

Thermodynamics

Question 1

One mole of an ideal gas at 300 K and 20 atm expands to 2 atm under isothermal and reversible conditions. The work done by the gas is $-x \text{ kJ mol}^{-1}$. The value of x is ($R = 8.3 \text{ J K}^{-1} \text{ mol}^{-1}$)

TG EAPCET 2025 (Online) 2nd May Evening Shift

Options:

A.

5.73

B.

7.37

C.

3.75

D.

4.57

Answer: A

Solution:

Work done in isothermal reversible expansion,

$$W = -nRT \ln \left(\frac{p_2}{p_1} \right)$$

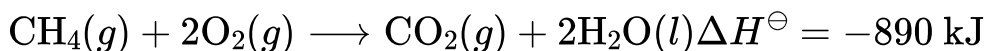
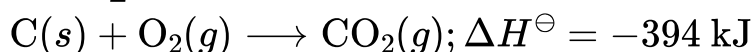
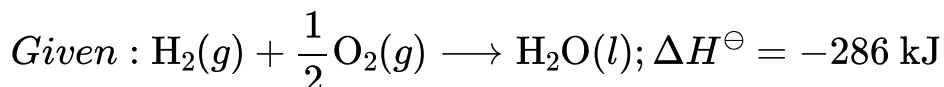
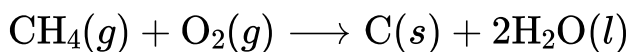
$$W = -(1)(8.3)(300) \ln \left(\frac{2}{20} \right) = 5730 \text{ J/mole or } 5.73 \text{ kJ/mol}$$

$$x = 5.73$$



Question2

At 298 K , the enthalpy change (in kJ) for the reaction given below is



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Options:

A.

+496

B.

-496

C.

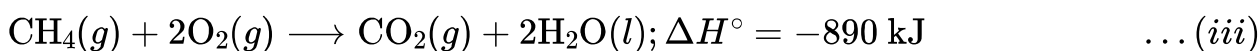
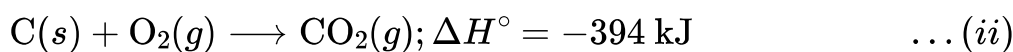
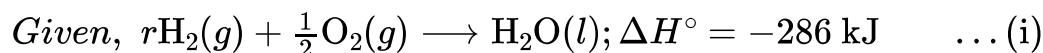
-1284

D.

+680

Answer: B

Solution:



Reverse reaction given in Eq. (ii),
 $\text{CO}_2(\text{g}) \longrightarrow \text{C}(\text{s}) + \text{O}_2(\text{g}); \Delta H^\circ = +394 \text{ kJ} \dots (\text{iv})$

Now, add Eq. (iii) and (iv),
 $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) + \text{CO}_2(\text{g}) \longrightarrow \text{CO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l}) + \text{C}(\text{s}) + \text{O}_2(\text{g})$
 $\text{CH}_4(\text{g}) + \text{O}_2(\text{g}) \longrightarrow \text{C}(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
So, $\Delta H^\circ = -890 + 394 = -496 \text{ kJ}$

Question3

At 300 K, 3.0 moles of an ideal gas at 3.0 atm pressure is compressed isothermally to one half of its volume by an external pressure of 6.0 atm . The work done (in kJ) is (Given, $R = 0.082 \text{ L atm K}^{-1} \text{ mol}^{-1}$) ($1 \text{ L atm} = 1013 \text{ J}$)

TG EAPCET 2024 (Online) 11th May Morning Shift

Options:

- A. 7.476
- B. 11.214
- C. 3.738
- D. 14.952

Answer: A

Solution:

Given, temperature = 300 K, number of moles (n) = 3

Change in volume (ΔV) = $-\frac{1}{2}(V)$ where

V is initial volume and external pressure, $p_{\text{ext}} = 6 \text{ atm}$.

Pressure at 300 K of gas, $p = 3 \text{ atm}$

Work done in isothermal compression is given by

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

$$W = -p_{\text{ext}}\left(-\frac{V}{2}\right)$$

Using ideal gas equation $pV = nRT$,

$$V = \frac{nRT}{p}$$

Substitute the value of V in Eq. (ii)

$$W = p_{\text{ext}} \frac{(nRT)}{2p}$$

Substitute all values in Eq. (iv)

$$W = \frac{6 \times 3 \times 0.082 \times 300}{2 \times 3} \times 101.3 \text{ J}$$

$$W = 7475.94 \text{ J or } W = 7.476 \text{ kJ}$$

Question4

The $\Delta_f H^\ominus$ of $\text{AO}(s)$, $\text{BO}_2(g)$ and $\text{ABO}_3(s)$ is -635 , x and $-1210 \text{ kJ mol}^{-1}$ respectively.

