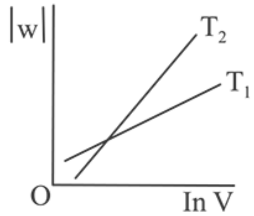
JEE Mains 2019 Chapter wise Question Bank

Chemical Thermodynamics - Questions

Q1

Consider the reversible isothermal expansion of an ideal gas in a closed system at two different temperatures T_1 and T_2 ($T_1 < T_2$). The correct graphical depiction of the dependence of work done (w) on the final volume (V) is:

(1)


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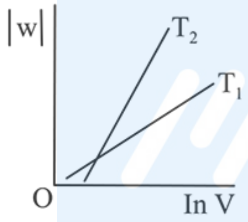

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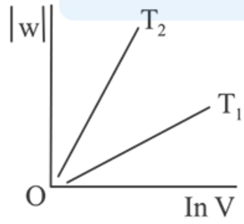

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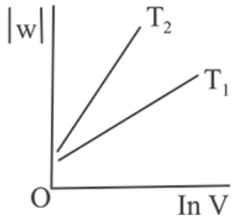

9 Jan Morning

Q2

The entropy change associated with the conversion of 1 kg of ice at 273 K to water vapours at 383 K is: (Specific heat of water liquid and water vapour are 4.2 kJ K^{-1} kg^{-1} and 2.0 kJ K^{-1} kg^{-1}; heat of liquid fusion and vapourisation of water are 334 kJ kg^{-1} and 2491 kJ kg^{-1}, respectively).

(log 273 = 2.436, log 373 = 2.572, log 383 = 2.583)

(1) 7.90 kJ kg^{-1} K^{-1}
(2) 2.64 kJ kg^{-1} K^{-1}
(3) 8.49 kJ kg^{-1} K^{-1}
(4) 9.26 kJ kg^{-1} K^{-1}

9 Jan Evening

Q3

A process has $\Delta H = 200$ J mol^{-1} and $\Delta S = 40$ $JK^{-1}mol^{-1}$. Out of the values given below, choose the minimum temperature above which the process will be spontaneous:

(1) 20 K
(2) 12 K
(3) 5 K
(4) 4 K

10 Jan Morning

Q4

The process with negative entropy change is:

(1) Dissociation of $CaSO_4(s)$ to $CaO(s)$ and $SO_3(g)$
(2) Sublimation of dry ice
(3) Dissolution of iodine in water
(4) Synthesis of ammonia from N_2 and H_2

10 Jan Evening

Q5

An ideal gas undergoes isothermal compression from 5 m^3 to 1 m^3 against a constant external pressure of 4 Nm^{-2}. Heat released in this process is used to increase the temperature of 1 mole of Al. If molar heat capacity of Al is 24 J mol^{-1} K^{-1}, the temperature of Al increases by:

(1) $\frac{3}{2}$ K (2) 2 K

(3) $\frac{2}{3}$ K (4) 1 K

11 Jan Morning

Q6

Two blocks of the same metal having same mass and at temperature T_1, and T_2, respectively, are brought in contact with each other and allowed to attain thermal equilibrium at constant pressure. The change in entropy, ΔS, for this process is :

(1) $C_P \ln\left[\frac{(T_1+T_2)^2}{4T_1T_2}\right]$

(2) $2C_P \ln\left[\frac{(T_1+T_2)^{\frac{1}{2}}}{T_1T_2}\right]$

(3) $2C_P \ln\left(\frac{T_1+T_2}{4T_1T_2}\right)$

(4) $2C_P \ln\left[\frac{T_1+T_2}{2T_1T_2}\right]$

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Q7

For the chemical reaction $X \rightleftharpoons Y$, the standard reaction Gibbs energy depends on temperature T (in K) as

$\Delta_r G°$ (in kJ mol^{-1}) $= 120 - \frac{3}{8}T$

The major component of the reaction mixture at T is :

(1) Y if T=300 K (2) Y if T= 280 K

(3) X if T=350 K (4) X if T=315 K

11 Jan Evening

Q8

The standard reaction Gibbs energy for a chemical reaction at an absolute temperature T is given by $\Delta G° = A - BT$ where A and B are non-zero constants. Which of the following is true about this reaction?

(1) Exothermic if $B < 0$

(2) Endothermic if $A > 0$

(3) Endothermic if $A < 0$ and $B > 0$

(4) Exothermic if $A > 0$ and $B < 0$

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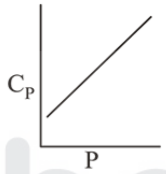
Q9

For a diatomic ideal gas in a closed system, which of the following plots does not correctly describe the relation between various thermodynamic quantities?

