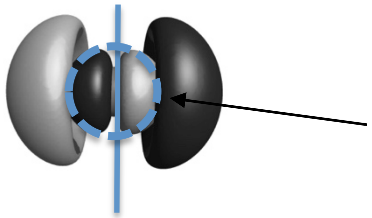
Marks

- A schematic representation of a *p* orbital is shown below. The central sphere (mostly obscured) represents the atomic nucleus. **2**



How many spherical and planar nodes does this orbital have? Label them on the diagram above.

Number of spherical nodes: **1**	Number of planar nodes: **1**

What is the principal quantum number, *n*, of this orbital? Explain your answer.

***n* = 3**

The total number of nodes is 1 fewer than the principal quantum number. As the total number of nodes is 1 + 1 = 2, the principal quantum number is 3.

- Shielding is important in multi-electron atoms. Briefly explain the concept of shielding. **3**

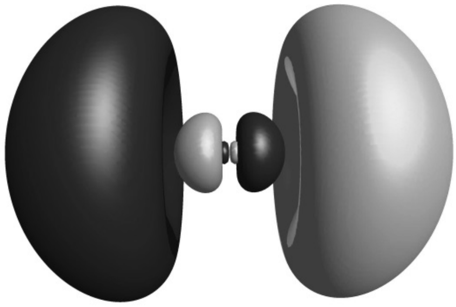
Electrons closer to the nucleus partially block the attractive force of the nucleus on the electrons that are further away, resulting in a lowering of the effective nuclear charge on such electrons.

Give one example of a consequence of shielding.

The elements in a group of the Periodic Table have similar reactivities but ionisation energies decrease and sizes increase.

Marks
3

- Consider the 4*p* orbital shown below. Note that, for clarity, the nucleus of the atom is not shown.



How many spherical and planar nodes does this orbital have?

Number of spherical nodes: **2**	Number of planar nodes: **1**

Complete the following table to give a set of quantum numbers that describes an electron in a 4*p* orbital.

Quantum number	n	l	m_l	m_s
Value	4	**1**	**-1, 0 or +1**	**½ or -½**

Marks
4

- Write down the ground state electron configurations for the following species. Na is given as an example.

Na	[Ne] $3s^1$
K	**[Ar] $4s^1$**
As	**[Ar] $4s^2$ $3d^{10}$ $4p^3$**
Sr	**[Kr] $5s^2$**
C^+	**[He] $2s^2$ $2p^1$**

Name the elements described by the following configurations.

[Kr] $5s^2$ $4d^6$	**ruthenium**
[Xe] $6s^2$ $5d^1$ $4f^{11}$	**erbium**

Marks

- Name the element described by the following configuration. **1**

[Kr] $5s^2\ 4d^{10}$	**Cadmium**

- Write out the valence electron configuration of the following anions and in each case explain why the anion is less stable than the separated atom and electron. **4**

Ne^-

$1s^2\ 2s^2\ 2p^6\ 3s^1$

Ne has a noble gas configuration. The extra electron needs to go into the 3*s* orbital which is in the next shell: it is high in energy as the electron is far from the nucleus.

N^-

$1s^2\ 2s^2\ 2p^4$

N has all 3 electrons in different *p* orbitals with parallel spins. Adding an extra electrons forces one of these electrons to become paired which is a higher energy situation.

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks 4

- The "Paschen" series of emission lines corresponds to emission from higher lying energy states to the $n = 3$ state in hydrogen-like atoms. Calculate the wavelength (in nm) of the lowest energy "Paschen" emission line in Li^{2+}.

The energy of an orbital in an 1-electron atom or ion is given by

$$E_n = -Z^2 E_R (1/n^2)$$

The energy difference between two levels is therefore:

$$\Delta E = E_{n1} - E_{n2} = [-Z^2 E_R (1/n_1^2)] - [-Z^2 E_R (1/n_2^2)] = Z^2 E_R (1/n_2^2 - 1/n_1^2)$$

The lowest energy line in the Paschen series corresponds to moving from $n = 4$ to $n = 3$. As Li^{2+} has $Z = 3$, the energy of this transition is therefore:

$$\Delta E = (3)^2 E_R (1/3^2 - 1/4^2) = 9.54 \times 10^{-19} \text{ J}$$

Using $E = hc / \lambda$, this corresponds to a wavelength of:

