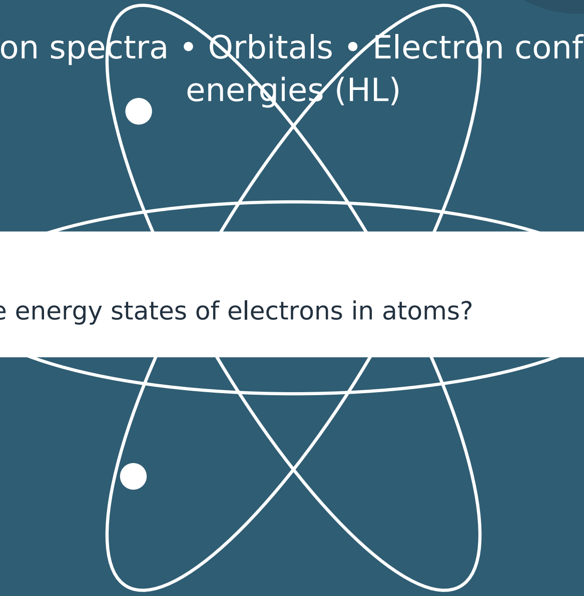
Electron Configurations

IB / A Level Exam Preparation Notes

Structure 1.3 • Emission spectra • Orbitals • Electron configurations • Ionization energies (HL)



Guiding Question
How can we model the energy states of electrons in atoms?

Exam route and syllabus checklist

Use these notes to revise the conceptual model, practise the calculations, and learn the wording expected in IB and A level mark schemes.

Core skill	What you must be able to do in an exam
Emission spectra	Explain why line spectra show that electrons occupy discrete energy levels; distinguish continuous and line spectra.
Electromagnetic radiation	Link wavelength, frequency and energy: shorter wavelength means higher frequency and higher photon energy.
Hydrogen spectrum	Relate transitions to n = 1, n = 2 and n = 3+ to UV, visible and IR regions; explain convergence.
Orbitals	Describe s and p orbital shapes; know that each orbital holds two electrons of opposite spin.
Configurations	Apply Aufbau, Hund and Pauli to atoms and ions up to Z = 36; use full and noble-gas notation.
Periodic table link	Connect period, group and block to the occupied sublevel of highest energy.
HL ionization energy	Calculate IE from spectral data and interpret first and successive IE trends.

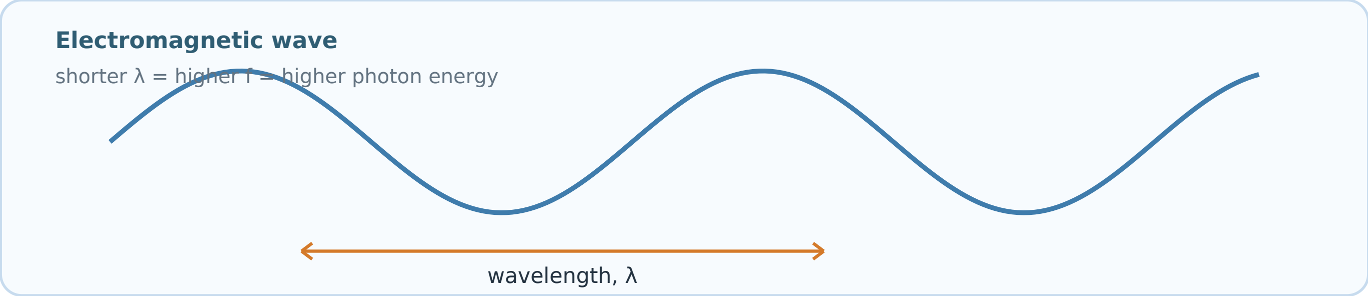
Study priority

For fast revision, master the three links: emission lines → quantized energy levels; orbital filling rules → electron configurations; ionization energy jumps → electron shell and sublevel structure.

Core equations

Equation	Use	Exam reminder
$c = f\lambda$	Relates speed, frequency and wavelength.	Convert nm to m before calculating.
$E = hf$	Energy of one photon.	Higher frequency means higher photon energy.
$E = hc/\lambda$	Photon energy from wavelength.	Shorter λ gives larger E.
$IE = hfN_A$	Ionization energy per mole from convergence frequency.	Divide by 1000 to convert J mol^{-1} to kJ mol^{-1}.
Maximum electrons = $2n^2$	Capacity of a main energy level.	n = 3 holds 18 electrons.

