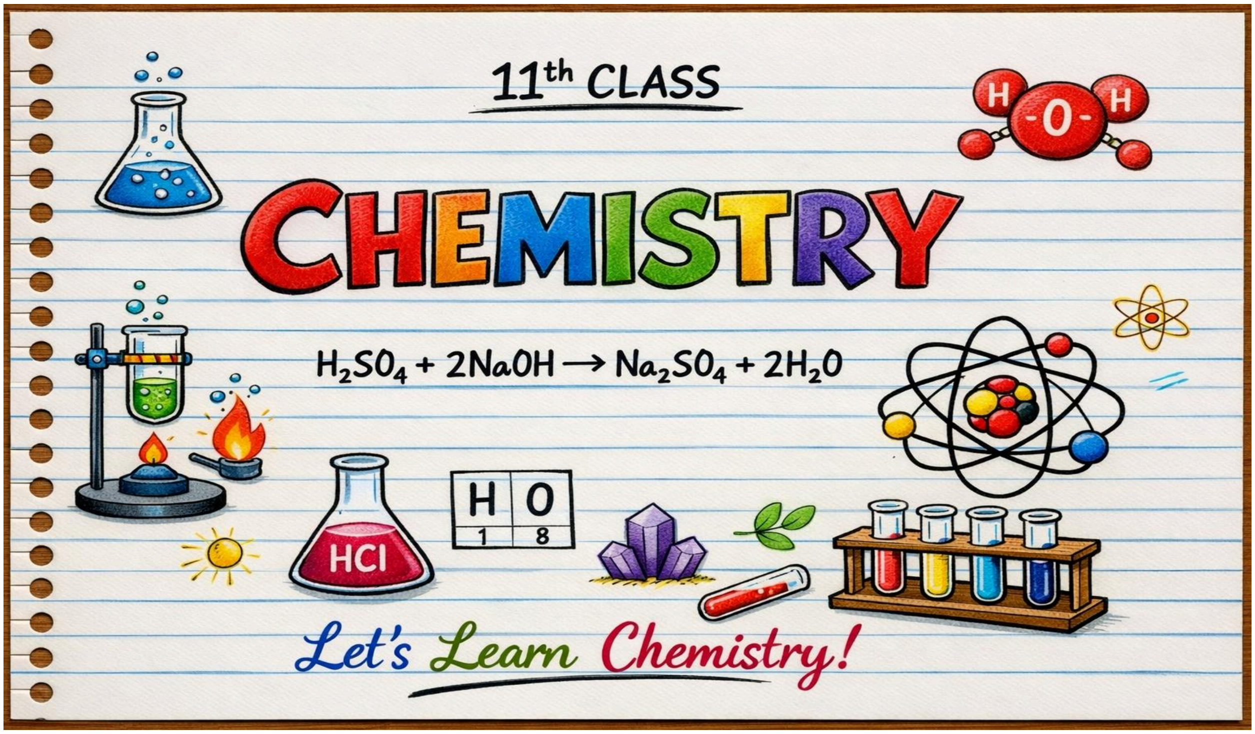

Class: 12th

Subject: chemistry

Unit 6: CHEMICAL ENERGETICS

❖ Important MCQs:

