PAPER 1
Section A

1. **D**

	no. of p	no. of n	no. of e⁻
$^{36}_{16}S^{2-}$	16	20	18
$^{37}_{17}Cl^-$	17	20	18

2. **C**

	1st	2nd	3rd	4th	5th	6th
I.E.	950	1800	2700	4800	6000	12300
Δ		850	900	2100	1200	6300

The big jump from the 5th I.E. to the 6th I.E. shows that the 6th electron is removed from the next inner electron shell.

Hence, element X is a Group V element with 5 valence electrons and can form compounds with formula XCl_3.

3. **D**

Cu^{2+} has configuration of [Ar] $3d^9$ and thus has a single unpaired electron.

4. **C**

The element in Option C contains 3 valence electrons ($3d^14s^2$) and would lose all 3 valence electrons most easily to form X^{3+} ion which has a stable electronic configuration.

5. **D**

When a neutron changes to a proton, the nucleon number does not change. The proton number increases by one.
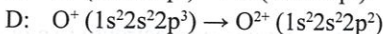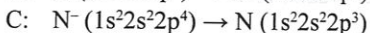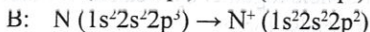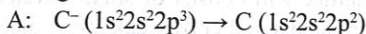
Tips

Nucleon No.(Mass No.) = No. of Proton + No. of Neutron

$^{40}_{19}K \longrightarrow {}^{40}_{20}Ca$
- 19 protons • 20 protons
- Nucleon No. = 40 • Nucleon No. = 40

6. **C**

A half-filled set of p orbitals has np^3 configuration.

A: $C^- (1s^22s^22p^3) \rightarrow C (1s^22s^22p^2)$
B: $N (1s^22s^22p^3) \rightarrow N^+ (1s^22s^22p^2)$
C: $N^- (1s^22s^22p^4) \rightarrow N (1s^22s^22p^3)$
D: $O^+ (1s^22s^22p^3) \rightarrow O^{2+} (1s^22s^22p^2)$
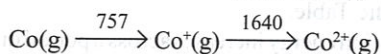
7. **A**

To form one mole of an ion with a 2+ charge, the 1st and 2nd I·E. must be involved.

$Al(g) \xrightarrow{577} Al^+(g) \xrightarrow{1820} Al^{2+}(g)$
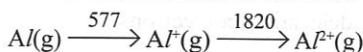
∴ To form Al^{2+}, $\Delta H = 577 + 1820$
= 2397 kJ mol⁻¹

$Co(g) \xrightarrow{757} Co^+(g) \xrightarrow{1640} Co^{2+}(g)$

∴ To form Co^{2+}, $\Delta H = 757 + 1640$
= 2397 kJ mol⁻¹

8. **D**

A – Cu atom has 29 electrons while Mg atom has 12 electrons
∴ Cu atom has more electrons than Mg atom

B – E.C. of Cu^{2+}: $1s^22s^22p^63s^23p^63d^9$
outer E.C.

E.C. of Mg^{2+}: $1s^22s^22p^6$
outer E.C.

C – E.C. of Cu^{2+}: $1s^22s^22p^63s^23p^63d^9$
E.C. of Mg: $1s^22s^22p^63s^2$
∴ Both particles have 3 electron shells which are occupied.

D – E.C. of Cu: $1s^22s^22p^63s^23p^63d^{10}4s^1$
E.C. of Mg^{2+}: $1s^22s^22p^6$
∴ Cu has 4 electron shells occupied whereas Mg^{2+} has only 2 electron shells that are occupied.

