

COLLATED QUESTIONS

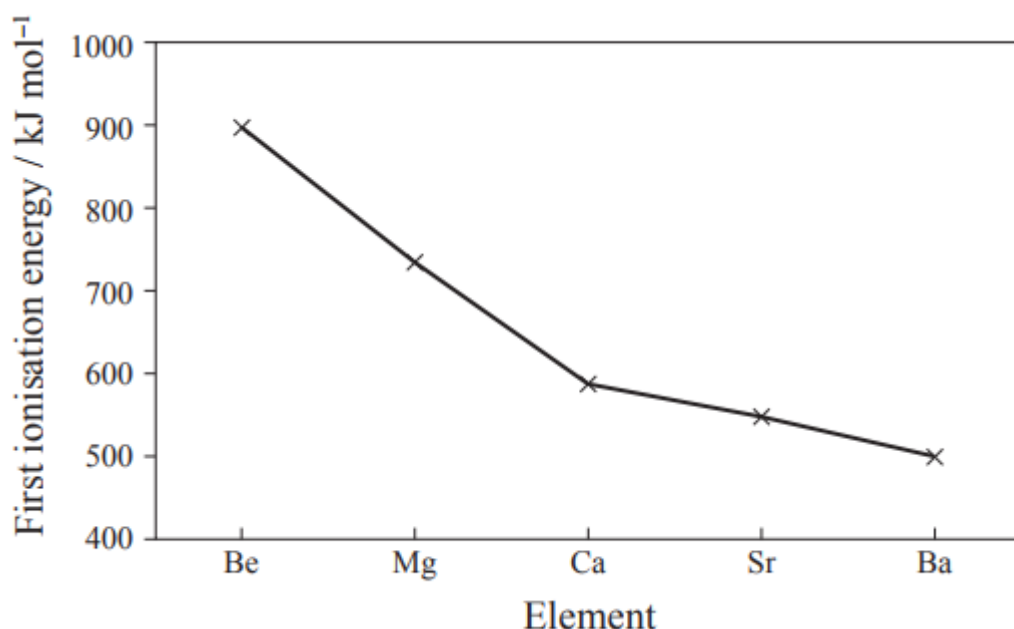
Electron configuration of atoms and ions of the first 36 elements (using *s, p, d* notation), periodic trends in atomic radius, ionisation energy, and electronegativity, and comparison of atomic and ionic radii

2017:1

(a) Complete the following table.

Symbol of particle	Electron configuration (use s, p, d notation)	Charge	Atomic number
Cl		0	
		+2	20
Mn ²⁺			

- (b) (i) Define the term electronegativity.
 (ii) Explain why the electronegativity of chlorine is greater than that of phosphorus.
- (c) The following graph shows the first ionisation energies of the Group 2 elements from Be to Ba.
 First ionisation energies of Group 2 elements



- (i) Write an equation to show the first ionisation energy for the element calcium.
 (ii) Explain the trend shown of first ionisation energies of the Group 2 elements.

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2016:1

(a) Complete the following table.

Symbol	Electron configuration
Cl	
Zn	
Cr ³⁺	

(b) (i) Explain why the radius of the Cl atom and the radius of the Cl⁻ ion are different.

	Radius (pm)
Cl atom	99
Cl ⁻ ion	181

(ii) Explain the factors influencing the trends in electronegativity and first ionisation energy down a group of the periodic table. In your answer you should:

- define both electronegativity and first ionisation energy
- explain the trend in both electronegativity and first ionisation energy down a group
- compare the trend in electronegativity and first ionisation energy down a group.

2015: 1

(a) Complete the electron configurations (s,p,d notation) for Al, Cu²⁺ and Sc

(b) Define the terms electronegativity and first ionisation energy.

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- (c) The following table shows the first ionisation energy values for elements in the third period of the periodic table.

Element	First ionisation energy/kJ mol ⁻¹
Na	502
Al	584
Si	793
Ar	1527

Justify the periodic trend of first ionisation energies shown by the data in the table above, and relate this to the expected trend in atomic radii across the third period.

