MEGA LECTURE

IB DP CHEMISTRY

Structure 2: Models of Bonding and Structure

Structure 2.4 From Models to Materials

Revision Notes · Standard and Higher Level

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Original notes prepared for the IB Diploma Programme Chemistry course (first assessment 2025)

What the syllabus requires

Structure 2.4 pulls together the three bonding models you have met (ionic 2.1, covalent 2.2, metallic 2.3) and shows that real substances rarely sit neatly in one box. Bonding is a **continuum**, and where a substance falls on that continuum decides its structure, and therefore its properties and its uses. Use this checklist before the exam.

Understanding	You should be able to...
The bonding continuum	Describe bonding as a spectrum from ionic to polar covalent to pure covalent, and recognise metallic bonding as a third limiting case.
Electronegativity and bond type	Use the difference in electronegativity (ΔEN) and the average electronegativity to predict and justify the dominant bond type.
The bonding triangle	Describe the van Arkel−Ketelaar triangle and calculate the coordinates that place a binary compound on it.
Alloys	Explain, using the metallic model, why alloys are usually harder and stronger than the pure metals they are made from.
Polymers	Draw the repeating unit of an addition polymer from its monomer (and the reverse); relate polymer structure to properties and uses.
Structure from properties (HL depth)	Deduce the structure type of an unknown substance from experimental data on melting point, conductivity and solubility.

> **Exam note:** The core ideas of 2.4 are common to SL and HL. Items flagged (HL) below go further into intermolecular forces and structure−property reasoning. Expect 2.4 to be combined with 2.1−2.3 and with Structure 3.2 (functional groups).

1. Bonding as a continuum: the bonding triangle

The three bonding models are idealisations. In an ideal ionic bond one atom transfers electrons completely to another; in an ideal covalent bond a pair of electrons is shared perfectly equally; in metallic bonding, positive ions sit in a shared sea of delocalised electrons. Most real substances lie **between** these extremes, so it is more honest to speak of the **ionic, covalent and metallic character** of a bond than to force a yes/no label.

The two controlling quantities

Electronegativity (EN) is the ability of an atom to attract the shared electrons in a bond. Two numbers built from EN place any binary substance on the continuum:

- **Electronegativity difference, $\Delta EN = |EN_A - EN_B|$.** A large difference means the electrons are pulled strongly toward one atom → more ionic character.
- **Average electronegativity, $EN_{avg} = \frac{1}{2}(EN_A + EN_B)$.** A high average means both atoms hold electrons tightly (non-metals) → covalent; a low average means both let electrons roam (metals) → metallic.

The van Arkel−Ketelaar triangle

Plotting ΔEN (vertical axis) against EN_{avg} (horizontal axis) for many binary substances produces a triangular map with the three ideal bonding types at the corners:

Corner	Location on the triangle	Example	EN_{avg}, ΔEN
Metallic	Bottom-left: low average EN, small ΔEN	Na, Cu, Mg	low, ≈ 0
Covalent	Bottom-right: high average EN, small ΔEN	Cl_2, F_2, diamond	high, ≈ 0
Ionic	Top: large ΔEN	CsF, NaCl, MgO	middling, large

The bottom edge (small ΔEN) runs from metallic on the left to covalent on the right as EN_{avg} increases: this is the metal−to-non-metal trend across a period. Moving up from the bottom edge (increasing ΔEN) adds ionic character. Polar covalent substances sit in the interior, below the ionic apex and to the right of centre.

> **How to place a compound:** (1) look up both EN values; (2) compute ΔEN and EN_{avg}; (3) read the point off the triangle. Large ΔEN plots high (ionic); small ΔEN plots low, and the average then decides metallic (left) versus covalent (right).

A short list of data-booklet electronegativities (Pauling) for quick placement:

Element	EN	Element	EN	Element	EN
K	0.8	Al	1.6	N	3.0
Na	0.9	Si	1.9	Cl	3.2
Ca	1.0	H	2.2	O	3.4
Mg	1.3	C / S	2.6	F	4.0

Worked coordinate: for magnesium oxide, EN(Mg) = 1.3 and EN(O) = 3.4, so the point sits at EN_{avg} = 2.35 (horizontal) and ΔEN = 2.1 (vertical) − high on the triangle, near the ionic apex. The same two-number recipe places any binary substance.

