**Tutorial 2 – Atomic Structure (Answers to Discussion Questions)**

**5. (a)** If $H^+$ where $\frac{q}{m}$ = 1 gives θ of +15°, since θ = k $(\frac{q}{m})$, k = 15
$D^-$: $\frac{q}{m}$ = ½ , hence θ = 15 x ½ = **–7.5°**
$T^+$: $\frac{q}{m}$ = ⅓, hence, θ = 15 x ⅓ = **+5.0°**
$He^{2+}$: $\frac{q}{m}$ = 2/4 = ½ , hence θ = 15 x ½ = **+7.5°**

**(b)(i)** From (a), θ = k $(\frac{q}{m})$ where k = 15
$5 = 15(\frac{q}{12}) \Rightarrow$ e = **4+**

**(ii)** Mass number = 12, hence, no. of neutrons = 12 – 6 = **6**.
Since charge of **R** is 4+, no. of electrons = 6 – 4 = **2**

**6. (a) (i) (A)** 1s **(B)** 2s **(C)** $2p_y$ **(D)** $2p_z$ **(E)** $2p_x$

**(ii)** Orbital (B) is bigger and more diffuse than orbital (A).
Orbital (B) is at a higher energy level than orbital (A).

**(iii)** Orbital (B) has a spherical shape and is non–directional**.**
Orbital (C) has a dumbbell shape and is directional as the electron density is concentrated along the *y* axis.

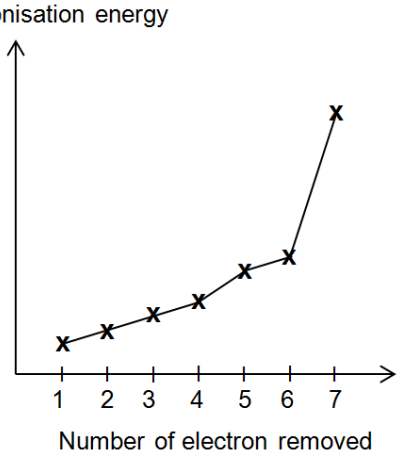
**(iv)** Orbitals (C), (D) and (E)

**(b) (i)** According to Hund's Rule, orbitals of the same energy must be occupied singly before pairing can occur. This is to minimise interelectronic repulsion to achieve greater stability.
Since (C) and (E) have the same energy and (C) is already doubly-filled, (E) cannot be empty.

**(b) (ii)** 2 electrons in (E)

**(c) (i)** Z = 2, 4 and 10 **(c) (ii)** Z = 7

**(d)**



**7.** **(a)** **(i)** $_{33}As^{3-}$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

**(ii)** $_{31}Ga^{3+}$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

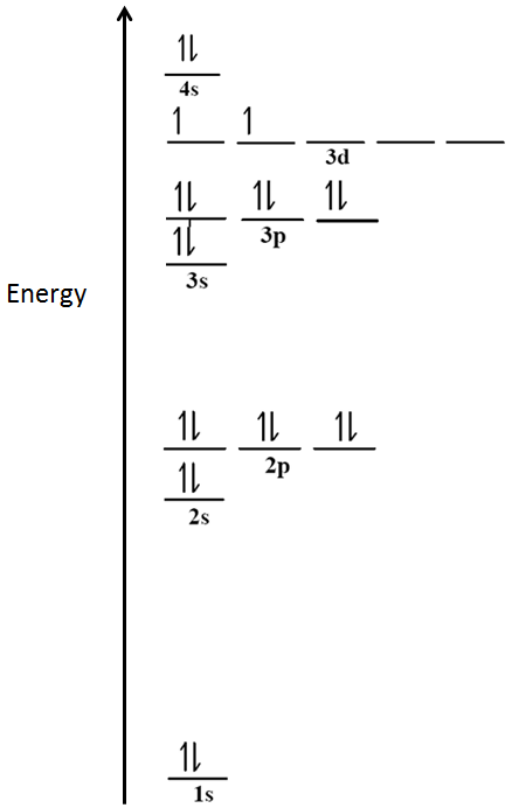
**(iii)** $_{36}Kr$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6$

**(iv)** $_{22}Ti^{2+}$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$

**(b)** $_{33}As^{3-}$ and $_{36}Kr$

**(c)** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5\ 4d^1$ or others

**(d)** Energy level diagram to illustrate the ground state electronic configuration of Ti:




**8.** **(a)** Li $1s^2\ 2s^1$ Na $1s^2\ 2s^2\ 2p^6\ 3s^1$

- Li has a **higher first ionisation energy** than Na.
- Na has **<u>one more electron shell</u>** than Li; the **<u>distance</u>** between the nucleus and valence electron in Na is **<u>greater</u>** than in Li.
- **<u>Shielding</u>** experienced by the valence electron of Na is **<u>greater</u>** than in Li.
- **<u>Despite the greater nuclear charge</u>** in Na, electrostatic attraction between the nucleus and the valence electron in Na is **<u>weaker</u>** than in Li.
- Hence, the 3s electron in Na require **<u>less energy</u>** for removal than the 2s electron in Li.

