Thermodynamics

EXAM DRILL

ANSWERS

1. **(d) :** In adiabatic process Q = constant, then $\Delta Q = 0$ and specific heat, $s = \dfrac{\Delta Q}{m\Delta T} = \dfrac{0}{m\Delta T} = 0$

2. **(d) :** The zeroth law of thermodynamics states that if two systems A and B are in thermal equilibrium with a third system C, then A, B and C will be in thermal equilibrium with each other.

3. **(d) :** All of these are thermodynamic variables.

4. **(a) :** The internal energy of an ideal gas depends only upon temperature of gas.

5. **(a) :** By first law of thermodynamics,
$dQ = dU + dW$
As $dW = 0$ and $dQ < 0$, so $dU < 0$
But for an ideal gas, $U \propto T$; $\quad \therefore \quad dT < 0$
Hence temperature of the gas decreases.

6. **(a)**

7. **(c) :** In the given P-V diagram pressure remains constant although volume increases, hence the process is an isobaric process.

8. **(c) :** Adiabatic process is the process in which no exchange of heat energy takes place between the gas and the surroundings, i.e., $\Delta Q = 0$.

9. **(d) :** $\Delta U = \Delta Q - \Delta W$
The internal energy is independent of path.

10. (c) : A thermodynamic process in which volume remains constant is known as isochoric process.

11. **(a) :** In adiabatic process, no heat transfer takes place between gas and surrounding i.e., $\Delta Q = 0$

From definition of specific heat, $C = \dfrac{\Delta Q}{m\Delta T} = 0$

Again, for isothermal process, $\Delta T = 0$; $\therefore$ $C = \dfrac{\Delta Q}{m\Delta T} = \infty$

12. **(b) :** When we heat the gas at constant volume, the heat is supplied to raise the temperature of gas only. When we heat the gas at constant pressure, the volume of the gas apart from temperature of the gas rises. When the gas expands, it does some external work for increasing the volume. So in this case, the heat is supplied.
(a)   to raise the temperature of the gas.
(b)   to do the mechanical work for expansion.
As a result more amount of heat is required to increase the temperature of the gas through the same amount when the pressure is kept constant. Hence C_P is always greater than the value of C_V..

13. **(c) :** In an adiabatic process, no exchange of heat is possible i.e., $dQ = 0$. $\therefore$ $dU = -dW$.

14.(i) (d) : According to first law of thermodynamics, $\delta Q = \delta U + \delta W$
Along the path adc
change in internal energy,
$\delta U_1 = \delta Q_1 - \delta W_1 = 50\,J - 20\,J = 30\,J$
Along the path abc
change in internal energy, $\delta U_2 = \delta Q_2 - \delta W_2 = 36\,J - \delta W_2$
As change in internal energy is path independent,
$\therefore \quad \delta U_1 = \delta U_2 \quad$ or $\; 30\,J = 36\,J - \delta W_2 \;$ or $\; \delta W_2 = 6\,J$

(ii) **(d) :** $\Delta U = \Delta Q - \Delta W$
The internal energy is independent of path.

(iii) **(d) :** According to first law of thermodynamics
$\Delta Q = \Delta U + \Delta W \;$ or $\Delta U = \Delta Q - \Delta W$
Here, $\Delta Q = 35\,J$, $\Delta W = -15\,J$
$\therefore \quad \Delta U = 35\,J - (-15\,J) = 50\,J$
Note : ΔW is negative because work is done on the system.