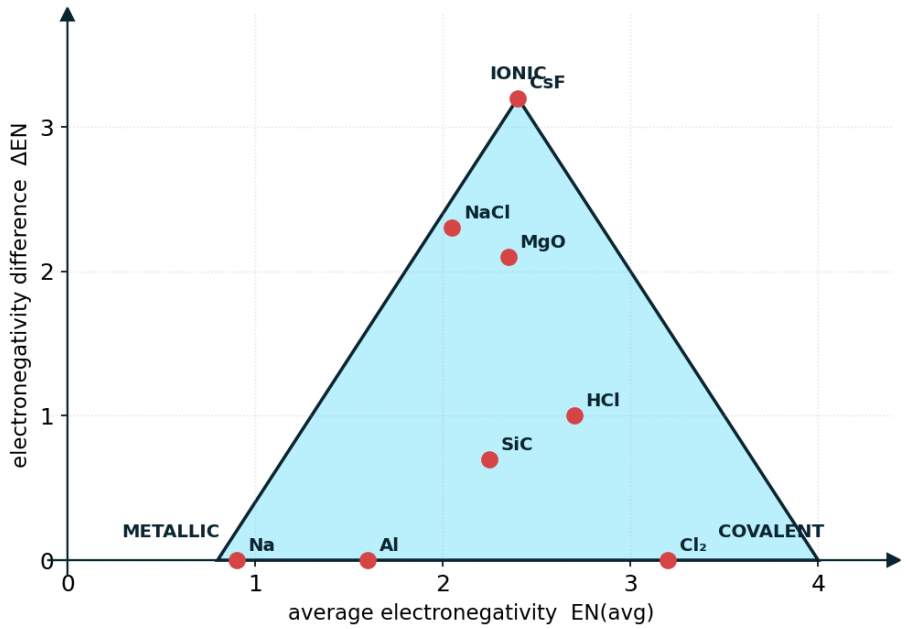

Figure 1. The van Arkel−Ketelaar bonding triangle: ΔEN (vertical) plotted against average electronegativity (horizontal). A large ΔEN plots near the ionic apex; along the base a low average is metallic (left) and a high average is covalent (right).

2. Predicting bond type from ΔEN

As a first pass, ΔEN alone gives a rough guide to the character of a bond between two *different* elements. Examiners accept these bands as approximate, not absolute:

ΔEN (Pauling)	Dominant character	Typical example
0	Pure (non-polar) covalent	Cl−Cl in Cl_2 (ΔEN = 0)
0 − about 0.4	Essentially non-polar covalent	C−H (ΔEN ≈ 0.4)
about 0.4 − 1.8	Polar covalent	H−Cl (ΔEN = 1.0)
greater than about 1.8	Predominantly ionic	Na−Cl (ΔEN = 2.3)

