AS91390 Demonstrate understanding of thermochemical principles and the properties of particles and substances

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

2023:2

(a) (i) Write the equation to show the reaction that has an enthalpy change equal to the first ionization energy for the element gallium, Ga.

(ii) Use your knowledge of periodic trends to identify and justify the difference in the first ionization energy of gallium, Ga, and selenium, Se.

(iii) Use your knowledge of periodic trends to explain why gallium, Ga, has a larger atomic radius than boron, B.

2023:3

(a) (i) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
S	
Co	
Cr^{3+}	

(ii) Explain why the radii of the Ca atom and Ca^{2+} ion are different.

	Radius / pm
Ca atom	197
Ca^{2+} ion	100

2022:2

(a) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
Br	
V	
Ni^{2+}	

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2022: 3

(a) Explain why the radii of the Cl atom and the Cl^- ion are different.

Radius of Cl atom = 99 pm Radius of Cl^- ion = 181 pm

(b) Justify why both first ionisation energy and electronegativity increase across a period, but atomic radius decreases across a period.

2021:1

(b) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
Sc	
Ga	
Fe^{3+}	

2021:2

(a) Explain the difference in the atomic radii of calcium and selenium.

	Radius / pm
Calcium, Ca	197
Selenium, Se	116

(b) Justify, with reference to the factors affecting periodic trends, why fluorine is the most electronegative element in Group 17.

2020:3

(c) (i) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
Mn	
As	
Cu^{2+}	

(ii) Explain why the radii of the Mg atom and the Mg^{2+} ion are different.

	Radius / pm
Mg atom	160
Mg^{2+} ion	72

(c) (i) Write the equation to show the reaction that has an enthalpy change equal to the first ionisation energy for the element arsenic, As.

(ii) Justify the difference in first ionisation energies for nitrogen, potassium, and arsenic.

Element	First ionisation energy/kJ mol^{-1}
Nitrogen, N	1407
Potassium, K	425
Arsenic, As	953

2019:1

(d) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
Cr	
Fe^{3+}	
Ge	

(c) (i) Explain why the radii of the S atom and the S^{2-} ion are different.

	Radius / pm
S atom	104
S^{2-} ion	184

(ii) Justify the difference in electronegativities for oxygen, sodium, and sulfur.

Element	Electronegativity
Oxygen, O	3.44
Sodium, Na	0.93
Sulfur, S	2.58

2018:1

(d) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
V	
Cu^{2+}	
Br^{-}	

(b) Explain the factors influencing the trends in first ionisation energy and atomic radius across the second period of the periodic table. In your answer, you should:

- describe the trends in both first ionisation energy and atomic radius across the second period
- explain the factors influencing the trends in first ionistion energy and atomic radius across the second period
- relate the trend in first ionisation energy to the trend in atomic radius.

2017:1

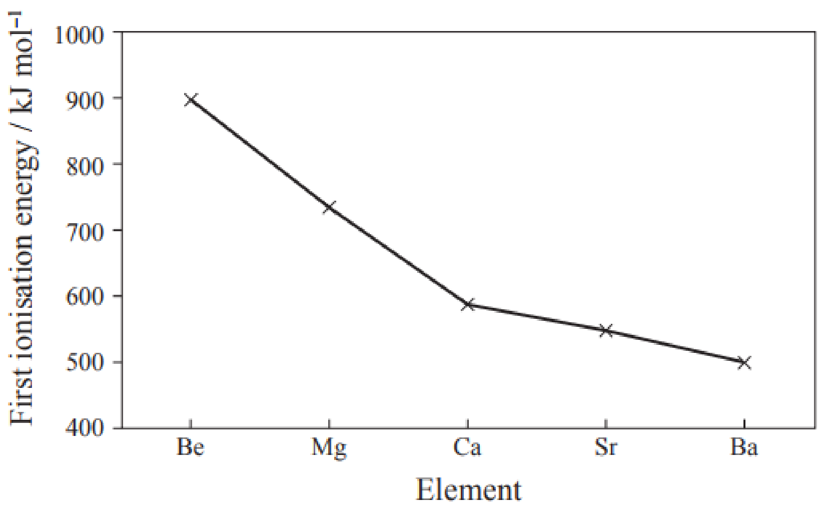
(a) Complete the following table.

