Year 3 Chemistry
Student Notes

ATOMIC STRUCTURE

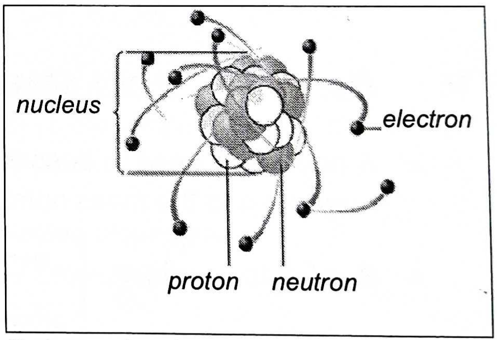
Learning Objectives

- state the relative charges and approximate relative masses of a proton, a neutron and an electron.
- describe, with aid of diagrams, the structure of an atom as containing protons and neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels), sub-shells and orbitals.
- state the electronic configuration of atoms and ions in terms of s, p, d orbitals for first 23 elements.
- define proton (atomic) number and nucleon (mass) number.
- interpret and make use of information presented in nuclides, e.g. $^{12}_{6}\mathrm{C}$.
- define the term isotopes.
- calculate relative atomic mass (A_r) from isotopic mass and relative isotopic abundances.
- deduce the number of protons, neutrons and electrons in atoms and ions, given the proton and nucleon numbers.

A. Structure and Particles

A1. Sub–Atomic Particles

- An **atom** is the smallest indivisible particle of an element that can take part in a chemical reaction.
- Atoms are made of three sub–atomic particles:
  - protons
  - neutrons
  - electrons



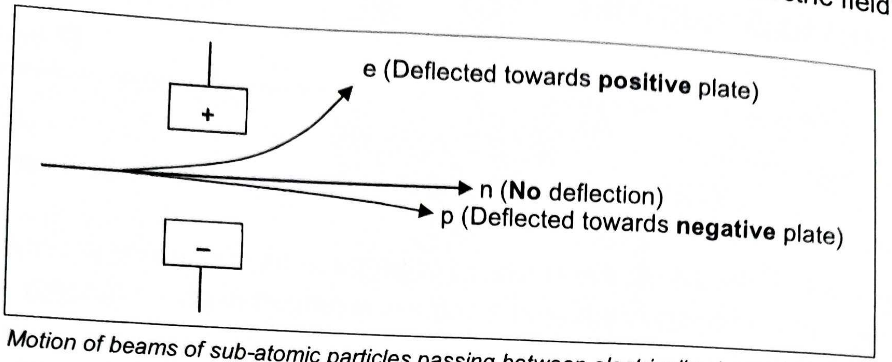
Protons and neutrons reside in the nucleus of the atoms.

Particle	Symbol	Position in atom	Relative charge	Mass / kg	Relative mass
Proton	*p*	In the nucleus	1+	1.67×10^{-27}	1
Neutron	*n*	In the nucleus	0	1.67×10^{-27}	1
Electron	*e*	Distributed around the nucleus	1-	9.11×10^{-31}	$\frac{1}{1840}$

Masses, charges and locations of sub-atomic particles

A2. Behaviour of particles in an electric field

How do protons, neutrons and electrons behave inside an electric field?



Motion of beams of sub-atomic particles passing between electrically charged plates.

- Protons and electrons are deflected in opposite directions.
    - protons are deflected towards the negative plate.
    - electrons are deflected towards the positive plate.
- neutrons (uncharged) are not deflected at all.

Why is the extent of deflection **different** for protons and electrons?

The extent of deflection is proportional to the charge over mass. As mass of an electron is much lower than that of a proton, the extent of deflection is greater for an electron.

B. Atomic Number and Mass Number

- A **nuclide** is used to describe an atomic species of which the atomic number (proton number) and the mass number (nucleon number) is specified.
- Symbol for nuclides: ${}^{\textbf{mass number}}_{\textbf{atomic number}}\textbf{Symbol}$

  E.g. carbon nuclide is represented as: ${}^{12}_{6}C$
- **Atomic number** (**proton number**) is the number of protons present in an atom.
- **Mass number** (**nucleon number**) is the total number of protons and neutrons in the atom.
- In an atom, the number of electrons is equal to the number of protons. Thus, atoms are said to be electrically neutral.
- The mass of the entire atom is largely attributed to the presence of protons and neutrons in the nucleus, as electrons have a negligible mass.