(1)



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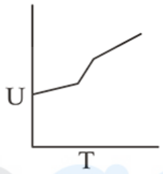

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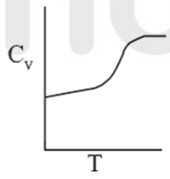

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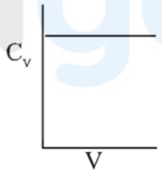

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Q10

Given :

(i) $C(\text{graphite}) + O_2(g) \rightarrow CO_2(g);\ \Delta_r H^0 = x\ \text{kJ mol}^{-1}$

(ii) $C(\text{graphite}) + \frac{1}{2}O_2(g) \rightarrow CO(g);\ \Delta_r H^0 = y\ \text{kJ mol}^{-1}$

(iii) $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g);\ \Delta_r H^0 = z\ \text{kJ mol}^{-1}$

Based on the above thermochemical equations, find out which one of the following algebraic relationships is correct?

(1) $x = y + z$ (2) $z = x + y$

(3) $y = 2z - x$ (4) $x = y - z$

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Q11

For silver, $C_p(JK^{-1}\ mol^{-1}) = 23 + 0.01T$. If the temperature (T) of 3 moles of silver is raised from 300 K to 1000 K at 1 atm pressure, the value of ΔH will be close to:
(1) 62 kJ (2) 16 kJ (3) 21 kJ (4) 13 kJ

8 April Morning

Q11

Which one of the following equations does not correctly represent the first law of thermodynamics for the given processes involving an ideal gas ? (Assume non-expansion work is zero)
(1) Cyclic process : $q = -w$
(2) Adiabatic process : $\Delta U = -w$
(3) Isochoric process: $\Delta U = q$
(4) Isothermal process: $q = -w$

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Q12

5 moles of an ideal gas at 100 K are allowed to undergo reversible compression till its temperature becomes 200 K. If $C_V = 28\ J\ K^{-1}\ mol^{-1}$, calculate ΔU and ΔpV for this process. ($R = 8.0\ J\ K^{-1}\ mol^{-1}$)
(1) $\Delta U = 14$ kJ; $\Delta(pV) = 18$ kJ
(2) $\Delta U = 14$ kJ; $\Delta(pV) = 0.8$ kJ
(3) $\Delta U = 14$ kJ; $\Delta(pV) = 4$ kJ
(4) $\Delta U = 14$ kJ; $\Delta(pV) = 8.0$ kJ

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Q13

Among the following, the set of parameters that represents path functions, is:
(1) q + w (2) q
(3) w (4) H – TS
(1) (B) and (C) (2) (B), (C) and (D)
(3) (A) and (D) (4) (A), (B) and (C)

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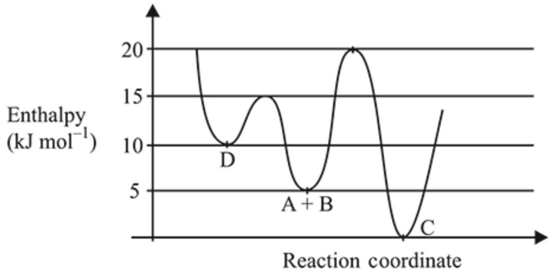
Q14

Consider the given plot of enthalpy of the following reaction between A and B, $A + B \rightarrow C + D$.
Identify the incorrect statement.



(1) Activation enthalpy to form C is 5kJ mol^{-1} less than that to form D.
(2) C is the thermodynamically stable product.
(3) D is kinetically stable product.
(4) Formation of A and B from C has highest enthalpy of activation.

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Q14

During compression of a spring the work done is 10 kJ and 2 kJ escaped to the surroundings as heat. The change in internal energy, ΔU (in kJ) is:
(1) –12 (2) –8 (3) 8 (4) 12

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Q15

A process will be spontaneous at all temperatures if :
(1) $\Delta H < 0$ and $\Delta S < 0$ (2) $\Delta H > 0$ and $\Delta S < 0$
(3) $\Delta H < 0$ and $\Delta S > 0$ (4) $\Delta H > 0$ and $\Delta S > 0$

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Q16

The difference between ΔH and ΔU $(\Delta H - \Delta U)$, when the combustion of one mole of heptane (I) is carried out at a temperature T, is equal to :
(1) – 4 RT (2) – 3 RT (3) 4RT (4) 3 RT

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Q17

An ideal gas is allowed to expand from 1 L to 10 L against a constant external pressure of 1 bar. The work done in kJ is :
(1) – 9.0 (2) + 10.0 (3) – 0.9 (4) – 2.0

