

Chapter	6 Chemical Energetics	Class	1 st Year	Subject	Chemistry
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Give Short Answers of the Following Questions

1. Differentiate between exothermic and endothermic reaction:

Ans.

Property	Exothermic Reactions	Endothermic Reactions
Energy Change	Release energy (heat) to surroundings	Absorb energy (heat) from surroundings
Enthalpy (ΔH)	Negative ($\Delta H < 0$)	Positive ($\Delta H > 0$)
Temperature	Increases in the surrounding (heats up)	Decreases in the surrounding (cools down)
Example	Combustion of fuels, respiration, freezing water	Photosynthesis, melting ice, boiling water
Reaction profile	Energy of product < energy of reactants	Energy of product > energy of reactants

2. What do you understand by the enthalpy of a system?

Ans. Enthalpy (H) of a system is the total heat content of the system at constant pressure.

It is defined as: $H = U + PV$

Where:

- H = enthalpy
- U = internal energy of the system
- P = pressure
- V = volume

3. Differentiate clearly between entropy (S) and Gibbs free energy (G).

Ans.

Property	Entropy (S)	Gibbs Free Energy (G)
Definition	Measure of disorder or randomness in a system	Measure of the usable energy available to do work
Symbol	S	G
Units	J. mol ⁻¹ .K ⁻¹	KJ. mol ⁻¹
Nature	Tells how dispersed energy is in a system	Tells whether a process is spontaneous
Formula	No single formula (but used in ΔS)	$\Delta G = \Delta H - T \Delta S$
Interpretation	Higher S = more disorder	Negative ΔG = spontaneous process
Temperature Dependency	Affects spontaneity when multiplied by T in ΔG	Directly determines spontaneity of a reaction

4. Distinguish clearly between standard enthalpy of reaction and standard enthalpy of formation.

Ans.

Property	Standard Enthalpy of Reaction ($\Delta H_{\text{reaction}}^{\circ}$)	Standard Enthalpy of Formation (ΔH_f°)
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Definition	Heat change during a chemical reaction under standard conditions	Heat change when 1 mole of a compound forms from its elements in their standard states
Reactants and Products	Any type of substances (elements or compounds)	Always starts from elements in standard states
Amount Involved	Varies depending on the reaction equation	Always forms exactly 1 mole of product
Symbol	ΔH_{rxn}°	ΔF_f°
Example	Combustion of CH₄: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$	Formation of H₂O: $\text{H}_{2(g)} + \frac{1}{2}\text{O}_{2(g)} \rightarrow \text{H}_2\text{O}_{(l)}$

5. Define the following enthalpies and give one example of each.

(i) Standard enthalpy of solution

(ii) Standard enthalpy of hydration

(iii) Standard enthalpy of atomization

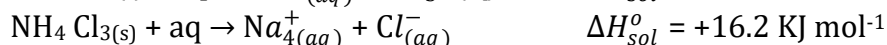
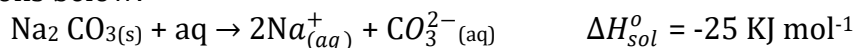
(iv) Standard enthalpy of combustion

Ans.

(i) Standard enthalpy of solution (ΔH_{sol}°)

"The standard enthalpy of solution is the amount of heat absorbed or evolved when one mole of a substance is dissolved in a solvent to give an infinitely dilute solution."

The enthalpy changes of solution of sodium carbonate and ammonium chloride are described by the equations below:



(ii) Standard enthalpy of hydration

"The enthalpy change involved when one mole of a solute is dissolved in excess of water to make infinitely dilute solution under standard conditions."

(iii) Enthalpy Change of Atomization (ΔH_{at}°)

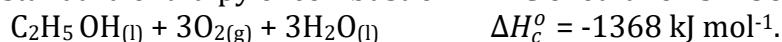
"The standard enthalpy change of atomization of an element is the enthalpy change involved when one mole of gaseous atoms is formed from the element, under standard conditions."



(iv) Enthalpy Change of Combustion (ΔH_c°)

The standard enthalpy change of combustion of a substance is the enthalpy change involved when one mole of the substance is completely burnt in excess of oxygen, under standard conditions.

For example, standard enthalpy of combustion ΔH_c° of ethanol is $-1368 \text{ kJ mol}^{-1}$.



6. Explain why the lattice enthalpy of an ionic compound is typically a large negative value.

Ans. The lattice enthalpy of an ionic compound is typically a large negative value because:

It represents the energy released when gaseous ions combine to form one mole of a solid ionic

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crystal. This process involves strong electrostatic attraction between oppositely charged ions. Since a large amount of energy is released during this bond formation, the lattice enthalpy is highly exothermic (negative value).

7. What factors influence the magnitude of the lattice enthalpy?

Ans. The magnitude of lattice enthalpy is mainly influenced by two key factors:

1. Ionic Charge

- Greater ionic charge $\Rightarrow$ Stronger electrostatic attraction between ions.
- This leads to a more negative (larger) lattice enthalpy.

Example:

Lattice enthalpy of MgO (Mg^{2+} and O^{2-}) is much higher than that of NaCl (Na^+ and Cl^-) due to higher charges.

2. Ionic Radius (Size of ions)

- Smaller ions $\Rightarrow$ Ions are closer together $\Rightarrow$ Stronger attraction.
- Smaller size = more negative lattice enthalpy.

Example:

Lattice enthalpy of LiF is greater than that of CsF, because Li^+ is smaller than Cs^+ .

8. Explain why the enthalpy of hydration is always an exothermic process for gaseous ions. What are the main interactions responsible for the release of energy during hydration?