15.(i) (d) : Suppose final temperature of gases be T.
Heat rejected by gas in lower compartment
$\qquad = nC_V \Delta T \qquad$ (as volume is fixed)

$\qquad = 2 \times \dfrac{3}{2} R(700 - T) = 3R(700 - T) \,...(i)$
Heat received by gas in upper compartment,
$\qquad = nC_P \Delta T$ (as piston is movable,
$\qquad\qquad$ so pressure is constant)
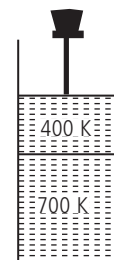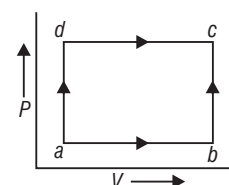$\qquad = 2 \times \dfrac{7}{2} R(T - 400)$ $\qquad\qquad$ ...(ii)
Equating equation (i) and (ii), we get
$\qquad 2 \times \dfrac{3}{2} R(700 - T) = 2 \times \dfrac{7}{2} R(T - 400)$
or $3(700 - T) = 7(T - 400)$
or $2100 - 3T = 7T - 2800$
or $10T = 4900$ or $T = 490$ K

(ii) **(d) :** $\Delta Q = 3R \times (700 - 490) = 210$ R

(iii) **(b) :** From equation (ii)

$\Delta Q = 7R(490 - 400) = 630$ R

16. From first law of thermodynamics

$\Delta Q = \Delta U + \Delta W$.

$\Rightarrow \quad \Delta Q = nC_v \Delta T + \Delta W$ \qquad ($\because \Delta U = nC_v\Delta T$)

If system has constant temperature in spite of heat supplied then $\Delta T = 0$ and $\Delta Q = \Delta W$. It means heat supplied (ΔQ) to the system is used in doing work (ΔW) against the surrounding.

17. Yes, the gas will do work at the expense of its internal energy.

18. Variables which indicate the size of the system are called extensive variables.

OR

Two objects in thermal equilibrium have the same temperature.

19. When solids and liquids are heated, there is only a slight change in their volume and as such they possess only one specific heat, *i.e.,* specific heat at constant volume. But in case of gases, pressure and volume both change and as such they possess two principal specific heats; one at constant pressure and one at constant volume.

20. A definite quantity of matter bounded by some closed surface is called a system and a collection of very large number of molecules of matter which are so arranged that these possess certain values of pressure, volume and temperature forms a thermodynamic system.

21. The total work done by the gas from D to E to F is equal to the area of ΔDEF.

$\therefore$ Area of $\Delta DEF = \dfrac{1}{2} DF \times EF$

Here, DF = change in pressure = $600 - 200 = 400$ N m^{-2}
and EF = change in volume = 7 m$^3 - 3$ m$^3 = 4$ m^3

$\therefore$ Area of $\Delta DEF = \dfrac{1}{2} \times 400 \times 4 = 800$ J

Thus, the total work done by the gas from D to E to F is 800 J.

22. Limitations of the second law of thermodynamics :

(i) The second law of thermodynamics cannot be proved directly. But its validity has not been contradicted by any machine designed so far.

(ii) It is applicable to a cyclic process in which the system returns to its original state after a complete cycle of changes.

(iii) It does not tell anything about the entropy.

23. Loss of mass, $m = 5$ kg

1 kg of fat is burnt on expending 7000 kilo calories.

7000×10^3 cal = $(7000 \times 10^3 \times 4.2)$J = 29.4×10^6 J.

$\therefore$ Work done to burn 5 kg of fat

$= 5 \times 29.4 \times 10^6$ J = 147×10^6 J

Work done towards burning of fat in one trip

$= mgh + \dfrac{1}{2} mgh$

$= \dfrac{3}{2} mgh = \dfrac{3}{2} \times 60 \times 10 \times 10 = 9 \times 10^3$ J

Number of trips required,

$N = \dfrac{147 \times 10^6}{9 \times 10^3} = 16.3 \times 10^3$ times

OR

Slope is given by dP/dV

For isothermal process, $PV = K$

Differentiating both sides, we get

$P.dV + V.dP = 0$, $V.dP = -P.dV$

$\dfrac{dP}{dV} = -P/V$ \hfill ...(i)

For adiabatic change $PV^\gamma = K$

Differentiating both sides

$\gamma PV^{\gamma-1}dV + V^\gamma dP = 0$

$dP = -(\gamma P/V)dV$ \hfill ...(ii)

$\dfrac{dP}{dV} = -\gamma \dfrac{P}{V}$

From eqn. (i) and (ii)

We find that slope (dP/dV) of an adiabatic curve is γ times the slope of (dP/dV) of an isothermal curve, as γ is greater than one for an isothermal process.

