

## 6. Thermodynamics

### Some Important Points and Terms of the Chapter

1. **System and the Surroundings:** A system in thermodynamics refers to that part of universe in which observations are made and remaining universe constitutes the surroundings. The surroundings include everything other than the system. System and the surroundings together constitute the universe. The wall that separates the system from the surroundings is called **boundary**
2. **Types of the System:**
  - **Open System:** In an open system, there is exchange of energy and matter between system and surroundings.
  - **Closed System** In a closed system, there is no exchange of matter, but exchange of energy is possible between system and the surroundings.
  - **Isolated system** :In an isolated system, there is no exchange of energy or matter between the system and the surroundings
3. **State of a System:** The state of a system means the condition of the system which is described in terms of certain observable properties such as temp(T), pressure(p), volume (v), etc. of the systems. These properties of a system are called state variables.
4. **State Functions:** A physical quantity is said to be state functions if its value depends only upon the state of the system and does not depend upon the path by which this state has been attained.
5. **Internal Energy:** a quantity which represents the total energy of the system. It may be chemical, electrical and mechanical or any other type of energy you may think of, the sum of all these is the energy of the system. In thermodynamics, we call it the internal energy,  $U$  of the system.

6. **Isothermal Process:** When a process is carried out in such a manner that the temp remains constant throughout this process, it is called an isothermal process.
7. **Adiabatic Process:** Process is carried out in such a manner that no heat can flow from the system to the surrounding or vice versa.
8. **Isochoric Process:** Process during which the volume of the system is kept constant.
9. **Isobaric Process:** Process during which the pressure of the system is kept constant. .
10. The positive sign expresses that Work ( $w_{ad}$ ) is positive when work is done **on** the system. Similarly, if the work is done **by** the system,  $w_{ad}$  will be negative.
11. The  $q$  is positive, when heat is transferred from the surroundings to the system and  $q$  is negative when heat is transferred from system to the surroundings.
12. **First law of Thermodynamics:** Statement: Energy can neither be created nor destroyed, however it may be converted from one form to another.

OR

The total energy of the universe remains constant although it may undergo transformation from one to another.

Mathematical expression =  $\Delta U = q + w$

13. A process or change is said to be **reversible**, if a change is brought out in such a way that the process could, at any moment, be reversed by an infinitesimal change. A **reversible process** proceeds infinitely slowly by a series of equilibrium states such that system and the surroundings are always in near equilibrium with each other. Processes other than reversible processes are known as **irreversible processes**.

#### 14. Difference between Reversible Process & Irreversible Process

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Reversible Process	Irreversible Process
1. The process is carried out infinitesimally slowly 2. At any stage, the equilibrium is not disturbed 3. It takes infinite time for completion. 4. Work obtained in this process is maximum.	1. It is carried out rapidly 2. Equilibrium may exist only after the completion of the process. 3. It takes a finite time for completion. 4. Work obtained in this process is not maximum

15. **Enthalpy,  $H$** : The enthalpy  $H$  [Greek word *enthalpien*, heat content] is defined as :  $H = U + pV$  For finite changes at constant pressure, we can write above equation as  $\Delta H = \Delta U + \Delta pV$  Since  $p$  is constant, we can write

$$\boxed{\Delta H = \Delta U + p\Delta V} \text{ since } p\Delta V = \Delta n_g RT, \text{ therefore } \boxed{\Delta H = \Delta U + \Delta n_g RT}$$

16. **Extensive property**: An extensive property is a property whose value depends on the quantity or size of matter present in the system. For example, mass, volume, internal energy, enthalpy, heat capacity, etc. are extensive properties.

17. **Intensive property**: Those properties which do not depend on the quantity or size of matter present are known as intensive properties. For example temperature, density, pressure etc. are intensive properties.

18. **Specific heat capacity** is the quantity of heat required to raise the temperature of one unit mass of a substance by one degree celsius (or one kelvin). For finding out the heat,  $q$ , required to raise the temperatures of a sample, we multiply the specific heat of the substance,  $c$ , by the mass  $m$ , and temperatures change,  $\Delta T$  as  $q = c \times m \times \Delta T = C\Delta T$ .