Example 1:

State the number of each type of sub-atomic particles represented by the nuclide, $^{27}_{13}Al$

No. of protons = 13

No. of neutrons = 27 − 13

= 14

No. of electrons = 13

C. Isotopes

- **Isotopes are atoms of the same element with different number of neutrons.**
    - Isotopes have the same atomic number, but different mass number.
- Isotopes of the same element have the same chemical properties, because they have the same number of electrons. However, they have slightly different physical properties (e.g. melting and boiling points, density), due to the different relative masses of the isotopes.

Example: ^{1}H, ^{2}H and ^{3}H are isotopes of hydrogen.

Isotope	**Hydrogen-1**	**Hydrogen-2**	**Hydrogen-3**
Common name	Hydrogen	Deuterium	Tritium
Mass number	1	2	3
Atomic number	1	1	1
Number of neutrons	0	1	2

- Most of the elements exist as many isotopes. Each of the isotopes exists in different **relative abundance.**
- Some common examples of isotopes (with their **relative abundance**):
    - Carbon – ^{12}C (98.9%), ^{13}C (1.1%), ^{14}C (trace amounts)
    - Oxygen – ^{16}O (99.76%), ^{17}O (0.04%), ^{18}O (0.2%)
    - Chlorine – ^{35}Cl (75.0%), ^{37}Cl (25.0%)
    - Silicon – ^{28}Si (92.2%), ^{29}Si (4.7%), ^{30}Si (3.1%)
    - Hydrogen -1H (99.9844%), ^{2}H (0.0156%), ^{3}H(trace amounts)
- Chlorine exists naturally as two isotopes, ^{35}Cl and ^{37}Cl. *Why is the relative atomic mass of chlorine in the Periodic Table 35.5?*
- Relative atomic mass, A_r, takes into consideration the mass of the natural occurring isotopes and their respective relative abundance.

Relative atomic mass of chlorine

$=\left(\frac{75}{100} \times 35\right) + \left(\frac{25}{100} \times 37\right)$

$= 35.5$

(Note: A_r has **no units** and is recorded to **1 decimal place**

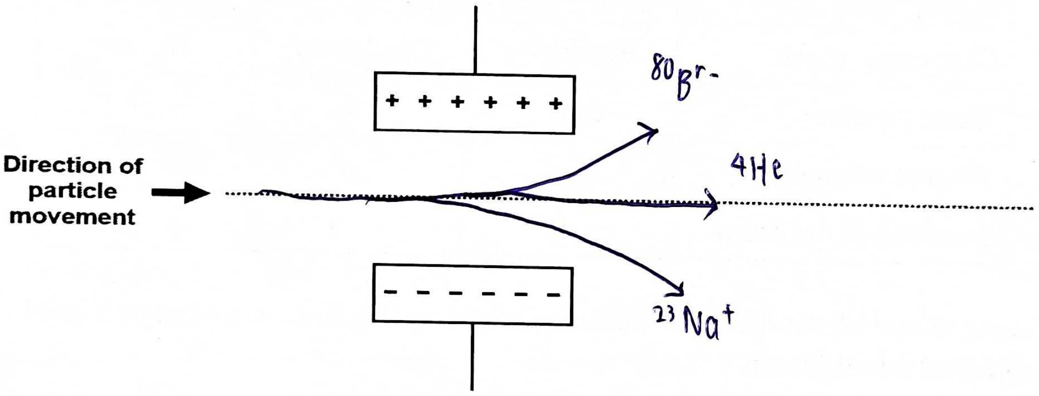
Quick Check Exercise (Part A)

1. Electrons and protons (subatomic particles) were fired with equal velocity into an electric field.
   Which row describe the behaviour of the electrons and protons?

	electrons	protons	relative amount of deflection
A	deflected towards (-) plate	deflected towards (+) plate	electrons deflected more than protons
B	deflected towards (-) plate	deflected towards (+) plate	protons deflected more than electrons
C	deflected towards (+) plate	deflected towards (-) plate	electrons deflected more than protons
D	deflected towards (+) plate	deflected towards (-) plate	protons deflected more than electrons

(C)

2. In the diagram below, draw the possible paths of the following particles if these particles were to separately move towards the metal plates in the direction as shown below:
   (a) a ^{4}He atom (b) a ^{23}Na^{+} ion (b) a ^{80}Br^{-} ion



3. Naturally occurring silver is 51.8% silver-107 and 48.2% silver-109. Calculate the relative atomic mass of silver.

Relative atomic mass of silver

$= (\frac{51.8}{100} \times 107) + (\frac{48.2}{100} \times 109)$

$= 107.964$

$= 108.0$ // (1d.p)

4. Copper consists of two isotopes, copper-63 and copper-65. Its relative atomic mass is 63.5. Find the **relative abundance** of each isotope.