Ans. Enthalpy of Hydration: Always an exothermic process

The enthalpy of hydration is the heat released when gaseous ions dissolve in water and become surrounded by water molecules. It is water exothermic because:

When gaseous ions dissolve, the water molecules surround and stabilize the ions, leading to a decrease in the system's energy.

Main Interactions Responsible for Energy Release

(i) Ion-Dipole Interactions:

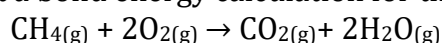
The polar water molecules interact with the charged ions through their partial charges (δ^- on oxygen and δ^+ on hydrogen). This creates a stable interaction between the ion and the water molecules, releasing energy.

(ii) Electrostatic Attraction:

Water molecules are strongly attracted to the ions, especially smaller, highly charged ions (like Mg^{2+} or Na^+), which increases the energy released during hydration.

9. For the reaction $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$, identify all the bonds that need to be broken and all the bonds that need to be formed to carry out a bond energy calculation of ΔH .

Ans. To carry out a bond energy calculation for the reaction:



Bonds that need to be broken (in reactants):

1. CH_4 (Methane):

4 C-H bonds (broken to form C and H atoms)

2. O_2 (Oxygen):

2 O=O bonds (broken to form 2 oxygen atoms)

Bonds that need to be formed (in products):

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- CO₂ (Carbon Dioxide):**
2 C=O bonds (formed between carbon and oxygen)
- H₂O (Water):**
2 O-H bonds in each H₂O molecule (2 molecules of H₂O, so 4 O-H bonds are formed).
To calculate ΔH we would use the bond energies (bond dissociation energies) for each of these bonds and apply the formula:
 $\Delta H = \text{Bonds Broken (Energy required)} - \text{Bonds Formed (Energy released)}$

10. For a reaction to be spontaneous, what is the required sign of the Gibbs free energy change (ΔG)? Under what conditions of enthalpy change (ΔH) and entropy change (ΔS) will a reaction always be spontaneous?

Ans. For a reaction to be **spontaneous**, the required sign of the **Gibbs free energy change (ΔG)** is:

$$\Delta G < 0$$

This means that the reaction must have a **negative Gibbs free energy change** for it to occur spontaneously under constant temperature and pressure.

In other words:

- $\Delta G < 0$ indicates a spontaneous reaction.
- $\Delta G > 0$ indicates a non-spontaneous reaction
- $\Delta G = 0$ indicates a system in equilibrium.

The Gibbs free energy change (ΔG) must be negative, i.e., $\Delta G < 0$.

The relationship between ΔG , ΔH (enthalpy change), and ΔS (entropy change) is given by the equation:

$$\Delta G = \Delta H - T\Delta S$$

Where:

- ΔG = Gibbs free energy change
- ΔH = Enthalpy change
- T = Temperature (in Kelvin)
- ΔS = Entropy change

Conditions for Always Spontaneous Reaction:

- $\Delta H < 0$ (Exothermic reaction)
- $\Delta S > 0$ (Increase in disorder)

Explanation:

- Exothermic ($\Delta H < 0$)** reactions release heat, making the system more stable.
- Increase in entropy ($\Delta S > 0$)** means the system becomes more disordered, which also favours spontaneity. These conditions ensure that ΔG will always be negative, regardless of the temperature. The negative ΔH and positive ΔS make both terms in the equation contribute to a negative value for ΔG .

11. The enthalpy of solution can be either positive or negative. Explain what a positive ΔH_{sol} and a negative ΔH_{sol} indicate about the energy changes during the dissolution process.

Ans. The enthalpy of solution (ΔH_{sol}) represents the heat change when a solute dissolves in a solvent. The sign of (ΔH_{sol}) gives us information about the energy changes that occur during the dissolution

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process.

(i) Positive (ΔH_{sol}) (Endothermic Dissolution):

- (ΔH_{sol}) means the dissolution process is endothermic.
- Energy is absorbed from the surroundings to dissolve the solute.
- This occurs when the energy required to break the interactions between solute particles (like ionic or covalent bonds) is greater than the energy released when solvent particles interact with the solute.
- **Example:** When ammonium nitrate (NH_4NO_3) dissolves in water, it absorbs heat, making solution feel cold.

(ii) Negative (ΔH_{sol}) (Exothermic Dissolution):

- (ΔH_{sol}) means the dissolution process is exothermic.
- Energy is released into the surroundings as the solute dissolves.
- This occurs when the energy released during the interaction of solute particles with solvent molecules exceeds the energy needed to separate the solute particles.
- **Example:** When sodium hydroxide (NaOH) dissolves in water, heat is released, making the solution feel warm

12. Consider two ions with similar charges but different sizes, or similar sizes but different charges. Explain how the concept of charge density can be used to predict which ion will have a more exothermic enthalpy of hydration and why.

Ans. The concept of charge density is essential for understanding which ion will have a more exothermic enthalpy of hydration.

- **Charge Density:**
- Charge density is the charge of an ion divided by its size (radius).
- Charge density = charge/radius.
- **Prediction of Enthalpy of Hydration:**
- Higher charge density means stronger ion-dipole interactions between the ion and water molecules leading to a more exothermic enthalpy of hydration i.e., more energy released during hydration.
- **For ions with similar charges but different sizes:**
- Smaller ions have a higher charge density because their charge is concentrated over a smaller area.
- Smaller ions (higher charge density) will have a more exothermic enthalpy of hydration because they create stronger interactions with water molecules.
- **For ions with similar sizes but different charges:**
- Ions with higher charges have a higher charge density, meaning they can interact more strongly with water molecules.
- Higher charge e.g., $\text{Al}^{3+} > \text{Mg}^{2+} > \text{Na}^+$ density results in a more exothermic enthalpy of hydration.