

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

1. Define electron affinity. Give example.

Ans. Electron Affinity (ΔH_{ea}°)

The first electron affinity is the enthalpy change involved when 1 mole of electrons is added to mole of gaseous atoms to form 1 mole of gaseous uni-negative ions under standard conditions.

Example:

Electron affinity of chlorine atom.



Since, energy is released, so first electron affinity carries negative sign.

2. What is calorie content?

Ans. Calorie content: The calorie content of food is a measure of the energy 'released' when the food is completely consumed in the body.

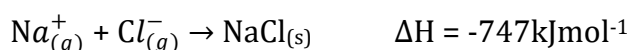
- This energy is typically expressed in units of kilocalories (k cal) or joules (J).

3. What is meant by lattice energy? Give example.

Ans. Lattice Energy (ΔH_{latt})

"The lattice energy of an ionic crystal is the enthalpy change involved when one mole of the ionic compound is formed from gaseous ions under standard conditions."

Example:



4. What do you know about entropy?

Ans. Entropy is a measure of the number of ways energy can be distributed within a system at a specific temperature. (OR)

Entropy can also be thought of as a measure of the randomness or disorder of a system.

- When the energy is distributed in more ways, a system is more stable.
- The higher the randomness or disorder, the greater the entropy of the system.

5. The sign of change in free energy (ΔG) of a process can be used to predict the spontaneity of that process. How?

Ans. The sign of change in free energy (ΔG) of a process can be used to predict the spontaneity of that process, when;

(i) If $\Delta G < 0$ (-ve), the given process may occur spontaneously

(ii) If $\Delta G > 0$ (+ve), the indicated process cannot occur spontaneously; instead, the reverse of it may occur.

(iii) If $\Delta G = 0$, neither the indicated process nor reverse of it can occur spontaneously. The system is in a state of equilibrium

6. Define: (i) Standard Enthalpy of reaction with one example.

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(i) Enthalpy of a system with an example.

Ans. (i) Standard Enthalpy of a Reaction:

"The enthalpy change which occurs when the certain number of moles of reactants react together completely to give the products under standard conditions i.e. 25°C (298K) and one atmosphere pressure is called standard enthalpy of reaction." It is represented by ΔH° . All the reactants and products must be in their standard physical states.

- Its unit is kJ mol^{-1} .

Example:



$-285.8 \text{ kJ mol}^{-1}$ is standard enthalpy of reaction.

(ii) Enthalpy of the system:

"The sum of internal energy (E) and the product of pressure and volume (PV) is called enthalpy of the system,"

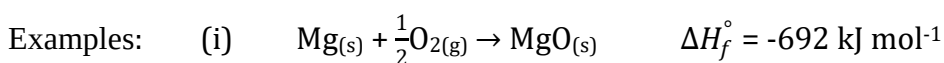
Formula: $H = E + PV$

7. Define enthalpy of formation (ΔH_f°). Give two examples.

Ans. Enthalpy of formation (ΔH_f°):

"The amount of heat evolved or absorbed when one mole of compound is formed from its elements under standard conditions of temperature and pressure is called standard enthalpy of formation (ΔH_f°) of a compound".

- All the substances involved are in their standard physical states.
- The reaction is carried at standard conditions i.e. at 25°C (298K) and one atmospheric pressure.



8. State with one example enthalpy of combustion.

Ans. Enthalpy of Combustion:

"The amount of heat evolved when one mole of a substance is completely burnt in excess of oxygen under standard conditions is called standard enthalpy of combustion of a substance (ΔH_c°)."

- Its unit is kJ mol^{-1} .

Example:

Standard enthalpy of ethanol is $-1368 \text{ kJ mol}^{-1}$.



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Enthalpy of Solution:

"The amount of heat absorbed or evolved when one mole of a substance is dissolved in so much solvent that further dilution results in no detectable heat change is called standard enthalpy of solution (ΔH_{sol}^o)."