Let the relative abundance of copper-63 be x

$(\frac{x}{100} \times 63) + (\frac{100-x}{100} \times 65) = 63.5$

$\frac{63x}{100} + \frac{6500-65x}{100} = 63.5$

$\frac{-2x+6500}{100} = 63.5$

$-2x + 6500 = 6350$

$2x = 150$

$x = 75$

Relative abundance of copper-63 = 75 //

Relative abundance of copper-65

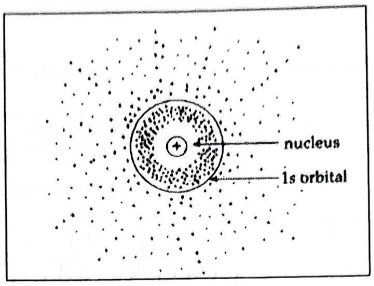
$= 100 - 75$

$= 25$ //

D. Atomic Orbitals

- Electrons do not circulate the nucleus in a fixed path. Instead, it occupies a certain region around the nucleus.

- An orbital is a <u>region of space</u> around the nucleus with the <u>highest probability</u> of locating the electrons.

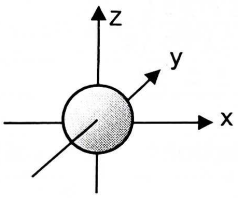


The position of the electron in a hydrogen atom is plotted every ten seconds over a period of time.

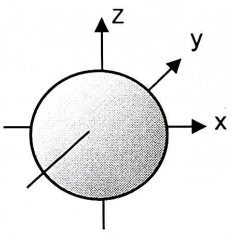
- Electrons can occupy four types of orbital, which differ from each other in <u>shape</u> and in their <u>orientation</u> in space. These are called s, p, d and f orbitals.

D1. s orbitals

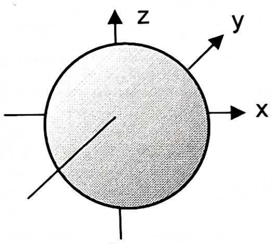
- All s orbitals are ...spherical... in shape. They only differ in size.



1s



2s

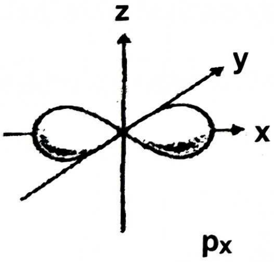


3s

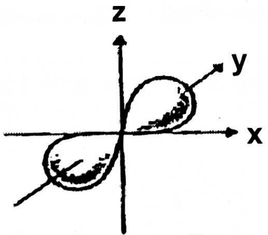
D2. p orbitals

- There are 3 types of p orbitals (p_x, p_y, p_z) with different **orientations** in space

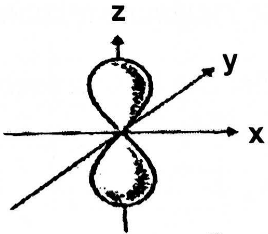
- Each p orbital has a ...dumbell.... shape.



p_x



p_y



p_z

E. Arrangement of Electrons in Atoms

- Electrons are found around the nucleus of an atom in definite energy levels (electron shells).
  - These energy levels are identified by a number, n, where n = 1,2,3...
  - The higher the value of n, the higher the energy level of the shell.
  - The maximum number of electrons that can occupy each shell is $2n^2$.

- These energy levels can consist of sub-energy levels (subshells), labelled as s,p,d,f

Electron shell, n	Number of subshells
1	1 (1s)
2	2 (2s and 2p)
3	3 (3s, 3p, 3d)
4	4 (4s, 4p, 4d, 4f)

- Each subshell contains a number of orbitals, in which electrons are placed.
  - Each orbital can hold a maximum of two electrons.

Type of subshell	Number of orbitals	Maximum number of electrons in each subshell
s	1	1 × 2 = 2
p	3	3x2=6
d	5	5x2=10
f	7	7x2=14

- The table below outlines how the maximum number of electrons in each shell is derived.