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Q18

Enthalpy of sublimation of iodine is 24 cal g^{-1} at 200°C. If specific heat of $I_2(s)$ and $I_2(vap)$ are 0.055 and 0.031 cal $g^{-1}K^{-1}$ respectively, then enthalpy of sublimation of iodine at 250°C in cal g^{-1} is :

(1) 2.85 (2) 5.7 (3) 22.8 (4) 11.4

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Q19

The INCORRECT match in the following is :

(1) $\Delta G^0 < 0, K > 1$ (2) $\Delta G^0 = 0, K = 1$

(3) $\Delta G^0 > 0, K < 1$ (4) $\Delta G^0 < 0, K < 1$

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JEE Mains 2019 Chapter wise Question Bank

Chemical Thermodynamics - Answers

Q1

$$w = -nRT \ln \frac{V_2}{V_1}$$

$$\Rightarrow \qquad |w| = nRT \ln \frac{V_2}{V_1}$$

$|w| = \text{nRT} (\ln V_2 - \ln V_1)$

$|w| = \text{nRT} \ln V_2 - \text{nRT } V_1$

$y = mx + c$

So, slope of curve 2 is more than curve 1 and intercept of curve 2 is more negative than curve 1.

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Q2

(4) As we know, $\Delta S_{\text{fus}} = \dfrac{\Delta H_{\text{fus}}}{273}$

$$= \frac{334}{273} = 1.22$$

Also, $\Delta S_{\text{vap}} = \dfrac{\Delta H_{\text{vap}}}{373} = \dfrac{2491}{373} = 6.67$

Also, $\Delta S_{\text{water}} = m\ C \ln\left(\dfrac{T_2}{T_1}\right)$

$$= 4.2 \times \ln\left(\frac{373}{273}\right) = 1.31$$

Now, $\Delta S_{\text{vap}} = m\ C \ln\left(\dfrac{T_2}{T_1}\right)$

$$= 2 \times \ln\left(\frac{383}{373}\right) = 0.05$$

$\therefore$ Total entropy change

$\Delta S = 9.26 \text{ kJ kg}^{-1} \text{ K}^{-1}$

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Q3

(3) $\Delta H = 200 \text{ J mol}^{-1}$

$\Delta S = 40 \text{ JK}^{-1} \text{ mol}^{-1}$

For spontaneous reaction,

$\Delta G < 0$

$\Delta H - T\Delta S < 0; \Delta H < T\Delta S$

$\dfrac{\Delta H}{\Delta S} < T; \quad \dfrac{200}{40} < T$

$5 < T$

So, minimum temperature is 5 K.

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Q4

(4) In the process of synthesis of ammonia from N_2 and H_2, number of moles decreases which implies that the change in entropy will be negative.

$$N_2(g) + 3H_2(g) \times 2NH_3(g)$$

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Q5

(3) We know that,

$w = -P_{\text{ext}}(V_f - V_i)$

$w = -4 \text{ Nm}^{-2}(1 - 5)\text{m}^3$

$w = 16 \text{ Nm} \Rightarrow 16 \text{ J}$

For isothermal compression,

$\Delta U = q + w$

$\Rightarrow q = -w = -16 \text{ J} \ (\because \Delta U = 0 \text{ for isothermal process})$

From calorimetry,

Heat given $= n\ C\ \Delta T$

So, $16 = \dfrac{1 \times 24 \text{ J} \times \Delta T}{\text{mol} K}$

$\therefore$ Change in temperature, $\Delta T = \dfrac{2}{3}$ K

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Q6

(1) Final temperature $= \dfrac{T_1 + T_2}{2}$, let $T_2 > T_1$

$$\therefore \quad dS = \frac{dq}{T} = \frac{C_p dT}{T}$$

$$\therefore \quad \Delta S = C_p \ln\left(\frac{T_f}{T_i}\right)$$

$$\therefore \quad \Delta S_{Total} = C_p \ln\left(\frac{T_1 + T_2}{2T_1}\right) + C_p \ln\left(\frac{T_1 + T_2}{2T_2}\right)$$

$$= C_p \ln\left[\frac{(T_1 + T_2)^2}{4T_1T_2}\right]$$

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Q7

(4) At 315K; $\Delta G° = 120 - \dfrac{3}{8}T$

$\Delta G° = 120 - 118.125 =$ positive

Since $\Delta G°$ is positive then $K_{eq} < 1$.

So $\dfrac{[Y]}{[X]} < 1$.

