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 (**<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 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 $ns^2 np^1$** in terms of most loosely held electron in a higher energy p subshell **(ii)** **$ns^2 np^3$ vs $ns^2 np^4$** in terms of interelectronic repulsion between paired electrons in p orbital	**(a)**: DQ4a, DQ5 **(b)(i)**: DQ4b(ii), **(b)(ii)**: DQ4b(iv)	

<u>General answering approaches to the questions</u>

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 **<u>continue</u>** with **Step 4**)
- If both e and p changes at the same time, **comment on the nuclear charge** (**<u>continue</u>** 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 C*l*

(a) Li: $1s^2\ 2s^1$ Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$

- Na has one more filled electronic shell than Li. The distance of its valence electron is further away from nucleus and experience higher shielding effect. This outweighs the higher nuclear charge in Na. Nuclear attraction for the outermost *electron* in Na is weaker. Thus, atomic radius of Na is larger.

(b) S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$ C*l*: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

- S has the same number of inner shell electrons as C*l* and hence same shielding effect. However, S has lower nuclear charge and hence nuclear attraction for the outermost *electron* in S is weaker. Thus, atomic radius of S is larger.

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

Mg^{2+}: $1s^2\ 2s^2\ 2p^6$ Al^{3+}: $1s^2\ 2s^2\ 2p^6$

- Mg^{2+} is isoelectronic with Al^{3+}. However, Mg^{2+} has less protons attracting the same number of electrons and hence nuclear attraction for the outermost electrons in Mg^{2+} is weaker. Thus, ionic radius for Mg^{2+} is larger.

Li^+: $1s^2$ Ne: $1s^2\ 2s^2\ 2p^6$

- Ne has one more filled electronic shell than Li^+. The distance of its outermost electron is further away from nucleus and experience higher shielding effect. This outweighs the higher nuclear charge in Ne. Nuclear attraction for the outermost *electron* in Ne is weaker. Thus, atomic radius of Ne is larger.

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.

- Nuclear charge remains the same for Mg^{2+} and Mg.
- Nuclear attraction for the outermost electrons in Mg^{2+} is stronger as the same number of protons are attracting fewer electrons.
- Mg^{2+} have smaller ionic radius.

- Nuclear charge remains the same for S^{2-} and S.
- Nuclear attraction for the outermost *electrons* in S^{2-} is weaker as the same number of protons attract more electrons.
- Thus, S^{2-} is larger than S.

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

S: $1s^22s^22p^63s^23p^4$ Mg: $1s^22s^22p^63s^2$

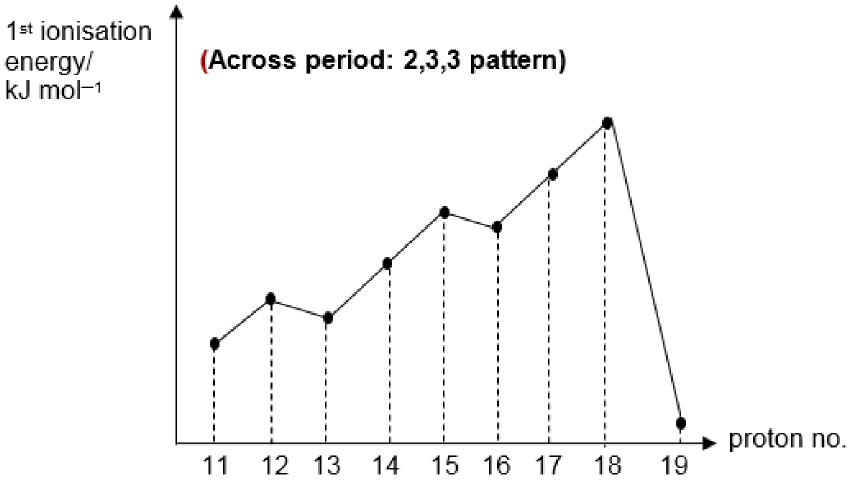
- S has the same number of inner shell electrons as Mg and hence same shielding effect. However, S has higher nuclear charge and hence nuclear attraction for the valence electrons in S is stronger. Thus, S has a smaller atomic radius than Mg.

