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Topic 3.2: Properties of Solids

Revision Notes · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

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Four Kinds of Solid, One Deciding Question

Every solid is classified by two questions: what are the particles, and what holds them together? Ions held by electrostatic attraction give an ionic solid. Metal cations in a sea of mobile electrons give a metallic solid. Atoms joined into one continuous covalent structure give a covalent network solid. Whole molecules held by intermolecular forces give a molecular solid.

Almost every property follows. Melting defeats whatever holds the solid together, so melting points rank the holding forces. Conductivity asks whether any charged particle can move. Brittleness and malleability ask what happens when the arrangement is forced to shift.

Ionic Solids: Locked Charges in a Lattice

An ionic solid is a three dimensional lattice of alternating cations and anions, every ion attracting its oppositely charged neighbours. Coulomb's law sets the strength: attraction grows with the charge product and falls as ions grow larger and sit farther apart.

Melting points are high: melting must loosen a lattice held by strong attractions in every direction. The solid never conducts: its charged ions are locked in place. Melt it or dissolve it, and the ions become mobile and the liquid conducts.

Ionic solids are hard but brittle. Slide one layer by a single ion spacing and identical charges come face to face; repulsion along that plane splits the crystal instead of letting it deform.

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Covalent Network Solids: One Continuous Bonded Structure

In a network solid there are no individual molecules. Every atom is covalently bonded into one continuous structure: carbon in diamond bonded tetrahedrally to four others, silicon and oxygen alternating through quartz.

Melting such a solid means breaking covalent bonds themselves, so melting points are extreme (quartz, 1713 °C) and they rank among the hardest materials known: SiC faces cutting tools. Electrons locked in bonds, and no ions: no conduction in any state, and no dissolving.

Graphite is the built in exception. Its carbons form flat layers, strongly bonded within each sheet, delocalized electrons free along them, and only weak dispersion forces between them. So graphite conducts along its layers, and slides layer over layer as a lubricant and a pencil lead.

Molecular Solids: Whole Molecules, Held Lightly

A molecular solid is a lattice of intact molecules: ice, dry ice, iodine, sucrose. The covalent bonds inside each molecule are strong, but melting never touches them; it overcomes only the forces between molecules, Topic 3.1's forces.

Melting points are low: sucrose at 186 °C, iodine at 114 °C, dry ice already subliming at -78 °C. And no state conducts: neutral molecules in solid, melt and solution alike, with no charged particle to carry a current.

Ice carries one extra lesson. Freezing locks water into an open hexagonal network, hydrogen bonds at full length, so ice takes more volume than the liquid: 0.917 g/cm^3 against 1.00 g/cm^3. Most substances pack tighter as solids; hydrogen bonding makes water the famous reverse.

Metallic Solids and Their Alloys

A metallic solid is a lattice of cations held together by a sea of delocalized electrons. The electrons explain conductivity twice over: they carry electric current and, moving fast, transfer thermal energy far quicker than vibrating atoms could.

The same structure explains malleability. Metallic attraction is not directional, and when cation layers slide, the electron sea shifts with them, keeping every cation attracted to its new neighbours. The metal deforms instead of shattering.

Alloys are metals with guest atoms built into the lattice, the two patterns of Topic 2.4. Guests matching the host radius, as zinc (134 pm) in copper (128 pm) in brass, take host sites:

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substitutional. Guests far smaller, as carbon in iron in steel, sit in the gaps: interstitial. Either way, guests block layer sliding: alloys end up harder and less malleable than the pure metals.

Crystalline and Amorphous: Order or None

A crystalline solid holds its particles in regular, repeating, long range order. An amorphous solid (glass, many plastics) has order only over short distances.

The difference shows on heating: a crystalline solid melts sharply at one temperature, while an amorphous solid has no single melting point and softens gradually over a range: why glass can be worked and blown.

