

St Andrew's Junior College H1 Chemistry 2021 Lecture Notes 2 Atomic Structure

Assessment Objectives:

Candidates should be able to:

- (a) identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses
- (b) deduce the behaviour of beams of protons, neutrons and electrons in an electric field
- (c) describe the distribution of mass and charges within an atom
- (d) deduce the numbers of protons, neutrons and electrons present in both atoms and ions given proton and nucleon numbers (and charge)
- (e) (i) describe the contribution of protons and neutrons to atomic nuclei in terms of proton number and nucleon number

(ii) distinguish between isotopes on the basis of different numbers of neutrons present

- (f) describe the number and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2 and 3 and also the 4s and 4p orbitals
- (g) describe the shapes of s and p orbitals[note : knowledge of wave functions is **not** required]
- (h) state the electronic configuration of atoms and ions given the proton number (and charge)
- (i) explain the factors influencing the ionisation energies of elements (see the *Data Booklet*) (see also topic: The Periodic Table)
- (j) deduce the electronic configurations of elements from successive ionisation energy data
- (k) interpret successive ionisation energy data of an element in terms of the position of that element within the Periodic Table

The Periodic Table

Assessment Objective:

For elements in the third period (sodium to chlorine), and in Group 2 (magnesium to barium) and Group 17 (chlorine to iodine) candidates should be able to:

- (a) recognise variation in the electronic configuration across a Period and down a Group
- (b) describe and explain qualitatively the general trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity
 - (i) across a Period in terms of shielding and nuclear charge
 - (ii) down a Group in terms of increasing number of electronic shells, shielding and nuclear charge

Content:

- 1. Atomic Structure
- 2. Principal Quantum Numbers, Subshells and Orbitals
- 3. Electronic Configurations
- 4. Periodic Properties
- 5. Experimental Determination of Ionisation Energies

St Andrew's Junior College

1 Atomic Structure

1.1 History of the Atomic Model



© COMPOUND INTEREST 2016 - WWW.COMPOUNDCHEM.COM | Twitter: @compoundchem | Facebook: www.facebook.com/compoundchem This graphic is shared under a Creative Commons Attribution-NonCommercial-NoDerivatives licence.



https://www.youtube.com/watch?v=xazQRcSCRaY

The development and revision of the atomic model over time is due to new evidence found through advances in theory and technology. This does not necessarily mean that the past atomic models are inaccurate, they are still useful for explaining science in a simpler manner.

1.2 The Fundamental Sub-Atomic Particles of Matter

- (a) identify and describe protons, neutrons and electrons in terms of their relative charges and relative masses
- (c) describe the distribution of mass and charges within an atom
 - An atom is the smallest particle of an element that retains its chemical properties.
 - An atom consists of a dense positively charged nucleus made up of **protons** and **neutrons**, collectively known as **nucleons**.
 - The nucleus is surrounded by negatively charged **electrons** occupying regions in space around it (known as atomic orbitals).
 - The electrons are attracted by protons via electrostatic forces.



Relative charges and masses of sub-atomic particles

Particle	Charge / C	Relative charge	Rest mass / kg	Relative mass	Location in atom
Proton (¹ ₁ p)	+1.60 x 10 ⁻¹⁹	+ 1	1.67 x 10 ⁻²⁷	1	In nucleus
Neutron (¹ ₀ n)	0	0	1.67 x 10 ⁻²⁷	1	In nucleus
Electron $\begin{pmatrix} 0 \\ -1 \end{pmatrix}$ e)	-1.60 x 10 ⁻¹⁹	- 1	9.11 x 10 ⁻³¹	1/1840	Around nucleus

1.3 Behaviour of beams of protons, neutrons and electrons in an electric field



Note: Deflection only occurs when the charged particles enter the electric field. Both p and e^- move in straight path as they leave the electric field.

• The angle of deflection is dependent on the relative **masses** and **charges** of the particles.



- The particle with a lighter mass and/or higher charge will have a greater angle of deflection, ie deflect more.
- Angle of deflection $\beta > \alpha$ as electrons have lighter mass than protons.

∽ Self- Check: Q1

1.4 Number of protons, neutrons and electrons in an atom

- (d) deduce the numbers of protons, neutrons and electrons present in both atoms and ions given proton and nucleon numbers (and charge)
- (e) (i) describe the contribution of protons and neutrons to atomic nuclei in terms of proton number and nucleon number
 - An atom of an element can be represented by



• The **proton number** (atomic number) is unique for each element. It gives the number of protons in the nucleus of an atom.

The **nucleon number** (**mass number**) gives the total number of protons and neutrons in the nucleus of an atom.

Eg. ${}_{1}^{1}H$ and ${}_{6}^{12}C$

In a neutral atom, the number of protons = the number of electrons.

When an atom gains electrons, a negatively charged ion or an **anion** is produced (number of protons<number of electrons).

When an atom loses electrons, a positively charged ion or a **cation** is produced (number of electrons<number of protons).

