



DPP SOLUTION

- **Subject – Physical Chemistry**
- **Chapter – Thermodynamics and Thermochemistry**

DPP No.- 08



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Question-



The equilibrium constant for a reaction is 100, what will be the value of ΔG° ?

$R = 8.314 \text{ JK}^{-1} \text{ mol}^{-1}$, $T = 300 \text{ K}$:

① -11488 kJ

~~②~~ -11.488 kJ

③ ~~X~~ -12 kJ

④ ~~X~~ -12000 kJ

$\frac{22011}{20}$

$K = 100$
 $T = 300 \text{ K}$

$$\begin{aligned}\Delta G^\circ &= -2.303 R T \log K \\ &= \frac{-2.303 \times 25 \times 300}{1000} \log 10 \\ &= -230.3 \times 25 \times 2 \log 10 \\ &= \frac{-230.3 \times 50}{1000} \text{ kJ} \\ &= -11.48 \text{ kJ}\end{aligned}$$

Ans. (2)

Question-



For the reaction: $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$; $\Delta U = 2.0 \text{ Kcal}$, $\Delta S = 50 \text{ cal K}^{-1}$ at 300 K .
Calculate ΔG .

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

$$\begin{aligned} \Delta H &= \Delta U + \Delta n_g RT \\ &= 2000 + 1 \times 2 \times 300 \\ &= 2600 \end{aligned}$$

$$\Delta G = 2600 - 300 \times 50$$

$$= 2600 - 15000$$

$$= -12400 \text{ Cal}$$

$$= \frac{-12400}{1000} \text{ Kcal}$$

$$= -12.4 \text{ Kcal}$$

① +12.4 kcal

② -12.4 kcal

③ -6.4 kcal

④ +6.4 kcal

Ans. (2)

Question-



Which relation is incorrect:

1 $\underline{\Delta G} = -T \underline{\Delta S_{\text{Total}}}$ ✗

2 $\underline{\Delta G^\circ} = \underline{-2.303 RT \log K}$ ✗

3 $\underline{\Delta H} = \underline{\Delta U + \Delta n_g RT}$ ✗

~~4 $W_{\text{useful}} = \Delta H$~~ ✗

Ans. (4)

Question-



Which is always correct at equilibrium

① ~~$\Delta G^\circ = 0$~~

✓ ② $\Delta G = 0$

③ ~~$\Delta S_{\text{system}} = 0$~~

④ ~~$\Delta E = 0$~~

Ans. (2)

Question-



$$\log 2 = 0.3$$

$$\log 3 = 0.48$$

For the water gas reaction: $\text{C(s)} + \text{H}_2\text{O(g)} \rightleftharpoons \text{CO(g)} + \text{H}_2\text{(g)}$. The standard Gibbs energy of reaction (at 1000 K) is -8.1 kJ mol^{-1} . Value of equilibrium constant is-

① 2.6 $\Delta G^\circ = -8.1 \text{ kJ/mol}$ $T = 1000 \text{ K}$ $K = ?$

② 6.2 $\Delta G^\circ = -2.303 RT \log K$

③ 8.2 $+8100 = +2.3 \times 8.3 \times 1000 \log K$

④ 10 $\frac{8100}{830 \times 23} = \log K$

$$\frac{8100}{19090} = \log K$$

$$\log K = 0.42$$

$$K = \text{antilog } 0.42$$

$$= 10^0 \times 10^{0.42}$$

$$= 2.6$$

Ans. (1)

Question-



For vaporization of water at 1 atm pressure, the values of ΔH and ΔS are $40.63 \text{ kJ mol}^{-1}$ and $108.8 \text{ J K}^{-1} \text{ mol}^{-1}$ respectively. The temp. at which Gibbs energy change (ΔG) for this transformation will be zero is:

(1) 273.4 K

(2) 393.4 K

~~(3) 373.4 K~~

(4) 293.4 K



$$= 40630 \text{ J/mol}$$

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

$$\Delta S = 108.8 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$0 = \Delta H - T\Delta S$$

$$T = ?$$

$$\Delta H = T\Delta S$$

$$\Delta G = 0$$

$$T = \frac{\Delta H}{\Delta S} = \frac{40630}{108.8}$$

$$= 373.4 \text{ K}$$

Ans. (3)

Question-



Calculate ΔG° for the conversion of oxygen to ozone, $\frac{3}{2} \text{O}_2(\text{g}) \rightarrow \text{O}_3(\text{g})$ at 298 K. If K_p for this conversion is 3×10^{-29} .

(1) $+175.3 \text{ kJ mol}^{-1}$ $\Delta G^\circ = -2.303 RT \log K_p$
~~(2) $+162.7 \text{ kJ mol}^{-1}$~~ $= \frac{-2.303 \times 25 \times 298}{3} \log 3 \times 10^{-29}$
 (3) $-162.7 \text{ kJ mol}^{-1}$ $= \frac{-2.303 \times 25 \times 298}{3} (\log 3 - 29 \log 10)$
 (4) $-140.5 \text{ kJ mol}^{-1}$ $= \frac{-2.303 \times 25 \times 298}{3} (0.48 - 29)$
 $= \frac{+2.303 \times 25 \times 298 \times 28.52}{3000} = \frac{489327}{3000} \approx 163 \text{ kJ/mol}$

Ans. (2)

Question-



Identify the correct statement regarding entropy.

- ① ~~At 0°C , the entropy of a perfectly crystalline substance is taken to be zero.~~
- ② ~~At absolute zero temp. the entropy of a perfectly crystalline solid is positive.~~
↓
ok at -273°C
- ③ ~~At absolute zero temp. the entropy of all crystalline substance is taken to be zero.~~
- ④ At absolute zero temp. the entropy of a perfectly crystalline solid is taken to be zero.

Ans. (4)

Question-



The following data is known about the melting of a compound AB. $\Delta H = 9.2 \text{ kJ mol}^{-1}$. $\Delta S = 0.008 \text{ kJ K}^{-1} \text{ mol}^{-1}$. Its melting point is:

- ① 736 K $\Delta H = 9.2 \text{ kJ/mol}$ $\Delta S = 0.008 \text{ kJ K}^{-1} \text{ mol}^{-1}$
- ② 1050 K $\Delta S = \frac{\Delta H}{T_m}$
- ③ 1150 K $T_m = \frac{\Delta H}{\Delta S} = \frac{9.2}{0.008} = 1150 \text{ K}$
- ④ 1150°C

Ans. (3)

Question-



The entropy change for the conversion of 1 mol of α -tin (at 13°C , 1 atm) to 1 mol of β -tin (13°C , 1 atm) if enthalpy of transition is $2.095 \text{ kJ mol}^{-1}$ is

~~1~~ $7.32 \text{ J mol}^{-1} \text{ K}^{-1}$

2 $14.62 \text{ J K}^{-1} \text{ mol}^{-1}$

3 $56.3 \text{ J mol}^{-1} \text{ K}^{-1}$

4 0



$$\Delta H = 2.095 \text{ kJ/mol} \\ = 2095 \text{ J/mol}$$

$$T = 273 + 13 \\ = 286 \text{ K}$$

$$\Delta S = \frac{\Delta H}{T} = \frac{2095}{286} = 7.32 \text{ J K}^{-1} \text{ mol}^{-1}$$

Ans. (1)



Thank

You...