- Its unit is kJ mol^{-1} .

Examples: (i) Enthalpy of a solution of $\text{NH}_4\text{Cl} = +16.2 \text{ kJ mol}^{-1}$

(ii) Enthalpy of a solution of $\text{Na}_2\text{CO}_3 = -25.0 \text{ kJ mol}^{-1}$

9. Define enthalpy of neutralization. Why enthalpy of neutralization of strong acid and strong base is always $-57.5 \text{ kJ mol}^{-1}$?

Ans. Enthalpy of Neutralization:

Enthalpy of neutralization is the amount of heat evolved when one mole of hydrogen ions H^+ from an acid react with one mole of hydroxide OH^- ions from a base to form one mole of water.



All strong acids and bases ionize completely in aqueous solutions providing equal number of H^+ and OH^- ions depending upon concentration of solutions.

So a net result of neutralization is just the formation of water from its ionic components. For all strong acids and bases with equimolar concentrations of ions, the value of heat of neutralization is same i.e., $-57.4 \text{ kJ mol}^{-1}$ or $13.2 \text{ Kcal mol}^{-1}$.

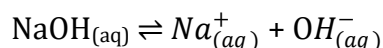
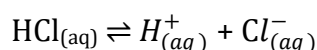
10. Define the terms standard enthalpy of neutralization and standard enthalpy of atomization.

Ans. Enthalpy of Neutralization:

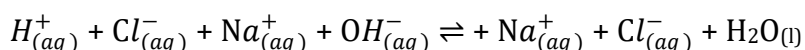
"The amount of heat evolved when one mole of hydrogen ions (H^+) from an acid, react with one mole of hydroxide ions (OH^-) from a base to form one mole of water under standard conditions is called standard enthalpy of neutralization (ΔH_n^o)."

Example:

A strong acid HCl and a strong base NaOH ionize completely in dilute solutions as follows:



When these solutions are mixed together during the process of neutralization, the only change that actually occurs is the formation of water molecules leaving the sodium ions and the chloride ions as free ions in solution. Thus, enthalpy of neutralization is merely the heat of formation of one mole of liquid water from its ionic components.



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Enthalpy of neutralization for any strong acid with a strong base is approximately the same. i.e. $-57.4 \text{ kJ mole}^{-1}$.

Enthalpy of Atomization:

"The amount of heat absorbed when one mole of gaseous atoms is formed from the element under standard conditions is called standard enthalpy of atomization of that element (ΔH_{at}°)."

- Its unit is kJ mol^{-1} .

Example: The standard enthalpy of atomization of hydrogen is



A wide range of experimental techniques are available for determining enthalpies of atomization of elements.

11. Define Hess's law of constant heat summation.

Ans. Hess's law:

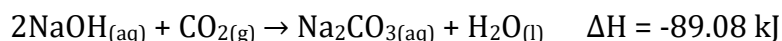
If a chemical change takes place by several different routes, the overall energy change is the same, regardless of the route by which the chemical change occurs, provided the initial and final conditions are the same."

Mathematical Form: $\Sigma \Delta H(\text{cycle}) = 0$

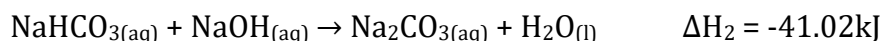
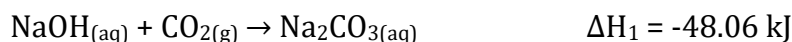
Examples:

Formation of Na_2CO_3 :

Single Step Process:



Two Steps Process:



According to Hess's law $\Delta H = \Delta H_1 + \Delta H_2$

Putting the values of ΔH , ΔH_1 and ΔH_2 in the equation.

$$-89.08 = -48.06 - 41.02$$

$$-89.08 = -89.08 \text{ kJ}$$

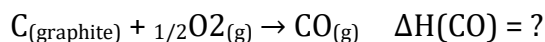
Enthalpy of Formation of CO:

Single Step Process:

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Two Step Process:



According to Hess's law

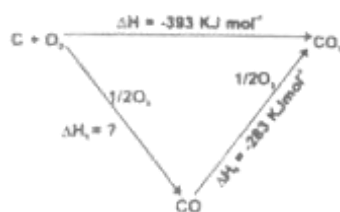
$$\Delta H = \Delta H_1 + \Delta H_2$$

$$\Delta H_1 = \Delta H_2 - \Delta H$$

$$= -393 - (-283)$$

$$= -110 \text{ kJ mol}^{-1}$$

So, the enthalpy change for the formation of $CO_{(g)} = -110.0 \text{ kJ mol}^{-1}$



12. Differentiate between Atomization energy and Lattice energy.

Ans. Atomization Energy:

The energy required to convert a substance in atomic state is called atomization energy. It is always endothermic in nature.



Lattice Energy:

"Enthalpy of formation of one mole of crystalline ionic compound from gaseous ions under standard conditions is called lattice energy."



OR

"Energy required to break one mole of crystalline ionic compound into its gaseous ions at standard conditions is called lattice energy."



13. Define heat capacity of a body. Give its mathematical expression.

Ans. Heat Capacity: "The product of mass and specific heat of water is called heat capacity of the whole system." It is represented by 'c'.

$$C = m \times s$$

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14. 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

15. 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

16. 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

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17. 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°)
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	$\Delta H_{\text{rxn}}^{\circ}$	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: $\text{H}_{2(\text{g})} + 1/2\text{O}_{2(\text{g})} \rightarrow \text{H}_2\text{O}_{(\text{l})}$

18. 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

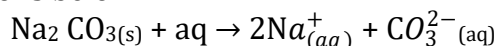
(iv) Standard enthalpy of combustion

Ans.

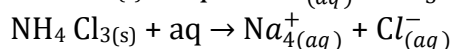
(i) Standard enthalpy of solution ($\Delta H_{\text{sol}}^{\circ}$)

"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:



$$\Delta H_{\text{sol}}^{\circ} = -25 \text{ KJ mol}^{-1}$$



$$\Delta H_{\text{sol}}^{\circ} = +16.2 \text{ KJ mol}^{-1}$$

(ii) Standard enthalpy of hydration

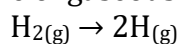
"The enthalpy change involved when one mole of a solute is dissolved in excess of water to make infinitely dilute solution under standard conditions."



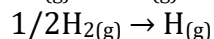
$$\Delta H_{(\text{Hyd})}^{\circ} = -406 \text{ KJ mol}^{-1}$$

(iii) Enthalpy Change of Atomization ($\Delta H_{\text{at}}^{\circ}$)

"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."



$$E_{\text{H-H}} = +436 \text{ kJ mol}^{-1}$$



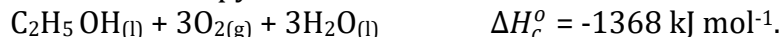
$$\Delta H_{\text{at}}^{\circ} = +218 \text{ kJ mol}^{-1}$$

(iv) Enthalpy Change of Combustion ($\Delta H_{\text{c}}^{\circ}$)

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.

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



19. 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 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).

20. 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^+ .

21. 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:

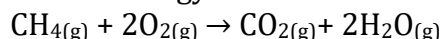
The polar water molecules interact with the charged ions through their partial charges (δ^- on oxygen and δ^+ on hydrogen). This creates a stable interaction between the ion and the water molecules, releasing energy.

(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.

22. For the reaction $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$, 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:



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Bonds that need to be broken (in reactants):

- CH₄ (Methane):**
4 C-H bonds (broken to form C and H atoms)
- O₂ (Oxygen):**
2 O=O bonds (broken to form 2 oxygen atoms)

Bonds that need to be formed (in products):

- 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 .

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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 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.



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