1. Emission spectra and electromagnetic radiation



Atoms can absorb energy and become excited. When excited electrons fall back to lower energy levels, energy is released as photons of electromagnetic radiation. The colour or region of the spectrum depends on the photon energy.

Key idea

A line emission spectrum contains only specific wavelengths because only specific electron energy changes are allowed. This is evidence that electron energy levels are discrete, not continuous.

Continuous spectrum vs line spectrum

Feature	Continuous spectrum	Line emission spectrum
Appearance	Unbroken range of colours or frequencies.	Separate bright lines at specific wavelengths.
Source	White light dispersed by a prism, or thermal radiation.	Excited atoms or ions returning to lower energy states.
Energy model	Energy appears over a continuous range.	Energy is emitted in discrete quanta.
Exam phrase	“Colours merge smoothly.”	“Specific frequencies show quantized energy levels.”

Energy, wavelength and colour

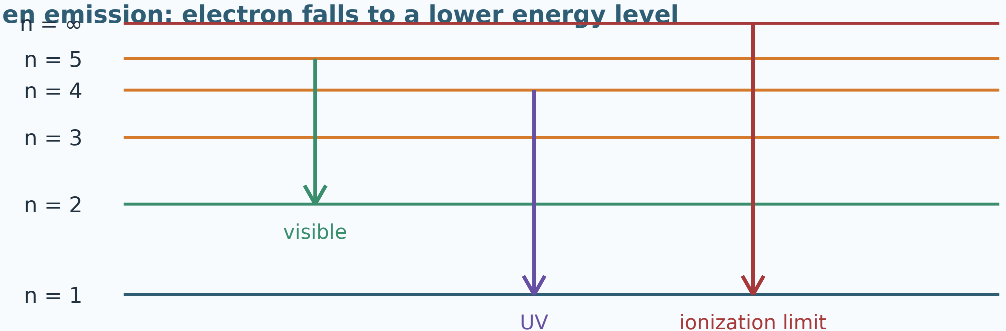
- All electromagnetic radiation travels at the same speed in a vacuum: $c = 3.00 \times 10^{8}$ m s^{-1}.
- Frequency is the number of waves passing a point per second; wavelength is the distance between successive crests.
- Short wavelength → high frequency → high photon energy.
- In visible light, red has a longer wavelength and lower energy than blue/violet.
- Infrared has longer wavelength and lower energy than visible red; ultraviolet has shorter wavelength and higher energy than violet.

Common exam trap

Do not say “blue light travels faster than red light”. In a vacuum, all electromagnetic waves travel at the same speed; they differ in wavelength, frequency and energy.

2. Hydrogen line spectrum and quantization

Hydrogen emission: electron falls to a lower energy level



Lines converge at high energy because the gaps between levels become smaller.

Hydrogen is important because it has only one electron. Bohr’s model successfully explains the pattern of lines in its emission spectrum by allowing the electron to occupy only certain energy levels.

When the electron falls from a higher level to a lower level, the energy difference is emitted as a photon:

$$\Delta E_{electron} = E_{photon} = hf$$

Final level	Region of spectrum	Relative energy change	What to say in exams
n = 1	Ultraviolet	Largest	Transitions to the first level have high energy and short wavelength.
n = 2	Visible	Intermediate	The visible hydrogen lines come from falls to n = 2.
n = 3 or higher	Infrared	Smaller	Transitions to higher final levels have lower energy.
n = ∞	Ionization limit	Electron removed	The convergence limit corresponds to ionization.

Why do the lines converge?

- bulletThe energy levels become closer together as n increases.
- bulletTherefore the differences in energy between adjacent high levels become smaller.
- bulletThe spectral lines get closer together at higher frequency and higher energy.
- bulletAt the convergence limit, the electron has enough energy to leave the atom completely.

Mark scheme wording

“The emission spectrum has discrete lines because electrons can only move between fixed energy levels. Each line corresponds to a particular transition and a photon with a specific frequency.”

3. Photon energy calculations

For calculations, decide whether the data give frequency or wavelength. Then calculate energy for one photon, and multiply by Avogadro's constant if the answer is required per mole.

