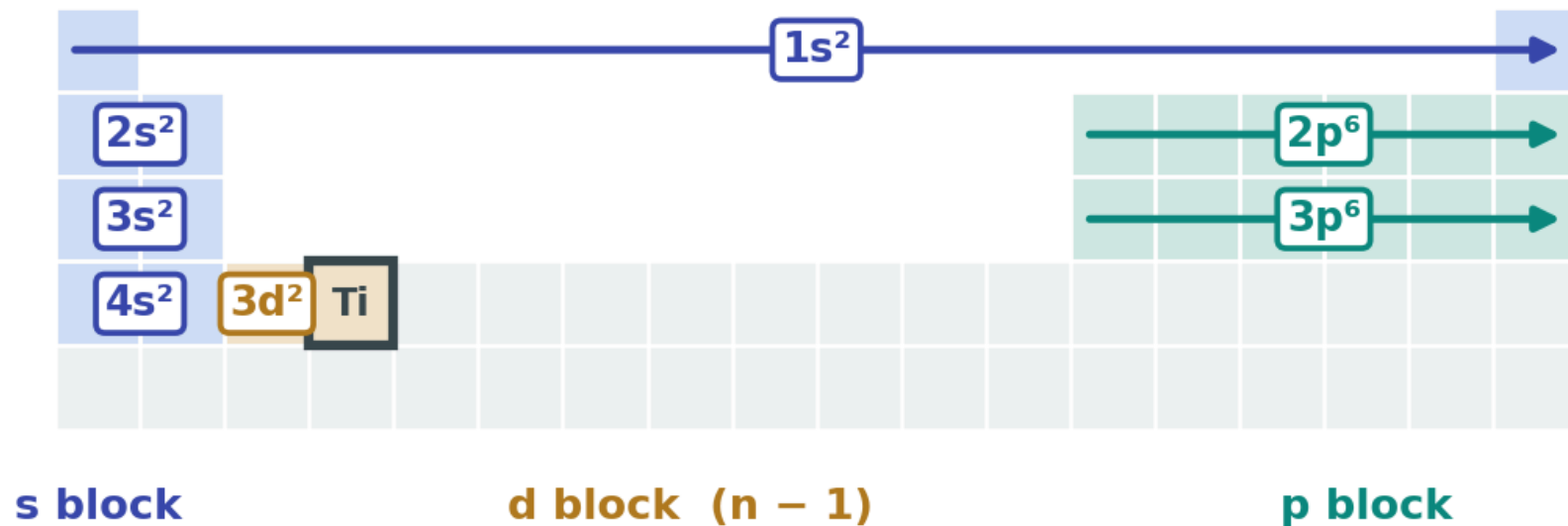
🧠 **Sense check:** [Ar] hides 18 electrons but keeps them: $18 + 2 + 2 = 22$, titanium's atomic number.

✓ Worked example 3: the route on the table



given: atomic number 22 · found: 18 + 2 + 2 = 22 electrons

the route to titanium: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d² · 22 electrons



🧠 **Sense check:** [Ar] covers periods 1 to 3. Period 4 adds 4s², then titanium, the second d-block cell: 3d². ✓

Practice 2

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 2 · answer: B



$18 + 2 + 8 + 4 = 32$ ✓ count · $\times$ 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.



Practice 3

V: atomic number 23
a d-block metal in period 4

Which is the correct ground-state configuration of vanadium?

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

✓ Practice 3 · answer: A



$$18 + 2 + 3 = 23 \checkmark \cdot \text{answer A, vanadium's 23 electrons}$$

After [Ar], 4s fills before 3d: $4s^2$ then $3d^3$. B fills 3d before 4s; it sums to 23 ($18 + 5$) but breaks the order. C stops one electron short: $18 + 2 + 2 = 22$, which is titanium. D skips 3d and fills 4p; it also sums to 23, but 3d is lower in energy than 4p.

🧠 **Sense check:** The count alone passes B and D. The building-up order decides: 4s, then 3d, then 4p.



Check yourself

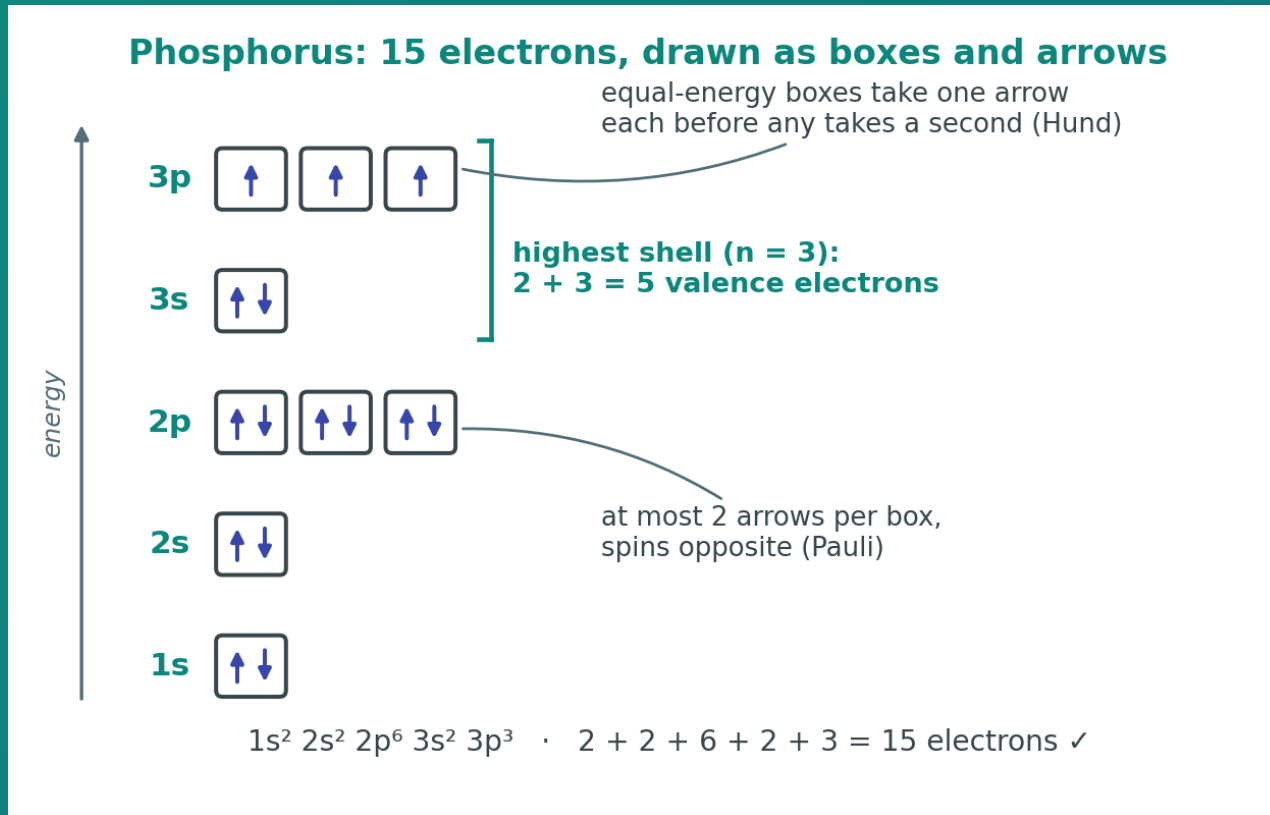
1. Write the ground-state configuration of sulfur (Z 16), and confirm the superscripts sum to 16.
2. In period 4, which fills first, 4s or 3d? Give the noble-gas-core configuration of calcium (Z 20).

A configuration lists each sublevel's total, and a total is a summary. Drawing every orbital as its own box, arrow by arrow, shows what the superscripts hide: which electrons pair up and which sit alone, and which few outermost ones do the atom's chemistry.

8 · Orbital Diagrams & Valence Electrons

Draw the orbital diagram of any element through calcium, 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

Lowest energy fills first

Boxes fill from the lowest energy up: 1s, 2s, 2p, 3s, 3p, then 4s. Through calcium, these six sublevels hold every electron.



capacities $2 + 2 + 6 + 2 + 6 + 2 = 20$ · enough through $Z = 20$

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

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.
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.



Worked example 2: reading a diagram

1s: $\uparrow\downarrow$ · 2s: $\uparrow\downarrow$ · 2p: $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ · 3s: $\uparrow\downarrow$ · 3p: $\uparrow\downarrow$ $\uparrow$ $\uparrow$

given: a complete orbital diagram · wanted: the element and its valence count

Read the diagram back: name the element, then count its valence electrons and its unpaired electrons.

✓ Worked example 2: solution

1s: $\uparrow\downarrow$ · 2s: $\uparrow\downarrow$ · 2p: $\uparrow\downarrow \uparrow\downarrow \uparrow\downarrow$ · 3s: $\uparrow\downarrow$ · 3p: $\uparrow\downarrow \uparrow \uparrow$

each sublevel's arrows become one superscript

Read the arrows into superscripts



$2 + 2 + 6 + 2 + 4 = 16$ electrons · atomic number 16: sulfur

Mark the highest shell

valence: $2 + 4 = 6$ · group check: $16 - 10 = 6$ ✓

core: $2 + 2 + 6 = 10$ · unpaired in 3p: $6 - 4 = 2$

 **Sense check:** Sixteen electrons, six of them valence, two of them unpaired. The diagram carries all three counts at once; the configuration alone hides the last one.

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

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 2 · 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. ✓

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 $3s^23p^1$.

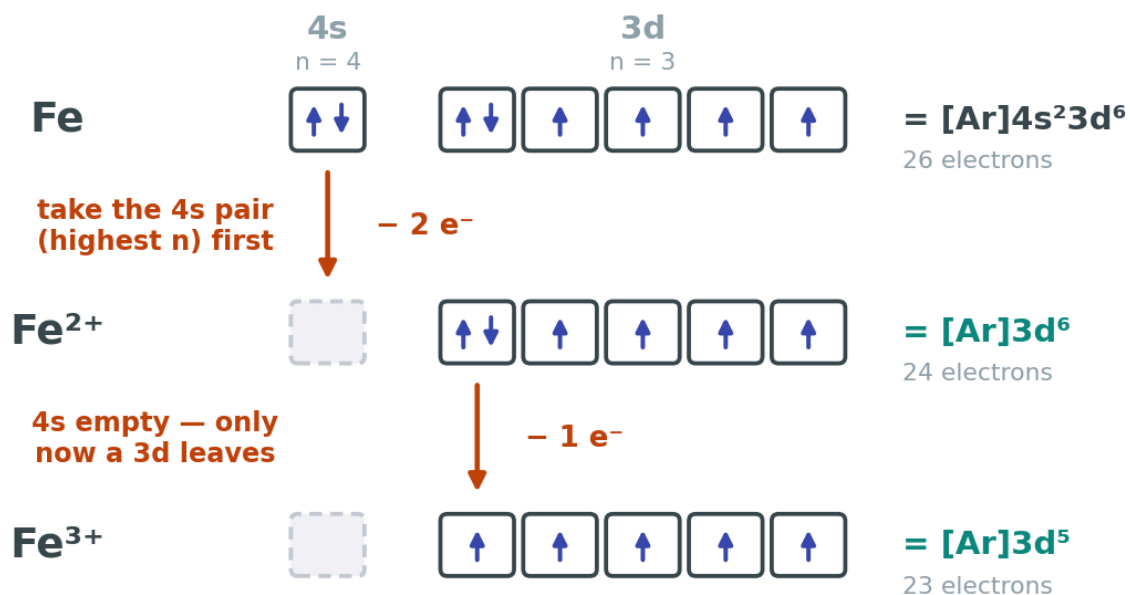
A configuration compresses the diagram; the diagram shows what the superscripts hide: which electrons sit alone. The valence count is the number to keep. The valence electrons sit outermost, so they are what every neighbor meets first: what an atom holds tight, gives up, or shares is decided by exactly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

9 • Electron Configurations of Ions

Write the ground-state configuration of any ion by removing electrons from the highest n first (ns before $(n-1)d$), and use the stable stopping points d^5 and d^{10} to explain a variable-charge metal's common charges.

