# Thermodynamics

## Question1

**Consider the following.**

**Statement -I Both internal energy $(U)$ and work $(W)$ are state functions.**

**Statement-II During the free expansion of an ideal gas into vacuum, the work done is zero.**

**The correct answer is**

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**Options:**

A.

Both statement-I and statement-II are correct.

B.

Both statement-I and statement-II are not correct.

C.

Statement-I is correct, but statement-II is not correct.

D.

Statement-I is not correct, but statement-II is correct.

**Answer: D**

## Solution:

**Statement-I**

Both internal energy (U) and work (W) are state functions.

- **Internal energy (U)** is indeed a **state function** — it depends only on the state (e.g., temperature, pressure, volume) of the system, not on the path taken.

- **Work (W)**, on the other hand, **is not** a state function; it depends on the path taken during a process. Different paths between the same initial and final states can result in different amounts of work.

✅ So, **Statement-I is incorrect**.

### Statement-II

During the free expansion of an ideal gas into vacuum, the work done is zero.

In free expansion:

- The gas expands into a vacuum.
- There is **no external pressure** acting on the gas.
- Therefore, $W = \int P_{\text{ext}}\ dV = 0$.

✅ Hence, **Statement-II is correct**.

### ✅ Correct answer: Option D

**Statement-I is not correct, but Statement-II is correct.**

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# Question2

## The signs of $\Delta_r H^\circ$ and $\Delta_r S^\circ$ for a reaction to be spontaneous at all temperature respectively are

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### Options:

A.

positive, positive

B.

positive, negative

C.

negative, negative

D.

negative, positive

**Answer: D**

## Solution:

For a reaction to be spontaneous at all temperature, the enthalpy change ( $\Delta H^{\circ}$ ) must be negative (exothermic) and the entropy change ( $\Delta S^{\circ}$ ) must be positive (increase in disorderness).

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# Question3

**The number of extensive and intensive properties in the list given below is respectively, density, enthalpy, mass, temperature, volume, pressure**

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**Options:**

A.

4,2

B.

1,5

C.

2,4

D.

3,3

**Answer: D**

## Solution:

Extensive property : mass, volume and enthalpy.

Intensive property : density, temperature and pressure.

Thus there are 3 extensive property and intensive property.

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# Question4

**One mole of ethanol ( $l$ ) was completely burnt in oxygen to form $\mathrm{CO_2(g)}$ and $\mathrm{H_2O}(l)$. What is the $\Delta_r H^{\circ}$ (in $\mathrm{kJmol^{-1}}$ ) for this reaction?**

**(The standard enthalpy of formation $(\Delta_f H^{\circ})$ of $\mathrm{C_2H_5OH}(l), \mathrm{CO_2}(g)$ and $\mathrm{H_2O}(l)$ is respectively $-277, -393$ and $-286\ \mathrm{kJ\ mol^{-1}}$ )**

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**Options:**

A.

+1921

B.

-1921

C.

+1367

D.

-1367

**Answer: D**

## Solution:

The complete combustion of ethanol is as follows

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$
$$\Delta_r H^\ominus = \Sigma n \Delta_f H^\ominus_{\text{product}} \quad - \Sigma m \Delta_f H^\ominus_{\text{reactant}}$$
$$= [(2 \times -393) + (3 \times -286)] - [(1 \times -277) + (3 \times 0)]$$

$$= -1367 \text{ kJ}$$

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# Question5

## If $\Delta_r H^\ominus$ and $\Delta_r S^\ominus$ are standard enthalpy change and standard entropy change respectively for a reaction, the incorrect option is

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**Options:**

A.

$\Delta_r H^\ominus =$ negative; $\Delta_r S^\ominus =$ positive: spontaneous at all temperatures

B.

$\Delta_1 H^{\ominus}$ = negative; $\Delta_1 S^{\ominus}$ = negative; non-spontaneous at low temperatures

C.

$\Delta_r H^{\ominus}$ = positive; $\Delta_r S^{\ominus}$ = positive; non-spontaneous at low temperatures

D.

$\Delta_r H^{\ominus}$ = negative; $\Delta_r S^{\ominus}$ = negative: spontaneous at low temperatures

**Answer: B**

## Solution:

Statement given in option (b) is incorrect. Its correct form is,

The spontaneity of a reaction with a negative enthalpy change (exothermic) and a negative entropy change (decreasing disorder) depends on temperature.

When $\Delta H$ is negative and $\Delta S$ is negative the reaction is non-spontaneous at high temperature.

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# Question6

## The $\mathrm{C}_p$ of $\mathrm{H}_2\mathrm{O}(l)$ is $75.3 \mathrm{~J} \mathrm{~mol}^{-1} \mathrm{~K}^{-1}$. What is the energy (in J ) required to raise 180 g of liquid water from $10^{\circ}\mathrm{C}$ to $15^{\circ}\mathrm{C}$ ? $(\mathrm{H}_2\mathrm{O}=18\mathrm{u})$

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**Options:**

A.

3.765

B.

3765

C.

753

D.

376.5

**Answer: B**

## Solution:

First, we need to find out how many moles of water we have.

$$n = \frac{m}{M} = \frac{180}{18}$$

$$= 10 \text{ mol}$$
$\Delta T = T_2 - T_1 = 15 - 10 = 5 \text{ K}$

The energy needed to heat the water is given by the formula: $Q = nC_p\Delta T$

Here, $n$ is number of moles (10), $C_p$ is the specific heat ($75.3 \text{ J mol}^{-1} \text{ K}^{-1}$), and $\Delta T$ is the temperature change (5 K).

So, $Q = 10 \times 75.3 \times 5 = 3765 \text{ J}$

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# Question7

### Identify the incorrect statements from the following.

### I. For adiabatic process, $\Delta U = w_{\text{ad}}$

### II. Enthalpy is an intensive property

### III. For the process, $\text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(s)$, the entropy increases

### The correct answer is

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**Options:**

A.

I, II only

B.

I, II, III

C.

I, III only

D.

II, III only

**Answer: D**

## Solution:

The statement given in option II and III are incorrect.

While statement I is correct.

The correct of form statement II and III is,

- Enthalpy is an extensive property.
- For the process, $H_2O(l) \longrightarrow H_2O(s)$ the entropy decreases.

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# Question8

**Enthalpy of formation of $CO_2(g), H_2O(l)$ and $C_6H_{12}O_6(s)$ are $-393, -286$ and $-1170 \text{ kJ mol}^{-1}$ respectively. The quantity of heat liberated when 18 g of $C_6H_{12}O_6(s)$ is burnt completely in oxygen is**

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**Options:**

A.

520 kJ

B.

145 kJ

C.

290 kJ

D.

420 kJ

**Answer: C**

## Solution:

The complete combustion reaction is

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_20(n)$$
$$\Delta H^\circ_{\text{comb}} = \Sigma\Delta H^\circ_{\text{froduct}} - \Sigma\Delta H^\circ_{\text{frectant}}$$
$$= [6 \times (-393) + 6 \times (-286)] - [1 \times (-1170) + 6 \times 0]$$

$$= -2904 \text{ kJ}$$

$$\text{Moles of glucose} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{18 \text{ g}}{180}$$

$$= 0.1 \text{ mol}$$

$$\text{Heat liberated} = \text{Moles } \times \Delta H^\circ_{\text{combustion}}$$

$$= 0.1 \times (-2904)$$

$$\approx -290 \text{ kJ}$$

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# Question9

## For which reaction $\Delta H \neq \Delta U$?

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**Options:**

A.

$H_2(g) + I_2(g) \longrightarrow 2HI(g)$

B.

$2NO(g) \longrightarrow N_2(g) + O_2(g)$

C.

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

D.

$C(s) + O_2(g) \longrightarrow CO_2(g)$

**Answer: C**

## Solution:

The reaction in which $n_{\text{product}} \neq n_{\text{reactant}}$ will have $\Delta H \neq \Delta U$ So, among the given reactions, in reaction,

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$
$$\Delta U \neq \Delta H$$
$$\text{As, } n_{\text{product}} = 2, n_{\text{reactant}} = 4$$
$$\Delta n_g = 2 - 4 = -2$$
$$\Delta H = \Delta U + \Delta n_g RT$$
$$\Delta H = \Delta U - 2RT$$

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# Question10

**At 298 K, $\Delta_r U^\ominus$ and $\Delta_r S^\ominus$ for the following reaction are -10.5 kJ and $+44.1 \mathrm{JK}^{-1}$; $2X(\mathrm{~g})+Y(\mathrm{~g}) \longrightarrow 2Z(\mathrm{~g})$ What is $\Delta_r G^\ominus$ (in kJ ) for this reaction? $\left(R=8.314 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}\right)$**

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**Options:**

A.