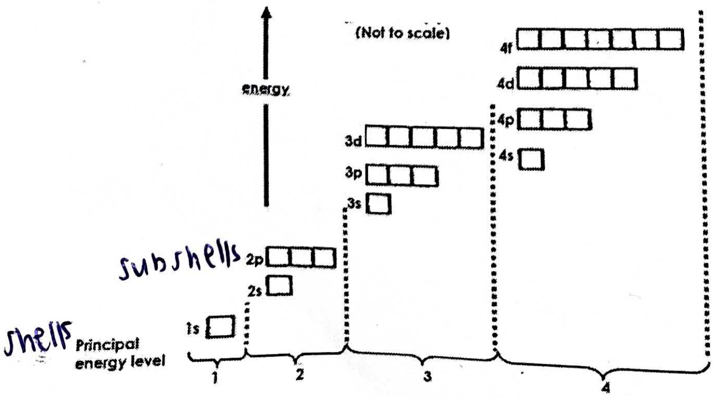
Electron Shell, n	Number of subshells	Type of subshells	Maximum number of electrons in each shell
1	1	1s	2
2	2	2s, 2p	2+6=8
3	3	3s, 3p, 3d	2+6+10=18
4	4	4s, 4p, 4d, 4f	2+6+10+14=32

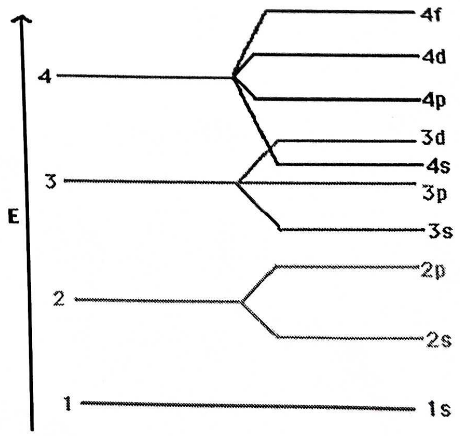
  - The maximum number of electrons that can occupy each shell is $2n^2$.

- Each orbital can be represented by a square box. Hence the three 2p orbitals are represented as:

2p ☐☐☐

Relative energy levels of the subshells and order of filling of subshells

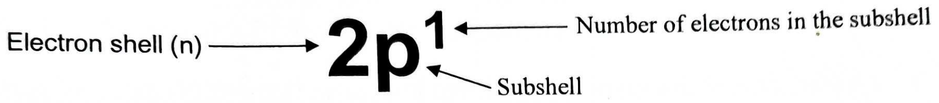




- The orbitals are filled in the order of ..increasing energy....... (starting with the 1s orbital).
- Electrons always go into an empty orbital with the ......lowest.... energy first.
- **IMPORTANT:** 4s has a **lower** energy level than 3d. Thus, when adding electrons, the 4s orbital is filled up before the 3d orbitals.

F. Electronic Configuration of Atoms

- The electronic configuration of a particle provides information about how the electrons are arranged in their electron shells and subshells.
- Notation for electronic configuration:



Element	Atomic number	Number of electrons	Electron shells and subshells							Electronic configuration
			1	2		3			4	
			1s	2s	2p	3s	3p	3d	4s	
H	1	1	1							$1s^{1}$
He	2	2	2							$1s^{2}$
Li	3	3	2	1						$1s^{2}\ 2s^{1}$
Be	4	4	2	2						$1s^{2}2s^{2}$
B	5	5	2	2	1					$1s^{2}2s^{2}2p^{1}$
C	6	6	2	2	2					$1s^{2}2s^{2}2p^{2}$
N	7	7	2	2	3					$1s^{2}2s^{2}2p^{3}$
O	8	8	2	2	4					$1s^{2}2s^{2}2p^{4}$
F	9	9	2	2	5					$1s^{2}2s^{2}2p^{5}$
Ne	10	10	2	2	6					$1s^{2}2s^{2}2p^{6}$
Na	11	11	2	2	6	1				$1s^{2}2s^{2}2p^{6}3s^{1}$
Mg	12	12	2	2	6	2				$1s^{2}2s^{2}2p^{6}3s^{2}$
Al	13	13	2	2	6	2	1			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{1}$
Si	14	14	2	2	6	2	2			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{2}$
P	15	15	2	2	6	2	3			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{3}$
S	16	16	2	2	6	2	4			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{4}$
Cl	17	17	2	2	6	2	5			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{5}$
Ar	18	18	2	2	6	2	6			$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}$
*K	19	19	2	2	6	2	6		1	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{1}$
*Ca	20	20	2	2	6	2	6		2	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}4s^{2}$
Sc	21	21	2	2	6	2	6	1	2	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{1}4s^{2}$
Ti	22	22	2	2	6	2	6	2	2	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{2}4s^{2}$
V	23	23	2	2	6	2	6	3	2	$1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{3}4s^{2}$

4s has a* ***lower* *energy level than 3d. Thus, when adding electrons, the 4s orbital is filled up before the 3d orbitals.*

G. The Periodic Table

- The elements in the Periodic Table are arranged in order of increasing proton number
- The vertical columns of elements in the Periodic Table are called groups.
  - Numbers from 1 to 18 are to be used as the Group Number.
  - Elements in the same group (1, 2, 13 to 17) have the same number of outer shell electrons (valence electrons).

