

Previous Year Questions

Subject	Chemistry
Class	11
Chapter	6
Topic Name	THERMODYNAMICS

1. If an endothermic reaction is non-spontaneous at freezing point of water and becomes feasible at its boiling point, then

- (a) ΔH is $-ve$, ΔS is $+ve$
- (b) ΔH and ΔS both are $+ve$
- (c) ΔH and ΔS both are $-ve$
- (d) ΔH is $+ve$, ΔS is $-ve$

Solution:

$\Delta H = +ve$ for endothermic reaction.

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

ΔG should be positive for a non-spontaneous reaction.

So ΔG is positive at low temperatures if ΔH is positive.

ΔG is negative at high temperatures if ΔS is positive.

Hence option (b) is the answer.

2. Identify the correct statement regarding a spontaneous process:

- (a) Lowering of energy in the reaction process is the only criterion for spontaneity.
- (b) For a spontaneous process in an isolated system, the change in entropy is positive.
- (c) Endothermic processes are never spontaneous.
- (d) Exothermic processes are always spontaneous.

Solution:

In an isolated system, there is neither exchange of energy nor matter between the system and the surroundings. For a spontaneous process in an isolated system, the change in entropy is positive, i.e. $S > 0$. The criteria for spontaneity is the lowering of energy and the increase of entropy.

Hence option (b) is the answer.

3. In an irreversible process taking place at constant T and P and in which only pressure-volume work is being done, the change in Gibbs free energy (dG) and change in entropy (dS), satisfy the criteria

- (a) $(dS)_{V,E} < 0$, $(dG)_{T,P} < 0$
- (b) $(dS)_{V,E} > 0$, $(dG)_{T,P} < 0$
- (c) $(dS)_{V,E} = 0$, $(dG)_{T,P} = 0$
- (d) $(dS)_{V,E} = 0$, $(dG)_{T,P} > 0$

Solution:

For spontaneity, change in entropy (dS) must be positive i.e., greater than zero. Change in Gibbs free energy (dG) must be negative. i.e less than zero. $(dS)_{V,E} > 0$, $(dG)_{T,P} < 0$.

Hence option (b) is the answer.

4. A heat engine absorbs heat Q_1 at temperature T_1 and heat Q_2 at temperature T_2 . The work done by the engine is $J(Q_1 + Q_2)$. This data

- (a) violates 1st law of thermodynamics
- (b) violates 1st law of thermodynamics if Q_1 is -ve
- (c) violates 1st law of thermodynamics if Q_2 is -ve
- (d) does not violate the 1st law of thermodynamics.

Solution:

It does not violate the first law of thermodynamics but violates the second law of thermodynamics.

Hence option (d) is the answer.

5. For a reaction, $A(g) \rightarrow A(l)$; $\Delta H = -3RT$. The correct statement for the reaction is

- (a) $|\Delta H| < |\Delta U|$
- (b) $\Delta H = \Delta U \neq 0$
- (c) $|\Delta H| > |\Delta U|$
- (d) $\Delta H = \Delta U = 0$

Solution:

Given $\Delta H = -3RT$

$\Delta U = \Delta H - \Delta n_g RT$

$\Delta n_g = n_p - n_r = -1$

$\Delta U = -3RT - (-RT) = -2RT$

So $|\Delta H| > |\Delta U|$

Hence option (c) is the answer.

6. If the bond dissociation energies of XY , X_2 and Y_2 (all diatomic molecules) are in the ratio of 1: 1: 0.5 and ΔH_f for the formation of XY is -200 kJ mol^{-1} . The bond dissociation energy of X_2 will be

- (a) 100 kJ mol^{-1}
- (b) 200 kJ mol^{-1}
- (c) 800 kJ mol^{-1}
- (d) 400 kJ mol^{-1}

Solution:

Given $\Delta H_f = -200 \text{ kJ mol}^{-1}$

Let the bond dissociation energy of XY , X_2 and Y_2 be $x \text{ kJ mol}^{-1}$, $x \text{ kJ mol}^{-1}$ and $0.5x \text{ kJ mol}^{-1}$ respectively.

$\frac{1}{2} X_2 + \frac{1}{2} Y_2 \rightarrow XY$

$\Delta H_f = \text{sum of bond dissociation energy of all reactants} - \text{the sum of bond dissociation energy of product.}$

$-200 = \frac{1}{2} \Delta H_{X_2} + \frac{1}{2} \Delta H_{Y_2} - \Delta H_{XY}$

$-200 = (x/2) + (0.5x/2) - x$

$-200 = 0.5x/2$

So $x = 200(2/0.5)$

$= 400/0.5$

$= 800 \text{ kJ mol}^{-1}$

Hence option (c) is the answer.

7. The enthalpy change for a reaction does not depend upon the

- (a) physical states of reactants and products
- (b) use of different reactants for the same product
- (c) nature of intermediate reaction steps
- (d) difference in initial or final temperatures of involved substances.