+0.164

B.

-26.119

C.

-2.6119

D.

-0.082

**Answer: B**

## Solution:

Given $\Delta_r U^{\circ}=-10.5 \mathrm{~kJ}$

$\Delta_r S^{\circ}=+44.1 \mathrm{~J} / \mathrm{K}$
$\Delta_r G^{\circ}=?$

Using formula

$\Delta H^{\circ}=\Delta U^{\circ}+\Delta(p V)$
$\Delta H^{\circ}=\Delta U^{\circ}+p \Delta V+V \Delta p$

as $\Delta p=0$

So, $\Delta H^{\circ}=\Delta U^{\circ}+p \Delta V$

Now, as $\Delta G^{\circ}=\Delta H^{\circ}-T \Delta S^{\circ}$

$\Delta G^{\circ}=\Delta U^{\circ}+p \Delta V-T \Delta S^{\circ}$

$$
\begin{aligned}
& p \Delta V=-R T \\
\because \quad & p \Delta V=\Delta n_g R T \\
& \Delta n_g=(2-3)=-1
\end{aligned}
$$

So, $\Delta G^{\circ}=\Delta U^{\circ}-R T-T \Delta S^{\circ}$

$$=-10.5-\left(8.314 \times 10^{-3} \times 298\right)-\left(298 \times 44.1 \times 10^{-3}\right)$$

$\approx-26.119 \mathrm{~kJ}$

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# Question11

**Consider the following reaction**

$$A(g)+3B(g)\longrightarrow 2C(g);\Delta H^{\ominus}=-24\text{ kJ}$$

**At $25^{\circ}\mathrm{C}$, if $\Delta G^{\ominus}$ of the reaction is -9 kJ , the standard entropy change (in $\mathrm{JK}^{-1}$ ) of the same reaction at same temperature is**

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**Options:**

A.

-5.33

B.

-50.33

C.

-500.33

D.

-0.533

**Answer: B**

## Solution:

The Gibbs free energy is given by

$$\Delta G=\Delta H-T\Delta S$$
$$\Delta S=\frac{\Delta H-\Delta G}{T}$$

Substituting the values,

$$=\frac{-24000\text{ J}-(-9000\text{ J})}{298\text{ K}}$$
$$\Delta S=-50.33\text{ J}/\text{K}$$

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# Question12

**One mole of $C_2H_5OH(l)$ was completely burnt in oxygen to form $CO_2(g)$ and $H_2O(l)$. The standard enthalpy of formation $(\Delta_f H^\ominus)$ of $C_2H_5OH(l)$, $CO_2(g)$ and $H_2O(l)$ is $x, y, z$ kJ mol$^{-1}$ respectively. What is $\Delta_r H^\ominus$ ( in kJmol$^{-1}$) for this reaction?**

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**Options:**

A.

$(2y + 3z + x)$

B.

$(2y - 3z + x)$

C.

$(x - 2y - 3z)$

D.

$(2y + 3z - x)$

**Answer: D**

### Solution:

The balanced equation for the combustion of ethanol is,

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

$$\begin{aligned}\Delta H_{\text{reaction}} &= \Sigma \Delta H_{f\text{ product}} - \Sigma \Delta H_{\text{reactant}} \\ &= [2\Delta H_{fCO_2} + 3\Delta H_{fH_2O}] - [\Delta H_{fC_2H_5OH} + 3\Delta H_{fO_2}] \\ &= [2y + 3z] - [x + 0] \\ &= 2y + 3z - x\end{aligned}$$

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# Question13

**Identify the correct statements from the following.**

**I. Work is a path function.**

**II. Enthalpy is an extensive property.**

**III. Lattice enthalpy of ionic compounds can be obtained from Born-Haber cycle.**

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**Options:**

A.

I and II

B.

I and III

C.

II and III

D.

I, II and III

**Answer: D**

## Solution:

**Statement I:** *Work is a path function.*

- **True.**

Work depends on the path taken between initial and final states, not just on the states themselves. Hence, work is a **path function**.

**Statement II:** *Enthalpy is an extensive property.*

- **True.**

Enthalpy ($H$) depends on the amount of substance. If the system size doubles, enthalpy doubles. So it is an **extensive property**.

**Statement III:** *Lattice enthalpy of ionic compounds can be obtained from Born–Haber cycle.*

- **True.**

The Born–Haber cycle uses Hess's law to calculate lattice enthalpy indirectly from measurable thermodynamic quantities.

✅ **Correct statements:** II and III are correct, and actually, I is also correct.

Wait, let's double-check statement I:

Work is **NOT** a state function (true), but **it is a path function** — the statement says "Work is a path function" — **That is correct.**

Therefore all three I, II, and III are correct.

✅ **Final Answer: Option D — I, II and III**

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# Question14

## Which of the following processes entropy change $(\Delta S)$ is negative?

## I. Sublimation of dry ice

## II. Freezing of water

## III. Crystallisation of the dissolved substance

## IV. Burning of rocket fuel

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**Options:**

A.

I and II only

B.

II and III only

C.

III and IV only

D.

I and IV only

**Answer: B**

## Solution:

Let's analyze each process in terms of entropy change ($\Delta S$):

**I. Sublimation of dry ice (solid $CO_2$ → gas $CO_2$)**

- A solid goes directly to a gas.
- The disorder (randomness) increases significantly.
- **Entropy increases** $\rightarrow \Delta S > 0$

**II. Freezing of water (liquid → solid)**

- A liquid becomes a more ordered solid structure.
- Disorder decreases.
- **Entropy decreases** → $\Delta S < 0$

**III. Crystallisation of the dissolved substance**

- Dissolved ions/molecules (disordered) form a regular crystal lattice (ordered).
- Disorder decreases.
- **Entropy decreases** → $\Delta S < 0$

**IV. Burning of rocket fuel (combustion)**

- Chemical reaction produces gases and increases randomness overall.
- **Entropy increases** → $\Delta S > 0$

✅ **Processes with negative entropy change:**

**II. Freezing of water**

**III. Crystallisation of the dissolved substance**

✅ **Correct answer: Option B — II and III only**

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# Question15

### Consider the following :

### Statement I : During isothermal expansion of an ideal gas its enthalpy decreases.

### Statement II : When 2.0 L of an ideal gas expands isothermally into vaccum, $\Delta U = 0$.

### The correct answer is :

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**Options:**

A. both statement-I and statement-II are correct

B. both statement-I and statement-II are not correct

C. statement-I is correct but statement-II is not correct

D. statement-I is not correct but statement-II is correct

**Answer: D**

## Solution:

### Statement I

> During isothermal expansion of an ideal gas its enthalpy decreases.

For an *ideal gas*, enthalpy $H = U + pV$.

For an ideal gas:

$U = nC_VT$

and since internal energy $U$ depends **only on temperature** for an ideal gas,

$dU = 0 \quad$ if temperature is constant (isothermal process).

Also, $pV = nRT$.

Therefore,

$H = U + pV = nC_VT + nRT = n(C_V + R)T = nC_PT$

Since $T$ is constant (isothermal), **enthalpy $H$ is also constant.**

Hence, **enthalpy does not decrease**, it remains the same.

✅ **Statement I is incorrect.**

### Statement II

> When 2.0 L of an ideal gas expands isothermally into vacuum, $\Delta U = 0$.

For **free expansion into vacuum**:

- External pressure $= 0$
- So $w = 0$
- No heat exchange ($q = 0$) because the system is isolated (expanding into vacuum)
- Thus, $\Delta U = q + w = 0 + 0 = 0$

Also, for an ideal gas, $\Delta U$ depends only on temperature, and since it is *isothermal*, $\Delta U = 0$.

✅ **Statement II is correct.**

### ✅ Final Answer

**Statement I is not correct but Statement II is correct.**

**Correct option: D**

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# Question16

## The energy required to increase the temperature of 180 g of liquid water from $10^\circ\mathrm{C}$ to $15^\circ\mathrm{C}$ is 3765 J . What is $C_p$ of water in

$\mathrm{Jmol}^{-1}\ \mathrm{K}^{-1}$? ($H_2O = 18u$)

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**Options:**

A.

75.3

B.

376.5

C.

753

D.

37.65

**Answer: A**

## Solution:

**Given:**

$m = 180\ \mathrm{g}$

$\Delta T = 15 - 10 = 5\ \mathrm{K}$

$q = 3765\ \mathrm{J}$

$M(\mathrm{H_2O}) = 18\ \mathrm{g/mol}$

We need to find $C_p$ in **J mol⁻¹ K⁻¹**.