19. **Relationship between  $C_p$  and  $C_v$  for an ideal gas**:  $\boxed{C_p - C_v = R}$

20. **Bomb calorimeter**: For chemical reactions, heat absorbed at constant volume, is measured in a bomb calorimeter . Here, a steel vessel (the bomb)

is immersed in a water bath. The whole device is called calorimeter. The steel vessel is immersed in water bath to ensure that no heat is lost to the surroundings. A combustible substance is burnt in pure dioxygen supplied in the steel bomb. Heat evolved during the reaction is transferred to the water around the bomb and its temperature is monitored. Since the bomb calorimeter is sealed, its volume does not change i.e., the energy changes associated with reactions are measured at constant volume. Under these conditions, no work is done as the reaction is carried out at constant volume in the bomb calorimeter. Even for reactions involving gases, there is no work done as  $\Delta V = 0$ . Temperature change of the calorimeter produced by the completed reaction is then converted to  $q_V$ , by using the known heat capacity of the calorimeter with the help of equation (18).

21. **Enthalpy Change of a reaction:** The enthalpy change accompanying a reaction is called the reaction enthalpy. The enthalpy change of a chemical reaction is given by the symbol  $\Delta_r H$ .

$\Delta_r H = (\text{sum of enthalpies of products}) - (\text{sum of enthalpies of reactants}).$

$$\Delta_r H = \sum a_i H_{\text{Products}} - \sum b_i H_{\text{Reactants}}$$

22. The **standard enthalpy of reaction** is the enthalpy change for a reaction when all the participating substances are in their standard states. The standard state of a substance at a specified temperature is its pure form at 1 bar. Standard conditions are denoted by adding the superscript  $^\circ$  to the symbol  $\Delta H$ , e.g.,  $\Delta H^\circ$
23. The enthalpy change that accompanies melting of one mole of a solid substance in standard state is called **standard enthalpy of fusion or molar enthalpy of fusion**,  $\Delta_{\text{fus}} H^\circ$ . e.g.  $\text{H}_2\text{O}(\text{s}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta_{\text{fus}} H^\circ = 6.00 \text{ kJ mol}^{-1}$ .
24. Amount of heat required to vaporize one mole of a liquid at constant temperature and under standard pressure (1 bar) is called its **standard enthalpy of vaporization or molar enthalpy of vaporization**,  $\Delta_{\text{vap}} H^\circ$ . e.g.  $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g}); \Delta_{\text{vap}} H^\circ = 40.79 \text{ kJ mol}^{-1}$

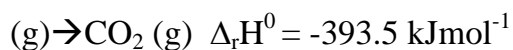
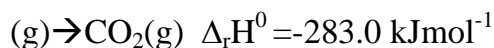
25. **Standard enthalpy of sublimation**,  $\Delta_{sub}H^0$  is the change in enthalpy when one mole of a solid substance sublimates at a constant temperature and under standard pressure (1bar). Sublimation is direct conversion of a solid into its vapour. Solid  $CO_2$  (dry ice) sublimates at 195K with  $\Delta_{sub}H^0=25.2 \text{ kJ mol}^{-1}$ ; naphthalene sublimates slowly and for this  $\Delta_{sub}H^0=73.0 \text{ kJ mol}^{-1}$
26. The **standard enthalpy change for the formation** of one mole of a compound from its elements in their most stable states of aggregation (also known as reference states) is called Standard Molar Enthalpy of Formation. Its symbol is  $\Delta_f H^0$ .
27. **Standard enthalpy of combustion** is defined as the enthalpy change per mole (or per unit amount) of a substance, when it undergoes combustion and all the **reactants and products** being in their standard states at the specified temperature.
28. **Enthalpy of atomization (symbol:  $\Delta_a H^0$ )**: It is the enthalpy change on breaking one mole of bonds completely to obtain atoms in the gas phase. In case of diatomic molecules, like dihydrogen the enthalpy of atomization is also the bond dissociation enthalpy. In some cases, the enthalpy of atomization is same as the enthalpy of sublimation.
29. **Bond Enthalpy (symbol:  $\Delta_{bond} H^0$ )**:
- $$\Delta_r H = \sum \text{bond enthalpies}_{\text{reactants}} - \sum \text{bond enthalpies}_{\text{products}}$$
30. **Enthalpy of Solution (symbol :  $\Delta_{sol} H^0$  )**: Enthalpy of solution of a substance is the enthalpy change when one mole of it dissolves in a specified amount of solvent. The enthalpy of solution at infinite dilution is the enthalpy change observed on dissolving the substance in an infinite amount of solvent when the interactions between the ions (or solute molecules) are negligible.
31. **Lattice Enthalpy** The lattice enthalpy of an ionic compound is the enthalpy change which occurs when one mole of an ionic compound dissociates into its ions in gaseous state.