∽ Self- Check: Q2 to Q5

1.5 Isotopes

(e) (ii) distinguish between isotopes on the basis of different numbers of neutrons present

 Isotopes are atoms of the same element with the same number of protons (same proton number) but different number of neutrons (different nucleon numbers).
 For example,

	Isotopes of carbon		Isotopes of Hydrogen		
	¹² ₆ C	¹³ ₆ C	${}^1_1\mathbf{H}$	${}^{2}_{1}\mathbf{H}$	${}^{3}_{1}\mathbf{H}$
Number of nucleons	12	13	1	2	3
Number of protons	6	6	1	1	1
Number of neutrons	6	7	0	1	2

• Relative atomic mass of an element is often not a whole number. This arises due to the presence of isotopes occurring in different percentage abundance. (*Refer to Topic 1: The Mole Concept and Stoichiometry*)

∽ Self- Check: Q6

2 Principal Quantum Numbers, Subshells and Orbitals

- In an atom or ion, the electrons are arranged in different **shells**, which consist of one or more **subshells**.
- Each subshell contains one or more orbitals.

2.1 Principal Quantum Number, n (Shell)

- Electrons are arranged in **shells** with certain main energy levels around the nucleus.
- Each shell is described by a number known as the principal quantum number, n.

The value of n ranges from n = 1 to $n = \infty$



The first shell (n=1) is closest to the nucleus and it is of the lowest energy level.
 When the value of n increases, the electrons are in a shell that is further away from the nucleus and of greater energy. Hence, the electron is less tightly bound to the nucleus.

Principal quantum number, n	Maximum no. of electrons
1	2
2	2 + 6 = 8
3	2 + 6 + 10 = 18
4	2 + 6 + 10 + 14 = 32

• Some examples from n = 1 to n = 4 are listed below:

2.2 Subshell

- Each quantum shell contains one or more subshells.
- In the shell of principal quantum number, n, there are n subshells.
- The subshells are designated by n and a letter (s, p, d, f, etc.)
- Within the same shell, the relative energies of the subshells increase in the following order:

s

• Some examples from n = 1 to n = 4 are listed below:

Principal quantum	Number of	Names of
number, n	Subshells	Subshells
1	1	1s
2	2	2s 2p
3	3	3s 3p 3d
4	4	4s 4p 4d 4f

2.3 Orbital

- Each subshell consists of a number of orbitals.
- An orbital represents a **region in space** where there is a **high probability** (> 95%) **of finding an electron**.
- Each orbital can contain a maximum of **2 electrons** of opposite spins.
- The number of orbitals in each subshell depends on the type of subshell and is shown below:

Subshell	Number of Orbitals	Type of orbitals	Maximum number of electrons contained
S	1	S	2
р	3	p _x p _y p _z	2 x 3 = 6
d	5	d_{xy} d_{xz} d_{yz} d_{z}^{2} $d_{x^{2}-y^{2}}^{2}$	2 x 5 = 10

2.3.1 Shapes of s, p, d orbitals

(g) describe the shapes of s, p and d orbitals (see also topic : An Introduction to the Chemistry of Transition Elements)

s orbital

- For every shell, there is only **one s orbital**.
- All s orbitals are **spherical**, symmetrical and **non-directional**.





p orbital

- For shells with principal quantum number n = 2 and above, each p subshell has three p orbitals, namely p_x, p_y and p_z. The subscript provides information on the orientation of the orbital with respect to the orbital.
- All p orbitals are of **dumb-bell shape** and are **directional** (different orientation along the x, y and z axes).
- p orbitals in the same subshell (same principal quantum number) have the same energy i.e. they are **degenerate**.



• p orbitals from different shells have the same dumb-bell shape but different sizes. The size of $2p_x$ orbital is smaller than the size of $3p_x$ orbital which is more diffused.

Principal Quantum Number (n)	Type. of subshells	No. of orbitals	Maximum no. of electrons in shell (2n ²)
1	1s	1	2
2	2s 2p	1 + 3 = 4	8
3	3s 3p 3d	1 + 3 + 5 = 9	18
4	4s 4p 4d 4f	1 + 3 + 5 + 7 = 16	32

Summary to principal quantum numbers, subshells and orbitals:

2.4 Electronic Energy Level

(f) describe the number and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2 and 3 and also the 4s and 4p orbitals.

- Each electron occupies an orbital and possesses a certain amount of energy.
- A schematic representation of the energy levels of each orbital is shown below, indicating their relative energies:



In general,

- Energy level increases as principal quantum number increases.
- Within the same shell, the relative energies of the subshells increase in the following order:
 s
- Orbitals in the same subshell are degenerate.
- An electron in a lower energy level orbital is closer to the nucleus.

* Note: the 4s orbital is of a lower energy than the 3d orbitals

3 Electronic Configurations

(h) state the electronic configuration of atoms and ions given the proton number (and charge)

The **electronic configuration** of an element refers to the way in which the electrons are distributed among the various orbitals.