1. Write the equation: $c = f\lambda$, $E = hf$, or $E = hc/\lambda$.
2. Convert wavelength to metres if it is given in nm: $1\ \text{nm} = 1 \times 10^{-9}\ \text{m}$.
3. Calculate energy for one photon in joules.
4. For one mole of photons or atoms, multiply by N_A.
5. Convert J mol^{-1} to kJ mol^{-1} by dividing by 1000.

Worked example: sodium lamp

A sodium lamp emits orange light with wavelength 590 nm. Calculate the energy released by one mole of excited sodium atoms.

$\lambda = 590\ \text{nm} = 590 \times 10^{-9}\ \text{m}$

$E_{mol} = hcN_A/\lambda$

$E_{mol} = (6.63 \times 10^{-34} \times 3.00 \times 10^{8} \times 6.02 \times 10^{23}) / (590 \times 10^{-9})$

$E_{mol} \approx 2.03 \times 10^{5}\ \text{J mol}^{-1} = 203\ \text{kJ mol}^{-1}$

HL worked example: ionization from convergence frequency

The hydrogen convergence frequency is approximately $3.23 \times 10^{15}\ \text{s}^{-1}$.

$IE = hfN_A = 6.63 \times 10^{-34} \times 3.23 \times 10^{15} \times 6.02 \times 10^{23}$

$IE \approx 1.29 \times 10^{6}\ \text{J mol}^{-1} = 1290\ \text{kJ mol}^{-1}$.

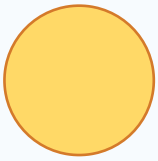
Calculation traps

Use SI units, keep enough significant figures, and remember that $E = hf$ gives energy for one photon. Exam questions often ask for one mole, so N_A must be included.

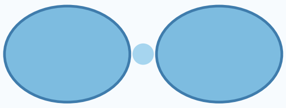
4. From Bohr orbits to atomic orbitals

The Bohr model pictures electrons in fixed circular energy levels. This works well for hydrogen but fails for multi-electron atoms. A more detailed quantum model uses orbitals: regions where there is a high probability of finding an electron.

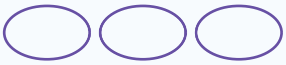
Atomic orbitals: regions of high probability



s orbital: spherical



p orbital: dumbbell



3 p orbitals per p sublevel

Orbitals and sublevels

- An orbital is a region around the nucleus where there is a high probability of finding an electron.
- Each orbital can hold a maximum of two electrons with opposite spin.
- Orbitals of equal energy are called degenerate orbitals and form a sublevel.
- s orbitals are spherical; p orbitals are dumbbell-shaped and occur in three orientations: p_x, p_y and p_z.
- You should recognize s and p orbital shapes; detailed d and f shapes are usually not required.

Pauli exclusion principle

No more than two electrons can occupy one orbital, and if two electrons are in the same orbital they must have opposite spins.

Main energy levels and sublevels

Main level	Sublevels present	Maximum electrons
n = 1	1s	2
n = 2	2s, 2p	8
n = 3	3s, 3p, 3d	18
n = 4	4s, 4p, 4d, 4f	32

Sublevel	Number of orbitals	Maximum electrons
s	1	2
p	3	6
d	5	10
f	7	14

5. Electron configurations: rules and method

A ground-state electron configuration describes the lowest-energy arrangement of electrons in an atom. Unless the question says “excited state”, assume ground state.

Rule	Meaning	How it appears in answers
Aufbau principle	Electrons fill the lowest available energy orbitals first.	Use the filling order correctly.
Hund’s rule	In degenerate orbitals, electrons occupy separate orbitals singly before pairing.	For p^2 or p^3, place arrows in separate boxes first.
Pauli exclusion principle	An orbital holds a maximum of two electrons with opposite spins.	Use ↑↓, not ↑↑, in the same orbital.

Order of filling

Memorize this sequence

1s → 2s → 2p → 3s → 3p → 4s → 3d → 4p → 5s → 4d → 5p → 6s → 4f → 5d → 6p → 7s → 5f → 6d → 7p

Worked example: sulfur atom

Sulfur has atomic number 16, so a neutral sulfur atom has 16 electrons. Fill the orbitals in order:

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$$

The 3p orbital diagram is:

Common configurations to know quickly

Element / ion	Configuration	Comment
Ne	$1s^2\ 2s^2\ 2p^6$	Noble gas core used as [Ne].
Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	Noble gas core used as [Ar].
Ca	$[Ar]\ 4s^2$	4s fills before 3d.
V	$[Ar]\ 3d^3\ 4s^2$	Often written with 3d before 4s after Ca.
Cr	$[Ar]\ 3d^5\ 4s^1$	Exception: half-filled d sublevel is stable.
Cu	$[Ar]\ 3d^{10}\ 4s^1$	Exception: filled d sublevel is stable.