32. **Hess's Law:** Enthalpy change for a reaction is the same whether it occurs in one step or in a series of steps. This may be stated as follows in the form of:

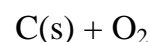
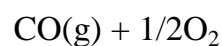
If a reaction takes place in several steps then its standard reaction enthalpy is the sum of the standard enthalpies of the intermediate reactions into

which the overall reaction may be divided at the same temperature.  $\Delta_r H^0$

$$\Delta_r H^0 = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3$$



For example: C(s) +



33. **Spontaneous Process:** A process which occurs by its own i.e. Without the intervention of an outside agency.

34. **Non Spontaneous Process:** A process which can neither take place by itself nor by initiation is called a non spontaneous process.

35. **Driving Force:** The force which is responsible for spontaneity of a process is called the driving force.

36. **Entropy(S):** Entropy is a measure of randomness or disorder of the system.

37. **Entropy change ( $\Delta S$ ):** It is defined as the amount of heat (q) observed isothermally and reversibly divided by the absolute temp(T) at which the heat is absorbed.

$$\Delta S = \frac{q}{T}$$

38. Entropy of fusion :  $\Delta S_{\text{fusion}} = \frac{\Delta H_{\text{fusion}}}{T_m}$

Where,  $\Delta H_{\text{fusion}}$  = Entropy of fusion per mole

$T_m$  = Melting point

$$\Delta S_{\text{VAP}} = \frac{\Delta H_{\text{VAP}}}{T_b}$$

$\Delta H_{\text{VAP}}$  = Entropy of vaporization per mole

$T_b$  = Boiling point

39. Second Law Of Thermodynamics

The entropy of universe is continuously increasing due to spontaneous process taking place in it.

$$\Delta S_{\text{system}} + \Delta S_{\text{surrounding}} > 0$$

$$\text{i.e. } \Delta S_{\text{total}} > 0$$

40. **Gibbs Free Energy (G)** : It is defined as max<sup>m</sup> amount of energy available to a system during the process that can be converted into useful work. It is a measure of capacity to do useful work.

$$G = H - TS$$

this equation is also known as Gibb's Helmholtz equation

G is a state function.  $\Delta G$  is change in free energy.

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

41. **If  $\Delta G$  is negative, process is spontaneous** when  $\Delta G = 0$ , the process is in equilibrium if  $\Delta G$  is positive, the process does not take place.
42. **Standard Free Energy Change ( $\Delta G^{\circ}$ )** : It is defined as free energy change measured at 298 K and 1 atm Pressure.
43. **Standard Free energy of formation: ( $\Delta_f G^{\circ}$ )** It is free energy change when 1 mole of compound is formed from its constituting elements in their standard state.

$$\Delta_r G^{\circ} = \sum \nu_p \Delta_f G^{\circ}(\text{products}) - \sum \nu_R \Delta_f G^{\circ}(\text{reactants})$$

#### 44. Gibbs Energy Change and Work

In case of galvanic cell, Gibbs energy change  $\Delta_r G$ , is related to the electrical work done by the cell.

$$\text{Thus } \Delta_r G = -nFE_{\text{cell}}$$

If reactants and products are in their standard states

$$\Delta_r G^{\circ} = -nFE^{\circ}_{\text{cell}}, \text{ Here } E^{\circ}_{\text{cell}} \text{ is the standard cell potential.}$$

#### 45. Gibbs Energy Change and Equilibrium Constant

$$\Delta_r G^{\circ} = -RT \ln K$$

$$\Delta_r G^{\circ} = -2.303RT \log K$$

### Some Important Formulas(Thermodynamics)

- $\Delta U = q + w$

- $\Delta H = \Delta U + p\Delta V$

- $\Delta H = \Delta U + \Delta ngRT$

- $C_p - C_v = R$

- $\Delta_r H = \sum a_i H_{\text{Products}} - \sum b_i H_{\text{Reactants}}$

- $\Delta_r H = \sum \text{bond enthalpies}_{\text{reactants}} - \sum \text{bond enthalpies}_{\text{products}}$

