Solutions

S1. Ans.(b)

When a liquid evaporates to vapour entropy increases.

$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$

Number of gaseous product molecules increases so entropy increases.

$Cl_{2(g)} \rightarrow 2Cl_{(g)}$

1 mole $Cl_{2(g)}$ form 2 mol $Cl_{(g)}$.

So, entropy increases.

S2. Ans.(c)

(A) Isothermal process ⇒ Temperature is constant throughout the process.

(B) Isochoric process ⇒ Volume is constant throughout the process.

(C) Isobaric process ⇒ Pressure is constant throughout the process.

(D) Adiabatic process ⇒ No exchange of heat (q) between system and surrounding.

S3. Ans.(a)

$W_{rev,\ iso} = -2.303\ nRT \log \frac{P_i}{P_f}$

$= -2.303 \times 1 \times 2 \times 298 \times \log \frac{20}{10}$

$= -2.303 \times 1 \times 2 \times 298 \times \log 2$

$= -2.303 \times 1 \times 2 \times 298 \times 0.3$

$= -413.14$ calories

S4. Ans.(b)

Decomposition for 1 mole of water

$H_2O(g) \longrightarrow H_2(g) + \frac{1}{2}O_2(g);\ \Delta H = +\frac{483.64}{2}$

$\Delta H = +241.82$ kJ

S5. Ans.(b)

$A + B \rightleftharpoons C + D$

$[A] = 2\ mol\ L^{-1}$

$[B] = 3\ mol\ L^{-1}$

$[C] = 10\ mol\ L^{-1}$

$[D] = 6\ mol\ L^{-1}$

$\Delta G° = -2.303\ RT \log K_{eq}$

$= -2.303\ RT \log \frac{[C][D]}{[A][B]}$

$= -2.303 \times 2 \times 300 \times \log \frac{10\times6}{2\times3}$

$= -2.303 \times 2 \times 300 \times \log 10$

$= -1381.8$ cal

S6. Ans.(a)

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

S7. Ans.(a)

$\Delta U = nC_v\Delta T$

For isothermal condition; $\Delta T = 0$

$\therefore \Delta U = 0$

S8. Ans.(b)

At constant temperature and amount

$P_1V_1 = P_2V_2$

$P_1V_1 = \frac{P_1}{3}V_2\ [\because P_2 = \frac{P_1}{3}]$

$V_2 = 3V_1$

mole of $O_2(g) = 3.2/32 = 0.1$ mole

Volume of $O_2(g) = (0.1 \times 22.4)$ L = 2.24 L

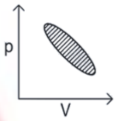
At STP (V_1)

$V_2 = 3V_1 = 3 \times 2.24 = 6.72$ L

S9. Ans.(b)

Work done under any thermodynamic process can be determined by area under the 'p-V' graph. As it can be observed maximum area is covered in option '2'.

1.



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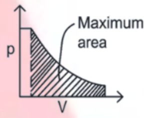

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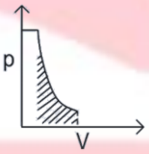

S10. Ans.(a)

For one mole of an ideal gas $C_P - C_V = R$

S11. Ans.(b)

For Irreversible expansion of an ideal gas under Isothermal condition

$\Delta U = 0$

$\Delta S_{Total} \neq 0$

S12. Ans.(c)

Given reaction, $2Cl(g) \rightarrow Cl_2(g)$

The reaction is exothermic since a bond is being formed in the given reaction, i.e., ΔH is also –ve. So $\Delta_r H < 0$ and Entropy is decreasing (–ve) in the reaction as no. of molecules are decreasing which reduces the randomness of the system.

Thus, $\Delta_r S < 0$

So, (c) option is correct.

S13. Ans.(d)

Free expansion, since $P_{ext} = 0$

$w = -P_{ex}\Delta V = 0$; So, work done = 0

For adiabatic condition, q = 0

From first law of thermodynamics

$\Delta E = q + w$

$\therefore\ \Delta E = 0$

Internal energy of an ideal gas is a function of temperature

$\therefore$ If internal energy remains constant

$\therefore\ \Delta T = 0$

Thus (d) option is correct.