1. Thermodynamics helps to predict:

(a) Rate of reaction

(b) Color of compounds

(c) Spontaneity of a reaction ✅

(d) Molecular shape

2. Thermochemistry is mainly concerned with:

(a) Work done

(b) Heat energy changes during reactions ✅

(c) Pressure variation

(d) Volume expansion

3. Why is thermochemistry called energetics?

(a) It studies motion

(b) It deals with energy changes ✅

(c) It studies pressure

(d) It studies volume

4. According to first law of thermodynamics, energy:

(a) Is created

(b) Is destroyed

(c) Is conserved ✅

(d) Is constant only in gases

5. Hess's law is useful because it:

(a) Measures pressure

(b) Calculates enthalpy change indirectly ✅

(c) Determines rate

(d) Measures volume

6. Enthalpy (H) includes:

(a) Only kinetic energy

(b) Only potential energy

(c) Both kinetic and potential energies ✅

(d) Only heat loss

7. Why is ΔH produced during a reaction?

(a) Because pressure changes

(b) Because temperature is constant

(c) Because enthalpy of reactants and products differ ✅

(d) Because volume changes

8. If Hp > HR, the reaction will be:

(a) Exothermic

(b) Endothermic ✅

(c) Neutral

(d) Reversible

9. In an exothermic reaction, heat flows:

(a) From surroundings to system

(b) From system to surroundings ✅

(c) Within system only

(d) No heat transfer

10. Which condition shows an endothermic process?

(a) ΔH < 0

(b) ΔH > 0 ✅

(c) ΔH = 0

(d) ΔH is constant

11. The enthalpy change (ΔH) represents:

(a) Total energy of system

(b) Net heat absorbed or evolved ✅

(c) Only work done

(d) Only temperature change

12. Combustion of carbon has ΔH = –393.7 kJ mol^{-1}. This indicates:

(a) Heat is absorbed

(b) Reaction is endothermic

(c) Heat is released ✅

(d) Reaction is neutral

13. Which statement is correct about enthalpy?

(a) Same for all substances

(b) Depends on reaction path

(c) Characteristic for each substance ✅

(d) Always zero

14. Dissolution of NH_4Cl in water causes:

(a) Release of heat

(b) Absorption of heat ✅

(c) No heat change

(d) Combustion

15. If ΔH = HP − HR is negative, it means:

(a) HP > HR

(b) HP = HR

(c) HP < HR ✅

(d) No reaction occurs

16. The minimum energy required to start a reaction is called:

(a) Enthalpy

(b) Activation energy ✅

(c) Kinetic energy

(d) Potential energy

17. Activation energy is required to:

(a) Break bonds in products

(b) Form new bonds

(c) Overcome energy barrier by reactants ✅

(d) Increase temperature

18. A reaction will occur only when molecules collide with:

(a) Low energy

(b) No energy

(c) Sufficient energy ✅

(d) Constant speed

19. Energy profile diagram shows:

(a) Pressure change

(b) Volume change

(c) Energy changes during reaction ✅

(d) Color change

20. Standard conditions include temperature:

(a) 0°C

(b) 25°C (298 K) ✅

(c) 100°C

(d) 273°C

21. Standard pressure is:

(a) 2 atm

(b) 1 atm ✅

(c) 5 atm

(d) 10 atm

22. Standard enthalpy change of reaction is measured when:

(a) Any conditions are used

(b) Reactants are not complete

(c) Reactants in standard states react completely ✅

(d) Temperature is zero

23. Enthalpy of combustion is always:

(a) Positive

(b) Zero

(c) Negative (exothermic) ✅

(d) Variable

24. Enthalpy of formation is defined for:

(a) One mole of compound formed from elements ✅

(b) Two moles of compound

(c) Breaking of bonds

(d) Only gases

25. Enthalpy of neutralization is always:

(a) Endothermic

(b) Exothermic ✅

(c) Neutral

(d) Reversible

26. Bond dissociation energy is defined as:

(a) Energy released in bond formation

(b) Energy required to break one mole of a bond ✅

(c) Energy stored in atoms

(d) Heat of reaction

27. When a chemical bond is formed:

(a) Energy is absorbed

(b) Energy is released ✅

(c) No energy change

(d) Temperature decreases

28. When a bond is broken:

(a) Energy is released

(b) Energy is absorbed ✅

(c) No energy change

(d) Pressure increases

29. Bond energy is affected by:

(a) Temperature only

(b) Nature of surrounding atoms ✅

(c) Pressure only

(d) Volume only

30. C–C bond energy is approximately:

(a) 100–150 kJ/mol

(b) 200–250 kJ/mol

(c) 350–380 kJ/mol ✅

(d) 500–600 kJ/mol

31. Chemical bond represents:

(a) Mechanical energy

(b) Chemical energy ✅

(c) Light energy

(d) Nuclear energy

32. A reaction is exothermic when:

(a) Energy absorbed > energy released

(b) Energy released > energy absorbed ✅

(c) Both are equal

(d) No bonds are formed

33. A reaction is endothermic when:

(a) Energy released > absorbed

(b) Energy absorbed > released ✅

(c) No energy change

(d) Bonds are not broken

34. Formation of water is exothermic because:

(a) Less energy is released

(b) More energy is absorbed

(c) More energy is released in bond formation than absorbed in breaking bonds ✅

(d) No bonds are formed

35. A calorimeter is used to measure:

(a) Pressure

(b) Volume

(c) Heat change in reactions ✅

(d) Density

