Physical Periodicity of Elements

Content

Periodicity of physical properties of the elements: variation with proton number across the third period (sodium to chlorine) and down the group of:

(i) atomic radius and ionic radius
(ii) ionisation energy
(iii) electronegativity

Learning outcomes

Trends and variations in atomic and physical properties

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) describe and explain qualitatively the 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 filled electronic shells, shielding and nuclear charge

References

1. Hill, G. & Holman, J. (1992). *Chemistry in Context, 6th edition.* New York: Prentice Hall.
2. Ramsden, E. N. (1994). *A Level Chemistry, 3rd edition.* UK: Stanley Thornes (Publishers) Ltd.
3. Chemistry for Advanced Level, Peter Cann, Peter Hughes
4. A-level Chemistry, E.N. Ramsden

1 Introduction to Periodic Table

Success criteria
• I can recognize the variations in the valence electronic configuration across the Period and down a Group.

Mankind had always been seeking to discover new chemical elements since centuries ago. As of January 2021, a total of 118 elements have been identified. Studying these hundreds of elements with different physical and chemical properties will surely be a daunting task, unless there is a systematic and organized way to do it. Much of what we will cover in this topic is credited to the Russian chemist Dmitri Ivanovich Mendeleev (1834-1907), whose seminal hypothesis that the *chemical and physical properties of the elements vary in a periodic way* laid the foundations of the modern Periodic Table.

The Periodic Table is a table of elements arranged by order of their atomic numbers in such a way that the periodic properties (chemical periodicity) of the elements are made clear. In fact, the Periodic Table is a masterpiece of organised chemical information.

The standard form of Periodic Table includes **periods** (horizontal rows) and **groups** (vertical columns). The elements in the same period have the same number of core electrons as that of the noble gas from previous period. **Elements in the same group have the same number of valence electrons and hence similar chemical properties.**

Physical periodicity is the study of periodic trends which indicate the tendencies of certain elemental characteristics to change as one progresses along a row or column of the Periodic Table.

Checkpoint 1:
Write the electronic configurations of (i) Li and Na and (ii) F and Cl. Observe their relative positions on the Periodic Table. Underline the valence electronic configurations.

(i) Li : Na :

(ii) F : Cl :

The valence (outermost) shell electronic configuration of the elements can be used to determine the location of the element in the Periodic Table.

- The principal quantum number, "n", of the valence shell indicates the **period** of the element.
- The number of valence electrons indicates the **group** of the element. (Note that elements with 3-8 valence electrons are in Group 13 to 18)

Group	Valence shell electronic configuration
1	ns^1
2	ns^2
13	$ns^2 np^1$
14	$ns^2 np^2$
15	$ns^2 np^3$
16	$ns^2 np^4$
17	$ns^2 np^5$
18	$ns^2 np^6$

n = period in which the element is found in the Periodic Table

For example, an element with the electronic configuration of $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$ is located at Period 4 of the Periodic Table and is in Group 17. The element is Br.

Group																	
1	2											13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0		Key atomic number atomic symbol name relative atomic mass									5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3	3	4	5	6	7	8	9	10	11	12	13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –		114 Fl flerovium –		116 Lv livermorium –		

lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.2	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
actinoids	89 Ac actinium –	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –

2 Fundamental Considerations in Physical Periodicity

Success criteria
• I am able to discuss the effect of nuclear charge, shielding effect and number of filled electronic shells on the nuclear attraction.

Apart from the valence electrons, we will now look at other factors which determine the characteristic physical properties of an atom.

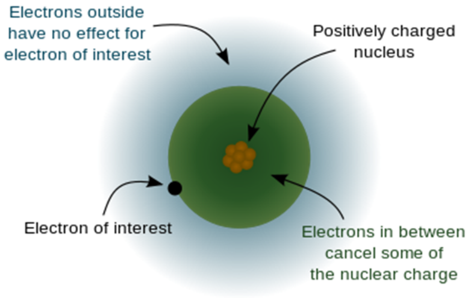
2.1.1 Nuclear charge

Nuclear charge is the total charge of all the protons in the nucleus. The larger the number of protons in the nucleus of an atom, the greater the nuclear charge. Hence, across a period or down a group, nuclear charge increases.

2.1.2 Shielding effect

Note: Electrons in the same electronic shell exert negligible shielding effect on each other.