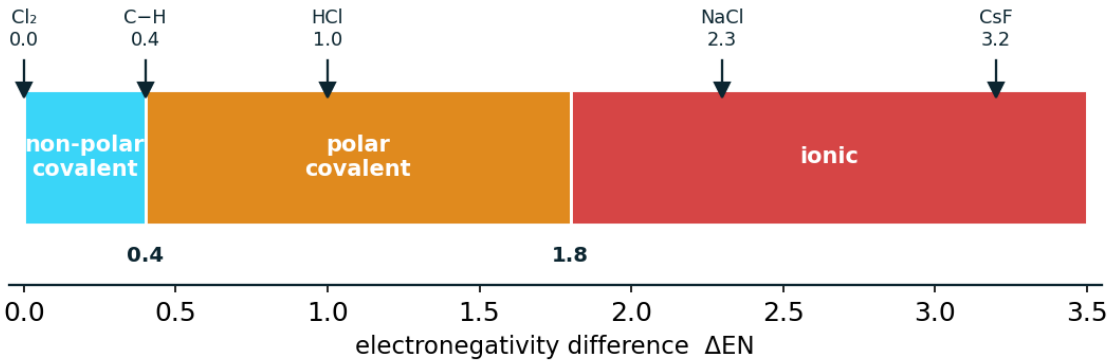

Figure 2. The bonding continuum keyed to ΔEN: pure (non-polar) covalent below ≈ 0.4, polar covalent from ≈ 0.4 to ≈ 1.8, and predominantly ionic above ≈ 1.8. Markers show the computed ΔEN of Cl_2 (0), C−H (0.4), HCl (1.0), NaCl (2.3) and CsF (3.2); the cut-offs are approximate.

The chemical logic behind the bands

- **Two non-metals, similar EN** → small ΔEN → electrons shared → covalent (polar if the atoms differ).
- **A metal with a non-metal** → large ΔEN → electrons effectively transferred → ionic.
- **Two metals (or a metal on its own)** → low average EN → electrons delocalised → metallic.

A polar covalent bond is the true intermediate: the shared pair is displaced toward the more electronegative atom, giving partial charges δ^+ and δ^- but no full transfer. Water's O−H bonds (ΔEN = 1.2) are a classic example: strongly polar, yet unmistakably covalent.

> **Limitation − cutoffs are not rigid:** The 1.8 boundary is a convenience, not a law. HF has ΔEN = 1.8 but is a covalent molecular gas; $AlCl_3$ (ΔEN = 1.5) is covalent enough to sublime. Always support a ΔEN estimate with the average EN and with physical evidence such as melting point or conductivity.

3. Alloys: engineering the metallic model

An **alloy** is a mixture (not a compound) of a metal with one or more other elements, usually other metals. The metallic bonding model carries straight over: a lattice of cations in a sea of delocalised electrons. Because the mixture is not a fixed compound, alloys keep the general metallic properties − electrical and thermal conductivity, lustre − but their mechanical properties can be tuned far beyond those of the parent metals.

Why alloys are usually harder and stronger

In a pure metal the identical cations form regular layers that can slip over one another when a force is applied – this is why pure metals are relatively soft and malleable. Adding atoms of a **different size** distorts the regular layers, so the layers can no longer slide easily past each other. The delocalised electrons still bind the whole structure, so the alloy remains a good conductor while becoming harder, stronger and often more resistant to corrosion.

Alloy	Main components	Property gained	Typical use
Bronze	Cu + Sn	Harder than copper, corrosion-resistant	Bearings, statues, coins
Brass	Cu + Zn	Hard, machinable, decorative	Fittings, instruments
Steel	Fe + C	Much harder and stronger than iron	Construction, tools
Stainless steel	Fe + Cr + Ni	Corrosion resistance	Cutlery, surgical tools
Duralumin	Al + Cu + Mg	Strong yet low density	Aircraft frames

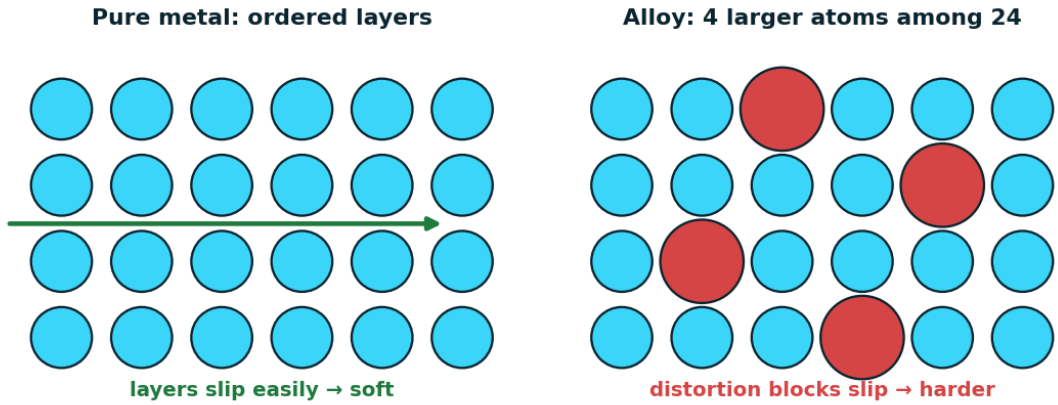

Figure 3. The metallic model of an alloy. Identical cations in a pure metal form regular layers that slip easily; introducing 4 larger atoms into the 24-site lattice distorts the layers so they can no longer slide, while the delocalised electron sea keeps the alloy conducting.