**Step 1: Use the relation $q = nC_p\Delta T$**

$n = \frac{m}{M} = \frac{180}{18} = 10\ \mathrm{mol}$

Substitute values:

$3765 = 10 \times C_p \times 5$

**Step 2: Solve for $C_p$**

$C_p = \frac{3765}{10\times5} = \frac{3765}{50} = 75.3\ \mathrm{J\ mol^{-1}\ K^{-1}}$

✅ **Final Answer:**

$\boxed{C_p = 75.3\ \mathrm{J\ mol^{-1}\ K^{-1}}}$

**Correct Option: A (75.3)**

# Question17

**At 273 K the maximum work done when pressure on 10 g of hydrogen is reduced from 10 atm to 1 atm under isothermal, reversible conditions is**

**(Assume the gas behaves ideally)**

$$\left(R=83\mathrm{Jk}^{-1}\ \mathrm{mol}^{-1}\right)$$

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**Options:**

A.

-52.18 kJ

B.

+26.09 kJ

C.

-26.09 kJ

D.

+52.18 kJ

**Answer: C**

## Solution:

Number of moles of hydrogen

$$\begin{aligned} & =\frac{\text{mass}}{\text{molar mass}} \\ & =\frac{10\ \mathrm{g}}{2\ \mathrm{g/mol}}=5\ \mathrm{mol} \end{aligned}$$

Using work done in isothermal condition,

$$W=-nRT\ln\left[\frac{p_1}{p_2}\right]$$

$$W=-5\times 8.3\times 273\times \ln\left[\frac{10}{1}\right]$$

$$W\simeq -26090\text{ J}=-26.09\text{ kJ}$$

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# Question18

## Observe the following reactions.I. $\mathrm{CaCO_3(\,s)\longrightarrow CaO(s)+CO_2(\,g)}$II. $\mathrm{Cl_2(\,g)\longrightarrow 2Cl(g)}$III. $\mathrm{H_2O}(l)\longrightarrow \mathrm{H_2O}(s)$Identify the reaction in which entropy increases

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**Options:**

A. I, I, III

B. I, II only

C. I, III only

D. II, III only

**Answer: B**

## Solution:

For each of the given reactions, we will analyze how entropy is affected:

**Reaction I:**

$\mathrm{CaCO_3}(s)\rightarrow \mathrm{CaO}(s)+\mathrm{CO_2}(g)$

The change in the number of gas molecules ($\Delta n_g$) is calculated as follows:

$\Delta n_g=(0+1)-0=1$

Since $\Delta n_g>0$, this indicates an increase in entropy because the production of gas molecules increases disorder.

**Reaction II:**

$\mathrm{Cl_2}(g)\rightarrow 2\mathrm{Cl}(g)$

The change in the number of gas molecules is:

$\Delta n_g=2-1=1$

Here, $\Delta n_g>0$, pointing to an increase in entropy because more gas molecules are produced, enhancing disorder.

**Reaction III:**

$\mathrm{H_2O}(l)\rightarrow \mathrm{H_2O}(s)$

This reaction involves water transitioning from a liquid to a solid state. As water solidifies into ice, the molecular arrangement becomes more ordered, resulting in a decrease in entropy.

Therefore, entropy increases in reactions I and II, but decreases in reaction III.

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# Question19

### Observe the following reaction.

$$A B \mathrm{O}_3(\mathrm{~s}) \xrightarrow{1000 \mathrm{~K}} A \mathrm{O}(\mathrm{~s})+B \mathrm{O}_2(\mathrm{~g})$$

**$\Delta_r H$ for this reaction is $x \mathrm{~kJ} \mathrm{~mol}^{-1}$. What is its $\Delta_r U$ (in $\mathrm{kJmol}^{-1}$ ) at the same temperature?**

$$\left(R=8.3 \mathrm{~J} \mathrm{~mol}^{-1} \mathrm{~K}^{-1}\right)$$

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**Options:**

A. $x-8300$

B. $x+8.3$

C. $x+8300$

D. $x-83$

**Answer: D**

## Solution:

To find $\Delta_r U$ for the given reaction at 1000 K, we start with the thermodynamic relationship:

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

where $\Delta_r H$ is the enthalpy change for the reaction, $\Delta_r U$ is the internal energy change, $R$ is the gas constant $(8.3 \mathrm{~J} \mathrm{~mol}^{-1} \mathrm{~K}^{-1})$, and $T$ is the temperature in Kelvin.

For the reaction:

$A B \mathrm{O}_3(s) \xrightarrow{1000 \mathrm{~K}} \mathrm{AO}(s)+B \mathrm{O}_2(g)$

$\Delta n_g$ represents the change in the number of moles of gas. Initially, there are 0 moles of gas because the reactant is a solid, while the product is 1 mole of gas $(B \mathrm{O}_2(g))$. Therefore, $\Delta n_g=1$.

Substitute the values into the thermodynamic equation:

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

$\Delta_r U = x - (1) \times (8.3 \times 10^3)$

$\Delta_r U = x - 8.3 \times 10^3$

Thus, $\Delta_r U$ in $\mathrm{kJ/mol}$ is:

$\Delta_r U = (x - 8.3)\,\mathrm{kJ/mol}$

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# Question20

### At 300 K, $\Delta_r G^\ominus$ for the reaction $A_2(g) \rightleftharpoons B_2(g)$ is $-11.5\ \mathrm{kJ\ mol}^{-1}$. The Equilibrium constant at 300 K is approximately ( $R = 8314\ \mathrm{J\ mol}^{-1}\ \mathrm{K}^{-1}$ )

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**Options:**

A. 10

B. 100

C. 1000

D. 25

**Answer: B**

## Solution:

To find the equilibrium constant $K$ for the reaction $A_2(g) \rightleftharpoons B_2(g)$ at 300 K, we are given:

$\Delta_r G^\ominus = -11.5\,\mathrm{kJ/mol} = -11,500\,\mathrm{J/mol}$

$T = 300\,\mathrm{K}$

$R = 8.314\,\mathrm{J/mol\ K}$

The relationship between $\Delta_r G^\ominus$ and the equilibrium constant $K$ is expressed as:

$\Delta_r G^\ominus = -RT \ln K$

You can also express this in terms of a logarithm base 10:

$\Delta_r G^\ominus = -2.303\,RT\ \log_{10} K$

Rearranging the equation to solve for $\log_{10} K$, we get:

$\log_{10} K = \frac{-\Delta_r G^\ominus}{2.303 \times R \times T}$

Substituting the given values:

$\log_{10} K = \frac{-(-11{,}500)}{2.303 \times 8.314 \times 300}$

Calculating this:

$= \frac{11{,}500}{2.303 \times 8.314 \times 300}$

$= \frac{11{,}500}{5744.202}$

$\approx 2.0$

Thus, solving for $K$, we have:

$K = 10^{2.0} \approx 100$

Therefore, the equilibrium constant at 300 K is approximately 100.

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# Question21

### Two statements are given below.

### Statement I : The reaction $\mathrm{Cr_2O_3} + 2\mathrm{Al} \longrightarrow \mathrm{Al_2O_3} + 2\mathrm{Cr}$ $(\Delta G^{\ominus} = -421\ \mathrm{kJ})$ is thermodynamically feasible.

### Statement II : The above reaction occurs at room temperature.

### The correct answer is

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**Options:**

A. Both Statement I and Statement II are correct.

B. Both Statements I and II are not correct.

C. Statement I is correct but Statement II is not correct.

D. Statement I is not correct but Statement II is correct.

**Answer: C**

## Solution:

Statement I is accurate, while Statement II is not. Here's the clarification:

In the thermite process described by the reaction:

$\mathrm{Cr_2O_3} + 2\mathrm{Al} \longrightarrow \mathrm{Al_2O_3} + 2\mathrm{Cr} + \mathrm{Q}$

This reaction is both exothermic and thermodynamically feasible, as indicated by the negative Gibbs free energy change ( $\Delta G^{\ominus}=-421$ kJ). This signifies that the reaction is spontaneous under standard conditions.

However, despite being thermodynamically feasible, the reaction requires a significant amount of activation energy to proceed. At room temperature, the available activation energy is insufficient, which means that the reaction will not occur spontaneously without the application of additional heat or a catalyst to overcome this energy barrier. Thus, the reaction does not naturally occur at room temperature.

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# Question22

### What is the enthalpy change (in J ) for converting 98 of $\mathrm{H_2O}(t)+10^{\circ}\mathrm{C}$ to $\mathrm{H_2O}(l)$ at $+20^{\circ}\mathrm{C}$ ?

$$\left(C_p\left(\mathrm{H_2O}(\eta)\right)=75 \mathrm{Jmol}^{-1} \mathrm{~K}^{-1}\right)$$

**(density of $\mathrm{H_2O}(l)=1 \mathrm{gmL}^{-1}$)**

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**Options:**

A.

750

B.

75

C.

37.5

D.