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.

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

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

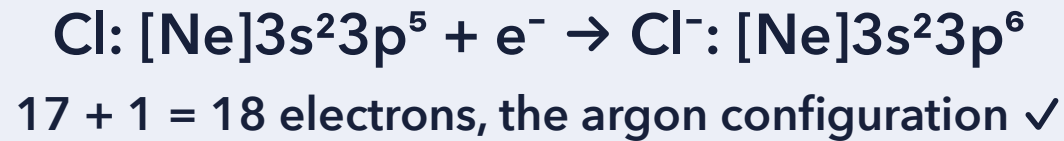
The building-up order fills 4s before 3d. Once 3d is occupied it sinks below 4s, leaving 4s as the outermost shell: the highest n , and the first to leave.

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 $(n-1)d$

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:** ns before (n-1)d. Anions add in filling order.
4. **Check:** protons - electrons = charge.

Worked example 1: iron's cations

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

18 (argon core) + 2 + 6 = 26 electrons $\checkmark$

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 → [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 → 2+ ✓

Step 3 · Remove from the highest n first

Fe³⁺: [Ar]3d⁵

18 + 5 = 23 electrons · 26 – 23 = 3 → 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 2: copper's cations

Cu: atomic number 29 $\rightarrow$ [Ar]4s¹3d¹⁰ (measured ground state)

18 + 1 + 10 = 29 electrons $\checkmark$ · one 4s electron sits in 3d, completing the d subshell

Copper's memorized charges are 1+ and 2+. Build Cu⁺ and Cu²⁺, and state which ion keeps a completely filled 3d.

✓ Worked example 2: solution

Cu: atomic number 29 → [Ar]4s¹3d¹⁰ (measured ground state)

$$18 + 1 + 10 = 29 \text{ electrons } \checkmark$$

Step 1 · Write the neutral atom

Step 2 · Count the charge

Step 3 · Remove from the highest n first

Cu⁺: [Ar]3d¹⁰

$$18 + 10 = 28 \text{ electrons} \cdot 29 - 28 = 1 \rightarrow 1+ \checkmark \cdot \text{completely filled d subshell}$$

Step 3 · Remove from the highest n first

Cu²⁺: [Ar]3d⁹

$$18 + 9 = 27 \text{ electrons} \cdot 29 - 27 = 2 \rightarrow 2+ \checkmark$$

Step 4 · Check

 **Sense check:** Both charges check against 29 protons; Cu⁺ stops where it keeps the extra-stable filled 3d¹⁰.

Take-home: stable configurations set the charges

A variable-charge metal's common ions sit at stable stopping points: an emptied ns, a half-filled d⁵, or a filled d¹⁰. The memorized charges from naming follow from the configurations.

iron: $\text{Fe}^{2+} = [\text{Ar}]3\text{d}^6$ (4s emptied) · $\text{Fe}^{3+} = [\text{Ar}]3\text{d}^5$ (half-filled d)

2+ comes from losing the 4s pair · 3+ reaches the extra-stable d⁵

copper: $\text{Cu}^+ = [\text{Ar}]3\text{d}^{10}$ (filled d) · manganese: $\text{Mn}^{2+} = [\text{Ar}]3\text{d}^5$ (half-filled d)

common charges land on filled or half-filled d subshells

🧩 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+ ✓ · half-filled 3d⁵

⚠ Where this goes wrong

Removing 3d before 4s. $[\text{Ar}]4s^23d^4$ treats "last filled" as "first removed." Highest n leaves first: 4s before 3d. Fe^{2+} is $[\text{Ar}]3d^6$.

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. Fe^{3+} is $[\text{Ar}]3d^5$, five electrons past argon. The stable stops are d^5 and d^{10} , not an octet.

 Practice 1

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 1 · 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. ✓



Practice 2



iron's two memorized charges from the naming rules

Which statement explains why 3+ is one of iron's common charges?

- A. Fe^{3+} has the electron configuration of argon, the nearest noble gas.
- B. Losing the 4s pair and one 3d electron leaves $[\text{Ar}]3d^5$, a half-filled d subshell with extra stability.
- C. Iron loses three 3d electrons first, and 3+ marks the point where 3d empties.
- D. A 3+ charge gives iron exactly eight valence electrons.

✓ Practice 2 · answer: B

Fe^{3+} : $[\text{Ar}]\text{3d}^5$, a half-filled d subshell (answer B)

$26 - 3 = 23$ electrons = $18 + 5$ ✓ · d^5 carries extra stability

A miscounts badly: argon holds 18 electrons, Fe^{3+} holds 23; transition-metal cations do not reach a noble gas. C removes from the wrong shell: 4s empties before 3d, and 3d is far from empty in Fe^{3+} . D invokes an octet that does not apply: the stability here is the half-filled 3d^5 , not eight valence electrons.

 **Sense check:** The memorized 2+ and 3+ are not arbitrary: 2+ empties 4s, and 3+ lands on the half-filled d^5 . ✓



Check yourself

1. Cobalt is $[\text{Ar}]4s^23d^7$. Write the configuration of Co^{2+} and name which electrons left first.
2. Zinc forms only Zn^{2+} , and copper forms Cu^+ . What do the configurations $[\text{Ar}]3d^{10}$ of both ions have in common?

Two neutral atoms use the same stability in their ground states: chromium and copper each move a 4s electron into 3d to reach a half-filled or filled d subshell.

10 · Exceptions: Chromium & Copper

Explain and apply the two building-up exceptions, Cr and Cu, where one 4s electron shifts into 3d to reach an extra-stable half-filled or filled d subshell.

When the order bends

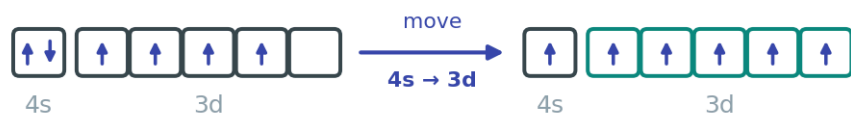
A half-filled or filled d subshell is extra stable

predicted (building-up order)

actual (ground state)

Cr predicted: $[\text{Ar}] 4s^2 3d^4$

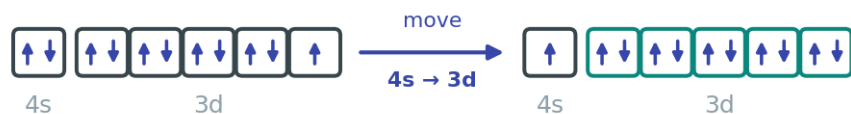
Cr actual: $[\text{Ar}] 4s^1 3d^5$



half-filled 3d — all five singly filled

Cu predicted: $[\text{Ar}] 4s^2 3d^9$

Cu actual: $[\text{Ar}] 4s^1 3d^{10}$



completely filled 3d

The building-up order predicts chromium and copper, but each atom moves one 4s electron into 3d, reaching an extra-stable half-filled or filled d subshell.

The building-up order across the 3d row

Fe: $[\text{Ar}]4s^23d^6$ · Ni: $[\text{Ar}]4s^23d^8$

$18 + 2 + 6 = 26$ · $18 + 2 + 8 = 28$ · 4s fills first, then 3d gains one electron per element

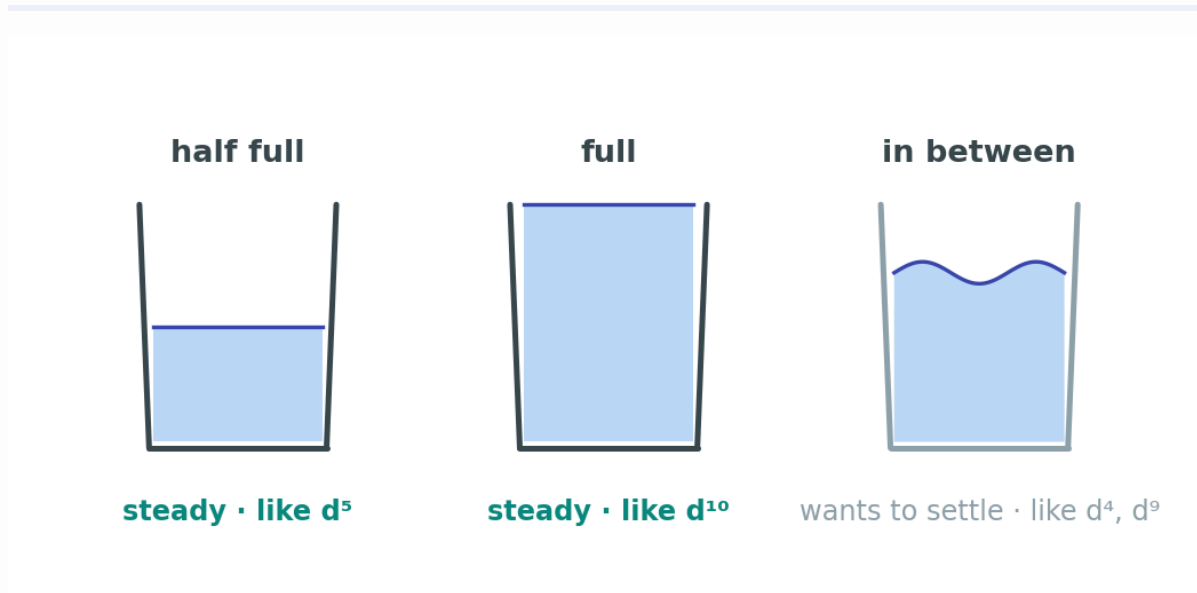
Both follow the building-up order. From Sc to Zn, the order predicts every 3d-row configuration. Eight of the ten measured ground states match. Cr and Cu do not.