> **Exam phrasing:** To score the mark, say that *differently sized atoms disrupt the orderly layers of cations so the layers cannot slide, while the sea of delocalised electrons is retained*. Naming this mechanism is worth more than simply asserting the alloy is stronger.

4. Polymers: many small molecules, one giant chain

A **polymer** is a very large molecule built by joining many small repeating units called **monomers**. Polymers are the workhorse materials of modern life (plastics, fibres, rubbers), and their properties follow directly from the covalent chains and the intermolecular forces between them.

Addition polymerisation

In **addition polymerisation**, monomers containing a C=C double bond add to one another with **no other product**. Conceptually, the π component of each double bond opens up; the freed electrons form new single bonds to the neighbouring monomers, linking hundreds or thousands of units into one chain. Every atom of the monomer ends up in the polymer, so the empirical formula of the repeating unit matches that of the monomer.

The general scheme for an alkene monomer $CH_2{=}CHX$ is:

$$n\ CH_2{=}CHX \rightarrow {-}(\ CH_2{-}CHX\ ){-}_n$$

The square-bracket unit, written with a bond passing out of each end and a subscript *n*, is the **repeating unit**. Note how the double bond has become a single bond and the bonds now extend beyond the brackets to show the chain continues.

Monomer	Repeating unit	Polymer	Uses
Ethene, $CH_2{=}CH_2$	$-(\ CH_2{-}CH_2\ )-$	Poly(ethene)	Bags, bottles, film
Propene, $CH_2{=}CH(CH_3)$	$-(\ CH_2{-}CH(CH_3)\ )-$	Poly(propene)	Ropes, containers, crates
Chloroethene, $CH_2{=}CHCl$	$-(\ CH_2{-}CHCl\ )-$	Poly(chloroethene), PVC	Pipes, insulation, flooring
Tetrafluoroethene, $CF_2{=}CF_2$	$-(\ CF_2{-}CF_2\ )-$	PTFE	Non-stick coatings, seals

Monomer ↔ repeating unit: the two skills

- **Monomer → repeating unit:** change the C=C to C−C and draw a bond leaving each of the two former double-bond carbons, then enclose in brackets with subscript *n*.
- **Repeating unit → monomer:** find the two carbons carrying the external bonds, join them with a double bond, and remove the brackets. Everything else stays put.

Properties, uses and the environmental cost

Addition polymers have non-polar or weakly polar chains held together only by London (dispersion) forces, so they are typically flexible, waterproof, electrically insulating and unreactive. That same inertness makes them **non-biodegradable**: the strong C−C backbone is not broken down by micro-organisms, which is why plastic waste persists in the environment.

Thermoplastic versus thermoset (context)

Most addition polymers are **thermoplastics**: their separate chains are held to one another only by intermolecular forces, so gentle heating loosens those forces and lets the material soften and be remoulded. By contrast a **thermoset** has covalent cross-links locking the chains into one rigid network, so it cannot be re-melted. This single structural difference – weak forces between chains versus strong covalent cross-links – decides whether a plastic can be recycled by melting, and is a neat illustration of the model-to-material theme.

> **Context – condensation polymers:** The other main route, **condensation polymerisation**, joins monomers that carry two functional groups (for example a diol and a dicarboxylic acid → a polyester), eliminating a small molecule such as water at each link. You meet these in Structure 2.4/Reactivity 3.4; here just note that they release a by-product whereas addition polymers do not.