3.1 Representations of Electronic Configurations

There are a few ways to represent electronic configuration of an atom, namely:

- Quantum shell electrons notation:
 - e.g. ₁₁Na 2,8,1

(NOT accepted in A-Level)

In the A-Level, you must learn how to write the following:

• Electronic configuration using **spdf notation** (default):

e.g. ₁₁Na 1s²2s²2p⁶3s¹

- * Note : the number of electrons in the subshell is denoted as a superscript
- Electronic configuration using noble gas core:

e.g. 11Na [Ne] 3s1

where [Ne] represents 1s² 2s² 2p⁶

• Electrons-in-box notation:

e.g. ₁₁ Na	1s	2s	2р	3s
	1,	1,	1, 1, 1,	1

The orbitals are shown in order of their principal quantum number, followed by the subshells.

3.2 Rules of working out electronic configurations

When electrons in the atom occupy the lowest energy levels, the atom is said to be in its **ground state**. When an electron is promoted to one of the higher energy levels, the atom is unstable and is said to be in an **excited state**.

In order to determine the ground state electronic configuration of an atom of any element, there are three rules to follow:

Rule 1: Aufbau ('Build-Up') Principle

- Electrons are filled from the orbitals of lowest energy level to those of a higher energy level.
- Hence, the order in which the orbitals are filled are:

 $1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow \textbf{4s} \rightarrow \textbf{3d} \rightarrow 4p \quad etc.$

* Note: 4s has a lower energy than the 3d orbital

For example, the electronic configuration of boron is
 1s² 2s² 2p¹.



Boron atom has 5 electrons. The first 2 electrons must

occupy the 1s orbital as it has the lowest energy, the next two will occupy 2s orbital of higher energy and the remaining electron will occupy the 2p orbital

Exercise 1:		Criteria of Success Checklist
Write the electronic	c configuration to show the	The electronic configuration is written
electron distribution i	In $_{10}$ Ne and $_{17}$ Cl.	in order of principal quantum number, then
10 Ne	17C <i>l</i>	according to the subshell (s, p, d)
1s ² 2s ² 2p ⁶	1s ² 2s ² 2p ⁶ 3s ² 3p ⁵	with superscript to indicate the number of
		electrons in the subshell (e.g. 1s ²)

Rule 2: Hund's Rule

• The orbitals in the same subshell must be occupied **singly** and with **parallel** spins before they can be occupied in pairs. This is to ensure electrons are as far apart as possible to **minimise inter-electronic repulsion**.



Rule 3: Pauli Exclusion Principle

• An orbital can contain a maximum two electrons, and they must have opposite spins.



Exercise 2:		
Draw the electron-in box diag	ram to show the electron distribution in	10Ne and 17C <i>l</i> .
10Ne	17Cl	Thinking process: What are the 3 rules in writing electronic configurations?
		\circ Self- Check: Q7 and Q8

In writing the electronic configurations, **3d** should be written **before 4s** even though the 4s subshell is filled before 3d.

Example: Vanadium ($_{23}$ V) has the electronic configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

3.3 Anomalous electronic configurations

Chromium (24Cr) and copper (29Cu) have **anomalous** electronic configurations.

Chromium (24Cr)

Actual configuration: 1s² 2s² 2p⁶ 3s² 3p⁶ 3d⁵ 4s¹ or [Ar] 3d⁵ 4s¹



The energy difference between the 4s orbitals and 3d orbitals is minimal, so electrons occupy the 3d orbitals singly to **minimise inter-electronic repulsion** in the 4s orbital.

Copper (29Cu)

Actual configuration: 1s²2s²2p⁶3s²3p⁶3d¹⁰4s¹ or [Ar] 3d¹⁰4s¹



The half filled (e.g. Cr) and fully-filled 3d subshells (e.g. Cu) are unusually stable due to a **symmetrical distribution of electron density** around the nucleus.

3.4 Electronic configuration of ions

lons are formed when atoms gain or lose electrons, usually to attain the stable electronic configuration of a noble gas.

3.4.1 Formation of cations

To obtain a cation, electron(s) are **removed** from the orbitals with the **highest** energy.

Example:

particle	₁₂ Mg	₁₂ Mg ⁺	₁₂ Mg ²⁺
electronic configuration	1s ² 2s ² 2p ⁶ 3s ²	1s ² 2s ² 2p ⁶ 3s ¹	1s ² 2s ² 2p ⁶

For atoms/ions with 21 to 30 electrons (d-block), the **4s** electrons are **removed before** the **3d** electrons. This is because once the 3d orbitals are occupied by electrons, the electrons in 3d orbitals repel the 4s orbital to a higher energy level.

Example:

particle	₂₆ Fe	₂₆ Fe ³⁺	
electronic	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$	[Ar] 3d⁵	
configuration			
electron in box	3d 4s	3d 4s	
diagram of 3d and 4s subshell	161 1 1 1		

3.4.2 Formation of anions

To obtain an anion, electron(s) are **added** to the next available orbital.