24. Heat supplied to the gas, $\Delta Q = 1500$ J

Work done by the gas,

$\Delta W = P\Delta V = (2.1 \times 10^5$ N m$^{-2})(2.5 \times 10^{-3}$ m$^3)$

$= 5.25 \times 10^2$ N m = 525 J

According to first law of thermodynamics,

$\Delta Q = \Delta W + \Delta U$

$\therefore$ $\Delta U = \Delta Q - \Delta W = 1500$ J $- 525$ J $= 975$ J

25. C_P is greater than C_V. When a gas is heated at constant volume, all the heat supplied is used to increase the temperature or internal energy of the gas.

26. (a) Here $V_2 = \dfrac{1}{4} V_1$, $P_1 = 75$ cm of Hg,
$T_1 = 50 + 273 = 323$ K

(i) When the gas is compressed slowly, the process is isothermal.

$\therefore$ $P_1 V_1 = P_2 V_2$ or $75 \times V_1 = P_2 \times \dfrac{1}{4} V_1$

or $P_2 = 75 \times 4 = 300$ cm of Hg

As the process is isothermal, so $T_2 = 50$ °C

(ii) When the gas is compressed suddenly, the process is adiabatic.

$\therefore$ $P_1 V_1^\gamma = P_2 V_2^\gamma$

or $P_2 = P_1 \left(\dfrac{V_1}{V_2}\right)^\gamma = 75 \left(\dfrac{V_1}{V_1/4}\right)^{1.5} = 75 \times 4^{1.5}$

$= 75 \times 4 \times 4^{1/2} = 75 \times 4 \times 2 = 600$ cm of Hg

Also, $T_1 V_1^{\gamma-1} = T_2 V_2^{\gamma-1}$

or $T_2 = T_1 \left(\dfrac{V_1}{V_2}\right)^{\gamma-1}$

$= 323 \times (4)^{0.5} = 323 \times 2 = 646$ K $= 373°$C.

(b) Work done in an isothermal process

$W = nRT \ln\left(\dfrac{V_2}{V_1}\right)$

Here, $n = 1$, $T = 500$ K, $V_1 = 2V$ and $V_2 = 4V$

$\therefore \quad W = 1 \times R \times 500 \ln\left(\dfrac{4V}{2V}\right) = 600R \ln 2$

27. The amount of heat required at constant pressure is

$\Delta Q = nC_P \Delta T \quad \therefore \quad C_P = \dfrac{\Delta Q}{n\Delta T} = \dfrac{70}{2(35-30)} = 7$ cal

As $C_P - C_V = R$

$\therefore \quad C_V = C_P - R = 7 - 2 = 5$ cal mol^{-1} K^{-1}

The amount of heat required at constant volume is

$\Delta Q = nC_V \Delta T = 2 \times 5 \times (35 - 30) = 50$ cal

28. In a process $PV^x = $ constant, molar heat capacity is given by

$C = \dfrac{R}{\gamma-1} + \dfrac{R}{1-x}$

As the process is $\dfrac{P}{V} = $ constant, *i.e.* $PV^{-1} = $ constant, therefore,

$x = -1$

For a monatomic gas, $\gamma = \dfrac{5}{3}$

$\therefore \quad C = \dfrac{R}{\frac{5}{3}-1} + \dfrac{R}{1-(-1)} = \dfrac{3}{2}R + \dfrac{R}{2} = 2R$

Heat supplied to the gas is

$\Delta Q = nC(\Delta T) = 1(2R)(2T_0 - T_0) = 2RT_0$

29. Given : $V = KT^{2/3}$ ...(i)

Differentiating both sides,

$\Delta V = \dfrac{2}{3} KT^{-1/3} \Delta T$ (As K is a constant) ...(ii)