Valence electrons in an atom do not only 'feel' attraction from the protons in the nucleus, but also repulsion from inner shells electrons. The 'feel' of repulsion from inner-shell electrons is called **shielding effect.** The inner-shell electrons shield the outer electron from the full charge of the nucleus.



Hence, shielding effect increases as the number of inner shell electrons increases.

2.1.3 Number of filled electronic shells (distance of electrons from nucleus)

Increase in number of filled electronic shell results in:
1) Increase in distance between the valence electrons and the nucleus
2) Increase in shielding effect

Hence attractive force between the valence electrons and the nucleus is weaker as the valence electrons is further away from the nucleus.

2.2 Nuclear attraction ("net attraction electrons experienced from the nucleus")

Nuclear attraction refers to "attraction between the nucleus and the electrons (nuclear charge) – "repulsion between electrons (shielding effect) + distance of electrons from nucleus (no. of filled electronic shells)"

Nuclear attraction = "nuclear charge – (shielding effect + distance of electrons from nucleus" *calculations are not required

3 Periodicity Trend

Success criteria

- I can describe and explain qualitatively the trends and variations in atomic radius, ionic radius, first ionisation energy and electronegativity:
  - (i) across a Period (**<u>same</u>** valence electronic shell) in terms of shielding effect and nuclear charge;
  - (ii) down a Group (**<u>increasing number of filled electronic shell</u>**) in terms of distance away from nucleus, shielding effect and nuclear charge.
- I can predict and explain the variations in ionic radius for species with either same number of electrons (e.g. isoelectronic species, Na^+, Mg^{2+}) or same number of protons (e.g. Na vs Na^+) using proton to electron ratio to account the nuclear attraction for the outermost electrons.

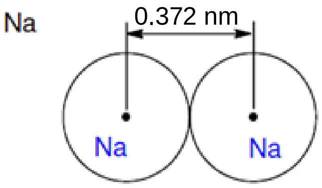
3.1 Atomic radius (atomic size)

The **atomic radius** of an element is a measure of the size of its atoms, usually the mean distance from the nucleus to the boundary of space occupied by the valence electrons. Since the boundary is not a well-defined physical entity, internuclear distance is being measured instead. Depending on the type of bonding and structure the atoms are involved, there are three commonly known atomic radius, namely metallic radius, covalent radius and van der Waals radius.

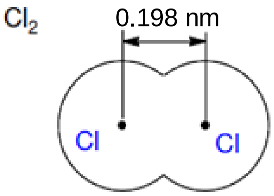
The ***metallic radius*** is defined as half the internuclei distance between the two adjacent metal ions in the metallic lattice.

The ***covalent radius*** is defined as half the internuclei distance between two atoms which are covalently bonded. In principle, the sum of the two covalent radii should equal the covalent bond length between two atoms.

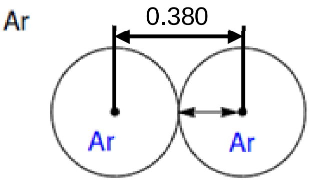
The ***van der Waals radius*** is defined as half the internuclei distance between two atoms that do not overlap to form covalent bond. E.g. Noble gas such as He, Ne and Ar do not form covalent bond.