Example:

particle	₁₅ P	15 P ^{3–}
electronic configuration	1s ² 2s ² 2p ⁶ 3s ² 3p ³	[Ne] 3s ² 3p ⁶
electron in box diagram of	3s 3p	3s 3p
3d and 4s subshell	11 1 1	16 16 16 16

Exerci Write down t	se 3: he electronic configuration of the ions listed below.	Th 1. 2.	inking Process : What is the electronic configuration of the neutral atom? Are electrons being
Particle	electronic configuration	3.	removed or added? Which orbitals are
₈ 0 ₂₀ Ca ²⁺	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶		they removed
₂₅ Mn ²⁺	1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁵		

3.5 Isoelectronic configurations

• Different species with the **same number of electrons** are said to be **isoelectronic** with each other.

- For example ${}_{11}Na^+$, ${}_{10}Ne$, ${}_{9}F$, ${}_{8}O^{2-}$ and ${}_{7}N^{3-}$ are isoelectronic with each other. They have 10 electrons.
- Examples of isoelectronic compounds include CO, N₂ and NO⁺.

3.6 Electronic Configurations and the Periodic Table

(a) recognise variation in the electronic configuration across a Period and down a Group

Elements in the Periodic Table are arranged in vertical columns (groups) and horizontal rows (periods). Atoms of elements belonging to the **same Group** have **similar electronic configurations** with the same number of valence electrons. The valence electrons are largely responsible for the atom's chemistry. Hence, elements in the same Group have similar chemical properties. In the Periodic table, elements are classified according to their orbitals in the outermost quantum shell. There are

- s block [Groups 1 & 2]
- d block [Transition elements]
- p block [Groups 13 to 18]
- f block [Lanthanoid & Actinold elements]

1s ¹																	1s²
2s¹	2s²											2 p ¹	2p²	2p ³	2 p ⁴	2p⁵	2p ⁶
3s¹	3s²											3p1	3p²	3p ³	3p⁴	3p⁵	3p ⁶
4s¹	4s²	3d¹	3d²	3d³	3d⁵	3d⁵	3d⁵	3d7	3d ⁸	3d ¹⁰	3d ¹⁰	4 p ¹	4p²	4p³	4p⁴	4p⁵	4p ⁶
5s1	5s²	4 d ¹	4 d²	$4d^4$	4d ⁵	4d ⁵	$4d^7$	4d ^s	4d ¹⁰	4 d ¹⁰	4 d ¹⁰	5p¹	5p²	5p³	5p4	5p⁵	5p ⁶
6s¹	<mark>6s</mark> ²	\setminus	5d²	5d ³	5d4	5d⁵	5d⁵	5d7	5d ⁹	5d ¹⁰	5d ¹⁰	6p¹	6p²	6p³	6p⁴	6p⁵	6p ⁶
7s¹	7s²	$ \rangle$	6d²	6d³	6d4	6d⁵	6de	6d ⁷	6d ⁸	6d ¹⁰	6d ¹⁰	7p¹	7p²	7p³	7p4	7p⁵	7p⁵
			5d1	4 f ¹	4f ³	4 f ⁴	4f ⁵	4f ⁶	4 f ⁷	4f7	4f ⁹	4f ¹⁰	4f ¹¹	4f ¹²	4f ¹³	4f ¹⁴	4f ¹⁴
		_ \	6d ¹	6d²	5f ²	5f ³	5f ⁴	5f ⁶	5f7	5f7	5f ⁹	5f ¹⁰	5f ¹¹	5f ¹²	5f ¹³	5f ¹⁴	5f ¹⁴

BExercise 4:

Write down the electronic configuration of the following atoms:

Element	Period no	Group No	Electronic configuration
Sodium, Na	3	1	[Ne] 3s ¹
Gallium, Ga	4	13	[Ar] 3d ¹⁰ 4s ² 4p ¹
lodine, I	5	17	[Kr] 4d ¹⁰ 5s ² 5p ⁵



Who came up with the Periodic Table and what is the significance of its discovery and development? https://www.youtube.com/watch?v=fPnwBITSmgU

4 Periodic Properties

Atomic Structure

(i) explain the factors influencing the ionisation energies of elements (see the *Data Booklet*) (see also topic: The Periodic Table)

The Periodic Table

- (b) describe and explain qualitatively the general trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity
 - (i) across a Period in terms of shielding and nuclear charge
 - (ii) down a Group in terms of increasing number of electronic shells, shielding and nuclear charge

4.1 Factors affecting trends in the Periodic Table

- (a) Nuclear Charge
 - Nuclear charge is the total positive charge of all the protons in the nucleus.
 - Hence nuclear charge increases with increasing number of protons.
 - The greater the nuclear charge, the stronger is the attraction for the valence electrons.

(b) Shielding Effect (Screening Effect)

• Shielding electrons are the inner shell electrons found in between the nucleus and the outermost electrons.



Shielding effect refers to the shielding of electrons from the electrostatic attraction of the positively-charged nucleus.

- The shielding electrons repel the outermost electrons which decreases the electrostatic attraction experienced by the outermost electrons.
- Electrons in the same shell exert a smaller shielding effect on one another.
- Hence shielding effect increases with increasing number of inner shell electrons.