9. **B**

From the Data Booklet:
A – 1st I.E. is 799, 1090, 1400 and 1310 kJ mol⁻¹ respectively
⇒ does not follow the trend in the bar chart

B – 1st I.E. is 757, 736, 745 and 908 kJ mol⁻¹ respectively
⇒ follows the trend in the bar chart

C – 1st I.E. is 736, 590, 548 and 502 kJ mol⁻¹ respectively
⇒ does not follow the trend in the bar chart

D – 1st I.E. is 786, 1060, 1000 and 1260 kJ mol⁻¹ respectively
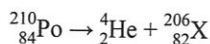⇒ does not follow the trend in the bar chart

10. **B**

No. of protons of Po – 84
No. of nucleons of Po = 210
∴ No. of neutrons of Po = 210 – 84 = 126
When an α-particle is emitted;
No. of protons of X = 84 – 2 = 82
No. of neutrons of X = 126 – 2 = 124

$$^{210}_{84}Po \rightarrow {}^4_2He + {}^{206}_{82}X$$

11. **D**

Neon and fluorine are in the same period of the Periodic Table.

Ionisation energy increases across a period due to increasing nuclear charge and decreasing atomic radius.

Since the electrons all go into the same shell, the shielding of the ionising electrons is about the same. The outer electrons are, therefore, increasingly more strongly attracted by the positive nucleus, and so, more energy is required to remove an electron.

☠ **Common Mistakes**

C – The atomic radius of fluorine is less than that of neon.

PAPER 3

1. (a) Nucleon number refers to the total number of neutrons and protons in the nucleus of an atom.

   (b) ① Electrons which are negatively charged are deflected towards the positive terminal, while protons being positively charged are deflected towards the negative terminal.

   ② Electrons (being lighter than protons) are deflected more than protons.

PAPER 1
Section A

1. **B**

$$\overset{\longleftrightarrow}{\underset{\delta-}{O}} = \overset{+}{\underset{\delta+}{C}} \overset{CH_3}{\underset{CH_3}{\diagdown}}$$
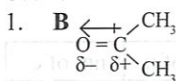
Molecule B has the largest dipole due to presence of a very electronegative O atom.

☠ **Common Mistakes**

Molecule C has a smaller dipole than Molecule B because the C=O and C–Cl bonds have very similar dipoles and tends to cancel each other out.

$$\overset{\longleftrightarrow}{O} = C \overset{Cl}{\diagup} \overset{}{\diagdown} Cl$$

2. **B**

In ice, each water molecule is hydrogen bonded to four other water molecules.

2 lone pairs (or 4 electrons) from each oxygen atom are involved in forming hydrogen bonds.

💡 **Tips**

A – All the angles surrounding oxygen atoms are 109.5° (tetrahedral shape around oxygen atom).

C – The hydrogen bonds (intermolecular bonding) are weaker than covalent bonds (intramolecular bonding).

D – The open structure causes ice to be less dense than water i.e. ice floats on water.

3. **C**

N–1 and N–3 exists as N⁻ ion and form ionic bonds with Mg^{2+} ion.

N–2 and N–4 form co-ordinate (dative covalent) bonds with Mg^{2+} ion.

$$\begin{array}{ccc} \overset{|}{-N^-} & & \overset{||}{N-} \\ & Mg^{2+} & \\ -N & & N^- \\ \overset{||}{} & & \overset{|}{} \end{array}$$
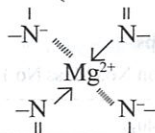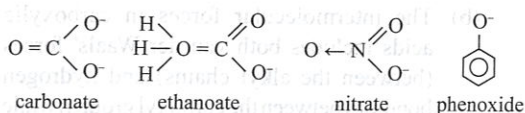
4. **D**

Copper has a metallic structure and its giant lattice consists of positive metal ions surrounded by 'sea of delocalised valence electrons'.

Iodine has a simple/discrete molecular structure and its lattice will consist of small covalent I_2 molecules held together by weak intermolecular (van der Waals) forces.

5. **C**

Delocalised electrons are electrons in a molecule that are not associated with a particular atom but extend over several adjacent atoms.