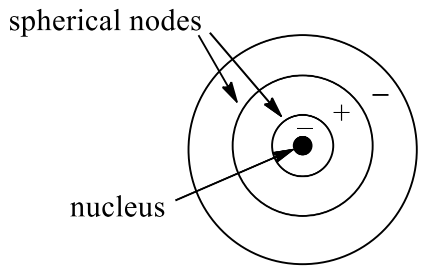
$$\lambda = hc / E = (6.626 \times 10^{-34} \text{ J s})(2.998 \times 10^8 \text{ m s}^{-1}) / (9.54 \times 10^{-19} \text{ J}) = 2.08 \times 10^{-7} \text{ m} = 208 \text{ nm}$$

Answer: **208 nm**

What are the possible l states for the $n = 4$ level of Li^{2+}? ***l* = 0, 1, 2 and 3**

Sketch the atomic orbital with $n = 3$ and the lowest value of l.

The orbital is 3*s*:

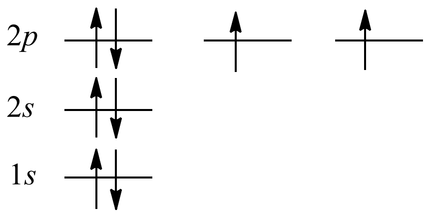


Marks
5

e) The oxygen atom in the reaction in part d) is formed in its ground electronic state. What is the ground state electronic configuration for O?

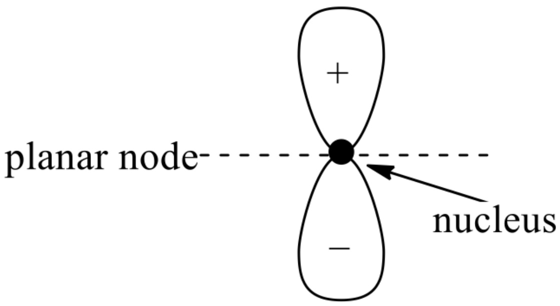
$\mathbf{1s^2\ 2s^2\ 2p^4}$

Draw an atomic orbital energy level diagram for the ground state O atom. Name the orbitals and show all electrons.



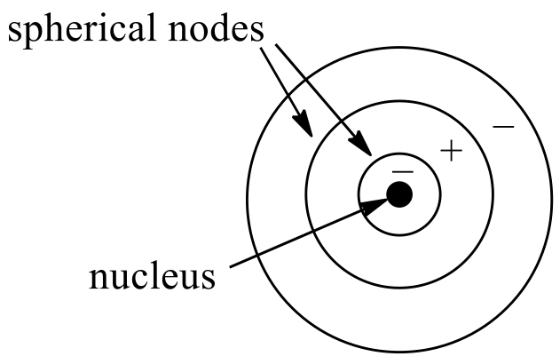
Name and sketch the atomic orbitals for the highest occupied atomic orbital and the lowest unoccupied atomic orbital in the ground state O atom. Make sure all nodes are clearly identified in your sketch.

sketch of highest occupied orbital



Name: **2*p* orbital**

sketch of lowest unoccupied orbital



Name: **3*s* orbital**

Marks
3

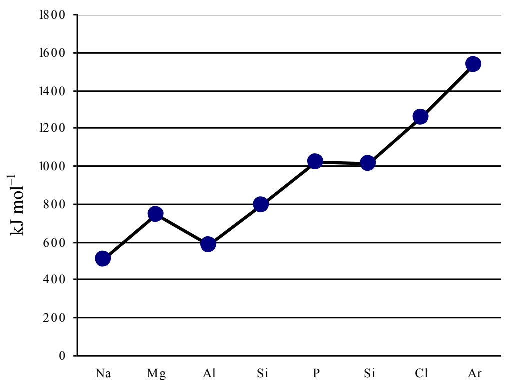
- Consider the values of the electronic energy levels of an He atom. State which interactions would be expected to increase the energies of the electrons and which would decrease them.

There are 3 electrostatic interactions:

- **Interaction between one electron and the 2 protons in the nucleus. This is attractive and lowers the energy of the electron.**
- **Interaction between the second electron and the 2 protons in the nucleus. This is attractive and lowers the energy of the electron.**
- **Interaction between the 2 electrons. This is repulsive and increases the energy of the electrons.**

Marks
4

- The graph shows the first ionisation energies for third row elements of the periodic table.



Explain the general trend and both anomalies.