36. Specific heat capacity is defined as:

(a) Heat required to raise 1 g by 1°C

(b) Heat required to raise 1 kg by 1 K ✅

(c) Heat released in reaction

(d) Heat absorbed only

37. Specific heat capacity of water is approximately:

(a) 2.1 J g^{-1} K^{-1}

(b) 3.0 J g^{-1} K^{-1}

(c) 4.2 J g^{-1} K^{-1} ✅

(d) 5.0 J g^{-1} K^{-1}

38. Glass calorimeter is NOT suitable for:

(a) Reactions in solutions

(b) Low temperature reactions

(c) Reactions involving gases ✅

(d) Neutralization reactions

39. Calorie content of food represents:

(a) Mass of food

(b) Volume of food

(c) Energy released on digestion ✅

(d) Temperature of food

40. Energy obtained from food is usually expressed in:

(a) Pascal

(b) Kelvin

(c) Kilocalories (kcal) or Joules (J) ✅

(d) Atmosphere

41. According to Hess’s law, enthalpy change depends on:

(a) Reaction path

(b) Intermediate steps

(c) Initial and final states only ✅

(d) Catalyst used

42. Hess’s law is based on:

(a) Law of conservation of mass

(b) Law of conservation of energy ✅

(c) Boyle’s law

(d) Charles law

43. Hess’s law is useful when:

(a) ΔH can be measured directly

(b) ΔH cannot be measured directly ✅

(c) Temperature is zero

(d) Pressure is high

44. Reaction intermediates are:

(a) Final products

(b) Initial reactants

(c) Substances formed in intermediate steps ✅

(d) Catalysts

45. In energetics of solution, dissolving process involves:

(a) One step only

(b) Two steps only

(c) Three steps ✅

(d) Four steps

46. Standard enthalpy of solution is defined for:

(a) One mole dissolved to give infinitely dilute solution ✅

(b) Two moles dissolved

(c) Only gases

(d) Only solids

47. Hydration is the process in which:

(a) Ions lose electrons

(b) Water surrounds solute ions ✅

(c) Bonds are broken

(d) Heat is absorbed only

48. Enthalpy of hydration is always:

(a) Positive

(b) Negative (energy released) ✅

(c) Zero

(d) Variable

49. Hydration energy increases with:

(a) Large size and low charge

(b) Small size and high charge ✅

(c) Large size and low pressure

(d) Small size and low temperature

50. Born-Haber cycle is used to calculate:

(a) Enthalpy of solution

(b) Lattice energy ✅

(c) Activation energy

(d) Specific heat

51. Entropy of a system is primarily related to:

(a) Mass of particles

(b) Number of possible arrangements of energy ✅

(c) Color of substance

(d) Density only

52. A system with maximum entropy is:

(a) Highly ordered

(b) Completely stable with no motion

(c) Highly disordered with maximum energy distribution ways ✅

(d) Always in solid state

53. Diffusion occurs mainly due to:

(a) External pressure

(b) Random motion of particles and probability of arrangements ✅

(c) Gravitational force

(d) Magnetic force

54. If the number of molecules in a system increases, entropy:

(a) Decreases

(b) Remains constant

(c) Increases due to more possible arrangements ✅

(d) Becomes zero

55. The formula used to calculate possible arrangements in diffusion is:

(a) x + y

(b) x × y

(c) x^y ✅

(d) x − y

56. If 4 particles are distributed in 2 jars, the number of arrangements will be:

(a) 4^2

(b) 2^4 ✅

(c) 4 × 2

(d) 2 × 4

57. Entropy depends on which of the following factors?

(a) Only temperature

(b) Number of particles, state, and physical properties ✅

(c) Only pressure

(d) Only volume

58. Which of the following has the lowest entropy?

(a) Gas

(b) Liquid

(c) Solid ✅

(d) Plasma

59. Among solids, which has lower entropy?

(a) Graphite

(b) Diamond ✅

(c) Liquid mercury

(d) Water vapor

60. The entropy of a substance increases when:

(a) It becomes more ordered

(b) It changes from gas to solid

(c) It changes from solid to liquid or gas ✅

(d) Temperature decreases

61. A reaction in which gas is produced will have:

(a) Decrease in entropy

(b) No change in entropy

(c) Increase in entropy ✅

(d) Zero entropy change always

62. Entropy change is positive when:

(a) Products are more ordered

(b) Number of gaseous molecules decreases

(c) Number of gaseous molecules increases ✅

(d) Temperature is constant

63. In the reaction $CaCO_3 \longrightarrow CaO + CO_2$, entropy:

(a) Decreases

(b) Remains constant

(c) Increases due to formation of gas ✅

(d) Becomes zero

64. Which substance has the highest entropy among the following?

(a) Ice

(b) Liquid water

(c) Water vapor ✅

(d) Diamond

65. Standard entropy is measured under:

(a) 2 atm pressure

(b) 1 atm pressure and specified temperature ✅

(c) Vacuum conditions

(d) High temperature only

66. Gibbs free energy change is given by:

(a) $\Delta G = \Delta H + T\Delta S$

(b) $\Delta G = \Delta H - T\Delta S$ ✅

(c) $\Delta G = T\Delta H - \Delta S$

(d) $\Delta G = \Delta S - \Delta H$

67. A reaction is spontaneous when:

(a) $\Delta G > 0$

(b) $\Delta G = 0$

(c) $\Delta G < 0$ ✅

(d) ΔG is infinite

68. The sign of temperature in Gibbs equation is always:

(a) Negative

(b) Zero