$\text{ABO}_3(s) \rightarrow \text{AO}(s) + \text{BO}_2(g) : \Delta_r H^\ominus = 175 \text{ kJ mol}^{-1}$). What is the value of x (in kJ mol^{-1})?

TG EAPCET 2024 (Online) 10th May Evening Shift

Options:

A. -750

B. +400

C. -400

D. +750

Answer: C

Solution:

Given values:

$$\Delta_f H^\ominus \text{ of } \text{AO}(s) = -635 \text{ kJ/mol}$$

$$\Delta_f H^\ominus \text{ of } \text{BO}_2(g) = x \text{ kJ/mol}$$

$$\Delta_f H^\ominus \text{ of } \text{ABO}_3(s) = -1210 \text{ kJ/mol}$$

$$\Delta_r H^\ominus = 175 \text{ kJ/mol for the reaction:}$$



The enthalpy change for the reaction, $\Delta_r H^\ominus$, is calculated using the formula:

$$\Delta_r H^\ominus = [\Delta_f H^\ominus(AO) + \Delta_f H^\ominus(BO_2)] - \Delta_f H^\ominus(ABO_3)$$

Substitute the known values:

$$175 = [-635 + x] - [-1210]$$

Simplifying:

$$175 = x + 1210 - 635$$

$$175 = x + 575$$

$$x = 175 - 575$$

$$x = -400 \text{ kJ/mol}$$

Therefore, the value of x is -400 kJ/mol .

Question5

The standard enthalpy of combustion of C (graphite), $H_2(g)$ and $CH_3OH(l)$ respectively are -393 , -286 and -726 kJ mol^{-1} . What is the standard enthalpy of formation of methanol?

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Options:

A. -726 kJ mol^{-1}

B. -239 kJ mol^{-1}

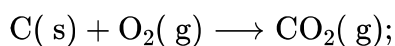
C. -96 kJ mol^{-1}

D. $+96 \text{ kJ mol}^{-1}$

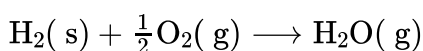
Answer: B

Solution:

According to given data

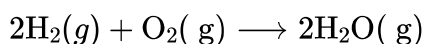


$$\Delta H_C = -393 \text{ kJ/mol}$$



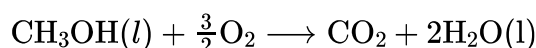
$$\Delta H_C = -286 \text{ kJ/mol}$$

Multiply Eq. (ii) by (2),

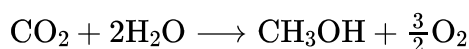


$$\Delta H_C = -572 \text{ kJ/mol.}$$

For combustion of $\text{CH}_3\text{OH}(\text{l})$

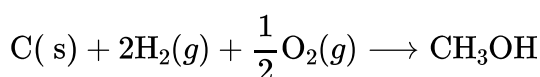


$$\Delta H_C = -726 \text{ kJ/mol}$$



$$\Delta H_C = +726 \text{ kJ/mol}$$

Add Eqs. (i), (iii) and (iv), we get



$$\Delta H_C = -393 + (-572) - (-726)$$

$$= -239 \text{ kJ/mol}$$

Question6

At 61 K , one mole of an ideal gas of 1.0 L volume expands isothermally and reversibly to a final volume of 10.0 L . What is the work done in the expansion?

TG EAPCET 2024 (Online) 9th May Evening Shift

Options:

A. -11.52 L atm

B. -23.04 L atm



C. -46.08 L atm

D. -5.76 L atm

Answer: A

Solution:

Given,

Initial volume (V_i) = 1.0 L

Final volume (V_f) = 10.0 L

n (number of moles) = 1

Temperature = 61 K

Now, workdone for isothermal process is given by

$$W = -nRT \times 2.303 \log_{10} \left(\frac{V_f}{V_i} \right)$$

Substitute all the values,

$$W = -2.303 \times 1 \times 0.0821 \times 61 \times \log_{10} \left(\frac{10}{1} \right)$$

$$= -11.52 \text{ L atm}$$

Question7

The standard enthalpy of formation ($\Delta_f H^\ominus$) of ammonia is $-46.2 \text{ kJ mol}^{-1}$, What is the $\Delta_r H^\ominus$ of the following reaction?