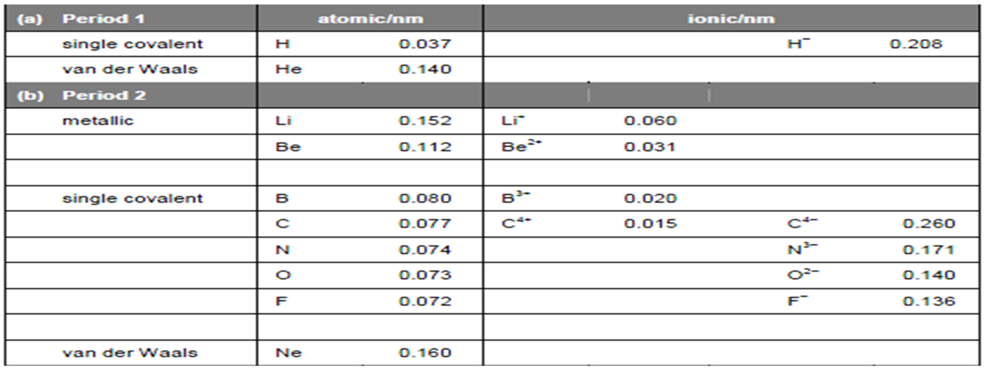
Metallic radius of Na
= 0.186 nm



Covalent radius of C*l*
= 0.099 nm



van der Waals radius of Ar
= 0.190 nm

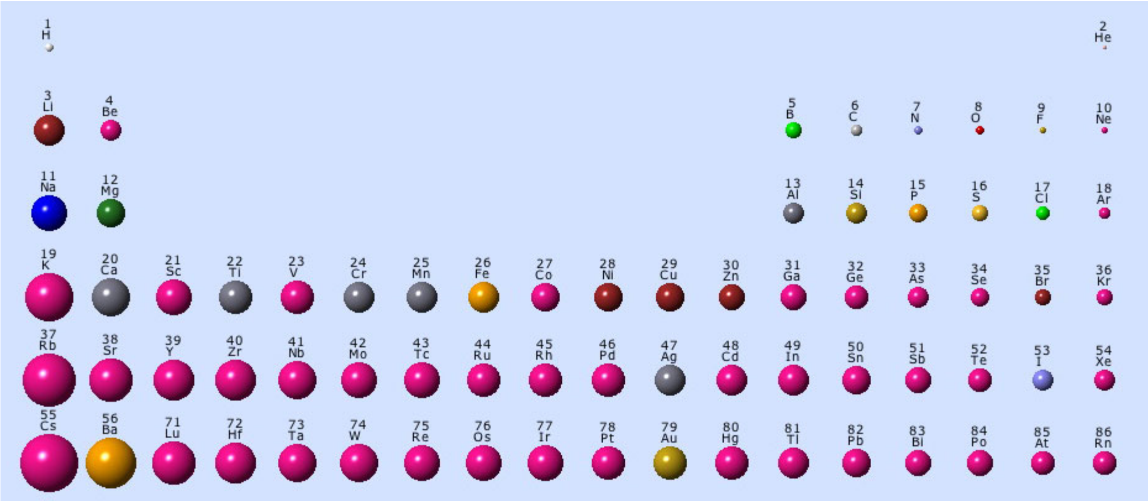


(a) Period 1	atomic/nm		ionic/nm			
single covalent	H	0.037			H^-	0.208
van der Waals	He	0.140				
(b) Period 2						
metallic	Li	0.152	Li^+	0.060		
	Be	0.112	Be^{2+}	0.031		
single covalent	B	0.080	B^{3+}	0.020		
	C	0.077	C^{4+}	0.015	C^{4-}	0.260
	N	0.074			N^{3-}	0.171
	O	0.073			O^{2-}	0.140
	F	0.072			F^-	0.136
van der Waals	Ne	0.160				

Atomic and ionic radii from Data Booklet

In this section we refer atomic radius to the size of its atoms in general.

Atomic Size Trends



3.1.1 Across a period, atomic radius decreases

Note:
In the data booklet, the radius of Ar is given as 0.190nm whereas that of C*l* is 0.099nm, which contradicts the trends shown in the above figure. Can you explain?

Refer to pg 4

Explanations:

- **nuclear charge increases** (due to increasing number of protons in the nucleus)
- **number of inner shell electrons is the same** (as additional electrons are added to the same outermost shell) and **shielding effect** remains relatively constant
- **nuclear attraction for the valence electrons increases**.
- atomic radius **decreases**.

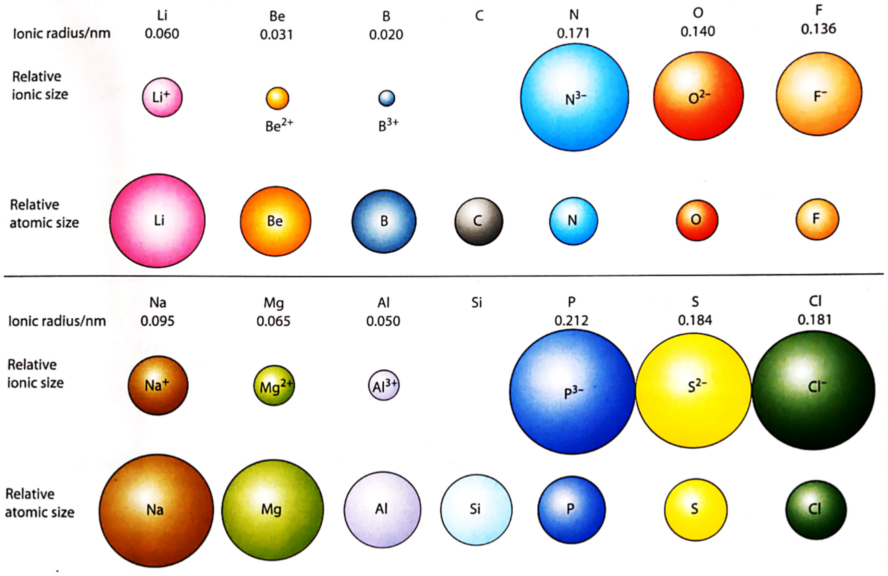
3.1.2 Down a group, atomic radius increases