(c) Positive (Kelvin scale) ✅

(d) Variable

69. Gibbs free energy depends strongly on:

(a) Pressure only

(b) Temperature only

(c) Temperature and entropy changes ✅

(d) Volume only

70. At higher temperature, spontaneity of a reaction depends mainly on:

(a) Enthalpy only

(b) Entropy contribution (TΔS term) ✅

(c) Mass of reactants

(d) Color of products

❖ Important Short Questions:

1. Define thermodynamics and mention its scope in science.

Answer:

Thermodynamics is the branch of science that deals with the study of energy, its transformations, and the laws governing these changes in physical and chemical processes.

Scope: It helps in understanding heat, work, energy exchange, and the feasibility of reactions in chemistry, physics, engineering, and biological systems.

Example: Studying how heat is produced during combustion of fuels.

2. Explain why thermochemistry is also called energetics.

Answer:

Thermochemistry is called energetics because it focuses on the study of energy changes, especially heat, during chemical reactions and physical changes.

Example: Measuring heat released during the burning of methane gas.

3. State the first law of thermodynamics.

Answer:

The first law of thermodynamics states that energy can neither be created nor destroyed, but can only be converted from one form to another.

Example: In a heater, electrical energy is converted into heat energy.

4. Define enthalpy and explain what it represents in a system.

Answer:

Enthalpy is the total heat content of a system at constant pressure. It represents the sum of internal energy and the product of pressure and volume of the system.

Example: Heat absorbed or released during a chemical reaction at constant pressure.

5. Write the mathematical expression for enthalpy change and explain its terms.

Answer:

ΔH = H_products − H_reactants

Where:

ΔH = change in enthalpy

H_products = enthalpy of products

H_reactants = enthalpy of reactants

Example: If products have less enthalpy than reactants, ΔH is negative (exothermic reaction).

6. Differentiate between exothermic and endothermic reactions with respect to ΔH.

Answer:

1. Exothermic reactions: Release heat energy; ΔH is negative.

Example: Combustion of fuels.

2. Endothermic reactions: Absorb heat energy; ΔH is positive.

Example: Photosynthesis in plants.

7. Why is ΔH negative for exothermic reactions? Explain.

Answer:

ΔH is negative in exothermic reactions because the system releases heat to the surroundings, resulting in a decrease in the system's enthalpy.

Example: Burning of coal releases heat, so $\Delta H < 0$.

8. State Hess's law and explain its significance in calculating enthalpy change.

Answer:

Hess's law states that the total enthalpy change of a reaction is independent of the path taken and depends only on the initial and final states.

Significance: It allows calculation of enthalpy changes for reactions that cannot be measured directly.

Example: Calculating enthalpy of formation using intermediate reactions.

9. What is activation energy and why is it required for a reaction?

Answer:

Activation energy is the minimum energy required to initiate a chemical reaction. It is needed to break bonds in reactants before new bonds can form.

Example: A matchstick requires initial friction energy to ignite.

10. Describe an energy profile diagram for a chemical reaction.

Answer:

An energy profile diagram shows the change in energy during a reaction. It plots potential energy against reaction progress. It includes:

- Reactants energy level
- Products energy level
- Activation energy (peak)
- Enthalpy change (ΔH) between reactants and products

Example: In an exothermic reaction, products are at lower energy than reactants, while in endothermic reactions, products are at higher energy.

11. Define standard conditions used in thermodynamics.

Answer:

Standard conditions refer to a temperature of 298 K (25°C) and a pressure of 1 atm (or 1 bar), along with standard concentration (1 M for solutions).

Example: Thermodynamic data like ΔH° are measured at these conditions.

12. What is standard enthalpy of reaction (ΔH°r)?

Answer:

Standard enthalpy of reaction is the enthalpy change when reactants in their standard states are converted into products in their standard states under standard conditions.

Example: Combustion of hydrogen gas under standard conditions has a specific ΔH°r value.

13. Define standard enthalpy of combustion and give one example.

Answer:

Standard enthalpy of combustion is the enthalpy change when one mole of a substance is completely burned in oxygen under standard conditions.

Example:

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ (releases heat)

14. What is standard enthalpy of formation? Explain with an example.

Answer:

It is the enthalpy change when one mole of a compound is formed from its elements in their standard states.

