18 - Lattice Energy

Q-1) What is the enthalpy change of formation?

> It's the enthalpy change when 1 mol of a compound is formed from its elements under standard conditions.

* The reactants and products must be in their standard states.

↓ exothermic.

Q-2) What is standard enthalpy change of atomisation?

> It's the enthalpy change when 1 mol of gaseous atoms are formed from its elements under standard conditions.

↑ endothermic

Q-3) What is ionisation energy 1/2?

> It's the energy needed to remove one e^- from each atom / +1 ion in one mole of atoms / +1 ions in the gaseous state to form one mole of +1/+2 ions.

↑ endothermic.

Q-4) What is electron affinity 1/2?

> It's the enthalpy change when 1mol of e^- is added to 1 mol of gaseous atoms/-1 ions to form 1mol of gaseous -1/-2 ions under standard conditions.

EA1 = ↓ exothermic

EA2+ = ↑ endothermic.

Q-5) What is bond energy?

> It's the energy needed to break 1mol of a particular bond in 1 mol of gaseous molecules

↑ endothermic

Q-6) What is lattice energy?

> It's the enthalpy change when 1 mol of an ionic compound is formed from its gaseous ions under standard conditions.

↓ exothermic

Q-7) Relationship between bondenergy and ΔH_{at}^{θ}.

> BE : $Cl_{2(g)} \longrightarrow 2Cl_{(g)}$

ΔH_{at}^{θ} : $\frac{1}{2}Cl_{2(g)} \longrightarrow Cl_{(g)}$.

$\therefore \Delta H_{at}^{\theta} = \frac{1}{2} BE$.

Q-8) Factors affecting lattice energy.

> Size of ions:

- as size increases, ΔH_{latt} becomes less exothermic (less -ve) as there are less forces of attraction.

> Charge on ions:

- as charge increases, ΔH_{latt} becomes more exothermic (more -ve) as there are more forces of attraction.

ΔH_{latt} more exothermic (more -ve) when:

- small size of ion
- high charge on ion.

Q-9) Thermal stability of group 2 carbonates and nitrates.

> $XCO_3 \longrightarrow XO + CO_2$.

$2X(NO_3)_2 \longrightarrow 2XO + 4NO_2 + O_2$

↓ brown fumes

Ion polarisation is when the high charge density cation pulls the e^- cloud of large anion towards itself.

Down the group 2
- cation size increases.
- charge density decreases
- ion polarisation decreases.

$\therefore$ → thermal stability increases
→ decomposition temperature increases.

Q-10) What is standard enthalpy change of solution?

> It's the enthalpy change when 1mol of an ionic solid dissolves completely in a solvent to give an infinitely ~~dilute~~ dilute solution

→ exothermic or endothermic.

Q-11) What is standard enthalpy change of hydration?

> It's the enthalpy change when 1 mol of a gaseous ion dissolves in a solvent to form an infinitely dilute solution.

↓ exothermic.

Q-12) Solubility of group 2 sulphates and hydroxides.

> A compound is soluble in water when:

$\Delta H_{hyd} > \Delta H_{latt}$. → since $\Delta H_{sol} = \Delta H_{hyd} - \Delta H_{latt}$

Down the group 2: SULPHATES:
- ΔH_{hyd} and ΔH_{latt} decreases.
- BUT ΔH_{hyd} decreases more rapidly

$\therefore$ $\Delta H_{hyd} < \Delta H_{latt}$.

$\therefore$ Solubility decreases. (ΔH_{sol} becomes more endothermic)

HYDROXIDES:
- BUT ΔH_{latt} decreases more rapidly

$\therefore$ $\Delta H_{hyd} > \Delta H_{latt}$

$\therefore$ solubility increases (ΔH_{sol} becomes more exothermic)

Q-13) What is entropy?

> Entropy is the measure of the disorder of a system. The lower the entropy, the more ordered a compound.

unit : $kJmol^{-1}K^{-1}$

$$\Delta S^\circ = \Delta S^\circ_{products} - \Delta S^\circ_{reactants}$$

* (s) → (l) → (g) → (aq).

entropy increases →

* As temperature increases, entropy increases.

* As no. of (g) molecules increases, entropy increases.

Q-14) What is standard Gibbs free energy change?

$$\underset{kJmol^{-1}}{\Delta G^\circ} = \underset{kJmol^{-1}}{\Delta H^\circ} - \underset{kJmol^{-1}K^{-1}}{T^{K}\Delta S^\circ}$$

ΔH° = enthalpy change
T = temperature
ΔS° = entropy

* If ΔS is -ve ; reaction is spontaneous.
* If ΔS is +ve ; reaction is non-spontaneous.

if:	$\Delta G^\circ = \Delta H^\circ$	$- T\Delta S^\circ$	then ΔG°:
1	−	+	−ve
2	+	−	+ve
3	+	+	+ve at low temp −ve at high temp
4	−	−	−ve at low temp