Two exceptions to memorize: Cr and Cu

The building-up order is a reliable guide, but two atoms in our range do better by rearranging: chromium and copper. In each, one electron leaves 4s for 3d.

predicted Cr: $[\text{Ar}]4s^23d^4 \rightarrow$ actual: $[\text{Ar}]4s^13d^5$
 $18 + 2 + 4 = 24$ and $18 + 1 + 5 = 24$ · only $4s^13d^5$ is observed

Half full or full sits steady



A glass poured exactly half full or exactly full sits steady; anything between sloshes. A d subshell is the same: d^5 and d^{10} carry extra stability, worth one 4s electron.

memory hook: the glass, half full or full, sits steady
 d^5 (half-filled) and d^{10} (filled) are the steady stopping points

A move, not a new electron

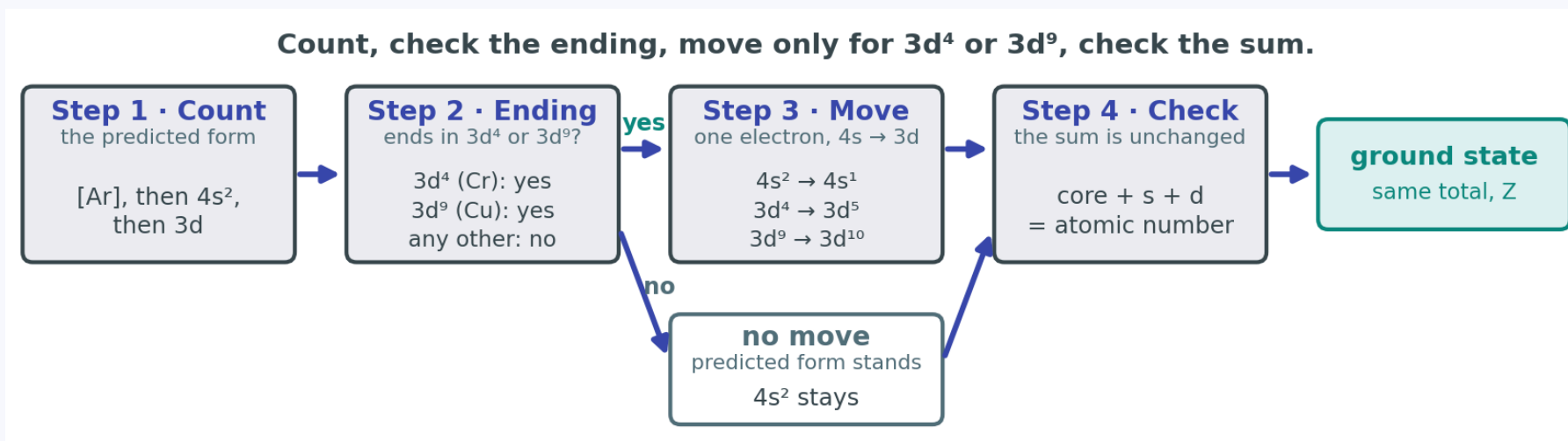
The electron relocates; the count never changes. It always comes out of 4s, never the core, and 4s drops to 4s¹, never to 4s⁰.

$[\text{Ar}]4s^23d^4 \rightarrow [\text{Ar}]4s^13d^5$: still 24 electrons
one electron changes seats · the total and the element stay the same



The method

1. **Count the electrons** and write the predicted form.
2. **Ends in $3d^4$ or $3d^9$?** The exception applies: Cr, Cu.
3. **Move one electron, $4s \rightarrow 3d$.** Only on a yes: $4s^2$ becomes $4s^1$.
4. **Check the sum is unchanged.**





Worked example 1: chromium

Step 1 · Count the electrons

Cr: atomic number 24

24 electrons to place · a d-block metal in period 4

Write what the building-up order predicts, then apply the exception. After [Ar], 6 electrons remain.

✓ Worked example 1: solution

Step 1 · Write the predicted form

predicted: $[\text{Ar}]4s^23d^4$

$18 + 2 + 4 = 24$ · correct count, but not the ground state

Step 2 · Ends in $3d^4$: the exception applies

Step 3 · Move one electron, $4s \rightarrow 3d$

actual: $[\text{Ar}]4s^13d^5$

six unpaired electrons: 1 in 4s, 5 in 3d

Step 4 · Check the sum is unchanged

$18 + 1 + 5 = 24$ ✓

same 24 electrons as the predicted form · chromium either way

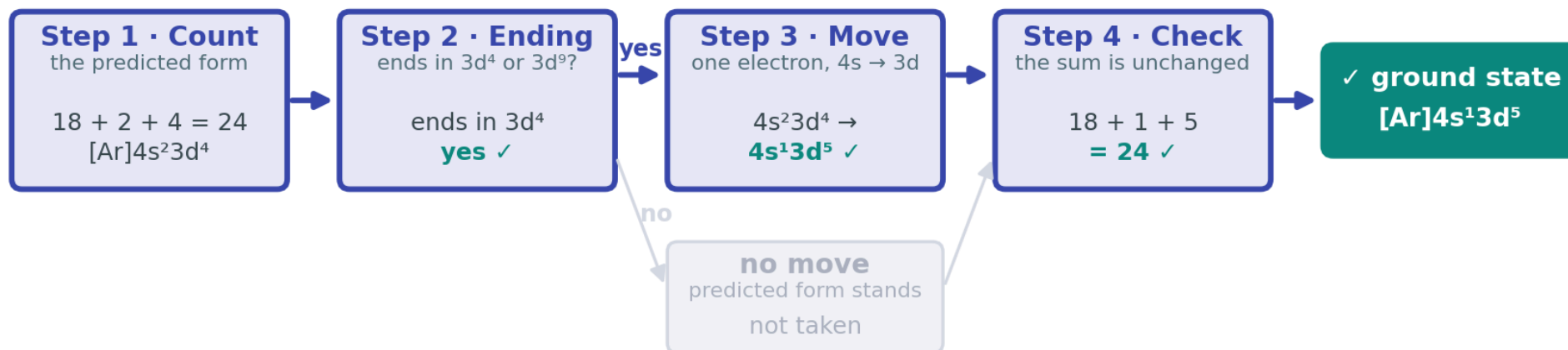
🧠 **Sense check:** Both forms total 24; only the rule flags the swap: $3d^5$ beats $3d^4$.

✓ Worked example 1: the route



given: atomic number 24 · found: $18 + 1 + 5 = 24$ electrons

chromium, $Z = 24$



🧠 **Sense check:** The predicted $3d^4$ answers yes at Step 2, so the route runs through Step 3. One 4s electron moves, and the sum stays 24. ✓



Worked example 2: copper

Step 1 · Count the electrons

Cu: atomic number 29

29 electrons to place · a d-block metal in period 4

Copper does what chromium does, one step further along. Write the predicted form, spot the ending, and apply the exception.

✓ Worked example 2: solution

Step 1 · Write the predicted form

predicted: $[\text{Ar}]4s^23d^9$

$18 + 2 + 9 = 29$ · one electron short of a filled d

Step 2 · Ends in $3d^9$: the exception applies

Step 3 · Move one electron, $4s \rightarrow 3d$

actual: $[\text{Ar}]4s^13d^{10}$

the glass fills to the brim · 4s drops to $4s^1$, never to $4s^0$

Step 4 · Check the sum is unchanged

$$18 + 1 + 10 = 29 \checkmark$$

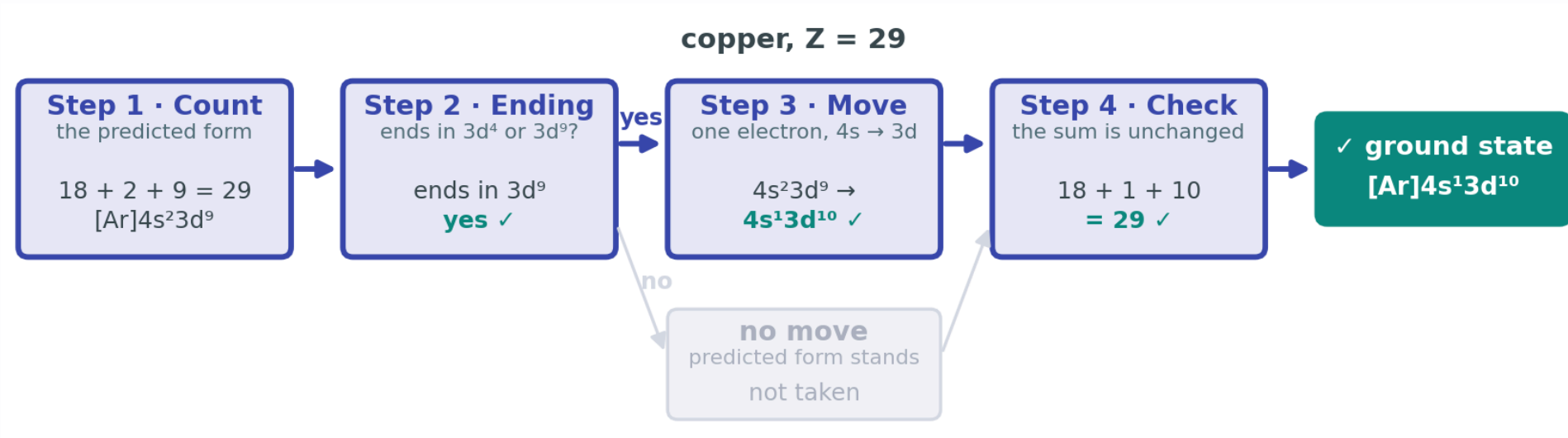
a move, not a change in count

🧠 **Sense check:** Keeping $4s^2$ and filling $3d$ would claim 30 electrons, one too many: the electron must come from $4s$.

✓ Worked example 2: the route



given: atomic number 29 · found: $18 + 1 + 10 = 29$ electrons



Sense check: The predicted $3d^9$ answers yes at Step 2, the same route chromium takes. The moved electron fills $3d$ to $3d^{10}$. ✓



Guided example: iron on a study sheet

Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^7$

iron, atomic number 26 · a study sheet's ground-state configuration

Run the four method steps on this entry. Decide whether it is iron's measured ground state.

✓ Guided example: solution

Step 1 • Count the electrons

sheet: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^7$ • predicted: $[\text{Ar}] 4s^2 3d^6$
 sheet $2 + 2 + 6 + 2 + 6 + 1 + 7 = 26$ ✓ • predicted $18 + 2 + 6 = 26$ ✓

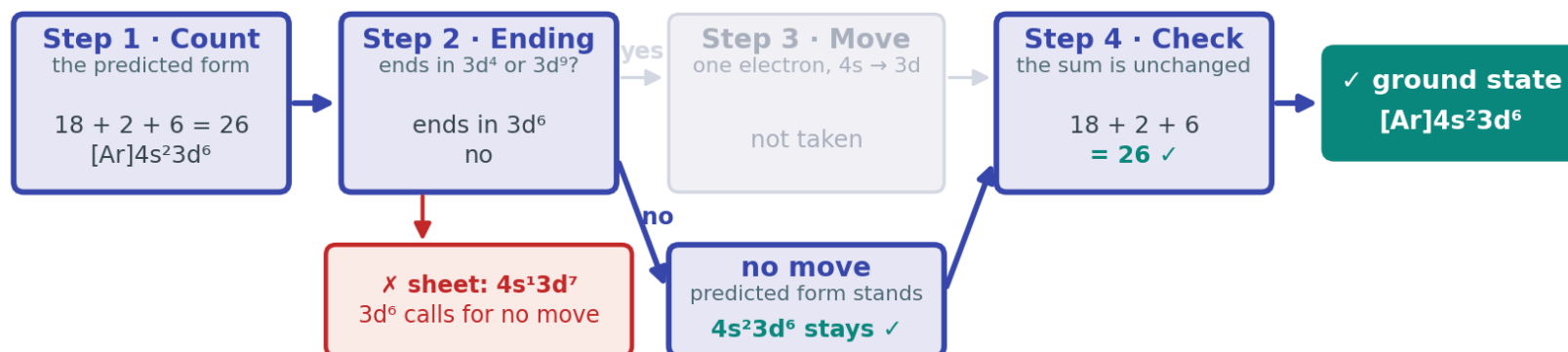
Step 2 • Ends in $3d^4$ or $3d^9$?

