3.1 The Periodic Table & Periodic Trends

Mark Schemes

Course	DP IB Chemistry
Section	3. Periodicity
Topic	3.1 The Periodic Table & Periodic Trends
Difficulty	Hard

To be used by all students preparing for DP IB Chemistry SL
Students of other boards may also find this useful

1

The correct answer is **B** because:

- All three species (K^+, Cl^- and Ar) have the same number of electrons (18) and hence the same number of electron orbitals
- Therefore the size of each species is determined by the **proton number**
- The species with the most protons will have the **most positive** nuclear charge, therefore exerting the biggest "pull' on the negatively charged electrons
- K^+ has 19 protons, Ar has 18 and Cl^- has 17, so K^+ will have the smallest radii whereas Cl^- will have the large

A, **C** & **D** are incorrect as	these options do not place the species in the correct order of decreasing radii

2

The correct answer is **C** because:

- Moving across the period the **atomic number** increases by 1 for each element, hence the effective nuclear charge becomes greater
- The electrons are added but to the **same energy level**, hence the increase in **shielding** is negligible

A, **B** & **D** are incorrect as	statements 2 and 3 are correct. Statement 1 is incorrect as the electron shells are added moving down a group, not across a period

3

The correct answer is **C** because:

- The electronic configuration of beryllium is $1s^2\ 2s^2$ and therefore the first ionisation requires removing an electron from a **full 2s subshell**
- The electronic configuration of boron is $1s^2\ 2s^2\ 2p^1$ and therefore the first electron will be removed from the 2p subshell, which is higher in energy than 2s and is **further away** from the nucleus than the 2s orbital

A is incorrect as	the outer shell of boron is not full
B is incorrect as	atomic radius **decreases** moving across the table
D is incorrect as	the atomic number of boron is one more than beryllium

4

The correct answer is **C** because:

- The electronic configuration of option C indicates that the element is **neon**
- First ionisation energies generally increase moving across the periodic table from left to right
- Neon is a noble gas and has a **full outer shell** which makes it very stable and difficult to ionise
- More energy is required to remove an electron from a complete shell than one with unpaired electrons or an incomplete shell

A is incorrect as	its outer electron is located in 3s, which will be easier to remove than an electron from the second shell

B & **D** are incorrect as	these atoms have incomplete shells hence their ionisation energies are not as high as an atom with a full outer shell

5

The correct answer is **D** because:

- The first ionisation energy of an element is the energy required to remove one mole of electrons from one mole of gaseous atoms to produce one mole of gaseous ions: X(g) ? $X^-(g) + e^-$
- **Group 2 elements** have the 3rd lowest ionisation energy in their Period because despite the overall increase in ionisation energy across a Period, Group 3 elements have **lower** first ionisation energy than group 2 elements because the electron in the **p orbital** is more **shielded** and therefore the **nuclear pull** is reduced
- Therefore, the element is either **magnesium or calcium**
- The atomic radius is the distance between the nucleus and the outer electron orbital and increases down a group because an extra shell is being added each time
- Therefore the 3rd smallest atomic radius in a Group would belong to the 3rd element down the Group, which in this case is **calcium**

6

The correct answer is **D** because:

- As you go down a group (Na → K) the ionisation energy decreases
- This is due to the atomic radius increasing as there are more inner shells resulting in a greater shielding effect
- Nuclear attraction for the outer electron decreases
- Resulting in Na (second element) having a higher first ionisation energy than K (first element)

A is incorrect as	Mg → Al is going across a period and the ionisation energy value for Al is lower than Mg as the 13th electron in aluminium is in the 3p subshell which is further from the nucleus than the 3s subshell of magnesium and will have extra shielding from the 3s subshell
B is incorrect as	N → O is going across a period and the ionisation energy value for O is lower than N as this marks the start of electron pairing in the p orbitals, whereby electron repulsion causes a decrease in ionisation energy
C is incorrect as	Ne → Na is the start of a new period and the ionisation energy of Na will be lower than Ne as there will be more electron shells and shielding, resulting in a lower ionisation energy for Na

7

The correct answer is **A** because:

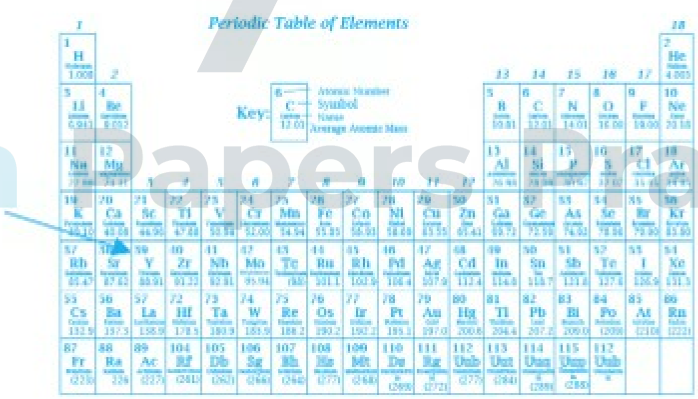
- There is a **general rise** in first ionisation energy as you go **across a period** because the **nuclear charge increases** and therefore the pull on the electrons increases
- When electrons are added to a **new orbital** (at the start of a new period) the first ionisation energy **drops**
- This is because of:
  - the increase in **shielding,** (due to an extra orbital between the electron and the nucleus)
  - the increased **distance** and therefore the **reduced attraction** to the nucleus outweighs the pull from the extra proton
- The means the 4 consecutive elements are from Group 7 (A), Group 8 (B), Group 1 (C), Group 2 (D)
- For an element to form a covalent compound with hydrogen with the formula HX, we know it needs to have **7 electrons** in its outer ring (so it

can bond with hydrogen to get the extra 1 electron needed for a full orbital) and therefore we know A is a **Group 7** element

8

The correct answer is **D** because:

- The order in increasing energy of subshells is s < p < d < f

A is incorrect as	the fourth energy shell has four subshells, s,p,d and f which can hold 2,6,10 and 14 electrons, so the maximum number of electrons is 32
B is incorrect as	one orbital contains a pair of electrons; the d-subshell contains 5 orbitals and up to 10 electrons
C is incorrect as	the location of yttrium shows it is the first element in the 4d sublevel 

9

The correct answer is **B** because:

- First ionisation energy
  - Decreases in first ionisation energy across a period are found between groups 2 and 3 and groups 5 and 6
- Melting point
  - Melting point increases from groups 2 to 3 as there are stronger metallic bonds between the positive metal cations and delocalised electrons as you go across the period
  - The melting points between groups 5 and 6 in period 3 increase as a sulfur molecule S_8 has more electrons than a phosphorus molecule P_4, resulting in stronger van der Waals forces and a larger melting point
- Isoelectronic arrangement
  - The electron configuration of a sulfur ion is $1s^2 2s^2 2p^6 3s^2 3p^4$ which is isoelectronic with argon
- Sulfur is in period 3 in the p block of the periodic table

10

The correct answer is **C** because:

- Electron affinity and electronegativity follow the same trends
- Both properties are influenced by the effective nuclear charge
- The effective nuclear charge is the net positive charge experienced by an electron; it is reduced by the presence of filled shells
- This is why electron affinity and electronegativity both decrease down a group

Electron affinity is defined as the energy change when 1 mole of electrons is added to 1 mole of atoms in the gaseous state. It can be thought of as the opposite process to first ionization energy and is exothermic

$$M(g) + e^- \rightarrow M^-(g) \qquad \Delta H = -ve$$

This is not to be confused with **electronegativity** which is the tendency of atoms to attract a bonding pair of electrons in a covalent bond.