

Q1. Which of the following is a state function?

- A. Work
- B. Heat
- C. Internal energy
- D. None of the above

Answer: C. Internal energy

Explanation:

State functions depend only on the initial and final states of a system, not the path taken. Internal energy (U) is a state function, whereas work and heat are path functions.

Q2. Which of the following equations correctly represents the First Law of Thermodynamics?

- A. $\Delta U = q + w$
- B. $\Delta U = q - w$
- C. $q = \Delta U + w$
- D. $\Delta U = -q + w$

Answer: A. $\Delta U = q + w$

Explanation:

The first law of thermodynamics states that the change in internal energy (ΔU) is equal to the heat added to the system (q) plus the work done on the system (w).

Q3. An ideal gas is compressed adiabatically. Which of the following is true?

- A. $q = 0$, $\Delta U = -w$
- B. $q = 0$, $\Delta U = w$
- C. $q \neq 0$, $\Delta U = w$
- D. $q \neq 0$, $\Delta U = -w$

Answer: A. $q = 0$, $\Delta U = -w$

Explanation:

In adiabatic processes, no heat is exchanged ($q = 0$). Thus, $\Delta U = w$. If work is done on the system (compression), w is negative, hence $\Delta U = -w$.

Q4. During an isothermal expansion of an ideal gas, which of the following is true?

- A. $\Delta U = 0$, $q = -w$
- B. $\Delta U = 0$, $q = w$
- C. $\Delta U \neq 0$, $q \neq w$
- D. $\Delta U \neq 0$, $q = w$

Answer: B. $\Delta U = 0$, $q = w$

Explanation:

In an isothermal process (constant temperature), internal energy does not change ($\Delta U = 0$), so heat absorbed is equal to the work done by the system.

Q5. Match the following:

Column I

- A. Enthalpy
- B. Entropy
- C. Heat
- D. Free Energy

Column II

- 1. Path function
- 2. Spontaneity
- 3. Heat content at constant pressure
- 4. Disorder

A–B–C–D:

- A. 3–4–1–2
- B. 4–3–1–2
- C. 2–1–4–3
- D. 3–2–4–1

Answer: A. 3–4–1–2

Explanation:

Enthalpy \rightarrow Heat content at constant pressure

Entropy → Measure of disorder

Heat → Path function

Free energy → Determines spontaneity

Q6. In a spontaneous process, the total entropy of the system and surroundings:

- A. Always increases
- B. Always decreases
- C. Remains constant
- D. Becomes zero

Answer: A. Always increases

Explanation:

According to the Second Law of Thermodynamics, for a spontaneous process, the total entropy change of the system and surroundings is positive.

Q7. Which of the following conditions must be satisfied for a process to be spontaneous at constant T and P?

- A. $\Delta G > 0$
- B. $\Delta G < 0$
- C. $\Delta G = 0$
- D. $\Delta H = T\Delta S$

Answer: B. $\Delta G < 0$

Explanation:

A process is spontaneous at constant temperature and pressure when the change in Gibbs free energy (ΔG) is negative.

Q8. The value of ΔH for the vaporization of water is 40.7 kJ/mol at 100°C. What is ΔS for this process?

- A. 373 J/mol·K
- B. 109 J/mol·K
- C. 22 J/mol·K
- D. 40.7 J/mol·K

Answer: B. 109 J/mol·K

Explanation:

$$\Delta S = \Delta H / T = (40.7 \times 10^3 \text{ J/mol}) / 373 \text{ K} \approx 109 \text{ J/mol}\cdot\text{K}$$

Q9. For a reaction, $\Delta H = -20 \text{ kJ}$ and $\Delta S = -100 \text{ J/mol}\cdot\text{K}$. At what temperature does the reaction become non-spontaneous?

- A. 200 K
- B. 300 K
- C. 400 K
- D. 500 K

Answer: C. 400 K

Explanation:

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

$$\text{Set } \Delta G = 0 \Rightarrow 0 = -20,000 - T(-100) \Rightarrow T = 20000 / 100 = 200 \text{ K}$$

Above 200 K, ΔG becomes positive \Rightarrow Non-spontaneous.

Correction: This means 200 K is the turning point. For non-spontaneity, $T > 200 \text{ K}$.
So option C. 400 K satisfies non-spontaneous condition.

Q10. Which of the following statements is NOT true?

- A. Heat is a path function
- B. Enthalpy is a state function
- C. Work is a state function
- D. Internal energy is a state function

Answer: C. Work is a state function

Explanation:

Work depends on the path taken, not just the initial and final states, hence it is a path function, not a state function.