**(b)** Be $1s^2\ 2s^2$ B $1s^2\ 2s^2\ 2p^1$

- Be has **higher first ionisation energy** than B.
- The 2s electron to be removed from Be is at a **lower energy level** than the 2p electron to be removed from B and is hence **more strongly attracted by nucleus.**
- Hence **more energy** is needed to remove the 2s electron in Be.

**(c)** N $1s^2\ 2s^2\ 2p^3$ O $1s^2\ 2s^2\ 2p^4$

- N has **higher first ionisation energy** than O.
- The valence electron to be removed from N is an **unpaired 2p electron** while that to be removed from O is a **paired 2p electron**.
- The unpaired 2p electron in N experiences **less electron-electron repulsion / interelectronic repulsion** and hence requires **more energy** for removal.

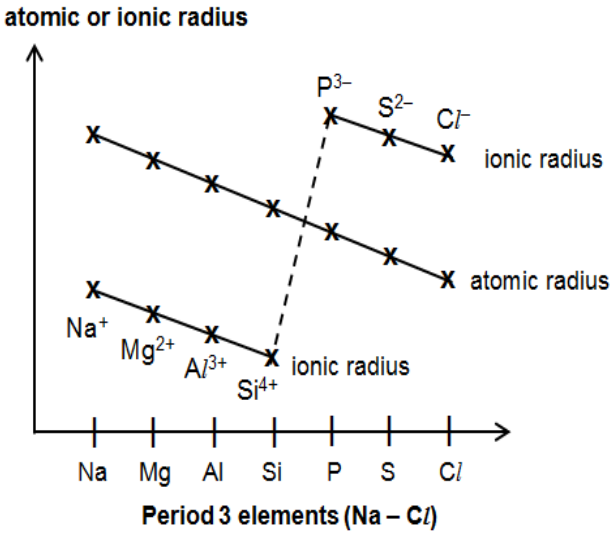
**(d)** F $1s^2\ 2s^2\ 2p^5$ Ne $1s^2\ 2s^2\ 2p^6$

- Ne has **higher first ionisation energy** than F.
- The Ne atom has a **higher nuclear charge**.
- While the **shielding effect remains approximately constant.**
- Hence electrostatic attraction between the nucleus and the valence electrons of Ne is **stronger**, resulting in a **greater amount of energy required** to remove the valence electron from an Ne atom.

**(e)** Ne $1s^2\ 2s^2\ 2p^6$ Na $1s^2\ 2s^2\ 2p^6\ 3s^1$

- Ne has a **higher first ionisation energy** than Na.
- Na has **one more electron shell** than Ne; the **distance** between the nucleus and valence electron in Na is **greater** than Ne.
- **Shielding** experienced by the valence electron in Na is **greater** than in Ne.
- Despite the **greater nuclear charge** in Na, electrostatic attraction between the nucleus and the valence electron in Na is **weaker** than in Ne.
- Hence, the 3s electron in Na require **less energy** for removal than the 2p electron in Ne.

**9. (a)**




**(b)** P $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$ $P^{3-}$ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

- Both $P^{3-}$ and P have the **same nuclear charge**.
- However $P^{3-}$ has **more electrons** than P leading to **greater shielding** (or electron-electron / interelectronic repulsion) and hence the electrostatic attraction between the nucleus and the outermost electron is **weaker** in $P^{3-}$, resulting in $P^{3-}$ having a larger electron cloud size, and a larger ionic radius than P.

**(c)** $Na^+$ $1s^2\ 2s^2\ 2p^6$ $Mg^{2+}$ $1s^2\ 2s^2\ 2p^6$

- $Na^+$ and $Mg^{2+}$ have the **same number of electrons/isoelectronic** and hence their outermost electrons experience the **same shielding effect**.
- However, $Mg^{2+}$ has a **higher nuclear charge** than $Na^+$.
- Consequently, the electrostatic attraction between the nucleus and outermost electrons is **stronger** in $Mg^{2+}$, hence the ionic radius of $Mg^{2+}$ is less than that of $Na^+$.

**10. (a)** There is a **large jump** in the **4th and 5th ionisation energies**.
This indicates that significantly more energy is needed to remove the 5th electron. Thus this **5th electron is located in an inner electron shell** that is **nearer to the nucleus**, experiencing less shielding and is attracted more strongly by the nucleus.
Therefore there are **4 valence electrons**.
Hence **Q** belongs to **Group 14** of the Periodic Table.