5. From structure to property to material choice (HL)

(HL) The reason 2.4 is titled *From Models to Materials* is that a material is chosen for a job because of the properties its bonding and structure give it. The chain of reasoning is always the same: **bonding → structure → intermolecular forces → bulk properties → use.**

Intermolecular forces set the properties of molecular materials

For molecular substances (including polymers), the strong covalent bonds hold each molecule together, but it is the weaker **intermolecular forces** between molecules that decide melting point, flexibility and solubility. Stronger intermolecular forces raise melting and boiling points and stiffen a material:

Intermolecular force	When it operates	Effect on the material
London (dispersion)	Between all molecules; grows with molar mass and chain length	Longer polymer chains → higher melting point, greater strength
Dipole−dipole	Between polar molecules (e.g. C=O, C−Cl groups)	Raises melting point and rigidity versus a non-polar chain of similar mass
Hydrogen bonding	N−H, O−H or F−H groups	Strong, directional; gives fibres such as nylon high strength and water affinity

(HL) Structure−property reasoning in practice

This lets a chemist explain, or predict, material behaviour. For example, high-density poly(ethene) has straight chains that pack closely, maximising London forces, so it is rigid and used for bottles; low-density poly(ethene) has branched chains that pack poorly, weakening the forces, so it is soft and used for film. Same monomer, different structure, different material – the essence of materials design.

(HL) Two everyday examples make the point. **Nylon** for ropes and clothing has N−H and C=O groups that hydrogen-bond between chains; the strong, directional forces give high tensile strength and let the fibre absorb some water. **PVC** for pipes has polar C−Cl groups giving dipole−dipole attraction plus London forces; it is rigid, weatherproof and electrically insulating. The same backbone chemistry, dressed with different side groups, yields materials matched to very different jobs – which is exactly what 'from models to materials' means.

> **Exam link:** Whenever you are asked to justify a use, work down the chain: identify the structure type, name the force holding the units together, state the property it produces, then connect that property to the application.

6. Deducing structure from properties

A frequent exam task gives you data on an unknown substance and asks you to classify its structure. Four physical clues – melting point, electrical conductivity (solid and molten/aqueous), and solubility – are usually enough. Work through them systematically.

Structure type	Melting point	Conducts?	Solubility	Example
Ionic lattice	High	Only when molten or dissolved (mobile ions)	Often soluble in water	NaCl, MgO
Molecular covalent	Low	Never (no free charges)	Often soluble in non-polar solvents	I_2, CO_2, sucrose

Structure type	Melting point	Conducts?	Solubility	Example
Covalent network	Very high	Usually not (graphite is the exception)	Insoluble	Diamond, SiO_2
Metallic	Moderate to high	Yes, when solid and molten	Insoluble (reacts with acids)	Cu, Fe, alloys

A decision approach

- **Does the solid conduct electricity?** If yes → metallic (or graphite). If no, go on.
- **Does it conduct only when molten or in solution?** If yes → ionic lattice (mobile ions once the lattice is broken).
- **Is the melting point very high and does it never conduct?** → covalent network.
- **Is the melting point low and does it never conduct?** → molecular covalent.

> **Watch the exceptions:** Graphite is a covalent network that conducts (delocalised electrons between layers). Some ionic solids are only slightly soluble. Never rely on a single property – combine at least two before committing to an answer.

Bringing the models together

The whole of Structure 2 can be read as one connected story: the bonding model fixes the structure, the structure fixes the properties, and the properties decide what the material is good for. The table below traces that story across the four structure types – it is worth committing to memory as a synthesis of Sections 2.1 to 2.4.

Bond / model	Structure	Key properties	Material use
Ionic (large ΔEN)	Giant ionic lattice	Hard, brittle, high m.p., conducts molten	Salts, refractory oxides
Polar/pure covalent (small ΔEN, high avg)	Small molecules or polymers	Low m.p., insulating, often soluble in like solvents	Plastics, solvents, gases
Covalent (small ΔEN, high avg)	Covalent network	Very hard, very high m.p., insulating	Abrasives, semiconductors
Metallic (low avg EN)	Cations in electron sea; alloys	Conducting, malleable, tunable by alloying	Structural metals, wiring

> **Big-picture mark:** If a question asks you to *relate bonding to use*, quote this chain explicitly – model → structure → property → application – and you will hit every marking point.