(c) Effective Nuclear Charge

- The effective nuclear charge is the **net attraction** felt by the outermost electrons after taking into account **both** the nuclear charge and the shielding effect.
- The term "effective" is used because the shielding effect by the inner shell electrons prevents the valence electrons from experiencing the full nuclear charge of the nucleus.
 When increase in nuclear charge outweighs the increase in shielding effect, effective nuclear charge increases. Hence, there is stronger attraction between the nucleus and the valence electrons and vice versa.
- The greater the effective nuclear charge, the stronger is the net attraction of the outermost electrons to the positively charged nucleus and vice versa.

(d) Number of electronic shells

• As number of filled inner shells increases, the distance between the valence electrons and nucleus increases, resulting in weaker attraction between the valence electrons and nucleus.

4.2 Ionisation Energy (I.E.)

• 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 singly charged cation**.

first I.E.:	M (g)	\rightarrow	M+ (g)	+	e⁻
second I.E.:	M+ (g)	\rightarrow	M ²⁺ (g)	+	e⁻

Remember the definition and learn how to write the equation with state symbols. **Link to Energetics.**

• Ionisation energies are **endothermic** (i.e. have positive values) as energy is **required** to overcome the force of attraction between the electrons and nucleus.



4.2.1 Trend of I.E. across the period

- The number of protons increases across the period, hence nuclear charge increases.
- Successive electrons are being added to the same outermost shell thus increase in shielding effect is negligible.
- Since the increase in nuclear charge outweighs the increase in shielding effect, effective nuclear charge increases. Hence, the valence electrons are more strongly attracted to the nucleus.
- More energy is required to remove the valence electrons.
- Ionisation energies generally increase across the period.

4.2.2 Irregularities in I.E. trend

1^{st} I.E. of A*l* is lower than that of Mg

 $_{12}$ Mg : 1s² 2s² 2p⁶ 3s² $_{13}$ Al : 1s² 2s² 2p⁶ 3s² 3p¹

- In Mg, the first electron is removed from the **3s** orbital, whereas for A*l*, the first electron is removed from the **3p** orbital.
- The valence electron in the 3p orbital is **further** away from the nucleus than the 3s and experiences **additional shielding effect** by the 3s electrons.
- This factor **outweigh the effect of increase in nuclear charge** from Mg to Al.
- The valence electron in 3p orbital of A*l* experience weaker attraction to the nucleus and requires less energy to be removed.

(A similar explanation can be applied to Boron & Beryllium in Period 2.)

1st I.E. of S is lower than that of P



₁₆ S: 1s ² 2s ² 2p ⁶ 3s ² 3p ⁴							
3s		Зр					
11	16	1	1				

- In S, electron is removed from a paired electron occupying the same 3p orbital. Thus the electron experience additional inter-electronic repulsion.
- This factor outweighs the effect of increase in nuclear charge from P to S.
- The paired electron in 3p orbital of S experience weaker attraction to the nucleus and requires less energy to be removed.

(A similar explanation can be applied to Oxygen & Nitrogen in Period 2.)

4.2.3 Trend of I.E. down the Group

- Nuclear charge increases as the number of protons increases down the Group,
- With an increase in the number of inner shells, the valence electrons are further away from the nucleus and shielding effect increases.
- The valence electrons are **less strongly attracted** to the nucleus and **less energy is** required to remove the valence electron.
- Ionisation energies decreases down the group.

Exercise 5:						
Why is the first ionisation energy of sodium much lower than that of neon?						
Ne: 1s ² 2s ² 2p ⁶	Thinking process :					
Na: 1s ² 2s ² 2p ⁶ 3s ¹	 What is first I.E.? What are the factors that affect I.E? 					
	Use its electronic configuration to help					
Nuclear charge of Na is higher than Ne as it has more	protons.					
The electron to be removed from Ne is from the 2p orb	<u>ital</u> , while that of Na is from the <mark>3s</mark> orbital.					
With an increase in the number of inner shells, the elec	tron in the <u>3s orbital is <mark>further</mark> away from</u>					
the nucleus and shielding effect increases. The elect	ron in the 3s orbital is <u>less strongly</u>					
attracted to the nucleus and <u>requires less energy to be</u>	e be removed.					

Criteria of Success Checklist

When comparing ionisation energies / radii between particles

 $\hfill\square$ mentioned the factors that increase and decrease the attraction of valence electrons by the nucleus

Clearly outline which factor is more significant

mentioned that more/less energy is required to remove electron (for I.E.)

4.3 Atomic Radii

The atomic radius of an atom can be interpreted as the distance from the centre of the nucleus to the outermost shell of electrons, i.e.



 Al^{3+}

Al

4.3.1 Trend of atomic radius across the period

- The number of protons increases across the period, hence nuclear charge increases.
- Successive electrons are being added to the **same outermost shell** thus **increase in shielding effect is negligible** (since there is no change in the no. of inner shell electrons).
- Since the increase in nuclear charge outweighs the increase in shielding effect, effective nuclear charge increases. Hence, the valence electrons are **more strongly attracted** to the nucleus.
- Atomic radius **decreases** across the period.