375

**Answer: D**

## Solution:

To find the enthalpy change for converting 9 grams of $\mathrm{H_2O}(l)$ from 10°C to 20°C, we follow these steps:

**Calculate the Number of Moles:**

The formula for the number of moles is:

Number of moles of $\mathrm{H_2O}=\frac{\text{Mass}}{\text{Molar mass}}$

Given the mass is 9 g and the molar mass of water is 18 g/mol, we have:

Number of moles $= \frac{9}{18} = 0.5\,\text{mol}$

**Determine the Change in Temperature:**

The change in temperature is:

$\Delta T = T_2 - T_1 = 20^\circ\text{C} - 10^\circ\text{C} = 10\,\text{K}$

**Calculate the Enthalpy Change:**

The enthalpy change is given by the formula:

$\Delta H = n \cdot C_p \cdot \Delta T$

Substitute the known values ($n = 0.5$, $C_p = 75\,\text{J/mol K}$, and $\Delta T = 10\,\text{K}$):

$\Delta H = 0.5 \times 75 \times 10 = 375\,\text{J}$

Therefore, the enthalpy change for this process is 375 Joules.

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# Question23

### $A, B, C$ and $D$ are some compounds. The entnalpy of formation of $A(g), B(g), C(g)$ and $D(g)$ is $9.7, -110, 81$ and $-393\ \text{kJ mol}^{-1}$ respectively. What is $\Delta_r H$

### (in $\text{kJmol}^{-1}$ ) for the given reaction ?

$$A(g) + 3B(g) \longrightarrow C(g) + 3D(g)$$

## AP EAPCET 2024 - 22th May Morning Shift

**Options:**

A. -777.7

B. +777.7

C. -14183

D. +14183

**Answer: A**

## Solution:

To calculate the reaction enthalpy change $\Delta_r H$ for the given reaction, we use the following reaction:

$A(g) + 3B(g) \longrightarrow C(g) + 3D(g)$

The reaction enthalpy change is calculated using the formula:

$$\Delta_r H = \Delta_{\text{formation of C}} + 3 \times \Delta_{\text{formation of D}} - \Delta_{\text{formation of A}} - 3 \times \Delta_{\text{formation of B}}$$

Given the enthalpies of formation are:

$\Delta_{\text{formation of A}} = 9.7\,\text{kJ/mol}$

$\Delta_{\text{formation of B}} = -110\,\text{kJ/mol}$

$\Delta_{\text{formation of C}} = 81\,\text{kJ/mol}$

$\Delta_{\text{formation of D}} = -393\,\text{kJ/mol}$

Substituting these values into the formula gives:

$\Delta_r H = 81 + (-3 \times 393) - 9.7 - (3 \times -110)$

Calculate each part:

$-3 \times 393 = -1179$

$3 \times -110 = -330$

Now plug these values back into the formula:

$\Delta_r H = 81 - 1179 - 9.7 + 330$

Breaking it down:

$81 - 1179 = -1098$

$-1098 - 9.7 = -1107.7$

$-1107.7 + 330 = -777.7$

Thus, the reaction enthalpy change is:

$\Delta_r H = -777.7\,\text{kJ/mol}$

---

# Question24

## Identify the correct equation relating $\Delta H, \Delta U$ and $\Delta T$ for 1 mole of an ideal gas from the following. ( $R =$ gas constant)

## AP EAPCET 2024 - 21th May Evening Shift

**Options:**

A. $(\Delta H)^2 = \Delta U + R\Delta T$

B. $\Delta H = (\Delta U)^2 + R\Delta T$

C. $\Delta U = \Delta H - R\Delta T$

D. $\Delta U = \Delta H + R\Delta T$

**Answer: C**

## Solution:

To relate the change in enthalpy ($\Delta H$), change in internal energy ($\Delta U$), and change in temperature ($\Delta T$) for 1 mole of an ideal gas, we start with the ideal gas law:

$pV = nRT$

Assuming pressure $p$ is constant, the change in volume $\Delta V$ can be expressed as:

$p\Delta V = R\Delta T$

The change in enthalpy $\Delta H$ is related to the change in internal energy $\Delta U$ by:

$\Delta H = \Delta U + p\Delta V$

Substituting the expression for $p\Delta V$ from the ideal gas law into the enthalpy equation:

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

Rearranging the equation to express $\Delta U$ in terms of $\Delta H$ gives:

$\Delta U = \Delta H - R\Delta T$

This equation shows the relationship between the change in internal energy, the change in enthalpy, and the change in temperature for an ideal gas.

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# Question25

## The number of extensive properties in the following list is enthalpy, density, volume, internal energy, temperature.

## AP EAPCET 2024 - 21th May Evening Shift

**Options:**

A. 4

B. 2

C. 3

D. 5

**Answer: C**

## Solution:

Let's review the properties:

Extensive properties depend on the size or amount of the system. Examples include enthalpy, volume, and internal energy.

Intensive properties do not depend on the system's size. Examples include density and temperature.

Looking at the list:

Enthalpy: Extensive

Density: Intensive

Volume: Extensive

Internal Energy: Extensive

Temperature: Intensive

Thus, there are 3 extensive properties in the list (enthalpy, volume, and internal energy).

Answer: Option C (3)

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# Question26

### Identify the correct statements from the following.I. $\Delta_r G$ is zero for $A \rightleftharpoons B$ reaction.II. The entropy of pure crystalline solids approaches. zero as the temperature approaches absolute zero.III. $\Delta U$ of a reaction can be determined using bomb calorimeter.

### AP EAPCET 2024 - 21th May Morning Shift

**Options:**

A. I, II only

B. I, III only

C. II, III only

D. I, II, III

**Answer: D**

## Solution:

All the given statements are correct.

----------------------------------------------------------------------------------------------

# Question27

### Observe the following reactions.

$$AB(g) + 25\mathrm{H_2O}(l) \longrightarrow (25\mathrm{H_2O})AB; \quad \Delta H = x \text{ kJ mol}^{-1}$$
$$AB(g) + 50\mathrm{H_2O}(l) \longrightarrow (50\mathrm{H_2O})AB; \quad \Delta H = y \text{ kJ mol}^{-1}$$

## AP EAPCET 2024 - 21th May Morning Shift

**Options:**

A. $(y-x)$

B. $(y+x)$

C. $\frac{y}{x}$

D. $\frac{x}{y}$

**Answer: A**

## Solution:

We're dissolving a gas $AB$ in water $(H_2O)$.

The enthalpy change $(\Delta H)$ tells us how much heat is absorbed or released during the dissolving process.

Here's what the reactions tell us:

When $AB$ dissolves in 25 moles of water, the heat change is $x$ kJ/mol.

When $AB$ dissolves in 50 moles of water, the heat change is $y$ kJ/mol.

Now, let's consider what $y-x$ would represent:

$y-x$ is the difference in enthalpy when we dissolve $AB$ in 50 moles of water compared to when we dissolve it in 25 moles of water. This value tells us about the additional heat change when we add another 25 moles of water to the solution.

Therefore, the correct answer is:

**Option A** $(y-x)$

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# Question28

## The equation that represents 'coal gasification' is

## AP EAPCET 2024 - 21th May Morning Shift

**Options:**

A. $\mathrm{CO(g)+H_2O(g)\xrightarrow[\text{catalyst}]{673\ K}CO_2(g)+H_2(g)}$

B. $\mathrm{C(s)+H_2O(g)\xrightarrow{1270\ K}CO(g)+H_2(g)}$

C. $\mathrm{2C(s)+O_2(g)+4\,N_2(g)\xrightarrow{1273\ K}2CO(g)+4\,N_2(g)}$

D. $CH_4(g)+H_2O(g)\xrightarrow[\text{Ni}]{1270\text{ K}}CO(g)+3H_2(g)$

**Answer: B**

## Solution:

Coal gasification is the method of transforming coal and water into sm gas, which is a mixture of CO and $H_2$

$C(s)+H_2O(g)\xrightarrow{1270\text{ K}}CO(g)+H_2(g)$

---

# Question29

## The standard free energy change $(\Delta G^{\circ})$ for the following reaction (in kJ ) at $25^{\circ}C$ is$3Ca(s)+2Au^{3+}(aq,1M)\longrightarrow 3Ca^{2+}(aq,1M)+2Au(s)$

## AP EAPCET 2024 - 21th May Morning Shift

**Options:**

A. $-2.53\times 10^3$

B. $+2.53\times 10^3$

C. $-2.53\times 10^4$

D. $+2.53\times 10^4$

**Answer: A**

## Solution:

To determine the standard free energy change $(\Delta G^{\circ})$ for the given reaction at $25^{\circ}C$, we follow these steps:

**Identify the Reaction**:

$3Ca(s)+2Au^{3+}(aq,1\text{ M})\rightarrow 3Ca^{2+}(aq,1\text{ M})+2Au(s)$

**Calculate Electron Exchange**:

The electron exchange number, $n$, is 6.

