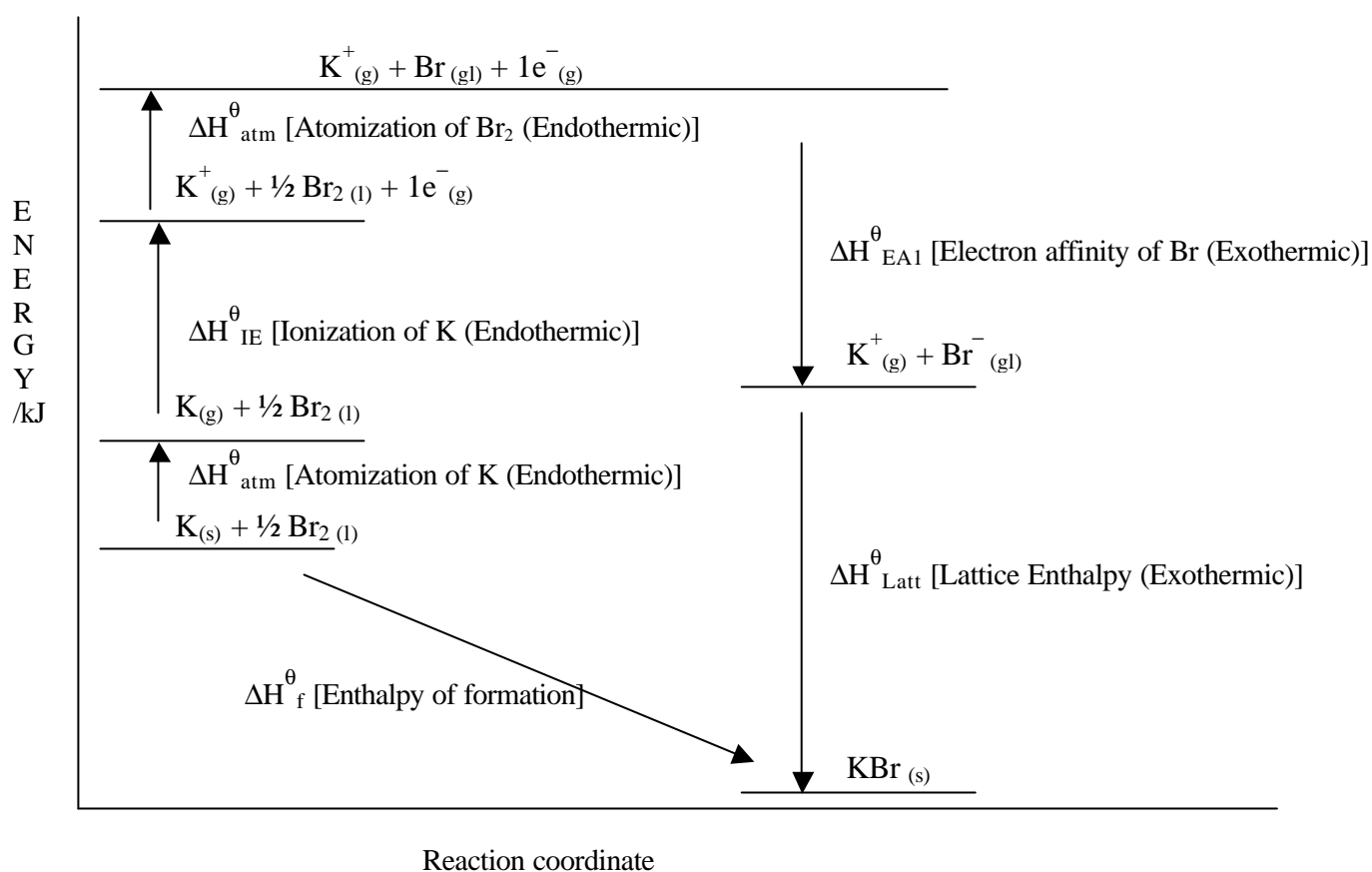


First-Year SL & HL Chemistry

6. Bonding I: Ionic, Covalent and Metallic

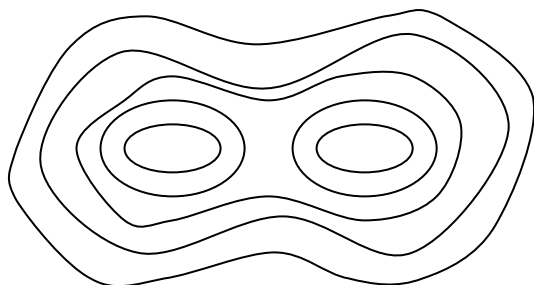
Solutions to Further problems:

- (a) Li^+ (b) Ba^{2+} (c) Mg^{2+} (d) Al^{3+} (e) K^+
- (a) S^{2-} (b) Cl^- (c) N^{3-} (d) I^- (e) O^{2-}
- (a) LiBr lithium bromide (b) MgSe_2 (c) SrCl (d) Al_2S_3 Aluminum sulfide (e) Ca_3P_2 calcium phosphide
- (a) $(\text{NH}_4)_2\text{SO}_4$, (b) $\text{Al}_2(\text{CO}_3)_3$
- (a) CaCO_3 : calcium carbonate (b) NH_4HCO_3 : ammonium hydrogencarbonate (bicarbonate) (c) Na_3PO_4 : sodium phosphate (d) KHSO_4 : potassium hydrogensulphate (e) $\text{Mg}(\text{NO}_3)_2$: magnesium nitrate (f) $\text{Ba}(\text{OH})_2$: barium hydroxide.
- (a) $\text{Al}(\text{NO}_3)_3$ (b) $\text{NH}_4\text{H}_2\text{PO}_4$, (c) $\text{Mg}(\text{OH})_2$, (d) HI
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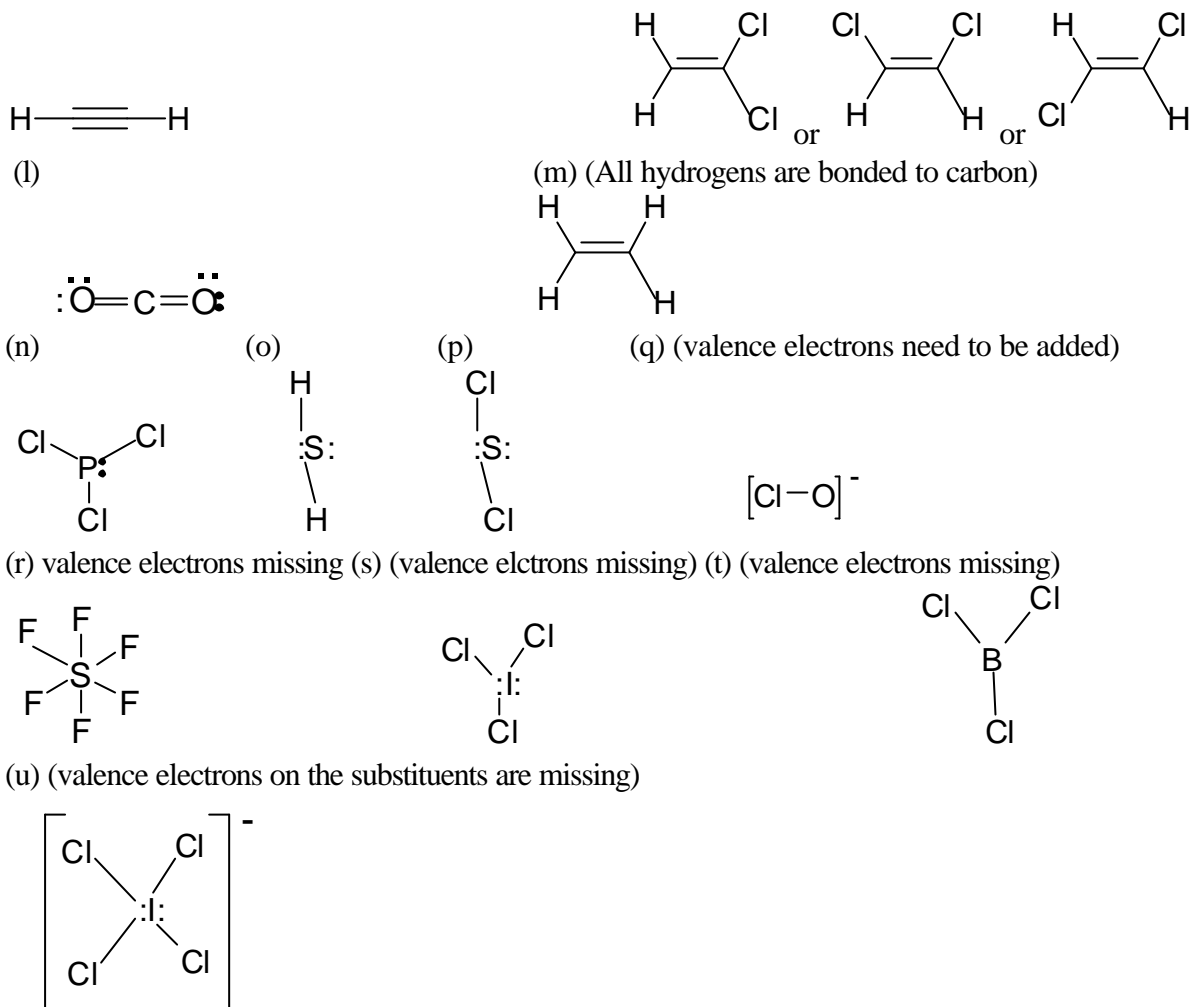


