

(a) - Principal quantum shells are the main energy levels in an atom where electrons are found and are labeled by the principal quantum number from $n = 1$ to $n = 7$. ✓

- Sub-energy levels (subshells) are subdivisions of the main energy level and are labeled s, p, d and f. ✓

The 1st energy level ($n=1$) has the s-sub energy level, the 2nd energy level ($n=2$) has two sub energy levels; s and p, the 3rd energy level ($n=3$) has three sub energy levels; s, p and d, while the 4th energy level ($n=4$) until the 7th energy level have four sub energy levels s, p, d and f.

- Orbitals are regions or volumes of space around the nucleus within which there is a high probability of finding electrons. ✓

(b)(i) Aufbau Principle states that, "electrons are arranged in a free atom in such a way that they fill lowest orbitals, sub energy levels and energy levels first before they fill higher ones". ✓

In nitrogen's atom, the 1s orbital is lowest in energy and filled first, followed by the 2s orbital and the 2p orbitals are filled last. ✓ $1s^2 2s^2 2p^3$

Pauli Exclusion Principle states that, "electrons are arranged in a free atom in such a way that an orbital can accommodate a maximum of two electrons which must spin in opposite directions." ✓ or "electrons are arranged in a free atom such that no two electrons can have the same values for quantum numbers."

In nitrogen's atom, the electrons in the 1s orbital are arranged as  but not  ✓

Hund's Rule states that, "electrons are arranged in a free atom where there are orbitals of equal or similar energy in such a way that electrons are filled in all the orbitals singly with parallel spins before pairing them up with opposite spins." ✓

In nitrogen's atom, the electrons in the p-orbitals are arranged as  but not  ✓

(b)(ii) Some configurations are more stable due to having either half-filled sub-energy levels or fully filled sub-energy levels.

Chromium (24) has an electronic structure of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ but not $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$ as expected. ✓

This is because the 4s and 3d orbitals are half-filled ✓ and hence stable. ✓

Copper (24) has an electronic structure of $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ but not $1s^2 2s^2 2p^6 3s^2 3p^6 3d^9 4s^2$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$ as expected. ✓

This is because it has a half-filled 4s orbital and fully-filled 3d orbitals ✓ which are stable. ✓

(c) Aluminium (Al, $Z = 13$) is $1s^2 2s^2 2p^6 3s^2 3p^1$ ✓

Phosphide ion (P^{3-} , $15 + 3 = 18 e^-$) is $1s^2 2s^2 2p^6 3s^2 3p^6$ ✓

Nickel (Ni, $Z = 28$) is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$ or $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$ ✓

Vanadium(III) ion (V^{3+} , $Z = 23 \rightarrow$ loses $3 e^-$) is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2$ ✓

(d)(i) Alkaline earth metals (Group II elements).

- The reactivity increases down the group II elements. ✓ This is because, the atomic radius (size) increases and the valence (outermost) electrons become further from nucleus. ✓ Thus, the valence electrons are lost more easily.

(ii) Halogens (Group VII elements).

- The reactivity decreases down the group. ✓ This is because, the atomic radius (size) increases and the gained electrons become further from nucleus. ✓ Thus, the electrons are gained less easily.

(iii) Noble gases (Group VIII elements).

- They are not reactive (are inert). ✓ This is because their electronic configuration has completely (fully) filled outermost energy levels and hence very stable. ✓

(25 scores)