2014: 1

- (a) Complete the electron configurations (s,p, d notation) for K, Cr and As.
 (b) Explain the difference between the radii of the K atom and the K⁺ ion.
 (c) The following table shows the electron configurations of four atoms, He, B, N, and Ne.

Arrange these atoms in order of increasing first ionisation energy by writing the symbol of the appropriate atom in the boxes below.

Atom	He	B	N	Ne
Electron configuration	1s ²	1s ² 2s ² 2p ¹	1s ² 2s ² 2p ³	1s ² 2s ² 2p ⁶

	→		→		→	
lowest ionisation energy						highest ionisation energy

2013: 1

- (a) Complete the electron configurations (s,p, d notation) for Se, V and V³⁺.
 (b) Discuss the data for each of the following pairs of particles.
 (i)

Atom	Electronegativity
O	3.44
Se	2.55

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(ii)

Atom or ion	Radius / pm
Cl	99
Cl ⁻	181

(iii)

Atom	First Ionization energy / kJ mol ⁻¹
Li	526
Cl	1257

2012: 1 (From expired AS 90780) – material no longer examined has not been included

(a) Complete the following table.

Symbol	Electron Configuration
Ge	
Cu	
Cu ⁺	

(c) Match the atoms and ions in the table below to the given radii.

Radii: **77 pm 123 pm 128 pm**

Symbol	Radii
Ge	
Cu	
Cu ⁺	

Justify your answer.

2011: 1 (From expired AS 90780)

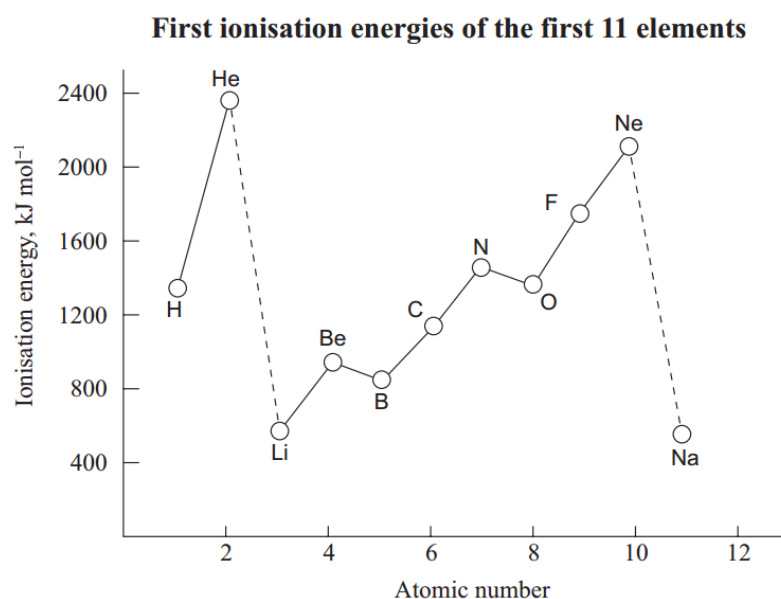
(a) Complete the following table.

Symbol	Electron Configuration
Fe	
Al	
Al ³⁺	
Na	

(b) State which has the larger radius, Al or Al³⁺. Justify your answer.

(c) (i) Write a balanced ion-electron equation to show the first ionisation of lithium.

(ii) With reference to the graph below, discuss the general trends in ionisation energies from lithium to sodium, and account for any anomalies.



2010:1 (From expired AS 90780)

(a) Complete the following table.

Symbol	Electron Configuration
Ca	
Cr	
Mn ²⁺	

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- (c) Match the atoms and ions in the table below to the radii given.

Radii: **99 pm 137 pm 197 pm**

Justify your answer.