Step 3 • Move one electron, $4s \rightarrow 3d$

Step 4 • Check the sum is unchanged

🧠 **Sense check:** Iron's ground state is the predicted $[\text{Ar}] 4s^2 3d^6$. The sheet's $3d^7$ is no steady point: only a predicted $3d^4$ or $3d^9$ moves an electron.

iron, $Z = 26$: the sheet gives $\dots 4s^1 3d^7$



Your turn: spot the exceptions

predicted: V [Ar]4s²3d³ · Cr [Ar]4s²3d⁴ · Ni [Ar]4s²3d⁸ · Cu [Ar]4s²3d⁹
which endings trigger the move? only 3d⁴ and 3d⁹

element	predicted ending	actual ground state
V	3d ³	_____
Cr	3d ⁴	_____
Ni	3d ⁸	_____
Cu	3d ⁹	_____

V: [Ar]4s²3d³ (no move) · Cr: [Ar]4s¹3d⁵ · Ni: [Ar]4s²3d⁸ (no move) · Cu: [Ar]4s¹3d¹⁰
d³ and d⁸ are not one short of steady · only 3d⁴ and 3d⁹ borrow from 4s

⚠ Where this goes wrong

Adding an electron instead of moving one. $[\text{Ar}]4s^23d^5$ keeps 4s full: that is $18 + 2 + 5 = 25$ electrons, one too many. Empty 4s to $4s^1$; the electron relocates, it is not created.

Deleting the 4s electron. $[\text{Ar}]3d^5$ drops 4s entirely: $18 + 5 = 23$, one short. The electron moves into 3d, it does not vanish.

Trusting the predicted form. $[\text{Ar}]4s^23d^4$ has the right count, but Cr's measured ground state is $[\text{Ar}]4s^13d^5$.

Over-applying the rule. Only a d subshell one short of half or full borrows from 4s: V ($3d^3$) and Ni ($3d^8$) keep $4s^2$.



Practice 1

Mn: atomic number 25

a d-block metal in period 4 · 25 electrons to place

Which is the ground-state configuration of manganese?

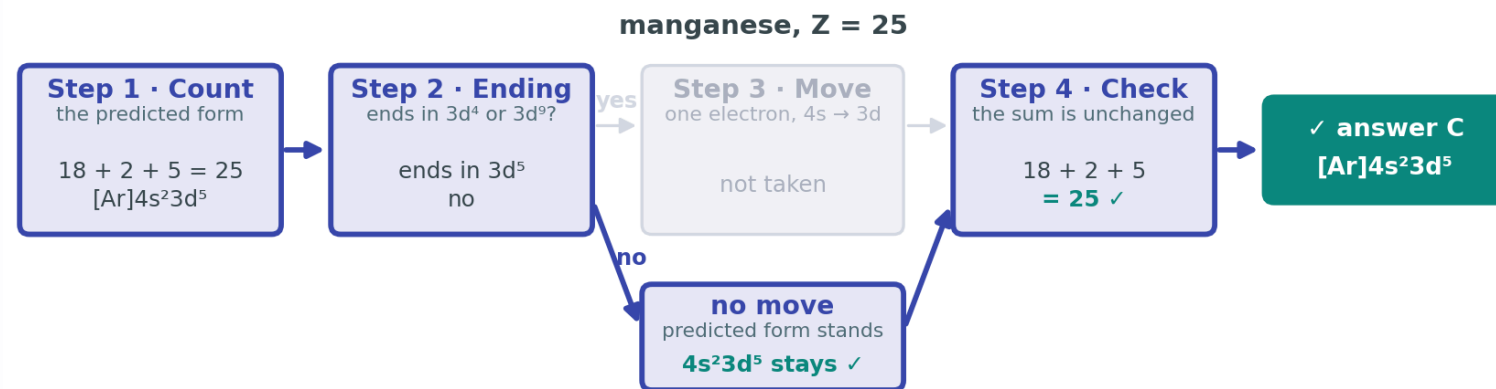
- A. $[\text{Ar}] 4s^1 3d^5$
- B. $[\text{Ar}] 4s^1 3d^6$
- C. $[\text{Ar}] 4s^2 3d^5$
- D. $[\text{Ar}] 3d^7$

✓ Practice 1 · answer: C

[Ar]4s²3d⁵ (answer C)

$18 + 2 + 5 = 25$ ✓ · the plain order lands 3d exactly half full, nothing to borrow

🧠 **Sense check:** Manganese reaches d⁵ by the plain order. The exception fires only for a predicted 3d⁴ or 3d⁹. ✓





Practice 2

Cu ($Z = 29$) · Zn ($Z = 30$)
two period-4 d-block metals, side by side

In which atom's ground state does a 4s electron actually move into 3d, and why?

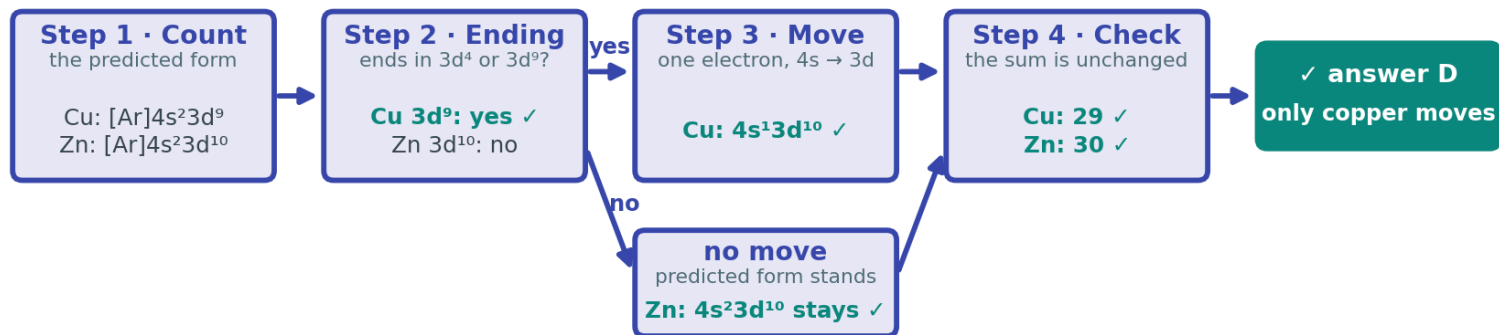
- A. Zinc: its 3d is full, so one 4s electron moves to keep it that way.
- B. Both: every period-4 metal past chromium borrows from 4s.
- C. Copper: one 4s electron moves so that 3d reaches the half-filled $3d^5$.
- D. Copper: its predicted $3d^9$ is one short of full, so one 4s electron fills 3d to $3d^{10}$.

✓ Practice 2 · answer: D

Cu: $[\text{Ar}]4s^23d^9 \rightarrow [\text{Ar}]4s^13d^{10}$ · **Zn:** $[\text{Ar}]4s^23d^{10}$ stays
 $18 + 1 + 10 = 29 \checkmark$ · $18 + 2 + 10 = 30 \checkmark$ · only a predicted $3d^4$ or $3d^9$ borrows (answer D)

🧠 **Sense check:** Write the predicted ending first. $3d^9$ is one short of full, so copper moves; $3d^{10}$ is already there, so zinc does not. ✓

copper (Z = 29) and zinc (Z = 30)





Practice 3

Ti (Z = 22) · Cr (Z = 24) · Ni (Z = 28) · Cu (Z = 29) · Zn (Z = 30)

five period-4 d-block metals, each written in full

A study sheet lists these ground-state configurations. Which entry contains an error?

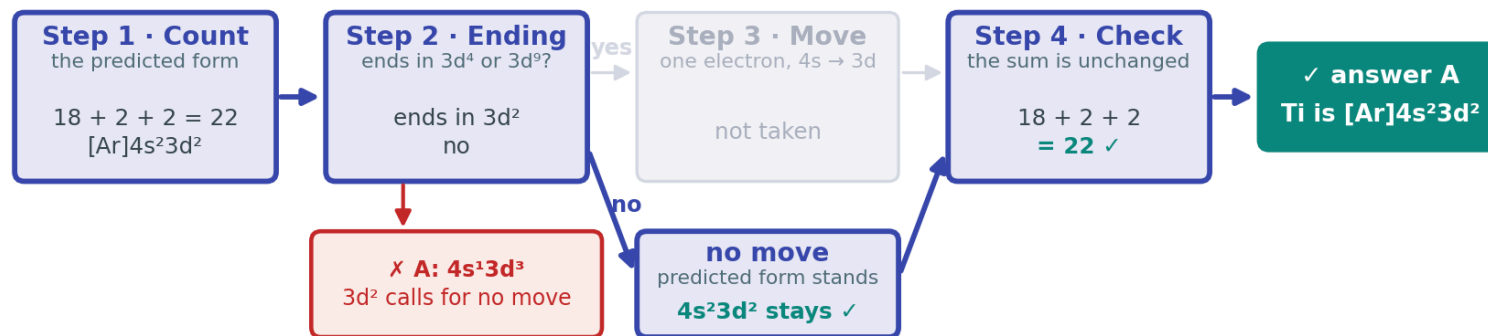
- A. Ti: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^3$
- B. Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$
- C. Ni: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$
- D. Cu: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$
- E. Zn: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$

✓ Practice 3 · answer: A

Ti: ...4s¹3d³ is the error · ground state [Ar]4s²3d² (answer A)
 2 + 2 + 6 + 2 + 6 + 1 + 3 = 22 counts right · predicted 18 + 2 + 2 = 22 ends 3d²: no move

🧠 **Sense check:** Every entry sums to its Z, so the count clears all five. The ending decides: only a predicted 3d⁴ or 3d⁹ moves an electron, and Ti's 3d² does not. ✓

titanium, Z = 22: choice A gives ...4s¹3d³





Practice 4

Cr ($Z = 24$)

a period-4 d-block metal, in its ground state

How many electrons in a ground-state chromium atom have $n = 3$?

