Solutions

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

2.1

- sp^2

2.2

- the bonding is delocalised / identical round the ring (not alternating single and double bonds)

3.1 C

- The p orbitals at right angles to the ring overlap sideways to form the delocalised π system

3.2 D

- Three C=C $\times$ 120 $= -360\ \text{kJ mol}^{-1}$

3.3

- each carbon is sp^2 hybridised, forming three σ bonds (to two C and one H) in a planar hexagon; each carbon has one electron in a p orbital perpendicular to the ring; these p orbitals overlap sideways all round to give one delocalised π system above and below the plane

3.4

- benzene is about 152 kJ mol^{-1} more stable than the Kekulé model predicts; this extra stability is due to delocalisation of the π electrons

3.5

(a)

- Each carbon uses three sp^2 hybrid orbitals, in a plane at 120°; these overlap end-on to form σ bonds to two carbons and one hydrogen, giving the planar hexagonal ring; each carbon keeps one electron in an unhybridised p orbital perpendicular to the ring; these six p orbitals overlap sideways all round the ring, forming a π system **delocalised** above and below the plane

(b)

- Kekulé's structure needs alternating single (0.154 nm) and double (0.134 nm) bonds; the measured 0.139 nm is the same for all six and lies **between** the two, which is what a delocalised, evenly shared π system predicts

(c)

- Three isolated C=C bonds would give $3 \times (-120) = -360$ kJ mol^{-1}; benzene releases only −208, i.e. 152 kJ mol^{-1} **less**; so real benzene lies 152 kJ mol^{-1} lower in energy than the Kekulé model — the delocalisation (resonance) energy

(d)

- Addition would destroy the delocalised π system and lose that stabilisation; substitution replaces a hydrogen and leaves the ring intact; so the substitution route is energetically far more favourable