**Apply the Free Energy Formula**:

$\Delta G^{\circ}=-nFE_{\text{Cell}}^{\circ}$

Here, $F$ is Faraday's constant (96500 C/mol).

**Determine $E_{\text{Cell}}^{\circ}$** :

$E_{\text{Cell}}^{\circ}=E_{\text{Cathode}}^{\circ}-E_{\text{Anode}}^{\circ}$

For this reaction:

$E^\circ_{\text{Cell}} = E^\circ_{\text{Au}^{3+}/\text{Au}} - E^\circ_{\text{Ca}^{2+}/\text{Ca}}$

Using given electrode potentials:

$E^\circ_{\text{Cell}} = +1.50\,\text{V} - (-2.87\,\text{V}) = +4.37\,\text{V}$

**Compute $\Delta G^\circ$:**

$\Delta G^\circ = -6 \times 96500 \times 4.37 = -2530230\,\text{J}$

**Convert $\Delta G^\circ$ to kJ:**

$\Delta G^\circ = -2530.23\,\text{kJ}$

Thus, the standard free energy change for the reaction is $-2.53 \times 10^3\,\text{kJ}$.

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# Question30

## What is the ratio of kinetic energies of 3 g of hydrogen and 4 g of oxygen at a certain temperature?

## AP EAPCET 2024 - 20th May Evening Shift

**Options:**

A. 3 : 4

B. 6 : 1

C. 12 : 1

D. 1 : 12

**Answer: C**

## Solution:

Given:

Mass of hydrogen = 3 g

Mass of oxygen = 4 g

**Number of Moles Calculation:**

**Hydrogen:**

Number of moles of hydrogen $= \frac{3}{2} = 1.5\,\text{mol}$

**Oxygen:**

Number of moles of oxygen $= \frac{4}{32} = 0.125\,\text{mol}$

*Formula used:* $\frac{\text{Mass}}{\text{Molar mass}}$

**Kinetic Energy at a Given Temperature:**

The kinetic energy (KE) of a gas is given by the formula:

$\mathrm{KE} = \frac{3}{2} nRT$

where:

$n$ is the number of moles,

$R$ is the ideal gas constant,

$T$ is the temperature.

**Ratio of Kinetic Energies:**

To find the ratio of kinetic energies of hydrogen ($H_2$) and oxygen ($O_2$):

$\frac{(\mathrm{KE})_{\mathrm{H}_2}}{(\mathrm{KE})_{\mathrm{O}_2}} = \frac{n_{\mathrm{H}_2}}{n_{\mathrm{O}_2}} = \frac{1.5}{0.125} = \frac{12}{1}$

Thus, the ratio of the kinetic energies of hydrogen to oxygen is:

$(\mathrm{KE})_{\mathrm{H}_2} : (\mathrm{KE})_{\mathrm{O}_2} = 12 : 1$

---

# Question31

## The volume of an ideal gas contracts from 10.0 L to 2.0 L under an applied pressure of 2.0 atm . During contraction the system also evolved 900 J of heat. The change in internal energy (in J ) involved in the system is ( $1 \mathrm{~L} \text{ atm} = 101.3 \mathrm{~J}$ )

## AP EAPCET 2024 - 20th May Evening Shift

**Options:**

A. 720.8

B. 360.4

C. 1620.8

D. 810.4

**Answer: A**

## Solution:

Given:

Heat released by the system, $Q = -900 \mathrm{~J}$

Change in volume, $\Delta V = 10 \mathrm{~L} - 2 \mathrm{~L} = 8 \mathrm{~L}$

Pressure, $p = 2.0 \mathrm{~atm}$

According to the first law of thermodynamics:

$\Delta U=Q-W$

where $W=-p\Delta V$. Thus, the equation modifies to:

$\Delta U=Q+p\Delta V$

Now substituting the given values:

$\Delta U=-900+2\times 8\times 101.3$

Calculating further:

$\Delta U=-900+2\times 8\times 101.3=-900+1620.8=720.8\text{ J}$

Therefore, the change in internal energy, $\Delta U$, is 720.8 J.

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# Question32

## The molar heat of fusion and vaporisation of benzene are 10.9 and $31.0\text{ kJ mol}^{-1}$ respectively. The changes in entropy for the solid $\rightarrow$ liquid and liquid $\rightarrow$ vapour trasitions for benzene are $x$ and $y$, $\text{JK}^{-1}\text{ mol}^{-1}$, respectively. The value of $(y-x)$ (in $\text{JK}^{-1}\text{ mol}^{-1}$ ) is (At 1 atm benzene melts at $5.5^{\circ}\text{C}$ and boils at $80^{\circ}\text{C}$ )

## AP EAPCET 2024 - 20th May Evening Shift

**Options:**

A. 87.8

B. 48.7

C. 39.1

D. 28.7

**Answer: B**

## Solution:

To determine the change in entropy for the transitions of benzene, we need to calculate the entropy changes for both the fusion (solid to liquid) and vaporization (liquid to vapor) processes.

### Change in Entropy for Fusion (Solid to Liquid)

The formula for the entropy change due to fusion is:

$$\Delta S_{\text{fusion}}=\frac{\Delta H_{\text{fusion}}}{T_{\text{melting point}}}$$

Given:

$\Delta H_{\text{fusion}}=10.9\text{ kJ/mol}=10,900\text{ J/mol}$

Melting point for benzene = $5.5^{\circ}\text{C}=278.5\text{ K}$

Calculate:

$x = \Delta S_{\text{fusion}} = \frac{10{,}900}{278.5} = 39.11\,\text{J/K/mol}$

**Change in Entropy for Vaporization (Liquid to Vapor)**

The formula for the entropy change due to vaporization is:

$\Delta S_{\text{vaporization}} = \frac{\Delta H_{\text{vaporization}}}{T_{\text{boiling point}}}$

Given:

$\Delta H_{\text{vaporization}} = 31.0\,\text{kJ/mol} = 31,000\,\text{J/mol}$

Boiling point for benzene = $80^{\circ}\text{C} = 353\,\text{K}$

Calculate:

$y = \Delta S_{\text{vaporization}} = \frac{31{,}000}{353} = 87.81\,\text{J/K/mol}$

**Difference in Entropy Change**

Finally, the difference between the entropy changes for vaporization and fusion is:

$y - x = 87.81 - 39.11 = 48.7\,\text{J/K/mol}$

---

# Question33

## Observe the following reaction,

$$2A_2(g) + B_2(g) \xrightarrow{T(\text{ K})} 2A_2B(g) + 600\text{ kJ}$$

## The standard enthalpy of formation $(\Delta_f H^{\ominus})$ of $A_2B(g)$ is

## AP EAPCET 2024 - 20th May Morning Shift

**Options:**

A. $600\text{ kJ mol}^{-1}$

B. $300\text{ kJ mol}^{-1}$

C. $-300\text{ kJ mol}^{-1}$

D. $-600\text{ kJ mol}^{-1}$

**Answer: C**

## Solution:

Given the reaction:

$2A_2(g) + B_2(g) \longrightarrow 2A_2B(g) + 600\,\mathrm{kJ}$

This reaction releases energy, indicating it is exothermic. To find the standard enthalpy of formation $\Delta_f H^{\ominus}$ for $A_2B(g)$, we consider that the total energy released is distributed over the formation of two moles of $A_2B(g)$.

Therefore, the calculation for $\Delta_f H^{\ominus}$ is:

$\Delta_f H^{\ominus}$ for $A_2B(g) = \frac{-600\,\mathrm{kJ}}{2}$

This results in:

$\Delta_f H^{\ominus} = -300\,\mathrm{kJ\,mol}^{-1}$

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# Question34

## Identify the molecule for which the enthalpy of atomisation ( $\Delta_a H^{\ominus}$ ) and bond dissociation enthalpy $\left(\Delta_{\text{bond}} \quad H^{\ominus}\right)$ are not equal

## AP EAPCET 2024 - 20th May Morning Shift

**Options:**

A. $\mathrm{H}_2$

B. $\mathrm{Cl}_2$

C. $\mathrm{F}_2$

D. $\mathrm{CH}_4$

**Answer: D**

## Solution:

Enthalpy of atomisation $(\Delta_a H^{\ominus})$ and bond dissociation enthalpy ( $\Delta_{\text{bond}} \quad H^{\ominus}$ ) is not same for polyatomic molecules.

This is because in every dissociation steps, bonds are broken into different species.

Hence, $\mathrm{CH}_4$ is correct answer.