6. Electron configurations of ions

Positive ions form when atoms lose electrons. Negative ions form when atoms gain electrons. The key question is always: which electrons are removed or added?

Most important ion rule

For transition metal atoms, remove electrons from the 4s sublevel before the 3d sublevel, even though 4s fills before 3d in neutral atoms.

Species	How to think	Electron configuration
O^{2-}	O gains 2 electrons: 8 + 2 = 10 electrons.	$1s^2$ $2s^2$ $2p^6$ = [Ne]
Cl^-	Cl gains 1 electron: 17 + 1 = 18 electrons.	[Ar]
Al^{3+}	Al loses 3 valence electrons.	[Ne]
Ti^{3+}	Ti is [Ar] $3d^2$ $4s^2$; remove $4s^2$ then one 3d.	[Ar] $3d^1$
Fe^{3+}	Fe is [Ar] $3d^6$ $4s^2$; remove $4s^2$ then one 3d.	[Ar] $3d^5$
Cu^{2+}	Cu is [Ar] $3d^{10}$ $4s^1$; remove 4s then one 3d.	[Ar] $3d^9$

Step-by-step method for ion configurations

1. Write the neutral atom configuration first.
2. For a negative ion, add electrons to the next available orbital using Aufbau and Hund.
3. For a positive main-group ion, remove electrons from the outer main energy level.
4. For a transition-metal ion, remove 4s electrons before 3d electrons.
5. Check that the total number of electrons matches the charge.

Common exam trap

Do not remove 3d electrons before 4s electrons when forming transition metal ions. For example, Fe^{2+} is [Ar] $3d^6$, not [Ar] $3d^4 4s^2$.

7. Electron configuration and the periodic table

The periodic table is organized by electron configuration. The block tells you which type of sublevel is being filled, and the period tells you the highest occupied main energy level for s and p block elements.

Block	Valence / highest-energy sublevel being filled	Examples
s block	ns	Group 1 and 2 elements, plus H and He by configuration.
p block	np	Groups 13-18. Example: Cl = [Ne] $3s^2$ $3p^5$.
d block	(n - 1)d	Transition metals. Example: Fe = [Ar] $3d^6$ $4s^2$.
f block	(n - 2)f	Lanthanides and actinides.

Deducing configuration from position

Examples

Caesium is in group 1, period 6 → [Xe] $6s^1$.

Iodine is in group 17, period 5 → [Kr] $4d^{10}$ $5s^2$ $5p^5$.

Zinc is the end of the 3d block in period 4 → [Ar] $3d^{10}$ $4s^2$.

Valence electrons

- bullet Valence electrons are the outer electrons most involved in bonding and ion formation.
- bullet For main-group elements, the group pattern helps predict the number of valence electrons.
- bullet Elements in the same group have similar outer electron configurations and therefore similar chemical properties.
- bullet For d-block elements, both ns and (n - 1)d electrons can be involved in chemical behaviour, which helps explain variable oxidation states.

Exam language

Use the phrase “occupied sublevel of highest energy” when explaining why the periodic table has s, p, d and f blocks.

8. HL: ionization energies and spectral convergence

The first ionization energy is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms in their ground state.

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

Definition to memorize

First ionization energy: the energy required to remove one mole of electrons from one mole of gaseous atoms in their ground state to form one mole of gaseous 1+ ions.

Calculating IE from convergence data

- bullet At the convergence limit, the electron has reached n = ∞ and the atom is ionized.
- bullet If frequency is given: $IE = hfN_A$.
- bullet If wavelength is given: $IE = hcN_A/\lambda$.
- bullet The final answer is usually reported in kJ mol^{-1}.