- $\Delta_r H^0 = \Delta_r H_1 + \Delta_r H_2 + \Delta_r H_3$

- $\Delta S_{\text{fusion}} = \frac{\Delta H_{\text{fusion}}}{T_m}$

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

- $\Delta_r G^0 = \sum \nu_p \Delta_f G^0(\text{products}) - \sum \nu_R \Delta_f G^0(\text{reactants})$

- $\Delta_r G = -nFE_{\text{cell}}$

- $\Delta_r G = -2.303RT \log K$



**Unit-6**  
**THERMODYNAMICS**

**1. Questions based on system, different types of system surroundings, First law of thermodynamics, internal energy**

1. Define the term system and surroundings. Explain the different types of system.
2. Explain the terms: state variables, adiabatic process, work, heat & internal energy.
3. Define First law of thermodynamics. Give its mathematical expression.
4. (a) In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process? (b) Calculate the internal energy change when the system absorbs 5 KJ of heat and 1KJ of work
5. Express the change in internal energy of a system when (i) No heat is absorbed by the system from the surroundings, but work ( $w$ ) is done on the system. What type of wall does the system have? (ii) No work is done on the system, but  $q$  amount of heat is taken out from the system and given to the surroundings. What type of wall does the system have? (iii)  $w$  amount of work is done by the system and  $q$  amount of heat is Supplied to the system. What type of system would it be?

**2. Questions based on system Enthalpy & its types.**

6. Explain the term Enthalpy. Give its mathematical expression.
7. If water vapour is assumed to be a perfect gas, molar enthalpy change for vapourisation of 1 mol of water at 1 bar and 100°C is 41kJ mol<sup>-1</sup>. Calculate the internal energy change, when (i) 1 mol of water is vaporized at 1 bar pressure and 100°C. (ii) 1 mol of water is converted into ice.

8. The reaction of cyanamide,  $\text{NH}_2\text{CN}$  (s), with Dioxygen was carried out in a bomb calorimeter, and  $\Delta U$  was found to be  $-742.7 \text{ KJ/mol}$  at  $298\text{K}$ . Calculate Enthalpy change for the reaction at  $298\text{K}$
- $$\text{NH}_2\text{CN}(\text{s}) + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
9. Enthalpies of formation of  $\text{CO}$  (g),  $\text{CO}_2$  (g),  $\text{N}_2\text{O}$  (g) and  $\text{N}_2\text{O}_4$  (g) are  $-110$ ,  $-393$ ,  $81$  and  $9.7 \text{ KJ/mol}$  respectively. Find the value of  $\Delta_r H$  for the reaction;
- $$\text{N}_2\text{O}_4(\text{g}) + 3 \text{CO}(\text{g}) \rightarrow \text{N}_2\text{O}(\text{g}) + 3\text{CO}_2(\text{g})$$
10. Given:  $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g}); \Delta_r H^\circ = -92.4 \text{ KJ/mol}$ . Calculate  $\Delta_f H^\circ \text{NH}_3(\text{g})$ .
11. The enthalpy of combustion of methane, graphite and dihydrogen at  $298 \text{ K}$  are,  $-890.3 \text{ kJ mol}^{-1}$ ,  $-393.5 \text{ kJ mol}^{-1}$ , and  $-285.8 \text{ kJ mol}^{-1}$  respectively. Calculate the Enthalpy of formation of  $\text{CH}_4(\text{g})$ .
12. For the process to occur under adiabatic conditions, the correct condition is (i)  $\Delta T = 0$  (ii)  $\Delta p = 0$  (iii)  $q = 0$  (iv)  $w = 0$
13. . What is the enthalpies of all elements in their standard states.
14. Enthalpy of combustion of carbon to  $\text{CO}_2$  is  $-393.5 \text{ kJ mol}^{-1}$ . Calculate the heat released upon formation of  $35.2 \text{ g}$  of  $\text{CO}_2$  from carbon and dioxygen gas.
15. The combustion of  $1 \text{ mol}$  of benzene takes place at  $298\text{K}$ . After combustion  $\text{CO}_2$  and  $\text{H}_2\text{O}$  are formed and  $3267\text{KJ/mol}$  of heat is liberated. Calculate  $\Delta_f H^\circ(\text{C}_6\text{H}_6)$  Given:  $\Delta_f H^\circ = -286 \text{ KJ/mol}$ ,  $\Delta_f H^\circ = -393 \text{ KJ/mol}$
16. Calculate the standard enthalpy of formation of  $\text{CH}_3\text{OH}(\text{l})$  from the following data:  $\text{CH}_3\text{OH}(\text{l}) + \frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}); \Delta_r H^\circ = -726 \text{ KJ/mol}$
- $$\text{C}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}); \Delta_c H^\circ = -393 \text{ KJ/mol}$$
- $$\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}); \Delta_f H^\circ = -286 \text{ KJ/mol}$$
17. Calculate the enthalpy change for the process :
- $$\text{CCl}_4(\text{g}) \rightarrow \text{C}(\text{g}) + 4\text{Cl}(\text{g})$$
- and calculate the bond enthalpy of C-Cl in  $\text{CCl}_4(\text{g})$
- $$\Delta_{\text{vap}} H^\circ(\text{CCl}_4) = 30.5 \text{ KJ/mol}, \Delta_f H^\circ(\text{CCl}_4) = -135.5 \text{ KJ/mol}$$
- $$\Delta_a H^\circ(\text{C}) = 715 \text{ KJ/mol}, \Delta_a H^\circ(\text{Cl}_2) = 242 \text{ KJ/mol}$$
18. Define the Extensive, intensive properties & Heat capacity.