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 radium 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, Cu^{2+} and Sc

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

(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.

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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	$1s^2$	$1s^22s^22p^1$	$1s^22s^22p^3$	$1s^22s^22p^6$

☐ → ☐ → ☐ → ☐

lowest ionisation energy … highest ionisation energy

2013: 1

(a) Complete the electron configurations (s,p, d notation) for Se, V and V^{3+}.

(b) Discuss the data for each of the following pairs of particles.

(i)

Atom	Electronegativity
O	3.44
Se	2.55

(ii)

Atom or ion	Radius / pm
Cl	99
Cl^-	181

(iii)

Atom	First Ionization energy / kJ mol^{-1}
Li	526
Cl	1257

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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)

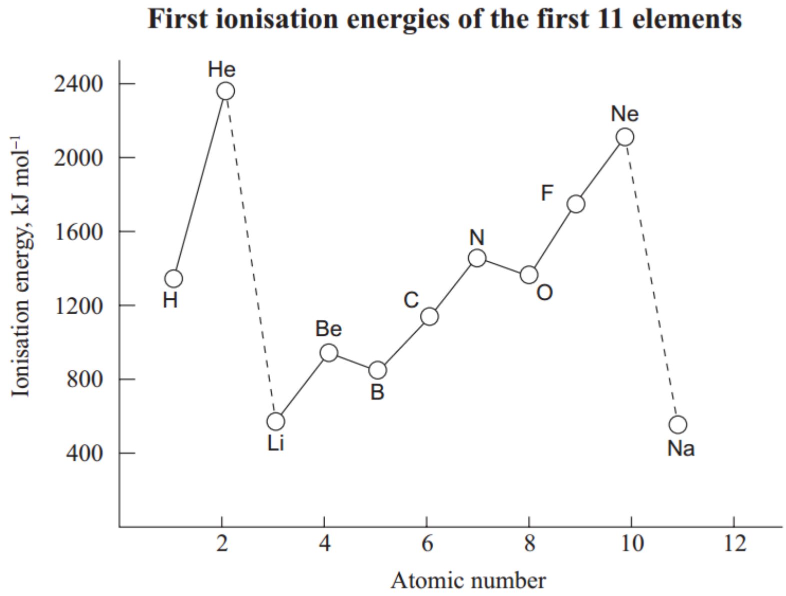
(a) Complete the following table.

Symbol	Electron Configuration
Fe	
Al	
Al^{3+}	
Na	

(b) State which has the larger radius, Al or Al^{3+}. 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 <u>general trends</u> 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+}	

(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	

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2009: 1 (From expired AS 90780)

(a) Write the electron configuration using s, p, d notation for: Ca^{2+}, Br, Fe^{2+}.

(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

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Answers

2023:2

(a) (i) $Ga(g) \rightarrow Ga^+(g) + e^-$.

(ii) Se has a greater first ionisation energy than Ga. Although the valence electrons are in the same energy level with the same repulsion / shielding from inner shells, the number of protons increases across a period. So, the electrostatic attraction between the positive nucleus and the valence electrons increases, and therefore more energy is required to remove the outermost electron, so Se has a greater first ionisation energy than Ga.

(iii) Going down a group, valence electrons are added to an energy level further from the nucleus with increased repulsion / shielding from inner shells. Although the number of protons increases down a group, this attraction is offset by the increasing distance between the nucleus and the valence electrons. So, the electrostatic attraction between the positive nucleus and its valence electrons decreases, and therefore the atomic radius of Ga is bigger than the atomic radius of B.