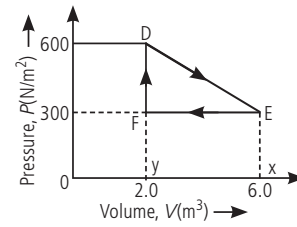According to ideal gas equation for 1 mole

$PV = RT$ or $P = \dfrac{RT}{V}$ ...(iii)

Work done by the gas is

$W = P\Delta V = \dfrac{RT}{V} \dfrac{2}{3} KT^{-1/3}\Delta T$ (Using (ii) and (iii))

$= \dfrac{2}{3} \dfrac{RKT^{2/3}}{V}\Delta T = \dfrac{2}{3}\dfrac{RKT^{2/3}}{KT^{2/3}}\Delta T$ (Using (i))

$= \dfrac{2}{3}R\Delta T = \dfrac{2}{3} \times R \times 40 = 26.6R$

OR

Process from D to E represents expansion

$\therefore$ Work done = [area $DEXYFD$]

$W_{DE} = $ [area of $\Delta DEF + $ area of rectangle $EXYF$]

$W_{DE} = \dfrac{1}{2} \times (EF) \times (FD) + (EF) \times (EX)$

$= \dfrac{1}{2} \times (4\ m^3) \times (300\ N/m^2) + (4\ m^2) \times (300\ N/m^2)$

$\Rightarrow \quad W_{DE} = 1800$ J

Process from E to F represent compression

$\therefore$ Work done = $-$[area $EFYX$]

$W_{DE} = -$[area of rectangle $EFYX$]

$= -(EF) \times (EX) = -(4\ m^3) \times (300\ N/m^2)$

$\Rightarrow \quad W_{DE} = -1200$ J

Process from F to D represents no change in volume

$\therefore$ Work done, $W_{FD} = 0$

Hence, work done in the complete cycle is

$W = W_{DE} + W_{DF} + W_{FD} = 1800$ J $+ (-1200$ J$) + 0$

$\Rightarrow \quad W = 600$ J

30. γ stands for atomicity of gas.

When the piston moves up through a small distance dx, the work done by the gas will be

$dW = PA\ dx = P\ dV$

where A is the cross-sectional area of the piston and $dV = A\ dx$ is the increase in the volume of the gas.

As the gas expands adiabatically and changes from the initial state (P_1, V_1, T_1) to the final state P_2, V_2, T_2), the total work done by the gas will be

$W_{\text{adia}} = \displaystyle\int_{V_1}^{V_2} P\ dV$
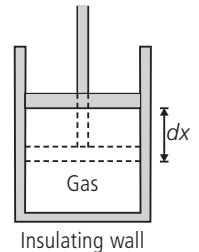

For an adiabatic change

$PV^\gamma = K$ or $P = KV^{-\gamma}$

$\therefore \quad W_{\text{adia}} = \displaystyle\int_{V_1}^{V_2} KV^{-\gamma}\ dV$

$= K\displaystyle\int_{V_1}^{V_2} V^{-\gamma}\ dV = K\left[\dfrac{V^{1-\gamma}}{1-\gamma}\right]_{V_1}^{V_2}$

$= \dfrac{K}{1-\gamma}[V_2^{1-\gamma} - V_1^{1-\gamma}] = \dfrac{1}{\gamma-1}[KV_1^{1-\gamma} - KV_2^{1-\gamma}]$

But $K = P_1V_1^\gamma = P_2V_2^\gamma$

$$\therefore \quad W_{adia} = \frac{1}{\gamma - 1}[P_1V_1^\gamma V_1^{1-\gamma} - P_2V_2^\gamma V_2^{1-\gamma}]$$

$$W_{adia} = \frac{1}{\gamma - 1}[P_1V_1 - P_2V_2]$$

31. (a) No. As the thermosflask is insulated, heat has not been added to the coffee ($\Delta Q = 0$).