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# Question35

## Given below are two statements :

## Statement I For isothermal irreversible change of an ideal gas, $q = -w = p_{\text{ext}} \quad \left(V_{\text{final}} \quad - V_{\text{initial}} \quad\right)$

**Statement II For adiabatic change, $\Delta U = w_{\text{adiabatic}}$**

**The correct answer is :**

## AP EAPCET 2024 - 19th May Evening Shift

**Options:**

A. Both Statement I and Statement II are correct.

B. Both Statement I and Statement II are not correct.

C. Statement I is correct but Statement II is not correct.

D. Statement I is not correct but Statement II is correct.

**Answer: A**

## Solution:

Let's analyze each statement step by step.

For an isothermal process involving an ideal gas, the internal energy change is zero, i.e.,

$\Delta U = 0.$

Since the first law of thermodynamics states that

$\Delta U = q + w,$

it follows that

$q = -w.$

For an irreversible process done against a constant external pressure, the work done by the gas is given by

$w = -p_{\text{ext}}\left(V_{\text{final}} - V_{\text{initial}}\right).$

Thus, the heat absorbed is

$q = -w = p_{\text{ext}}\left(V_{\text{final}} - V_{\text{initial}}\right).$

So, Statement I is correct (assuming the process is performed against a constant external pressure).

For an adiabatic process, by definition there is no heat exchange, so

$q = 0.$

Then, the first law reduces to

$\Delta U = q + w = 0 + w = w.$

Expressing the work performed in the adiabatic process as $w_{\text{adiabatic}}$, we have

$\Delta U = w_{\text{adiabatic}}.$

Therefore, Statement II is also correct.

Since both statements are valid, the correct answer is:

Option A

Both Statement I and Statement II are correct.

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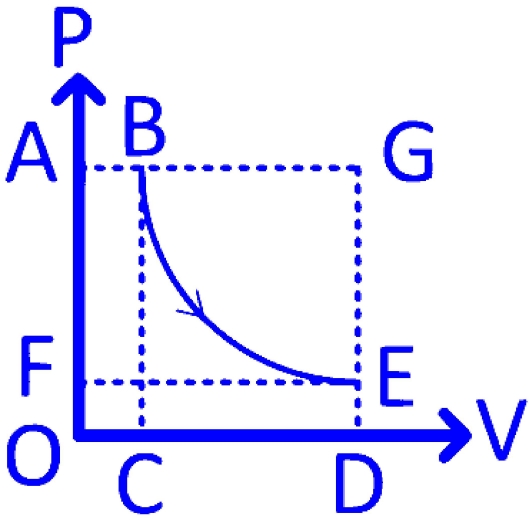
# Question36

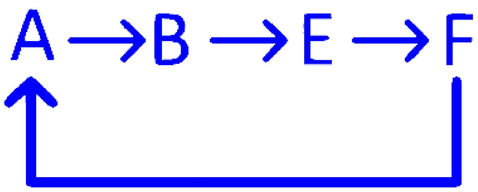
**A thermodynamic process $(B \rightarrow E)$ was completed as shown below. The work done is equal to area under the limits**



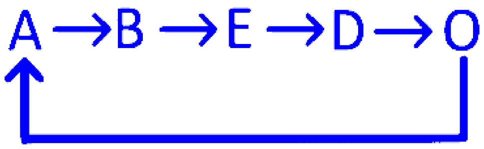
## AP EAPCET 2024 - 19th May Evening Shift

**Options:**

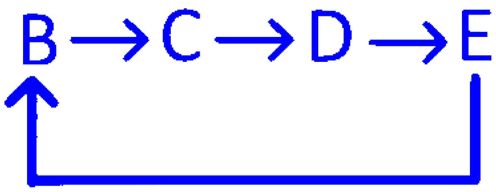
A.



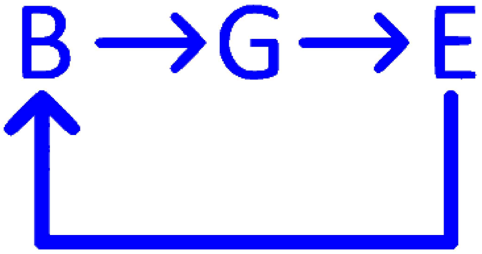
B.



C.



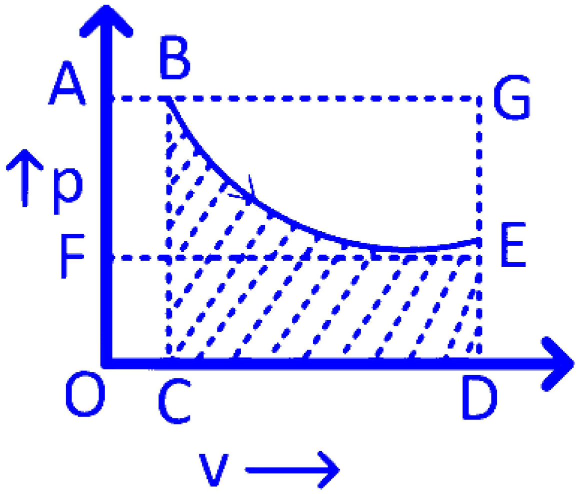
D. Image



**Answer: C**

## Solution:

From the $pV$ graph, we can observe that it is an isothermal process.



The graph is of isothermal expansion.

Initial state, $p_A$ and $V_C$ as point $B$

Final state, $p_F$ and $V_D$ at point $E$.

So, work done is area under curve of isothermal expansion.

$\therefore$ Workdone $=$ Area of $B \rightarrow C \rightarrow D \rightarrow E$

---

# Question37

**At 300 K for the reaction. $A \rightarrow P$. The $\Delta S_p$ is $5 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}$, What is the heat absorbed (in kJ $\mathrm{mol}^{-1}$ ) by the system?**

## AP EAPCET 2024 - 18th May Morning Shift

**Options:**

A. 1.5

B. 15

C. 1500

D. 0.6

**Answer: A**

## Solution:

Given for the reaction $A \rightarrow P$:

Temperature $T=300 \mathrm{~K}$

Entropy change $\Delta S_{\text {system }}=5 \mathrm{~J} / \mathrm{K} / \mathrm{mol}$

The heat absorbed by the system can be calculated using the formula:

$\Delta H_{\text {system }}=T \times \Delta S_{\text {system }}$

Substitute the given values:

$\Delta H_{\text {system }}=300 \mathrm{~K} \times 5 \mathrm{~J} / \mathrm{K} / \mathrm{mol}=1500 \mathrm{~J} / \mathrm{mol}$

Since the answer is needed in kilojoules per mole, convert joules to kilojoules:

$1500 \mathrm{~J} / \mathrm{mol}=1.5 \mathrm{~kJ} / \mathrm{mol}$

Thus, the heat absorbed by the system is $1.5 \mathrm{~kJ} / \mathrm{mol}$.

---

# Question38

**Identify the incorrect statements form the following.**

**I.** $\Delta S_{\text {pum }}=\left(\Delta S_{\text {nal }}+\Delta S_{\text {um }}\right)$

**II.** $A(\bar{i} \rightarrow A(\phi)$ **: For this process entropy change decreases.**

**III. Entropy units are** $\mathrm{JKmol}^{-1}$**.**

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**Options:**

A. I,III only

B. I,II only

C. I,II,III

D. II,III only

**Answer: A**

## Solution:

Statement I and statement III are incorrect, while statement II is correct. The corrections for the incorrect statements are as follows:

The total entropy change that accompanies a process is given by:

$\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surrounding}}$

The SI units of entropy are $\mathrm{J\,K^{-1}\,mol^{-1}}$ or $\mathrm{J\,K^{-1}}$.

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# Question39

### Identify the correct statements from the following.

### I. At 0 K , the entropy of pure crystalline materials approach zero.

### II. Entropy for the process, $\mathrm{H_2O(l)} \longrightarrow \mathrm{H_2O(g)}$ decreases.

### III. Gibb's energy is a state function.

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**Options:**

A. I, II, III

B. I, III

C. II, III

D. I, II

**Answer: B**

## Solution:

(b) At 0 K , the entropy of pure crystalline materials is zero.

$\therefore$ The statement (I) is correct as entropy approach zero, at absolute zero.

Statement (II) entropy for the process, $H_2O(t) \longrightarrow H_2O(g)$ decreases is incorrect as entropy increases. The degree of randomness when water vaporises to gaseous state increases.

(III) Gibbs' energy is a state function is correct as the Gibbs' free energy depends on initial and final state only.

Hence, statement (I) and (III) are correct.