$$\text{N}_2(g) + 3\text{H}_2(g) \longrightarrow 2\text{NH}_3(g)$$

TG EAPCET 2024 (Online) 9th May Morning Shift

Options:

A. -46.2 kJ

B. +46.2 kJ

C. -92.4 kJ

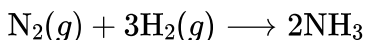
D. -184.8 kJ



Answer: C

Solution:

Given reaction is



Two moles of ammonia are formed during the reaction. Thus, the reaction for formation of one mole of ammonia is given

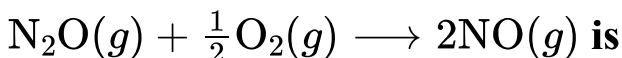
$$\Delta_f H^\ominus = -46.2 \text{ kJ/mol}$$

The standard enthalpy of reaction on is

$$\begin{aligned}\Delta_r H^\ominus &= 2 \times \Delta_f H^\ominus \\ &= 2 \times -46.2 \\ &= -92.4 \text{ kJ}\end{aligned}$$

Question8

The enthalpies of formation of gaseous N_2O and NO at 298 K are 82.0 and 90.0 kJ mol^{-1} respectively. The enthalpy change of the reaction



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Options:

- A. -74 kJ
- B. +98 kJ
- C. +89 kJ
- D. -47 kJ

Answer: B

Solution:

Given the following information about enthalpies of formation at 298 K:

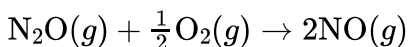
For the formation of N_2O :



For the formation of NO :



To find the enthalpy change for the reaction:



We use the formula for calculating the enthalpy change of the reaction based on the enthalpies of formation:

$$\Delta H = \sum \Delta H_f(\text{products}) - \sum \Delta H_f(\text{reactants})$$

Applying this to the given reaction:

$$\Delta H = (2 \times \Delta H_{f(\text{NO})}) - (\Delta H_{f(\text{N}_2\text{O})} + \frac{1}{2}\Delta H_{f(\text{O}_2)})$$

Since the formation enthalpy of a pure element in its standard state, like O_2 , is zero:

$$\Delta H = 2 \times 90 \text{ kJ/mol} - 82 \text{ kJ/mol}$$

Calculating gives:

$$\Delta H = 180 \text{ kJ/mol} - 82 \text{ kJ/mol} = 98 \text{ kJ/mol}$$

Therefore, the enthalpy change of the reaction is $+98 \text{ kJ/mol}$.

Question9

If 2.5 moles of an ideal gas at a certain temperature are allowed to expand isothermally and reversibly from an initial volume of 2dm^3 to 20dm^3 , the work done by the gas is -16.5 kJ . The temperature (in **K) of the gas is (Round off to the nearest value) (**

$$R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \text{)}$$

TS EAMCET 2023 (Online) 12th May Morning Shift

Options:

A. 445

B. 245

C. 345

D. 745

Answer: C

Solution:

We can solve this using the formula for the work done by an ideal gas during an isothermal, reversible expansion:

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

where:

- W is the work done by the gas (given as $-16.5 \text{ kJ} = -16500 \text{ J}$),
- $n = 2.5 \text{ mol}$,
- $R = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$,
- $V_1 = 2 \text{ dm}^3$ and $V_2 = 20 \text{ dm}^3$.

Follow these steps:

Write down the work formula:

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

Solve for the temperature T :

$$T = -\frac{W}{nR \ln \left(\frac{V_2}{V_1} \right)}$$

Substitute the known values:

$$T = -\frac{-16500}{2.5 \times 8.314 \times \ln \left(\frac{20}{2} \right)}$$

Calculate the volume ratio and its natural logarithm:

$$\frac{V_2}{V_1} = \frac{20}{2} = 10, \quad \text{so} \quad \ln(10) \approx 2.3026.$$

Now compute the temperature:

$$T = \frac{16500}{2.5 \times 8.314 \times 2.3026}.$$

First, compute the product in the denominator:

$$2.5 \times 8.314 = 20.785,$$

and then

$$20.785 \times 2.3026 \approx 47.88.$$

Finally, calculate T :

$$T \approx \frac{16500}{47.88} \approx 344.8 \text{ K}.$$



Rounding off to the nearest value gives approximately 345 K.

Thus, the correct option is Option C: 345.