(b) Yes, some work is done by the man in shaking the coffee against the forces of viscosity *i.e.,* ΔW is negative.

(c) By first law of thermodynamics, $\Delta Q = \Delta U + \Delta W$.

As $\Delta Q = 0$ and ΔW is negative, so ΔU is positive *i.e.,* internal energy of the coffee increases.

(d) Because of the increase in internal energy of the coffee, the temperature of the coffee will also increase.

32. Consider one mole of an ideal gas. Heat the gas to raise its temperature by ΔT. According to the first law of thermodynamics, the heat supplied ΔQ is used partly to increase the internal energy and partly in doing the work of expansion. That is, $\Delta Q = \Delta U + P\Delta V$

If the heat ΔQ is absorbed at constant volume, then $\Delta V = 0$ and we have

$$C_V = \left(\frac{\Delta Q}{\Delta T}\right)_V = \left(\frac{\Delta U}{\Delta T}\right)_V = \left(\frac{\Delta U}{\Delta T}\right)$$

We have dropped the subscript V because the internal energy U of an ideal gas depends only on its temperature T.

If now the heat ΔQ is absorbed at constant pressure, then

$$C_P = \left(\frac{\Delta Q}{\Delta T}\right)_P = \left(\frac{\Delta U}{\Delta T}\right)_P + P\left(\frac{\Delta V}{\Delta T}\right)_P$$

$$= \left(\frac{\Delta U}{\Delta T}\right) + P\left(\frac{\Delta V}{\Delta T}\right)_P$$

Again, we have dropped the subscript P from the first term because U of an ideal gas depends only on T.

$$\therefore \quad C_P - C_V = P\left(\frac{\Delta V}{\Delta T}\right)_P$$

But for one mole of an ideal gas, $PV = RT$.

Differentiating both sides w.r.t. T for constant pressure P,

$$\frac{\Delta(PV)}{\Delta T} = \frac{\Delta(RT)}{\Delta T}$$

or $\quad P\left(\frac{\Delta V}{\Delta T}\right)_P = R$

Hence, $C_P - C_V = R$

This is the required relation between C_P and C_V. It is also known as Mayer's formula.

33. (a) : Change in internal energy is path independent and depends only on the initial and final states.

As the initial and final states in the three processes are same, therefore,
$\Delta U_1 = \Delta U_2 = \Delta U_3$

Workdone, W = Area under *P-V* graph

As area under curve 1 > area under curve 2 > area under curve 3

$\therefore \quad W_1 > W_2 > W_3$

According to first law of thermodynamics,

$\quad Q = W + \Delta U$

As $W_1 > W_2 > W_3$ and $\Delta U_1 = \Delta U_2 = \Delta U_3$

$\therefore \quad Q_1 > Q_2 > Q_3$

(b) : For the process at constant pressure

$\quad W = RdT$ ...(i)

For the process at constant volume,

$$dQ = C_V dT = \frac{R}{\gamma - 1}dT = \frac{W}{\gamma - 1}$$ (Using (i))

$$\therefore \quad dQ = \frac{W}{\left(\frac{5}{3} - 1\right)} = \frac{3}{2}W$$

OR

Given, cross-sectional area of the piston

$A = 1$ m^2

Heat supplied $= Q$.

Initial Volume $= V_0$, Final volume $= V$

Atmospheric pressure $= P_a$

(a) Since system is in equilibrium so, initial pressure of the system, $P_i = P_a$.

(b) When the gas expands from V_0 to V after heat Q is supplied.

Change in volume of the gas $= V - V_0$

Extension in the spring, $x = \frac{(V - V_0)}{A} = V - V_0$ ($\because A = 1$ m^2)

Force applied by the spring on the piston,

$F = kx = k(V - V_0)$

So, pressure exerted by the extended spring on the piston of unit cross-section,

$$P = \frac{F}{A} = \frac{k}{A}(V - V_0) = k(V - V_0)$$ ($\because A = 1$ m^2)

Hence final pressure of the system,

$P_f = P_a + P = P_a + k(V - V_0)$.