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

Mg^{2+} : $1s^2\ 2s^2\ 2p^6$ vs S^{2-}: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

- S^{2-} has one more filled electronic shell than Mg^{2+}.The distance of its outermost electron is further away from nucleus and experiences higher shielding effect. This outweighs the higher nuclear charge in S^{2-} . Nuclear attraction for the *valence electron* in S^{2-} is weaker. Thus, ionic radius of S^{2-} is larger than Mg^{2+}.

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.

Steps to answering questions relating to ionisation energy trend:

1. Write down the **electronic configuration** of the respective species
2. Identify the **position of the most loosely held electron** of the respective species.
3. If the most loosely held electron of the respective species are removed from

(i) **Same** electronic shell, **n** (**Use nuclear charge & shielding effect**)

(ii) **Different** electronic shell, **n** vs **n+1** (**Use nuclear charge, no of filled electronic shells & shielding effect**)

Exception

(i) Same principal quantum shell with configuration $\mathbf{ns^2}$ vs $\mathbf{ns^2\,np^1}$
(Most loosely held electron in p subshell has higher energy than that of s subshell, thus requires less energy to remove → Lower 1st IE)

(ii) Same principal quantum shell with $\mathbf{ns^2\,np^3}$ vs $\mathbf{ns^2\,np^4}$
(Most loosely held electron in a doubly filled p orbital experiences interelectronic repulsion, thus requires less energy to remove → Lower 1st IE.)

(i) Na and K

Na: $1s^2\,2s^2\,2p^6\,\mathbf{3s^1}$ K: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,\mathbf{4s^1}$

- K has **1 more filled principal quantum shell** than Na.
- The **most loosely held** electron in K is **further away from the nucleus** and experiences **greater shielding effect**.
- These factors **outweigh the higher nuclear charge in K.**
- The nuclear attraction for the most loosely held electron is weaker in K, and less energy is required to remove it, thus 1st I.E **is lower for K.**

(ii) Mg and Al

Mg: $1s^2\ 2s^2\ 2p^6\ \mathbf{3s^2}$ Al: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^1}$

- The most loosely held electron of Al is in the **higher energy 3p subshell** while that of Mg is in the **3s subshell**.
- This outweighs the **<u>higher nuclear charge</u>** in Al.
- Nuclear attraction for the *most loosely held electron* in Al is **weaker** and <u>less</u> energy is required to remove it, hence 1st IE is lower for Al.

(iii) Si and P

Si: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^2}$ P: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^3}$

- The most loosely held electron is removed from the **<u>same 3p</u>** subshell for both Si and P.
- **<u>Nuclear charge</u>** for P **<u>larger</u>** than Si.
- **<u>Shielding effect</u>** for both P and Si are **<u>relatively constant</u>**,
- **<u>Nuclear attraction</u>** for the most loosely held electron in P is **<u>larger</u>** than Si.
- **<u>More</u>** energy is required to remove the most loosely held electron in P and 1st IE higher.

(iv) P and S

P: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^3}$ 3p $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

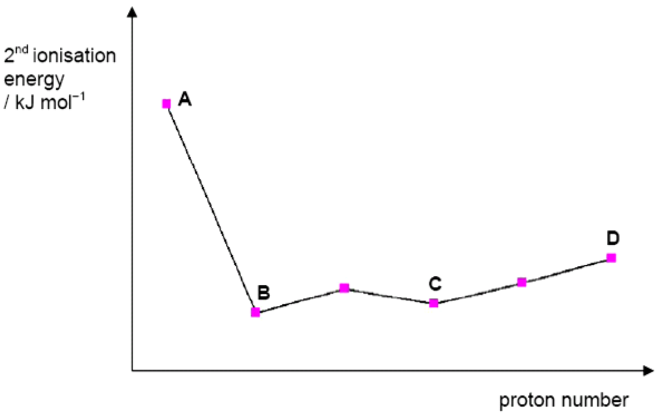
S: $1s^2\ 2s^2\ 2p^6\ 3s^2\ \mathbf{3p^4}$ 3p $\boxed{\uparrow\downarrow}\boxed{\uparrow}\boxed{\uparrow}$

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

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

$M^+(g) \rightarrow M^{2+}(g) + e$

(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.