Q11. A gas expands from volume V_1 to V_2 at constant pressure P . The work done is given by:

- A. $W = -P(V_2 - V_1)$
- B. $W = P(V_2 - V_1)$
- C. $W = nRT \ln(V_2/V_1)$
- D. $W = -nRT \ln(V_2/V_1)$

Answer: B. $W = P(V_2 - V_1)$

Explanation:

At constant pressure, work done by the gas during expansion is:

$$W = P\Delta V = P(V_2 - V_1)$$

Q12. For a certain reaction, $\Delta H = -100 \text{ kJ/mol}$ and $\Delta S = -200 \text{ J/mol}\cdot\text{K}$. The reaction will be spontaneous at:

- A. High temperatures only
- B. Low temperatures only
- C. All temperatures
- D. Never spontaneous

Answer: B. Low temperatures only

Explanation:

$$\Delta G = \Delta H - T\Delta S = -100,000 + 200T$$

$\Delta G < 0$ (spontaneous) when T is small. At high T, ΔG becomes positive.

Q13. When an ideal gas undergoes free expansion in a vacuum:

- A. $q \neq 0, w = 0, \Delta U = 0$
- B. $q = 0, w \neq 0, \Delta U = 0$
- C. $q = 0, w = 0, \Delta U = 0$
- D. $q = 0, w = 0, \Delta U \neq 0$

Answer: C. $q = 0, w = 0, \Delta U = 0$

Explanation:

In free expansion:

No heat exchange ($q = 0$)

No work done ($w = 0$)

For an ideal gas, internal energy depends on temperature, which remains constant $\rightarrow \Delta U = 0$

Q14. In a cyclic process, the net change in internal energy is:

- A. Equal to the net work done
- B. Zero
- C. Equal to net heat supplied
- D. Infinite

Answer: B. Zero

Explanation:

In a cyclic process, the system returns to its initial state. Internal energy is a state function, so $\Delta U = 0$.

Q15. Assertion (A): Enthalpy of formation of elements in their standard state is zero.

Reason (R): Enthalpy is a state function.

- A. Both A and R are true, and R is the correct explanation of A
- B. Both A and R are true, but R is not the correct explanation of A
- C. A is true, R is false
- D. A is false, R is true

Answer: B. Both A and R are true, but R is not the correct explanation of A

Explanation:

Standard enthalpy of formation of an element in its standard state is taken as zero by convention. Though enthalpy is a state function, that's not the reason for the value being zero.

Q16. A gas expands isothermally and reversibly. Which of the following is correct?

- A. $\Delta U = 0$, $q = -w$
- B. $\Delta U \neq 0$, $q = w$
- C. $\Delta U = 0$, $q = 0$
- D. $q = 0$, $w = 0$

Answer: A. $\Delta U = 0$, $q = -w$

Explanation:

Isothermal process $\rightarrow \Delta U = 0$ (for ideal gas). From First Law:

$$\Delta U = q + w \Rightarrow 0 = q + w \Rightarrow q = -w$$

Q17. The relationship between C_p and C_v for an ideal gas is:

- A. $C_p = C_v$
- B. $C_p > C_v$
- C. $C_p < C_v$
- D. $C_p = C_v + R/2$

Answer: B. $C_p > C_v$

Explanation:

$C_p - C_v = R$ (ideal gas relation). Therefore, C_p is always greater than C_v .

Q18. Which of the following thermodynamic quantities is not a state function?

- A. Enthalpy
- B. Internal energy
- C. Heat
- D. Entropy

Answer: C. Heat

Explanation:

Heat and work are path functions. Internal energy, enthalpy, and entropy are state functions.

Q19. The change in Gibbs free energy (ΔG) for a process is zero. This means the process is:

- A. Irreversible
- B. At equilibrium
- C. Spontaneous
- D. Non-spontaneous

Answer: B. At equilibrium

Explanation:

At equilibrium, $\Delta G = 0$. This indicates no net change in free energy and no net forward or backward progress.

Q20. Match the thermodynamic processes with their characteristics:

Column I

- A. Isothermal
- B. Adiabatic
- C. Isochoric
- D. Isobaric

Column II

- 1. $q = 0$
- 2. $\Delta U = 0$
- 3. $\Delta V = 0$
- 4. $\Delta P = 0$

A–B–C–D:

- A. 2–1–3–4
- B. 1–2–3–4
- C. 4–3–2–1
- D. 2–3–1–4

Answer: A. 2–1–3–4

Explanation:

Isothermal $\rightarrow \Delta U = 0$ (temperature constant)

Adiabatic $\rightarrow q = 0$ (no heat exchange)

Isochoric $\rightarrow \Delta V = 0$ (constant volume)

Isobaric $\rightarrow \Delta P = 0$ (constant pressure)

Q21. The enthalpy change for the combustion of glucose is -2800 kJ/mol . What is the enthalpy change for the combustion of 90 g of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$)?