Explanations:

- **nuclear charge increases** (due to increasing number of protons in the nucleus)
- **number of filled electronic shells increases**.
- the valence electrons are **further away** from the nucleus and experience **greater shielding effect** which outweighs the **increasing nuclear charge**
- **nuclear attraction for the valence electrons decreases**
- atomic radius **increases**.

3.2 Ionic radius (ionic size)

Ionic radius is the radius of an ion (cation or anion).

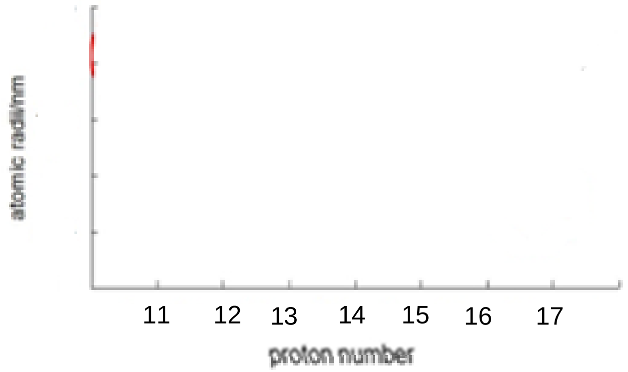


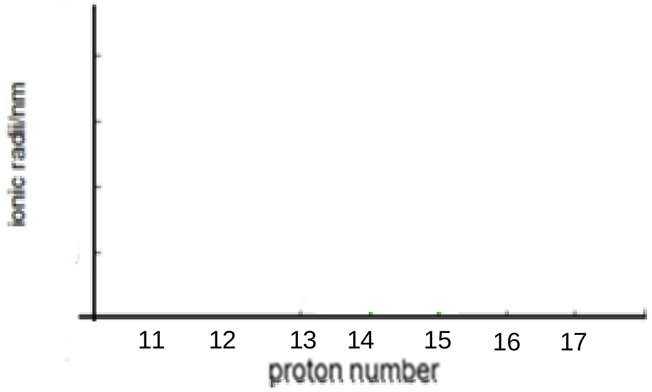
Observations	**Explanations**
Cations are <u>smaller</u> than its corresponding atom (trend A) Eg. Na vs Na^+, Mg vs Mg^{2+}	Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ vs Na^+: $1s^2\ 2s^2\ 2p^6$ • **Nuclear charge remains the same** for the cations and the atoms. • **Nuclear attraction for the valence electrons** in cations is **stronger** as the same number of protons are attracting fewer number of electrons. • cations have **smaller** ionic radius.
Anions are <u>bigger</u> than its corresponding atom (trend B) Eg. C*l* vs C*l*⁻, S vs S^{2-}	C*l*: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ vs C*l*⁻: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ • **Nuclear charge remains the same** for the anions and the atoms. • **Nuclear attraction for the valence electrons** in anions is **weaker** as the same number of protons are attracting more number of electrons. • Anions have **bigger** ionic radius.

Ionic radius <u>decreases</u> across isoelectronic series. Eg Na^+ vs Mg^{2+}, S^{2-} vs Cl^-	Na^+: $1s^2\ 2s^2\ 2p^6$ vs Mg^{2+}: $1s^2\ 2s^2\ 2p^6$ • **nuclear charge** across isoelectronic series **increases** while **number of electrons remains the same**. • **nuclear attraction for the valence electrons** increases, as more protons are attracting the same number of electrons. • Ionic radius **decreases**.
Anions are <u>bigger</u> than the cations in the same period. Eg. P^{3-} vs Na^+	P^{3-}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ vs Na^+: $1s^2\ 2s^2\ 2p^6$ • Anions has **higher nuclear charge and one additional filled electronic shell** than cations in the same period. • **Nuclear attraction for the valence electrons** in anions is **weaker** as the electrons are **further away** from nucleus and have greater **shielding effect** which outweighs the **higher nuclear charge** in **anions**. • Anions have **bigger** ionic radius.

