Solutions

Each answer shows the steps, then the **result**.

2.1

- it decreases across the period

2.2

- silicon; giant covalent / giant molecular

3.1 B

- Giant metallic structures have delocalised electrons that conduct

3.2 B

- Silicon is a giant covalent structure with strong bonds throughout

3.3

- from Na to Al each atom releases more delocalised electrons (1
- 3); and the metal ions get smaller with higher charge; so the metallic bonding is stronger, needing more energy to melt

3.4

- sulfur exists as larger S_8 molecules with more electrons than P_4; so it has stronger London (van der Waals) forces, giving a higher melting point

3.5

(a)

- All three are metallic lattices of cations in a sea of delocalised electrons; across the period each atom releases more electrons (1, 2, then 3) and the cation charge rises while its radius falls; so the electrostatic attraction between cations and delocalised electrons strengthens and more energy is needed to melt the lattice

(b)

- Silicon is a **giant covalent** (macromolecular) lattice; melting means breaking covalent bonds right through the structure, not merely loosening a lattice; covalent bonds are far stronger than the metallic bonding or the intermolecular forces in the neighbouring elements

(c)

- Sulfur is **simple molecular**, so only weak induced-dipole (London) forces act between the S_8 molecules and little energy is needed; S_8 has more electrons than P_4; a larger, more polarisable electron cloud gives stronger induced dipoles, hence the higher melting point

(d)

- The atomic radius **decreases**; the nuclear charge rises while the added electrons enter the same outer shell; shielding is essentially unchanged, so the outer electrons are pulled in more strongly