Exam Papers Practice

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	Medium

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

1

The correct answer is **C** because:

- Elements A-C are in the same period as one another and yet in C the p subshell is full, unlike in **A** and **B**
- **Remember** the trend;
  - Ionisation energies increase across a period
    - Nuclear charge increases
    - Atomic radius decreases
    - Stronger attractive forces between the nucleus and outer electrons
    - Larger first ionisation energies

A is incorrect as	fewer shells of electrons does not necessarily mean more energy is required to remove the electrons, remember that ionisation energy increases across a period because of the above factors
B is incorrect as	there are four 3p electrons - this means that one of the 3p orbitals will have a pair of electrons repelling, which leads to a slightly lower ionisation energy than might be expected
D is incorrect as	D is in the period below elements A-C and the outer electrons are in the 4s subshell which is further away from the nucleus, so easier to remove

2

The correct answer is **B** because:

- The question states that it is the **second** ionisation energy of magnesium, hence the electron is being removed from the Mg^+ ion
- The result is the Mg^{2+} ion plus and electron
- The sign of the enthalpy change is positive as energy is **required** to remove the electron, hence it is an **endothermic** process

A & **D** are incorrect as	the sign of the enthalpy change is incorrect
C is incorrect as	the equation shows two electrons being removed from a magnesium atom

Exam tip: Don't forget to include state symbols in ionization energy equations as they are part of the definition

3

The correct answer is **A** because:

- There is a **general increase** in first ionisation energy **across a period** because of the **nuclear charge** increases
- However, there is a slight reduction in first ionisation energy between **Groups 2 and 3** and between **Groups 5 and 6**
- This is due the loss of a single p^1 electron in the case of Group 3 and in the case of Group 6 due to repulsion experienced by a paired electron in p orbital

↑↓	↑	↑

- Therefore, sulfur has a slightly lower first ionisation energy than phosphorus
- Silicon has a **high melting point** because it has a **giant covalent structure** so strong covalent bonds must be broken to melt it
- Phosphorus (P_4) and sulfur (S_8) are both **simple covalent molecules,** but sulfur has a higher melting point because it contains **more electrons,** so the **London dispersion forces** are stronger
- The information given is consistent with **X, Y** and **Z** being silicon, phosphorus, sulfur respectively

B is incorrect as	the melting points would be increasing from Na to Al, so **Y** cannot be Mg, and the ionization energy of Al would be slightly less than Mg
C is incorrect as	**Y** cannot be silicon as it would have the highest melting point due to its giant covalent structure, and the ionization energy would be increasing from Al to P
D is incorrect as	**Z** cannot be silicon as it would have the highest first ionisation energy

4

The correct answer is **C**

- The first thing to do is to add +4 electrons onto [Ar] $3d^1$, which gives you
  [Ar] $3d^3$ $4s^2$
  - **Remember:** The 4s sublevel is the first to fill and first to empty So, make sure you add two electrons to 4s first and then fill 3d with the remaining two electrons
- This gives an atom with 23 electrons and it is in the d block (as stated in the question), so checking for an element with an atomic number of 23 and gives you the answer, vanadium

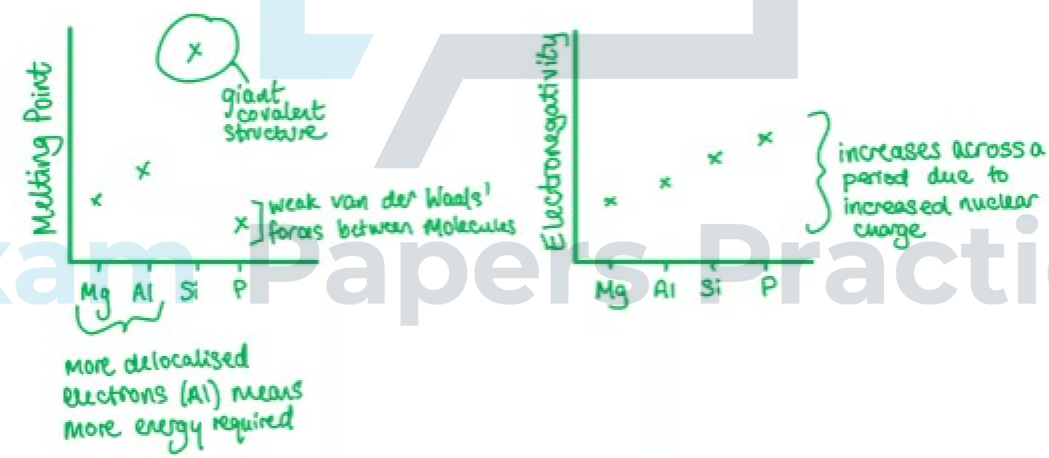
A & **B** are incorrect as	although copper and chromium are in the d block of the periodic table, copper does not form a +4 ion and chromium's +4 ion configuration is [Ar] $3d^2$
D is incorrect as	although silicon forms a +4 ion, it is not in the d block of the periodic table and it's ion electron configuration would be $1s^2 2s^2 2p^6$ or [Ne]

5

The correct answer is **C**

Melting point:

- Silicon has the highest melting point because it is a **giant covalent structure**, so to turn silicon from a solid to a liquid you have to break many strong covalent bonds which require lots of **energy**
- The melting point of the **metals** (Na, Mg, *Al*) increases across the period because the number of electrons **delocalised** from each metal atom increases (1 from Na, 2 from Mg, 3 from *Al*).
  - Therefore, more energy is required to force the delocalised electrons back into the electron shell
- Phosphorus and sulfur exist as **simple covalent molecules**, which means they have strong covalent bonds within molecules but only low amounts of energy are required to overcome the **Van der Waals' forces** between molecules in order to melt the solid



Electronegativity:

- Electronegativity is a measure of the tendency of an atom to **attract** a **bonding pair** of electrons
- Electronegativity **increases across a period** between group 1 and group 7
- This is because the number of electron orbitals remains the same, but the **nuclear charge increases** (because the number of protons increases)
- This means that phosphorus has the highest electronegativity

6

The correct answer is **C** because:

- A positive ion is smaller than a neutral atom whilst a negative ion is larger
  - Li^+ < Li and Na^+ < Na
  - Br^- > Br and Cl^- > Cl
- Ionic radius increases across a period hence we are looking for the smallest positive ion and largest negative ion.
- Ionic radius also increases down a group
  - Cl^- < Br^-
  - Li^+ < Na^+

A is incorrect as	Li^+ is smaller than Li as Li^+ has one less shell as it has lost its single valence electron. Cl^- is larger than Cl as Cl^- has gained an extra electron but has the same number of protons reducing its effective nuclear charge
B is incorrect as	Na is larger than Li as atomic radius increases down Group 1 as Na has an extra electron shell. Br is smaller than Br^- as Br^- has an extra electron but the same number of protons reducing its effective nuclear charge
D is incorrect as	Na^+ is larger than Li^+ as Na is below Li in Group 1 and Cl^- is smaller than Br^- as Br is below Cl in Group 17. Down a group, the number of energy levels (n) increases, so there is a greater distance between the nucleus and the outermost orbital

7

The correct answer is **D** because:

- Chlorine has the smallest **atomic radius** in period 3 (other than argon) because it has the largest number of **protons** and therefore the highest **nuclear charge**
- A bigger nuclear charge means a greater **pull** on the electrons to the nucleus, reducing the atomic radius, so atomic radius **decreases** across a period
- As argon is unreactive, **X** must be chlorine
- Phosphorus (P_4) and sulfur (S_8) are **simple covalent molecules,** so the melting point is determined by the energy required to break **van der Waals'** forces between molecules
- However, there are more electrons in sulfur molecules so phosphorus has a lower melting point than sulfur. Only chlorine and argon have lower melting points in Period 3
- Therefore, the information is consistent with **Y** being **phosphorous** and **Z** being **PCl_3**

8

The correct answer is **C** because:

- Arsenic has the electron configuration [Ar] $3d^{10}$ $4s^2$ $4p^3$
- Arsenic is in Group 15 so needs to gain 3 electrons to give it a full outer shell forming As^{3-}
- The electron configuration of As^{3-} is $3d^{10}$ $4s^2$ $4p^6$
    - This is the electron configuration of the Noble gas, krypton
- A useful way to visualise this is as As gains 3 electrons it is equivalent to it moving the places to the right of As which is Kr. This means As^{3-} has the same electron configuration as Kr
- Negative ions are larger than positive ions
    - Gallium and rubidium form positive ions
    - Arsenic and bromine form negative ions
- The more negative the ion, the larger the ionic radius
    - As^{3-} > Br^-

A is incorrect as	gallium is in group 13 so forms the Ga^{3+} ion as it needs to lose 3 extra electrons for a full outer shell.
B is incorrect as	bromine is in group 17 so forms the Br^- ion as it needs to gain an extra electron for a full outer shell.
D is incorrect as	rubidium is in group 1 so forms the Rb^+ ion as it needs to lose 1 extra electron for a full outer shell.

9

The correct answer is **C** because:

- Tin is located in Group 14 and period 5
- The fifth Period corresponds to elements with five energy levels
- The shorthand electron configuration of tin shows that the outermost shell is in Period 5

$$[Kr]\ 4d^{10}\ 5s^2\ 5p^2$$

- An element in Group 14 has four outer electrons. For the Group numbers that are 'teens' subtract 10 to get the number of outermost electrons.

Short versions of electron configurations are a handy way to save writing out the whole configuration for higher elements. Take the nearest noble gas, write its symbol in square brackets and continue the electron configuration from there.

$$[Ar] = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$$
$$[Kr] = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$$

10

The correct answer is **B** because:

- The elements are arranged in order of increasing atomic number or proton number
- The charge of the nucleus arises from the positively charged protons, so the nuclear charge will be the same as the atomic number

A is incorrect as	although most elements are also in order of increasing mass, there are a few exceptions like iodine and tellurium which are the other way around
C is incorrect as	reactivity varies according to the group, so Group I increases in reactivity down the group, but Group 17 decreases down the group
D is incorrect as	electronegativity increases left to right across the periodic table, but decreases down a group. This is a trend not an order