Checkpoint 2:

1 Sketch the trends in (i) atomic radius and (ii) ionic radius for elements with atomic number 11 to 17.





2 In which of the following pairs is the radius of the second atom greater than that of the first atom?

A Na, Mg **B** Sr, Ca **C** P, N **D** C*l*, Br

3 The following species contain the same number of electrons. In which order do their radii increase?

	smallest radius ----------------------> largest radius		
A	Ar	K^+	Ca^{2+}
B	Ca^{2+}	Ar	K^+
C	Ca^{2+}	K^+	Ar
D	K^+	Ar	Ca^{2+}

4 The following represents the electronic configuration of both a Group 2 cation and Group 17 anion.

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$$

The radius of the anion is about twice that of the cation.

Which reasons explain the difference in their sizes?

1 The cation has a greater nuclear charge than the anion.
2 Anion has a greater shielding effect than the cation.
3 On forming the anion from its atom, the extra electron repulsion makes the ion much bigger.

A 1, 2 and 3 **B** 1 and 2 only **C** 2 and 3 only **D** 1 only

Answer to Checkpoint 2(Q2-4): DCD

4 Ionisation Energy (IE)

Success criteria

- I can sketch, describe and explain the first ionisation energy trend across ~~a~~ Period 2 or 3 and down a group by referring to their valence electronic configuration.
- I can explain the anomalous trend in first ionisation energy for elements with valence electronic configuration of
  - (i) **ns^2 vs ns^2np^1** in terms of most loosely held electron in a higher energy p subshell
  - (ii) **ns^2np^3 vs ns^2np^4** in terms of interelectronic repulsion between paired electrons in p orbital

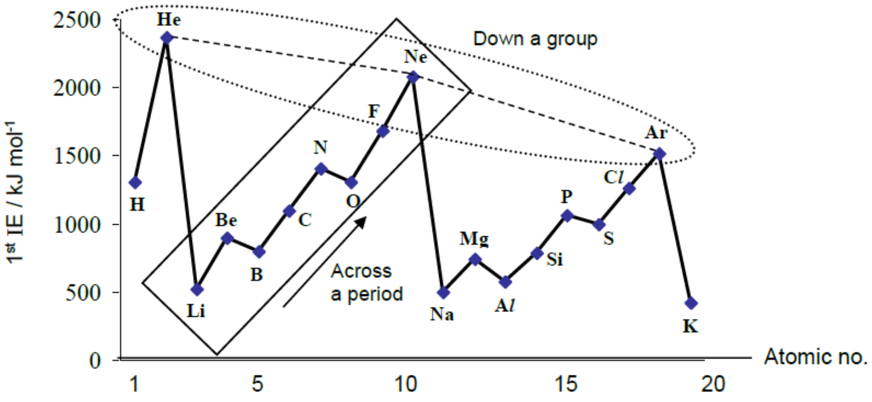
Note the difference of the term “valence electrons” and “most loosely held electrons”:

- Valence electrons are found in the outermost (valence) shell.
- The most loosely held electron is the valence electron that requires the least amount of energy to be removed.

The first ionisation energy of an element is the **energy required** to remove **1 mole of most loosely held electrons** from **1 mole of gaseous atoms of the element** to form **1 mole** of **singly-charged gaseous cations**.

Eqn: $X(g) \rightarrow X^+(g) + e^-$ ΔH = positive

4.1 Trend in first ionisation energies of different elements



It is observed that the first ionisation energy for elements across a period increases generally while first ionisation energy decreases down a group.

Down a group (eg. He → Ne → Ar)
First ionisation energy of elements **decreases**.
Explanations:

- **nuclear charge** increases,
- **number of filled electronic shells** increases,
- the most loosely held electron is **further away** from the nucleus and experiences **greater shielding effect** which **outweighs** the **increasing nuclear charge**
- **nuclear attraction** for the most loosely held electron **decreases**
- **less energy** is required to remove the electron and 1st IE **decreases**.

Across a period (eg. Li to Ne)
First ionisation energy of elements shows a general **increase**.
Explanations:

- **nuclear charge** increases,
- number of inner shell electrons remains the same and **shielding effect** remains constant.
- **nuclear attraction** for the most loosely held electron is **stronger**
- **more energy** is required to remove it and 1st IE **increases.**

4.2 Anomalies in first ionisation energies trend

There are **two anomalies** in the first IE of elements across a period. Let's look at the electronic configurations of the elements showing the anomalies.

