3.1 Electron configurations and spectra

Name: _____ Class: _____ Date: _____ **Total: 14 marks**

KEY WORDS emission spectrum • energy level • sublevel • orbital • electron configuration • convergence limit

Objective

Build the skills to answer exam questions on **electron configurations** : emission spectra, energy levels, sublevels and orbitals.

You must be able to:

- explain a line emission spectrum, and why the lines of hydrogen converge at high frequency
- write full and condensed electron configurations for atoms and ions up to $Z = 36$, including Cr and Cu
- draw orbital diagrams using the Aufbau principle, Hund's rule and the Pauli principle

1 Key knowledge and worked examples

Spectra and energy levels

A **continuous spectrum** has all wavelengths; a **line spectrum** has only certain ones. An excited electron falls to a lower level and emits a photon with $\Delta E = hf = \frac{hc}{\lambda}$. Levels are fixed, so only certain photon energies appear. In hydrogen, falls to $n = 1$ give ultraviolet lines and falls to $n = 2$ give visible lines. Levels get closer at higher energy, so the lines **converge**.

level n	sublevels	orbitals	max electrons
1	1s	1	2
2	2s, 2p	$1 + 3$	8
3	3s, 3p, 3d	$1 + 3 + 5$	18
4	4s, 4p, 4d, 4f	$1 + 3 + 5 + 7$	32

Rules: fill lowest energy first (1s 2s 2p 3s 3p **4s 3d** 4p); each orbital holds two electrons of opposite spin (Pauli); in a sublevel, electrons fill orbitals singly before pairing (Hund). Exceptions: Cr $[Ar]3d^5 4s^1$, Cu $[Ar]3d^{10}4s^1$. Transition-metal ions lose **4s electrons first**.

Worked example 1: an atom. Ti ($Z = 22$): $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$, or $\mathbf{[Ar]3d^2 4s^2}$.

Worked example 2: an ion. Mn^{2+}: Mn is $[Ar]3d^5 4s^2$; remove the two 4s electrons: $\mathbf{[Ar]3d^5}$.

Answer shapes: count electrons first; for ions, add or remove electrons last.

2 Practice

2.1 Write the full electron configuration of sulfur and the condensed configuration of Fe^{2+}. [2]

S $\quad$ Fe^{2+}

$Z = 16$ $\quad$ $Z = 26$

2.2 Complete the orbital diagram for a nitrogen atom, and state the number of unpaired electrons and the rule that explains it. [3]

1s ☐ 2s ☐ 2p ☐☐☐

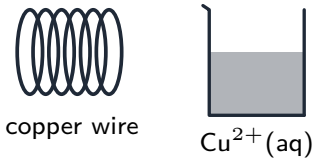
2.3 The diagram shows part of the emission spectrum of hydrogen in the visible region. Explain why the spectrum is a series of lines and why the lines get closer together towards the violet end. [3]



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3 Exam-style questions

3.1 Copper is used in electrical wiring, and copper(II) compounds are coloured.



(a) Write the condensed electron configuration of a copper atom. [1]

(b) Write the condensed electron configuration of the Cu^{2+} ion. [1]

(c) State the maximum number of electrons in the third energy level and in a 3d sub-level. [2]

(d) Copper salts give a green flame. Explain how the flame produces light of only certain colours. [2]

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Solutions

Marking: **M1** a correct method, **A1** the correct answer, **B1** an independent mark.

Each answer shows the steps, then the **result**.

2.1

S: $1s^2 2s^2 2p^6 3s^2 3p^4$; Fe^{2+}: $[Ar]3d^6$

2.2

1s and 2s each have a pair of opposite arrows; each 2p box has one arrow, all the same direction; **3** unpaired electrons; Hund's rule: electrons occupy orbitals of a sublevel singly before pairing

2.3

electrons fall from higher levels to $n = 2$, each emitting a photon of one energy / frequency; energy levels are fixed (quantized), so only certain frequencies appear; the levels get closer together at higher energy, so the lines converge

3.1

(a) $[Ar]3d^{10}4s^1$

(b) $[Ar]3d^9$

(c) **18**; **10**

(d) heat excites electrons to higher energy levels; they fall back and emit photons; the levels are fixed, so the photons have only certain energies / frequencies / colours

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