2nd I.E. is $M^+(g) \rightarrow M^{2+}(g) + e^-$

Large dip in 2nd IE from **A^+** to **B^+** implies that

- the **removal of 2nd electron for B is from an outer electronic shell which is further away from the nucleus** and experiences weaker nuclear attraction hence needs much less energy to remove **OR**
- **the removal of 2nd electron for A is from an inner electronic shell** and experience more nuclear attraction hence needs much more energy to remove

Valence shell electronic configuration of A^+ : ns^2np^6
Valence shell electronic configuration of B^+ : $ns^2np^6\ (n+1)s^1$

Hence, B's configuration is $ns^2np^6\ (n+1)s^2$. It is a group 2 element.

Answer: **B**

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
  F & G (isotopes of the same element must have same proton number, but different number of neutron)
  E is excluded as it carries a negative charge.
- the cation with the same electronic configuration as an atom of argon
  D (cation has more proton than electron and must have same no. of electron as Argon = 18)

- the pair of atoms of different elements with the same nucleon number
  A & C (atom of different elements have different number of proton; and have same no. of proton as electrons)

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

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

Both C and D have same no. of proton, but different no. of electron. In particle C, the proton is attracting more electrons than that of particle D. Nuclear attraction on the outermost electron is weaker and hence the size of particle C is larger.

(ii) D and **E**

Both D and E are isoelectronic. Particle E has less proton than D attracting the same no. of electrons. Nuclear attraction on the outermost electron is weaker and hence the size of particle E is larger.

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

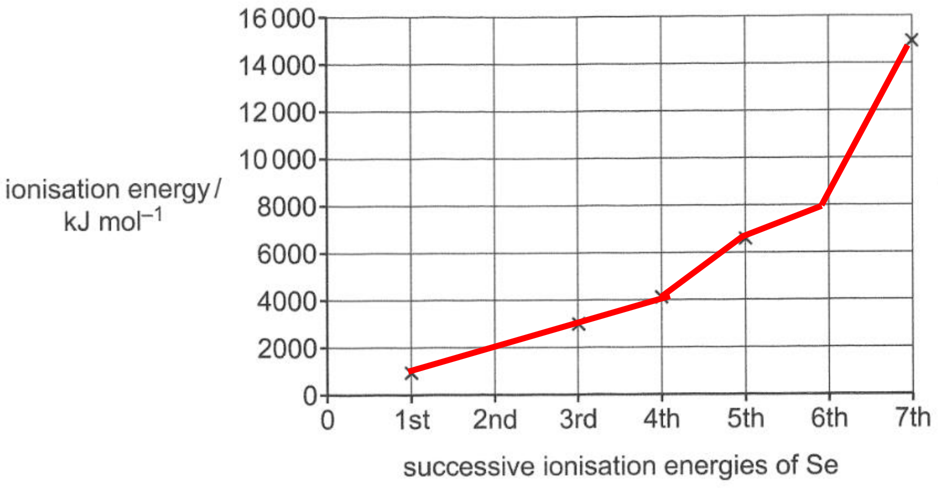


Fig. 2.1

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

- **With each successive removal of electron from the atom, the protons are attracting fewer no. of electron.**
- **nuclear attraction on the remaining electron increases.**
- **Increasing energy is required to remove the subsequent most loosely held electron, thus giving rise to the general increase in successive ionization for any atom.**

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

In an atom of Se, there are 34 proton and 34 electrons.
The number of electron pairs = 34/2 – 1 = 16
(There are two unpaired electrons in $4p^4$)

(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.

The electronegativity value for Se could be any value in-between 2 to 3.5. Se has more filled electronic shell than O. The distance of <u>shared pair of electrons</u> in the covalent bond is <u>further away</u> from nucleus in Se and experience <u>greater shielding effect</u>. This outweighs the higher <u>nuclear charge</u> in Se. Nuclear attraction for the electrons in Se is <u>weaker</u> and Se has lower electronegativity.

Electronegativity trend:

- **Increases across the period**
- **Decreases down a group**