7. Worked examples

EN values used below are the IB data booklet (Pauling) values: Na 0.9, Mg 1.3, Al 1.6, Si 1.9, H 2.2, C 2.6, S 2.6, N 3.0, Cl 3.2, O 3.4, F 4.0.

Worked example 1 — placing NaCl on the triangle

Use ΔEN and EN_{avg} to place sodium chloride and state its bonding type.

ΔEN = |3.2 − 0.9| = **2.3**. EN_{avg} = ½(3.2 + 0.9) = **2.05**.

The large ΔEN (well above 1.8) plots the point high on the triangle, near the ionic apex. Conclusion: **predominantly ionic** — consistent with NaCl being a high-melting solid that conducts only when molten or dissolved.

Worked example 2 — polar covalent, not ionic

Classify the bonding in hydrogen chloride, HCl.

ΔEN = |3.2 − 2.2| = **1.0**; EN_{avg} = ½(3.2 + 2.2) = **2.7**.

ΔEN of 1.0 falls in the polar-covalent band, and the high average EN (both are non-metals) places the point in the lower-right interior of the triangle. HCl is a **polar covalent** molecule with partial charges $H\delta^{+}$ and $Cl\delta^{-}$. It is a gas that does *not* conduct — confirming it is molecular, not ionic, despite the sizeable ΔEN.

Worked example 3 — a borderline oxide

Aluminium oxide, Al_2O_3, is often called ionic. Test this with EN.

ΔEN = |3.4 − 1.6| = **1.8**; EN_{avg} = ½(3.4 + 1.6) = **2.5**.

A ΔEN of 1.8 sits right on the ionic/covalent borderline, so the point plots in the upper-interior of the triangle — ionic with appreciable covalent character. This is why the simple ΔEN > 1.8 rule must be used with care: the compound has substantial polarisation of the oxide ions and is far less 'perfectly ionic' than, say, NaCl.

Worked example 4 — monomer to repeating unit

Propene is $CH_2{=}CH(CH_3)$. Draw the repeating unit of poly(propene).

Open the double bond to a single bond and extend a bond from each former double-bond carbon out of the brackets:

$$\mathbf{—(\,CH_2—CH(CH_3)\,)—_{\mathit{n}}}$$

The methyl (CH_3) branch is unchanged; only the C=C is consumed. The equation is *n* $CH_2{=}CH(CH_3)$ → the unit above.

Worked example 5 — repeating unit back to monomer

A polymer has the repeating unit —(CH_2—CHCl)—. Identify the monomer.

Locate the two carbons with the external bonds (the CH_2 and the CHCl), join them with a double bond and drop the brackets:

Monomer = **CH_2=CHCl** (chloroethene). The polymer is PVC.

Worked example 6 — classify an unknown from data

Substance Z: melting point 801 °C; does not conduct as a solid; conducts when molten; dissolves in water to give a conducting solution. Deduce its structure.

High melting point rules out molecular covalent. Conducting only when molten or dissolved (not as a solid) is the signature of an **ionic lattice**: ions are fixed in the solid but become mobile once the lattice is broken. The water-soluble conducting solution confirms it. (Z is in fact NaCl.)

Worked example 7 — why average EN matters

Silicon carbide, SiC, has EN(Si) = 1.9 and EN(C) = 2.6. Predict its bonding and structure.

$\Delta EN = |2.6 - 1.9| =$ **0.7** (small); $EN_{avg} = \frac{1}{2}(2.6 + 1.9) =$ **2.25** (fairly high).

Small ΔEN plots the point low on the triangle, and the high average EN places it toward the covalent (right) corner rather than the metallic (left) corner. SiC is therefore **covalent** — in fact a covalent network solid, extremely hard and very high-melting, used as an abrasive. Without the average EN, the small ΔEN alone could not distinguish covalent from metallic.