Solution:

The enthalpy change for a reaction does not depend upon the nature of intermediate reaction steps. This is according to Hess's law.

Hence option (c) is the answer.

8. The incorrect match in the following is

- (a) $\Delta G^\circ = 0$, $K = 1$
- (b) $\Delta G^\circ < 0$, $K < 1$
- (c) $\Delta G^\circ < 0$, $K > 1$
- (d) $\Delta G^\circ > 0$, $K < 1$

Solution:

$$\Delta G^\circ = -2.303 RT \log K$$

If $\Delta G^\circ = 0$, $K = 1$.

If $\Delta G^\circ < 0$, K is greater than 1.

If $\Delta G^\circ > 0$, K is less than 1.

Hence option (b) is the answer.

9. The internal energy change when a system goes from state A to B is 40 kJ/mol. If the system goes from A to B by a reversible path and returns to state A by an irreversible path what would be the net change in internal energy?

- (a) 40 kJ
- (b) > 40 kJ
- (c) < 40 kJ
- (d) zero

Solution:

For a cyclic process, the net change in internal energy is equal to zero. The change in the internal energy does not depend on the path by which the final state is reached.

Hence option (d) is the answer.

10. During compression of a spring, the work done is 10 kJ and 2 kJ escapes to the surrounding as heat. The change in internal energy ΔU (in kJ) is

- (a) -8
- (b) 12
- (c) 8
- (d) -12

Solution:

$$w = 10 \text{ kJ}$$

$$q = -2 \text{ kJ}$$

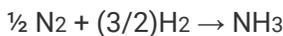
According to the first law of thermodynamics, $\Delta U = q + w$

$$= -2 + 10 = 8 \text{ kJ}$$

Hence option (c) is the answer.

11. The standard enthalpy of formation of NH_3 is -46 kJ mol^{-1} . If the enthalpy of formation of H_2 from its atoms is -436 kJ mol^{-1} and that of N_2 is -712 kJ mol^{-1} , the average bond enthalpy of N-H bond in NH_3 is

- (a) $-1102 \text{ kJ mol}^{-1}$
- (b) -964 kJ mol^{-1}
- (c) $+352 \text{ kJ mol}^{-1}$
- (d) $+1056 \text{ kJ mol}^{-1}$

Solution:

$$(\Delta H_f) \text{NH}_3 = [\frac{1}{2} \text{B.E N}_2 + (3/2) \text{B.E H}_2 - 3 \text{B.E N-H}]$$

$$-46 = [\frac{1}{2} 712 + (3/2) 436 - 3 \text{B.E N-H}]$$

$$-46 = 356 + 654 - 3 \text{B.E N-H}$$

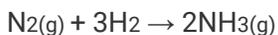
$$3 \text{B.E N-H} = 1056$$

$$\text{B.E N-H} = 1056/3 = 352 \text{ kJ mol}^{-1}$$

Hence option (c) is the answer.

12. The process with negative entropy change is

- (a) dissolution of iodine in water
- (b) sublimation of dry ice
- (c) synthesis of ammonia from N₂ and H₂
- (d) dissociation of CaSO₄(s) to CaO(s) and SO₃(g).

Solution:

$$\Delta s = 2 - 4 = -2 < 0$$

So entropy is negative.

Hence option (c) is the answer.

13. ($\Delta H - \Delta U$) for the formation of carbon monoxide (CO) from its elements at 298 K is ($R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$)

- (a) $-1238.78 \text{ J mol}^{-1}$
- (b) $1238.78 \text{ J mol}^{-1}$
- (c) $-2477.57 \text{ J mol}^{-1}$
- (d) $2477.57 \text{ J mol}^{-1}$

Solution:

Given $T = 298$

$$\Delta H - \Delta U = \Delta n_g RT$$



$$\Delta n_g = 1 - \frac{1}{2} = \frac{1}{2}$$

$$\Delta n_g RT = \frac{1}{2} \times 8.314 \times 298 = 1238.78 \text{ J mol}^{-1}$$

Hence option (b) is the answer.

14. An ideal gas undergoes isothermal expansion at constant pressure. During the process

- (a) enthalpy increases but entropy decreases
- (b) enthalpy remains constant but entropy increases
- (c) enthalpy decreases but entropy increases
- (d) both enthalpy and entropy remain constant.

Solution:

During isothermal expansion at constant pressure, $\Delta H = nC_p \Delta T = 0$

$$\Delta S > 0$$

Hence option (b) is the answer.

15. ΔU is equal to

- (a) adiabatic work
- (b) isothermal work
- (c) isochoric work
- (d) isobaric work.

Solution:

From the first law of thermodynamics, $\Delta U = q + w$

For adiabatic process, $q = 0$

So $\Delta U = w$

Change in internal energy is equal to adiabatic work.

Hence option (a) is the answer.

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