- A. 5
- B. 8
- C. 12
- D. 13

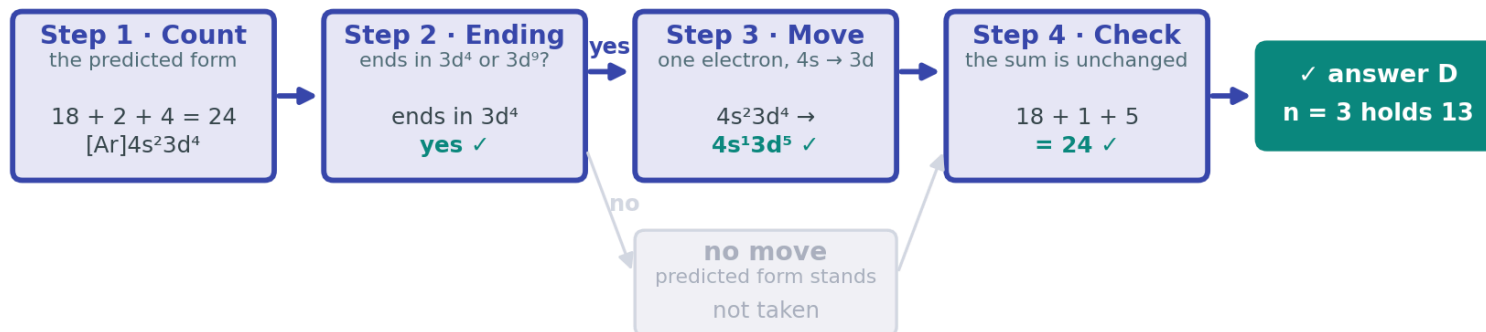
✓ Practice 4 · answer: D



$$n = 3: 3s^2 + 3p^6 + 3d^5 = 2 + 6 + 5 = 13 \text{ (answer D) · shells: } 2 + 8 + 13 + 1 = 24 \checkmark$$

 **Sense check:** The digit in front of each subshell is n. Chromium's move leaves only 4s¹ in n = 4 and puts 13 electrons in the third shell. ✓

chromium, Z = 24, then the n = 3 electrons





Practice 5

Cu ($Z = 29$)

a period-4 d-block metal, in its ground state

In ground-state copper, how many electrons occupy subshells with $\ell = 0$?

- A. 1
- B. 6
- C. 7
- D. 8

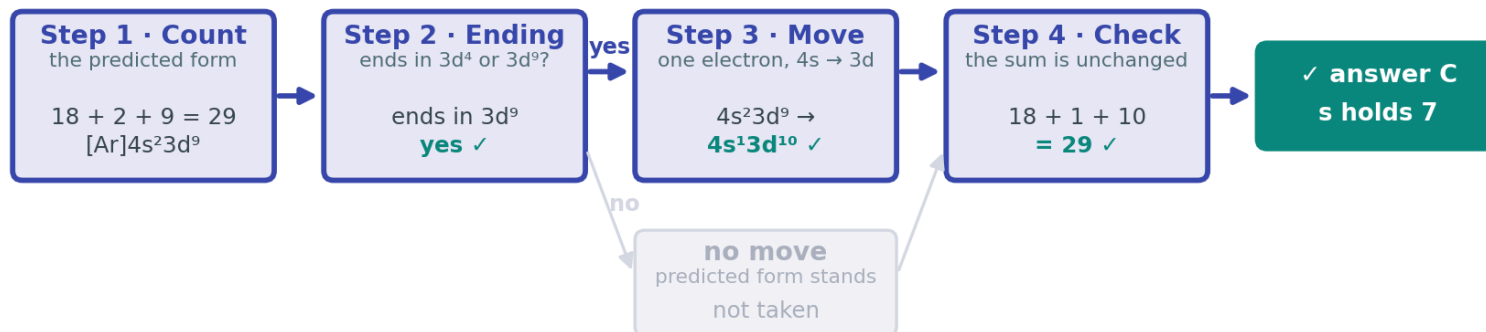
✓ Practice 5 · answer: C



$$\ell = 0 \text{ is s: } 1s^2 + 2s^2 + 3s^2 + 4s^1 = 2 + 2 + 2 + 1 = 7 \text{ (answer C) } \cdot s 7 + p 12 + d 10 = 29 \checkmark$$

🧠 **Sense check:** $\ell = 0$ labels every s subshell, core ones included. Copper's move takes one electron out of 4s, so the s count is 7, not 8. ✓

copper, $Z = 29$, then the s ($\ell = 0$) electrons





Practice 6

Cr (Z = 24) · Co (Z = 27) · Ga (Z = 31) · As (Z = 33)

four period-4 elements

Which element and ground-state electron configuration pair contains an error?

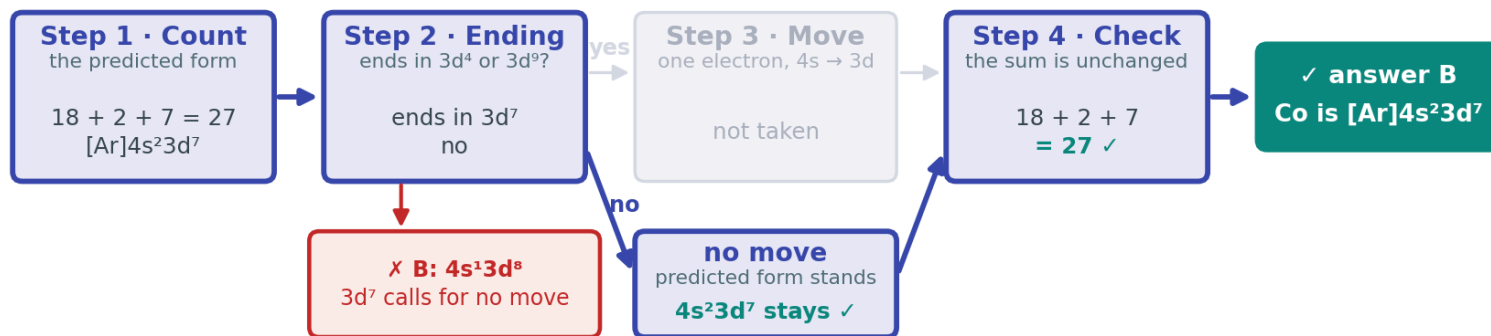
- A. Cr: [Ar] 4s¹ 3d⁵
- B. Co: [Ar] 4s¹ 3d⁸
- C. Ga: [Ar] 4s² 3d¹⁰ 4p¹
- D. As: [Ar] 4s² 3d¹⁰ 4p³

✓ Practice 6 · answer: B

Co: predicted $[\text{Ar}]4s^23d^7$ is the ground state · $[\text{Ar}]4s^13d^8$ is the error (answer B)
 $18 + 2 + 7 = 27$ ✓ · $18 + 1 + 8 = 27$ counts right, but $3d^7$ is not one short of half or full

🧠 **Sense check:** Every pair counts right, so the sum cannot flag the error. The ending does: only a predicted $3d^4$ or $3d^9$ moves an electron. ✓

cobalt, $Z = 27$: choice B gives $\dots 4s^13d^8$



Check yourself

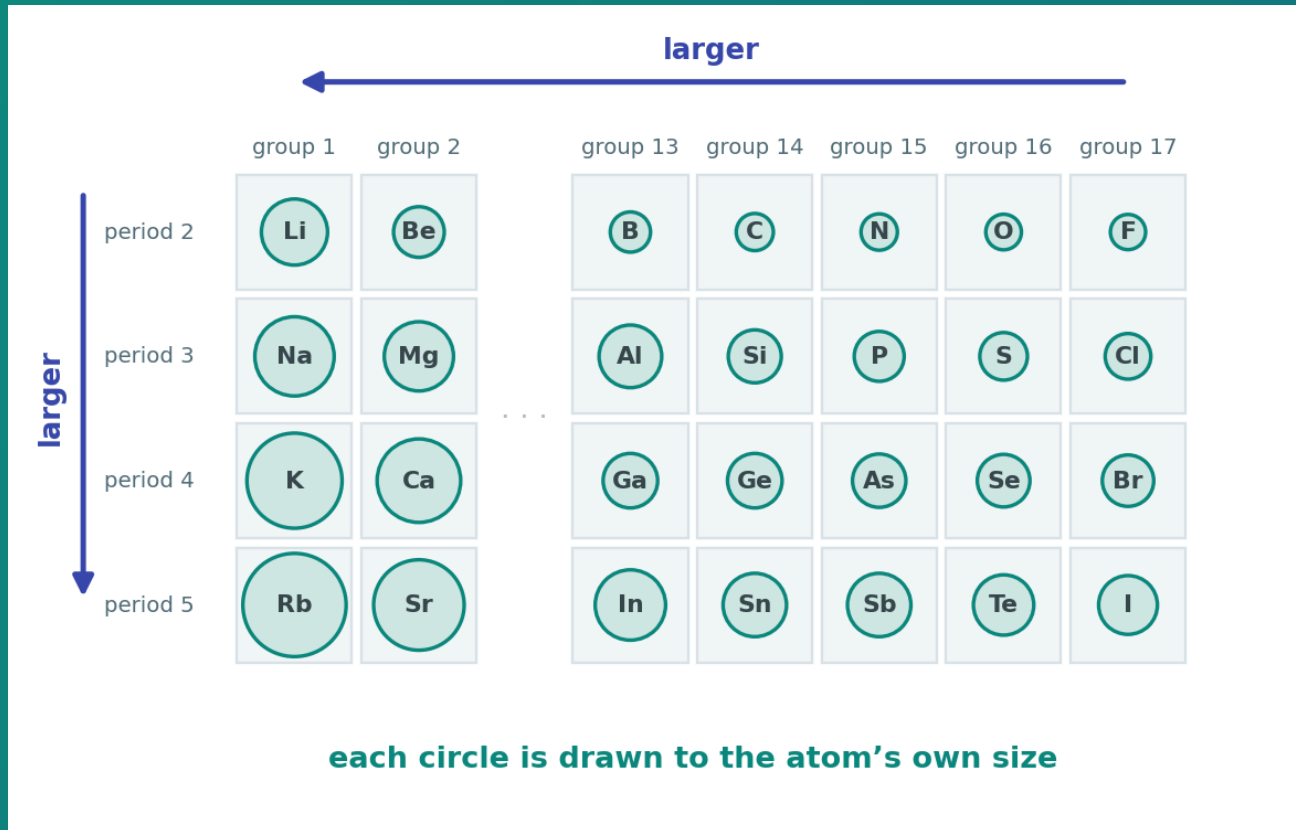
1. Write chromium's ground-state configuration, and confirm the superscripts sum to 24.
2. Copper's $[\text{Ar}]4s^13d^{10}$: how many unpaired electrons does it have, and why is that fewer than chromium's?

These two exceptions, Cr and Cu, are the ones to memorize for this course. The stability behind them, a d subshell exactly half full or exactly full, is the same d^5 and d^{10} steadiness that steers a variable-charge metal toward its common ions.

11 • 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.

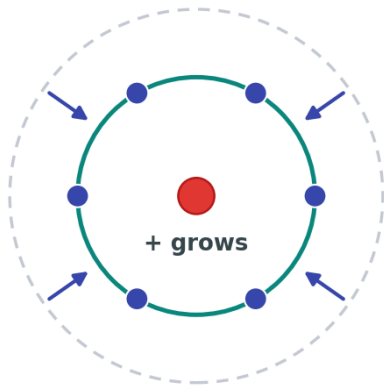
An atom's place predicts its size



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.