19(a) Give the relationship between  $C_p$  and  $C_v$ .

(b) Write a note on Bomb Calorimeter.

20. Explain the following terms with suitable examples: (a) Standard enthalpy of reaction (b) standard enthalpy of formation (c) enthalpy of fusion (d) enthalpy of vaporization (e) enthalpy of Sublimation (f) enthalpy of Combustion (g) enthalpy of Hydration (h) enthalpy of Atomization (i) Bond enthalpy

### **3. Questions based on Hess's law of constant heat summation, Born Haber cycle, Entropy, Gibbs Energy**

21. Explain Hess's law of constant heat summation with an example.

22. Explain Born Haber cycle & lattice enthalpy.

23. Define Entropy. Give mathematical expressions related to it.

24. Predict in which of the following, entropy increases/decreases:

(i) A liquid crystallizes into a solid.

(ii) Temperature of a crystalline solid is raised from 0 K to 115 K.

(iii)  $2\text{NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$

(iv)  $\text{H}_2(\text{g}) \rightarrow 2\text{H}(\text{g})$

25. Define Gibbs Energy. Give its mathematical expression. What is Gibbs energy criteria of Spontaneity.

26. For the reaction at 298K,  $2\text{A} + \text{B} \rightarrow \text{C}$ ,  $\Delta H = 400 \text{ KJ/mol}$  and  $\Delta S = 0.2 \text{ KJ/mol K}$ . At what temperature will the reaction become spontaneous?

27. For the reaction,  $2\text{Cl}(\text{g}) \rightarrow \text{Cl}_2(\text{g})$ , What are the signs of  $\Delta H$  and  $\Delta S$ ?

28. For the reaction:  $2\text{A}(\text{g}) + \text{B}(\text{g}) \rightarrow 2\text{D}(\text{g})$ ,  $\Delta U^0 = -10.5 \text{ KJ}$  and  $\Delta S^0 = -44.1 \text{ J/K}$

Calculate  $\Delta G^0$  for the reaction, and predict whether the reaction will occur spontaneously.

29. The equilibrium constant for a reaction is 10. Calculate  $\Delta G^0$ ,  $T = 300\text{K}$ ,

$$R = 8.314 \text{ J/K mol}$$

30. Calculate the value of  $\Delta G^0$  for the conversion of Oxygen to Ozone,



31. Find out the value of equilibrium constant for the following reaction at 298 K.



Standard Gibbs energy change,  $\Delta_r G^0$  at the given temperature is  $-13.6 \text{ kJ mol}^{-1}$ .

32. Calculate the entropy change in surroundings when 1.00 mol of  $\text{H}_2\text{O}(\text{l})$  is formed under standard conditions.  $\Delta_f H^0 = -286 \text{ kJ mol}^{-1}$ .