There are a number of factors to explain:

- **There is an overall *increase* in ionization energy across the period. This is due to the increase in effective nuclear charge (Z_{eff}) which leads to greater pull from the nucleus and an increase in the energy required to remove an electron.**
- **There is a *decrease* in ionization between Mg and Al *despite* the latter having a higher Z_{eff}. Mg has an atomic configuration of [Ne] $3s^2$ and Al has an atomic configuration of [Ne] $3s^2\ 3p^1$. Thus, the electron being ionized in Mg is a $3s$ electron and the electron being ionized in Al is a $3p$ electron. The orbit of the $3p$ electron is, on average, greater than that of a $3s$ electron so the electron is held less tightly. Thus, Al has a lower ionization energy than Mg despite having a higher nuclear charge. Although a $3p$ electron is also being removed in Si, it has a higher ionization energy than Mg because the increase in nuclear charge is more than enough to compensate for the sub-shell change.**
- **Between P and S there is also a drop in ionization *despite* the latter having a higher Z_{eff}. Between P and S, the extra electron *has* to pair up:**

***p*-orbitals**

P	↑	↑	↑
S	↑↓	↑	↑

Electrons which are paired and occupy the same orbital are forced to be close to one another in space, leading to higher repulsion. Because of this, the electron which is paired up is easier to remove and the ionization energy is lower.

Marks

- Moseley discovered experimentally in 1913 that the atomic number, Z, of an element is inversely proportional to the square root of the wavelength, λ, of fluorescent X-rays emitted when an electron drops from the $n = 2$ to the $n = 1$ shell. **3**

$$i.e. \qquad \frac{1}{\sqrt{\lambda}} = kZ$$

What element would emit such X-rays with a wavelength one-quarter that of zirconium?

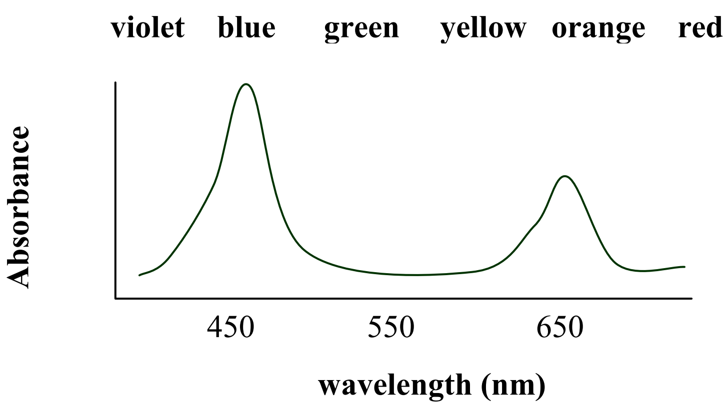
From Moseley's equation, the atomic number is inversely proportional to the square root of the atomic number. The ratio of the wavelengths emitted by two elements is therefore:

$$\frac{Z_1}{Z_2} = \sqrt{\frac{\lambda_2}{\lambda_1}}$$

As zirconium has $Z_1 = 40$ and $\frac{\lambda_2}{\lambda_1} = \frac{1}{4}$, the element has $Z = 80$. This corresponds to mercury.

Answer: **mercury**

- Many plants are green due to their high chlorophyll content. Draw on the diagram below the absorption spectrum of a green pigment such as chlorophyll. **2**



Marks

- Provide a brief explanation of each of the following terms. (You may include an equation or a diagram where appropriate). **4**

(a) Pauli exclusion principle

No two electrons may occupy the same orbital with the same spin, thereby having the same set of quantum numbers, *n*, *l*, m_l, *s* and m_s.

(b) the Bohr model of the atom

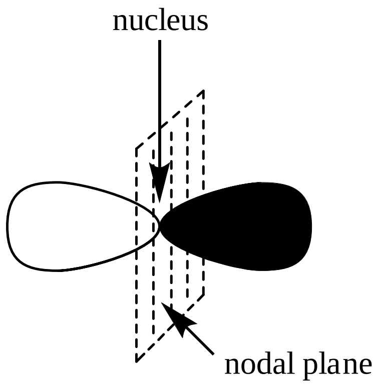
In the Bohr model, electrons in atoms occupy only discrete circular orbits. By being restricted to certain orbits, the energy of the electron can only have certain discrete values. The orbit occupied is labelled by a quantum number, *n*, which can only take integer values: *n* = 1, 2, 3….

- Write down the ground state electron configurations for the following elements. The configuration of lithium is given as an example. **2**

Li	$1s^2$ $2s^1$
Ne	**$1s^2$ $2s^2$ $2p^6$ or [He] $2p^6$**
Br	**$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^5$ or [Ar] $4s^2$ $3d^{10}$ $4p^5$**

- Sketch the following wave functions as lobe representations. Clearly mark all nodal surfaces and nuclear positions. **4**

(a) a 2*p* orbital



(b) a π molecular orbital