Trends in first ionization energy

Trend or discontinuity	Explanation expected in exams
Across a period: generally increases	Nuclear charge increases while electrons are added to the same main energy level, so attraction to the outer electron increases.
Down a group: decreases	The outer electron is in a higher main energy level, further from the nucleus and more shielded.
Group 2 → group 13 drop	The electron removed from group 13 is in a higher-energy p orbital, so it is easier to remove than an s electron.
Group 15 → group 16 drop	The electron removed from group 16 is paired in a p orbital and is repelled by its partner, so it is easier to remove.

Do not just say “shielding increases” across a period

Across a period, shielding is similar because electrons are added to the same main shell. The main reason for the increase is increasing nuclear charge and stronger attraction to the outer electron.

9. HL: successive ionization energies

Successive ionization energies measure the energy needed to remove electrons one after another from the same element. They provide evidence for shells and sublevels.

How to interpret jumps

1. Look for a very large jump between two successive ionization energies.
2. The jump occurs after all valence electrons have been removed.
3. The number of electrons removed before the jump gives the number of valence electrons.
4. For main-group elements, this usually identifies the group number.
5. Smaller jumps can show sublevel changes, such as from p to s within the same main level.

Worked example: group deduction

First four ionization energies: 738, 1450, 7730, 10550 kJ mol^{-1}.

There is a large jump between the second and third ionization energies.

Therefore two electrons are removed relatively easily before an inner-shell electron would be removed.

The element has two valence electrons and is in group 2.

Using IE evidence for electron configuration

bullet A 2, 8, 3 pattern in successive IE jumps suggests an arrangement such as aluminium: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^1$.

bullet A jump after six p-electrons have been removed suggests a lower-energy s sublevel is being reached.

bullet Transition metals often show more gradual increases for removal of 4s and 3d electrons because these sublevels are close in energy.

bullet A very large jump signals removal from a lower main energy level closer to the nucleus.

Marking language

When explaining a jump, name the orbital/sublevel being removed before and after the jump, then compare distance from the nucleus, shielding and electrostatic attraction.

10. Exam command words and answer patterns

Command	What to include
Distinguish	State clear differences, usually in paired sentences. Example: continuous spectrum has an unbroken range; line spectrum has discrete lines.
Explain	Give a reason using scientific principles. Example: first IE increases because nuclear charge increases and attraction to outer electrons becomes stronger.
Deduce	Use data or given information to reach a conclusion. Show electron count, jump position or transition energy.
Calculate	Write formula, substitute with units, show conversion and final value.
Justify	Support your answer with evidence from configuration, graph, spectrum or data.

High-value phrases for marks

- “The electron is removed from a higher-energy orbital.”
- “The orbital is further from the nucleus and experiences weaker electrostatic attraction.”
- “The electron is paired and repelled by another electron in the same orbital.”
- “The large jump shows that the next electron is in a lower main energy level.”
- “The lines converge because the energy levels become closer together at higher energy.”
- “Each line corresponds to a photon of a specific frequency produced by a specific electron transition.”

Typical one-mark errors

Missing “gaseous” in the ionization energy definition; forgetting to remove 4s before 3d for transition metal ions; saying atoms “emit electrons” instead of photons in emission spectra; confusing wavelength and frequency trends.

Revision checklist

- Can you write configurations up to Kr without using a periodic table?
- Can you draw p^2, p^3, d^5 and d^{10} orbital diagrams?
- Can you calculate E from λ and from f?
- Can you explain both first IE discontinuities across a period?
- Can you identify a group from successive IE data?
- Can you explain why Cr and Cu are exceptions?

11. Practice questions

Try these without notes first. The answer key and marking guidance are on the following pages.

1 1. State the relationship between wavelength, frequency and energy for electromagnetic radiation.

2 2. Distinguish between a continuous spectrum and a line emission spectrum.

3 3. Explain why the line emission spectrum of hydrogen provides evidence for discrete energy levels.

4 4. A transition from n = 4 to n = 2 in hydrogen is compared with n = 3 to n = 2. Which transition has the shorter wavelength? Explain.

5 5. Calculate the energy, in kJ mol^{-1}, of light with wavelength 486 nm.

6 6. State the maximum number of electrons in the n = 3 main energy level and explain your answer.

7 7. State the number of orbitals and maximum electrons in an f sublevel.

8 8. Draw or describe the orbital diagram for nitrogen, $1s^2\ 2s^2\ 2p^3$.

9 9. Write the full electron configuration of selenium, Se (Z = 34).