**33 What is meant by entropy? Predict the sign of entropy change ( $\Delta S$ ) in each of the following:**

- (a) Temperature of crystalline solid is raised from 0K to 115 K
- (b) A liquid crystallizes into solid
- (c)  $2 \text{ NaHCO}_3(\text{s}) \rightarrow \text{Na}_2\text{CO}_3(\text{s}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g})$
- (d)  $2 \text{ SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$
- (e)  $\text{H}_2(\text{at } 298\text{K}, 1 \text{ atm}) \rightarrow \text{H}_2(\text{at } 298\text{K}, 10 \text{ atm})$
- (f)  $\text{H}_2\text{O}(\text{at } 298\text{K}, 1 \text{ atm}) \rightarrow \text{H}_2\text{O}(\text{at } 330\text{K}, 1 \text{ atm})$
- (g)  $2 \text{ NH}_4\text{NO}_3(\text{s}) \text{ at } 1 \text{ atm \& } 373\text{K} \rightarrow 2 \text{ N}_2(\text{g}) + 4 \text{ H}_2\text{O} + \text{O}_2(\text{g})$
- (h) When rubber band is stretched.
- (i) When an egg is boiled
- (j)  $\text{C}(\text{graphite}) \rightarrow \text{C}(\text{diamond})$
- (k)  $\text{I}_2(\text{g}) \rightarrow \text{I}_2(\text{s})$
- (l)  $\text{Hg}(\text{l}) \rightarrow \text{Hg}(\text{g})$
- (m)  $\text{AgNO}_3(\text{s}) \rightarrow \text{AgNO}_3(\text{aq})$
- (n) Dissolution of iodine in a solvent
- (o) A partition is removed to allow two gases to mix
- (p)  $\text{HCl}$  is added to  $\text{AgNO}_3$  solution and precipitate of  $\text{AgCl}$  is obtained
- (q) crystallization of copper sulphate from its saturated solution

**34 Give reasons for the following:**

- (a) A real crystal has more entropy than an ideal crystal
- (b) The dissolution of  $\text{NH}_4\text{Cl}$  in water is endothermic still it dissolves in water .
- (c) Why does a mole of water at  $0^\circ\text{C}$  have greater entropy than a mole of ice at  $0^\circ\text{C}$  ?
- (d) Neither “q” nor “w” is a state function but “q+w” is a state function.
- (e) Thermodynamically an exothermic reaction is sometimes not spontaneous.
- (f) The entropy of steam is more than that of water at its boiling point.

- (g) The equilibrium constant for a reaction is greater than one if  $\Delta_r G^0$  for it is less than zero .
- (h) Endothermic reactions are carried out at higher temperature.
- (i) Evaporation of water is endothermic process but spontaneous .
- (j) When an ideal gas expands in vacuum, there is neither absorption nor evolution of heat .
- (k) Why does entropy of a solid increase on fusion ?
- (l) Why a non-spontaneous reaction becomes spontaneous when coupled with a suitable spontaneous reaction ?
- (m) why for predicting the spontaneity of a reaction , free energy criteria is better than the entropy criteria ?
- (n) Why internal energy is a state function but work is not ?
- (o) Why is standard heat of formation of diamond not zero although it is an element ?
- (p) Why is entropy of a solution higher than that of pure liquid ?
- (q) acetic acid and hydrochloric acid react with KOH solution . The enthalpy of neutralization of acetic acid is -55.8 kJ per mole while that of hydrochloric acid is -57.3 kJ/mol. Why?

**35 Justify the following statements**

- (a) An exothermic reaction is always thermodynamically spontaneous .why?
- (b)The entropy of a substance increases on going from liquid to vapour state at any temperature.
- ( c) At low temperatures enthalpy change dominates  $\Delta G$  expression and at high temperatures it is the entropy which dominate the value of  $\Delta G$  .
- (d) Many thermodynamically feasible reactions do not occur under ordinary conditions.
- (e) Reactions with  $\Delta G^0 < 0$  always have an equilibrium constant greater than one.

**36 Define.../ Discuss the following terms:(Give examples and chemical equation wherever necessary )**

Standard enthalpy of combustion

(b) Lattice enthalpy (c) Enthalpy of solution

(d) Standard enthalpy fusion / vapourisation / sublimation (e) Enthalpy of atomization

(f) Bond enthalpy for diatomic and polyatomic molecule. (g) Calorific value (h) Enthalpy of Neutralisation .

(i) State variables/ state functions (j) System (k) work (l) isobaric and isochoric process.

**37 Differentiate between the following (with examples )**

(a) Intensive and extensive properties

(b) Enthalpy of formation and Enthalpy of reaction.

(c) Enthalpy and Internal energy (d) Heat capacity and specific heat capacity.

(e) Reversible and Irreversible process (f) Adiabatic and Isothermal process.

(g) State function and path function (h) Exothermic and Endothermic reaction.

(i) Isolated, Open, Closed, Adiabatic Systems (j) Heat and Work.