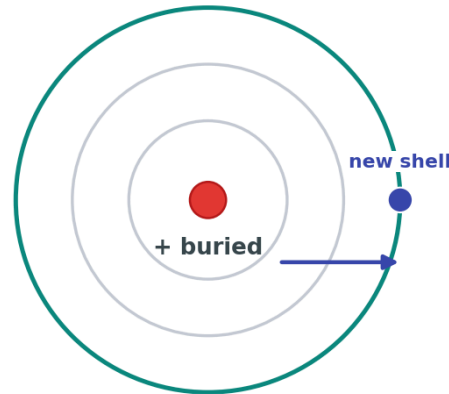
Two forces set every trend

Across a period



more protons pull the same shell in → smaller

Down a group



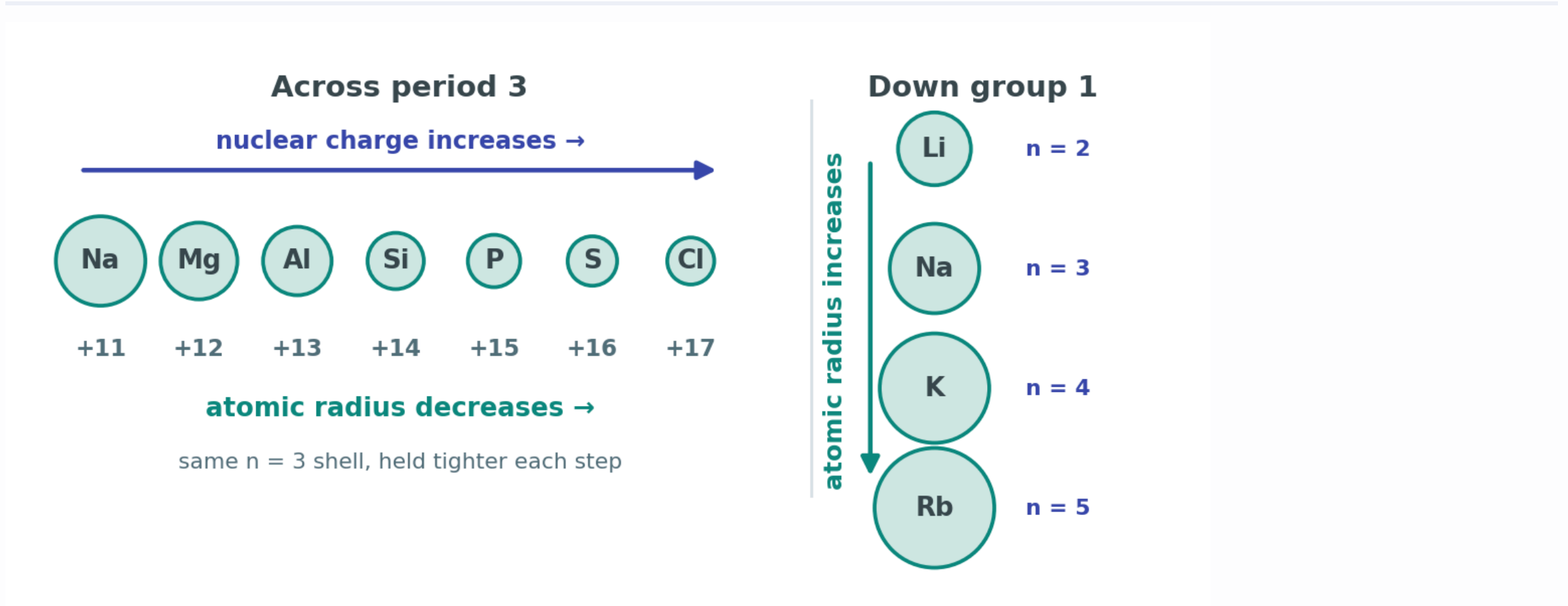
each new shell seats electrons farther out → larger

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.

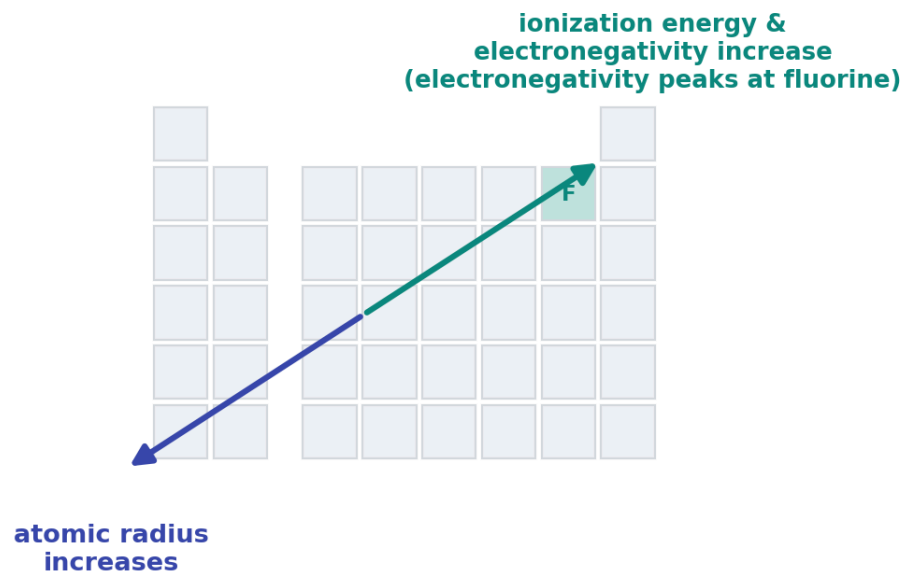
across period 2: Li 520 → C 1086 → F 1681 kJ/mol

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

Electron affinity

Electron affinity is the energy change when a gaseous atom gains an electron. Most nonmetals release energy, and the release grows toward the halogens, one electron short of an octet.

$\text{Na} + \text{e}^- \rightarrow \text{Na}^-$ releases 53 kJ/mol · $\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$ releases 349 kJ/mol
across period 3 the release grows: the added electron lands closer to a stronger nucleus



The method

1. **Place the two elements.** Same period, or same group?
2. **Name the trend** for that direction: radius, ionization energy, electronegativity, metallic character, or electron affinity.
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. ✓

Take-home: radius and ionization energy run opposite

down group 1: $\text{Li} \rightarrow \text{Na} \rightarrow \text{K}$, the atoms grow larger
atomic radius: $152 \rightarrow 186 \rightarrow 227$ pm: each new shell adds size

down group 1: $\text{Li} \rightarrow \text{Na} \rightarrow \text{K}$, the electron leaves more easily
ionization energy: $520 \rightarrow 496 \rightarrow 419$ kJ/mol: the outer electron sits farther out

The larger atom holds its outer electron more loosely. A big radius and a high ionization energy never belong to the same atom.

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

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 1 · 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 2

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 2 · 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. ✓

Exceptions to the trends

Two dips break the ionization-energy climb: where p starts (group 13) and where p first pairs (group 16). Noble gases and group 2 metals accept an electron poorly: those shells and s sublevels are full.

period 2 climb, with dips: Be 899 > B 801 and N 1402 > O 1314 kJ/mol

B's lone 2p electron leaves more easily than Be's full 2s · O's first 2p pair repels, one leaves more easily

gaining an electron: Cl releases 349 kJ/mol · Ne and Mg release none

Ne: full $n = 2$ shell · Mg: full 3s · the next electron must start a higher sublevel



Check yourself

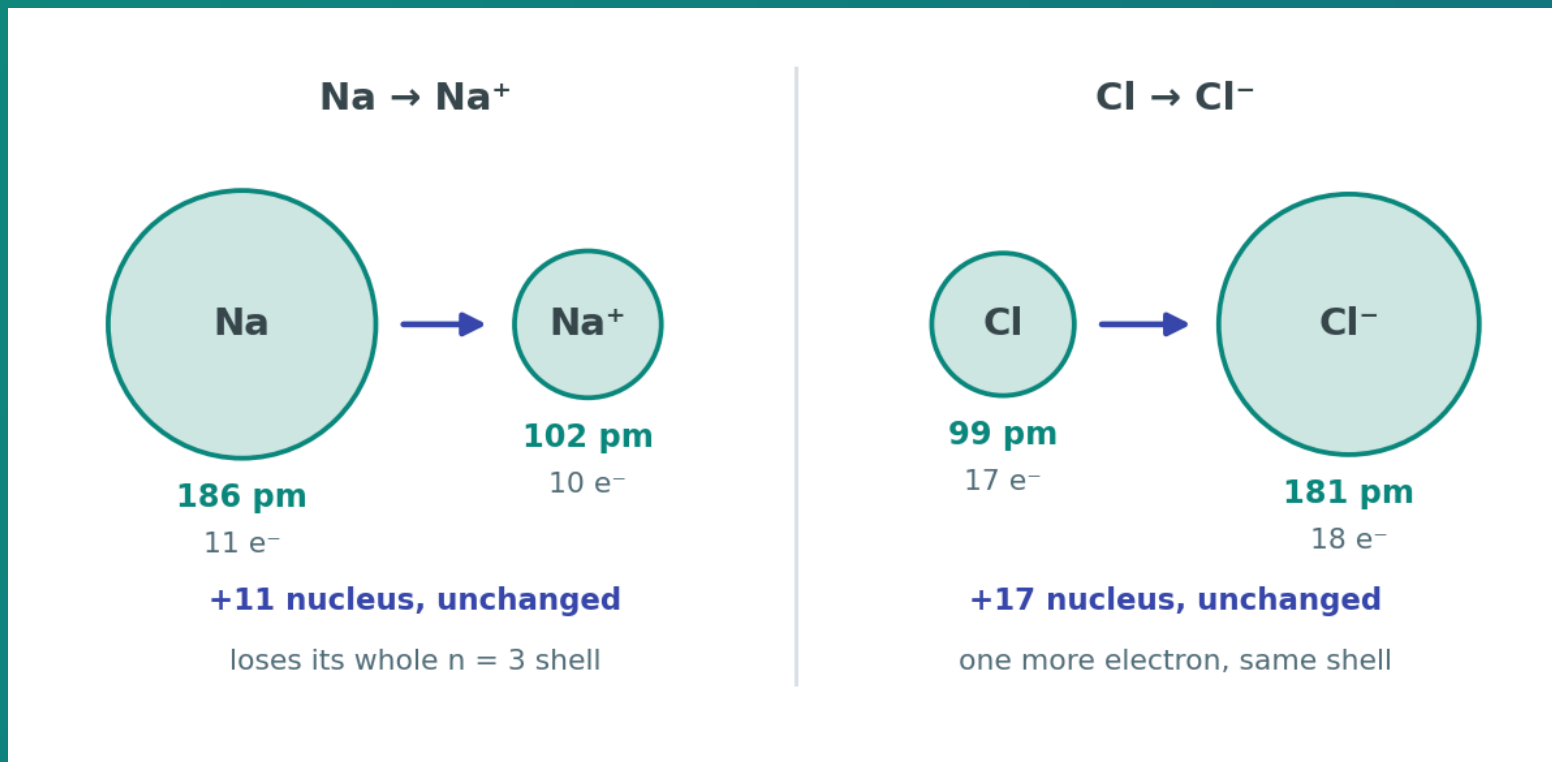
1. Chlorine and iodine sit in the same group. Which holds its electrons more tightly, and does that make it the larger or the smaller atom?
2. Across period 3, phosphorus lies left of chlorine. Rank their atomic radius and their ionization energy. Do the two rankings point the same way?

