

SCIENCE EXAM SCORE-SAVER

Particle Nature of Matter | Lessons 11-14

Use this guide to protect easy marks

Do not memorize isolated facts. Lock in the relationships that the lesson slides repeatedly connect.

Focus first on distinctions that are easy to mix up: valence electrons vs total electrons, electron transfer vs ionic bond, shell vs subshell vs orbital, and the three filling rules.

If you cannot explain a box below in one clear sentence, mark it for immediate review.

THE TWO MASTER CHAINS

Periodic position group pattern	Valence electrons outer electrons	Stability full shell	Ion formation gain / lose e ⁻	Ionic compound charge balance
Shell energy level	Subshell sub-level	Orbital probable region	Filling rules Aufbau / Pauli / Hund	Configuration electron arrangement

THE 6 DISTINCTIONS MOST WORTH PROTECTING

1 ELECTRON TRANSFER ≠ IONIC BOND Transfer creates ions. The ionic bond is the attraction between positive and negative ions.	2 IONIC COMPOUND ≠ MOLECULE The lesson states that ionic compounds form large crystal lattices, not individual molecules.
3 CHARGE ≠ SUBSCRIPT Charges tell you how many positive and negative charges must balance. The final formula is written with subscripts, not charge signs.	4 SHELL ≠ SUBSHELL ≠ ORBITAL Shell = energy level. Subshell = sub-level within a shell. Orbital = region where an electron will probably be found.
5 AUFBAU ≠ PAULI ≠ HUND Aufbau: lowest energy first. Pauli: max 2 electrons per orbital. Hund: occupy equal-energy orbitals singly before pairing.	6 FULL ≠ ABBREVIATED CONFIGURATION Full notation writes all occupied subshells. Abbreviated notation replaces the completed inner part with a noble-gas symbol.

Source lessons: Lesson 11 - Nuclide Notation and Ions | Lesson 12-14 - Electron Orbitals and the Periodic Table

01 | VALENCE ELECTRONS, STABILITY & IONS

High-risk checkpoints from Lesson 11. Read the left column as the trap; memorize the right column as the score-saving correction.

HIGH-RISK POINT	WHAT TO REMEMBER
Same group = similar properties	The lesson links similar group behavior to repeating valence-electron structure.
Noble gases	Group 18 is very stable and unreactive because the valence shell is full: 8 electrons, or 2 in helium's first energy level.
Why atoms react	Atoms may gain, lose, or share electrons to achieve a full valence shell and greater stability.
Group 1 / 2 / 13 ions	Lose 1 / 2 / 3 electrons -> 1+ / 2+ / 3+.
Group 15 / 16 / 17 ions	Gain 3 / 2 / 1 electrons -> 3- / 2- / 1-.
Group 18	The lesson chart shows no ions for Group 18.
Group 14	The lesson chart marks this region "XX" rather than giving a simple ion-charge pattern.
Transition-metal note	The slide warns that added electrons in the transition-metal region are not added to the valence shell in the same simple way; it notes that many transition metals have 2 valence electrons in this simplified treatment.

PRECISION LANGUAGE TO USE

- "loses electrons" / "gains electrons"
- "forms a positive ion" / "forms a negative ion"
- "has a full valence shell"
- "elements in the same group have similar properties"

ONE-LINE LOGIC

Periodic position -> valence-electron pattern -> ion behavior -> chemical behavior.

02 | IONIC BONDING, NAMES & FORMULAS

Most avoidable errors here come from mixing up the process, the bond, the name, and the formula.

HIGH-RISK POINT	WHAT TO REMEMBER
What is ionic bonding?	Metal + nonmetal; electron transfer produces positive and negative ions; the ionic bond is the attraction that holds them together.
Structure of ionic compounds	Large crystal lattices, not individual molecules.
Naming order	Metal first. Nonmetal second, with its ending changed to -ide.
Overall charge	An ionic compound has no overall charge because positive and negative charges are balanced.
Na ⁺ + Cl ⁻	1+ and 1- balance in a 1:1 ratio -> NaCl -> sodium chloride.
Mg ²⁺ + Cl ⁻	One Mg ²⁺ needs two Cl ⁻ ions -> MgCl ₂ -> magnesium chloride.
Criss-cross shortcut	Move the charge numbers to the opposite ion as subscripts.
What to ignore in the final formula	Ignore the charge signs and omit 1s when writing subscripts.

FORMULA CHECK: DO THIS IN THE SAME ORDER EVERY TIME

1	2	3	4
Find elements	Determine charges	Balance / criss-cross	Write symbols + subscripts

EXAMPLES FROM THE LESSON

Al₂O₃ = aluminium oxide
Be₃N₂ = beryllium nitride
Li₂S = lithium sulfide
NaBr = sodium bromide

MORE LESSON PRACTICE ANSWERS

CaF₂ = calcium fluoride
Mg₃P₂ = magnesium phosphide
K₂O = potassium oxide
AlP = aluminium phosphide
CaO = calcium oxide

DO NOT CONFUSE THESE TWO SENTENCES

Electron transfer explains how the ions are formed.
Attraction between the positive and negative ions is the ionic bond.