(i) Small dip between Group 2 and Group 13 elements

We will consider the first IE of Period 3 elements. Notice that A*l* has a lower first IE than Mg.

Let's consider the electronic configurations of Mg and A*l*.

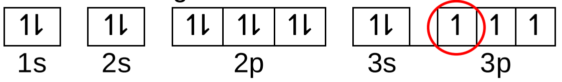
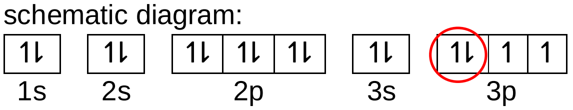
Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$ A*l*: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

- The most loosely held electron of A*l* is in the **higher energy 3p subshell** while that of Mg is in the 3s subshell.
- This **outweighs** the higher **nuclear charge** in A*l*.
- **Nuclear attraction** for the most loosely held electron of A*l* is **weaker**.
- **Less energy** is required to remove this electron, resulting in lower 1st IE.

(ii) Small dip between Group 15 and Group 16 elements

In S has lower first IE than P.

Let's consider the electronic configurations of P and S.

P: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	schematic diagram: 
S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	schematic diagram: 

- The most loosely held electron of S is in a **doubly filled 3p orbital** while that of P is in a singly filled 3p subshell.
- The most loosely held electron in S experiences **inter-electronic repulsion with its paired electron** which outweighs the **higher nuclear charge** in S.
- **Nuclear attraction** for this electron in S is **weaker**.
- **Less energy** is required to remove this electron in S, resulting in lower 1st IE.

Note:
How about between group 1 and group 2? And from group 16 to group 17 to group 18? Aren't there also increase in interelectronic repulsion?
Why aren't there a dip in 1st IE?

4.3 Sharp drop in first IE between Group 18 element and Group 1 element of the next period

Considering Ar ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$) and K ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$).

- The most loosely held electron of K is removed from electronic shell **n = 4** while that of Ar is from n =3.
- The most loosely held electron is **further away** from the nucleus and experience greater **shielding effect**.
- This outweighs the **higher nuclear charge** in K and the **nuclear attraction** for this electron is **weaker**.
- **Less energy** is required to remove this electron and 1st IE of K is much lower.

Checkpoint 3:

Compare the difference in 1st ionisation energy between the following pairs of elements:

(i) Be and B

Electronic Configuration of Be: ____________ B: ____________________

The most loosely held electron of B is in a ____________________ subshell while that of Be is in the ________ subshell. This outweighs the ______________________________ in B. Nuclear attraction for the ________________________ *electron* in B is ______________. ____________ energy is required to remove this electron, hence 1st IE of B is lower

(ii) S and C*l*

Electronic Configuration of S: _______________________ C*l*:_____________________

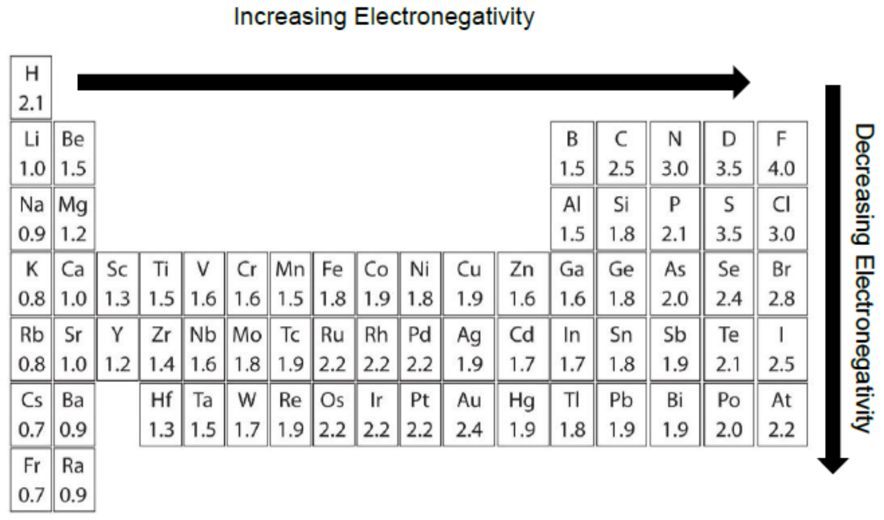
Both S and C*l* has the same number of ______________________ electrons and hence _______________________________. C*l*, with ________________ nuclear charge, has _____________ nuclear attraction for the *most loosely held electron*. _______________ energy is required to remove this electron, hence 1st IE of C*l* is higher.