Example:

Group 1	Proton number	Electronic configuration	Number of valence electrons
Lithium	3	$1s^2 2s^1$	1
Sodium	11	$1s^2 2s^2 2p^6 3s^1$	1
Potassium	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$	1

- The horizontal rows of the elements in the Periodic Table are called periods.
- All elements in the same period have the same number of occupied electron shells. The period number is given by the number of occupied electron shells.

Example:

Period 2	Li	Be	B	C	N	O	F	Ne
Proton number	3	4	5	6	7	8	9	10
Electronic configuration	$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p^5$	$1s^2 2s^2 2p^6$

H. Electronic Configuration of the Noble Gases

- The noble gases in Group 18 (Helium, Neon, Argon, Krypton, Xenon and Radon) tend to be unreactive.
- As the electron arrangements of noble gases are relatively stable, they exist naturally as monatomic elements.
- Atoms of noble gases (except helium) have eight electrons in their outer shells (stable octet configuration).

$$\text{He - } \mathbf{2} \longrightarrow \text{duplet configuration}$$

$$\left.\begin{array}{l} \text{Ne - } 2,\mathbf{8} \\ \text{Ar - } 2,8,\mathbf{8} \\ \text{Kr - } 2,8,18,\mathbf{8} \\ \text{Xe - } 2,8,18,18,\mathbf{8} \\ \text{Rn - } 2,8,18,32,18,\mathbf{8} \end{array}\right\} \text{octet configuration}$$

- In general, when atoms react to form compounds, they tend to combine in such a way that they each have eight electrons in their outer shell, giving them the same electronic configuration as a noble gas. This is called octet rule.

I. Formation of Ions

- An ion is formed when an atom loses or gains electron(s), such that it acquires an electrical charge (as the number of protons and electrons are no longer equal).
- Atoms lose or gain electrons to obtain the duplet or octet configuration in order to achieve greater stability.
- When atoms gain electrons, ..negative...... ions (...anions.........) are formed.
    - Electrons are first added into available orbitals with the ...lowest............ energy.

Symbol	Particle	Number of electrons	Electron shells and subshells							Electronic configuration
			1	2		3			4	
			1s	2s	2p	3s	3p	3d	4s	
F	Atom	9	2	2	5					$1s^2\ 2s^2\ 2p^5$
F^-	Ion	10	2	2	6					$1s^2\ 2s^2\ 2p^6$
Mg	Atom	12	2	2	6	2				$1s^2\ 2s^2\ 2p^6\ 3s^2$
Mg^{2+}	Ion	10	2	2	6					$1s^2\ 2s^2\ 2p^6$

- When atoms lose electrons, ...positive............ ions (...cations..........) are formed.
    - Electrons are first removed from orbitals with the ...highest...... energy.
    - **IMPORTANT:** Once the 4s orbital is filled, the energy level of the 4s orbital will increase above that of the 3d orbitals. Therefore, during the formation of cations, the electrons will be removed from the 4s orbital before the 3d orbitals.

Symbol	Particle	Number of electrons	Electron shells and subshells							Electronic configuration
			1	2		3			4	
			1s	2s	2p	3s	3p	3d	4s	
S	Atom	16	2	2	6	2	4			$1s^2 2s^2 2p^6 3s^2 3p^4$
S^{2-}	Ion	18	2	2	6	2	6			$1s^2 2s^2 2p^6 3s^2 3p^6$
K	Atom	19	2	2	6	2	6		1	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
K^+	Ion	18	2	2	6	2	6			$1s^2 2s^2 2p^6 3s^2 3p^6$
Sc	Atom	21	2	2	6	2	6	1	2	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$
^Sc^{2+}	Ion	19	2	2	6	2	6	1		$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1$
^Sc^{3+}	Ion	18	2	2	6	2	6			$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

^ The electrons will be removed from the 4s orbital before the 3d orbitals.

J. Isoelectronic Species

- Any atom or ions with .the same number of electrons.......... are said to be isoelectronic.

Examples:

Isoelectronic species	Number of electrons	Electronic configuration
F^-, Ne and Mg^{2+}	10	$1s^2 2s^2 2p^6$
S^{2-}, Ar, K^+ and Sc^{3+}	18	$1s^2 2s^2 2p^6 3s^2 3p^6$