All four ions contain C=C double bond(s) and so, have delocalised π electrons.

$$O = C \overset{\displaystyle O^-}{\underset{\displaystyle O^-}{\Big<}} \qquad H-\overset{\displaystyle H}{\underset{\displaystyle H}{\overset{\displaystyle |}{\underset{\displaystyle |}{C}}}} \overset{\displaystyle O}{-} O=C \overset{\displaystyle O}{\underset{\displaystyle O^-}{\Big<}} \qquad O \leftarrow N \overset{\displaystyle O^-}{\underset{\displaystyle O^-}{\Big<}} \qquad \overset{\displaystyle O^-}{\bighexagon}$$

carbonate   ethanoate   nitrate   phenoxide

6. **B**

Dry hydrogen chloride does not ionise in methylbenzene. Hence, HCl exists as covalent molecules in the resultant solution and does not show any acidic properties since no H^+ ions are dissociated.

A – pH less than 7 due to presence of H^+ ions.

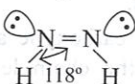B – HCl does not dissociate into ions, thus no mobile ions to conduct electricity.

C – No H^+ ions to have any acidic properties.

D – No H^+ ions to have any acidic properties.

7. **B**

Each nitrogen atom in di-imine has one lone pair and two bond pairs.

It has a bent shape with bond angle of $118°$ since bond pair – lone pair repulsion is greater than bond pair – bond pair repulsion.

$$\overset{\displaystyle \overset{\frown}{N}=\overset{\frown}{N}}{\underset{\displaystyle H \quad 118° \quad H}{\diagup \quad \diagup}}$$

8. **D**

Weak van der Waals' forces between the microscopic hairs and the non-polar surface of glass allows interaction to be formed and broken easily.
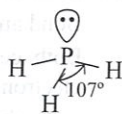
9. **B**

No. of electron pairs = 4

No. of bond pairs = 3

No. of lone pairs = 1

Molecular Shape = Trigonal Pyramidal

Bond $\triangle$ = $107°$

$$H-\overset{\displaystyle \overset{\displaystyle \cdot\cdot}{P}}{\underset{\displaystyle H}{\diagdown}}\overset{\displaystyle }{\underset{\displaystyle 107°}{\diagup}} H$$

10. **C**

The lattice structure of MgO is similar to that of NaCl. Each Mg^{2+} ion is surrounded by 6 other O^{2-} ions, and every O^{2-} ion is surrounded by 6 other Mg^{2+} ions.

11. **B**

C1 and C2 are sp hybridised while C3 is sp^3 hybridised. Hence, the sigma bond between C2 and C3 is a sp–sp^3 overlap.

12. **C**

A – Carbon 2 and 3 are sp hybridised.

B – The alcohol will be oxidised to a carboxylic acid. $KMnO_4$ will decolorised from purple colour, as it is being reduced to Mn^{2+} ions.

C – Carbon 1 and 4 are sp^3 hybridised.

D – Carbon 1 and 4 are sp^3 hybridised which has tetrahedral geometry.

Section B

13. **C**

In graphite, each carbon uses 3 valence electrons for covalent bonding. The fourth valence electron is used for π bonding between adjacent carbon atoms, resulting in delocalisation.

In the giant metallic structure of sodium, positive sodium ions are surrounded by a 'sea of delocalised valence electrons' throughout the giant lattice.

Tips

In cyclohexene, all the valence electrons are used for covalent bonding between the atoms.

14. **B**

CO_2, COS and COSe are all linear molecules. Electronegativity of Group VI elements decreases down the group.

Electronegativity: O > S > Se

Bond Polarity   : C=O > C=S > C=Se

15. **D**

1 – Diamond has tetrahedral shape and C–C–C bond angle is $109.5°$, whereas graphite has trigonal planar shape with C–C–C bond angle of $120°$.

2 – The C–C bond length in diamond is 0.154 nm and in graphite is 0.142 nm.

3 – Diamond has a giant covalent structure and all C–C covalent bonds are of the same strength.

Graphite has a giant structure of hexagonal rings layer over another layer held by weak van der Waals' forces. The C–C covalent bonds within the layer are of the same strength.

16. **B**

1 & 2 – Dative covalent bonds are formed between Mg and two of the N atoms (shown by ⋯⋯ line). The other two N atoms form sigma bonds to the Mg atom (shown by —— line).