- A. -1400 kJ
- B. -2800 kJ

- C. -12600 kJ
- D. -5600 kJ

Answer: A. -1400 kJ

Explanation:

Molar mass of glucose = 180 g/mol

Given: $\Delta H = -2800$ kJ/mol

For 90 g: $(90/180) \times (-2800) = -1400$ kJ

Q22. In which of the following processes, entropy decreases?

- A. Evaporation of water
- B. Sublimation of iodine
- C. Freezing of water
- D. Melting of ice

Answer: C. Freezing of water

Explanation:

Freezing involves conversion from a more disordered (liquid) to a more ordered (solid) state \rightarrow entropy decreases.

Q23. The temperature at which a reaction becomes spontaneous when $\Delta H = +40$ kJ and $\Delta S = +100$ J/mol·K is:

- A. 200 K
- B. 300 K
- C. 400 K
- D. 500 K

Answer: D. 500 K

Explanation:

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

Set $\Delta G = 0$:

$$0 = 40,000 - T(100)$$

$$T = 40000/100 = 400 \text{ K} \rightarrow \text{above this, } \Delta G < 0$$

So reaction is spontaneous at $T > 400 \text{ K} \rightarrow 500 \text{ K}$

Q24. Assertion (A): Entropy of a perfect crystal at 0 K is zero.

Reason (R): Third law of thermodynamics states that $S = 0$ at 0 K for a perfect crystalline substance.

- A. Both A and R are true, and R is the correct explanation of A
- B. Both A and R are true, but R is not the correct explanation of A
- C. A is true, R is false
- D. A is false, R is true

Answer: A. Both A and R are true, and R is the correct explanation of A

Q25. For an endothermic process to be spontaneous, the entropy change of the system must be:

- A. Zero
- B. Negative
- C. Positive and large
- D. Negative and large

Answer: C. Positive and large

Explanation:

Endothermic: $\Delta H > 0 \rightarrow$ to make $\Delta G < 0$, $T\Delta S$ must be large and positive.

Q26. Which one of the following is a path function?

- A. Enthalpy
- B. Internal energy
- C. Work
- D. Entropy

Answer: C. Work

Explanation:

Work depends on the path followed during a process. It's not a state function.

Q27. Calculate work done in joules when 1 mole of an ideal gas expands reversibly and isothermally from 1 L to 10 L at 300 K. ($R = 8.314 \text{ J/mol}\cdot\text{K}$)

- A. 5744J
- B. 1716 J
- C. 1915 J
- D. 2191 J

Answer: A . 5744J

Explanation:

$$\begin{aligned}W &= nRT \ln(V_2/V_1) \\&= 1 \times 8.314 \times 300 \times \ln(10/1) \\&= 8.314 \times 300 \times 2.303 = \sim 5744 \text{ J}\end{aligned}$$

Q28. Which of the following combinations is correct?

Match the terms in Column I with appropriate values or descriptions from Column II:

Column I

- A. Isobaric process
- B. Isochoric process
- C. Isothermal process
- D. Adiabatic process

Column II

- 1. Constant pressure
- 2. Constant volume
- 3. No heat exchange
- 4. Constant temperature

A–B–C–D:

- A. 1–2–4–3
- B. 1–2–3–4
- C. 2–1–4–3
- D. 1–2–3–4

Answer: A. 1–2–4–3

Q29. In which of the following conditions will a gas show the most deviation from ideal behavior?

- A. High temperature and low pressure
- B. Low temperature and high pressure
- C. Low temperature and low pressure
- D. High temperature and high pressure

Answer: B. Low temperature and high pressure

Explanation:

Under these conditions, intermolecular forces and volume of particles become significant → deviation from ideality increases.

Q30. One mole of a gas is compressed isothermally from 10 L to 1 L at 300 K. The change in internal energy is:

- A. Zero
- B. Positive
- C. Negative
- D. Infinite

Answer: A. Zero

Explanation:

Isothermal process for ideal gas → $\Delta U = 0$ (internal energy depends only on temperature, which is constant)

Q31. Which of the following equations correctly represents the first law of thermodynamics?

- A. $\Delta H = q - w$
- B. $\Delta U = q + w$
- C. $\Delta S = q/T$
- D. $\Delta U = q - w$

Answer: B. $\Delta U = q + w$

Explanation:

The first law states that the change in internal energy is equal to the heat added to the system plus work done on the system.

Q32. Which condition must be met for a chemical reaction to be spontaneous at all temperatures?

- A. $\Delta H < 0$ and $\Delta S < 0$
- B. $\Delta H > 0$ and $\Delta S < 0$
- C. $\Delta H < 0$ and $\Delta S > 0$
- D. $\Delta H > 0$ and $\Delta S > 0$

Answer: C. $\Delta H < 0$ and $\Delta S > 0$

Explanation:

For spontaneity at all temperatures, $\Delta G = \Delta H - T\Delta S$ must be negative regardless of T. That happens when ΔH is negative and ΔS is positive.