4.3.2 Trend of atomic radius down the Group

- Nuclear charge increases as the number of protons increases down the Group,
- With an increase in the number of inner shells, the valence electrons are further away from the nucleus and shielding effect increases.
- The valence electrons are **less strongly attracted** to the nucleus.
- Atomic radius **increases** down the Group.

In w	Exercise 6:	irs is t	he radius d	of the second atom greater than that of the first atom?
				(J1994)
A	Na, Mg	в	Sr, Ca	
С	P, N	D	C <i>l</i> , Br	

4.4 Ionic Radius

The ionic radius is defined as the radius of a spherical ion in an ionic lattice.

4.4.1 Cationic Radius

 $Al : 1s^2 2s^2 2p^6 3s^2 3p^1$

Al³⁺: 1s² 2s² 2p⁶

- The radius of a **cation** is **smaller** than that of its parent atom.
- The number of protons in both the atom and its cation is the same, hence nuclear charge remains constant.
- The cation has one less shell of electrons than its atom.
- Hence, its valence electrons experience lower shielding effect and are more strongly attracted to the nucleus.

4.4.2 Anionic Radius

Particle	No. of protons	No. of electrons	0 o ²⁻
0 _{>} +2e ⁻	8 constant	8 jincreases	
0 ²⁻⁴	8	10	0.73 1.40

- The radius of an **anion** is **larger** than that of its parent atom.
- The number of protons in both the atom and its anion is the same, hence **nuclear charge remains constant**.
- The anion has more electrons than its atom.
- Hence, the **attraction** between the nucleus and its valence electrons **decreases**.

Group	1	2	13	14	15	16	17
Element	Na	Mg	Al	Si	Р	S	Cl
lon	Na⁺	Mg ²⁺	A <i>l</i> ³⁺	Si ⁴⁺	P ^{3–}	S ^{2–}	Cl⁻
Ionic radius / nm	0.095	0.065	0.050	0.041	0.212	0.184	0.181

4.4.3 Trend of Ionic Radius across the period

Ionic radius of cations: $Na^+ > Mg^{2+} > Al^{3+} > Si^{4+}$

Ionic radius of anions: $P^{3-} > S^{2-} > Cl^{-}$

- The nuclear charge increases due to increasing number of protons.
- The cations (or anions) are **isoelectronic**. Hence, **shielding effect** is the **same**.
- Effective nuclear charge increases and the valence electrons are more strongly attracted to the nucleus. Hence, ionic radius decreases.

lonic radius of cations < anions

- For example, ionic radius of $P^{3-} > Al^{3+}$ $Al^{3+}: 1s^2 2s^2 2p^6 \qquad P^{3-}: 1s^2 2s^2 2p^6 3s^2 3p^6$
- P^{3-} has 2 more proton than Al^{3+} and hence a greater nuclear charge.
- P³⁻ has **one more shell of electrons** than Al³⁺. The valence electrons of P³⁻ are **further** away from the nucleus. Hence, **shielding effect increases**.
- The increase in shielding effect outweighs the increase in nuclear charge. Hence, effective nuclear charge decreases.
- The valence electrons are less strongly attracted to the nucleus. Hence, ionic radius of P³⁻ is larger than that of Al³⁺.



Graph of Radius against Atomic Number for Lithium to Calcium

4.4.4 Trend of Ionic Radius down the Group

lonic radius of anions: $Li^+ < Na^+ < K^+$

- Nuclear charge increases from Li⁺ to K⁺ as the number of protons increases.
- From Li⁺ to K⁺, the number of inner shell electrons increases. Hence, the shielding effect increases, and the distance between the nucleus and the valence shell electrons increases.
- Thus, overall, ionic radius increases from Li⁺ to K⁺

Why does the ionic radius of $F^- < Cl^- < Br^- < I^-$?

4.5 Electronegativity

Electronegativity is the measure of the **tendency of an atom to attract a bond pair of electrons** (shared paired of electrons) **in a covalent bond**. The electronegativity of an atom is the **ability to attract bonding electrons toward itself**.

н 2.1	Pauling Electronegativity Values										He						
Li 1.0	Be 1.6											В 2.0	С 2.5	N 3.0	0 3.5	F 4.0	Ne
Na 0.9	Mg 1.3											Al 1.5	Si 1.9	Р 2.2	S 2.6	C1 3.0	Ar
К 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 2.8	Kr
Rb 0.8	Sr 0.9	Y 1.2	Zr 1.3	Nb 1.6	M o 2.2	Тс 1.9	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.1	Te 2.1	I 2.5	Xe

Link to Chemical Bonding

Charles E. Sundin, University of Wisconsin-Platteville

An atom that is *more electronegative* is likely to have a *higher electron density* when it is bonded in a molecule. For example, the higher electron density on F atom in the covalent molecule of hydrogen fluoride as F is more electronegative.



