Atomic Structure

AQA A-LEVEL CHEMISTRY NOTES

ChemPrep

Background

Atomic structure is a fundamental topic in A-Level Chemistry as it forms a foundation for your understanding of the behaviour, chemical properties and interactions of elements, which will aid you in your understanding of complex topics.

Below are some <u>important</u> definitions for you to learn!

Atom: The smallest unit of an element that retains its chemical properties.

Nucleus: The central part of an atom containing protons and neutrons.

Proton: A positively charged subatomic particle found in the nucleus.

Neutron: A neutral subatomic particle found in the nucleus.

Electron: A negatively charged subatomic particle found in orbitals around the nucleus.

Isotope: Atoms of the same element with different numbers of neutrons.

Ionisation Energy: The energy required to remove one electron from an atom in its gaseous state.

Relative Atomic Mass (Ar): The weighted mean mass of an atom compared to 1/12th of the mass of a carbon-12 atom.

Mass Number (A): The total number of protons and neutrons in an atom.

Atomic Number (Z): The number of protons in the nucleus of an atom.

Within an atom there are three sub-atomic (fundamental) particles it is important for you to know the relative charges and masses for these particles

- An atom is composed of a nucleus which contains protons and electrons.
- Electrons move in fixed orbitals around the nucleus as shown in figure 1.

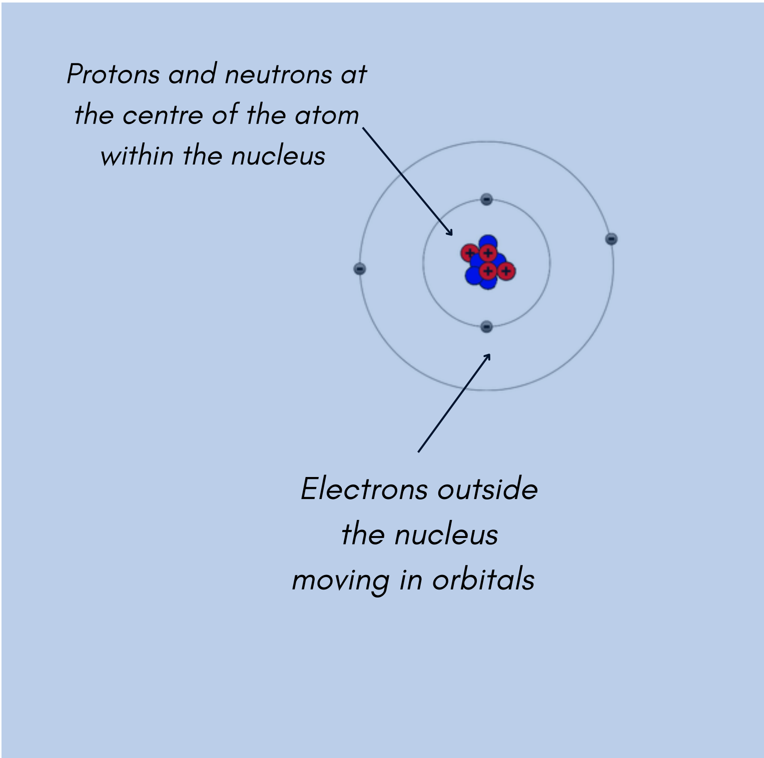


Figure 1

*Properties including **relative mass and charge** of the three fundamental particles*

Particle	Position	Relative Mass	Relative Charge
proton	nucleus	1	+1
neutron	nucleus	1	0
electron	orbitals	1/1840	-1

An element is usually displayed alongside its mass number and atomic number in the periodic table. It is important for you to understand what these terms mean and for you to be able to explain the existance of isotpes.

- The **mass number (A)** is the total number of protons and neutrons in the atom.
- The **atomic number (Z)** represents the number of protons, which also equals the number of electrons in a neutral atom.

Mass number (A):	24.3	**Mg**
Atomic number (Z):	12	

- **Isotopes** are atoms of the same element with the same number of protons, but different numbers of neutrons.
- As mass number is dependent on the total number of neutrons and protons present, isotopes of the same element also have different mass numbers.

2 isotopes of Hydrogen:

^{1}H ^{2}H

The difference in the number of neutrons is what causes the difference in mass number. The atomic number remains the same as this is only dependent on the number of protons

Remember: Don't confuse mass number and atomic number!.

Electron configuration is the arrangement of electrons in an atom. It is important for you to understand the idea of principle quantum shells, and how they are numbered by principle quantum numbers.