Q33. In an adiabatic process, heat exchanged with the surroundings is:

- A. Always positive
- B. Always negative
- C. Zero
- D. Infinite

Answer: C. Zero

Explanation:

Adiabatic processes occur without any heat exchange ($q = 0$).

Q34. For a given process at constant pressure, the heat change (q_p) is equal to:

- A. ΔU
- B. ΔS
- C. ΔH
- D. ΔG

Answer: C. ΔH

Explanation:

At constant pressure, $q_p = \Delta H$.

Q35. Assertion (A): A gas expands adiabatically and its temperature decreases.

Reason (R): In adiabatic expansion, work is done by the gas at the cost of internal energy.

- A. Both A and R are true, and R is the correct explanation of A
- B. Both A and R are true, but R is not the correct explanation of A
- C. A is true, R is false
- D. A is false, R is true

Answer: A. Both A and R are true, and R is the correct explanation of A

Q36. Which of the following is NOT a state function?

- A. Internal energy
- B. Enthalpy
- C. Work
- D. Entropy

Answer: C. Work

Explanation:

Work depends on the path, not the state. Others depend only on initial and final states.

Q37. Which process will have maximum entropy change?

- A. Melting of ice
- B. Freezing of water
- C. Condensation of steam
- D. Vaporization of water

Answer: D. Vaporization of water

Explanation:

Vaporization leads to gas phase → greatest increase in randomness → maximum entropy change.

Q38. 2 mol of an ideal gas is expanded from 2 L to 6 L at constant temperature. What is the work done?

- A. 3.46 kJ
- B. 2.2 kJ

C. 5.48 kJ

D. 8.3 kJ

Answer: C. 5.48 kJ

Explanation:

$$W = nRT \ln(V_2/V_1)$$

$$= 2 \times 8.314 \times 300 \times \ln(6/2)$$

$$= 2 \times 8.314 \times 300 \times \ln(3)$$

$$\approx 5.48 \text{ kJ}$$

Q39. Match the following:

Column I

A. Isothermal

B. Adiabatic

C. Isochoric

D. Isobaric

Column II

1. $\Delta T = 0$

2. $q = 0$

3. $\Delta V = 0$

4. $\Delta P = 0$

A – B – C – D:

A. 1 – 2 – 3 – 4

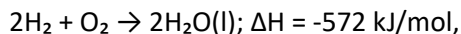
B. 2 – 1 – 4 – 3

C. 1 – 3 – 2 – 4

D. 4 – 3 – 2 – 1

Answer: A. 1 – 2 – 3 – 4

Q40. For the reaction:



The amount of heat evolved when 4 g of hydrogen reacts is:

- A. 572 kJ
- B. 286 kJ
- C. 1144 kJ
- D. 143 kJ

Answer: B. 286 kJ

Explanation:

Molar mass $H_2 = 2 \text{ g/mol}$

4 g = 2 mol \rightarrow From equation, 2 mol gives -572 kJ

So for 2 mol $H_2 \rightarrow -572 \text{ kJ}$

\therefore Answer = 572 kJ

Q41. Which of the following statements is correct?

- A. ΔG is always zero for a spontaneous process
- B. ΔG is always negative for a non-spontaneous process
- C. ΔG is zero at equilibrium
- D. ΔG is always positive for a spontaneous process

Answer: C. ΔG is zero at equilibrium

Q42. In a spontaneous reaction:

- A. $\Delta H > 0$ and $\Delta S > 0$
- B. $\Delta H < 0$ and $\Delta S < 0$
- C. $\Delta G < 0$
- D. $\Delta G > 0$

Answer: C. $\Delta G < 0$

Q43. Which of the following is the correct condition for equilibrium in terms of Gibbs free energy?

- A. $\Delta G > 0$
- B. $\Delta G = 0$
- C. $\Delta G < 0$
- D. $\Delta G = \infty$

Answer: B. $\Delta G = 0$

Q44. The standard entropy of formation of any element in its most stable form is:

- A. 0
- B. 1
- C. 298
- D. Infinite

Answer: A. 0

Explanation:

Standard enthalpy of formation = 0 for elements in standard state

But entropy is not zero.

So correction: The entropy of formation is not necessarily zero,

Q45. Gibbs free energy change is related to equilibrium constant (K) by the equation:

- A. $\Delta G = -RT \ln K$
- B. $\Delta G = RT \ln K$
- C. $\Delta G = R \ln K$
- D. $\Delta G = -R \ln K$

Answer: A. $\Delta G = -RT \ln K$

Explanation:

This is the thermodynamic relation connecting ΔG and equilibrium constant.