Worked Examples

Example 1: Pricing two ionic lattices with Coulomb's law. NaCl melts at 801 °C, MgO at 2852 °C. Charge products: (1 × 1) = 1 against (2 × 2) = 4. Ion separations (Mg^{2+} 72 pm, O^{2-} 140 pm, Na^{+} 102 pm, Cl^{-} 181 pm): 212 pm against 283 pm. Pairwise attraction runs as $q_1q_2 \div r^2$, so MgO scores $4 \div 212^2$ against NaCl's $1 \div 283^2$, a ratio of about 7.1. Both effects push the same way: a 2,051 °C gap.

Example 2: The ice expansion, in numbers. Freeze 250 g of water. As a liquid it fills 250 ÷ 1.00 = 250 cm^3. As ice it needs 250 ÷ 0.917 = 273 cm^3, an expansion of about 9.1%. The open network is the cause: it holds molecules at full hydrogen bond length.

Example 3: Which alloy pattern, decided by radii. Zinc's metallic radius is 134 pm against copper's 128 pm, a difference of (134 - 128) ÷ 128 × 100 = 4.7%, small enough for zinc to take copper sites: brass is substitutional. Carbon's radius, about 77 pm, is (126 - 77) ÷ 126 × 100 = 38.9% smaller than iron's 126 pm, far too small to hold an iron site, so carbon takes the gaps: steel is interstitial.

Example 4: Three unknowns from properties alone. Solid P melts at 1538 °C, conducts as a solid, and bends without breaking: metallic, the iron pattern. Solid Q melts at 186 °C and conducts in no state at all: molecular, the sucrose pattern. Solid R melts at 1713 °C, is extremely hard, and conducts in no state: covalent network, the quartz pattern. Each call uses two properties, because melting point alone misleads: ionic MgO out-melts network quartz.

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Traps That Cost Marks in Topic 3.2

Saying a molten ionic compound conducts by electron flow. In a melt or solution, moving ions carry the current. Electron flow is the metallic mechanism; swapping the two reverses the chemistry.

Claiming a high melting point proves a network solid. Melting point measures holding strength, not identity: ionic MgO melts at 2852 °C, above quartz at 1713 °C. Classify by particles and forces, not by melting point.

Treating graphite as a typical network solid. Graphite conducts and is soft enough to write with: its bonding runs in layers, with delocalized electrons and weak forces between sheets. Diamond, the same element in a full network, is the typical case.

Explaining malleability by weak metallic bonding. Iron's bonding survives to 1538 °C. Metals bend because the attraction is not directional and the electron sea moves with the sliding layers, not because the bonding is feeble.

Forgetting that melting a molecular solid breaks no covalent bonds. Sucrose melts at 186 °C with every $C_{12}H_{22}O_{11}$ molecule intact; only the forces between molecules are overcome. The reverse error fails too: molten NaCl is still Na^+ and Cl^-, only mobile.

Calling an alloy a compound. Brass has no fixed formula: a metal lattice with guest atoms, substituted or interstitial, in proportions that vary.

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Key Relationships to Remember

- Classification first: particles and holding force decide every property (ions, atoms, molecules, or cations in an electron sea)
- Ionic solids: melting points set by Coulomb's law (charge product up, ion size down); conduct only when ions can move (molten or aqueous); brittle, because a shifted lattice aligns like charges
- Covalent network solids: melting breaks covalent bonds, so melting points and hardness are extreme (quartz, 1713 °C); no conduction in any state, graphite excepted (layers, delocalized electrons)
- Molecular solids: melting overcomes intermolecular forces only, so melting points are low (sucrose 186 °C, iodine 114 °C) and no state conducts
- Metallic solids: the electron sea carries current and heat, and lets layers slide (malleable, ductile) instead of shattering
- Alloys: substitutional when radii nearly match (brass, 4.7% apart), interstitial when the guest is far smaller (steel, carbon 38.9% smaller than iron); guests block sliding, so alloys are harder than pure metals
- Crystalline solids melt sharply; amorphous solids such as glass soften over a range, with no long range lattice
- Ice is less dense than liquid water (0.917 against 1.00 g/cm^3): freezing locks an open hexagonal hydrogen bonded network

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