2023: 3

(a) (i) S: $1s^22s^22p^63s^23p^4$
Co: $[Ar]3d^74s^2$
Cr^{3+}: $[Ar]3d^3$

(ii) The Ca atom loses its two valence/outer shell electrons when it forms the Ca^{2+} ion. Therefore the Ca atom has four occupied energy levels, whereas the Ca^{2+} ion only has three occupied energy levels. As a result, the Ca^{2+} ion has a smaller radius.

2022:2

(a) Br: $[Ar]\ 3d^{10}4s^24p^5$
V: $[Ar]\ 3d^34s^2$
$Ni^{2+:}$ $[Ar]\ 3d^8$

2022:3

(a) When the Cl atom gains one electron in its valence energy level to become the Cl^- ion, there is increased electron-electron repulsion in the valence energy level while nuclear charge / number of protons remains the same. This causes the valence electrons to move further away from the nucleus, so the Cl^- ion is larger than the Cl atom.

(b) Across a period, the valence electrons / bonding electrons are both found in the same energy level with the same repulsion (shielding) from inner energy levels. The number of protons increases across a period / nuclear charge increases. This means the electrostatic attraction between the positive nucleus and the valence electrons / bonding electrons increases across a period, and therefore

- the atomic radius decreases.
- more energy is required to remove the outermost valence electrons, so the first ionisation energy increases.
- bonding electrons are more strongly attracted to the nucleus, so the electronegativity increases.

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2021:1

(a) Sc: [Ar] $3d^1$ $4s^2$ Ga: [Ar] $3d^{10}$ $4s^2$ $4p^1$ Fe^{3+}: [Ar] $3d^5$

2021:2

(a) Calcium and selenium are in the same period. Although the valence electrons are in the same energy level with the same repulsion / shielding from inner shells, the number of protons increases across a period. So, the electrostatic attraction between the positive nucleus and its valence electrons increases, and therefore the atomic radius of Se is smaller than the atomic radius of Ca.

(b) Fluorine is at the top of Group 17. Going up a group, the number of energy levels decreases, as does repulsion / shielding from inner energy levels. Fluorine is the smallest atom with the valence electrons closest to the nucleus. This means there is a strong electrostatic attraction between the positive nucleus of fluorine and any bonding electrons, so it is the most electronegative element in Group 17.

2020:3

(a) (i) Mn: [Ar] $3d^5$ $4s^2$ As: [Ar] $3d^{10}$ $4s^2$ $4p^3$ Cu^{2+}: [Ar] $3d^9$

(ii) The Mg atom and Mg^{2+} ion both have the same number of protons. The Mg atom loses its two valence electrons when it forms the Mg^{2+} ion. Whereas the Mg atom has three occupied energy levels, the Mg^{2+} ion only has two occupied energy levels. As a result, the Mg^{2+} ion is smaller.

2019:1

(a) Cr: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^1$ $3d^5$ or [Ar] $4s^1$ $3d^5$
Fe^{3+}: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^5$ or [Ar] $3d^5$
Ge: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^{10}$ $4s^2$ $4p^2$ or [Ar]$3d^{10}$ $4s^2$ $4p^2$

(c) (i) Sulfur and the sulfur ion have the same number of proton/nuclear charge but when the sulfur atom gains two electrons to form the sulfur ion, there is increased electron-electron repulsion in the valence energy level. As a result, the electrons move further apart, and therefore the S^{2-} ion has a larger radius than the S atom.

(ii) Electronegativity increases across a period, i.e. from Na to S. Both Na and S have the same number of energy levels and therefore the same shielding/electron-electron repulsion from inner levels. S has more protons/greater nuclear charge and therefore a greater attraction for valence/bonding electrons therefore greater electronegativity than Na.

Electronegativity decreases down a group. Sulfur has one more energy level and therefore increased shielding/electron-electron repulsion. Even though S has greater nuclear charge/more protons than O, because the valence electrons are further from the nucleus, electronegativity is lower.

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2018:1

(a) Complete the following table.