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.

12 · 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 $n = 2$

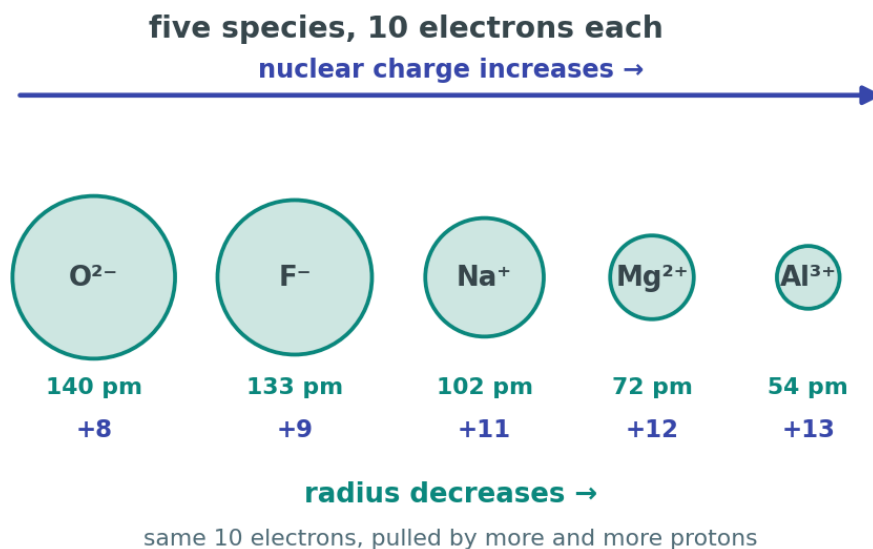
An anion is larger than its atom

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

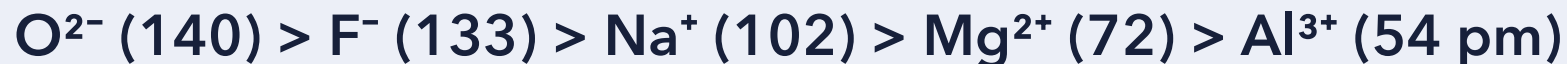


electrons: $8 + 2 \cdot 9 + 1 \cdot 11 - 1 \cdot 12 - 2 \cdot 13 - 3 = 10$ each

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. ✓



Practice 3

Fe^{2+} and Fe^{3+}

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

Iron forms both ions. Which ion is larger, and why?

- A. Fe^{3+} : a higher positive charge makes a larger ion.
- B. Fe^{2+} : the same +26 nucleus holds 24 electrons instead of 23.
- C. Neither: both keep 26 protons, and the nucleus alone sets the size.
- D. Neither: both lost their 4s electrons, so both end in the $n = 3$ shell.

✓ Practice 3 · answer: B

Fe (126) > Fe²⁺ (78) > Fe³⁺ (65 pm) (answer B)

Fe²⁺: [Ar]3d⁶, 26 – 2 = 24 e⁻ · Fe³⁺: [Ar]3d⁵, 26 – 3 = 23 e⁻ · same +26 nucleus

A ran the pull backwards: the extra positive charge means one fewer electron under the same +26 pull, so Fe³⁺ is held tighter and is smaller. C held size to the nucleus alone: the protons match, but size follows pull per electron, and 23 electrons feel more of it than 24. D stopped at the shell: both ions do end in n = 3, yet Fe³⁺ has also lost a 3d electron, so the same pull grips one fewer electron.

 **Sense check:** Same element, fewer electrons, smaller species: Fe > Fe²⁺ > Fe³⁺. ✓



Practice 4



atomic numbers: Te 52 · Cs 55 · Ba 56 · Mg 12

Rank these four ions by radius, largest first.

- A. $\text{Mg}^{2+} > \text{Te}^{2-} > \text{Cs}^+ > \text{Ba}^{2+}$
- B. $\text{Cs}^+ > \text{Ba}^{2+} > \text{Mg}^{2+} > \text{Te}^{2-}$
- C. $\text{Te}^{2-} > \text{Cs}^+ > \text{Ba}^{2+} > \text{Mg}^{2+}$
- D. $\text{Ba}^{2+} > \text{Cs}^+ > \text{Te}^{2-} > \text{Mg}^{2+}$
- E. $\text{Te}^{2-} = \text{Cs}^+ = \text{Ba}^{2+} > \text{Mg}^{2+}$

✓ Practice 4 · answer: C

Te^{2-} (221) > Cs^+ (167) > Ba^{2+} (135) > Mg^{2+} (72 pm) (answer C)

electrons: $52 + 2 \cdot 55 - 1 \cdot 56 - 2 = 54$ each, outer shell $n = 5$ · Mg^{2+} : $12 - 2 = 10$, outer shell $n = 2$

A ranked all four by proton count: fewer protons means larger only when the electron count matches, and Mg^{2+} sits three shells inside the others. B ranked the parent atoms (Cs 265 > Ba 222 > Mg 160 > Te 138 pm), but ionization rewrites those sizes. D ran the pull backwards: +56 grips the same 54 electrons tightest, so Ba^{2+} is the smallest of the trio. E called the isoelectronic trio equal: same electron count, different nuclei, different sizes.

 **Sense check:** Sort by electron count first, then rank each isoelectronic group by nuclear charge. ✓



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.

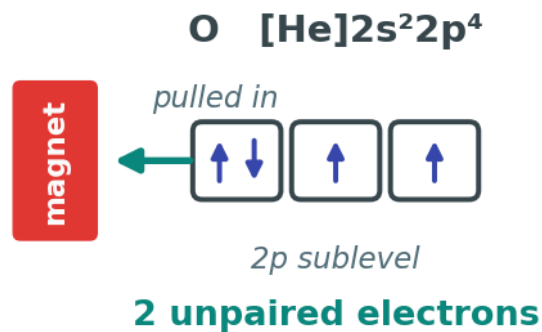
13 · Paramagnetism & Diamagnetism

Decide whether an atom or ion is paramagnetic or diamagnetic by writing its electron configuration, filling the orbital boxes, and counting unpaired electrons.

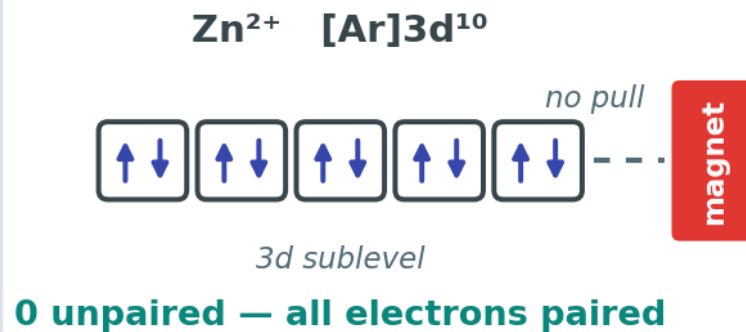
A magnet can sort atoms

One unpaired electron makes an atom paramagnetic

paramagnetic — attracted



diamagnetic — not attracted



Bring a magnet near a sample and some atoms pull toward it while others ignore it. The difference is one question: is any electron left unpaired?

Orbital diagrams you already draw

N: 2p $\uparrow$ $\uparrow$ $\uparrow$ $\cdot$ **Ne: 2p** $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$

nitrogen: three electrons sit alone $\cdot$ **neon: every electron has a partner**

A box holds at most two electrons, spinning opposite ways. Hund's rule seats one per box before any pair forms. Magnetism reads one more fact off the diagram: whether any arrow stands alone.

An unpaired electron is a tiny magnet

Every electron spins, and a spinning charge is a magnet. Paired electrons spin opposite ways and cancel. An unpaired electron has nothing to cancel it, so the atom keeps a net magnetic pull.

≥ 1 unpaired electron $\rightarrow$ paramagnetic: attracted to a magnet

all electrons paired $\rightarrow$ diamagnetic: weakly repelled

memory hook: hear the di in diamagnetic as two

diamagnetic: every electron sits in a pair of two · paramagnetic: at least one sits alone

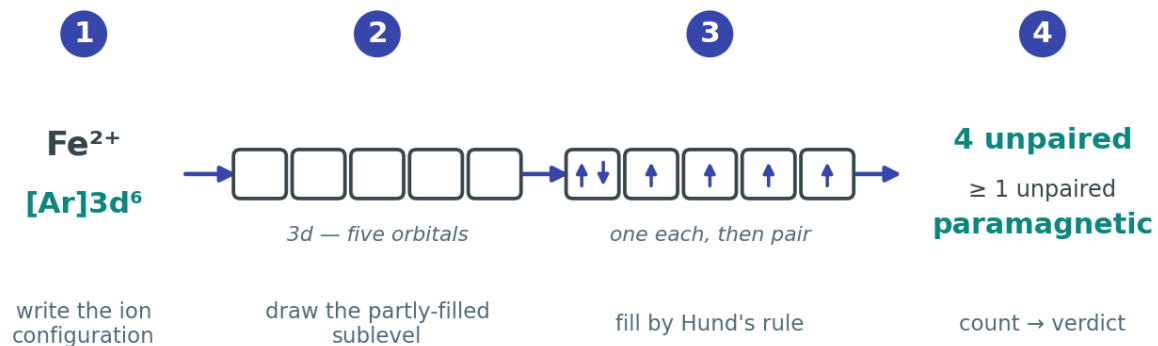
Diamagnetism is the strict case

A paramagnetic species needs only a single unpaired electron. A diamagnetic one needs every electron paired: no exceptions. So the real work is counting: fill the orbital boxes and look for any lone arrow.

any lone arrow → paramagnetic
not one lone arrow anywhere → diamagnetic

The method

From configuration to magnetic verdict — count the unpaired electrons



1. **Write the configuration** (ions: neutral atom first, then remove highest n).
2. **Draw the partly-filled sublevel** as boxes.
3. **Fill singly before pairing** (Hund's rule).
4. **Count lone arrows**: any $\rightarrow$ paramagnetic; none $\rightarrow$ diamagnetic.



Worked example 1: oxygen

Step 1 • Write the configuration

O: atomic number 8

$$1s^2 2s^2 2p^4 = [\text{He}] 2s^2 2p^4 \cdot 2 + 2 + 4 = 8 \text{ electrons}$$

Oxygen is a neutral atom, so no electrons are removed. Find the highest partly-filled sublevel, fill it by Hund's rule, and count the unpaired electrons.