**(b)** $ns^2\ np^2$

**(c)** The 6th ionisation energy of **Q** is the energy required to remove one mole of electrons from one mole of gaseous $\mathbf{Q^{5+}}$ ions to form one mole of gaseous $\mathbf{Q^{6+}}$ ions.

$$\mathbf{Q}^{5+}(g) \longrightarrow \mathbf{Q}^{6+}(g) + e^-$$

**(d)** The 6th ionisation energy (IE) would be greater than the 5th.

- The number of protons of $\mathbf{Q}^{5+}$ and $\mathbf{Q}^{4+}$ are the same, hence nuclear charge is the same
- There are fewer electrons in the outermost shell of $\mathbf{Q}^{5+}$, thus shielding experienced by the remaining outermost electrons in $\mathbf{Q}^{5+}$ is less than that in $\mathbf{Q}^{4+}$
- Electrostatic attraction between the nucleus and remaining outermost electrons is stronger in $\mathbf{Q}^{5+}$ than in $\mathbf{Q}^{4+}$, hence the 6th ionisation energy of **Q** would be greater than the 5th.

**11.** **(a) (i)** Energy difference between the 1st and the 2nd ionisation energies (IEs) is smaller because both involve the removal of electrons from the **same outermost electron shell**.

The difference between the 2nd and 3rd IEs is greater because the 3rd electron to be removed is located in **an inner electron shell** / an electron shell of a lower principal quantum number, which is **closer to the nucleus**, experiencing less shielding and is attracted more strongly by the nucleus.

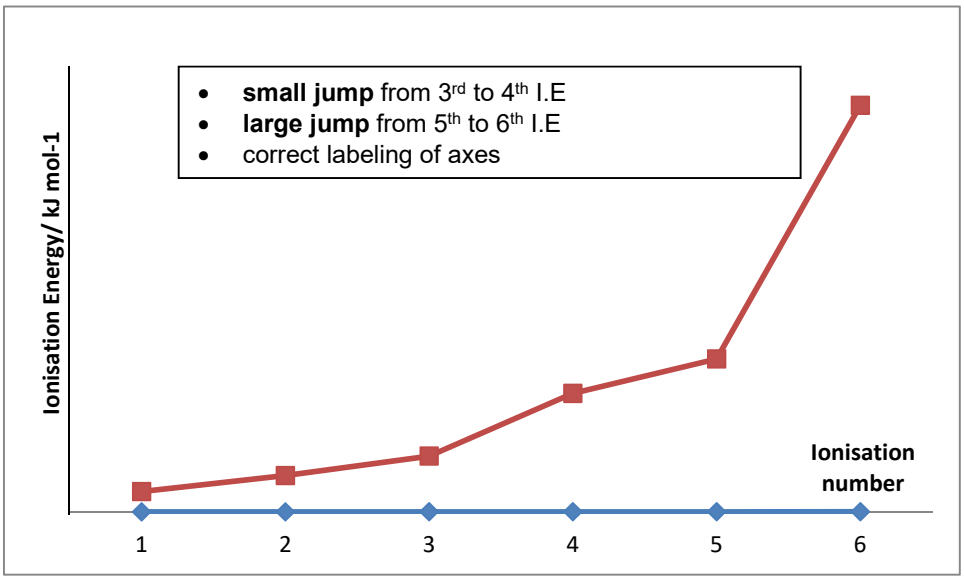
Thus, the 3rd IE is very much higher than the 2nd IE, resulting in a much greater energy difference.

**(ii)** No. Element **Q** is from **Group 1** due to the big jump in the 1st and 2nd IEs. Given that 5 IEs are given, element **Q** must have **at least 5 electrons**. Group 1 element in Period 2 (i.e. Li) **has only 3 electrons**.

**(iii)** Energy required = 0.02 (420 + 3050 + 4420) = +157.8 kJ

**(b) (i)** $\mathbf{A}^+(g) \rightarrow \mathbf{A}^{2+}(g) + e^-$ [state symbols must be present]

**(ii)**




**(iii)** Element **G** is from Group 2 since its 2nd IE is much lower than that of the preceding element. Hence, **D** is from Group 17.
Since all the atomic numbers are less than 18, **D** belongs to Period 2.
We can use this to deduce the electronic configurations of the species responsible for the 2nd IE of **C** and **D**.

Outermost shell electronic configuration of $\mathbf{C}^+$: $2s^2\ 2p^3$
Outermost shell electronic configuration of $\mathbf{D}^+$: $2s^2\ 2p^4$

The electron to be removed from $\mathbf{D}^+$ is a **paired 2p electron** while that to be removed from $\mathbf{C}^+$ is an **unpaired 2p electron**.
Due to **greater electron-electron repulsion / interelectronic repulsion** between paired electrons in the same orbital, less energy is required to remove the paired 2p electron from $\mathbf{D}^+$.
Hence the 2nd IE of **D** is lower than that of **C**.

**(iv)** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$