Symbol	Electron configuration (use *s, p, d* notation)
V	[Ar] $3d^3$ $4s^2$ or $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^3$ $4s^2$
Cu^{2+}	[Ar] $3d^9$ or $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^9$
Br^-	[Ar] $3d^{10}$ $4s^2$ $4p^6$ or $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^{10}$ $4s^2$ $4p^6$

(b) The first ionisation energy increases across the second period. There is an increase in the number of protons therefore the nuclear charge / attractive force of the nucleus increases. As the electrons are added to the same energy level across the second period, the electrostatic attraction for the valence electrons increases. This means more energy is required to remove an electron from the valence shell.
The atomic radius decreases across the second period. There is an increase in the number of protons therefore the nuclear charge / attractive force of the nucleus increases. Electrons are added to the same energy level as well. This causes the electrostatic attraction between the positive nucleus and the valence electrons to increase across the period pulling the valence electrons closer to the nucleus, so the atomic radius decreases.
As the ionisation energy increases, the atomic radius decreases, this is due to the same factor of increased nuclear charge due to more protons in the nucleus going across the period whilst electrons are adding to the same energy level. This decreased radius means more energy is required to remove the valence electron due to stronger attractive forces.

2017:1

(a)

Particle symbol	Electron configuration	Charge	Atomic number
	$1s^22s^22p^63s^23p^5$		17
Ca^{2+}	$1s^22s^22p^63s^23p^6$		
	$1s^22s^22p^63s^23p^63d^5$	+2	25

(b) (i) Electronegativity is the ability of an atom to attract a bonding pair of electrons to itself.

(ii) Electronegativity increases as you go across a period. Both Cl and P are row 3 elements and have valence electrons in their 3rd shell. The electrons are in the same shell so experience the same shielding effect. Chlorine has more protons in its nucleus than phosphorus so its nuclear charge is greater. This means that chlorine will have more attraction for the bonding pair of electrons so its electronegativity is greater.

(c) (i) $Ca(g) \rightarrow Ca^+ (g) + e^-$

(ii) The first ionisation energy is the energy required to remove one mole of the most loosely held electrons from one mole of gaseous atoms. The trend is that the ionisation energy decreases going down the group two elements. Although the nuclear charge increases due to more protons in the atoms going down a group, it is offset by the increasing distance of the outer electrons from the nucleus as the atomic radius increases due to more energy levels being added. The full inner energy levels shield the outer electrons from the protons in the nucleus so the electrostatic attraction is less. Additional energy levels result in greater shielding / repulsion between energy levels. The further the outer electron is from the nucleus, the less energy needed to remove it. *The trend is important, not the 'kink' at Ca, which requires no explanation.*

2016:1

(a) Cl: $1s^2 2s^2 2p^6 3s^2 3p^5$ Zn: [Ar] $3d^{10} 4s^2$ Cr^{3+}: [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.

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2015:1

(a) Al = [Ne] $3s^2\ 3p^1$ where [Ne] = $1s^2\ 2s^2\ 2p^6$
Cu^{2+} = [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.

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 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. Cu> Ge > Cu^+.

2011: 1

(a) Fe [Ar] $3d^6\ 4s^2$

Al [Ne] $3s^2\ 3p^1$

Al^{3+} [Ne]

Na [Ne] $3s^1$

(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.

(c) (i) $Li(g) \rightarrow Li^+(g) + e^-$

(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 ½ 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.

2010:1

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

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

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

(c) Ca = 197 pm, Ca^{2+} = 99 pm Mn = 137 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.

Ca > Ca^{2+} or Mn > 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) Ca^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ OR [Ar]
Br $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$ OR $[Ar]3d^{10}\ 4s^2\ 4p^5$
Fe^{2+} $1s^2\ 2s\ 2p^6\ 3s^2\ 3p^6\ 3d^6$ OR $[Ar]3d^6$

(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) Br^- > Br added electron increases electron-electron repulsion, increasing size of the electron cloud so Br^- larger.

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

Br^- > 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.