8. Common pitfalls

- Treating the ΔEN cutoffs as exact — they are approximate bands; always back them with average EN and real properties.
- Calling a strongly polar molecule 'ionic'. A large ΔEN raises ionic *character* but polar covalent ≈/= ionic (HCl, HF are molecular).
- Forgetting the average EN. Two substances can share a ΔEN yet differ (a low average points to metallic, a high average to covalent).
- Drawing a polymer repeating unit with the double bond still present, or leaving off the external bonds and the subscript *n*.
- Saying an alloy is a compound. It is a mixture, so it has no fixed formula and keeps metallic bonding.
- Classifying a structure from one property alone — graphite conducts yet is a covalent network; combine at least two clues.

9. Quick reference

Idea	Statement
ΔEN	$\Delta EN = \lvert EN_A - EN_B \rvert$; large ΔEN → ionic character
Average EN	$EN_{avg} = \frac{1}{2}(EN_A + EN_B)$; low → metallic, high → covalent
Triangle corners	metallic (bottom-left), covalent (bottom-right), ionic (top)
Rough bands	ΔEN ≈ 0 non-polar; 0.4−1.8 polar covalent; >1.8 ionic
Alloys	differently sized atoms disrupt slipping layers → harder; electrons stay delocalised
Addition polymer	$n\ CH_2{=}CHX \rightarrow {-}(\ CH_2{-}CHX\ ){-}_n$; no by-product
Classify structure	conducts solid → metallic; conducts molten only → ionic; very high m.p. + no conduction → network; low m.p. + no conduction → molecular

10. Test yourself

Attempt these without notes; full answers follow. EN values as listed in Section 7.

1. Calculate ΔEN and EN_{avg} for MgO and state where it plots on the bonding triangle.

2. Cl_2 and NaCl have very different bonding. Using EN, explain the difference and name each bond type.

3. Explain, in terms of structure, why brass (Cu + Zn) is harder than pure copper but still conducts electricity.

4. Draw the repeating unit of poly(ethene) and write the equation for its formation from ethene.

5. A polymer has repeating unit $-(\ CF_2{-}CF_2\ )-$. Identify the monomer and name one use of the polymer.

6. An unknown solid has a very high melting point, does not dissolve in water and does not conduct in any state. Classify its structure and give an example.

7. (HL) Explain why high-density poly(ethene) is more rigid than low-density poly(ethene), although both are made from ethene.

8. Silicon has EN 1.9 and chlorine 3.2. Predict the bond type in $SiCl_4$ and comment on whether the simple ΔEN rule is reliable here.

Answers

1. $\Delta EN = \lvert 3.4 - 1.3 \rvert = 2.1$; $EN_{avg} = \frac{1}{2}(3.4 + 1.3) = 2.35$. Large ΔEN plots the point high on the triangle, near the ionic apex − MgO is predominantly ionic (a very high-melting lattice).

2. Cl_2: two identical atoms, ΔEN = 0 → pure (non-polar) covalent, electrons shared equally. NaCl: ΔEN = 2.3 → ionic, electrons effectively transferred from Na to Cl. High ΔEN moves the point up the triangle from covalent to ionic.

3. Zinc atoms differ in size from copper atoms, so they distort the regular layers of cations and stop them sliding, making brass harder. The delocalised sea of electrons is retained, so brass still conducts.

4. Repeating unit $-(\ CH_2{-}CH_2\ )-_n$. Equation: $n\ CH_2{=}CH_2 \rightarrow -(\ CH_2{-}CH_2\ )-_n$.

5. Monomer $CF_2{=}CF_2$ (tetrafluoroethene); the polymer is PTFE, used for non-stick coatings or as an inert seal/gasket.

6. Very high melting point, insoluble and non-conducting in every state → covalent network (e.g. diamond or silicon dioxide, SiO_2).

7. (HL) HDPE has unbranched chains that pack closely together, maximising the London (dispersion) forces between chains, so more energy is needed to move them and the material is rigid. LDPE has branched chains that cannot pack tightly, so the intermolecular forces are weaker and the material is softer.

8. $\Delta EN = |3.2 - 1.9| = 1.3$ → the rule predicts polar covalent, which is correct: $SiCl_4$ is a covalent molecular liquid. Here the simple rule works, but with a high average EN (2.55, both fairly electronegative) it is best confirmed by the low boiling point and lack of conductivity – properties that rule out an ionic lattice.