$\therefore [X] > [Y]$

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Q8

(3) $MgO (s) + C(s) \rightarrow Mg (s) + CO (g)$

For a reaction to be spontaneous

$\Delta G < 0$

$\Delta H° - T\Delta S° < 0$

$$\Rightarrow \quad T > \frac{\Delta H°}{\Delta S°}$$

$$T > \frac{491.1 \times 1000}{198}$$

$T > 2480.3$ K

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Q9

(2) $\Delta G° = A - BT$

A and B are non-zero constants

$\because \quad \Delta G° = \Delta H° - T\Delta S° = A - BT$

$\therefore$ Reaction will be endothermic if $\Delta G° > 0$.

Hence, A > O and B < O.

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Q10

(1) $(\Delta T_f)_x = (\Delta T_f)_y$

$K_f.m_x = K_f.m_y$

$$\frac{4 \times 100}{A \times 96} = \frac{12 \times 100}{m_y \times 88}$$

$m_y = 3.27$ A; y ≈ 3 A

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Q11

(1) **Given:** $n = 3$

$T_1 = 300$; $T_2 = 1000$

$C_p = 23 + 0.01T$

The relation between ΔH and C_p is

$$\Delta H = \int_{T_1}^{T_2} nC_p dT \qquad \text{... (i)}$$

After putting all variable values in eq. (i)

$$\Delta H = n \int_{300}^{1000} (23 + 0.01T)\, dT$$

$$= 3\left[23T + \frac{0.01T^2}{2}\right]_{300}^{1000}$$

$= 3[16100 + 4550] = 3 \times 20650 = 61950$ J

$= 61.95$ kJ

≈ 62 kJ

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Q11

(2) From first law of thermodynamics, $\Delta U = q + w$

For adiabatic process, $q = 0$

$\therefore \Delta U = w$

For isothermal process, $\Delta U = 0$

For cyclic process, $\Delta U = 0$

For isochoric process, $w = 0$

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Q12

(3) $\Delta U = n\,C_v\,\Delta T = 5 \times 28 \times 100 = 14$ kJ

$\Delta(PV) = nR\,(T_2 - T_1) = 5 \times 8 \times 100 = 4$ kJ

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Q13

(1) We know that heat and work are not state functions but $q + w = \Delta U$ is a state function. $H - TS$ (i.e. G) is also a state function.

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Q14

(1) As we can see from the graph that activation enthalpy to form D from A + B is $15 - 5 = 10$ kJ mol^{-1}, whereas, to form C from A + B is $20 - 5 = 15$ kJ mol^{-1}. Therefore, activation enthalpy to form C is 5 kJ more than that to form D.

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Q14

(3) $w = 10$ kJ

$q = -2$ kJ

$\Delta U = q + w = -2 + 10 = 8$ kJ

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Q15

(3) A reaction is spontaneous if ΔG_{sys} is negative.

$\Delta G_{sys} = \Delta H_{sys} - T\Delta S_{sys}$

A reaction will be spontaneous at all temperatures if ΔH_{sys} is negative and ΔS_{sys} is positive.

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Q16

(1) $C_7H_{16}(l) + 11O_2(g) \xrightarrow{\Delta} 7CO_2(g) + 8H_2O(l)$

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

Δn_g = no. of moles of product in gaseous state - no. of moles of reactant in gaseous state.

$\because \Delta n_g = -4$

$\therefore \Delta H - \Delta U = -4RT$

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Q17

(3) $w = -P\Delta V$

$= -(1 \text{ bar}) \times (9 \text{ L})$

$= -(10^5 \text{ Pa}) \times (9 \times 10^{-3})\text{ m}^3$

$= -9 \times 10^2$ N.m

$= -900$ J $= -0.9$ kJ

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Q18

(3) $I_2(s) \longrightarrow I_2(g)$

Heat of reaction depend upon temperature i.e., it varies with temperature, as given by Kirchoff's equation,

$$\Delta H_{T_2} = \Delta H_{T_1} + \int_{T_1}^{T_2} \Delta C_p dT$$

where $\Delta C_p = C_p$ of product $- C_p$ of reactant

$\therefore \quad \Delta C_p = 0.031 - 0.055 = -0.024$ cal/g

Now, $\Delta H_{T_2} - \Delta H_{T_1} = \Delta C_p\left(T_2 - T_1\right)$

$\Delta H_{(250)} - \Delta H_{(200)} = -0.024\,(523 - 473)$

$\Delta H_{(250)} = 24 - 50 \times 0.024 = 22.8$ cal/g

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Q19

(4) $\Delta G° = -RT \ln K$

$\therefore$ If $K > 1$ then $\Delta G° < 0$

If $K < 1$ then $\Delta G° > 0$

If $K = 1$ then $\Delta G° = 0$

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