(c) Using the first law of thermodynamics,

$Q = U + W$ ...(i)

Here, $U = C_V(T - T_0)$

$$W = P_a(V - V_0) + \frac{1}{2}kx^2$$ ...(ii)

Also, $T_0 = \frac{P_0V_0}{R} = \frac{P_aV_0}{R}$

$$T = \frac{P_fV}{R} = \frac{[P_a + k(V - V_0)]V}{R}$$

$$\therefore U = C_V \left[(P_a + k(V - V_0)) \frac{V}{R} - \frac{P_a V_0}{R} \right] \qquad \text{...(iii)}$$

From (i), (ii) and (iii)

$$Q = \frac{1}{(\gamma - 1)} [(P_a + k(V - V_0))V - P_a V_0] + P_a(V - V_0) + \frac{1}{2} kx^2$$

$$\left(\because \frac{C_V}{R} = \frac{1}{(\gamma - 1)} \right)$$

$$= \frac{1}{(\gamma - 1)} [(P_a + k(V - V_0))V - P_a V_0] + P_a(V - V_0) + \frac{1}{2} k(V - V_0)^2$$

$$(\because \quad x = V - V_0)$$

34. Given, $PV^{1/2} = $ constant $= C$

(a) Work done by the gas

$$\Delta W = \int_{V_1}^{V_2} P dV = \int_{V_1}^{V_2} \frac{C}{\sqrt{V}} dV$$

$$= C \left[\frac{V^{1/2}}{\frac{1}{2}} \right]_{V_1}^{V_2} = 2C(\sqrt{V_2} - \sqrt{V_1})$$

$$\therefore \Delta W = 2P_1 \sqrt{V_1}(\sqrt{V_2} - \sqrt{V_1}) \qquad \text{...(i)}$$

(b) $\because PV = nRT$

$$\Rightarrow \frac{C}{\sqrt{V}} V = nRT \Rightarrow T = \frac{C}{nR} \sqrt{V} \qquad \text{...(ii)}$$

$$\Rightarrow T \propto \sqrt{V}$$

$$\therefore \frac{T_1}{T_2} = \sqrt{\frac{V_1}{V_2}} = \sqrt{\frac{V_1}{2V_1}} = \frac{1}{\sqrt{2}} \qquad \text{(as } V_2 = 2V_1) \qquad \text{...(iii)}$$

(c) Given, $U = \frac{3}{2} RT$

$$\Rightarrow \Delta U = \frac{3}{2} R \Delta T = \frac{3}{2} R(T_2 - T_1)$$

$$\Rightarrow \Delta U = \frac{3}{2} RT_1(\sqrt{2} - 1) \qquad \text{[using eqn (iii)]}$$

From eqn. (i)

$$\Delta W = 2C(\sqrt{V_2} - \sqrt{V_1}) = 2C\sqrt{V_1}(\sqrt{2} - 1) \qquad \text{[using eqn (iii)]}$$
$$\text{(as } PV^{1/2} = C)$$

$$= 2RT_1(\sqrt{2} - 1) \text{ [Using eqn (ii) for one mole of gas]}$$

$\therefore$ From first law of thermodynamics,

$$\Delta Q = \Delta U + \Delta W = \frac{7}{2} RT_1(\sqrt{2} - 1).$$

OR

Heat capacity is given by $C = C_V + \frac{RT}{V} \frac{dV}{dT}$

(i) Given $C = C_V + \alpha T$

So, $C_V + \alpha T = C_V + \frac{RT}{V} \frac{dV}{dT}$ or $\alpha T = \frac{RT}{V} \frac{dV}{dT}$

or $\frac{\alpha}{R} dT = \frac{dV}{V}$

Integrating both sides, we get $\frac{\alpha}{R} T = \ln V + \ln C = \ln VC$

or $V \cdot C = e^{-\alpha T/R}$ or $V \cdot e^{\alpha T/R} = \frac{1}{C} = $ constant