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# Question40

## Use the data from table to estimate the enthalpy of formation of $CH_3CHO$.

| Bond enthalpy | Bond | Enthalpy of formation |
|---|---|---|
| $400 \text{ kJ mol}^{-1}$ | $C-H$ | $C(g) 700 \text{ kJ mol}^{-1}$ |
| $350 \text{ kJ mol}^{-1}$ | $C-C$ | $H(g) 200 \text{ kJ mol}^{-1}$ |
| $700 \text{ kJ mol}^{-1}$ | $C=O$ | $O(g) 250 \text{ kJ mol}^{-1}$ |

## AP EAPCET 2022 - 5th July Morning Shift

**Options:**

A. $-200 \text{ kJ mol}^{-1}$

B. $-400 \text{ kJ mol}^{-1}$

C. $-350 \text{ kJ mol}^{-1}$

D. $-150 \text{ kJ mol}^{-1}$

**Answer: A**

## Solution:

$2C + 2H_2 + \frac{1}{2}O_2 \longrightarrow CH_3CHO$

$$\Delta H_f = \Sigma \mathrm{BE}_{\text{(reactants)}} \qquad - \Sigma \mathrm{BE}_{\text{(products)}}$$

$$= \left[(2 \times 700) + (2 \times 400) + \left(\tfrac{1}{2} \times 500\right)\right] - [(4 \times 400) + (250 \times 1) + (700 \times 1)]$$

$$\begin{aligned} &= 2450 - 2650 \text{ J} \\ &= -200 \text{ kJ/mol} \end{aligned}$$

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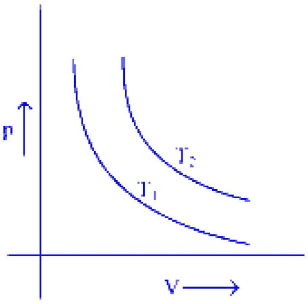
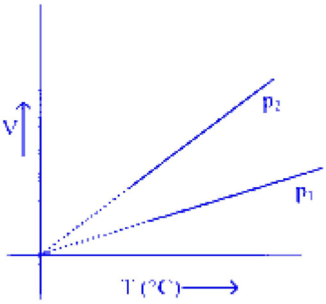
# Question41

## From the following plots, find the correct option.





### AP EAPCET 2022 - 4th July Evening Shift

**Options:**

A. $T_1 > T_2 : p_1 > p_2$

B. $T_1 > T_2 : p_2 > p_1$

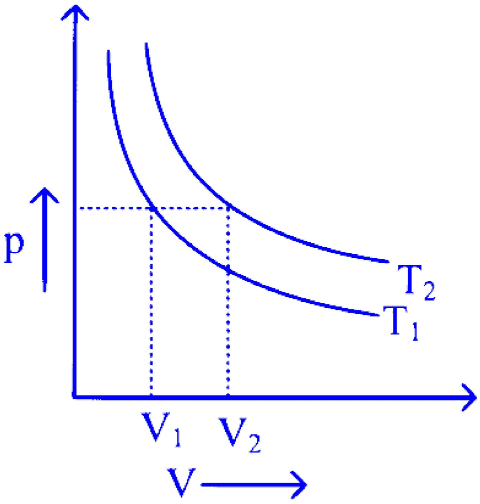
C. $T_2 > T_1 : p_2 > p_1$

D. $T_2 > T_1 : p_1 > p_2$

**Answer: D**

## Solution:

From first graph,

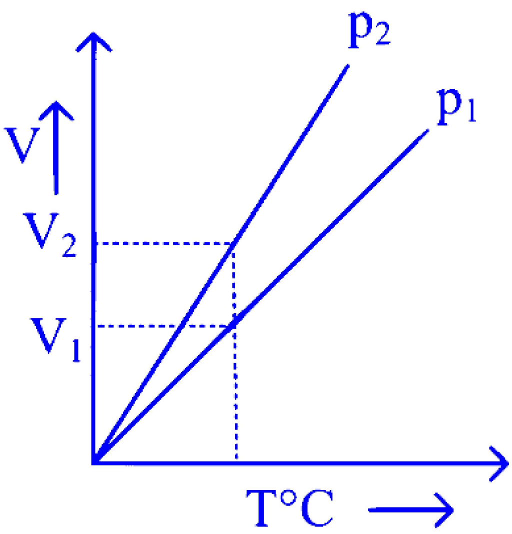


If we draw a line parallel to $x$-axis, then at that point pressure will be constant.

So, temperature will vary according to volume i.e, $V_2 > V_1$.

So, $T_2$ will be greater than 7, i.e., $T_2 > T_1$

Similarly, for second graph,



If we draw a line parallel to $y$-axis. Then, pressure will vary according to volume i.e., $V_2 > V_1$, but since pressure and volume are inversely proportional i.e., $pV = k$ hence $p_1 > p_2$.

So, the correct option is $T_2 > T_1$ and $p_1 > p_2$.

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# Question42

## Observe the following properties :Volume, enthalpy, density, temperature, heat capacity, pressure and internal energy. The number of extensive properties in the above list is

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**Options:**

A. 4

B. 5

C. 6

D. 3

**Answer: A**

## Solution:

Extensive properties are the properties that depends on the amount or mass of the matter present in a sample. Examples of extensive property are volume, enthalpy, heat capacity, internal energy, etc.

Hence, out of the given properties, only four properties, i.e. volume, enthalpy, heat capacity and internal energy are extensive properties.

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# Question43

## Match the following.

| | | | |
|---|---|---|---|
| A. | Isothermal process | i. | $q = \Delta U$ |
| B. | Adiabatic process | ii. | $W = -p \times \Delta V$ |
| C. | Isobaric process | iii. | $W = \Delta U$ |
| D. | Isochoric process | iv. | $W = -nRT \ln\left(\frac{V_t}{V_i}\right)$ |

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**Options:**

A. A-(iv), B-(iii), C-(ii), D-(i)

B. A-(iii), B-(iv), C-(i), D-(ii)

C. A-(i), B-(ii), C-(iii), D-(iv)

D. A-(ii), B-(i), C-(iv), D-(iii)

**Answer: A**

## Solution:

(A) Work done in an isothermal process can be calculated by using formula, $W = -nRT \ln\left(\frac{V_f}{V_i}\right)$

(B) For an adiabatic process, $q = 0$.

So, $\Delta U = W$
$\Delta U = q + W \Rightarrow \Delta U = W$

(C) Work done for an isobaric process is

$W = -p_{\text{ext}} \quad \times \Delta V$

(D) For an isochoric process, $\Delta V = 0$

So, $W = -p\Delta V = 0$

Hence, $q = \Delta U$

So, the correct match is

| | | | |
|---|---|---|---|
| A. | Isothermal process | (iv) | $W = -nRT \ln\left(\frac{V_1}{V_i}\right)$ |
| B. | Adiabatic process | (iii) | $W = \Delta U$ |
| C. | Isobaric process | (ii) | $W = -p \times \Delta V$ |
| D. | Isochoric process | (i) | $q = \Delta U$ |

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# Question44

## Which of the following expression is correct?

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**Options:**

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

B. $\Delta G = \frac{1}{RT^2 \ln K}$

C. $\Delta G^{\circ}=-RT\ln K$

D. $\Delta G^{\circ}=-\frac{1}{RT^{2}\ln K}$

**Answer: C**

## Solution:

Using equation,

$\Delta G=\Delta G^{\circ}+RT\ln Q$

where, $\Delta G\rightarrow$ Change in free energy

$\Delta G^{\circ}\rightarrow$ Change in standard free-energy

$R\rightarrow$ Gas constant

$T\rightarrow$ Temperature

$Q\rightarrow$ Reaction quotient

At equilibrium,

$\Delta G=0$ and $Q=K$

So, $0=\Delta G^{\circ}+RT\ln K$
$\Rightarrow\Delta G^{\circ}=-RT\ln K$

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# Question45

## Identify the reaction/process in which the entropy increases.

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**Options:**

A. $\mathrm{H(g)}+\mathrm{H}(g)\longrightarrow\mathrm{H}_2(g)$

B. $\mathrm{H_2O(g)}\longrightarrow\mathrm{H_2O(s)}$

C. $\mathrm{H_2O(l)}\longrightarrow\mathrm{H_2O(g)}$

D. $A(g)+B(g)+C(s)\longrightarrow 2D(s)$

**Answer: C**

## Solution:

Entropy order for solid, liquid and gas is :

Gas > Liquid > Solid

(a) $\mathrm{H(g)} + \mathrm{H(g)} \longrightarrow \mathrm{H_2(\,g)}$

Two gaseous atoms combine to form one gaseous molecule. Hence, entropy decreases.

(b) $\mathrm{H_2O(l)} \longrightarrow \mathrm{H_2O}(s)$

Entropy decreases when liquid freezes to solid form.

(c) $\mathrm{H_2O}(l) \longrightarrow \mathrm{H_2O}(g)$

Conversion of liquid to gas increases the entropy.