Example:

$H_2(g) + \frac{1}{2}O_2(g) \longrightarrow H_2O(l)$

This reaction represents the standard enthalpy of formation of water.

15. Define bond dissociation energy and explain its significance.

Answer:

Bond dissociation energy is the energy required to break one mole of a particular bond in the gaseous state.

Significance: It helps in calculating enthalpy changes and understanding bond strength.

Example: Breaking H–H bond in hydrogen gas requires energy.

16. How does bond breaking and bond formation affect enthalpy change?

Answer:

- Bond breaking requires energy (endothermic).
- Bond formation releases energy (exothermic).

If energy released in bond formation is greater than energy required for bond breaking, the reaction is exothermic, and vice versa.

Example: In combustion, formation of strong bonds in CO_2 and H_2O releases more energy.

17. What is a calorimeter and how is it used to measure heat change?

Answer:

A calorimeter is a device used to measure heat changes during chemical or physical processes. It works by measuring the temperature change of a known mass of substance (usually water).

Example: Used to determine heat released in neutralization reactions.

18. Define specific heat capacity and give its SI unit.

Answer:

Specific heat capacity is the amount of heat required to raise the temperature of 1 gram of a substance by 1°C (or 1 K).

SI Unit: $J\ g^{-1}\ K^{-1}$ (or $J\ kg^{-1}\ K^{-1}$).

Example: Water has a high specific heat capacity, so it heats up slowly.

19. What is entropy and how does it relate to disorder in a system?

Answer:

Entropy is a measure of the randomness or disorder in a system. Higher entropy means greater disorder.

Example: Gas molecules have higher entropy than solids because they move freely.

20. Explain why entropy increases when a solid changes into a liquid or gas.

Answer:

When a solid melts or evaporates, its particles gain freedom of movement, increasing randomness and disorder. Therefore, entropy increases from solid → liquid → gas.

Example: Ice melting into water increases molecular motion and disorder, hence entropy increases.

❖ Important Long Questions:

🌟 Q1. Explain the First Law of Thermodynamics with an example.

Explanation:

The First Law of Thermodynamics states that energy cannot be created or destroyed, but it can be transformed from one form to another. In other words, the total energy of a system and its surroundings remains constant. This law is based on the principle of conservation of energy.

Mathematically, it is expressed as:

$\Delta U = Q - W$

Where:

- ΔU = change in internal energy of the system
- Q = heat added to the system
- W = work done by the system

When heat is supplied to a system, part of that energy may increase the internal energy of the system, and part may be used in doing work on the surroundings.

System and Surroundings:

- A system is the part of the universe under study (e.g., water in a container).
- Surroundings are everything outside the system.
- Energy can be exchanged between the system and surroundings in the form of heat and work.

Example:

When heat is supplied to water in a container, the water absorbs this energy. As a result, its internal energy increases, causing a rise in temperature. If the container allows expansion, some energy may also be used to do work by expanding against the atmosphere. This demonstrates how energy is conserved but changes form.

✦ Summary:

The First Law of Thermodynamics states that energy is conserved and cannot be created or destroyed. It can only be converted from one form to another. When heat is added to a system, it increases the internal energy or is used in doing work. This law is mathematically written as $\Delta U = Q - W$ and is clearly observed when heating water, where supplied heat raises its temperature and internal energy.

🌟 Q2. Differentiate between exothermic and endothermic reactions with examples.

Explanation:

Chemical reactions involve energy changes in the form of heat. Based on heat exchange, reactions are classified into exothermic and endothermic reactions.

Exothermic Reactions:

These reactions release heat energy into the surroundings. In such reactions, the energy of the products is less than the energy of the reactants, so the change in enthalpy (ΔH) is negative. Because heat is released, the surroundings become warmer.

Example:

1. Combustion of methane:

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + \text{heat}$

- This reaction releases heat energy, so it is exothermic.

2. Endothermic Reactions:

- These reactions absorb heat energy from the surroundings. In such reactions, the energy of the products is higher than that of the reactants, so the change in enthalpy (ΔH) is positive. Because heat is absorbed, the surroundings become cooler.

Example:

Photosynthesis:

$6CO_2 + 6H_2O + \text{energy (sunlight)} \longrightarrow C_6H_{12}O_6 + 6O_2$

This reaction absorbs energy from sunlight, so it is endothermic.

✦ Summary:

Exothermic reactions release heat to the surroundings and have a negative ΔH, while endothermic reactions absorb heat from the surroundings and have a positive ΔH. Exothermic reactions make the surroundings warmer (e.g., combustion of methane), whereas endothermic reactions make the surroundings cooler (e.g., photosynthesis).