(ii) Given, $C = C_V + \beta V$

and $C = C_V + \frac{RT}{V} \frac{dV}{dT}$ so, $C_V + \frac{RT}{V} \frac{dV}{dT} = C_V + \beta$

or $\frac{RT}{V} \frac{dV}{dT} = \beta V$ or $\frac{dV}{V_2} = \frac{\beta}{R} \frac{dT}{T}$

or $\frac{R}{\beta} V^{-2} = \frac{dT}{T}$

Integrating both sides, we get

$$\frac{R}{\beta} \frac{V^{-1}}{-1} = \ln T + \ln C = \ln T \cdot C$$

So, $\ln T \cdot C = -\frac{R}{\beta V} \Rightarrow T \cdot C = e^{-R/\beta V}$

or $Te^{+R/\beta V} = \frac{1}{C} = $ constant

(iii) $C = C_V + aP$ and $C = C_V + \frac{RT}{V} \frac{dV}{dT}$

So, $C_V + aP = C_V + \frac{RT}{V} \frac{dV}{dT}$ so, $aP = \frac{RT}{V} \frac{dV}{dT}$

or $a\frac{RT}{V} = \frac{RT}{V} \frac{dV}{dT}$ (as $P = \frac{RT}{V}$ for one mole of gas)

or $\frac{dV}{dT} = a$ or $dV = a \, dT$ or $dT = \frac{dV}{a}$

So, $T = \frac{V}{an} + $ constant

As $n = 1, T = \frac{V}{a} + $ constant or $V - aT = $ constant

35. (a) : Heat supplied to the gas at constant pressure is

$$\Delta Q = nC_P \Delta T$$

Change in internal energy, $\Delta U = nC_V \Delta T$

According to first law of thermodynamics

$$\Delta Q = \Delta U + \Delta W$$

$\therefore \quad \Delta W = \Delta Q - \Delta U = nC_P \Delta T - nC_V \Delta T = n\Delta T(C_P - C_V)$

For an ideal gas, $C_P - C_V = R$; $\therefore \Delta W = nR\Delta T$

Here, $n = 1, R = 8.31$ J mol^{-1} K^{-1},

$$\Delta T = T_2 - T_1 = 100°C - 0°C = 100°C = 100 \text{ K}$$

$\therefore \quad \Delta W = 1 \times 8.31 \times 100 = 8.31 \times 10^2 \text{ J}$

(b) : Given process is isobaric.

$\therefore \quad dQ = nC_p dT$; where C_p is specific heat at constant pressure.

or $dQ = n\left(\dfrac{5}{2}R\right)dT$

Also, $dW = PdV = nRdT$ ($\because PV = nRT$)

Required ratio $= \dfrac{dW}{dQ} = \dfrac{nRdT}{n\left(\dfrac{5}{2}R\right)dT} = \dfrac{2}{5}$

OR

(a) The specific heat of a gas depends upon the conditions of volume and pressure under which it is heated. Since a gas can be heated under a large number of such conditions, its specific heat has innumerable values.

For simplicity, we confine ourselves to only two specific heats of a gas, one, when gas is heated at constant volume (C_V) and the other when it is heated at constant pressure (C_P).

$\therefore$ A gas has two principal specific heat capacities

$C_P - C_V = R =$ gas constant

$\dfrac{C_P}{C_V} = \gamma =$ constant of expansion

Their significances are :

(i) The difference between the two specific heats is equal to the amount of heat equivalent to the work performed by the gas during the expansion at constant pressure.

(ii) R is always positive, $C_P > C_V$

(iii) The value of γ depends on the atomicity of a gas and hence explains the nature of gas.

(iv) Knowing γ, we can determine the number of degrees of freedom.