- Electrons are arranged in energy levels, **shells**, around the nucleus.
- These shells are numbered using principle quantum numbers, where the number of the shell corresponds to the energy of the shell.
- The **greater the energy** of a shell the **higher the principal quantum number**, as the **further away the shell sits from the nucleus**.

n = 1, n = 2 , n = 3, n = 4 , n = 5

n = principal quantum number of energy level, where n = 1 is the energy level closest to the nucleus and n = 5 is the energy level furthest from the nucleus.

- Shells are further split into sub-shells
- These are labelled: **s, p, d, f**
- Each sub-shell contains one or more atomic orbitals which exist at specific energy levels.
- Electrons can **only exist** within these specific levels
- Each orbital can hold up to **2 electrons** within each sub-shell having a certain number of orbitals

s = 1 orbital = holds up to 2 electrons
p = 3 orbitals = holds up to 6 electrons
d = 5 orbitals = holds up to 10 electrons
f = 7 orbitals = holds up to 14 electrons

Energy Level (n)	Sub-shell	Orbitals	Number of electrons
1	s	1,	2,
2	s, p	1, 3	2, 6
3	s, p , d	1, 3, 5	2, 6 , 10
4	s, p, d, f	1, 3 , 5 , 7	2, 6, 10, 14

Table showing the arrangement of electrons within an atom

Orbitals are filled in order of increasing energy. When filling atomic orbitals it is important for you to follow a set of rules.

1. **The Aufbau Principle**

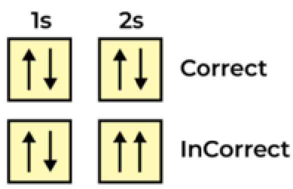
- The Aufbau principle states that electrons **fill the lowest energy orbitals first before moving to higher energy orbitals**

1s, 2s, 2p, 3s 3p, 4s (lower energy than 3d, so it fills first!), 3d, 4p, 5s, 4d

Order of filling of orbitals

2. **The Pauli Exclusion Principle**

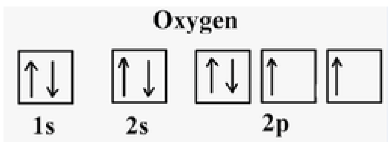
- The Pauli exclusion principle states that each orbital can **hold a maximum of two electrons** with opposite spins
- Arrows are used to represent electrons
- Arrows must drawn in **opposite directions, one up and one down, to display the opposite spins of the electrons.**



3. **Hunds rule**

- Hunds rule states that electrons will occupy **singly first before pairing**
- In the example of oxygen shown below, the electrons fill all empty orbitals singly before beginning to pair in the 1s orbital

Example: Oxygen contains the atomic orbitals: 1s, 2s, and 2p. The diagram below shows electrons are arranged to fill the orbitals following the rules we've discussed.



- *The 1s fills first then the 2s then the 2p*
- *The electrons show opposite spins once they pair up*
- *The p orbitals are half-filled as the 3 electrons in the 2p sub-shel fill all the empty 2p orbitals before pairing*

Below are two examples of how to show the full written electron configuration for an element:

Sodium (Na, Z=11): $1s^2$ $2s^2$ $2p^6$ $3s^1$
Chlorine (Cl, Z=17): $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$

First Ionisation Energy

- The first ionisation energy is defined as the energy required to remove one mole of electrons from one mole of gaseous atoms to form one mole of gaseous +1 ions.

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

Factors affecting ionisation energy

1. **Nuclear charge**

- The more protons within the nucleus the stronger the attraction between the outer electrons and the nucleus
- Therefore, the higher the nuclear charge the higher the ionisation energy as the more difficult it is to remove one mole of electrons from an atom

2. **Atomic radius**

- The greater the distance between the nucleus and the outermost electron the weaker the attraction
- Therefore, the larger the atomic radius the lower the ionisation energy as the more easier it is to remove one mole of electrons from an atom

3. **Shielding effect**

- The inner electron shells shield the outer electrons from the nuclear attraction
- The higher the number of electron shells the larger the shielding effect and the weaker the attraction between the outer electrons and the nucleus
- Therefore, the more shielding the lower the ionisation energy

4. **Pairing effect**

- Electrons in the same orbital repel one another
- Therefore the electrons in orbials that are paired are easier to remove and thus the lower the ionisation energy

Trends in Ionisation energy

- Ionisation energy increases across a period due to increased nuclear charge.
- Ionisation energy decreases down a group due to increased shielding and atomic radius.

Exam-Style Questions

Define the term first ionisation energy. (2 marks)

Write the full electron configuration for a sulfur atom (Z=16). (1 mark)

Explain why the first ionisation energy of oxygen is lower than that of nitrogen. (3 marks)

Magnesium has three isotopes: Mg-24 (79%), Mg-25 (10%), and Mg-26 (11%). Calculate its relative atomic mass. (3 marks)