🌟 Q3. Explain enthalpy and enthalpy change (ΔH).

Explanation:

Enthalpy is the total heat content of a system at constant pressure. It represents the energy stored in a system due to the internal energy of particles and the energy associated with pressure and volume. Enthalpy is denoted by H.

Mathematically, enthalpy is expressed as:

$H = U + PV$

Where:

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

Enthalpy Change (ΔH):

Enthalpy change is the heat absorbed or released during a chemical reaction at constant pressure. It is denoted by ΔH and calculated as:

- $\Delta H = H\_products - H\_reactants$
- **If ΔH is negative,** the reaction is exothermic (heat is released).
- **If ΔH is positive,** the reaction is endothermic (heat is absorbed).

Example:

In the combustion of methane:

$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ + heat

Heat is released, so ΔH is negative. This means the products have lower enthalpy than the reactants.

✦ Summary:

Enthalpy (H) is the total heat content of a system, while enthalpy change (ΔH) represents the heat change during a reaction at constant pressure. It is calculated as the difference between enthalpy of products and reactants. A negative ΔH indicates an exothermic reaction, whereas a positive ΔH indicates an endothermic reaction.

🌟 Q4. State Hess's Law and explain its importance.

Explanation:

Hess's Law states that the total enthalpy change of a chemical reaction is independent of the route taken and depends only on the initial and final states of the reaction.

In simple words, whether a reaction occurs in one step or multiple steps, the overall enthalpy change (ΔH) remains the same.

This is because enthalpy is a state function, meaning it depends only on the state of reactants and products, not on the path followed.

How Hess's Law is Used:

Hess's Law is used to calculate enthalpy changes of reactions that cannot be measured directly. In such cases, the reaction is broken into a series of intermediate steps whose enthalpy changes are known. The total enthalpy change of the main reaction is obtained by adding the enthalpy changes of these steps.

Steps involved:

- Write the target reaction.
- Identify known reactions that lead to the same products and reactants.

- Adjust the equations (reverse or multiply) if necessary.
- Add the equations and their ΔH values to get the overall enthalpy change.

Example:

To determine the enthalpy of formation of a compound like CO_2, we can use multiple steps:

- Instead of measuring directly, we use combustion reactions of carbon and carbon monoxide and combine their enthalpy changes to calculate the desired value.

Importance of Hess's Law:

- It allows calculation of enthalpy changes for reactions that are difficult to measure experimentally.
- It helps in determining standard enthalpy of formation and combustion.
- It is widely used in thermochemical calculations in chemistry.
- It proves that enthalpy is independent of the reaction pathway.

✦ Summary:

Hess's Law states that the total enthalpy change of a reaction is independent of the path taken and depends only on initial and final states. It is used to calculate enthalpy changes indirectly by combining known reactions. This law is important because it helps determine enthalpy values for reactions that cannot be measured directly, using stepwise calculations.

🌟 Q5. Define activation energy and explain the energy profile diagram of a reaction.

Activation Energy:

Activation energy is the minimum amount of energy required by reactant molecules to start a chemical reaction. It is needed to overcome the energy barrier and break the existing bonds in reactants so that new bonds can form.

Without sufficient activation energy, reactants cannot convert into products, even if the reaction is thermodynamically possible.

Example:

A matchstick does not burn on its own. It requires initial energy (friction) to reach activation energy, after which the combustion reaction starts.

Energy Profile Diagram:

An energy profile diagram shows the change in energy during a chemical reaction. It is a graph plotted between potential energy (y-axis) and reaction progress (x-axis).

Key features of the diagram include:

- **Reactants energy level:** Initial energy of the system

- **Products energy level:** Final energy after reaction
- Activation energy (Ea): The energy difference between reactants and the peak of the curve
- **Transition state (activated complex):** The highest energy point where bonds are breaking and forming

Enthalpy change (ΔH): Difference between energy of reactants and products

In an exothermic reaction, products have lower energy than reactants, so energy is released. **In an endothermic reaction,** products have higher energy than reactants, so energy is absorbed.

Role of Catalyst:

A catalyst lowers the activation energy by providing an alternative pathway, making the reaction faster without changing ΔH.

✦ Summary:

Activation energy is the minimum energy required to initiate a reaction. The energy profile diagram represents energy changes during a reaction, showing reactants, products, activation energy, and enthalpy change. In this diagram, the peak represents the transition state, and catalysts reduce the activation energy, speeding up the reaction without affecting the overall energy change.

📘 Exercise:

Q.1 Four choices are given for each question. Select the correct choice.