S14. Ans.(d)

For spontaneous reaction, $\Delta_r G$ must be less than zero

So, $\Delta_r G = \Delta_r H - T\Delta_r S < 0$

or, $\Delta_r S > \frac{\Delta_r H}{T}$

$> \frac{30,000}{450} = 66.67\ JK^{-1}\ mol^{-1}$

For reaction to be spontaneous the value of $\Delta_r S$ must be greater than 66.67 J that is 70 $JK^{-1}\ mol^{-1}$

S15. Ans.(d)

$H_2(g) + Br_2(g) \to 2HBr(g)$;

$\Delta H = -109 kJ\ mol^{-1}$

ΔH= sum of bond energy of reactants – sum of bond energy of products

$(BE_{H-H}) + (BE_{Br-Br}) - 2(BE_{H-Br})$

$-109 = (435) + (192) - 2(BE_{H-Br})$

$BE_{H-Br} = 368\ kJ\ mol^{-1}$

S16. Ans.(a)

$W_{irr} = -P_{ext}\Delta V = -2\ bar \times (0.25 - 0.1)L$

= –0.30 L-bar = –0.30 × 100 J

⇒ –30 J

S17. Ans.(d)

(i) For evaporation of water $\Delta S > 0$;

ΔS = +ve

(ii) In expansion of gas at constant T

$\Delta S > 0$; ΔS = +ve

(iii) Solid → gas ; ΔS = +ve

(iv) $2Hg \to H_2(g), \Delta S < 0$ $(\because \Delta n_g < 0)$

No. of particle decreases from reactant to product side.

S18. Ans.(d)

The reaction for $\Delta_f H°(XY)$

$\frac{1}{2}X_2(g) + \frac{1}{2}Y_2(g) \to XY(g)$

Bond energies of X_2, Y_2 and XY are x, $\frac{x}{2}$, x respectively

$\therefore\ \Delta H = \left(\frac{x}{2} + \frac{x}{4}\right) - \frac{x}{2} = -200$

On solving, we get

$\Rightarrow\ \frac{x}{4} = -200$

⇒ x = 800 kJ/mole

S19. Ans.(d)

$\Delta U = q + W$

For adiabatic process, q = 0

$\therefore\ \Delta U = W = -P.\ \Delta V$

= –2.5 atm × (4.50 – 2.50) L

= –2.5 × 2 L-atm ⇒ –5 × 101.3 J

= –506.5 J ⇒ –505 J

S20. Ans.(c)

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

For spontaneous reaction, ΔG must be negative. For negative value of ΔG, ΔH should be less than TΔS. It is possible when T > 425 K.

ΔH = 35.5 KJ/mol = 35500 J/mol

TΔS = (425) (83.6) = 35530

TΔS > ΔH

Hence, the reaction is spontaneous.

S21. Ans.(c)

The molecule having more number of bonds have largest value of entropy.

$\therefore$ C_2H_6 have large value of entropy.

S22. Ans.(b)

Under isothermal reversible conditions, the term "free-energy" in thermodynamics signifies that "No expansion work done by the system".

$-\Delta G_{system} = W_{non-expansion}$

S23. Ans.(d)

For isothermal process

$\Delta S = nRln\frac{P_i}{P_f}$

S24. Ans.(d)

For any spontaneous process, ΔG = (–ve) and ΔS = (+ve) that is increase in entropy. So, at ΔH < 0 and ΔS > 0 at all temperatures according the reaction will be spontaneous.

S25. Ans.(b)

$\frac{dlnP}{dT} = \frac{-\Delta H_v}{RT^2}$ (this is Clausis Clapeyron equation).

S26. Ans.(a)

$C(s) + O_2 \rightarrow CO_2(g) \quad \Delta H = -393.5\ KJ/mole$

$\Delta H_r = \Delta H f_{CO_2}$

So heat of formation of CO_2

= 393.5 kJ/mole

or we can say

Heat released on formation of 44 gm CO_2 = 393.5 kJ

Heat released on formation of 35.2 gm of $CO_2 = \frac{393.5}{44} \times 35.2 = 314.8$

$\Rightarrow \ \simeq 315$ kJ

S27. Ans.(a)

$X_2O_4 \rightleftharpoons 2XO_2(g)$

$\Delta H = \Delta U + \Delta n_g R \times T$

$\Delta n_g =$ No. of moles of gaseous product – No. of moles of gaseous reactant

= 2 – 0

= 2

$\Delta H = 2.1 + \frac{2(2)\times 300}{1000}$

$= 2.1 + 1.2 \Rightarrow 3.3\ K\ Cal$

$\Delta G = \Delta H - T\Delta S = 3300 - 300 \times 20$

$= -2700$ cal = -2.7 kcal