03 | SHELLS, SUBSHELLS & ORBITALS

Lesson 12-14 moves from simple electron shells to the quantum model. The safest strategy is to keep the hierarchy and the definitions separate.

SHELL An energy level that electrons may occupy. Labelled by the principal quantum number $n = 1, 2, 3, \dots$	SUBSHELL A sub-level of energy within an electron shell.
ORBITAL A region of space around the nucleus in which an electron will probably be found. Maximum: 2 electrons per orbital.	SCHRODINGER MODEL The lesson states that electrons behave like waves and can be described using probability functions.

ORBITAL SHAPES: KEEP THE BASIC PICTURE CLEAN

s ORBITAL Spherical shape.	p ORBITALS Dumbbell shape. Oriented in x, y, and z directions. The lesson notes that shape affects orbital overlap in bonding.
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HIERARCHY LOCK

Shell → contains subshells → contain orbitals → orbitals hold electrons.

HIGH-RISK ERROR

Do not use shell, subshell, and orbital as if they are interchangeable words. The lesson gives each a different definition.

04 | THE 3 ELECTRON-FILLING RULES

These three names are easy to memorize but easy to apply incorrectly. Match each rule to exactly one job.

RULE	THE ONE JOB TO REMEMBER
AUFBAU PRINCIPLE	Electrons fill the lowest available energy orbitals before higher-energy orbitals.
PAULI EXCLUSION PRINCIPLE	No two identical electrons can occupy the exact same quantum state; the lesson-level practical check is a maximum of 2 electrons per orbital.
HUND'S RULE	Within equal-energy orbitals of a subshell, occupy each orbital singly before any orbital is doubly occupied; the singly occupied orbitals have the same spin.

FAST ERROR-DIAGNOSIS

Wrong energy order? Check AUFBAU.	More than 2 in one orbital? Check PAULI.
Paired too early in equal-energy orbitals? Check HUND.	Not sure which rule? Ask: energy order, orbital capacity, or pairing pattern?

IRON EXAMPLE IN THE LESSON

The lesson explains that the 2p orbitals are filled according to Hund's rule: one electron enters each equal-energy orbital before pairing.

It also notes that the final electrons in iron are placed in 3d orbitals in a pattern that obeys Hund's rule.

05 | ELECTRON CONFIGURATION: PROTECT THE NOTATION

The lesson tests not just the idea of electrons, but how their arrangement is written.

FULL CONFIGURATION	ABBREVIATED FORM
Potassium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ Phosphorus: $1s^2 2s^2 2p^6 3s^2 3p^3$ The superscript tells how many electrons are in that subshell.	Potassium: $[\text{Ar}] 4s^1$ Argon: $[\text{Ne}] 3s^2 3p^6$ The noble-gas symbol replaces the completed inner electron configuration.

4 CHECKS BEFORE YOU MOVE ON

CHECK	QUESTION TO ASK
1 Total electrons	Does the notation account for the intended number of electrons?
2 Energy order	Did you follow the lesson's lowest-energy-first filling rule?
3 Orbital capacity	No orbital may contain more than 2 electrons.
4 Equal-energy orbitals	If an orbital diagram is shown, apply Hund's rule before pairing.

THE NOTATION ITSELF CAN COST MARKS

Keep the number, subshell letter, and electron count in the correct places.

Do not confuse the subshell label (for example, $3p$) with the superscript electron count (for example, 3 electrons in $3p^3$).

06 | FINAL 3-MINUTE PRE-EXAM SCAN

If every statement below feels automatic, the core content in the supplied lesson slides is under control.

01 I can explain why elements in the same group have similar properties using valence electrons.	02 I know why Group 18 is stable and unreactive in the lesson model.
03 I remember the helium exception: a full first shell has 2 electrons.	04 I can predict the simple main-group ion charges shown in the lesson chart.
05 I know that Group 14 is not given a simple ion-charge pattern in the chart.	06 I can distinguish electron transfer from the ionic bond itself.
07 I remember that ionic compounds form large crystal lattices, not individual molecules.	08 I can name a simple ionic compound: metal first, nonmetal ending in -ide.
09 I can balance ionic charges so the compound has no overall charge.	10 I know exactly what the criss-cross shortcut changes and what it ignores.
11 I can define shell, subshell, and orbital without mixing them up.	12 I know the maximum number of electrons in one orbital.
13 I know the basic shapes of s and p orbitals from the lesson.	14 I can tell Aufbau, Pauli, and Hund apart by function.
15 I can write and read a full electron configuration.	16 I can read a superscript as the number of electrons in a subshell.
17 I can recognize the abbreviated noble-gas form.	

LAST RULE

If one line above is not automatic, review that exact line first. Do not reread the entire unit.