I. Which of the following equations represents standard heat of formation of C_2H_4?

(a) 2C (diamond) + $2H_2(g) \longrightarrow C_2H_4(g)$

(b) 2C (graphite) + $2H_2(g) \longrightarrow C_2H_4(g)$ ✅

(c) C(graphite) + $H_2(g) \longrightarrow C_2H_4(g)$

(d) 2C (diamond) + $4H(g) \longrightarrow C_2H_4(g)$

II. Which of the following equations correctly defines lattice energy of $MgCl_2$?

(a) $Mg(s) + Cl_2(g) \longrightarrow MgCl_2$

(b) $Mg^{2+}(g) + 2Cl^-(g) \longrightarrow MgCl_2(s)$ ✅

(c) $Mg(s) + 2Cl(g) \longrightarrow MgCl_2(s)$

(d) $Mg^{2+}(g) + 2Cl^-(g) \longrightarrow MgCl_2(g)$

III. Suppose there are 100 molecules of a gas initially in jar A, connected to an evacuated jar B. The possible ways of arrangement will be:

(a) 100

(b) 1/100

(c) 2^100 ✅

IV. For a reaction to occur spontaneously:

(a) (ΔH − TΔS) must be negative ✅

(b) (ΔH + TΔS) must be negative

(c) ΔH must be negative

(d) ΔS must be negative

V. The calorie content of food is fundamentally related to:

(a) Entropy change (ΔS)

(b) Gibbs free energy change (ΔG)

(c) Enthalpy change (ΔH) ✅

(d) Specific heat capacity (c)

VI. Which of the following is NOT typically determined using Hess's Law?

(a) Enthalpy change of formation

(b) Enthalpy change of combustion

(c) Activation energy ✅

(d) Enthalpy change of reaction

VII. Which factor leads to greater enthalpy of hydration (more exothermic)?

(a) Larger ionic radius and smaller charge

(b) Smaller ionic radius and smaller charge

(c) Larger ionic radius and larger charge

(d) Smaller ionic radius and larger charge ✅

VIII. The enthalpy of solution can be expressed as:

(a) ΔH_lattice + ΔH_hydration

(b) ΔH_lattice − ΔH_hydration

(c) −ΔH_lattice + ΔH_hydration

(d) ΔH_lattice − ΔH_hydration ✅

IX. Which reaction represents the standard enthalpy of formation of water?

(a) $2H(g) + O(g) \longrightarrow H_2O(l)$

(b) $H_2(g) + ½O_2(g) \longrightarrow H_2O(l)$ ✅

(c) $H_2(g) + O_2(g) \longrightarrow H_2O(l)$

(d) $2H^+(aq) + O^{2-}(aq) \longrightarrow H_2O(l)$

X. The enthalpy change for a reaction depends on:

(a) Pathway taken from reactants to products

(b) Presence of a catalyst

(c) Initial and final states of the reactants and products ✅

(d) Rate of the reaction

XI. Which of the following processes would typically result in an increase in entropy of the system?

(a) Freezing of water

(b) Condensation of steam

(c) Dissolving a solid in a liquid ✅

(d) Formation of a crystal from a saturated solution

XII. Consider a reaction with $\Delta H > 0$ and $\Delta S < 0$. This reaction will be:

(a) Spontaneous at all temperatures

(b) Non-spontaneous at all temperatures ✅

(c) Spontaneous only at high temperatures

(d) Spontaneous only at low temperatures

Q.2 Attempt the following short-answer questions:

a. Differentiate between exothermic and endothermic reactions

Exothermic reaction:

- A chemical reaction in which heat is evolved from the system to the surroundings. ΔH is negative.

Endothermic reaction:

- A chemical reaction in which heat is absorbed by the system from the surroundings. ΔH is positive.

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

Enthalpy (H) is the heat content of a system. It is the sum of all possible kinetic and potential energies of the system.

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

Entropy (S):

- Entropy is a measure of the randomness or disorder of a system. Greater disorder means higher entropy.

Gibbs free energy (G):

- Gibbs free energy is the thermodynamic quantity used to predict spontaneity of a reaction and is given by:

👉 $G = H - TS$

d. Distinguish between standard enthalpy of reaction and standard enthalpy of formation

Standard enthalpy of reaction ($\Delta H^\circ r$):

- It is the enthalpy change when reactants in their standard states react completely under standard conditions.

Standard enthalpy of formation (ΔH°f):

- It is the enthalpy change when one mole of a compound is formed from its elements in their standard states under standard conditions.

e. Define the following enthalpies with one example

(i) Standard enthalpy of solution (ΔH°sol):

- It is the heat absorbed or evolved when one mole of a substance is dissolved in a solvent to form an infinitely dilute solution under standard conditions.

Example: Dissolution of NH_4Cl in water (endothermic).