3 – Since Mg atom is in a planar arrangement with the four N atoms, it cannot be sp^3 hybridised since sp^3 hybridisation will give molecule a tetrahedral arrangement/shape.

PAPER 2

1. (a)

**(b)** $Cl–C–Cl$ bond angle is 120°C.

> 💡 **Tips**
> Molecular shape is Trigonal Planar

2. (a) Co-ordination number refers to the maximum number of neighbouring ions that can be placed around another ion of the opposite charge in a crystal lattice.
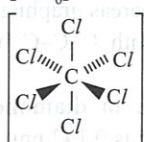
(b) The ionic radii of $Cs^+ > K^+ > Na^+$.

The larger ionic radii of Cs^+ ion allows more Cl^- ions to be packed around it. Thus, the co-ordination number in $CsCl$ lattice is larger than those in $NaCl$ and KCl.
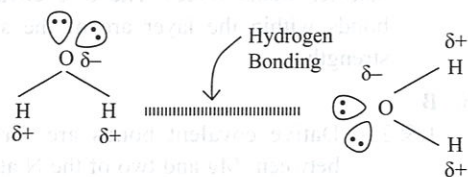
3. $[PCl_4]^+$
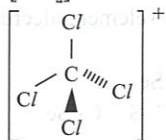
Shape: Tetrahedral

$[PCl_6]^-$

Shape: Octahedral

PAPER 3

1. (a) Example of Hydrogen Bonding:

Hydrogen bonding between H_2O molecules

Requirements for Hydrogen Bonding are:
① An electron-deficient hydrogen atom bonded to a very electronegative atom (such as N, O or F) in one molecule.
② A lone pair of electrons on a very electronegative atom (such as N, O or F) in the other molecule.

(b) The intermolecular forces in carboxylic acids includes both van der Waals' forces (between the alkyl chains) and hydrogen bonding (between the carboxyl groups) while that between water molecules are mainly hydrogen bonding.

The first four members of the series of carboxylic acids are fully soluble in water because they form hydrogen bonds with water molecules.

As the alkyl chain (of the carboxylic acids) gets longer, the acids become increasingly insoluble because there is stronger van der Waals' forces between the acid molecules.

As such, the energy released from the formation of hydrogen bonds between the carboxylic acids and water molecules is insufficient to compensate the energy required to break the stronger van der Waals' forces and hydrogen bonding between the acid molecules, as well as hydrogen bonding between water molecules.

Hence, the solubility of carboxylic acids in water decreases as the chain length increases.

2. (a)

(b) Bond angle of NO_2 : 118°
Bond angle of O_3 : 116°
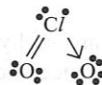Both central atoms are surrounded by 3 electron pairs.
∴ Basic angle is about 120°.
Since both has 1 'lone pair' and 2 bond pairs, Bond angle < 120°.
The single unpaired electron on NO_2 molecule will be less effective in causing lone pair – bond pair repulsion, and so the bond angle for NO_2 should be greater than that for O_3.

(c)

In the structure above, there are 9 electrons around the central chlorine atom. ClO_2 exists because Cl is a Period 3 element and has available low-lying 3d orbitals to expand its octet to accomodate more than 8 electrons.

The above structure of FO_2 does not exist because F is a Period 2 element and can only accomodate a maximum of 8 electrons (octet).

The above structure also does not exist because F cannot form 2 dative bonds to oxygen atoms since F is too electronegative for dative bonding.