(d) $A(g) + B(g) + C(s) \longrightarrow 2D(s)$

In the above reaction, the product formed is a solid. Thus, entropy will decreases.

So, option (c) is the correct option.

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# Question46

$$\textbf{State } 1 \rightleftharpoons \textbf{State } 2 \rightleftharpoons \textbf{State 3}$$
$$\begin{pmatrix} T = 300\text{ K} \\ p = 15\text{bar} \\ 1\text{ mol} \end{pmatrix} \begin{pmatrix} T = 300\text{ K} \\ p = 10\text{bar} \\ 1\text{ mol} \end{pmatrix} \begin{pmatrix} T = 300\text{ K} \\ p = 5\text{bar} \\ 1\text{ mol} \end{pmatrix}$$

**Above shows a cyclic process. Calculate the total work done during one complete cycle. [Assume a single step to reach the next state].**

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**Options:**

A. 25/3 L bar

B. −25/3 L bar

C. 50/3 L bar

D. −50/3 L bar

**Answer: A**

## Solution:

The given process is isothermal reversible reaction.

State 1 to $2 p_1 = 15$ bar and $p_2 = 10$ bar

$$\begin{aligned} n &= 1 \text{ mol} \\ w_1 &= -2.303 nRT \log \frac{p_1}{p_2} \\ &= -2.303 \times 1 \times 0.083 \times 300 \log \frac{15}{10} \\ &= -10.09 \text{ L bar} \quad \text{.... (i)} \end{aligned}$$

State 2 to $3 p_1 = 10$ bar and $p_2 = 5$ bar

$$\begin{aligned} n &= 1 \text{ mol} \\ w_2 &= -2.303 nRT \log \frac{p_1}{p_2} \\ &= -2.303 \times 1 \times 0.083 \times 300 \log \frac{10}{5} \\ &= -17.26 \text{ Lbar} \quad \text{... (ii)} \end{aligned}$$

State 3 to $1 p_1 = 5$ bar and $p_2 = 15$ bar

$$\begin{aligned} n &= 1 \text{ mol} \\ w_3 &= -2.303 nRT \log \frac{p_1}{p_2} \\ &= -2.303 \times 1 \times 0.083 \times 300 \log \frac{5}{15} \\ &= +27.36 \text{ L bar} \quad \text{.... (iii)} \end{aligned}$$

$\therefore$ Total workdone,

$$\begin{aligned} &w = w_1 + w_2 + w_3 \\ &w = -10.09 - 17.26 + 27.36 = 0 \end{aligned}$$

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# Question47

**When an ideal gas expands isothermally from $5 \mathrm{~m}^3$ to $10 \mathrm{~m}^3$ at $25^{\circ} \mathrm{C}$ against a constant pressure of $10^7 \mathrm{~Nm}^{-2}$, then the work done on the gas is**

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**Options:**

A. $-100$ MJ

B. $-50$ MJ

C. $-0.5$ MJ

D. $-10^5$ MJ

**Answer: B**

## Solution:

Work done on isothermal irreversible expansion for ideal gas

$$\begin{aligned} W &= -p_{\text{ext}} \quad (V_2 - V_1) \\ &= -10^7(10-5) = -10^7 \times 5 = -50 \times 10^6 \text{ J} = -50 \text{ MJ}. \end{aligned}$$

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# Question48

**Find the approximate value of $(\Delta H - \Delta U)$ in $\mathrm{Jmol}^{-1}$, for the formation of CO from its elements at $298$ K. $\left(R = 8.314 \mathrm{JK}^{-1} \mathrm{~mol}^{-1}\right)$**

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**Options:**

A. $-1238$

B. $1238$

C. $2477$

D. $-2477$

**Answer: B**

### Solution:

$$\mathrm{C}(s) + \frac{1}{2}\mathrm{O}_2(g) \longrightarrow \mathrm{CO}(g)$$

$$\begin{aligned} &\Delta n_g = 1 - \frac{1}{2} = 0.5 \\ &\Delta H = \Delta U + \Delta n_g RT \\ &\Delta H - \Delta U = \Delta n_g \times R \times T = 0.5 \times 8.314 \times 298 \\ &\Delta H - \Delta U = 1238.7 \text{ J/mol.} \end{aligned}$$

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# Question49

**For the reaction, $\mathrm{H_2O}(l) \longrightarrow \mathrm{H_2O(g)}$ at $T = 100^{\circ}\mathrm{C}$ and $p = 1$ atm, choose the correct option.**

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**Options:**

A. $\Delta S_{\text{system}} > 0$ and $\Delta S_{\text{surrounding}} > 0$

B. $\Delta S_{\text{system}} > 0$ and $\Delta S_{\text{surrounding}} < 0$

C. $\Delta S_{\text{system}} < 0$ and $\Delta S_{\text{Surrounding}} > 0$

D. $\Delta S_{\text{system}} < 0$ and $\Delta S_{\text{surrounding}} < 0$

**Answer: B**

## Solution:

$H_2O(l) \longrightarrow H_2O(g)$
$T = 100\Upsilon\text{C} \Rightarrow p = 1\text{ atm}$

The process is an endothermic process, entropy increases during this change

$$\begin{aligned} & \Delta S = +ve \\ \therefore \quad & \Delta S_{\text{system}} > 0 \\ & \Delta S_{\text{surrounding}} < 0. \end{aligned}$$

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# Question50

### Two flasks $A$ and $B$ have equal volumes. $A$ is maintained at $300\text{ K}$ and $B$ at $600\text{ K}$. Equal masses of $\text{H}_2$ and $\text{CO}_2$ are taken in flasks $A$ and $B$ respectively. Find the ratio of total KE of gases in flask $A$ to that of $B$.

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**Options:**

A. 1 : 2

B. 11 : 1

C. 33 : 2

D. 55 : 7

**Answer: B**

## Solution:

Formula of kinetic energy $= \frac{3nRT}{2}$

Total kinetic energy expression for $A$ and $B$ flask,

$\mathrm{KE}_A=\frac{3n_ART_A}{2}$ and $\mathrm{KE}_B=\frac{3n_BRT_B}{2}$. ...... (i)

Let us consider common mass $= m$

$n_A=\frac{m}{2}$ and $n_B=\frac{m}{44}$

Taking ratio (i) and (ii) we get,

$$\frac{\mathrm{KE}_A}{\mathrm{KE}_B}=\frac{n_AT_A}{n_BT_B}$$
$$\frac{\mathrm{KE}_A}{\mathrm{KE}_B}=\frac{\frac{m}{2}\times 300}{\frac{m}{44}\times 600}$$
$[\mathrm{KE}_A:\mathrm{KE}_B=11:1]$

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# Question51

## When the temperature of 2 moles of an idealgas is increased by 20°C at constant pressure.Find the work involved in the process.

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**Options:**

A. 5R

B. 40R

C. 15R

D. 20R

**Answer: B**

## Solution:

Temperature of an ideal gas increased by $=20^{\circ}\mathrm{C}$ i.e. $T_2$

Initial temperature of gas, $T_1=0^{\circ}\mathrm{C}$

Number of moles, $n=2$

Formula of work done, $W=-nR\left(T_2-T_1\right)$

$$W=-2R(20-0)$$
$$W=-40R$$

But work involved in the process, $W=+\text{ ve}$

$W=40R$

Here, $R =$ Rate constant i.e. $8.314 \text{ JK}^{-1} \text{ mol}^{-1}$

$W = 40 \times 8.314 = 332.56 \text{ JK}^{-1} \text{ mol}^{-1}$

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# Question52

## If a chemical reaction is known to benon-spontaneous at 298 K but spontaneousat 350 K, then which among the followingconditions is true for the reaction?

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**Options:**

A. $\Delta G = -ve, \Delta H = -ve, \Delta S = +ve$

B. $\Delta G = +ve, \Delta H = +ve, \Delta S = +ve$

C. $\Delta G = -ve, \Delta H = +ve, \Delta S = +ve$

D. $\Delta G = +ve, \Delta H = +ve, \Delta S = -ve$

**Answer: C**

## Solution:

According to Gibb's free energy, $\Delta G = \Delta H - T\Delta S$

Reaction is non spontaneous i.e. $\Delta G > 0$ at 298 K.

As the temperature increases from 298 K to 350 K.

The reaction becomes spontaneous i.e. $\Delta G < 0$.

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

When $\Delta H$ and $\Delta S$ are + ve then, $\Delta G$ will be positive till $\Delta H > T\Delta S$.

When the temperature is increased further $\Delta H$ becomes less than $T\Delta S$ i.e.

$\Delta H < T\Delta S$ and $\Delta G$ becomes $-$ve.

Thus, reaction becomes spontaneous.

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