8. Enthalpy of formation of CaCl_2 = Enthalpy of atomization of Ca (enthalpy of vaporization of Ca) + 1st and 2nd ionization of Ca + Bond enthalpy of Cl_2 (or 2 X atomization of Cl_2) + 2 mol X
 Electron affinity of chlorine + Lattice enthalpy
 $(-795 \text{ kJ/mol}) = (192 \text{ kJ/mol}) + (589 \text{ kJ/mol}) + (1146 \text{ kJ/mol}) + (238 \text{ kJ/mol}) + (2)(-348 \text{ kJ/mol})$
 + Lattice enthalpy
 $-795 \text{ kJ/mol} = 192 \text{ kJ/mol} + 589 \text{ kJ/mol} + 1146 \text{ kJ/mol} + 238 \text{ kJ/mol} - 696 \text{ kJ/mol}$
 + Lattice enthalpy
 Lattice enthalpy = $-795 \text{ kJ/mol} - 192 \text{ kJ/mol} - 589 \text{ kJ/mol} - 1146 \text{ kJ/mol} - 238 \text{ kJ/mol}$
 + 696 kJ/mol
 Lattice enthalpy = $-2960 \text{ kJ/mol} + 696 \text{ kJ/mol}$
 Lattice enthalpy = -2264 kJ/mol

9. (a) NaF has a higher lattice enthalpy than NaCl because of the smaller size of the F^- relative to Cl^- . The smaller the ions, the closer they can get and the stronger the electrostatic force of attraction between the ions leading to a higher lattice enthalpy.
- (b) LiBr has a higher lattice enthalpy than CsBr because of the smaller size of Li^+ relative to Cs^+ . The smaller the ions, the smaller the internuclear distance between the oppositely charged ions and therefore stronger the electrostatic force of attraction leading to a higher lattice enthalpy.
- (c) SrS has a higher lattice enthalpy than $SrCl_2$ because of the fact that the lattice of $SrCl_2$ would contain a cation to anion ratio of 1:2 thus increasing the distance between the oppositely charged ions in the lattice. Though the Cl^- is smaller than S^{2-} , the inter-nuclear distance between the ions in SrS is most likely to be smaller than in $SrCl_2$ on account of the stoichiometry and therefore a higher lattice enthalpy.
10. The type of compounds that show the greatest discrepancies between experimental and theoretical values of lattice enthalpy are those that have a small difference in the electronegativities of the elements involved.
- A possible explanation is the nature of the actual bonds. The actual bonds in these compounds are not completely ionic as assumed in the theoretical determination of the lattice enthalpies of those compounds. The bonds in those compounds have significant covalent characters resulting from the polarization of the anion by the cation in the compound, and therefore the greatest discrepancies between their theoretical and the experimental lattice enthalpies.
11. When the atoms have incomplete valence shells and if the formation of the bond leads to energetic stabilization.
- 12.