3. For ClF_5:
   No. of electron pairs = 6
   No. of bond pairs = 5
   No. of lone pairs = 1
   Molecular Geometry = Octahedral

Based on VSEPR Theory;
LP–LP repulsion > LP–BP repulsion > BP–BP repulsion
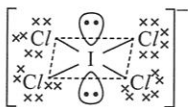∴ Molecular Shape = Square Pyramidal

For ICl_4^-:
No. of electron pairs = 6
No. of bond pairs = 4
No. of lone pairs = 2
Molecular Geometry = Octahedral

Based on VSEPR Theory;
LP–LP repulsion > LP–BP repulsion > BP–BP repulsion
∴ Molecular Shape = Square Planar

PAPER 1

Section A

1. **D**

   Using the Ideal Gas Equation:

   $PV = nRT$ whereby: P in Pa

   V in m^3

   T in K

   R = 8.31 J K^{-1} mol^{-1}

   $\therefore P = \dfrac{nRT}{V}$

   $= \dfrac{(1.6 \times 10^{-3})(8.31)(273 + 273)}{(3.0 \times 10^{-3})}$

☠ **Common Mistakes**
- Units of T must be in K.
- Units of V must be in m^3.

2. **A**

   The stronger the intermolecular forces of attraction between the molecules, the greater the deviation from ideality.

   Methane, ethane and nitrogen are non-polar molecules which are held together by the weaker induced dipole – induced dipole attractions.

   Ammonia molecules are held together by the stronger hydrogen bonding.

3. **D**

   For Ideal gas, $PV = nRT$

   $\therefore \dfrac{P}{T} = \dfrac{nR}{V}$ = constant; since same mass & same volume

   $\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2}$

   $T_2 = P_2 \times \dfrac{T_1}{P_1}$

   $= 2p \times \dfrac{(27 + 273)}{p}$

   $= 600$ K

☠ **Common Mistakes**

Temperature must be in Kelvin (K) when using the Ideal Gas Equation. If it is used as °C, you will end up with a different answer.

i.e. $T_2 = 2p \times \dfrac{27°C}{p}$

$= 54°C$ (which is equivalent to 327K)

$T/K = T/°C + 273$

PAPER 2

1. (a) $n = \dfrac{PV}{RT}$

$$= \dfrac{(6 \times 100 \times 10^3) \times (670 \times 10^{-6})}{8.31 \times (20 + 273)}$$

$$= 0.165 \text{ mol}$$

(b) $P = \dfrac{nRT}{V}$

$$= \dfrac{0.165 \times 8.31 \times (5 + 273)}{670 \times 10^{-6}}$$

$$= 569\,000 \text{ Pa}$$

$$= 5.69 \text{ bar}$$

(c) Maximum safe pressure difference between internal and external pressure $= 800 - 101$

$$= 699 \text{ kPa}$$

At 10,000 m, the pressure difference across the wall of the tyre in flight

$$= (5.69 \times 10^2 - 0.28 \times 10^2) \text{ kPa}$$

$$= 541 \text{ kPa}$$

Therefore, it is not necessary to reduce the air pressure inside the tyre since difference between the internal and external pressure (541 kPa) at 10,000 m is well within the safe pressure range (699 kPa) that the tyre can withstand before it burst.

PAPER 3

1. For an ideal gas:
   ① Gas particles have negligible volume (or size)
   ② Gas particles have negligible intermolecular forces of attraction
   ③ Collisions of gas particles are perfectly elastic.

2. (a) Total mass of gas mixture
   $$= 1.00 - 0.438$$
   $$= 0.562 \text{ g}$$

   For an Ideal Gas; $PV = nRT$

   $$PV = \dfrac{m}{M_r}RT$$

   $$\therefore M_r = \dfrac{mRT}{PV}$$

   $$= \dfrac{(0.562)(8.31)(298)}{(1.00 \times 10^5)(386 \times 10^{-6})}$$

   $$= 36.055$$

   $$= 36.1 \text{ (3 s.f.)}$$

   (b) Carbon dioxide

(c) M_r of both gases

$$= \dfrac{M_r \text{ of } CO_2 + M_r \text{ of other gas}}{2}$$

$\therefore$ M_r of other gas
$$= (2 \times M_r \text{ of both gas}) - M_r \text{ of } CO_2$$
$$= (2)(36.1) - 44$$
$$= 28.2$$
$$\approx 28.0$$

Carbon Monoxide.

(d) E is CaC_2O_4.
$$CaC_2O_4 \rightarrow CaO + CO_2 + CO$$