4.5.1 Trend of electronegativity across the period

- The number of protons increases across the period, hence nuclear charge increases.
- Successive electrons are being added to the same outermost shell thus increase in shielding effect is negligible.
- Since the increase in nuclear charge outweighs the increase in shielding effect, effective nuclear charge increases. Hence, the **ability of an atom to attract bonding electrons** to itself **increases**.
- Electronegativity increases across the period.

4.5.2 Trend of electronegativity down the group

- Nuclear charge increases as the number of protons increases down the Group,
- With an **increase in the number of inner shells**, the **valence electrons** are **further away** from the nucleus and **shielding effect increases**.
- The ability of an atom to attract bonding electrons to itself decreases.
- Electronegativity decreases down the Group.

☞ Self- Check: Q10

St Andrew's Junior College

5 Experimental Determination of Ionisation Energies

- (j) deduce the electronic configurations of elements from successive ionisation energy data
- (k) interpret successive ionisation energy data of an element in terms of the position of that element within the Periodic Table.



Copyright © MOE, Singapore

5.1 Successive ionisation energy

- The successive ionisation energies of an element increase.
- The number of electrons decreases but number of protons remain the same.
- Stronger attraction between the nucleus and remaining electrons.
- More energy is required to remove an electron from increasingly positively-charged ion.

5.2 Large increase in ionisation energy

- There is a large increase in I.E. from the first to second I.E.
- Na: 1s²2s²2p⁶3s¹ Na⁺: 1s²2s²2p⁶

The first electron is removed from an outer shell, while the second electron is from an **inner shell** which is much **closer** to the nucleus.

- Stronger attraction between the nucleus and the electron.
- Much more energy is required to remove this electron.
 How would you explain the large increase between the 9th I.E. and 10th I.E?

5.3 Determining the Group number from successive ionisation data

From the successive ionisation energies values for an element, one can determine the Group of the element.

Example:

A **large** increase from the 1st IE to the 2nd IE indicates that the element has 1 valence electron (valence electronic configuration is ns¹) as the second electron is removed from an inner shell. The second electron is more strongly attracted to the nucleus and requires a larger amount of energy to remove. Hence, the element is from Group 1.

The	Exercise 7:	n energies, in ł	⟨J mol⁻¹, of an e	element X are given below.
870	1800	3000	3600	5800 7000 13200
Wha	t group does X belor	ng to?		Thinking process :
Α	Group 14			 How to make sense of the
В	Group 15			successive I.E data and the group?
С	Group 16			
D	Group 17			
				∽ Self- Check: Q

SUMMARY of Periodic Trends



Across the Period,

- The number of protons increases across the period, hence nuclear charge increases.
- Successive electrons are being added to the same outermost shell thus increase in shielding effect is negligible.
- Since the increase in nuclear charge outweighs the increase in shielding effect, effective nuclear charge increases.
 - o The valence electrons are more strongly attracted to the nucleus.
 - Atomic radius **decreases** across the period.
 - The ability of an atom to attract bonding electrons to itself increases.
 - Electronegativity increases across the period.
 - More energy is required to remove the valence electrons.
 - Ionisation energies generally increase across the period.

Down the Group,

- Nuclear charge increases as the number of protons increases down the Group,
 - With an increase in the number of inner shells, the valence electrons are further away from the nucleus and shielding effect increases.
 - The valence electrons are **less strongly attracted** to the nucleus.
 - Atomic radius **increases** down the group.
 - The ability of an atom to attract bonding electrons to itself decreases.
 - Electronegativity decreases down the Group.
 - The valence electrons are less strongly attracted to the nucleus and less energy is required to remove the valence electron.
 - Ionisation energies decreases down the group.

Appendix A: Uses of Radioactive Isotopes

The advantage of a radioisotope is that its position can be detected very easily by the radiation which it emits. It has wide applications in various fields like medicine, agriculture, etc. In agriculture, radioactive phosphorus (³²P) is used in the study of metabolism of plants. Radioactive sulfur (³⁵S) helps to study advantages and disadvantages of fungicides. Carbon dating involves using the naturally occurring radioisotope carbon-14 (¹⁴C) and one of the most frequent uses is to estimate the age of organic remains from archaeological sites.

Appendix B: Recent discoveries

As you may have learnt recently, 30 December 2015 IUPAC (the International Unions of Pure and Applied Chemistry) announced the claims for the discovery of four elements (proton numbers 113, 115, 117 and 118) had been accepted. This is in spite of the fact that the Periodic Table had these empty spots and that even as early 1976 suggestions of such elements had been proposed and yet it took close to 40 years before this could be announced by IUPAC to the rest of the world.

Was this due to laziness of scientists in these boards or that of the personal reasons? This was not so. Science is an evidence-based enterprise that is concerned with the natural world and needed the consensus within a community of scientists through critical debate and peer review before the knowledge could be certified as reliable and durable until proven otherwise. Independent teams of scientists in both Russia and Japan were tasked assess the validity of the claim from the years of 2004 to 2012, before IUPAC made the official announcement.