Symbol	Radii
Ca	
Ca ²⁺	
Mn	

2009: 1 (From expired AS 90780)

- (a) Write the electron configuration using s, p, d notation for: Ca²⁺, Br, Fe²⁺.
- (c) Account for the differences in the atomic or ionic properties given below.
- (i) Atom Ionisation energy (kJ mol⁻¹) Ca 596, Br 1146
- (ii) Atom/ion Radius (pm) Br 114, Br⁻ 196, I 133

Answers

2016:1

(a) Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$

Zn: $[Ar] 3d^{10} 4s^2$

Cr³⁺: $[Ar] 3d^3$

- (b) (i) The Cl atom gains one electron to complete its valence shell to form the Cl⁻ ion, the nuclear charge remains the same. The increased inter-electron repulsion in the outer energy level causes the valence electrons to move further from the nucleus, so the Cl⁻ ion is larger than the Cl atom.
- (ii) **Electronegativity** decreases down a group. Electronegativity is a measure of how strongly an atom attracts bonding electrons. Although the nucleus will become increasingly positive down a group (number of protons increases), the atomic radius increases down a group as more energy levels are added and shielding / repulsion from inner shells increases. Therefore, the bonding electrons in the valence shell will be further from the positive nucleus, resulting in a weaker electrostatic attraction between the nucleus and the bonding electrons.

First ionisation energy is a measure of how easily the first mole of electrons is removed from one mole of gaseous atoms. It becomes easier to remove an electron down a group / first IE decreases down a group as the valence electrons are further from nucleus with greater repulsion / shielding from inner shells, so there is less electrostatic attraction between protons in the nucleus and valence electron to be removed.

For both EN and first IE, the attraction between the positive nucleus and bonding / valence electrons in the outer shell is decreasing down a group, so both EN and first IE decrease down a group.

2015:1

(a) Al = $[Ne] 3s^2 3p^1$ where $[Ne] = 1s^2 2s^2 2p^6$

Cu²⁺ = $[Ar] 3d^9$

Sc = $[Ar] 3d^1 4s^2$ where $[Ar] = 1s^2 2s^2 2p^6 3s^2 3p^6$

- (b) Electronegativity is the ability of an atom in a compound to attract electrons to itself. First ionisation energy is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms.
- (c) First ionisation energy increases from 502 in Na to 1527 in Ar. There is an increase in the number of protons and thus the nuclear charge / attractive force of the nucleus. As the electrons are added to the same energy level, there is no increase in repulsion between energy levels. The nuclei with a greater number of protons have a stronger electrostatic attraction for the valence electrons in the third shell, thus the first ionisation energy increases across a period.
- Both periodic trends are influenced by nuclear charge and the number of shells / distance, the ionisation energy increases while the atomic radii decrease. The larger the ionisation energy the more strongly the valence electrons are held. Thus atomic radii across Period 3 decrease.

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2014: 1

- (a) K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ [Ar] $4s^1$
 Cr $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ [Ar] $3d^5 4s^1$
 As $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$ [Ar] $3d^{10} 4s^2 4p^3$
- (b) The K^+ ion has a smaller radius than the K atom, as the ion has lost an electron from the valence/outer energy level, and therefore has fewer shells. This results in greater attraction between the nucleus and the valence electrons, as the outer electrons are now closer to the nucleus. There is less repulsion between the remaining electrons. Both species have the same number of protons / amount of nuclear charge.
- (c) lowest B N Ne He highest

2013: 1

- (a) Se: [Ar] $3d^{10}4s^24p^4$ or $4s^23d^{10}4p^4$
 V: [Ar] $3d^34s^2$ or $4s^23d^3$
 V^{3+} : [Ar] $3d^2$ where [Ar]: $1s^22s^22p^63s^23p^6$
- (b) (i) Se has more shells/electrons in energy levels further from the nucleus than O, with increased shielding from inner shells. This means there is a weaker electrostatic attraction between the nucleus and the bonded electrons, so Se has a lower electronegativity than O.
- (ii) Cl^- has an extra electron in its outermost/same energy level. This causes increased repulsion between electrons in the valence shell, so the electrons move further apart. This makes Cl^- bigger than Cl. Both Cl and Cl^- have the same number of protons/attractive force of the nucleus remains the same.
- (iii) Cl has more protons than Li. Therefore there is a greater attraction between the nucleus and outer electrons/electrons held more tightly so it is harder to remove an electron from Cl than Li.
 Even though the valence electrons of Cl are in the 3rd energy level/has an extra energy level the extra shielding is not as significant as the effect of the increased nuclear charge, so Cl has a higher first ionisation energy than Li.