13. (a) $\begin{array}{c} F \\ | \\ :N-F \\ | \\ F \end{array}$ (b) $\begin{array}{c} H & & H \\ & \backslash & / \\ & N-O \\ & / & \backslash \\ F & & H \end{array}$ (c) $[:N=O]^+$ (d)
- (e) $\left[\begin{array}{c} H \\ | \\ H-P-H \\ | \\ H \end{array} \right]^+$ (f) (the hydrogens are bonded to carbon) (g) $\begin{array}{c} H & H & H \\ | & | & | \\ H-C & -C & -C-H \\ | & | & | \\ H & H & H \end{array}$ (h) $\begin{array}{c} F & & F \\ & \backslash & / \\ & Si \\ & / & \backslash \\ F & & F \end{array}$ $:N \equiv N:$
- (i) $\left[\begin{array}{c} O \\ | \\ O-P-O \\ | \\ O \end{array} \right]^{3-}$ or $\left[\begin{array}{c} O & & O \\ & \backslash & / \\ & P \\ & / & \backslash \\ O & & O \end{array} \right]^{3-}$ or $\left[\begin{array}{c} O & & O \\ & \backslash & / \\ & P=O \\ & / & \backslash \\ O & & O \end{array} \right]^{3-}$ or $\left[\begin{array}{c} O & & O \\ & \backslash & / \\ & P=O \\ & / & \backslash \\ O & & O \end{array} \right]^{3-}$ or $\left[\begin{array}{c} O & & O \\ & \backslash & / \\ & P=O \\ & / & \backslash \\ O & & O \end{array} \right]^{3-}$ or...
- (j) (the hydrogens are bonded to carbon) (k) (the H's and Cl's are bonded to carbon; valence electrons on Cl are missing)



14. (a) the electrons are arranged tetrahedrally; the shape of the molecule is trigonal pyramidal.
 (b) the electrons are arranged tetrahedrally, the shape of the molecule is bent (or V-shaped).
 (c) the electrons are arranged tetrahedrally while the shape of the molecule is trigonal pyramidal.
 (d) the electrons are arranged in a trigonal planar fashion, the shape of the molecule is also trigonal planar.
15. Except for CS_2 and AlCl_3 all of them are polar.
16. CH_2Cl_2 is polar but CCl_4 is not because in CH_2Cl_2 the polarities of the four bonds are different and since the molecule is tetrahedral, a resultant dipole (in the direction of the chlorines) is created as a result. In CCl_4 however, the polarities are equal, the molecule is tetrahedral and so the resultant dipole moment of any three bonds will be equal and opposite to that of the fourth bond.
17. (a) polar covalent, (b) mainly ionic, (c) polar covalent, (d) mainly ionic (e) polar covalent, (f) polar covalent, (g) non-polar covalent, (h) mainly ionic.
18. (a) $[\text{O}=\text{N}-\text{O}]^- \leftrightarrow [\text{O}-\text{N}=\text{O}]^-$ (b) $[\text{N}=\text{N}=\text{N}]^- \leftrightarrow [\text{N}=\text{N}-\text{N}]^- \leftrightarrow [\text{N}-\text{N}=\text{N}]^-$ (The location of the charge is different between the resonance structures).
19. $\text{C}=\text{O}$, not only is it a double bond, it's a bond between two small atoms, two atoms of elements from period 2.
20. The pure (atomic) orbitals that combine to form sp^2 hybrid orbitals are one s and two p orbitals. The geometric arrangement of the hybrid orbitals can be explained by the fact that all three sp^2 orbitals are of equal energy (which means they are of the same size) and therefore arrange themselves so that each one occupies the same position relative to the other two. A trigonal planar arrangement does just that.
21. (a) 109 (b) 120 (c) 180 (d) axial = 180, equatorial = 120, between axial and equatorial bonds = 90.
22. 13 σ and 4 π bonds.