✓ Worked example 1: solution



the 2p sublevel holds 4 electrons in its 3 boxes

Step 2 • Draw the partly-filled sublevel

Step 3 • Fill singly before pairing



two boxes still carry a single, unpaired electron

Step 4 • Count lone arrows

2 unpaired → paramagnetic

a p sublevel with 4 electrons: $6 - 4 = 2$ unpaired

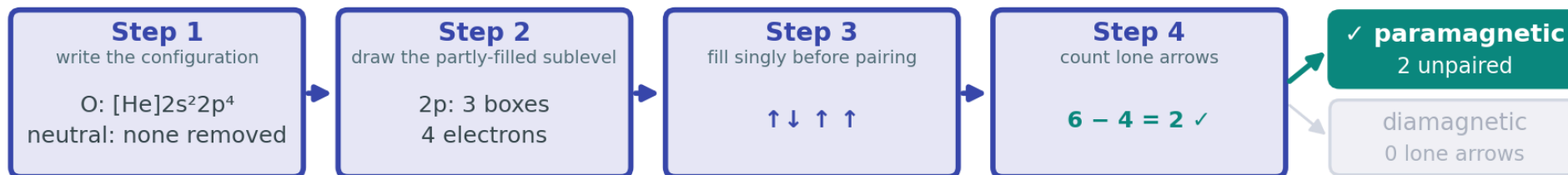
 **Sense check:** Oxygen gas is genuinely paramagnetic. Poured between the poles of a strong magnet, pale-blue liquid O₂ hangs in the gap instead of falling through.

✓ Worked example 1: the route



given: a neutral oxygen atom · found: $6 - 4 = 2$ unpaired · paramagnetic

the oxygen atom



 **Sense check:** The four steps run in order with no detour. Only the 2p sublevel decides the verdict: two lone arrows make oxygen paramagnetic.



Guided example: the vanadium(II) ion

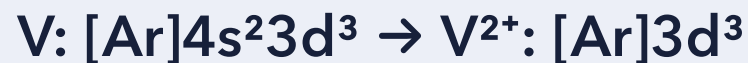
V: atomic number 23 $\rightarrow$ V^{2+}

a 2+ ion: two electrons removed from the neutral atom

Run all four method steps in order and name each move as you make it. Decide whether V^{2+} is paramagnetic or diamagnetic.

✓ Guided example: solution

Step 1 · Write the configuration



$$23 - 2 = 21 \text{ electrons} = 18 + 3$$

Step 2 · Draw the partly-filled sublevel

Step 3 · Fill singly before pairing

Step 4 · Count lone arrows

3d: $\uparrow \uparrow \uparrow _ _ \rightarrow 3 \text{ unpaired} \rightarrow \text{paramagnetic}$
 three singles · $5 - 3 = 2 \text{ boxes left empty}$

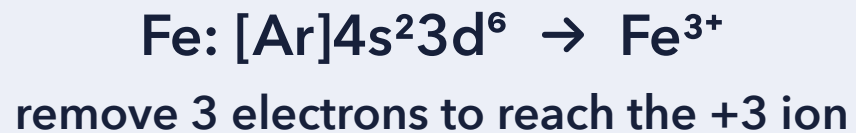
🧠 **Sense check:** The 4s pair that left carried no lone arrow. Neutral V keeps the same $3d^3$, so it also has 3 unpaired.

the vanadium(II) ion



Worked example 2: the iron(III) ion

Step 1 • Write the configuration



A common first attempt pulls all three electrons straight out of 3d. Build Fe^{3+} the right way, then count its unpaired electrons.

✓ Worked example 2: solution

A common first attempt



wrong order: the 4s electrons must leave before any 3d electron

Step 1 • Remove the highest-n electrons first



$26 - 3 = 23$ electrons remain

Steps 2-4 • Fill the boxes and count

3d: $\uparrow \uparrow \uparrow \uparrow \uparrow \rightarrow 5$ unpaired $\rightarrow$ paramagnetic

a half-filled d sublevel carries the maximum 5 unpaired electrons

🧠 **Sense check:** Strip 4s before 3d. Fe^{2+} keeps it: $[\text{Ar}]3d^6$, $10 - 6 = 4$ unpaired, still paramagnetic.

✓ Worked example 2: the route



given: remove 3 electrons · found: 3d $\uparrow \uparrow \uparrow \uparrow \uparrow$, 5 unpaired · paramagnetic

the iron(III) ion



Sense check: The one slip sits at Step 1: taking 3d before 4s gives $[\text{Ar}]4s^23d^3$. Box and count only after the 4s pair is gone.

Your turn: the manganese(II) ion

Mn: atomic number 25 $\rightarrow$ $[\text{Ar}]4s^23d^5 \rightarrow \text{Mn}^{2+}$
 remove 2 electrons: $25 - 2 = 23$ remain

step	answer
which electrons leave first	_____
Mn^{2+} configuration	$[\text{Ar}]3d$ _____
the 3d boxes	_____
unpaired count $\rightarrow$ verdict	_____

Run the method, then compare Mn^{2+} with Fe^{3+} .

Mn^{2+} : $[\text{Ar}]3d^5 \cdot 3d: \uparrow \uparrow \uparrow \uparrow \uparrow \cdot 5$ unpaired $\rightarrow$ paramagnetic
the 4s pair leaves $\cdot 23 - 18 = 5$ d electrons, one per box, none paired $\cdot$ isoelectronic with Fe^{3+}

⚠ Where this goes wrong

Counting orbitals or total electrons instead of unpaired ones. A full 3d sublevel has 5 orbitals and 10 electrons, yet 0 are unpaired. Only the lone, unpartnered arrows decide magnetism.

Pairing electrons too early. Writing $2p^4$ as $\uparrow\downarrow \uparrow\downarrow$ (two pairs, one empty box) gives 0 unpaired and the wrong verdict. Hund's rule fills each box singly first, so the true count is 2.

Removing (n-1)d before ns in an ion. Fe^{3+} is $[\text{Ar}]3d^5$, not $[\text{Ar}]4s^23d^3$. The 4s electrons leave first; only then does 3d give one up.

Assuming every ion is diamagnetic. Many transition-metal ions keep unpaired d electrons: Fe^{3+} has 5, Cr^{3+} has 3, both paramagnetic. Diamagnetic requires *all* electrons paired.

Practice 1

Which one of these species is diamagnetic?

- A. O (oxygen atom)
- B. Fe^{3+}
- C. Zn^{2+}
- D. Na (sodium atom)

✓ Practice 1 · answer: C



every 3d box paired · $10 - 10 = 0$ unpaired → diamagnetic

🧠 **Sense check:** Diamagnetic is the strict verdict: not a single electron may be left unpaired.

choice C: the zinc ion, Zn^{2+}





Practice 2

How many unpaired electrons does a ground-state arsenic atom have? Its electron configuration is $[\text{Ar}]4s^23d^{10}4p^3$.

- A. 13
- B. 3
- C. 15
- D. 5

✓ Practice 2 · answer: B



$4s^2$ and $3d^{10}$ are full, so only $4p$ is partly filled · $18 + 2 + 10 + 3 = 33$

$4p$: $\uparrow \uparrow \uparrow \rightarrow 3$ unpaired $\rightarrow$ paramagnetic (answer B)

three electrons, three boxes: one apiece, none forced to pair

🧠 **Sense check:** Full sublevels add no lone arrows. Arsenic ends in p^3 like N and P, with the same three singles.

the arsenic atom, $[\text{Ar}]4s^23d^{10}4p^3$





Practice 3

How many unpaired electrons does a ground-state Ni^{3+} ion have?

- A. 3
- B. 5
- C. 2
- D. 1
- E. 7

✓ Practice 3 · answer: A



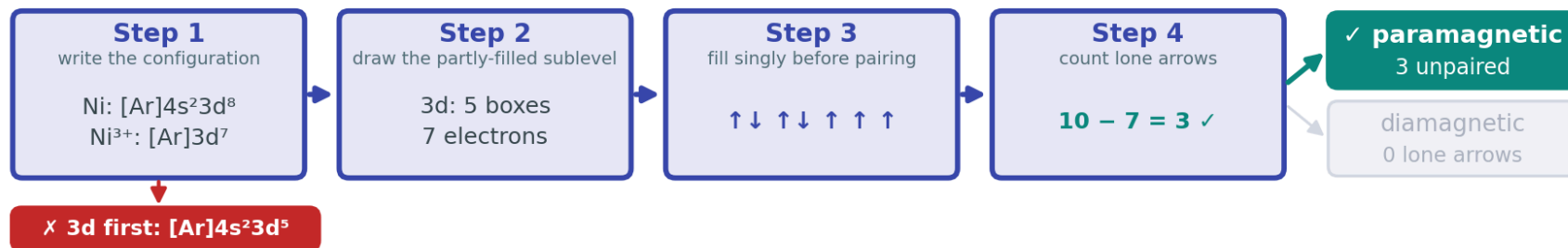
the 4s pair leaves first, then one 3d · $28 - 3 = 25$ electrons = $18 + 7$

3d: $\uparrow\downarrow \uparrow\downarrow \uparrow \uparrow \uparrow \rightarrow 3$ unpaired $\rightarrow$ paramagnetic (answer A)

one arrow per box first, then the last two pair · $10 - 7 = 3$

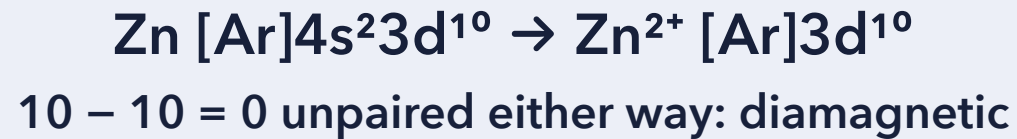
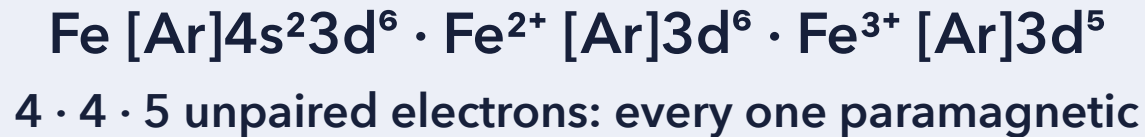
 **Sense check:** Neutral Ni and Ni^{2+} both hold $3d^8$ with 2 unpaired. Pulling the third electron out of a paired box raises the count to 3.

the nickel(III) ion



Ions shift the count, rarely the verdict

Removal empties the highest n first: 4s before 3d. The unpaired count changes, but one lone electron keeps the ion paramagnetic. The verdict flips only when removal leaves every electron paired.



Check yourself

1. Manganese is $[\text{Ar}]4s^23d^5$. How many unpaired electrons does it have, and is it paramagnetic or diamagnetic?
2. Write the configuration of Ca^{2+} and decide its magnetism.

An atom's magnetism is a direct read-out of its orbital boxes: fill them, count the lone arrows, and the verdict follows. The same unpaired-electron bookkeeping, carried into molecules and their bonding orbitals, is what explains why O_2 is magnetic while N_2 is not.

Can you...?

- ☐ relate a light wave's wavelength, frequency, and energy using $c = \lambda\nu$ and $E = h\nu$?
- ☐ calculate the energy of a photon from its wavelength or frequency?
- ☐ state how many electrons each sublevel (s, p, d, f) holds and why?
- ☐ write the ground-state electron configuration of an atom using the building-up order?
- ☐ predict periodic trends in atomic radius, ionization energy, electronegativity, metallic character, and electron affinity from an element's position, including the exceptions to the trends?
- ☐ explain the effect of ionization on size: a cation is smaller and an anion larger than its atom, and an isoelectronic series ranks by nuclear charge?

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