(iii) Ne and Na

(iv) N and O

5 Trends in electronegativity in the periodic table

Electronegativity of an atom is a measure of its ability to **attract** the **shared pair of electrons** in a **covalent bond** towards itself.



H 2.1																
Li 1.0	Be 1.5											B 1.5	C 2.5	N 3.0	D 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 3.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9		Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9															

Pauling Electronegativity Values

The higher the electronegativity value, the more electronegative the atom is.

Checkpoint 4:

Using the same fundamental considerations in Section 2, compare the difference in electronegativity between the following pairs of elements:

(i) C and N

Both C and N have the ______________________________ electrons and hence ______________________________. N has _______________ nuclear charge and hence nuclear attraction for the **<u>bonding electrons</u>** in N is _______________. Thus, N has higher electronegativity.

(ii) F and C*l*

C*l* has more ________________________ than F. The distance of __________________________ is **further away** from nucleus and experience ________________________________. This outweighs the __________________________________ in C*l*. Nuclear attraction for the electrons in C*l* is ______________ and C*l* has lower electronegativity.

Note:
The concept of electronegativity is **useful in predicting if a bond formed is covalent or ionic** in the next topic of chemical bonding.

- **Across a period**, electronegativity **<u>increases</u>**.
- **Down a group**, electronegativity **<u>decreases</u>**.

Success Criteria	Relevant Tutorial questions	What do you still struggle with? Write your queries here.
1. I can recognize the variations in the valence electronic configuration across the Period and down a Group.	DQ5	
2. I can apply an understanding of nuclear charge, shielding effect and number of filled electronic shells from nucleus to account for the nuclear attraction.		
3. (a) I can describe and explain qualitatively the trends and variations in (I)atomic radius, (II)ionic radius, (III)first ionisation energy and (IV)electronegativity: (i) across a Period (**same** valence electronic shell) in terms of shielding effect and nuclear charge; (ii) down a Group (**increasing number of filled electronic shell**) in terms of shielding, distance away from nucleus and nuclear charge. (b) I can predict and explain the variations in ionic radius for species with either same number of electrons (e.g. isoelectronic species, Na^+, Mg^{2+}) or same number of protons (e.g. Na vs Na^+) using proton to electron ratio to account the nuclear attraction for the outermost electrons.	**(a)(I)**:SAQ1, DQ3b **(II)**: DQ2b, DQ3c, DQ4b(i)/(iii) **(III)**: DQ5 **(IV)**: DQ7d **(b)**: DQ2a, DQ3a, DQ6b	
4. (a) I can sketch, describe and explain the first ionisation energy trend across ~~a~~ Period 2 or 3 and down a group by referring to their valence electronic configuration. (b) I can explain the anomalous trend in first ionisation energy for elements with valence electronic configuration of **(i)** **ns^2 vs ns^2np^1** in terms of most loosely held electron in a higher energy p subshell **(ii)** **ns^2np^3 vs ns^2np^4** in terms of interelectronic repulsion between paired electrons in p orbital	**(a)**: DQ4a, DQ5 **(b)(i)**: DQ4b(ii), **(b)(ii)**: DQ4b(iv)	

General answering approaches to the questions

Step 1: Write down the electronic configuration of the respective species

Step 2: Compare the no. of protons (p) and electrons (e) between the species:

- If either e or p has the same number, use p to e ratio to compare the nuclear attraction on the outermost electron. (skip Step 3 and **continue** with **Step 4**)
- If both e and p changes at the same time, **comment on the nuclear charge** (**continue** with **Step 3**)

Step 3: Compare the number of filled electronic shell to **comment on the shielding effect and the distance from the nucleus**

Step 4: **State the strength of the nuclear attraction on**

(i) outermost electrons (atomic radius / ionic radius)
(ii) most loosely held electron (ionisation energy)
(iii) bonding electrons (electronegativity)

Step 5: Relate to its physical properties

Note:
For ionisation energy, relate the strength of the nuclear attraction to the energy needed to remove the most loosely held electrons.

Tutorial - Physical Periodicity of Elements

Self Attempt Question

1 State and explain which species in the following pairs has a larger radius.

(a) Li and Na **(b)** S and Cl

Discussion Questions

2 State and explain which species in the following pairs has a larger radius.

(a) Mg^{2+} and Al^{3+} **(b)** Li^{+} and Ne

3

	Atomic radius / nm	Ionic radius / nm
Magnesium	0.160	0.065 (Mg^{2+})
Sulfur	0.104	0.184 (S^{2-})

(a) Explain why the atomic radii of magnesium and sulfur are different from the ionic radii of their respective ions.