(ii) Standard enthalpy of hydration (ΔH°hyd):

- It is the enthalpy change when one mole of gaseous ions is dissolved in excess water to form hydrated ions under standard conditions.

Example:

$Na^+(g) \longrightarrow Na^+(aq)$

(iii) Standard enthalpy of atomization (ΔH°at):

It is the enthalpy change when one mole of gaseous atoms is formed from an element in its standard state.

Example:

$\frac{1}{2}Cl_2(g) \longrightarrow Cl(g)$

(iv) Standard enthalpy of combustion (ΔH°c):

It is the enthalpy change when one mole of a substance is completely burnt in excess oxygen under standard conditions.

Example:

$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$

f. Why is lattice enthalpy usually large negative?

Lattice enthalpy is large negative because a large amount of energy is released when oppositely charged gaseous ions combine to form a solid ionic compound due to strong electrostatic attraction.

g. Factors affecting lattice enthalpy

The magnitude of lattice enthalpy depends on:

- Charge on the ions
- Size of the ions

Greater charge and smaller size result in higher lattice enthalpy.

h. Why is enthalpy of hydration always exothermic?

Enthalpy of hydration is always exothermic because energy is released when water molecules surround and interact with gaseous ions.

The main interaction responsible is:

- Ion–dipole attraction between ions and water molecules.
- This interaction releases energy, making the process exothermic.

i. For the reaction $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$, 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.

Answer:

Bonds broken:

- 4 × C–H bonds (in CH_4)
- 2 × O=O bonds (in O_2)

Bonds formed:

- 2 × C=O bonds (in CO_2)
- 4 × O–H bonds (in H_2O)

j. 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?

Answer:

For spontaneity:

ΔG must be negative ($\Delta G < 0$)

A reaction is always spontaneous when:

- ΔH is negative (exothermic)
- ΔS is positive (increase in disorder)

k. 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.

Answer:

Positive ΔH_{sol}: Heat is absorbed → endothermic process

Negative ΔH_{sol}: Heat is released → exothermic process

l. Consider two ions with similar charges but different sizes, or similar sizes but different charges. Explain how charge density predicts which ion has more exothermic enthalpy of hydration and why.

Answer:

Charge density = charge / size

- Smaller ions with higher charge have greater charge density
- Greater charge density causes stronger ion–dipole attraction with water

- Stronger attraction releases more energy ⟶ more exothermic hydration enthalpy

❖ DESCRIPTIVE QUESTIONS

🌟 Q.3 State and explain Hess' law. Give its two applications.

❖ Answer:

Hess's Law:

Hess's law states that the total enthalpy change in a chemical reaction is independent of the route by which the reaction takes place, provided the initial and final states are the same.

Explanation:

According to the law of conservation of energy, energy can neither be created nor destroyed. Therefore, whether a reaction occurs in one step or in several intermediate steps, the overall enthalpy change (ΔH) remains the same. This can be represented by energy cycles (Hess cycles), where the sum of enthalpy changes of individual steps equals the overall enthalpy change.

Applications of Hess's Law:

1. Calculation of enthalpy changes of reactions

- It is used to calculate ΔH of reactions that cannot be measured directly (e.g., formation of CCl_4).

2. Determination of lattice energy (Born-Haber cycle)

- It is used to calculate lattice energy of ionic compounds using indirect data.

🌟 Q.4 What is lattice energy? How does Born-Haber cycle help to calculate the lattice energy of NaCl?

❖ **Answer:**

Lattice Energy:

Lattice energy is the enthalpy change involved when one mole of an ionic compound is formed from its gaseous ions under standard conditions. It is usually a large negative value due to strong electrostatic attraction between oppositely charged ions.

Born-Haber Cycle (Explanation):

The Born-Haber cycle is an application of Hess's law used to calculate the lattice energy of an ionic compound like NaCl indirectly.

Since lattice energy cannot be measured directly, we use different known enthalpy changes and arrange them in an energy cycle:

Steps involved for NaCl:

1. Sublimation of Na(s) ⟶ Na(g)
2. Ionization of Na(g) ⟶ $Na^+(g) + e^-$
3. Dissociation of $Cl_2(g)$ ⟶ Cl(g)
4. Electron affinity: $Cl(g) + e^- \longrightarrow Cl^-(g)$

5. Formation of NaCl(s) from Na^+(g) and Cl^-(g) ⟶ (Lattice Energy)

By applying Hess's law:

Total enthalpy change = sum of all steps

Thus, lattice energy is calculated indirectly using known values of:

- Enthalpy of formation
- Ionization energy
- Electron affinity
- Bond dissociation energy

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