The extended Periodic Table (beyond proton number 118) : <u>https://en.wikipedia.org/wiki/Extended_periodic_table</u>

The discovery of new elements : http://earthsky.org/earth/four-new-elements-added-to-periodic-table

Further development of particle theory (Standard Model Theory): (Quarks, Leptons, Gluons, Higgs Boson, etc.) https://www.youtube.com/watch?v=ehHoOYqAT_U







E Constanting	B P	Prior Knowledge in the Pure O-Level Chemistry
2.2	Ato	mic structure
	(a)	state the relative charges and approximate relative masses of a proton, a neutron and an electron
	(b)	describe, with the aid of diagrams, the structure of an atom as containing protons and neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels)
		(knowledge of s, p, d and f classification is not required; a copy of the Periodic Table will be available in Papers 1 and 2)
	(c)	define proton (atomic) number and nucleon (mass) number
	(d)	interpret and use symbols such as ${}^{12}_6$ C
	(e)	define the term <i>isotopes</i>
	(f)	deduce the numbers of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers.
8.1	Per	iodic trends
	(a)	describe the Periodic Table as an arrangement of the elements in the order of increasing proton (atomic) number
	(b)	describe how the position of an element in the Periodic Table is related to proton number and electronic structure
	(c)	describe the relationship between group number and the ionic charge of an ion of an element
	(d)	explain the similarities between the elements in the same group of the Periodic Table in terms of their electronic structure
	(e)	describe the change from metallic to non-metallic character from left to right across a period of the Period Table
	(f)	describe the relationship between group number, number of valency electrons and metallic/ non-metallic character
	(g)	predict the properties of elements in Group I and Group VII using the Periodic Table.
Note	e: As	sessment Objective 8.1(c) is not in the Combined Science O-Level syllabus. Please do the
nece	essar	y bridging required.

Answer to Exercises:

6	7
D	D

The Periodic Table of Elements

Group																	
1	2											13	14	15	16	17	18
Let the second s													α.	2 He ^{helium} 4.0			
3	4	atomic number				10 d		•				5	6	7	8	9	10
Li	Be		atomic symbol									B	С	N	0	F	Ne
lithium	beryllium	name									boron	carbon	nitrogen	oxygen	fluorine	neon	
0.9	9.0				1 10						10.0	12.0	14.0	16.0	19.0	20.2	
Na	Ma											Δ1	Si	P	S	CI	Δr
sodium	magnesium											aluminium	silicon	phosphorus	sulfur	chlorine	argon
23.0	24.3	3	4	5	6	7	8	9	10	11	12	27.0	28.1	31.0	32.1	35.5	39.9
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
potassium	calcium	scandium	titanium	vanadium	chromium	manganese	iron	cobalt	nickel	copper	zinc	gallium	germanium	arsenic	selenium	bromine	krypton
39.1	40.1	45.0	47.9	50.9	52.0	54.9	55.8	58.9	58.7	63.5	65.4	69.7	72.6	74.9	79.0	79.9	83.8
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	Ι	Xe
RE 5	strontium 87.6	yttrium 88.9	ZIRCONIUM		molybdenum 95 9	technetium	ruthenium	rhodium	palladium	silver	112 /		tin 118 7	121 8	127 6	126 9	xenon
55	56	57 74	72	72	74	- 75	76	77	79	70	80	81	82	92	94	85	96
Cc	Ba	J/-/I				Po	06	Ir	D+	7.5 Au	Ha	T7	Dh	Di	Do Po	Δ+	Dn
caesium	barium	ananoido	hafnium	tantalum	tungsten	rhenium	osmium	iridium	nlatinum	nold	mercury	thallium	I lead	bismuth	nolonium	astatine	radon
132.9	137.3		178.5	180.9	183.8	186.2	190.2	192.2	195.1	197.0	200.6	204.4	207.2	209.0		-	-
87	88	89-103	104	105	106	107	108	109	110	111	112	2	114		116		
Fr	Ra	actinoids	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn		Fl		Lv		
francium	radium		rutherfordium	dubnium	seaborgium	bohrium	hassium	meitnerium	darmstadtium	roentgenium	copernicium		flerovium		livermorium		
	-		-		-		-	-	<u> </u>	_	-		-		_		
lanthanoids		57	58	59	60	61	62	63	64	65	66	67	68	69	70	71	
		La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	
		lanthanum	cerium	praseodymium	neodymium	promethium	samarium	europium	gadolinium	terbium	dysprosium	holmium	erbium	thulium	ytterbium	lutetium	
		138.9	140.1	140.9	144.2	-	150.4	152.0	157.3	158.9	162.5	164.9	167.3	168.9	173.1	175.0	
actinoids		89	90	91	92	93	94	95	96	97	98	99	100	101	102	103	
		Ac	Ih	Ра	U	Np	Pu	Am	Cm	BK	Ct	Es	Fm	Md	No	Lr	
		actinium	232 0	231 0	238 0	neptunium	plutonium	americium	curium	berkelium	californium	einsteinium	fermium	mendelevium	nobelium	lawrencium	
		_	202.0	201.0	200.0	-			_	-						-	

12 5