2012: 1

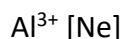
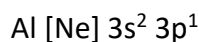
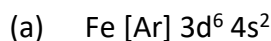
- (a) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^2$ or [Ar] $3d^{10} 4s^2 4p^2$
 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ or [Ar] $3d^{10} 4s^1$
 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ or [Ar] $3d^{10}$
- (c) Ge = 123 pm Cu = 128 pm Cu^+ = 77 pm

Both atoms have the same number of electron shells/ energy levels / shielding of outer electrons by inner electrons/ valence electrons in same energy level. Ge, however, has a

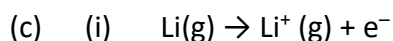
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greater nuclear charge / number of protons, compared to Cu, so there is a stronger attraction for the valence electrons, bringing them in closer, resulting in a smaller radius. Cu^+ has fewer electron shells than the Cu/Ge atoms (only 3 vs 4) and hence the electrons are closer to the nucleus meaning it is the smallest of the three particles. $\text{Cu} > \text{Ge} > \text{Cu}^+$.

2011: 1



- (b) Al has the larger radius. Al^{3+} has lost 3 electrons / valence shell. This means that there is one less energy level than in Al. The remaining electrons are drawn closer by nuclear charge / nuclear attraction greater causing smaller size.



- (ii) As you move across a period from Li to Ne, the ionisation energies increase. Electrons are added to the same valence shell / the same distance from the nucleus. Extra protons in the nucleus increase the nuclear charge, so the electrons in the valence shell are held more tightly and ionisation energy is greater.

As you go down a group, ionisation energy decreases. This is due to a new energy level being added, which is further from the nucleus. Electrons can be removed more easily and the ionisation energy is less/shielding increase explained

The drop Be and B is due to B having 1 electron in the p subshell ($2p^1$) and Be being $2s^2$. Although B has a greater nuclear charge, the electron in the p subshell is further from the nucleus/has less stability. Thus the p-electron in B's valence shell is not held so tightly/is more easily removed.

Drop N – O N has $\frac{1}{2}$ full subshell ($2p^3$) and O 1 more electron giving it a partly full subshell ($2p^4$). Added electron is going into suborbital already occupied by an electron – increased electron-electron repulsion so makes electron more easily removed / partly full subshell less stable so electron more easily removed.

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2010:1

(a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ or $[\text{Ar}] 4s^2$

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$ or $[\text{Ar}] 3d^5$

$4s^1 1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$ or $[\text{Ar}] 3d^5$

(c) $\text{Ca} = 197 \text{ pm}$, $\text{Ca}^{2+} = 99 \text{ pm}$ $\text{Mn} = 137 \text{ pm}$. Both Ca and Mn have the same number of electron shells/ energy levels / shielding of outer e's by inner e's / valence e's in same energy level (same orbital 4s). But Mn has a greater nuclear charge / no of protons so there is a stronger attraction for the valence electrons, bringing them in closer, resulting in a smaller radius.

$\text{Ca} > \text{Ca}^{2+}$ or $\text{Mn} > \text{Ca}^{2+}$ Ca^{2+} is smallest because it has lost electrons from an entire valence shell, so the electrons are in only 3 shells instead of 4 / less shells.

2009: 1

(a) $\text{Ca}^{2+} 1s^2 2s^2 2p^6 3s^2 3p^6$ OR $[\text{Ar}]$

$\text{Br} 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$ OR $[\text{Ar}]3d^{10} 4s^2 4p^5$

$\text{Fe}^{2+} 1s^2 2s 2p^6 3s^2 3p^6 3d^6$ OR $[\text{Ar}]3d^6$

(c) (i) Valence electrons are added to same shell / distance from nucleus similar. In Br, there is a greater number of protons / nuclear attraction greater (ENC), so valence electron more strongly held (implying IE).

(ii) $\text{Br}^- > \text{Br}$ added electron increases electron-electron repulsion, increasing size of the electron cloud so Br^- larger.

$\text{Br} < \text{I}$. I outer shell electrons are in an extra energy level / shell further from the nucleus and shielding of outer electrons so I larger.

$\text{Br}^- > \text{I}$ increase in repulsion when e- added to form the ion, has greater influence than the energy level difference for the valence e- So Br^- larger.