Correcting configurations

This resource accompanies the infographic poster **Electron configurations** in *Education in Chemistry* which can be viewed and downloaded at: rsc.li/3mi1eui

Learning objectives

1. Give the electron configurations of atoms and ions using subshells and orbitals including Hund's rules.
2. Use different representations of electron configurations including noble gas cores and 'electrons in boxes'.
3. Relate the subshell containing the outermost electrons to an element's position in the periodic table.
4. Explain electron configurations in terms of ionisation energy trends.

This resource can be used to check progress during teaching of the topic, or for revision later on in the course. It could be teacher marked or used as a peer marked formative assessment activity.

Learners who succeed with question 1 should hopefully make the link between pre-16 representations of electron configuration and post-16 electron configurations as illustrated in the infographic.

Question 4 tests learners' knowledge about the formation of transition metal ions (electrons removed from 4s before 3d) and you should ask your learners to miss this out if it has not been covered.

Learners who can give the correct answers to questions 1–5 will have demonstrated they have met learning objectives 1 and 2. Question 2 relates to learning objective 3. Question 5 tests Hund's rules from learning objective 1.

Good answers to question 6–9 show learners have made progress towards learning objective 4.

Answers

The correct answer for each question is given here in bold along with a full explanation. Learners would not be expected to give their explanation of the wrong answer to this level of detail but should identify the problem. Full notes are given here to support any discussion of the answers.

1. **$1s^2\ 2s^2\ 2p^6\ 3s^1$**
   The student answer was the correct electronic configuration in terms of shells for sodium but did not show the detail of the subshells asked for in the question. The two electrons in the first shell are in the 1s orbital, the eight electrons in the

second shell are in the 2s (two electrons) and 2p (six electrons) subshells. The single electron in the third shell is in the 3s orbital.

2. **$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$**
   Sulfur is in the p-block of the periodic table. Its outermost electrons are in the p subshell not the d subshell. The 3p subshell is filled before the 3d subshell when we build up the electron configuration.

3. **[Ar] $4s^2$**
   The student answer gives the wrong noble gas core. [Ne] represents $1s^2\ 2s^2\ 2p^6$ but for calcium we need a noble gas core of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$, which is represented by [Ar].

4. **[Ar] $3d^3$ ($4s^0$) or [Ar] $3d^3$**
   The electron configuration of the vanadium atom is [Ar] $3d^3$ ($4s^2$) and the 4s electrons are the highest energy electrons, which are removed first when the V^{2+} ion is formed. Because the 4s electrons are removed first, it is usual to list the subshells in order with 3d before 4s.

5. Carbon:

1s	2s	2p		
↑↓	↑↓	↑	↑	

Or:

1s	2s	2p		
↑↓	↑↓	↑		↑

But not (although this is not commonly tested at this level):

1s	2s	2p		
↑↓	↑↓	↑	↓	

The electron configuration of carbon is $1s^2\ 2s^2\ 2p^2$, but the two electrons would not occupy the same orbital if another orbital in the same subshell is unoccupied. Two electrons in the same orbital experience greater electron-electron repulsion than two electrons in different orbitals in the same subshell. Additionally, the two electrons in different orbitals are lower in energy if they have the same spin, rather than opposite spin. This is because the Pauli exclusion principle means they stay further apart if they have the same spin. Hund's rule states that the lowest energy electron configuration, the ground state, in any subshell is the one with the greatest number of parallel electron spins (although this is not commonly tested at this level).

6. The ionisation energy increases from H to He because the nuclear charge has increased (from 1+ to 2+) while the shielding has not substantially increased and

the electron removed is in the same shell and subshell. The increased nuclear charge causes a greater attraction between the electron removed and the nucleus so the electron is harder to remove, requiring more energy. The student answer, suggesting that there is an extra electron, is wrong because an extra electron would add to the electron-electron repulsion, making it easier to remove. It would be easier to remove an electron for $H^-(g)$ than H(g). The ionisation energy of H(g) is +1310 kJ mol^{-1} and $H^-(g)$ is +72 kJ mol^{-1}. The shell being full does not cause an increase in the ionisation energy.

7. The first ionisation energy decreases as you go from Li to Na because the electron removed from Na is from the third shell rather than the second shell. The nuclear charge has increased but this is counteracted by the increased shielding from the full second shell. The student answer suggests they are from different subshells, they are both from s subshells (2s and 3s), the crucial difference is the shell they are in.

8. The first ionisation energy decreases as you go from Be to B because the electron removed from B is from the p subshell as opposed to the s subshell in Be. The p subshell is higher in energy, so the electron is easier to remove. An unpaired electron would be harder to remove than a paired electron if everything else was the same.

9. The first ionisation energy decreases as you go from N to O because the electron removed from O is sharing an orbital (spin-paired) and therefore experiencing more electron-electron repulsion than the unpaired electron in N. The electron removed from O is from the same subshell as the electron in N.

   The answer above is entirely appropriate for students at this level but teachers who are interested might look at: https://en.wikipedia.org/wiki/Hund%27s_rules#Rule_1 for a more in-depth explanation.