(b) Explain why the atomic radius of magnesium and sulfur are different.

(c) Explain why the ionic radius of Mg^{2+} and S^{2-} are different.

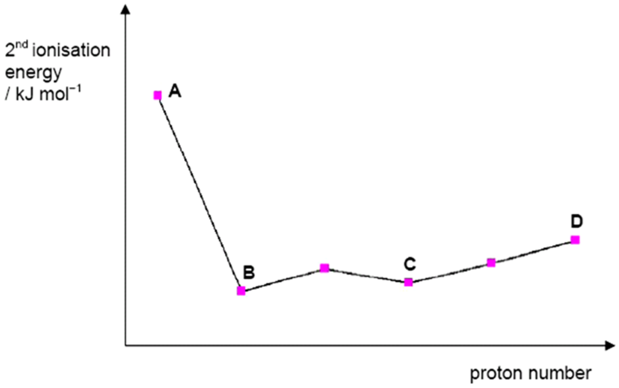
4 **(a)** Sketch a graph of the first IEs of the elements sodium to potassium against proton number.

(b) Explain the difference in the first IEs of the following pairs of elements.

- **(i)** Na and K
- **(ii)** Mg and Al
- **(iii)** Si and P
- **(iv)** P and S

5 **(a)** Write down the equation which represents the 2nd ionisation energy of element M.

(b) The variation in the second ionisation energy of six consecutive elements in the Periodic Table is shown in the graph.



Which of these elements is in Group 2? Explain your answer.

6 (2022 P2 Q1) Table 1.1 lists the number of protons, neutrons and electrons in seven different particles. Each particle may be an atom, an anion or a cation.

Particle	number of protons	number of neutrons	number of electrons
A	18	22	18
B	18	22	17
C	20	20	20
D	20	20	18
E	17	20	18
F	17	18	17
G	17	20	17

(a) Use the information in Table 1.1 to identify

- the two particles which are a pair of isotopes of the same element
- the cation with the same electronic configuration as an atom of argon
- the pair of atoms of different elements with the same nucleon number

(b) In **(i)** and **(ii)** deduce which particle is larger. Explain each answer.

(i) **C** and **D**

(ii) **D** and **E**

7 (2022 P2 Q2) Fig. 2.1 shows successive ionization energies of the Period 4 element selenium, Se.

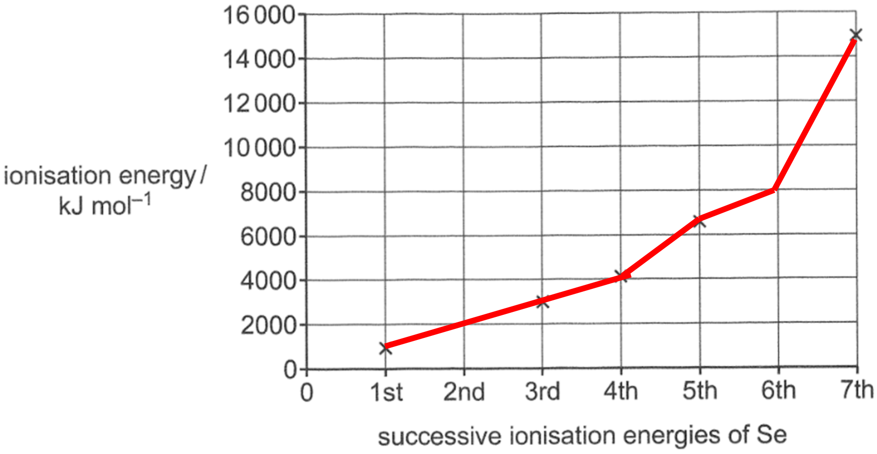


Fig. 2.1

(a) Explain the general increase in successive ionization energies for any atom.

(b) State the number of electron pairs in an atom of Se.

(c) Complete Fig 2.1 by plotting approximate values for the 2nd successive ionization energy and the 6th successive ionization energy of Se.

(d) All elements are assigned a value of electronegativity on the Pauling scale. Francium has the lowest electronegativity value of 0.7 and fluorine has the highest electronegativity value of 4.0.

Suggest the electronegativity value for Se by comparison with that given for oxygen. Explain your answer.