CHEM1611 Worksheet 2 – Answers to Critical Thinking Questions

The worksheets are available in the tutorials and form an integral part of the learning outcomes and experience for this unit.

Model 1: Atomic Orbitals

1. *s* orbitals consist of a single lobe. *p* orbitals consist of two large lobes. *d* orbitals consist of four lobes. (The shape of $3d_{z^2}$ is a little different to that of the other 3*d* orbitals. It consists of two large lobes with a ring around the centre.)

2. The orbital quantum number, *l*.

3. See table below.

n	Possible *l* values	Orbital labels
1	0 only	1*s*
2	0 and 1	2*s* and 2*p*
3	0, 1 and 2	3*s*, 3*p* and 3*d*

4. See table below.

5. See table below.

l	Possible m_l values	Number of values of m_l	Number of orbitals	Maximum number of electrons
0	0	1	1	1 × 2 = 2
1	1, 0 and -1	3	3	3 × 2 = 6
2	2, 1, 0, -1 and -2	5	5	5 × 2 = 10

Model 2: Electronic Configurations of Atoms

1. (a) $1s^2\ 2s^1$ or [He] $2s^1$ (b) $1s^2\ 2s^2$ or [He] $2s^2$
(c) $1s^2\ 2s^2\ 2p^1$ or [He] $2s^2\ 2p^1$ (d) $1s^2\ 2s^2\ 2p^2$ or [He] $2s^2\ 2p^2$
(e) $1s^2\ 2s^2\ 2p^3$ or [He] $2s^2\ 2p^3$ (f) $1s^2\ 2s^2\ 2p^4$ or [He] $2s^2\ 2p^4$
(g) $1s^2\ 2s^2\ 2p^5$ or [He] $2s^2\ 2p^5$ (h) $1s^2\ 2s^2\ 2p^6$ or [He] $2s^2\ 2p^6$

2. $1s^2\ 2s^2\ 2p^6\ 3s^1$ or [Ne] $3s^1$. Both Na and Li have 1 electron in an *s* orbital outside a noble gas configuration.

3. 2*s* is lower in energy than 2*p*. (2*s* is filled before 2*p* is.).

4. The arrows refer to m_s - the spin quantum number. According to the Pauli Principle, a maximum of one 'up' and one 'down' spin (arrow) can occupy each orbital (box).

5. Hund's Rule.

6. See below.

Atom	Electron configuration	Representation – only valence electrons
Be	$1s^2\ 2s^2$ or [He] $2s^2$	$\boxed{\uparrow\downarrow}\quad\boxed{}\boxed{}\boxed{}$
N	$1s^2\ 2s^2\ 2p^3$ or [He] $2s^2\ 2p^3$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$
Ne	$1s^2\ 2s^2\ 2p^6$ or [He] $2s^2\ 2p^6$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$ or [Ne] $3s^2\ 3p^1$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow}\boxed{}\boxed{}$
Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ or [Ne] $3s^2\ 3p^5$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow}$

7. (i) It is quicker, (ii) it makes the valence electrons obvious and (ii) it shows the relationship between the valence configurations of elements.

The X atoms are the noble gases.

8. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$ or [Ar] $4s^2\ 3d^6$.

9. $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$ or [Ar] $4s^2\ 3d^{10}\ 4p^3$

Model 2: Electronic Configurations of Atoms

1. The configurations of Mg^{2+} and O^{2-} are the same and match the preceding and proceeding noble gas respectively.

Atom / Ion	Electron configuration	Representation – only valence electrons
O	$1s^2\ 2s^2\ 2p^4$ or [He] $2s^2\ 2p^4$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$
O^{2-}	$1s^2\ 2s^2\ 2p^6$ or [He] $2s^2\ 2p^6$	$\boxed{\uparrow\downarrow}\quad\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}$
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$ or [Ne] $3s^2$	$\boxed{\uparrow\downarrow}\quad\boxed{}\boxed{}\boxed{}$
Mg^{2+}	$1s^2\ 2s^2\ 2p^6$ or [Ne]	$\boxed{}\quad\boxed{}\boxed{}\boxed{}$

2. (a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ or [Ne] $3s^2\ 3p^6$
(b) $1s^2\ 2s^2\ 2p^6\ 3s^2$ or [Ne]
(c) It does not have one! $1s^0$
(c) $1s^2$

Extension questions

1. As H is $1s^1$ and He is $1s^2$, a Periodic Table based purely on electronic configurations would have the elements in the *s*-block, above Li and Mg respectively.

   The Periodic Table, however, was worked out from experimental observations of chemical and physical properties. He is an unreactive gas and for this reason, it is placed in the Nobel Gases. This can perhaps be justified from its electronic configuration as it has the maximum number of electrons possible for an element with $n = 1$ just as Ne has the maximum number of electrons possible for an element with $n = 2$.

   H is more awkward! Chemically it behaves a little like the Group 1 elements – its commonest oxidation number is +1, just like the Group 1 elements. However, compounds containing it with this oxidation number are very different to those of Group 1 elements: LiOH, NaOH and KOH contain M^+ and OH^- ions and are strongly basic. HOH is a covalent molecule and is not basic.

   In its elemental form, hydrogen exists as a gas made up of H_2 molecules whereas the Group 1 elements exist as metallic solids. (Note however that at *very* high pressure such as in the cores of planets, hydrogen probably is also metallic).

   In forming the diatomic H_2 molecule, hydrogen acts more like the halogens (F_2, Cl_2, Br_2 and I_2). When it reacts with metals, it forms hydrides such as LiH which contain H^- ions. This is akin to the halogens which also forms anions in compounds like LiF and NaCl.

   There is no correct answer!

2. There are a number of trends:

   (a) Ionisation energies *decrease* down each group as the electron to be removed is in an orbital with a higher *n* quantum number and orbits further from the nucleus. As a result, the attraction to the nucleus is reduces.

   (b) Ionisation energies generally *increase* across a period as the *n* quantum number does not change but the positive charge of the nucleus increases. As a result, the attraction to the nucleus increases.

   The combination of (a) and (b) means that the first element in each row (H, Li, Na) have the lowest ionisation energy in that row whilst the last element in each row (He, Ne and Ar) have the highest ionisation energy in that row.

   (c) Trend (b) is broken between the between Be and B and between Mg and Al. This is due to the change in the orbital of the electron which is being removed. When Be ($2s^2$) is ionised, an *s*-electron is removed. When B ($2s^2 2p^1$) is ionised, a *p* electron is removed. As 2*p*-orbitals have higher energies than 2*s* orbitals, it is easier to ionise B than Be *despite* the nuclear charge being higher.

   (d) Trend (b) is also broken between N and O and between P and S. This is due to the change in the spin of the electron being ionised. In N ($2s^2 2p^3$), the *p*-electron being ionised has the same spin as the other *p*-electrons. In O, ($2s^2 2p^4$), the *p*-electron being ionised has the opposite spin as the other *p*-electrons. (Draw the electronic configurations using the box notation to see this.)

   Electrons with the opposite spin repel each other more than electrons with the same spin. It is then easier to remove an electron from O.