10 10. Write the condensed electron configuration of Cu and Cu^{2+}.

11 11. Explain why the first ionization energy decreases from Be to B.

12 12. Explain why the first ionization energy decreases from N to O.

13 13. The first four ionization energies of an element are 590, 1145, 4912 and 6491 kJ mol^{-1}. Deduce the group and identify the likely period-4 element.

14 14. The convergence frequency of an emission series is $5.27 \times 10^{16}\ s^{-1}$. Calculate the ionization energy in kJ mol^{-1}.

15 15. Explain why titanium can show several positive oxidation states but calcium usually forms only Ca^{2+}.

12. Answer key and marking guidance

Q	Answer
1	$c = f\lambda$ and $E = hf$, so shorter wavelength means higher frequency and higher energy.
2	A continuous spectrum has an unbroken range of wavelengths; a line emission spectrum has discrete bright lines at specific wavelengths.
3	Each line is caused by a photon from a specific electron transition. Since only certain photon energies are emitted, electrons can only occupy certain energy levels.
4	$n = 4 \rightarrow n = 2$ has the greater energy change than $n = 3 \rightarrow n = 2$, so it has a higher frequency and shorter wavelength.
5	$E = hcN_A/\lambda = (6.63 \times 10^{-34} \times 3.00 \times 10^{8} \times 6.02 \times 10^{23}) / (486 \times 10^{-9}) = 246$ kJ mol^{-1} approximately.
6	18 electrons. Maximum = $2n^2 = 2 \times 3^2 = 18$.
7	An f sublevel contains 7 orbitals and holds a maximum of 14 electrons.
8	1s and 2s are paired. The three 2p electrons occupy the three p orbitals singly with parallel spins.
9	Se: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^4$, or $[Ar]\ 3d^{10}\ 4s^2\ 4p^4$.
10	Cu = $[Ar]\ 3d^{10}\ 4s^1$; Cu^{2+} = $[Ar]\ 3d^9$. Remove 4s before 3d.
11	B loses a 2p electron, while Be loses a 2s electron. The 2p electron is higher in energy/further from the nucleus and is easier to remove.
12	O has a paired electron in a p orbital. Repulsion within the doubly occupied orbital makes this electron easier to remove than an unpaired p electron in N.
13	Large jump after the second IE, so the element has two valence electrons and is group 2. The period-4 group 2 element is Ca.
14	$IE = hfN_A = 6.63 \times 10^{-34} \times 5.27 \times 10^{16} \times 6.02 \times 10^{23} = 2.10 \times 10^{7}$ J mol^{-1} = 2.10×10^{4} kJ mol^{-1}.
15	Ti has 4s and 3d electrons close in energy, so removal of several electrons is energetically possible. Ca has a large jump after removing two 4s electrons because the third would come from an inner 3p level.

13. Quick marking guidance

Use this page to check whether your written answers contain enough chemistry for full marks.

Question type	Full-mark features
Spectrum explanation	Mentions excited electrons, falling to lower levels, photon emission, specific frequency/energy, and discrete energy levels.
Energy calculation	Correct equation, SI conversion, correct substitution, correct units, sensible significant figures.
Configuration answer	Correct electron count, correct filling order, Hund/Pauli applied, correct exception if Cr or Cu, correct ion rule if transition metal.
IE trend explanation	Nuclear charge, shielding, distance from nucleus, orbital energy and electron-pair repulsion used appropriately.
Successive IE interpretation	Identifies jump position, links to valence electrons, names the shell/sublevel change, deduces group or configuration.

Final one-page memory list

Core facts to memorize

- bullet Line spectra are evidence for discrete electron energy levels.
- bullet Visible hydrogen lines are transitions to n = 2; UV to n = 1; IR to n = 3 or higher.
- bullet s, p, d, f sublevels hold 2, 6, 10, 14 electrons respectively.
- bullet Each orbital holds two electrons with opposite spin.
- bullet Filling order includes 4s before 3d, but transition metal ions lose 4s before 3d.
- bullet Cr = [Ar] $3d^5 4s^1$; Cu = [Ar] $3d^{10} 4s^1$.
- bullet Large successive IE jumps reveal inner shells and the number of valence electrons.
- bullet First IE increases across a period, decreases down a group, with key dips at groups 13 and 16.

End of notes